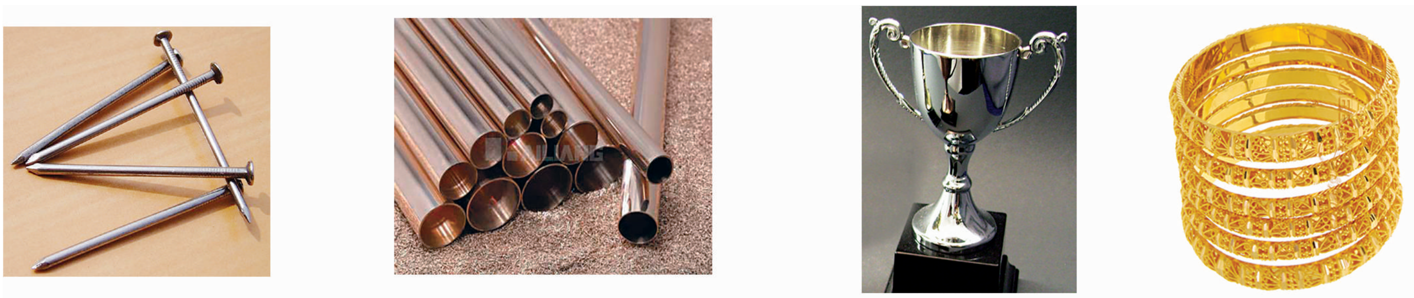
3

METALS AND NON-METALS

There are 115 chemical elements known at present. There are similarities as well as differences in the properties of these elements. **On the basis of their properties, all the elements can be divided into two main groups : metals and non-metals.** Both, metals as well as non-metals are used in our daily life. We also use a large number of compounds of metals and non-metals. Before we go further and give the definitions of metals and non-metals, we should know the meaning of some new terms such as malleable, ductile and brittle. **Malleable means which can be beaten with a hammer to form thin sheets (without breaking). Ductile means which can be stretched (or drawn) to form thin wires. And brittle means which breaks into pieces on hammering or stretching**. Keeping these points in mind, we will now write the definitions of metals and non-metals.

METALS

Metals are the elements that conduct heat and electricity, and are malleable and ductile. Metals are also lustrous (shiny), hard, strong, heavy and sonorous (which make ringing sound when struck). **Some of the examples of metals are : Iron, Aluminium, Copper, Silver, Gold, Platinum, Zinc, Tin, Lead, Mercury,**



(*a*) These nails are made of iron metal

(*b*) These pipes are made of copper metal

(*c*) This cup is made of silver metal

(*d*) These bangles are made of gold metal

Figure 1. Iron, copper, silver and gold are some of the metals.

Sodium, Potassium, Calcium and Magnesium (see Figure 1). A majority of the known elements are metals. All the metals are solids, except mercury which is a liquid metal.

During chemical reactions, metals can form positive ions by losing electrons. Based on this observation, we can write another definition of metals as follows : **Metals are the elements (except hydrogen) which form positive ions by losing electrons (or donating electrons).** For example, aluminium (Al) is a metal which forms positively charged aluminium ions (Al^{3+}) by losing electrons. In fact, **metals are known as electropositive elements because they can form positive ions by losing electrons.**

Metals are widely used in our daily life for a large number of purposes. The cooking utensils, electric fans, sewing machines, cars, buses, trucks, trains, ships and aeroplanes are all made of metals or mixtures of metals called alloys. In fact, the list of articles made of metals which we use in our daily life is unending. Metals are very important for the National economy of every country.

The most abundant metal in the earth's crust is aluminium, which constitutes about 7% of the earth's crust. The second most abundant metal in the earth's crust is iron, which constitutes about 4% of the earth's crust. The major metals in the earth's crust in the decreasing order of their abundance are : Aluminium, Iron, Calcium, Sodium, Potassium and Magnesium.

NON-METALS

Non-metals are the elements that do not conduct heat and electricity, and are neither malleable nor ductile. They are brittle. Non-metals are not lustrous (not shiny), they have dull appearance. Non-metals are generally soft, and not strong. They are light substances and non-sonorous (which do not make ringing sound when struck). **Some of the examples of non-metals are : Carbon, Sulphur, Phosphorus, Silicon, Hydrogen, Oxygen, Nitrogen, Chlorine, Bromine, Iodine, Helium, Neon and Argon** (see Figure 2)**.** The

(*a*) Carbon

(*b*) Sulphur

(*c*) This bottle contains chlorine gas

Figure 2. Carbon, sulphur and chlorine are some of the non-metals.

two allotropic forms of carbon element, **diamond and graphite, are also non-metals.** In fact, there are 22 non-metals (or non-metallic elements). Out of these, 10 non-metals are solids, 1 non-metal (bromine) is a liquid whereas the remaining 11 non-metals are gases. Thus, all the non-metals are solids or gases, except bromine which is a liquid non-metal at the room temperature.

During chemical reactions, non-metals can form negative ions by gaining electrons. Based on this observation, we can write another definition of non-metals as follows : **Non-metals are the elements which form negative ions by gaining electrons (or accepting electrons).** For example, oxygen (O) is a non-metal which forms negatively charged oxide ions (O^{2-}) by gaining electrons. In fact, **non-metals are known as electronegative elements because they can form negative ions by gaining electrons.** There is, however, an exception. Hydrogen (H) is the only non-metal element which loses electrons to form positive ions, hydrogen ions (H^+). We will discuss the reason for this in higher classes.

Though non-metals are small in number as compared to metals, but they play a very important role in our daily life. In fact, life would not have been possible without the presence of non-metals on the earth. For example, **carbon** is one of the most important non-metals because all the life on this earth is based on carbon compounds. This is because the carbon compounds like proteins, fats, carbohydrates, vitamins and enzymes, etc., are essential for the growth and development of living organisms. Another non-metal **oxygen**

is equally important for the existence of life. This is because the presence of oxygen gas in the air is essential for breathing to maintain life. It is also necessary for the combustion (or burning) of fuels which provide us energy for various purposes. **Nitrogen** is an inert gaseous non-metal whose presence in air reduces the rate of combustion and makes it safe. Another non-metal **sulphur** is present in many of the substances found in plants and animals. For example, sulphur is present in hair, onion, garlic and wool, etc. Non-metals are required to make vegetable *ghee*, fertilisers, acids, explosives and fungicides, etc.

The most abundant non-metal in the earth's crust is oxygen, which constitutes about 50% of the earth's crust. The second most abundant non-metal in the earth's crust is silicon, which constitutes about 26% of the earth's crust. The major non-metals in the earth's crust in the decreasing order of their abundance are: Oxygen, Silicon, Phosphorus and Sulphur. It should be noted that although non-metals are small in number (being only 22 in all), but they are the major constituents of earth, air and oceans (seas). For example, two non-metals, oxygen and nitrogen are the main constituents of air; two non-metals, oxygen and silicon, are the main constituents of earth; and two non-metals, hydrogen and oxygen are the main constituents of oceans (in the form of water). Another non-metal, chlorine, also occurs in oceans in the form of metal chlorides.

All the metals have similar properties. All the non-metals also have similar properties. But **the properties of non-metals are opposite to those of metals**. We will now describe the properties of metals and non-metals, one by one.

PHYSICAL PROPERTIES OF METALS

The important physical properties of metals are given below.

1. Metals are malleable, that is, metals can be beaten into thin sheets with a hammer (without breaking)

It we take a piece of aluminium metal, place it on a block of iron and beat it with a hammer four or five times, we will find that the piece of aluminium metal turns into a thin aluminium sheet, without breaking. And we say that aluminium metal is malleable or that it shows malleability. *The property which allows the metals to be hammered into thin sheets is called malleability.* **Malleability is an important characteristic property of metals.**

Figure 3. Metals can be beaten into thin sheets with a hammer. They are malleable.

Most of the metals are malleable. **Gold and silver metals are some of the best malleable metals.** Aluminium and copper metals are also highly malleable metals. All these metals can be beaten with a hammer to form very thin sheets called foils. For example, silver metal can be hammered into thin silver foils because of its high malleability. The silver foils are used for decorating sweets. Similarly, aluminium metal is quite malleable and can be converted into thin sheets called aluminium foils. **Aluminium foils are used for packing food items like biscuits, chocolates, medicines, cigarettes,** etc. Milk bottle caps are also made of aluminium foil. Aluminium sheets are used for making cooking utensils. Copper metal is also highly malleable. So, copper sheets are used to make utensils and other containers. Iron is also a quite malleable metal which can be hammered to form iron sheets. These iron sheets are used to make boxes, buckets, drums and water tanks, etc.

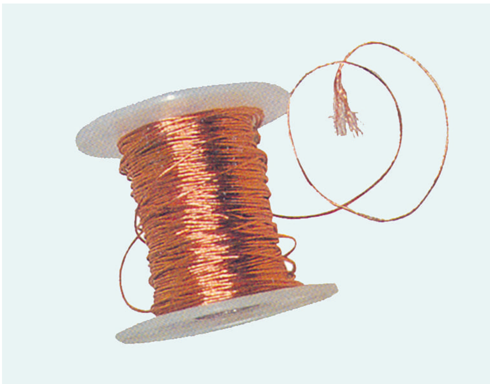


Figure 4. Metals can be drawn (or stretched) into thin wires. They are ductile.

2. Metals are ductile, that is, metals can be drawn or stretched) into thin wires

The metals such as copper, aluminium, magnesium and iron are available in the form of wires. *The property which allows the metals* *to be drawn into thin wires is called ductility.* **Ductility is another important characteristic property of metals.** Most of the metals are ductile. But all the metals are not equally ductile. Some metals are more ductile than the others. **Gold is the most ductile metal.** For example, just 1 gram of gold can be drawn into a thin wire about 2 kilometres long ! Silver is also among the best ductile metals. **Copper and aluminium metals are also very ductile and can be drawn into thin copper wires and aluminium wires (which are used as electric wires).** Iron, magnesium and tungsten metals are also quite ductile and can be drawn into thin wires. Iron wires are used for making wire gauzes. Magnesium wires are used in science experiments in the laboratory. And thin wires of tungsten metal are used for making the filaments of electric bulbs.

From the above discussion we conclude that **metals are malleable and ductile.** It is due to the properties of malleability and ductility that metals can be given different shapes to make various articles needed by us.

3. Metals are good conductors of heat

By saying that metals are good conductors of heat we mean that **metals allow heat to pass through them easily.** This can be demonstrated as follows.

We take a flat aluminium rod and fix some small iron nails on it with the help of wax. This rod (alongwith its iron nails) is clamped to a stand as shown in Figure 5. Let us heat the free end (left end) of the aluminium rod by keeping a burner below it. We will see that the iron nails attached to aluminium rod with wax start falling one by one. The nail attached nearest to the heated end of rod falls down first. And then the next ones fall. But the nail attached to the clamped end of the rod drops last of all. These observations can be explained as follows :

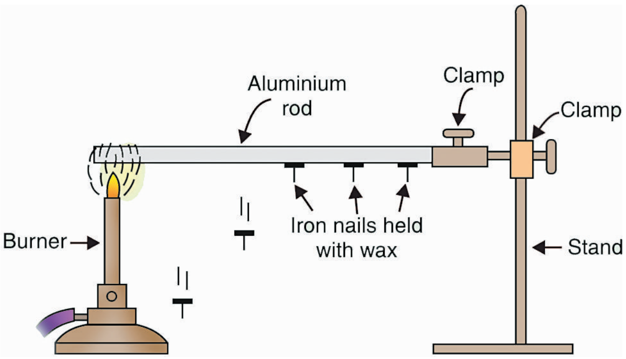


Figure 5. Experiment to show that a metal (here aluminium) conducts heat.

The burner is placed below the left end of aluminium rod. So, the left end of aluminium rod gets heated first. Now, the left end of aluminium rod is hot but the right end of rod is cold. So, heat now travels from the hotter left end of aluminium rod to its colder right end. As heat travels from the left side to the right side along the aluminium rod, it melts the wax which holds the nails. Due to this the nails fall down one by one. From this experiment we conclude that an aluminium metal rod conducts heat.

Metals are generally good conductors of heat (The conduction of heat is also called thermal conductivity). **Silver metal is the best conductor of heat.** It has the highest thermal conductivity. Copper and aluminium metals are also very good conductors of heat. **The cooking utensils and water boilers, etc., are usually made of copper or aluminium metals because they are very good conductors of heat.** The poorest conductor of heat among the metals is **lead.** Mercury metal is also a poor conductor of heat. We will now describe how a metal conducts heat. When a metal is heated, its atoms gain energy and vibrate more vigorously. This energy is transferred to the electrons present in the atoms. These electrons can move through the metal. When the energetic electrons move through the metal, they transfer energy to other electrons and atoms of the metal (some distance away from the end that is being heated). In this way, heat is conducted from one end of the metal to its other end. Thus, **heat conductivity (or thermal conductivity) is a characteristic property of metals.**

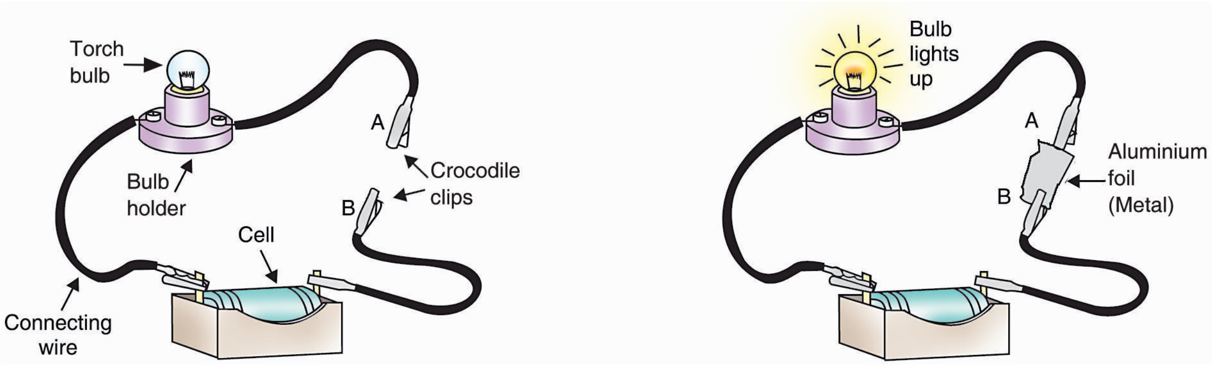
Figure 6. Metals conduct heat well. That is why this frying pan is made of a metal.

4. Metals are good conductors of electricity

By saying that metals are good conductors of electricity, we mean that **metals allow electricity (or**

electric current) to pass through them easily. This can be demonstrated as follows.

We take a dry cell, a torch bulb fitted in a holder and some connecting wires (copper wires) with crocodile clips, and connect them [as shown in Figure 7(*a*)] to make an electric circuit. There is a gap between the ends of the crocodile clips *A* and *B* so no current flows in the incomplete circuit shown in Figure 7(*a*) and hence the bulb does not light up. Let us now insert a piece of aluminium foil between the ends of crocodile clips *A* and *B* as shown in Figure 7(*b*). We will see that the bulb lights up at once. This means that the aluminium foil allows electric current to pass through it. This shows that aluminium metal conducts electric current (or electricity). In other words, aluminium metal is a good conductor of electricity. Please note that the connecting wires used in this experiment are made of copper metal. Since these copper connecting wires allow electric current to pass through them, therefore, copper metal is also a good conductor of electricity.



(*a*) (*b*)

Figure 7. Experiment to show that a metal (here aluminium) conducts electricity.

Metals are good conductors of electricity. The metals offer very little resistance to the flow of electric current and hence show high electrical conductivity. **Silver metal is the best conductor of electricity.** Copper metal is the next best conductor of electricity followed by gold, aluminium and tungsten. **The electric wires are made of copper and aluminium metals because they are very good conductors of electricity** (see Figure 8). The metals like iron and mercury offer comparatively greater resistance to the flow of current, so they have lower electrical conductivity. We will now describe how a metal conducts electricity. Metals are good conductors of electricity because they contain free electrons. These free electrons can move easily through the metal and conduct electric current. Thus, **electrical conductivity is another characteristic property of metals.** From the above discussion we conclude that **metals are good conductors of heat and electricity.**

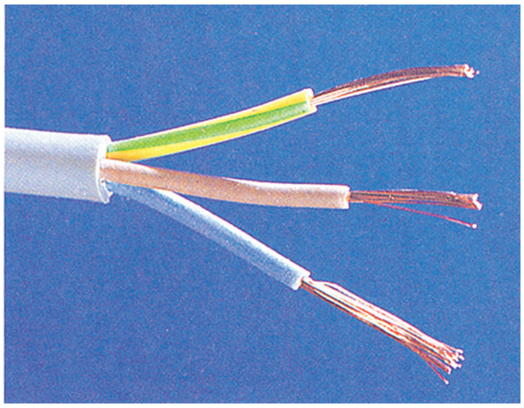


Figure 8. Because copper metal is a good conductor of electricity, it is used in making electric wires. The electric wires have a covering of plastic such as PVC (which is an insulator).

The electric wires that carry current in our homes have a covering of plastic such as Poly Vinyl Chloride (PVC). Polyvinyl chloride is an insulator. It does not allow electric current to pass through it. The electric wires have a covering of an insulating material (like PVC) around them so that even if we happen to touch them, the current will not pass through our body and hence we will not get an electric shock.

5. Metals are lustrous (or shiny), and can be polished

By saying that metals are lustrous, we mean that **they have a shining surface.** For example, gold, silver and copper are shiny metals and they can be polished. *The property of a metal of having a shining surface is called 'metallic lustre' (chamak).* The shiny appearance of metals makes them useful in making jewellery and decoration pieces. For example, gold and silver are used for making jewellery because they are bright and shiny (see Figure 9). The shiny appearance of metals makes them good reflectors of light. Silver metal is an excellent reflector of light. This is why it is used in making mirrors.

A metal has a shining surface only when it is fresh. When a metal has been kept exposed to air for a long time, then it gets a dull appearance. It loses most of its shine or brightness. **The metals lose their shine or brightness on keeping in air for a long time and acquire a dull appearance due to the formation of a thin layer of oxide, carbonate or sulphide on their surface (by the slow action of the various gases present in air).** We say that the metal surface has been corroded. If we rub the dull surface of a metal object with a sand paper, then the outer corroded layer is removed and the metal object becomes shiny and bright once again.

Figure 9. Metals are lustrous (or shiny). These bangles are made of 'gold' metal because it is a highly lustrous (or shiny) metal.

6. Metals are generally hard (except sodium and potassium which are soft metals)

Most of the metals are hard. But all the metals are not equally hard. The hardness varies from metal to metal. Most of the metals like iron, copper, aluminium, etc., are very hard. They cannot be cut with a knife. There are some exceptions. **Sodium and potassium are soft metals which can be easily cut with a knife.** We can perform the following experiment to study the hardness of metals.

Take a piece of iron (or copper) metal. Try to cut it with a knife. We will find that the piece of iron (or copper) metal cannot be cut with a knife. This tells us that iron (and copper) metals are very hard. Now, hold a piece of sodium metal carefully with a pair of tongs and dry it by pressing between the folds of a filter paper. Place it on a watch glass and try to cut it with a dry knife. We will find that the piece of sodium metal can be easily cut into small pieces (just like wax) (see Figure 10). This shows that sodium metal is soft.

Figure 10. Sodium metal is so soft that it can be easily cut with a knife.

7. Metals are strong (except sodium and potassium metals which are not strong)

By saying that metals are strong we mean that **they can hold large weights without snapping (without breaking).** For example, iron metal (in the form of steel) is very strong. Due to this iron metal is used in the construction of bridges, buildings, railway lines, girders, machines, vehicles and chains, etc. (see Figure 11). Though most of the metals are strong but some of the metals are not strong. For example, sodium and potassium metals are not strong.

Figure 11. Iron metal (as steel) is very strong. It is used to build bridges like this one.

8. Metals are solids at room temperature (except mercury which is a liquid metal)

Most of the metals like iron, copper, aluminium, silver and gold, etc., are solids at the room temperature. Only one metal, mercury, is in liquid state at the room temperature.

9. Metals have high melting points and boiling points (except sodium and potassium metals which have low melting and boiling points)

For example, iron metal has a high melting point of 1535°C. This means that solid iron melts and turns into liquid iron (or molten iron) on heating to a high temperature of 1535°C. Copper metal has also a high melting point of 1083°C. There are, however, some exceptions. For example, sodium and potassium metals have low melting points (of 98°C and 64°C respectively). Gallium and cesium metals also have low melting points (of 30°C and 28°C respectively). The melting points of gallium and cesium metals are so low that they start melting in hand (by the heat of our body).

10. Metals have high densities (except sodium and potassium metals which have low densities)

By saying that metals have high densities, we mean that **metals are heavy substances.** For example, the density of iron is 7.8 g/cm^3 which is quite high. So, iron metal is a heavy substance. There are, however, some exceptions. For example, sodium and potassium metals have low densities (of 0.97 g/cm^3 and 0.86 g/cm^3 respectively). They are very light metals.

Figure 12. Metals are sonorous. So they are used to make bells.

11. Metals are sonorous. That is, metals make sound when hit with an object

Sonorous means capable of producing a deep or ringing sound. If we suspend a big piece of a metal and strike it with an object, we will find that it makes a ringing sound. And we say that the metal is sonorous. *The property of metals of being sonorous is called sonorousness or sonority.* **It is due to the property of sonorousness (or sonority) that metals are used for making bells, and strings (wires) of musical instruments like *sitar* and violin** (see Figure 12)**.**

12. Metals usually have a silver or grey colour (except copper and gold)

Copper has a reddish-brown colour whereas gold has a yellow colour.

PHYSICAL PROPERTIES OF NON-METALS

The physical properties of non-metals are just the opposite of the physical properties of metals. The important physical properties of non-metals are given below :

1. Non-metals are neither malleable nor ductile. Non-metals are brittle (break easily)

Since non-metals are not malleable, they cannot be beaten with a hammer to form thin sheets. Again, since non-metals are not ductile, they cannot be stretched to form thin wires. Thus, **solid non-metals can neither be hammered into thin sheets nor drawn into thin wires.** Non-metals are brittle which means that non-metals break into pieces when hammered or stretched. For example, sulphur and phosphorus are solid non-metals which are non-malleable and non-ductile. When sulphur or phosphorus are beaten with a hammer or stretched, they break into pieces (they do not form thin sheets or wires). Carbon is also a solid non-metal which is brittle (see Figure 13). **The property of being brittle (breaking easily) is called brittleness.** Thus, **brittleness is a characteristic property of non-metals.** Please note that we can consider the brittleness of solid non-metals only. It is not applicable to liquid or gaseous non-metals.

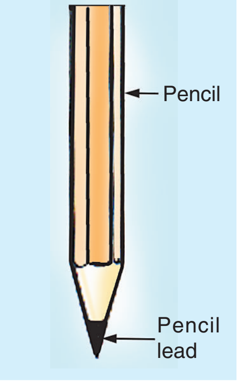
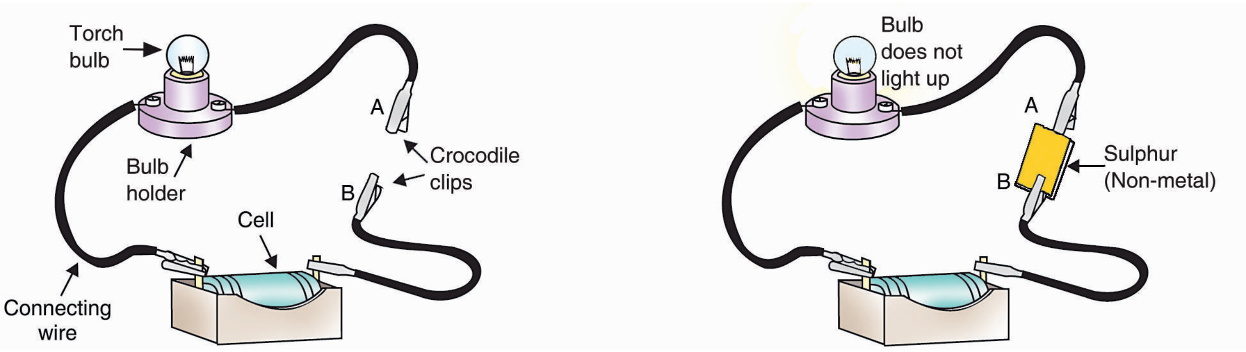


Figure 13. Pencil lead is made of graphite (which is a form of carbon non-metal). Pencil lead breaks easily. It is brittle.

2. Non-metals do not conduct heat and electricity

Non-metals do not conduct heat and electricity because unlike metals, they have no free electrons (which are necessary to conduct heat and electricity). For example, sulphur and phosphorus are non-metals which do not conduct heat and electricity. There is, however, one exception. **Carbon (in the form of graphite) is the only non-metal which is a good conductor of electricity.** Since graphite (which is an allotropic form of carbon) is a good conductor of electricity, it is used for making electrodes.

We will now describe an experiment to demonstrate that non-metals do not conduct electricity. This can be done as follows. We take a dry cell, a torch bulb fitted in a holder and some connecting wires (copper wires) with crocodile clips, and connect them [as shown in Figure 14(*a*)] to make an electric circuit. There is a gap between the ends of crocodile clips *A* and *B* so no current flows in the open circuit shown in



(*a*) (*b*)

Figure 14. Experiment to show that a non-metal (here sulphur) does not conduct electricity.

Figure 14(*a*). Let us now insert a piece of sulphur (which is a non-metal) between the crocodile clips *A* and *B* as shown in Figure 14(*b*). We will see that the bulb does not light up at all. This means that sulphur does not allow electric current to pass through it and no current flows in the circuit. This observation shows that sulphur (a non-metal) does not conduct electricity.

3. Non-metals are not lustrous (not shiny). They are dull

Non-metals do not have lustre (*chamak*) which means that **non-metals do not have a shining surface. The solid non-metals have a dull appearance.** For example, sulphur and phosphorus are non-metals which have no lustre, that is, they do not have a shining surface. They appear to be dull. There is, however, an exception. **Iodine is a non-metal having lustrous appearance.** It has a shining surface (like that of metals).

4. Non-metals are generally soft (except diamond which is an extremely hard non-metal)

Most of the solid non-metals are quite soft. For example, sulphur and phosphorus are solid non-metals which are quite soft. **Only one non-metal carbon (in the form of diamond) is very hard.** In fact, diamond (which is an allotropic form of carbon) is the hardest natural substance known.

5. Non-metals are not strong. They are easily broken

For example, graphite is a non-metal which is not strong. It has low strength. So, when a large weight is placed on a graphite sheet, it gets snapped (breaks).

6. Non-metals may be solid, liquid or gases at the room temperature

Non-metals can exist in all the three physical states : solid, liquid and gaseous. For example, carbon, sulphur and phosphorus are solid non-metals; bromine is a liquid non-metal; whereas hydrogen, oxygen, nitrogen and chlorine are gaseous non-metals.

7. Non-metals have comparatively low melting points and boiling points (except diamond which is a non-metal having a high melting point and boiling point)

For example, the melting point of sulphur is 115°C which is quite low. The melting point of diamond is, however, more than 3500°C, which is very high.

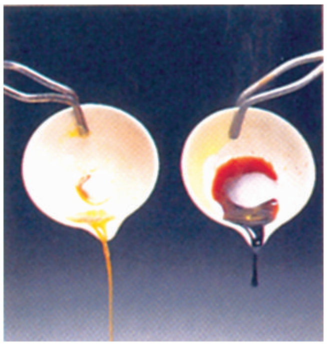


Figure 15. The non-metal 'sulphur' has a comparatively low melting point of 115°C. It can be easily melted by heating.

8. Non-metals have low densities, that is, non-metals are light substances

For example, the density of sulphur is 2 g/cm^3, which is quite low.

9. Non-metals are non-sonorous. They do not produce sound when hit with an object

10. Non-metals have many different colours

For example, sulphur is yellow, phosphorus is white or red, graphite is black, chlorine is yellowish-green whereas hydrogen and oxygen are colourless.

Exceptions in Physical Properties

We have just studied the physical properties of metals and non-metals. The physical properties of non-metals are different from those of metals but there are some exceptions. The important exceptions are given below :

1. Electrical Conductivity. We have studied that non-metals do not conduct electricity. But carbon non-metal (in the form of graphite) is an exception. **Carbon (in the form of graphite) is a non-metal which conducts electricity.** Thus, graphite is a good conductor of electricity (just like metals).

2. Lustre. We have studied that non-metals do not have lustre (*chamak*), they have dull appearance. But iodine is an exception. **Iodine is a non-metal which is lustrous, having a shining surface** (like that of metals).

3. Hardness and Softness. We have studied that metals are hard. But alkali metals (such as lithium, sodium and potassium) are exceptions. **Alkali metals (lithium, sodium and potassium) are soft** (just like solid non-metals). We have also studied that solid non-metals are soft. But carbon (in the form of diamond) is an exception. **Carbon (in the form of diamond) is a non-metal which is extremely hard** (just like metals).

4. Physical State. We have studied that metals are solids. But mercury metal is an exception. **Mercury metal is a liquid at room temperature.**

5. Melting Points and Boiling Points. We have studied that metals have high melting points and boiling points. But sodium, potassium, gallium and cesium metals are exceptions. **Sodium, potassium, cesium and gallium metals have low melting points (just like non-metals).** We have also studied that non-metals have low melting points and boiling points. But diamond is an exception. **Diamond is a non-metal which has a very high melting point and boiling point** (just like metals).

6. Density. We have studied that metals have high densities. But alkali metals (such as lithium, sodium and potassium) are exceptions. They have low densities (like that of non-metals).

From the above discussion it is obvious that **we cannot classify the elements as metals or non-metals clearly on the basis of their physical properties alone because there are many exceptions.** Elements can, however, be classified more clearly as metals and non-metals on the basis of their chemical properties. We will now study the chemical properties of metals and non-metals, one by one.

CHEMICAL PROPERTIES OF METALS

Metals and non-metals show different chemical properties. First we will describe the chemical properties of metals and then of non-metals. The important chemical properties of metals are given below :

1. Reaction of Metals with Oxygen (of Air)

When metals are burnt in air, they react with the oxygen of air to form metal oxides :

$$\underset{}{\textbf{Metal}} + \underset{\text{(From air)}}{\textbf{Oxygen}} \longrightarrow \underset{\text{(Basic oxide)}}{\textbf{Metal oxide}}$$

Thus, **metals react with oxygen to form metal oxides. Metal oxides are basic in nature.** Some of the metal oxides react with water to form alkalis. **Metal oxides, being basic, turn red litmus solution blue.**

The vigour of reaction with oxygen depends on the chemical reactivity of metal. Some metals react with oxygen even at room temperature, some react on heating, whereas still others react only on strong heating. Here are some examples :

(*i*) Sodium metal reacts with the oxygen of air at room temperature to form a basic oxide called sodium oxide :

$$\underset{\substack{\text{Sodium} \\ \text{(Metal)}}}{4Na\,(s)} + \underset{\substack{\text{Oxygen} \\ \text{(From air)}}}{O_2\,(g)} \longrightarrow \underset{\substack{\text{Sodium oxide} \\ \text{(Basic oxide)}}}{2Na_2O\,(s)}$$

Potassium metal (K) also reacts with the oxygen (O_2) of air at room temperature to form a basic oxide, called potassium oxide (K_2O). Please write the equation for this reaction yourself.

Potassium and sodium metals are so reactive that they react vigorously with the oxygen (of air). They catch fire and start burning when kept open in the air (see Figure 16). In fact, **potassium metal and sodium metal are stored under kerosene oil to prevent their reaction with the oxygen, moisture and carbon dioxide of air (so as to protect them).** Since potassium and sodium metals react with oxygen even at room temperature, therefore, potassium and sodium are very reactive metals.

Figure 16. Sodium burns in the oxygen (of air) to form sodium oxide.

Another metal which is very reactive is lithium (Li). Just like sodium and potassium metals, lithium metal is also stored under kerosene oil to prevent its reaction with oxygen, moisture and carbon dioxide of air (so as to protect it). Please note that lithium, sodium and potassium are all alkali metals (because they belong to a group of metals known as alkali metals).

Most of the metal oxides are insoluble in water. But some of the metal oxides dissolve in water to form alkalis. Sodium oxide and potassium oxide are the two metal oxides which are soluble in water. They dissolve in water to form alkalis. Sodium oxide and potassium oxide dissolve in water to form alkalis as follows :

Sodium oxide is a basic oxide which reacts with water to form an alkali called sodium hydroxide :

$$\underset{\substack{\text{Sodium oxide}\\\text{(Basic oxide)}}}{Na_2O\ (s)} + \underset{\text{Water}}{H_2O\ (l)} \longrightarrow \underset{\substack{\text{Sodium hydroxide}\\\text{(An alkali)}}}{2NaOH\ (aq)}$$

Due to the formation of sodium hydroxide alkali, a solution of sodium oxide in water turns red litmus to blue.

Potassium oxide is also a basic oxide which reacts with water to form an alkali called potassium hydroxide :

$$\underset{\substack{\text{Potassium oxide}\\\text{(Basic oxide)}}}{K_2O\ (s)} + \underset{\text{Water}}{H_2O\ (l)} \longrightarrow \underset{\substack{\text{Potassium hydroxide}\\\text{(An alkali)}}}{2KOH\ (aq)}$$

Due to the formation of potassium hydroxide alkali, a solution of potassium oxide in water turns red litmus to blue.

***(ii)* Magnesium metal** does not react with oxygen at room temperature. But on heating, magnesium metal burns in air giving intense heat and light to form a basic oxide called magnesium oxide (which is a white powder) :

$$\underset{\substack{\text{Magnesium}\\\text{(Metal)}}}{2Mg\ (s)} + \underset{\substack{\text{Oxygen}\\\text{(From air)}}}{O_2\ (g)} \longrightarrow \underset{\substack{\text{Magnesium oxide}\\\text{(Basic oxide)}}}{2MgO\ (s)}$$

Since heat is required for the reaction of magnesium with oxygen, it means magnesium is less reactive than sodium (or potassium).

Magnesium oxide dissolves in water partially to form magnesium hydroxide solution :

$$\underset{\text{Magnesium oxide}}{MgO\ (s)} + \underset{\text{Water}}{H_2O\ (l)} \longrightarrow \underset{\substack{\text{Magnesium hydroxide}\\\text{(A base)}}}{Mg(OH)_2\ (aq)}$$

This magnesium hydroxide turns red litmus solution to blue showing that it is a base and that magnesium oxide is basic in nature. We can perform the reaction of magnesium metal with oxygen (of air) as follows.

We take a magnesium ribbon, hold it with a pair of tongs and heat it over the flame of a burner. Magnesium ribbon burns vigorously in air producing a bright white light to form an ash (which is magnesium oxide). We put this magnesium oxide in a test-tube, add a little water and shake it. We will find that magnesium oxide dissolves in water partially. Let us divide this solution in two parts and test with blue litmus solution and red litmus solution, one by one. When blue litmus solution is added to magnesium oxide solution, there is no change in colour. On adding red litmus solution to magnesium oxide solution, the colour changes to blue. We know that only basic substances turn red litmus to blue. ***Since magnesium oxide solution turns red litmus to blue, it is basic in nature.***

(*iii*) Aluminium metal burns in air, on heating, to form aluminium oxide :

$4Al\ (s)$	$+$	$3O_2\ (g)$	$\longrightarrow$	$2Al_2O_3\ (s)$
Aluminium		Oxygen		Aluminium oxide
(Metal)		(From air)		(Amphoteric oxide)

Since the reaction of aluminium with oxygen takes place less readily than magnesium, so aluminium is less reactive than magnesium.

Figure 17. This is aluminium oxide. It is an amphoteric oxide.

Though most of the metal oxides are basic in nature but some of the metal oxides show basic as well as acidic nature. **Those metal oxides which show basic as well as acidic behaviour are known as amphoteric oxides.** Aluminium metal and zinc metal form amphoteric oxides. Thus, **aluminium oxide and zinc oxide are amphoteric in nature** (which show basic as well as acidic behaviour). **Amphoteric oxides react with both, acids as well as bases to form salts and water.** For example, aluminium oxide is an amphoteric oxide which reacts with acids as well as bases to form salt and water. This is described below.

(*a*) Aluminium oxide reacts with hydrochloric acid to form aluminium chloride (salt) and water :

$Al_2O_3\ (s)$	$+$	$6HCl\ (aq)$	$\longrightarrow$	$2AlCl_3\ (aq)$	$+$	$3H_2O\ (l)$
Aluminium oxide		Hydrochloric acid		Aluminium chloride		Water
		(Acid)		(Salt)		

In this reaction, aluminium oxide behaves as a basic oxide (because it reacts with an acid to form salt and water).

(*b*) Aluminium oxide reacts with sodium hydroxide to form sodium aluminate (salt) and water :

$Al_2O_3\ (s)$	$+$	$2NaOH\ (aq)$	$\longrightarrow$	$2NaAlO_2\ (aq)$	$+$	$H_2O\ (l)$
Aluminium oxide		Sodium hydroxide		Sodium aluminate		Water
		(Base)		(Salt)		

In this reaction, aluminium oxide behaves as an acidic oxide (because it reacts with a base to form salt and water).

(*iv*) Zinc metal burns in air only on strong heating to form zinc oxide :

$2Zn\ (s)$	$+$	$O_2\ (g)$	$\longrightarrow$	$2ZnO\ (s)$
Zinc		Oxygen		Zinc oxide
				(Amphoteric oxide)

Since the reaction of zinc with oxygen takes place less readily than aluminium, so zinc is less reactive than aluminium.

Zinc oxide is an amphoteric oxide which reacts with acids as well as with bases to form salt and water. This is described below.

(*a*) Zinc oxide reacts with hydrochloric acid to form zinc chloride (salt) and water :

$ZnO\ (s)$	$+$	$2HCl\ (aq)$	$\longrightarrow$	$ZnCl_2\ (aq)$	$+$	$H_2O\ (l)$
Zinc oxide		Hydrochloric acid		Zinc chloride		Water
		(Acid)		(Salt)		

In this reaction, zinc oxide behaves as a basic oxide (because it reacts with an acid to form salt and water).

(*b*) Zinc oxide reacts with sodium hydroxide to form sodium zincate (salt) and water :

ZnO (s)	+	$2NaOH$ (aq)	$\longrightarrow$	Na_2ZnO_2 (aq)	+	H_2O (l)
Zinc oxide		Sodium hydroxide (Base)		Sodium zincate (Salt)		Water

In this reaction, zinc oxide behaves as an acidic oxide (because it reacts with a base to form salt and water).

At ordinary temperature, the surfaces of the metals like magnesium, aluminium, zinc and lead, etc., are covered with a thin layer of their respective oxides. This oxide layer acts as a protective layer and prevents further oxidation (or corrosion) of the metal underneath.

(*v*) Iron metal does not burn in air even on strong heating. Iron reacts with the oxygen of air on heating to form iron (II, III) oxide :

$3Fe$ (s)	+	$2O_2$ (g)	$\longrightarrow$	Fe_3O_4 (s)
Iron		Oxygen		Iron (II, III) oxide

Thus, the reaction of iron with oxygen takes place less readily than that of zinc, so iron is less reactive than zinc. Please note that though a piece of iron metal does not burn in air but iron filings (small particles of iron) burn vigorously when sprinkled in the flame of a burner.

(*vi*) Copper metal also does not burn in air even on strong heating. Copper reacts with the oxygen of air on prolonged heating to form a black substance copper (II) oxide :

$2Cu$ (s)	+	O_2 (g)	$\longrightarrow$	$2CuO$ (s)
Copper		Oxygen		Copper (II) oxide

Since the reaction of copper with oxygen takes place even less readily than that of iron, so copper is less reactive than iron. Silver and gold metals do not react with oxygen even at high temperature, so they are still less reactive.

2. Reaction of Metals with Water

Metals react with water to form a metal hydroxide (or metal oxide) and hydrogen gas. All the metals, however, do not react with water. The intensity of reaction of a metal with water depends on its chemical reactivity. Some metals react even with cold water, some react with hot water, some react only with steam whereas others do not react even with steam (Steam is the gaseous form of water. It is very hot).

(*a*) When a **metal** reacts with **water** (cold water or hot water), then the products formed are **metal hydroxide** and **hydrogen gas** :

Metal + Water $\longrightarrow$ Metal hydroxide + Hydrogen

(*b*) When a **metal** reacts with **steam,** then the products formed are **metal oxide** and **hydrogen gas** :

Metal + Steam $\longrightarrow$ Metal oxide + Hydrogen

We will now describe the reactions of metals with water (or steam) by taking some examples. Potassium and sodium metals react violently even with cold water. For example :

(*i*) Potassium reacts violently with cold water to form potassium hydroxide and hydrogen gas :

$2K$ (s)	+	$2H_2O$ (l)	$\longrightarrow$	$2KOH$ (aq)	+	H_2 (g)	+	Heat
Potassium		Water (Cold)		Potassium hydroxide		Hydrogen		

Figure 18. Potassium metal reacts violently with water to form potassium hydroxide and hydrogen gas. So much heat is produced during this reaction that hydrogen gas formed catches fire and burns explosively.

The reaction of potassium metal with water is highly exothermic (heat producing) due to which the hydrogen gas formed during the reaction catches fire immediately. Thus, potassium is a very, very reactive metal.

***(ii)* Sodium** reacts vigorously with cold water forming sodium hydroxide and hydrogen gas :

$$\begin{array}{ccccccccc} 2Na\ (s) & + & 2H_2O\ (l) & \longrightarrow & 2NaOH\ (aq) & + & H_2\ (g) & + & \text{Heat} \\ \text{Sodium} & & \text{Water} & & \text{Sodium hydroxide} & & \text{Hydrogen} & & \\ & & \text{(Cold)} & & & & & & \end{array}$$

The reaction of sodium metal with water is also highly exothermic (heat producing) due to which the hydrogen gas formed during the reaction catches fire and burns causing little explosions. Thus, sodium is also a very reactive metal.

The reaction of sodium metal with water can be studied as follows : We cut a small piece of sodium metal carefully and dry it by using a filter paper. This piece of sodium metal is placed in water filled in a glass trough. We will find that the piece of sodium metal starts moving in water making a hissing sound and reacts with water causing little explosions. Soon the piece of sodium metal catches fire. This can be explained as follows. Sodium metal reacts with water to form sodium hydroxide and hydrogen gas. A lot of heat is also produced in this reaction. This heat burns the hydrogen gas as well as the sodium metal. The burning of hydrogen gas causes little explosions.

***(iii)* Calcium** reacts with cold water to form calcium hydroxide and hydrogen gas :

$$\begin{array}{ccccccc} Ca\ (s) & + & 2H_2O\ (l) & \longrightarrow & Ca(OH)_2\ (aq) & + & H_2\ (g) \\ \text{Calcium} & & \text{Water} & & \text{Calcium hydroxide} & & \text{Hydrogen} \\ & & \text{(Cold)} & & & & \end{array}$$

The heat produced in this reaction is less which is not sufficient to burn the hydrogen gas which is formed. The piece of calcium metal starts floating in water because the bubbles of hydrogen gas formed during the reaction stick to its surface. The reaction of calcium metal with water is less violent. So, calcium is less reactive than sodium.

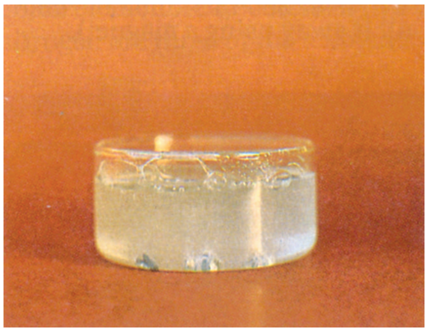


Figure 19. The calcium metal bubbles and fizzes in water to form calcium hydroxide solution and hydrogen gas. Much less heat is produced in this reaction due to which hydrogen gas formed does not catch fire and burn.

***(iv)* Magnesium metal** does not react with cold water. Magnesium reacts with hot water to form magnesium hydroxide and hydrogen :

$$\begin{array}{ccccccc} Mg\ (s) & + & 2H_2O\ (l) & \longrightarrow & Mg(OH)_2\ (aq) & + & H_2\ (g) \\ \text{Magnesium} & & \text{Water} & & \text{Magnesium hydroxide} & & \text{Hydrogen} \\ & & \text{(Hot)} & & & & \end{array}$$

In this reaction, the piece of magnesium metal starts floating in water due to the bubbles of hydrogen gas sticking to its surface. Calcium reacts with cold water but magnesium reacts only with hot water. This shows that magnesium is less reactive than calcium. We will now give the reaction of magnesium metal with steam.

Magnesium reacts very rapidly with steam to form magnesium oxide and hydrogen :

$$\begin{array}{ccccccc} Mg\ (s) & + & H_2O\ (g) & \longrightarrow & MgO\ (s) & + & H_2\ (g) \\ \text{Magnesium} & & \text{Steam} & & \text{Magnesium oxide} & & \text{Hydrogen} \end{array}$$

Please note that when magnesium reacts with hot water, it forms ***magnesium hydroxide*** and hydrogen. But when the same magnesium reacts with steam (at a much higher temperature), it forms ***magnesium oxide*** and hydrogen.

Metals like aluminium, zinc and iron do not react with either cold water or hot water. They react with steam to form a metal oxide and hydrogen. For example :

***(v)* Aluminium** reacts with steam to form aluminium oxide and hydrogen gas :

$$\begin{array}{ccccccc} 2Al\ (s) & + & 3H_2O\ (g) & \longrightarrow & Al_2O_3\ (s) & + & 3H_2\ (g) \\ \text{Aluminium} & & \text{Steam} & & \text{Aluminium oxide} & & \text{Hydrogen} \end{array}$$

Aluminium metal does not react with water under ordinary conditions because of the presence of a thin (but tough) layer of aluminium oxide on its surface.

(*vi*) Zinc reacts with steam to form zinc oxide and hydrogen :

$$\underset{\text{Zinc}}{Zn\ (s)} + \underset{\text{Steam}}{H_2O\ (g)} \longrightarrow \underset{\text{Zinc oxide}}{ZnO\ (s)} + \underset{\text{Hydrogen}}{H_2\ (g)}$$

(*vii*) Red-hot iron reacts with steam to form iron (II, III) oxide and hydrogen :

$$\underset{\text{Iron}}{3Fe\ (s)} + \underset{\text{Steam}}{4H_2O\ (g)} \longrightarrow \underset{\text{Iron (II, III) oxide}}{Fe_3O_4\ (s)} + \underset{\text{Hydrogen}}{4H_2\ (g)}$$

We can study the reaction of metals (like magnesium, aluminium, zinc and iron) with steam by using the apparatus shown in Figure 20.

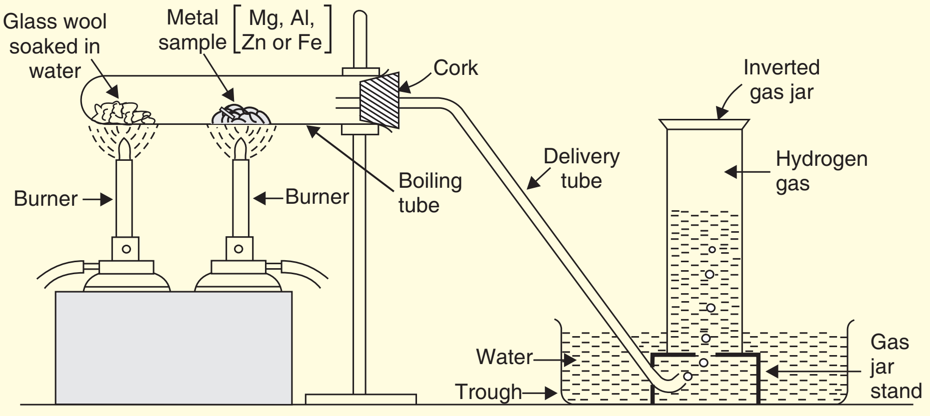


Figure 20. Experimental set-up to study the reaction of metals with steam.

A lump of glass wool soaked in water is placed at the bottom of a boiling tube (see Figure 20). The water present in glass wool will form steam on heating (but glass wool itself does not take part in the reaction). The metal sample (to be reacted with steam) is placed in the middle of the horizontally kept boiling tube. The boiling tube containing water soaked glass wool and metal sample is then arranged in the apparatus as shown in Figure 20.

To start the experiment, the metal sample is heated by using a burner. When the metal gets hot, then the glass wool is heated by using another burner. The water present in glass wool forms steam on heating. This steam then passes over the hot metal. The hot metal reacts with steam to form the corresponding metal oxide and hydrogen gas. The hydrogen gas comes out of the boiling tube and it is collected over water as shown in Figure 20. When a lighted match stick is applied to the gas collected in the gas jar, the gas burns with a 'pop' sound (making a little explosion), indicating that it is hydrogen (This is a dangerous test and should be performed carefully with the help of your teacher). The metal oxide formed remains behind in the boiling tube.

This experiment is performed by taking magnesium, aluminium, zinc, and iron as metal samples, one by one. It is found that the reaction of steam with magnesium is the most vigorous followed by the reactions with aluminium and zinc; but the reaction with iron is very slow. This shows that out of magnesium, aluminium, zinc and iron : magnesium is the most reactive whereas iron is the least reactive. On the basis of the vigour of their reaction with steam, we can arrange magnesium, aluminium, zinc and iron metals in the decreasing order of their reactivity as : Mg > Al > Zn > Fe. **Metals like lead, copper, silver and gold do not react with water (or even steam).**

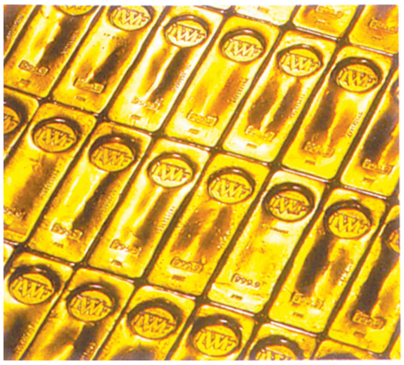


Figure 21. This picture shows gold metal in the form of 'gold biscuits'. Gold metal is very unreactive. It does not react with water or even steam.

We will now explain how metals displace hydrogen from water. Water (H_2O) is slightly ionised to give hydrogen ions (H^+) and hydroxide ions (OH^-). Now, when a reactive metal combines with water, it gives electrons to reduce the hydrogen ions of water to hydrogen atoms, which then form hydrogen gas. The unreactive metals like copper do not give electrons easily, so they are not able to reduce the hydrogen ions of water to hydrogen gas. Hence, unreactive metals like copper do not displace hydrogen from water. Please note that **only those metals displace hydrogen from water (or steam) which are above hydrogen in the reactivity series.**

3. Reaction of Metals with Dilute Acids

Metals usually displace hydrogen from dilute acids. Only the less reactive metals like copper, silver and gold do not displace hydrogen from dilute acids. When a metal reacts with a dilute acid, then a metal salt and hydrogen gas are formed :

$$\textbf{Metal} + \textbf{Dilute acid} \longrightarrow \textbf{Metal salt} + \textbf{Hydrogen}$$

All the metals, however, do not react with dilute acids.

The vigour of reaction of a metal with dilute acid depends on its chemical reactivity. Some metals react explosively (extremely rapidly) with dilute acids, some metals react rapidly, some metals react only on heating whereas others do not react at all. We will first describe the reactions of metals with dilute hydrochloric acid.

Metals react with dilute hydrochloric acid to give metal chlorides and hydrogen gas. The reactions of metals with dilute hydrochloric acid are given below :

(*i*) Sodium metal reacts violently with dilute hydrochloric acid to form sodium chloride and hydrogen :

$$\underset{\text{Sodium}}{2Na\,(s)} + \underset{\text{Hydrochloric acid}}{2HCl\,(aq)} \longrightarrow \underset{\text{Sodium chloride}}{2NaCl\,(aq)} + \underset{\text{Hydrogen}}{H_2\,(g)}$$

This reaction shows that sodium metal is very reactive.

(*ii*) Magnesium reacts quite rapidly with dilute hydrochloric acid forming magnesium chloride and hydrogen gas :

$$\underset{\text{Magnesium}}{Mg\,(s)} + \underset{\text{Hydrochloric acid}}{2HCl\,(aq)} \longrightarrow \underset{\text{Magnesium chloride}}{MgCl_2\,(aq)} + \underset{\text{Hydrogen}}{H_2\,(g)}$$

The reaction of magnesium with dilute hydrochloric acid is less vigorous than that of sodium, so magnesium is less reactive than sodium.

(*iii*) Aluminium metal at first reacts slowly with dilute hydrochloric acid due to the presence of a tough protective layer of aluminium oxide on its surface. But when the thin, outer oxide layer gets dissolved in acid, then fresh aluminium metal is exposed which reacts rapidly with dilute hydrochloric acid. Thus :

Aluminium metal reacts rapidly with dilute hydrochloric acid to form aluminium chloride and hydrogen gas :

$$\underset{\text{Aluminium}}{2Al\,(s)} + \underset{\text{Hydrochloric acid}}{6HCl\,(aq)} \longrightarrow \underset{\text{Aluminium chloride}}{2AlCl_3\,(aq)} + \underset{\text{Hydrogen}}{3H_2\,(g)}$$

The reaction of aluminium with dilute hydrochloric acid is less rapid than that of magnesium, so aluminium is less reactive than magnesium.

(*iv*) Zinc reacts with dilute hydrochloric acid to give zinc chloride and hydrogen gas (but the reaction is less rapid than that of aluminium) :

$$\underset{\text{Zinc}}{Zn\,(s)} + \underset{\text{Hydrochloric acid}}{2HCl\,(aq)} \longrightarrow \underset{\text{Zinc chloride}}{ZnCl_2\,(aq)} + \underset{\text{Hydrogen}}{H_2\,(g)}$$

This reaction shows that zinc is less reactive than aluminium.

(*v*) **Iron** reacts slowly with cold dilute hydrochloric acid to give iron (II) chloride and hydrogen gas :

$$\underset{\text{Iron}}{Fe\,(s)} + \underset{\text{Hydrochloric acid}}{2HCl\,(aq)} \longrightarrow \underset{\text{Iron (II) chloride}}{FeCl_2\,(aq)} + \underset{\text{Hydrogen}}{H_2\,(g)}$$

This shows that iron is less reactive than zinc.

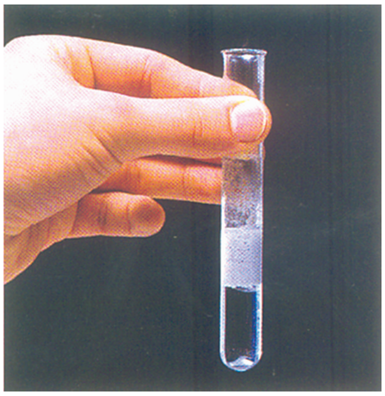


Figure 22. Magnesium metal reacts with dilute hydrochloric acid giving off hydrogen gas.

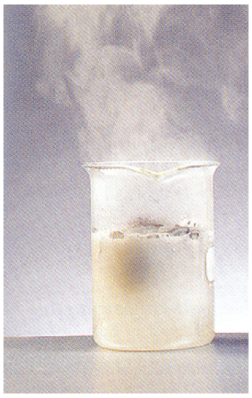


Figure 23. Aluminium metal reacts with dilute hydrochloric acid producing hydrogen gas

Figure 24. Copper metal does not react with dilute hydrochloric acid to give hydrogen gas.

(*vi*) **Copper** does not react with dilute hydrochloric acid (or dilute sulphuric acid) at all. This shows that copper is even less reactive than iron :

$$\underset{\text{Copper}}{Cu\,(s)} + \underset{\substack{\text{Hydrochloric acid}\\\text{(Dilute)}}}{HCl\,(aq)} \longrightarrow \text{No reaction}$$

We will now describe an experiment to show the relative reactivities of some metals with a dilute acid. We take small pieces of magnesium, aluminium, zinc, iron and copper metals and clean their surfaces by rubbing with a sand paper. Place these metal pieces in separate test-tubes and add equal volume of 10 mL of dilute hydrochloric acid to each test-tube. Observe the rate of formation of hydrogen gas bubbles carefully. We will find that the formation of bubbles of hydrogen is fastest in the case of magnesium showing that magnesium is the most reactive metal here. The rate of formation of hydrogen gas bubbles decreases in the order Magnesium > Aluminium > Zinc > Iron, showing the decreasing chemical reactivity of these metals with dilute hydrochloric acid. But no hydrogen gas bubbles are formed in the test-tube containing copper metal and dilute hydrochloric acid. This shows that copper does not react with dilute hydrochloric acid and hence it is the least reactive out of these metals. **Silver and gold metals also do not react with dilute acids.**

We will now discuss how metals displace hydrogen from dilute acids. All those metals which are more reactive than hydrogen, that is, those metals which lose electrons more easily than hydrogen, displace hydrogen from dilute acids to produce hydrogen gas. This is due to the fact that the more reactive metals give electrons easily and these electrons reduce the hydrogen ions of acids to hydrogen gas. **The metals like copper and silver which are less reactive than hydrogen, do not displace hydrogen from dilute acids.** Because they do not give out electrons required for the reduction of hydrogen ions present in acids. Thus, **all the metals which are above hydrogen in the activity series, displace hydrogen from dilute acids** (like dil. HCl and dil. H_2SO_4). Those metals which are below hydrogen in the activity series, do not displace hydrogen from dilute acids.

We will now give the reactions of metals with dilute sulphuric acid. **Metals react with dilute sulphuric acid to give metal sulphates and hydrogen gas.** The equations for the reactions of sodium, magnesium, aluminium, zinc and iron metals with dilute sulphuric acid are given on the next page.

2Na (s)	+	H_2SO_4 (aq)	$\longrightarrow$	Na_2SO_4 (aq)	+	H_2 (g)
Sodium		Sulphuric acid		Sodium sulphate		Hydrogen
Mg (s)	+	H_2SO_4 (aq)	$\longrightarrow$	$MgSO_4$ (aq)	+	H_2 (g)
Magnesium		Sulphuric acid		Magnesium sulphate		Hydrogen
2Al (s)	+	$3H_2SO_4$ (aq)	$\longrightarrow$	$Al_2(SO_4)_3$ (aq)	+	$3H_2$ (g)
Aluminium		Sulphuric acid		Aluminium sulphate		Hydrogen
Zn (s)	+	H_2SO_4 (aq)	$\longrightarrow$	$ZnSO_4$ (aq)	+	H_2 (g)
Zinc		Sulphuric acid		Zinc sulphate		Hydrogen
Fe (s)	+	H_2SO_4 (aq)	$\longrightarrow$	$FeSO_4$ (aq)	+	H_2 (g)
Iron		Sulphuric acid		Iron (II) sulphate		Hydrogen
Cu (s)	+	H_2SO_4 (aq)	$\longrightarrow$	No reaction		
Copper		Sulphuric acid (Dilute)				

We will now discuss the reactions of metals with dilute nitric acid. **When a metal reacts with dilute nitric acid, then hydrogen gas is *not* evolved.** This can be explained as follows : Nitric acid is a strong oxidising agent. So, **as soon as hydrogen gas is formed in the reaction between a metal and dilute nitric acid, the nitric acid oxidises this hydrogen to water.** So, in the reactions of metals with dilute nitric acid, no hydrogen gas is evolved. Now, when nitric acid oxidises hydrogen to water, then nitric acid itself is reduced to any of the nitrogen oxides (such as dinitrogen monoxide, N_2O; nitrogen monoxide, NO; or nitrogen dioxide, NO_2). The type of oxide formed depends on the nature of metal, the temperature of reaction and concentration of nitric acid.

Very dilute nitric acid, however, reacts with magnesium and manganese metals to evolve hydrogen gas. This is because the very dilute nitric acid is a weak oxidising agent which is not able to oxidise hydrogen to water. The reactions of magnesium and manganese metals with very dilute nitric acid are given below.

(*a*) Magnesium reacts with very dilute nitric acid to form magnesium nitrate and hydrogen gas :

Mg (s)	+	$2HNO_3$ (aq)	$\longrightarrow$	$Mg(NO_3)_2$ (aq)	+	H_2 (g)
Magnesium		Nitric acid (Very dilute)		Magnesium nitrate		Hydrogen

(*b*) Manganese reacts with very dilute nitric acid to form manganese nitrate and hydrogen gas :

Mn (s)	+	$2HNO_3$ (aq)	$\longrightarrow$	$Mn(NO_3)_2$ (aq)	+	H_2 (g)
Manganese		Nitric acid (Very dilute)		Manganese nitrate		Hydrogen

Aqua-Regia

Aqua-regia is a freshly prepared mixture of 1 part of concentrated nitric acid and 3 parts of concentrated hydrochloric acid. Thus, the ratio of conc. HNO_3 and conc. HCl in aqua-regia is 1 : 3. Aqua-regia is a highly corrosive, fuming liquid (Corrosive means which can cause corrosion). **Aqua-regia can dissolve all metals.** For example, **aqua-regia can dissolve even gold and platinum metals** (though concentrated nitric acid or concentrated hydrochloric acid alone cannot dissolve gold or platinum metals). Let us solve one problem now.

Figure 25. This is aqua-regia.

Sample Problem. Between copper and sodium, which metal is more reactive ? Explain with reasons.

Solution. Sodium metal is more reactive than copper, because :

(*i*) Sodium reacts with oxygen easily to form sodium oxide but copper does not react with oxygen easily.

(*ii*) Sodium reacts vigorously with cold water to form sodium hydroxide and hydrogen but copper does not react even with steam.

(*iii*) Sodium reacts rapidly with dilute hydrochloric acid to form sodium chloride and hydrogen, but copper does not react with dilute hydrochloric acid.

Before we describe the reactions of metals with salt solutions, we should know the meaning of 'the reactivity series of metals'. This is known as reactivity series of metals because it tells us the relative chemical reactivities of metals towards other elements. Please note that the 'reactivity series of metals' is also known as 'activity series of metals'. So, let us now discuss the reactivity series of metals or the activity series of metals.

The Reactivity Series of Metals (or Activity Series of Metals)

Some metals are chemically very reactive whereas others are less reactive or unreactive. For example, potassium and sodium react very, very rapidly with cold water, so they are very reactive metals. Zinc and iron react only with steam, so they are less reactive metals. On the other hand, copper and silver do not react even with steam, so they are quite unreactive metals. On the basis of vigour of reactions of various metals with oxygen, water and acids, as well as displacement reactions, the metals have been arranged in a group or series according to their chemical reactivity. **The arrangement of metals in a vertical column in the order of decreasing reactivities is called reactivity series of metals (or activity series of metals).** In reactivity series, the most reactive metal is placed at the top whereas the least reactive metal is placed at the bottom. The reactivity series of the common metals is given below.

Figure 26. Potassium is the most reactive metal, so it has been placed at the top in the reactivity series.

Reactivity Series (or Activity Series) of Metals

These metals are more reactive than hydrogen	**Potassium**	**K**	**(Most reactive metal)**
	Sodium	Na	↓
	Calcium	Ca	
	Magnesium	Mg	
	Aluminium	Al	
	Zinc	Zn	
	Iron	Fe	Decreasing chemical reactivity
	Tin	Sn	
	Lead	Pb	
	[Hydrogen]	**[H]**	
These metals are less reactive than hydrogen	Copper	Cu	
	Mercury	Hg	
	Silver	Ag	↓
	Gold	Au	**(Least reactive metal)**

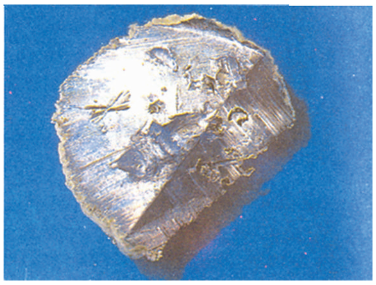


Figure 27. Sodium metal is less reactive than potassium, so sodium has been placed below potassium in the reactivity series.

Please note that potassium is the most reactive metal here, so it has been placed at the top in the reactivity series. As we come down in the series the chemical reactivity of metals decreases. Gold being least reactive metal has been placed at the bottom in the series. Since the metals placed at the bottom of the reactivity series (like silver and gold) are less reactive, so they are usually found in free state (native state) in nature. Though hydrogen is not a metal but even then it has been placed in the reactivity series of metals. This is due to the fact that like metals, hydrogen also loses electrons and forms positive ions, H^+.

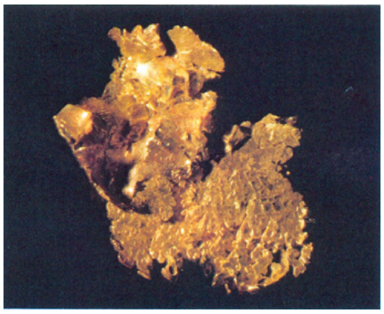


Figure 28. Gold being the least reactive metal, has been placed at the bottom of the reactivity series.

Why Some Metals are More Reactive and Others Less Reactive

We have just seen that some metals are more reactive and others are less reactive. Let us now find out the reason for this difference in the chemical reactivities of metals.

When metals react, they lose electrons to form positive ions. Now, if a metal atom can lose electrons easily to form positive ions, it will react rapidly with other substances and hence it will be a reactive metal. On the other hand, if a metal atom loses electrons less readily to form positive ions, it will react slowly with other substances. Such a metal will be less reactive. For example, sodium atoms lose electrons readily to form sodium ions, due to which sodium metal is very reactive. On the other hand, iron atoms lose electrons less readily to form positive ions, so iron metal is less reactive.

Metals Which are More Reactive Than Hydrogen

Those metals which lose electrons more readily than hydrogen are said to be more reactive than hydrogen. All the metals which have been placed above hydrogen in the reactivity series, lose electrons more readily than hydrogen, and hence they are more reactive than hydrogen. Thus, **the metals which are more reactive than hydrogen are: Potassium, Sodium, Calcium, Magnesium, Aluminium, Zinc, Iron, Tin and Lead.** These more reactive metals can displace hydrogen from its compounds like water and acids to form hydrogen gas.

Metals Which are Less Reactive Than Hydrogen

Those metals which lose electrons less readily than hydrogen are said to be less reactive than hydrogen. All the metals placed below hydrogen in the reactivity series lose electrons less readily than hydrogen, and hence they are less reactive than hydrogen. Thus, **the metals which are less reactive than hydrogen are: Copper, Mercury, Silver and Gold.** These less reactive metals cannot displace hydrogen from its compounds like water and acids to form hydrogen gas.

From this discussion we conclude that : If a metal is above hydrogen in the activity series, then it will displace hydrogen from water or acids, that is, it will react with water and acids to produce hydrogen gas. On the other hand, if a metal is below hydrogen in the activity series, then it will not displace hydrogen from water and acids, that is, it will not react with water and acids to produce hydrogen gas. **We should remember the reactivity series of metals to decide whether a particular displacement reaction will take place or not.**

4. Reaction of Metals with Salt Solutions

When a more reactive metal is put in the salt solution of a less reactive metal, then the more reactive metal displaces (pushes out) the less reactive metal from its salt solution. In other words : **A more reactive metal displaces a less reactive metal from its salt solution.** The more reactive metal takes the place of less reactive metal and forms its own salt solution. For example, if metal A is more reactive than metal B, then metal A will displace metal B from its salt solution to form salt solution of metal A, and metal B will be set free. That is :

Salt solution of metal B + Metal A ⟶ Salt solution of metal A + Metal B

Let us take some examples to make this point more clear.

(*i*) The Reaction of Zinc with Copper Sulphate Solution. When a strip of zinc metal is put in copper sulphate solution, then the blue colour of copper sulphate solution fades gradually due to the formation of colourless zinc sulphate solution, and red-brown copper metal is deposited on the zinc strip :

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

In this reaction zinc metal is displacing copper metal from its salt solution, copper sulphate solution. This displacement reaction occurs because zinc is more reactive than copper. **If, however, a strip of copper metal is placed in zinc sulphate solution, then no reaction occurs.** This is because copper metal is less reactive than zinc metal and hence cannot displace zinc from zinc sulphate solution.

If we put silver metal in copper sulphate solution, even then no reaction takes place. This is because silver metal is less reactive than copper metal and hence cannot displace copper from copper sulphate solution. Iron and magnesium metals are, however, more reactive than copper metal, so they can displace copper from copper sulphate solution.

(*ii*) Reaction of Iron with Copper Sulphate Solution. When a strip of iron metal (or iron nail) is placed in copper sulphate solution, then the blue colour of copper sulphate solution fades gradually and red-brown copper metal is formed :

$CuSO_4$ (aq)	+	Fe (s)	$\longrightarrow$	$FeSO_4$ (aq)	+	Cu (s)
Copper (II) sulphate		Iron		Iron (II) sulphate		Copper
(Blue solution)		(Grey)		(Greenish solution)		(Red-brown)

The copper metal produced in this reaction forms a red-brown layer on the iron strip (or iron nail) (see Figure 29). In this reaction, iron is displacing copper from copper sulphate solution. This displacement occurs because iron is more reactive than copper. **If, however, a strip of copper metal is placed in iron (II) sulphate solution, then no reaction occurs.** This is because copper is less reactive than iron and hence cannot displace iron from iron (II) sulphate solution.

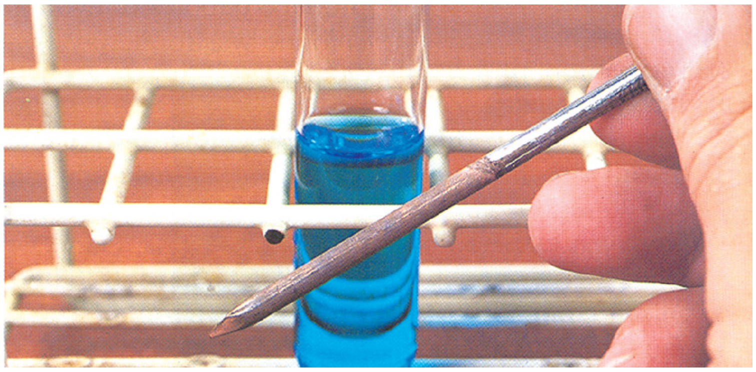


Figure 29. This iron nail was left in copper sulphate solution. It displaced copper from copper sulphate solution. The displaced copper forms a red-brown coating on the surface of iron nail.

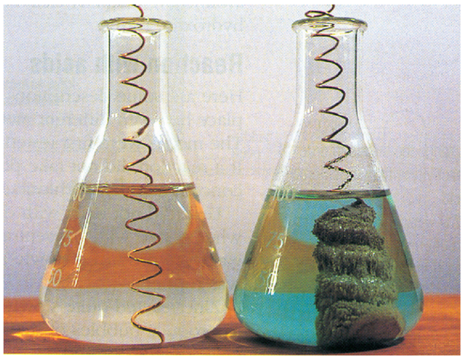


Figure 30. A copper wire coil has been placed in silver nitrate solution (left). Copper displaces silver from silver nitrate solution. This silver is deposited on the copper wire coil (right).

(*iii*) Reaction of Copper with Silver Nitrate Solution. When a strip of copper metal is kept immersed in silver nitrate solution for some time, the solution gradually becomes blue and a shining greyish-white deposit of silver metal is formed on copper strip :

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

In this reaction, copper metal is displacing silver from silver nitrate solution forming copper nitrate and silver metal. The solution becomes blue due to the formation of copper nitrate. Please note that this displacement occurs because copper is more reactive than silver. **If, however, we place a strip of silver metal in copper nitrate solution (or copper sulphate solution) then no reaction occurs.** This is because silver is less reactive than copper and hence cannot displace copper from copper nitrate solution (or copper sulphate solution). Before we go further, let us solve some problems now.

Sample Problem 1. In a solution of silver nitrate, a copper plate was dipped. After some time, silver from the solution was deposited on the copper plate. Which metal is more reactive—copper or silver ? How ?

Solution. We know that a more reactive metal displaces a less reactive metal from its salt solution. Here, copper metal is displacing silver from silver nitrate solution (which then gets deposited on copper plate), therefore, copper metal is more reactive than silver metal.

Sample Problem 2. A solution of $CuSO_4$ was kept in an iron pot. After a few days, the iron pot was found to have a number of holes in it. Write the equation of the reaction that took place. Explain this reaction in terms of reactivity.

Solution. We know that iron metal is more reactive than copper metal. So, when a solution of copper sulphate ($CuSO_4$) was kept in an iron pot, then iron being more reactive displaced copper of copper sulphate solution to form copper metal and iron (II) sulphate solution. The equation for this displacement reaction can be written as :

$$\underset{\substack{\text{Copper (II) sulphate} \\ \text{(Blue solution)}}}{CuSO_4\,(aq)} + \underset{\substack{\text{Iron metal} \\ \text{(From iron pot)}}}{Fe\,(s)} \longrightarrow \underset{\substack{\text{Iron (II) sulphate} \\ \text{(Greenish solution)}}}{FeSO_4\,(aq)} + \underset{\text{Copper metal}}{Cu\,(s)}$$

Since the iron metal taking part in this displacement reaction is being taken from the iron pot, so holes are formed at those places in the iron pot from where iron metal has dissolved to form iron (II) sulphate.

Sample Problem 3. What would you observe when zinc is added to a solution of iron (II) sulphate ? Write the chemical reaction that takes place. **(NCERT Book Question)**

Solution. When zinc is added to a solution of iron (II) sulphate, then the greenish colour of iron (II) sulphate solution fades gradually due to the formation of colourless zinc sulphate solution, and iron metal is deposited on zinc :

$$\underset{\substack{\text{Iron (II) sulphate} \\ \text{(Greenish solution)}}}{FeSO_4\,(aq)} + \underset{\substack{\text{Zinc} \\ \text{(Silvery-white)}}}{Zn\,(s)} \longrightarrow \underset{\substack{\text{Zinc sulphate} \\ \text{(Colourless solution)}}}{ZnSO_4\,(aq)} + \underset{\substack{\text{Iron} \\ \text{(Grey)}}}{Fe\,(s)}$$

Sample Problem 4. Which of the following pairs will give displacement reactions ?

(*a*) NaCl solution and copper metal

(*b*) $MgCl_2$ solution and aluminium metal.

(*c*) $FeSO_4$ solution and silver metal.

(*d*) $AgNO_3$ solution and copper metal. **(NCERT Book Question)**

Solution. (*a*) Copper metal is less reactive than sodium metal (Na), so no displacement reaction will occur between NaCl solution and copper metal.

(*b*) Aluminium metal is less reactive than magnesium metal (Mg), so no displacement reaction will take place between $MgCl_2$ solution and aluminium metal.

(*c*) Silver metal is less reactive than iron metal (Fe), so no displacement reaction will occur between $FeSO_4$ solution and silver metal.

(*d*) Copper metal is more reactive than silver metal (Ag), so a displacement reaction will take place between $AgNO_3$ solution and copper metal.

Sample Problem 5. Zinc oxide, magnesium oxide and copper oxide were heated, turn by turn, with zinc, magnesium and copper metals as shown in the following table :

Metal oxide	Zinc	Magnesium	Copper
1. Zinc oxide 2. Magnesium oxide 3. Copper oxide			

In which cases will you find displacement reactions taking place ? **(NCERT Book Question)**

Solution. We know that a more reactive metal can displace a less reactive metal from its oxide. Keeping in mind that out of zinc, magnesium and copper metals, magnesium is the most reactive, zinc is less reactive whereas copper is the least reactive metal, we will find that the displacement reactions will take place in the following cases :

Metal oxide	Zinc	Magnesium	Copper
1. Zinc oxide	—	Displacement	—
2. Magnesium oxide	—	—	—
3. Copper oxide	Displacement	Displacement	—

Sample Problem 6. Samples of four metals *A*, *B*, *C* and *D* were taken and added to the solutions given in the following table, one by one. The results obtained are 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 above table to answer the following questions about metals *A*, *B*, *C* and *D* :

(*i*) Which is the most reactive metal ?

(*ii*) What would you observe when metal *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. (**NCERT book Question**)

Solution. (*i*) *B* is the most reactive metal [because it gives displacement reaction with iron (II) sulphate].

(*ii*) When metal *B* is added to copper (II) sulphate solution, a displacement reaction will take place due to which the blue colour of copper (II) sulphate solution will fade and a red-brown deposit of copper will be formed on metal *B*.

(*iii*) Metal *B* is the most reactive (because it displaces iron from its salt solution) ; metal *A* is less reactive (because it displaces copper from its salt solution) ; metal *C* is still less reactive (because it can displace only silver from its salt solution); and metal *D* is the least reactive (because it cannot displace any metal from its salt solution). So, the decreasing order of reactivity of the metals is : $B > A > C > D$.

Please note that metal *B* is like zinc (Zn), metal *A* is like iron (Fe), metal *C* is like copper (Cu) whereas metal *D* is like silver (Ag).

5. Reaction of Metals with Chlorine

Metals react with chlorine to form ionic chlorides. In the formation of metal chlorides, the metal atoms lose electrons and become positively charged ions, whereas chlorine atoms gain electrons (given by metal atoms) and become negatively charged chloride ions. In other words, metals form ionic chlorides because they can give electrons to chlorine atoms to form ions. Metal chlorides are usually solid and conduct electricity in solution or in molten state. Thus, metal chlorides are electrolytes. Metal chlorides have high melting points and boiling points. So, metal chlorides are non-volatile. Here are some examples.

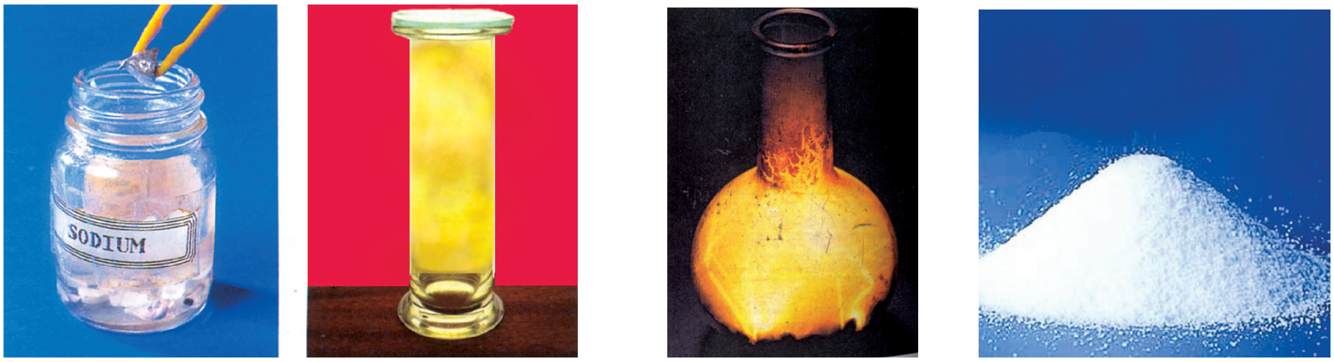
(*i*) Sodium is a metal. So, sodium readily reacts with chlorine to form an ionic chloride called sodium chloride :

$$\underset{\substack{\text{Sodium}\\\text{(A metal)}}}{2Na\ (s)} + \underset{\text{Chlorine}}{Cl_2\ (g)} \longrightarrow \underset{\substack{\text{Sodium chloride}\\\text{(Ionic chloride)}}}{2NaCl\ (s)}$$

Sodium chloride (NaCl) is an ionic compound or electrovalent compound containing sodium ions, Na^+, and chloride ions, Cl^- ($NaCl = Na^+Cl^-$). Sodium chloride solution conducts electricity. It is an electrolyte.

(*ii*) Calcium is a metal which reacts vigorously with chlorine to form an ionic chloride called calcium chloride :

$$\underset{\text{Calcium}}{Ca\ (s)} + \underset{\text{Chlorine}}{Cl_2\ (g)} \longrightarrow \underset{\text{Calcium chloride}}{CaCl_2\ (s)}$$



(*a*) Sodium metal (*b*) Chlorine gas (*c*) Sodium metal reacting with chlorine gas (*d*) Sodium chloride is formed

Figure 31. Reaction of sodium metal with chlorine.

(*iii*) Magnesium on heating with chlorine readily forms magnesium chloride, which is an ionic chloride :

$$\begin{array}{ccccc} Mg\ (s) & + & Cl_2\ (g) & \longrightarrow & MgCl_2\ (s) \\ \text{Magnesium} & & \text{Chlorine} & & \text{Magnesium chloride} \end{array}$$

(*iv*) Aluminium reacts with chlorine, on heating, to form aluminium chloride :

$$\begin{array}{ccccc} 2Al\ (s) & + & 3Cl_2\ (g) & \longrightarrow & 2AlCl_3\ (s) \\ \text{Aluminium} & & \text{Chlorine} & & \text{Aluminium chloride} \end{array}$$

(*v*) Zinc combines directly with chlorine to form zinc chloride :

$$\begin{array}{ccccc} Zn\ (s) & + & Cl_2\ (g) & \longrightarrow & ZnCl_2\ (s) \\ \text{Zinc} & & \text{Chlorine} & & \text{Zinc chloride} \end{array}$$

(*vi*) Iron combines with chlorine, when heated, to form iron (III) chloride :

$$\begin{array}{ccccc} 2Fe\ (s) & + & 3Cl_2\ (g) & \longrightarrow & 2FeCl_3\ (s) \\ \text{Iron} & & \text{Chlorine} & & \text{Iron (III) chloride} \\ & & & & \text{(or Ferric chloride)} \end{array}$$

(*vii*) On heating, **copper** reacts with chlorine to form copper (II) chloride :

$$\begin{array}{ccccc} Cu\ (s) & + & Cl_2\ (g) & \longrightarrow & CuCl_2\ (s) \\ \text{Copper} & & \text{Chlorine} & & \text{Copper (II) chloride} \\ & & & & \text{(or Cupric chloride)} \end{array}$$

All these metal chlorides are ionic compounds (or electrovalent compounds).

6. Reaction of Metals with Hydrogen

Metals generally do not react with hydrogen because metals form compounds by losing electrons (which are accepted by other elements) and hydrogen also forms compounds by losing electron (or by sharing of electrons). So, normally a hydrogen atom does not accept the electrons given by a metal atom to form a compound. But a few very reactive metals (like sodium, potassium, calcium and magnesium) can force the hydrogen atoms to accept electrons given by them and form salt-like solid compounds called metal hydrides. Thus, **most of the metals do not combine with hydrogen. Only a few reactive metals like sodium, potassium, calcium and magnesium react with hydrogen to form metal hydrides.** Metal hydrides are ionic compounds formed by the transfer of electrons from metal atoms to hydrogen atoms. In a metal hydride, the hydrogen is present in the form of a negative ion (anion) called hydride ion, H^-. Here are some examples.

(*i*) When hydrogen gas is passed over heated sodium, then sodium hydride is formed :

$$\begin{array}{ccccc} 2Na\ (s) & + & H_2\ (g) & \longrightarrow & 2NaH\ (s) \\ \text{Sodium} & & \text{Hydrogen} & & \text{Sodium hydride} \\ \text{(A metal)} & & & & \text{(Ionic hydride)} \end{array}$$

Sodium hydride, NaH, is an ionic compound containing sodium ions, Na^+, and hydride ions, H^-. When hydrogen gas is passed over heated potassium, then potassium hydride (KH) is formed. Write the equation for this reaction yourself. Potassium hydride is also an ionic hydride.

(*ii*) When hydrogen gas is passed over heated calcium, then calcium hydride is formed :

$$\underset{\substack{\text{Calcium}\\\text{(A metal)}}}{Ca\ (s)} + \underset{\text{Hydrogen}}{H_2\ (g)} \longrightarrow \underset{\substack{\text{Calcium hydride}\\\text{(Ionic hydride)}}}{CaH_2\ (s)}$$

Calcium hydride is an ionic compound containing calcium ions, Ca^{2+}, and hydride ions, $2H^-$. Similarly, when hydrogen gas is passed over heated magnesium, then magnesium hydride (MgH_2) is formed. Write the equation for this reaction yourself. Magnesium hydride is also an ionic hydride. The comparatively less reactive metals like zinc, copper and iron do not react with hydrogen to form hydrides.

CHEMICAL PROPERTIES OF NON-METALS

The important chemical properties of non-metals are given below :

1. Reaction of Non-Metals with Oxygen

Non-metals react with oxygen to form acidic oxides or neutral oxides. Carbon forms an acidic oxide CO_2, sulphur forms an acidic oxide SO_2, and hydrogen forms a neutral oxide, H_2O. The non-metal oxides are covalent in nature which are formed by the sharing of electrons. **The acidic oxides of non-metals dissolve in water to form acids.** The acidic oxides of non-metals turn blue litmus solution to red. Here are some examples.

(*i*) Carbon is a non-metal. When carbon burns in air it reacts with the oxygen of air to form an acidic oxide called carbon dioxide :

$$\underset{\substack{\text{Carbon}\\\text{(Non-metal)}}}{C\ (s)} + \underset{\substack{\text{Oxygen}\\\text{(From air)}}}{O_2\ (g)} \longrightarrow \underset{\substack{\text{Carbon dioxide}\\\text{(Acidic oxide)}}}{CO_2\ (g)}$$

The acidic oxide, carbon dioxide, dissolves in water to form an acid called carbonic acid :

$$\underset{\substack{\text{Carbon dioxide}\\\text{(Acidic oxide)}}}{CO_2\ (g)} + \underset{\text{Water}}{H_2O\ (l)} \longrightarrow \underset{\substack{\text{Carbonic acid}\\\text{(An acid)}}}{H_2CO_3\ (aq)}$$

A solution of carbon dioxide gas in water turns blue litmus to red, showing that it is acidic in nature.

(*ii*) Sulphur is a non-metal. When sulphur is burned in air, it reacts with the oxygen of air to form an acidic oxide called sulphur dioxide :

$$\underset{\substack{\text{Sulphur}\\\text{(Non-metal)}}}{S\ (s)} + \underset{\substack{\text{Oxygen}\\\text{(From air)}}}{O_2\ (g)} \longrightarrow \underset{\substack{\text{Sulphur dioxide}\\\text{(Acidic oxide)}}}{SO_2\ (g)}$$

Figure 32. This is sulphur. Sulphur burns in air to form an acidic gas sulphur dioxide.

The acidic oxide, sulphur dioxide, dissolves in water to form an acid called sulphurous acid :

$$\underset{\substack{\text{Sulphur dioxide}\\\text{(Acidic oxide)}}}{SO_2\ (g)} + \underset{\text{Water}}{H_2O\ (l)} \longrightarrow \underset{\substack{\text{Sulphurous acid}\\\text{(An acid)}}}{H_2SO_3\ (aq)}$$

A solution of sulphur dioxide in water turns blue litmus to red. This shows that sulphur dioxide is acidic in nature. We can perform the reaction of sulphur with oxygen of air as follows. Please note that sulphur is a *yellow* solid (see Figure 32).

We take a small amount of sulphur powder in a deflagrating spoon (combustion spoon) and heat it over the flame of a burner [see Figure 33(*a*)]. After some time, the sulphur will start burning with a blue flame. As soon as sulphur starts burning, we introduce the deflagrating spoon in a gas jar and allow the sulphur to burn inside the gas jar [see Figure 33(*b*)]. Sulphur burns in the air of gas jar to form a pungent smelling gas, sulphur dioxide. After all the sulphur has burnt, remove the deflagrating spoon from the gas jar and cover it with a lid. The gas jar now contains sulphur dioxide gas [see Figure 33(*c*)].

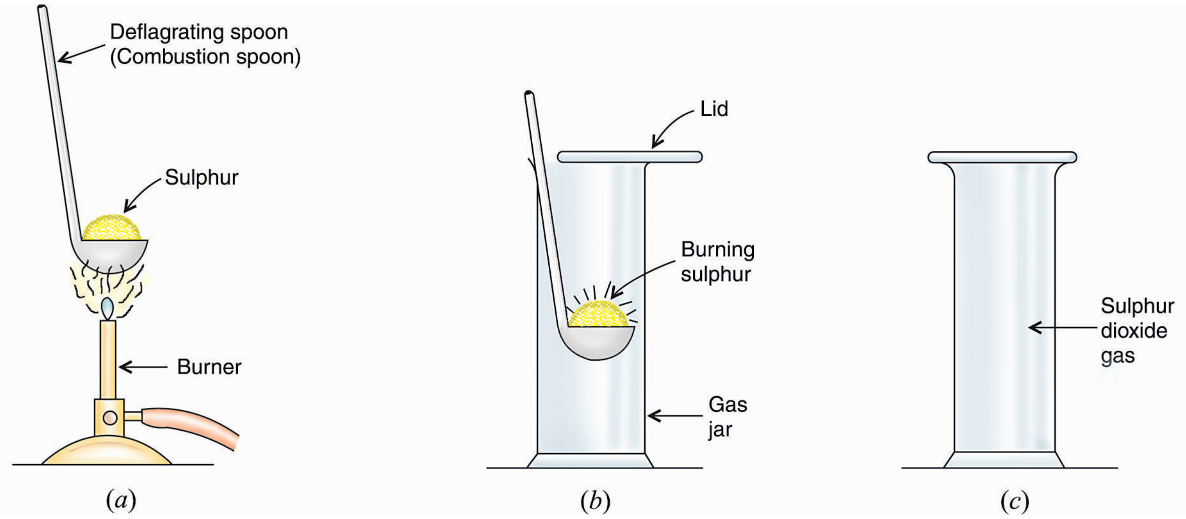


Figure 33. When sulphur is burned in air, it forms sulphur dioxide gas.

We now put some water in the gas jar, cover it with a lid and shake it to dissolve sulphur dioxide gas. Let us divide this solution into two parts by putting it in two test-tubes. We now test these solutions with blue litmus solution and red litmus solution, turn by turn. *When blue litmus solution is added to the sulphur dioxide solution, its colour changes to red*. We know that only acidic substances can turn blue litmus to red. Since sulphur dioxide solution turns blue litmus to red, it shows that sulphur dioxide is *acidic* in nature. When red litmus solution is added to sulphur dioxide solution, there is no change in colour.

Please note that instead of using litmus solutions, we can also use litmus papers for testing sulphur dioxide gas (or any other gas). Blue litmus paper and red litmus paper are available in every science laboratory. The solution of sulphur dioxide gas can be tested by using even dry litmus paper (because the gas is already dissolved in water). But **for testing sulphur dioxide gas directly, we have to use a moist litmus paper** (or wet litmus paper). The moist litmus paper contains some water which dissolves sulphur dioxide gas being tested to form acid. And this acid will then change the colour of litmus paper. **The sulphur dioxide gas has no action on a dry litmus paper**.

The non-metal oxides like CO_2 and SO_2 turn blue litmus solution red, showing that they are acidic in nature. These acidic oxides are called *acid anhydrides*. Please note that phosphorus is also a non-metal which reacts with the oxygen of air to form an acidic oxide, phosphorus pentoxide (P_2O_5) (see Figure 34).

We will now discuss some of the non-metal oxides which are neutral, being neither acidic nor basic. The neutral non-metal oxides are carbon monoxide, CO; water, H_2O; nitrogen monoxide, NO; and dinitrogen monoxide, N_2O. These oxides do not turn blue litmus solution red or red litmus solution blue. That is, these neutral non-metal oxides have

Figure 34. Phosphorus burns in air with a yellow flame to form phosphorus pentoxide.

Figure 35. White phosphorus is stored under water because it burns spontaneously in air but does not react with water.

no action on any type of litmus. Here are some examples.

(*i*) Carbon is a non-metal. When carbon burns in an insufficient supply of oxygen (of air), then it forms a neutral oxide called carbon monoxide :

$$\underset{\substack{\text{Carbon}\\ \text{(Non-metal)}}}{2C\ (s)} + \underset{\substack{\text{Oxygen}\\ \text{(Insufficient air)}}}{O_2\ (g)} \longrightarrow \underset{\substack{\text{Carbon monoxide}\\ \text{(Neutral oxide)}}}{2CO\ (g)}$$

This neutral non-metal oxide, carbon monoxide, does not produce an acid with water.

(*ii*) Hydrogen is a non-metal. When hydrogen combines with the oxygen of air, then it forms a neutral oxide called water :

$$\underset{\substack{\text{Hydrogen}\\ \text{(Non-metal)}}}{2H_2\ (g)} + \underset{\substack{\text{Oxygen}\\ \text{(From air)}}}{O_2\ (g)} \longrightarrow \underset{\substack{\text{Water}\\ \text{(Neutral oxide)}}}{2H_2O\ (l)}$$

Water (H_2O) is actually hydrogen oxide. Please note that non-metal oxides are formed by the sharing of electrons, so they are covalent compounds. They do not contain any oxide ions.

2. Reaction of Non-Metals with Water

Non-metals do not react with water (or steam) to evolve hydrogen gas. This is because non-metals cannot give electrons to reduce the hydrogen ions of water into hydrogen gas.

3. Reaction of Non-Metals with Dilute Acids

Non-metals do not react with dilute acids. In other words, non-metals do not displace hydrogen from acids. For example, the non-metals like carbon, sulphur and phosphorus do not react with dilute hydrochloric acid (HCl) or dilute sulphuric acid (H_2SO_4) to produce hydrogen gas. Let us see why non-metals are not able to displace hydrogen from acids. In order to displace hydrogen ions (H^+) of an acid and convert them into hydrogen gas, electrons should be supplied to the hydrogen ions (H^+) of the acid. Now, a non-metal, being itself an acceptor of electrons, cannot give electrons to the hydrogen ions of the acid to reduce them to hydrogen gas. And hence the non-metals are not able to displace hydrogen ions from acids to form hydrogen gas. Thus, **if non-metals like carbon, sulphur or phosphorus are put into a test-tube containing dilute sulphuric acid (or dilute hydrochloric acid), then no hydrogen gas is evolved.**

4. Reaction of Non-Metals with Salt Solutions

A more reactive non-metal displaces a less reactive non-metal from its salt solution. For example, when chlorine is passed through a solution of sodium bromide, then sodium chloride and bromine are formed :

$$\underset{\text{Sodium bromide}}{2NaBr\ (aq)} + \underset{\text{Chlorine}}{Cl_2\ (g)} \longrightarrow \underset{\text{Sodium chloride}}{2NaCl\ (aq)} + \underset{\text{Bromine}}{Br_2\ (aq)}$$

In this displacement reaction, a more reactive non-metal chlorine is displacing a less reactive non-metal bromine from its salt solution, sodium bromide solution.

5. Reaction of Non-Metals with Chlorine

Non-metals react with chlorine to form covalent chlorides which are non-electrolytes (do not conduct electricity). Non-metal chlorides are usually *liquids* or *gases*. Here are some examples.

(*i*) Hydrogen is a non-metal. So, hydrogen reacts with chlorine to form a covalent chloride called hydrogen chloride :

$$\underset{\substack{\text{Hydrogen}\\ \text{(Non-metal)}}}{H_2\ (g)} + \underset{\text{Chlorine}}{Cl_2\ (g)} \longrightarrow \underset{\substack{\text{Hydrogen chloride}\\ \text{(Covalent chloride)}}}{2HCl\ (g)}$$

***(ii)* Phosphorus** is a non-metal which reacts with chlorine to form a covalent chloride called phosphorus trichloride :

P_4 (s)	+	$6Cl_2$ (g)	$\longrightarrow$	$4PCl_3$ (l)
Phosphorus		Chlorine		Phosphorus trichloride
(Non-metal)				(Covalent chloride)

Some phosphorus pentachloride, PCl_5, is also formed in this reaction. Similarly, carbon (C) is a non-metal which reacts with chlorine to form a covalent chloride called carbon tetrachloride,CCl_4, which contains covalent bonds and does not conduct electricity. **Non-metals form covalent chlorides because they cannot give electrons to chlorine atoms to form chloride ions.**

6. Reaction of Non-Metals with Hydrogen

Non-metals react with hydrogen to form covalent hydrides. The non-metal hydrides are formed by the sharing of electrons, that is, non-metal hydrides are formed by covalent bonding. Here are some examples.

***(i)* Sulphur** is a non-metal which combines with hydrogen to form a covalent hydride called hydrogen sulphide, H_2S :

H_2 (g)	+	S (l)	$\longrightarrow$	H_2S (g)
Hydrogen		Sulphur		Hydrogen sulphide
		(Non-metal)		(Covalent hydride)

Figure 36. Hydrogen sulphide is a poisonous gas with an odour of rotten eggs.

The hydrogen sulphide gas has a characteristic smell resembling the smell of rotten eggs (see Figure 36).

***(ii)* Nitrogen** is a non-metal which combines with hydrogen in the presence of iron catalyst to form a covalent hydride called ammonia, NH_3 :

N_2 (g)	+	$3H_2$ (g)	$\xrightarrow{Fe}$	$2NH_3$ (g)
Nitrogen		Hydrogen		Ammonia
(Non-metal)				(Covalent hydride)

Oxygen is also a non-metal which combines with hydrogen to form a hydride called water, H_2O. Similarly, the hydride of carbon is methane (CH_4), and the hydride of chlorine is hydrogen chloride (HCl). The non-metal hydrides are covalent compounds formed by the sharing of electrons. **Non-metals form covalent hydrides because non-metal atoms cannot give electrons to hydrogen atoms to form hydride ions**. Non-metal hydrides are *liquids* or *gases*. Non-metal hydrides do not contain ions and hence they do not conduct electricity. Non-metal hydrides are stable compounds.

Comparison Among the Properties of Metals and Non-Metals

We have studied the characteristic properties of metals and non-metals. We will now give the main points of difference between the metals and non-metals.

Main Points of Difference between Metals and Non-Metals

Metals	*Non-Metals*
Differences in Physical Properties	
1. Metals are malleable and ductile. That is, metals can be hammered into thin sheets and drawn into thin wires.	1. Non-metals are brittle (break easily). They are neither malleable nor ductile.
2. Metals are good conductors of heat and electricity.	2. Non-metals are bad conductors of heat and electricity (except *graphite* which is a good conductor of electricity).
3. Metals are lustrous (shiny) and can be polished.	3. Non-metals are non-lustrous (dull) and cannot be polished (except *iodine* which is a lustrous non-metal).
4. Metals are solids at room temperature (except *mercury* which is a liquid metal).	4. Non-metals may be solid, liquid or gases at the room temperature.
5. Metals are strong and tough.	5. Non-metals are not strong or tough.

Differences in Chemical Properties	
1. Metals form basic oxides.	1. Non-metals form acidic oxides or neutral oxides
2. Metals displace hydrogen from water (or steam).	2. Non-metals do not react with water (or steam) and hence do not displace hydrogen from water (or steam).
3. Metals displace hydrogen from dilute acids.	3. Non-metals do not react with dilute acids and hence do not displace hydrogen from dilute acids.
4. Metals form ionic chlorides with chlorine. These ionic chlorides are electrolytes but non-volatile.	4. Non-metals form covalent chlorides with chlorine (which are non-electrolytes but volatile).
5. Metals usually do not combine with hydrogen. Only a few reactive metals combine with hydrogen to form ionic metal hydrides.	5. Non-metals react with hydrogen to form stable, covalent hydrides.

We have just given a large number of physical and chemical properties to distinguish metals from non-metals. The classification of elements into metals and non-metals is, however, not entirely satisfactory because there are exceptions to the rules given in the above table, particularly with the physical properties. So, we should keep these exceptions in mind while answering the questions. In most of the cases one or more of the following five points will be sufficient to decide whether the given substance is a metal or non-metal.

1. (*i*) If the substance is *malleable and ductile*, it will be a *metal*.
   (*ii*) If the substance is *brittle and non-ductile*, it will be a *non-metal.*
2. (*i*) If the substance is a *good conductor of heat and electricity*, it will be a metal.
   (*ii*) If the substance is a *non-conductor of heat and electricity*, it may be a *non-metal*.
3. (*i*) If the substance *reacts with a dilute acid to produce hydrogen*, it will be a *metal.*
   (*ii*) If the substance *does not react with a dilute acid*, it may be a *non-metal*.
4. (*i*) If the substance *forms a basic oxide*, it will be a *metal.*
   (*ii*) If the substance *forms an acidic oxide or neutral oxide*, it will be a *non-metal.*
5. (*i*) If the substance *forms an ionic chloride*, it will be a *metal*.
   (*ii*) If the substance *forms a covalent chloride*, it will be a *non-metal.*

Some less reactive metals like copper do not react with dilute acids to give hydrogen. So, we cannot use the dilute acid test in the case of such metals. Similarly, some non-metals like carbon (in the form of graphite) also conduct electricity. So, we cannot use the conductivity test in the case of such non-metals. **When in doubt, the nature of oxides and chlorides of the elements must be referred to for deciding whether it is a metal or a non-metal**. Apart from this, other properties like melting points, boiling points and densities, etc., are also sometimes helpful in distinguishing metals from non-metals. Let us solve some problems now.

Sample Problem 1. From amongst the following, choose the metals and non-metals and state one of the properties on the basis of which you have made your choice.

(*i*) Graphite (*ii*) Sodium (*iii*) Phosphorus (*iv*) Helium.

Solution. Out of graphite, sodium, phosphorus and helium, only sodium is a metal. All others are non-metals. This choice has been made on the basis of the nature of their oxides. This is because metals form basic oxides whereas non-metals form acidic oxides or neutral oxides.

(*i*) Graphite is actually carbon element. Graphite or carbon usually forms an acidic oxide, carbon dioxide. So, graphite is a non-metal.

(*ii*) Sodium forms a basic oxide, sodium oxide. So, sodium is a metal.

(*iii*) Phosphorus forms an acidic oxide, phosphorus pentoxide. So, phosphorus is a non-metal.

(*iv*) Helium is a gas, so it is a non-metal. Being an inert gas, helium does not form an oxide.

Sample Problem 2. An element reacts with oxygen to form an oxide which dissolves in dilute hydrochloric acid. The oxide formed also turns a solution of red litmus blue. Is the element a metal or a non-metal ? Explain your answer.

Solution. Here the oxide of given element dissolves in an acid, therefore, the oxide must be basic in nature. Moreover, since the oxide turns red litmus solution to blue, this also confirms that the oxide is basic in nature. Now, basic oxides are formed by metals, so the element in this case is a metal.

Sample Problem 3. Which of the following elements would yield a basic oxide ?

S, P, Ca, Si

Solution. We know that only metal elements yield basic oxides. Now, out of the above given elements only Ca is a metal (Ca = calcium), therefore, Ca will yield a basic oxide. The elements S (sulphur), P (phosphorus) and Si (silicon) are all non-metals.

Sample Problem 4. Which of the following will displace hydrogen from acids to form salts ?

S, P, Na, Si

Solution. The metals displace hydrogen from acids to form salts. Out of the above given elements only Na (sodium) is a metal. So, Na will displace hydrogen from acids to form salts. The other elements S, P and Si are all non-metals (which do not displace hydrogen from acids).

Sample Problem 5. Pratyush took sulphur powder on a spatula and heated it. He collected the gas evolved by inverting a test-tube over the burning sulphur.

(*a*) What will be the action of this gas on :

(*i*) dry litmus paper ?

(*ii*) moist litmus paper ?

(*b*) Write a balanced chemical equation for the reaction taking place. **(NCERT Book Question)**

Solution. (*a*) When sulphur is burnt in air then sulphur dioxide gas is formed.

(*i*) Sulphur dioxide gas has no action on dry litmus paper.

(*ii*) Sulphur dioxide gas turns moist blue litmus paper to red.

(*b*) $S\,(s) + O_2\,(g) \longrightarrow SO_2\,(g)$

USES OF METALS

Metals are used for a large number of purposes. Some of the uses of metals are given below :

1. Copper and aluminium metals are used to make wires to carry electric current. This is because copper and aluminium have very low electrical resistance and hence very good conductors of electricity.
2. Iron, copper and aluminium metals are used to make house-hold utensils and factory equipment.

(*a*) These bins made of steel have been galvanised (coated with zinc metal)

(*b*) Chromium metal is used for electroplating iron and steel objects

(*c*) Lead metal is used in making car batteries

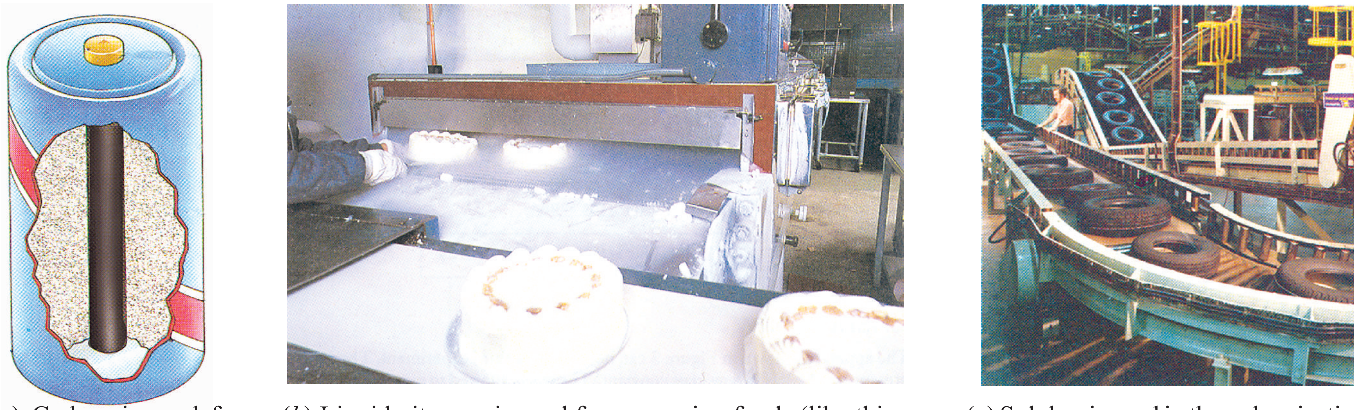
Figure 37. Some of the uses of metals.

3. Iron is used as a catalyst in the preparation of ammonia gas by Haber's process.
4. Zinc is used for galvanizing iron to protect it from rusting.
5. Chromium and nickel metals are used for electroplating and in the manufacture of stainless steel.
6. The aluminium foils are used in packaging of medicines, cigarettes and food materials.
7. Silver and gold metals are used to make jewellery. The thin foils made of silver and gold are used to decorate sweets.
8. The liquid metal 'mercury' is used in making thermometers.
9. Sodium, titanium and zirconium metals are used in atomic energy (nuclear energy) and space science projects.
10. Zirconium metal is used in making bullet-proof alloy steels.
11. Lead metal is used in making car batteries.

USES OF NON-METALS

The important uses of non-metals are as follows :

1. Hydrogen is used in the hydrogenation of vegetable oils to make vegetable *ghee* (or *vanaspati ghee*).
2. Hydrogen is used in the manufacture of ammonia (whose compounds are used as fertilisers).
3. Liquid hydrogen is used as a rocket fuel.
4. Carbon (in the form of graphite) is used for making the electrodes of electrolytic cells and dry cells.
5. Nitrogen is used in the manufacture of ammonia, nitric acid and fertilisers.



(*a*) Carbon is used for making electrodes of dry cells.

(*b*) Liquid nitrogen is used for preserving foods (like this cake) by freezing it quickly

(*c*) Sulphur is used in the vulcanisation of rubber (hardening of rubber) as in making tyres

Figure 38. Some of the uses of non-metals.

6. Due to its inertness, nitrogen is used to preserve food materials.
7. Compounds of nitrogen like Tri Nitro Toluene (TNT) and nitroglycerine are used as explosives.
8. Sulphur is used for manufacturing sulphuric acid.
9. Sulphur is used as a fungicide and in making gun powder.
10. Sulphur is used in the vulcanisation of rubber.

Before we go further and describe how metals and non-metals combine to form chemical bonds, **please answer the following questions** :

Very Short Answer Type Questions

1. Name one metal and one non-metal which exist in liquid state at room temperature.
2. Why are metals called electropositive elements whereas non-metals are called electronegative elements ?
3. (*a*) Name the most abundant metal in the earth's crust.
(*b*) Name the most abundant non-metal in the earth's crust.

4. Name one metal which has a low melting point.
5. Name the metal which is the poorest conductor of heat.
6. State whether the following statement is true or false :
   Non-metals react with dilute acids to produce a gas which burns with a 'pop' sound.
7. From amongst the metals sodium, calcium, aluminium, copper and magnesium, name the metal :
   (*i*) which reacts with water only on boiling, and
   (*ii*) another which does not react even with steam.
8. What changes in the colour of iron nails and copper sulphate solution do you observe after keeping the iron nails dipped in copper sulphate solution for about 30 minutes ?
9. What is aqua-regia ? Name two special metals which are insoluble in common reagents but dissolve in aqua-regia.
10. Give the names and formulae of (*a*) two acidic oxides, and (*b*) two basic oxides.
11. What name is given to those metal oxides which show basic as well as acidic behaviour ?
12. Name two metals which form amphoteric oxides.
13. A copper coin is kept immersed in a solution of silver nitrate for some time. What will happen to the coin and the colour of the solution ?
14. Which property of copper and aluminium makes them suitable :
    (*a*) for making cooking utensils and boilers ?
    (*b*) for making electric wires ?
15. Write the names and formulae of (*a*) a metal hydride, and (*b*) a non-metal hydride.
16. Name the metal which has been placed :
    (*a*) at the bottom of the reactivity series
    (*b*) at the top of the reactivity series
    (*c*) just below copper in the reactivity series
17. Which of the two metals is more reactive : copper or silver ?
18. (*a*) Name one metal which is stored in kerosene oil.
    (*b*) Name one non-metal which is stored under water.
19. Write equation for the reaction of :
    (*a*) sodium with oxygen
    (*b*) magnesium with oxygen
20. Name two metals which are used :
    (*a*) for making electric wires.
    (*b*) for making domestic utensils and factory equipment.
    (*c*) for making jewellery and to decorate sweets.
21. Which metal foil is used for packing some of the medicine tablets ?
22. Name the non-metal which is used :
    (*a*) to convert vegetable oil into vegetable *ghee* (solid fat).
    (*b*) as a rocket fuel (in liquid form).
    (*c*) to make electrodes of dry cells.
    (*d*) to preserve food materials.
    (*e*) in the vulcanisation of rubber.
23. Name one property which is characteristic of (*a*) metals, and (*b*) non-metals.
24. What is meant by "brittleness" ? Which type of elements usually show brittleness : metals or non-metals ?
25. What will happen if a strip of zinc is immersed in a solution of copper sulphate ?
26. What will happen if a strip of copper is kept immersed in a solution of silver nitrate ($AgNO_3$) ?
27. What happens when iron nails are put into copper sulphate solution ?
28. How would you show that silver is chemically less reactive than copper ?
29. Give reasons for the following :
    Blue colour of copper sulphate solution is destroyed when iron filings are added to it.

30. Name a non-metal having a very high melting point.

31. Which property of graphite is utilised in making electrodes ?

32. Name two non-metals which are both brittle and non-ductile.

33. Explain why, the surface of some metals acquires a dull appearance when exposed to air for a long time.

34. Complete and balance the following equations :

(*a*) $Na + O_2 \longrightarrow$

(*b*) $Na_2O + H_2O \longrightarrow$

(*c*) $Fe\ (s) + H_2O\ (g) \xrightarrow{\text{Red heat}}$

(*d*) $Cu(NO_3)_2\ (aq) + Zn\ (s) \longrightarrow$

35. Fill in the following blanks with suitable words :

(*a*) Magnesium liberates .... gas on reacting with hot boiling water.

(*b*) The white powder formed when magnesium ribbon burns in oxygen is of ....

(*c*) Ordinary aluminium strips are not attacked by water because of the presence of a layer of .... on the surface of aluminium.

(*d*) A metal having low melting point is ....but a non-metal having very high melting point is.............

(*e*) Calcium is a ....reactive metal than sodium.

Short Answer Type Questions

36. (*a*) What is meant by saying that the metals are malleable and ductile ? Explain with examples.

(*b*) Name two metals which are both malleable and ductile.

(*c*) Which property of iron metal is utilised in producing iron sheets required for making buckets ?

(*d*) Which property of copper metal is utilised in making thin wires ?

37. Name two metals which react violently with cold water. Write any three observations you would make when such a metal is dropped into water. How would you identify the gas evolved, if any, during the reaction ?

38. (*a*) With the help of examples, describe how metal oxides differ from non-metal oxides.

(*b*) Which of the following elements would yield : (*i*) an acidic oxide, (*ii*) a basic oxide, and (*iii*) a neutral oxide ?

Na, S, C, K, H

39. (*a*) What are amphoteric oxides ? Give two examples of amphoteric oxides.

(*b*) Choose the acidic oxides, basic oxides and neutral oxides from the following :

Na_2O ; CO_2 ; CO ; SO_2 ; MgO ; N_2O ; H_2O.

(*c*) Which of the following are amphoteric oxides :

MgO, ZnO, P_2O_3, Al_2O_3, NO_2

40. (*a*) What is the nature of the oxide SO_2 ? What happens when it is dissolved in water ? Write the chemical equation of the reaction involved.

(*b*) What is the nature of the oxide Na_2O ? What happens when it is dissolved in water ? Write the chemical equation of the reaction inolved.

41. (*a*) What type of oxides are formed when non-metals react with oxygen ? Explain with an example.

(*b*) What type of oxides are formed when metals combine with oxygen ? Explain with the help of an example.

42. (*a*) Explain why, metals usually do not liberate hydrogen gas with dilute nitric acid.

(*b*) Name two metals which can, however, liberate hydrogen gas from very dilute nitric acid.

43. (*a*) How do metals react with hydrogen ? Explain with an example.

(*b*) How do non-metals react with hydrogen ? Explain with an example.

44. (*a*) What happens when calcium reacts with chlorine ? Write an equation for the reaction which takes place.

(*b*) What happens when magnesium reacts with very dilute nitric acid ? Write an equation for the reaction involved.

45. (*a*) Arrange the following metals in order of their chemical reactivity, placing the most reactive metal first :

Magnesium, Copper, Iron, Sodium, Zinc, Lead, Calcium.

(*b*) What happens when a rod of zinc metal is dipped into a solution of copper sulphate ? Give chemical equation of the reaction involved.

46. A copper plate was dipped in $AgNO_3$ solution. After certain time, silver from the solution was deposited on the copper plate. State the reason why it happened. Give the chemical equation of the reaction involved.

47. State five uses of metals and five of non-metals.

48. State one use each of the following metals :
Copper, Aluminium, Iron, Silver, Gold, Mercury

49. (*a*) State one use each of the following non-metals :
Hydrogen, Carbon (as Graphite), Nitrogen, Sulphur
(*b*) Name the metal which is used in making thermometers.

50. (*a*) Why does aluminium not react with water under ordinary conditions ?
(*b*) Name two metals which can displace hydrogen from dilute acids.
(*c*) Name two metals which cannot displace hydrogen from dilute acids.

51. (*a*) Why is sodium kept immersed in kerosene oil ?
(*b*) Why is white phosphorus kept immersed under water ?
(*c*) Can we keep sodium immersed under water ? Why ?

52. (*a*) Describe the reaction of potassium with water. Write the equation of the reaction involved.
(*b*) Write an equation of the reaction of iron with steam. Indicate the physical states of all the reactants and products.
(*c*) Which gas is produced when dilute hydrochloric acid is added to a reactive metal ?

53. (*a*) Give one example, with equation, of the displacement of hydrogen by a metal from an acid.
(*b*) Name two metals (other than zinc and iron) which can displace hydrogen from dilute hydrochloric acid ?

54. What is the action of water on (*a*) sodium (*b*) magnesium, and (*c*) aluminium ? Write equations of the chemical reactions involved.

55. You are given samples of three metals — sodium, magnesium and copper. Suggest any two activities to arrange them in order of their decreasing reactivities.

56. (*a*) Write one reaction in which aluminium oxide behaves as a basic oxide and another in which it behaves as an acidic oxide.
(*b*) What special name is given to substances like aluminium oxide.
(*c*) Name another metal oxide which behaves like aluminium oxide.

57. (*a*) What happens when calcium reacts with water ? Write the chemical equation of the reaction of calcium with water.
(*b*) Write the chemical equation of the reaction which takes place when iron reacts with dilute sulphuric acid. What happens when the gas produced is ignited with a burning matchstick ?

58. You are given a dry cell, a torch bulb with holder, wires and crocodile clips. How would you use them to distinguish between samples of metals and non-metals ?

59. State any five physical properties of metals and five physical properties of non-metals.

60. (*a*) Name two physical properties each of sodium and carbon in which their behaviour is not as expected from their classification as metal and non-metal respectively.
(*b*) Name two metals whose melting points are so low that they melt when held in the hand.

61. Metals are said to be shiny. Why do metals generally appear to be dull ? How can their brightness be restored ?

Long Answer Type Questions

62. (*a*) What are metals ? Name five metals.
(*b*) Name a metal which is so soft that it can be cut with a knife.
(*c*) Name the metal which is the best conductor of heat and electricity.
(*d*) What happens when a metal reacts with dilute hydrochloric acid ? Explain with the help of an example.
(*e*) Write the equations for the reactions of :
(*i*) Magnesium with dilute hydrochloric acid
(*ii*) Aluminium with dilute hydrochloric acid
(*iii*) Zinc with dilute hydrochloric acid
(*iv*) Iron with dilute hydrochloric acid
Name the products formed in each case. Also indicate the physical states of all the substances involved.

63. (*a*) Define non-metals. Give five examples of non-metals.
(*b*) Name a non-metal which conducts electricity.
(*c*) Name a non-metal having lustre (shining surface).
(*d*) Name a non-metal which is extremely hard.
(*e*) How do non-metals react with oxygen ? Explain with an example. Give equation of the reaction involved. What is the nature of the product formed ? How will you demonstrate it ?

64. (*a*) What is meant by the reactivity series of metals ? Arrange the following metals in an increasing order of their reactivities towards water :
Zinc, Iron, Magnesium, Sodium
(*b*) Hydrogen is not a metal but still it has been assigned a place in the reactivity series of metals. Why ?
(*c*) Name one metal more reactive and another less reactive than hydrogen.
(*d*) Name one metal which displaces copper from copper sulphate solution and one which does not.
(*e*) Name one metal which displaces silver from silver nitrate solution and one which does not.

65. (*a*) State any three differences between the physical properties of metals and non-metals.
(*b*) Differentiate between metals and non-metals on the basis of their chemical properties.
(*c*) State three reasons (of which at least one must be chemical) for believing that sodium is a metal.
(*d*) State three reasons (of which at least one must be chemical) for believing that sulphur is a non-metal.
(*e*) Which non-metal has been placed in the reactivity series of metals ?

Multiple Choice Questions (MCQs)

66. The elements whose oxides can turn phenolphthalein solution pink are :
(*a*) Na and K (*b*) K and C (*c*) Na and S (*d*) K and P

67. "Is malleable and ductile". This best describes :
(*a*) a metal (*b*) a compound (*c*) a non-metal (*d*) a solution

68. One of the following is not a neutral oxide. This is :
(*a*) CO (*b*) H_2O (*c*) N_2O (*d*) Na_2O

69. A basic oxide will be formed by the element :
(*a*) K (*b*) S (*c*) P (*d*) Kr

70. An acidic oxide is produced by the element :
(*a*) Na (*b*) C (*c*) Ca (*d*) H

71. You are given a solution of $AgNO_3$. Which of the following do you think cannot displace Ag from $AgNO_3$ solution ?
(*a*) Magnesium (*b*) Zinc (*c*) Gold (*d*) Copper

72. Out of aluminium, copper, calcium and tin, the most reactive metal is :
(*a*) aluminium (*b*) copper (*c*) tin (*d*) calcium

73. The least reactive metal among the following is :
(*a*) sodium (*b*) silver (*c*) copper (*d*) lead

74. An element X reacts with hydrogen, when heated, to form a covalent hydride H_2X. If H_2X has a smell of rotten eggs, the element X is likely to be :
(*a*) carbon (*b*) sulphur (*c*) chlorine (*d*) phosphorus

75. Out of the following oxides, the amphoteric oxide is :
(*a*) Fe_2O_3 (*b*) Al_2O_3 (*c*) P_2O_5 (*d*) N_2O

76. The metals which can produce amphoteric oxides are :
(*a*) sodium and aluminium (*b*) zinc and potassium
(*c*) calcium and sodium (*d*) aluminium and zinc

77. An element X forms two oxides XO and XO_2. The oxide XO is neutral but XO_2 is acidic in nature. The element X is most likely to be :
(*a*) sulphur (*b*) carbon (*c*) calcium (*d*) hydrogen

78. The elements whose oxides can turn litmus solution blue are :
(*a*) carbon and sulphur (*b*) sodium and carbon
(*c*) potassium and magnesium (*d*) magnesium and sulphur

79. The elements whose oxides can turn litmus solution red are :
(*a*) lithium and sodium (*b*) copper and potassium
(*c*) carbon and hydrogen (*d*) phosphorus and sulphur

80. Zinc oxide is a metal oxide. Which of the following term best describes the nature of zinc oxide :
(*a*) an acidic oxide (*b*) a basic oxide (*c*) an amphoteric oxide (*d*) a neutral oxide

81. A metal less reactive and another metal more reactive than hydrogen are :
(*a*) aluminium and lead (*b*) iron and magnesium
(*c*) copper and tin (*d*) copper and mercury

82. An element E reacts with water to form a solution which turns phenolphthalein solution pink. The element E is most likely to be :
(*a*) S (*b*) Ca (*c*) C (*d*) Ag

83. 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

84. Which one of the following four metals would be displaced from the solution of its salt by the other three metals ?
(*a*) Zn (*b*) Ag (*c*) Cu (*d*) Mg

85. An element is soft and can be cut with a knife. It is very reactive and cannot be kept open in the air. It reacts vigorously with water. The element is most likely to be :
(*a*) Mg (*b*) S (*c*) P (*d*) Na

86. Which of the following metal exists in the liquid state ?
(*a*) Na (*b*) Ag (*c*) Cr (*d*) Hg

87. Which of the following non-metal is a liquid ?
(*a*) carbon (*b*) sulphur (*c*) bromine (*d*) iodine

88. Which of the following pair of reactants can undergo a displacement reaction under appropriate conditions ?
(*a*) $MgSO_4$ + Fe (*b*) $ZnSO_4$ + Fe (*c*) $MgSO_4$ + Pb (*d*) $CuSO_4$ + Fe

Questions Based on High Order Thinking Skills (HOTS)

89. An element E forms an oxide E_2O. An aqueous solution of E_2O turns red litmus paper blue.
(*a*) What is the nature of the oxide E_2O ?
(*b*) State whether element E is a metal or a non-metal.
(*c*) Give one example of an element like E.

90. Metal A burns in air, on heating, to form an oxide A_2O_3 whereas another metal B burns in air only on strong heating to form an oxide BO. The two oxides A_2O_3 and BO can react with hydrochloric acid as well as sodium hydroxide solution to form the corresponding salts and water.
(*a*) What is the nature of oxide A_2O_3 ?
(*b*) What is the nature of oxide BO ?
(*c*) Name one metal like A.
(*d*) Name one metal like B.

91. An element X forms two oxides XO and XO_2. The oxide XO has no action on litmus solution but oxide XO_2 turns litmus solution red.
(*a*) What is the nature of oxide XO ?
(*b*) What is the nature of oxide XO_2 ?
(*c*) Would you call element X a metal or a non-metal ? Give reason for your choice.
(*d*) Can you give an example of element like X ?

92. State and explain the reactions, if any, of the following metals with a solution of copper sulphate :
(*a*) Gold (*b*) Copper (*c*) Zinc (*d*) Mercury

93. (*a*) Give the names and formulae of one metal chloride and one non-metal chloride.
(*b*) State an important property in which these metal chloride and non-metal chloride differ.
(c) Why do they differ in this property ?

94. In a solution of lead acetate, a strip of metal M was dipped. After some time, lead from the solution was deposited on the metal strip. Which metal is more reactive, M or lead ?

95. $CuSO_4$ (aq) + Fe (s) $\longrightarrow$ $FeSO_4$ (aq) + Cu (s)

$FeSO_4$ (aq) + Zn (s) $\longrightarrow$ $ZnSO_4$ (aq) + Fe (s)

On the basis of the above reactions, indicate which is most reactive and which is least reactive metal out of zinc, copper and iron.

96. Which of the following reactions will not occur ? Why not ?

(*a*) $MgSO_4$ (aq) + Cu (s) $\longrightarrow$ $CuSO_4$ (aq) + Mg (s)

(*b*) $CuSO_4$ (aq) + Fe (s) $\longrightarrow$ $FeSO_4$ (aq) + Cu (s)

(*c*) $MgSO_4$ (aq) + Fe (s) $\longrightarrow$ $FeSO_4$ (aq) + Mg (s)

97. In nature, metal A is found in a free state while metal B is found in the form of its compounds. Which of these two will be nearer to the top of the activity series of metals ?

98. If A, B, C, D, E, F, G, H, I, J and K represent metals in the decreasing order of their reactivity, which one of them is most likely to occur in a free state in nature ?

99. (*a*) Name a metal for each case :

(*i*) It does not react with cold as well as hot water but reacts with steam.

(*ii*) It does not react with any physical state of water.

(*b*) When calcium metal is added to water, the gas evolved does not catch fire but the same gas evolved on adding sodium metal to water catches fire. Why is it so ?

100. A zinc plate was kept in a glass container having $CuSO_4$ solution. On examining it was found that the blue colour of the solution is getting lighter and lighter. After a few days, when the zinc plate was taken out of the solution, a number of small holes were noticed in it. State the reason and give chemical equation of the reaction involved.

ANSWERS

6. False **7.** (*i*) Aluminium (*ii*) Copper **8.** Iron nails get covered with a red-brown coating of copper metal ; The blue colour of copper sulphate solution fades gradually **11.** Amphoteric oxides **12.** Aluminium and Zinc **13.** Copper coin will get a coating of silver metal ; The colour of solution will turn blue **14.** (*a*) High thermal (heat) conductivity (*b*) High electrical conductivity **17.** Copper **18.** (*a*) Sodium (*b*) White phosphorus **21.** Aluminium foil **22.** (*a*) Hydrogen (*b*) Hydrogen (*c*) Carbon (as Graphite) (*d*) Nitrogen (*e*) Sulphur **24.** Non-metals **28.** If a strip of silver metal is kept immersed in copper sulphate solution for some time, silver is not able to displace copper from copper sulphate solution (to form a red-brown coating on silver strip) **32.** Sulphur and Phosphorus **35.** (*a*) hydrogen (*b*) magnesium oxide (*c*) aluminium oxide (*d*) sodium ; diamond (*e*) less **36.** (*c*) Malleability (*d*) Ductility **37.** Sodium and Potassium ; Metal moves over the surface of water causing little explosions and ultimately catches fire and starts burning ; The gas burns producing a 'pop' sound indicating that it is hydrogen **38.** (*b*) (*i*) Acidic oxide : S, C (*ii*) Basic oxide : Na, K (*iii*) Neutral oxide : H **42.** (*b*) Magnesium and Manganese **45.** (*a*) Sodium > Calcium > Magnesium > Zinc > Iron > Lead > Copper **49.** (*b*) Mercury **51.** (*c*) No. Because sodium reacts with water to form sodium hydroxide and hydrogen **52.** (*c*) Hydrogen **53.** (*b*) Magnesium and Aluminium **56.** (*b*) Amphoteric oxides (*c*) Zinc oxide **60.** (*a*) Sodium metal : Soft, Low melting point ; Carbon non-metal : Graphite conducts electricity, Diamond has a very high melting point (*b*) Gallium and Cesium **64.** (*a*) Iron < Zinc < Magnesium < Sodium (*c*) Zinc ; Copper (*d*) Zinc ; Silver (*e*) Copper ; Gold **65.** (*c*) Sodium: Solid ; Conducts electricity ; Forms basic oxides (*d*) Sulphur ; Brittle and non-ductile ; Non-conductor of electricity ; Forms acidic oxides (*e*) Hydrogen **66.** (*a*) **67.** (*a*) **68.** (*d*) **69.** (*a*) **70.** (*b*) **71.** (*c*) **72.** (*d*) **73.** (*b*) **74.** (*b*) **75.** (*b*) **76.** (*d*) **77.** (*b*) **78.** (*c*) **79.** (*d*) **80.** (*c*) **81.** (*c*) **82.** (*b*) **83.** (*a*) **84.** (*b*) **85.** (*d*) **86.** (*d*) **87.** (*c*) **88.** (*d*) **89.** (*a*) Basic oxide (*b*) Metal (*c*) Sodium, Na **90.** (*a*) Amphoteric oxide (*b*) Amphoteric oxide (*c*) Aluminium, Al (*d*) Zinc, Zn **91.** (*a*) Neutral oxide (*b*) Acidic oxide (*c*) Non-metal. Because it also forms an acidic oxide (*d*) Carbon, C **92.** (*a*) No displacement reaction with gold because gold is less reactive than copper (*b*) No reaction of copper with copper sulphate solution (*c*) Zinc displaces copper from copper sulphate solution to form zinc sulphate solution and copper metal because zinc is more reactive than copper (*d*) No displacement reaction with mercury because mercury is less reactive than copper **93.** (*a*) Metal chloride : Sodium chloride, NaCl ; Non-metal chloride : Carbon tetrachloride, CCl_4 (*b*) Sodium chloride solution conducts electricity whereas carbon tetrachloride does not conduct electricity (*c*) Sodium chloride is an ionic compound whereas carbon tetrachloride is a covalent compound **94.** M is more reactive **95.** Zinc is most reactive ; Copper is least reactive **96.** Reaction (*a*) will not occur because Cu is less reactive than Mg ; Reaction (*c*) will also not occur because Fe is less reactive than Mg **97.** B **98.** K **99.** (*a*) (*i*) Iron (*ii*) Copper (*b*) More heat is evolved during the reaction of sodium metal with water due to which the hydrogen gas formed catches fire. On the other hand, less heat is evolved during the reaction of calcium metal with water which cannot make the hydrogen gas burn. **100.** Zinc metal is more reactive than copper. Some of the zinc metal of zinc plate dissolves and displaces copper from copper sulphate solution. This dissolving of zinc metal forms tiny holes in zinc plate. Blue colour of copper sulphate solution gets lighter and lighter due to the formation of colourless zinc sulphate solution.

HOW DO METALS AND NON-METALS REACT

When metals react with non-metals, they form ionic compounds (which contain ionic bonds). On the other hand, **when non-metals react with other non-metals, they form covalent compounds** (which contain covalent bonds). Metals, however, do not react with other metals. Let us first see what is meant by a chemical bond. When atoms of the elements combine to form molecules, a force of attraction is developed between the atoms (or ions) which holds them together. **The force which links the atoms (or ions) in a molecule is called a chemical bond (or just 'bond')**. In order to understand the formation of chemical bonds 'between the atoms of metals and non-metals' or 'between the atoms of two non-metals', it is necessary to know the reason for the unreactive nature (or inertness) of noble gases which we will discuss now. Please note that ***noble gases*** are also called ***inert gases*** (because they are chemically very inert or unreactive).

Inertness of Noble Gases

There are some elements in group 18 of the periodic table which do not combine with other elements. These elements are : Helium, Neon, Argon, Krypton, Xenon and Radon. They are known as noble gases or inert gases because they are unreactive and do not react with other elements to form compounds. In other words, inert gases do not form chemical bonds. We know that only the outermost electrons of an atom take part in a chemical reaction. ***Since the noble gases are chemically unreactive, we must conclude that the electron arrangements in their atoms are very stable which do not allow the outermost electrons to take part in chemical reactions.*** We will now write down the electronic configurations of the noble gases to find out the exact reason for their inert nature.

Electronic Configurations of Noble Gases (or Inert Gases)

Noble gas (Inert gas)	*Symbol*	*Atomic number*	*Electronic configuration* K L M N O P	*Number of electrons in outermost shell (Valence shell)*
1. Helium	He	2	2	2
2. Neon	Ne	10	2, 8	8
3. Argon	Ar	18	2, 8, 8	8
4. Krypton	Kr	36	2, 8, 18, 8	8
5. Xenon	Xe	54	2, 8, 18, 18, 8	8
6. Radon	Rn	86	2, 8, 18, 32, 18, 8	8

If we look at the number of electrons in the outermost shells of the inert gases (in the table given above), we find that **only one inert gas helium has 2 electrons in its outermost shell, all other inert gases have 8 electrons in the outermost shells of their atoms.** Since the atoms of inert gases are very stable and have 8 electrons (or 2 electrons) in their outermost shells, therefore, to have 8 electrons (or 2 electrons) in the outermost shell of an atom is considered to be the most stable arrangement of electrons. From this discussion we conclude that the atoms having 8 electrons (or 2 electrons) in their outermost shells are very stable and unreactive. It is very important to note here that **though 8 electrons in the outermost shell always impart stability to an atom, but 2 electrons in the outermost shell impart stability only when the outermost shell is the first shell (K shell), and no other shells are present in the atom.** To have "8 electrons" in the outermost shell of an atom is known as "octet" of electrons. Most of the inert gases have octet of electrons in their valence shells. To have "2 electrons" in the outermost K shell is known as "duplet" of electrons. Helium is the only inert gas having duplet of electrons in its valence shell. Thus, **the usual**

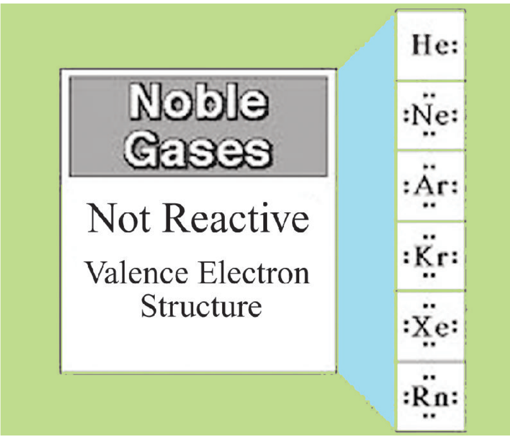


Figure 39. Only one noble gas helium (He) has 2 valence electrons (2 outermost electrons). All other noble gases have 8 electrons each in their valence shells.

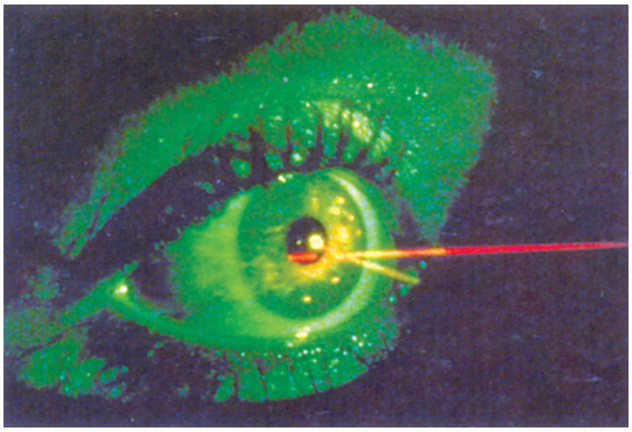


Figure 40. The noble gases are used in lasers. Here a helium-neon laser is being used in eye-surgery.

number of electrons in the outermost shell of the atom of a noble gas is 8. Only in the case of one noble gas helium, the number of outermost electrons is 2. So, helium is the only inert gas having less than 8 electrons in its outermost shell.

It should be noted that noble gases are unreactive because they have very stable electron arrangements with 8 (or 2) electrons in their outermost shells. In other words, **the noble gas atoms have completely filled outermost shells (or valence shells)**. It is not possible to remove electrons from the outermost shell of a noble gas atom or to add electrons to the outermost shell of a noble gas atom. Due to this the outermost electrons of a noble gas atom cannot take part in chemical reactions. Since the noble gases having completely filled outermost shells or valence shells are chemically unreactive, **we can explain the reactivity of elements as a tendency of their atoms to achieve a completely filled outermost shell or valence shell (just like those of noble gases) and become stable.**

Cause of Chemical Bonding (or Chemical Combination)

Everything in this world wants to become more stable. For atoms, stability means having the electron arrangement of an inert gas. **The atoms combine with one another to achieve the inert gas electron arrangement and become more stable.** In other words, **atoms form chemical bonds to achieve stability by acquiring the inert gas electron configuration.** So, when atoms combine to form chemical bonds (or chemical compounds), they do so in such a way that each atom gets 8 electrons in its outermost shell or 2 electrons in the outermost K shell. In other words, the atoms having less than 8 electrons (or less than 2 electrons) in their outermost shell are unstable. So, all the atoms have a tendency to achieve the inert gas electron arrangement of 8 electrons (or 2 electrons) in their outermost shells and become more stable. **An atom can achieve the inert gas electron arrangement (or noble gas electron arrangement) in three ways :**

(*i*) by losing one or more electrons (to another atom)
(*ii*) by gaining one or more electrons (from another atom)
(*iii*) by sharing one or more electrons (with another atom)

The chemical reactions in which the inert gas electron arrangement is achieved by the loss and gain of electrons (or transfer of electrons) between atoms, take place between *metals* and *non-metals*. On the other hand, the chemical reactions in which the inert gas electron configuration is achieved by the sharing of electrons between atoms, take place between *non-metals* and *non-metals*. We will discuss both these cases one by one.

In order to understand the reactions between metals and non-metals, it is necessary to know the meaning of the term 'ions' and how they are formed. So, let us discuss the ions first. Before we do that please note that an electron is represented by the symbol e^- (where e stands for electron and minus sign shows one unit negative charge on it). Another point to be noted is that the ***outermost electron shell*** of an atom is also known as its ***valence shell*** ; and ***outermost electrons*** are also known as ***valence electrons***.

IONS

An ion is an electrically charged atom (or group of atoms). Examples of the ions are : sodium ion, Na^+, magnesium ion, Mg^{2+}, chloride ion, Cl^-, and oxide ion, O^{2-}. **An ion is formed by the loss or gain of electrons by an atom, so it contains an unequal number of electrons and protons.** There are two types of ions : **cations and anions.**

1. A positively charged ion is known as cation. Sodium ion, Na^+, and magnesium ion, Mg^{2+}, are cations because they are positively charged ions. **A cation is formed by the loss of one or more electrons by an atom.** For example, sodium atom loses 1 electron to form a sodium ion, Na^+, which is a cation :

$$\underset{\text{Sodium atom}}{Na} \quad - \quad \underset{\text{Electron}}{e^-} \longrightarrow \underset{\substack{\text{Sodium ion}\\ \text{(A cation)}}}{Na^+}$$

Figure 41. A positively charged sodium ion (or sodium cation), Na^+.

Since a cation is formed by the removal of electrons from an atom, therefore, **a cation contains less electrons than a normal atom.** We also know that a normal atom (or a neutral atom) contains an equal number of protons and electrons. Now, since a cation is formed by the loss of one or more electrons by an atom, therefore, **a cation contains less electrons than protons.** The ions of all the metal elements are cations. Only the hydrogen ion, H^+, and ammonium ion, NH_4^+ , are the cations formed from non-metals.

2. A negatively charged ion is known as anion. Chloride ion, Cl^-, and oxide ion, O^{2-}, are anions because they are negatively charged ions. **An anion is formed by the gain of one or more electrons by an atom.** For example, a chlorine atom gains (accepts) 1 electron to form a chloride ion, Cl^-, which is an anion :

$$\underset{\text{Chlorine atom}}{Cl} \quad + \quad \underset{\text{Electron}}{e^-} \longrightarrow \underset{\substack{\text{Chloride ion}\\ \text{(An anion)}}}{Cl^-}$$

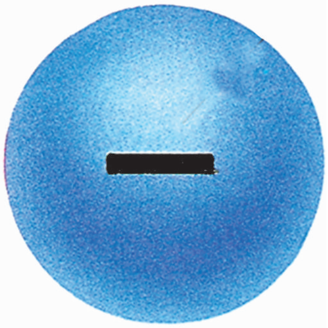


Figure 42. A negatively charged chloride ion (or chloride anion), Cl^-.

Since an anion is formed by the addition of electrons to an atom, therefore, **an anion contains more electrons than a normal atom.** We also know that a normal atom (or a neutral atom) contains an equal number of protons and electrons. Now, since an anion is formed by the addition of one or more electrons to an atom, therefore, **an anion contains more electrons than protons.** The ions of all the non-metal elements are anions (except hydrogen ion and ammonium ion). We will now discuss the formation of ions in detail.

FORMATION OF POSITIVE IONS (OR CATIONS)

If an element has 1, 2 or 3 electrons in the outermost shell of its atoms, then it loses these electrons to achieve the inert gas electron arrangement of eight valence electrons and forms positively charged ion or cation (It is not possible to add 7, 6 or 5 electrons to an atom due to energy considerations). Now, the metal atoms have usually 1, 2 or 3 electrons in the outermost shell, so **the metal atoms lose electrons to form positively charged ions or cations.** For example, lithium, sodium, potassium, magnesium, calcium and aluminium, etc., are all metals which donate their outermost electrons to form positive ions.

Please note that an atom having 1 electron in its outermost shell loses this 1 electron to form a cation having 1 unit positive charge. An atom having 2 outermost electrons loses these 2 electrons to form a cation having 2 units of positive charge. And an atom having 3 valence electrons loses these 3 electrons and forms a cation having 3 units of positive charge. We will now take some examples to understand how positive ions are formed and what changes take place in the electronic configuration during their formation.

A Point to Remember : Look at the electronic configurations of sodium (atomic number 11) and its nearest inert gas neon (atomic number 10) :

Sodium atom	*Neon atom*
K L M	K L
2, 8, 1	2, 8

The sodium atom has 1 electron more than a neon atom. So, if a sodium atom donates its 1 outermost electron (to some other atom), then it will achieve the electron arrangement of inert gas neon and become very stable. Keeping this point in mind, it will now be easier to understand the formation of a sodium ion.

1. Formation of a Sodium ion, Na^+

The atomic number of sodium is 11. So, one atom of sodium contains 11 electrons. The electronic configuration of sodium will be $\underset{2,\ 8,\ 1}{\text{K L M}}$ We find that sodium atom has 1 electron in its outermost shell (M shell). This is not a stable arrangement of electrons. A stable arrangement has usually 8 electrons in its outermost shell. Thus, a sodium atom is not very stable, it is very reactive. In order to become more stable, a sodium atom donates its 1 outermost electron to some other atom (like that of chlorine). In this way the whole M shell is removed and the L shell (having 8 electrons in it) becomes the outermost shell. By losing 1 electron, the sodium atom gets 1 unit of positive charge and becomes a sodium ion, Na^+

	Na	$-$	e^-	$\longrightarrow$	Na^+
	Sodium atom		Electron		Sodium ion
Electronic configurations :	K L M 2, 8, 1				K L 2, 8
	(Unstable electron arrangement)				(Stable, neon gas electron arrangement)

The sodium ion (Na^+) has the inert gas electron arrangement of 8 outermost electrons, so it is more stable than a sodium atom. **The electronic configuration of a sodium ion is the same as that of the nearest inert gas neon.**

A proton has 1 unit positive charge whereas an electron has 1 unit negative charge. A sodium atom (Na) contains 11 protons and 11 electrons. **Since the number of protons and electrons in a sodium atom is equal, therefore, it is electrically neutral having no overall charge.** In the sodium ion (Na^+) there are 11 protons but only 10 electrons (because 1 electron has been given out). This means that in a sodium ion there is 1 proton more than electrons. **Due to 1 more proton than electrons, a sodium ion has 1 unit positive charge** (and it is written as Na^+).

The formation of sodium ion can be represented by a diagram as follows :

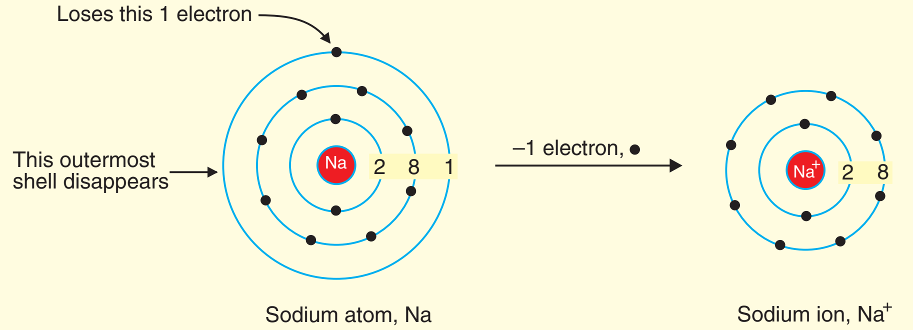


Figure 43. Diagram to show the formation of a sodium ion.

We can see from the above diagram that when a sodium atom loses 1 electron from its outermost shell to form a sodium ion, then its whole outermost shell is removed.

The formation of a potassium ion (K^+) is similar to the formation of a sodium ion because like sodium atom, the potassium atom (K) has also 1 electron in its outermost shell. Knowing that the atomic number of potassium is 19 and its electronic configuration is $\underset{2,\ 8,\ 8,\ 1}{\text{K L M N}}$, explain the formation of a potassium ion yourself. Remember that the noble gas nearest to potassium is argon having atomic number 18 and electronic configuration of $\underset{2,}{\mathrm{K}}\;\underset{8,}{\mathrm{L}}\;\underset{8}{\mathrm{M}}$.

The formation of a lithium ion (Li^+) is also similar to the formation of a sodium ion and a potassium ion. Knowing that the atomic number of lithium (Li) is 3 and its electronic configuration is $\underset{2,}{\mathrm{K}}\;\underset{1}{\mathrm{L}}$, explain the formation of a lithium ion yourself. Please note that the noble gas nearest to lithium is helium having the atomic number 2 and electronic configuration $\underset{2}{\mathrm{K}}$.

Figure 44. This is a rechargeable 'lithium ion battery'.

Figure 45. Lithium ion batteries are used in mobile phones.

A Point to Remember : Look at the electronic configurations of magnesium (atomic number 12), and its nearest inert gas neon (atomic number 10) :

Magnesium atom	*Neon atom*
K L M	K L
2, 8, 2	2, 8

The magnesium atom has 2 electrons more (in the M shell) than a neon atom. So, if a magnesium atom loses its 2 outermost electrons (to some other atom), then it will achieve the electron arrangement of inert gas neon and become more stable. Knowing this point, it will now be easier for us to understand the formation of a magnesium ion.

2. Formation of a Magnesium ion, Mg^{2+}

The atomic number of magnesium is 12, so its electronic configuration is $\underset{2,}{\mathrm{K}}\;\underset{8,}{\mathrm{L}}\;\underset{2}{\mathrm{M}}$. It has 2 electrons in its valence shell (M shell). The magnesium atom donates its 2 outermost electrons (to some other atom) and forms a magnesium ion, Mg^{2+}, having 2 units of positive charge :

$$\underset{\substack{\text{Magnesium atom}\\ \text{K L M}\\ 2,\ 8,\ 2}}{\mathrm{Mg}} \quad - \quad \underset{\text{Electrons}}{2e^-} \quad \longrightarrow \quad \underset{\substack{\text{Magnesium ion}\\ \text{K L}\\ 2,\ 8}}{\mathrm{Mg}^{2+}}$$

The magnesium ion has the inert gas electron structure of 8 electrons in the outermost shell, so it is more stable than a magnesium atom. **The electronic configuration of magnesium ion is the same as that of its nearest inert gas neon.** The number of protons and electrons in a magnesium atom is equal (12 each), so a magnesium atom is electrically neutral. A magnesium ion contains 12 protons but only 10 electrons so it has 2 protons more than electrons. **Since a magnesium ion has 2 protons more than electrons, it has 2 units of positive charge** (and it is written as Mg^{2+}).

The formation of a calcium ion (Ca^{2+}) is similar to the formation of a magnesium ion, because like a magnesium atom, a calcium atom (Ca) has also 2 electrons in its outermost shell. Knowing that the atomic number of calcium is 20, and its electronic configuration is $\underset{2,}{\mathrm{K}}\;\underset{8,}{\mathrm{L}}\;\underset{8,}{\mathrm{M}}\;\underset{2}{\mathrm{N}}$, explain the formation of calcium ion yourself. Remember that the noble gas nearest to calcium is argon having an atomic number of 18 and electronic configuration $\underset{2,}{\mathrm{K}}\;\underset{8,}{\mathrm{L}}\;\underset{8}{\mathrm{M}}$.

A Point to Remember : Look at the electron arrangements of aluminium atom (atomic number 13) and its nearest inert gas neon (atomic number 10) :

Aluminium atom	*Neon atom*
K L M	K L
2, 8, 3	2, 8

The aluminium atom has 3 electrons more than a neon atom, so if an aluminium atom gives its 3 outermost electrons (to some other atom), then it will achieve the electron arrangement of its nearest inert gas neon and become more stable. Keeping this point in mind, it will now be easier for us to understand the formation of an aluminium ion.

3. Formation of an Aluminium Ion, Al^{3+}

The atomic number of aluminium is 13, so its electronic configuration is $\underset{2,\ 8,\ 3}{\text{K L M}}$. The aluminium atom has 3 electrons in its outermost shell which it donates to some other atom and forms an aluminium ion, Al^{3+}, having 3 units of positive charge :

$$\underset{\substack{\text{Aluminium atom}\\ \text{K L M}\\ 2,\ 8,\ 3}}{Al} \quad - \quad \underset{\text{Electrons}}{3e^-} \longrightarrow \underset{\substack{\text{Aluminium ion}\\ \text{K L}\\ 2,\ 8}}{Al^{3+}}$$

The aluminium ion has an inert gas electron structure of 8 electrons in the outermost shell, so it is more stable than an aluminium atom. **The electronic configuration of an aluminium ion is the same as that of its nearest noble gas neon.** An aluminium atom has an equal number of protons and electrons (13 each), so it is electrically neutral. On the other hand, an aluminium ion has 13 protons but only 10 electrons. That is, it has 3 protons more than electrons. **Since an aluminium ion has 3 protons more than electrons, it has 3 units of positive charge** (and it is written as Al^{3+}).

Electron-Dot Representation

Only the outermost electrons of an atom take part in chemical bonding. These are known as valence electrons. **The valence electrons in an atom are represented by putting dots (·) on the symbol of the element, one dot for each valence electron.** For example, sodium atom has 1 valence electron in its outermost shell, so we put 1 dot with the symbol of sodium and write Na· for it. Sodium atom loses this 1 electron to form a sodium ion. Since the sodium ion does not have this valence electron, so we do not put a dot with the sodium ion. We just write Na^+ for it. Magnesium atom has 2 valence electrons so we write Mg**:** for it. Similarly, aluminium atom has 3 valence electrons and we write ·Al**:** for it.

In order to write the electron-dot structures, we should know the number of valence electrons in an atom. The number of valence electrons or outermost electrons can be obtained by writing the electronic configuration of the element. Some of the common metal elements that form positive ions or cations are given below :

Some Common Metal Elements that form Positive Ions (or Cations)

Metal element	*Symbol*	*Atomic number*	*Electronic configuration* K L M N	*No. of outermost electrons*	*Electron-dot structure*	*Ion formed*
1. Sodium	Na	11	2, 8, 1	1	Na·	Na^+
2. Magnesium	Mg	12	2, 8, 2	2	Mg:	Mg^{2+}
3. Aluminium	Al	13	2, 8, 3	3	·Al:	Al^{3+}
4. Potassium	K	19	2, 8, 8, 1	1	K·	K^+
5. Calcium	Ca	20	2, 8, 8, 2	2	Ca:	Ca^{2+}

We will now discuss the formation of negative ions or anions in detail.

FORMATION OF NEGATIVE IONS (OR ANIONS)

If an element has 5, 6 or 7 electrons in the outermost shell of its atom, then it gains (accepts) electrons to achieve the stable, inert gas electron configuration of 8 valence electrons, and forms negatively charged ion called anion (It is not possible to remove 5, 6 or 7 electrons from an atom due to very high energy required). Now, the non-metal atoms have usually 5, 6 or 7 electrons in their outermost shell, so **the non-metal atoms accept electrons to form negative ions or anions.** Fluorine, chlorine, bromine, iodine, oxygen, sulphur, nitrogen and phosphorus, etc., are all non-metals which accept electrons to form negative ions. The element carbon, having 4 electrons in its outermost shell, is also a non-metal but it can neither lose 4 electrons nor gain 4 electrons due to energy considerations. So, a carbon atom does not form ions.

Please note that an atom having 7 electrons in its outermost shell accepts 1 more electron to form an anion having one unit negative charge. An atom having 6 valence electrons accepts 2 more electrons to form an anion having two units negative charge. Similarly, an atom having 5 electrons in its outermost shell accepts 3 more electrons to form an anion having three units negative charge. We will now take some examples to understand how negative ions are formed and what changes take place in the electronic configuration during their formation.

A Point to Remember : Look at the electronic configurations of chlorine (atomic number 17), and its nearest inert gas argon (atomic number 18) :

Chlorine atom	*Argon atom*
K L M	K L M
2, 8, 7	2, 8, 8

The chlorine atom has 7 electrons in its outermost shell whereas an argon atom has 8 electrons in its outermost shell. That is, a chlorine atom has 1 electron less than its nearest inert gas argon. So, if a chlorine atom gains (accepts) 1 electron from some other atom, then it will achieve the 8-electron arrangement of argon and become more stable. Keeping this point in mind, it will now be easier for us to understand the formation of a chloride ion.

1. Formation of a Chloride Ion, Cl^-

The atomic number of chlorine is 17, so its electronic configuration is $\underset{2,\ 8,\ 7}{\text{K L M}}$. We find that chlorine atom has 7 electrons in its outermost shell (M shell). It needs 1 more electron to achieve the stable, 8-electron configuration of an inert gas. So, in order to become more stable, a chlorine atom accepts (gains) 1 electron from some other atom (like sodium atom) and achieves the argon gas configuration of $\underset{2,\ 8,\ 8}{\text{K L M}}$.

By gaining 1 electron, the chlorine atom gets 1 unit of negative charge and forms a chloride ion, Cl^-

	Cl	+	e^-	$\longrightarrow$	Cl^-
	Chlorine atom		Electron		Chloride ion
Electronic	K L M				K L M
configurations :	2, 8, 7				2, 8, 8
	(Unstable electron arrangement)				(Stable, argon gas electron arrangement)

The chloride ion, Cl^-, has an inert gas electronic configuration of 8-outermost electrons, so it is more stable than a chlorine atom. **The electronic configuration of a chloride ion is the same as that of its nearest inert gas argon.**

A chlorine atom (Cl) contains 17 protons and 17 electrons. **Since the number of protons and electrons in a chlorine atom is equal, therefore, it is electrically neutral, having no overall charge.** In the chloride ion (Cl^-), there are 17 protons but 18 electrons (because 1 extra electron has been added). This means that in a chloride ion, there is 1 electron more than protons. **Due to 1 more electron than protons, a chloride ion**

has 1 unit negative charge (and it is written as Cl^-).

The formation of a chloride ion can be represented by a diagram as follows :

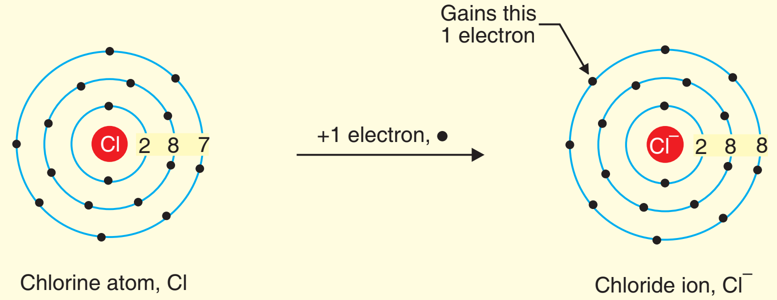


Figure 46. Diagram to show the formation of a chloride ion.

We can see from the above diagram that the extra electron is added to the outermost shell of the chlorine atom to form a chloride ion.

The formation of a fluoride ion (F^-) is similar to the formation of a chloride ion because like a chlorine atom, a fluorine atom (F) has also 7 electrons in its outermost shell. Knowing that the atomic number of fluorine is 9, and its electronic configuration is $\underset{2,\ 7}{\text{K L}}$, please explain the formation of a fluoride ion yourself. Remember that the inert gas nearest to fluorine is neon having an atomic number of 10 and electronic configuration $\underset{2,\ 8}{\text{K L}}$. The other halogens, bromine (Br) and iodine (I), have also 7 valence electrons each in their atoms, and accept 1 electron each to form bromide ion (Br^-), and iodide ion (I^-), respectively.

A Point to Remember : Look at the electronic configurations of an oxygen atom (atomic number 8) and its nearest noble gas neon (atomic number 10) :

Oxygen atom	*Neon atom*
K L	K L
2, 6	2, 8

Oxygen atom has 6 electrons in its valence shell whereas neon atom has 8 electrons in its valence shell. That is, oxygen atom has 2 electrons less than a neon atom. So, if an oxygen atom takes 2 electrons from some other atom, it will achieve the electron arrangement of inert gas neon and become more stable. Keeping this point in mind, it will now be easier for us to understand the formation of an oxide ion.

2. Formation of an Oxide Ion, O^{2-}

The atomic number of oxygen is 8, so its electronic configuration is $\underset{2,\ 6}{\text{K L}}$. The oxygen atom has 6 electrons in its outermost shell, so it needs 2 more electrons to achieve the stable, 8-electron inert gas structure. By taking 2 electrons from some other atom, the oxygen atom forms an oxide ion, O^{2-}, having 2 units of negative charge :

$$\underset{\substack{\text{Oxygen atom}\\ \text{K L}\\ 2,\ 6}}{O} \quad + \quad \underset{\text{Electrons}}{2e^-} \longrightarrow \underset{\substack{\text{Oxide ion}\\ \text{K L}\\ 2,\ 8}}{O^{2-}}$$

The oxide ion has an inert gas electron arrangement of 8 electrons in the outermost shell, so it is more stable than an oxygen atom. **The electronic configuration of an oxide ion is the same as that of a neon atom.** The oxygen atom has an equal number of protons and electrons (8 each), so it is electrically neutral. An oxide ion has 8 protons but 10 electrons. **Since an oxide ion has 2 electrons more than protons, it has 2 units of negative charge** (and it is written as O^{2-}).

The formation of a sulphide ion (S^{2-}) is similar to the formation of an oxide ion because like an oxygen atom, a sulphur atom (S) has also 6 electrons in its outermost shell. Knowing that the atomic number of sulphur is 16, and its electronic configuration is $\underset{2,\ 8,\ 6}{\text{K L M}}$, please explain the formation of a sulphide ion yourself. Remember that the inert gas nearest to sulphur is argon having an atomic number of 18 and electronic configuration $\underset{2,\ 8,\ 8}{\text{K L M}}$.

A Point to Remember : Compare the electronic configurations of a nitrogen atom (atomic number 7), and its nearest inert gas neon (atomic number 10) given below :

Nitrogen atom	*Neon atom*
K L	K L
2, 5	2, 8

The nitrogen atom has 5 electrons in its outermost shell whereas a neon atom has 8 electrons in its outermost shell. Thus, a nitrogen atom has 3 electrons less than its nearest inert gas neon. So, if a nitrogen atom accepts 3 electrons from some other atom, then it will achieve the electronic configuration of inert gas neon and become more stable. Knowing this point, it will now be easier for us to understand the formation of a nitride ion.

3. Formation of a Nitride ion, N^{3-}

The atomic number of nitrogen is 7 so its electronic configuration is $\underset{2,\ 5}{\text{K L}}$. Nitrogen atom has 5 valence electrons, so it needs 3 more electrons to achieve the 8-electron inert gas structure. So, by taking 3 electrons from some other atom, a nitrogen atom forms a nitride ion, N^{3-}, having 3 units of negative charge :

$$\begin{array}{ccccc} \text{N} & + & 3e^- & \longrightarrow & \text{N}^{3-} \\ \text{Nitrogen atom} & & \text{Electrons} & & \text{Nitride ion} \\ \text{K L} & & & & \text{K L} \\ 2,\ 5 & & & & 2,\ 8 \end{array}$$

The nitride ion has an inert gas electron arrangement of 8 electrons in the outermost shell, so it is more stable than a nitrogen atom. **The electronic configuration of a nitride ion is the same as that of a neon atom.** A nitrogen atom contains an equal number of protons and electrons (7 each), so it is electrically neutral. On the other hand, a nitride ion contains 7 protons but 10 electrons. **Since a nitride ion contains 3 electrons more than protons, it has 3 units of negative charge** (and it is written as N^{3-}).

Figure 47. Because of its metallic gold colour, titanium nitride (TiN) is used for giving decorative coatings to iron and steel objects.

The formation of a phosphide ion (P^{3-}) is similar to the formation of a nitride ion because like a nitrogen atom, a phosphorus atom (P) has also 5 electrons in its outermost shell. Knowing that the atomic number of phosphorus is 15 and its electronic configuration is $\underset{2,\ 8,\ 5}{\text{K L M}}$, please explain the formation of a phosphide ion yourself.

Electron-Dot Representation

A chlorine atom has 7 electrons in its outermost shell, so we put 7 dots with its symbol and write $\cdot\underset{\cdot\cdot}{\ddot{\text{Cl}}}:$ for it. When a chlorine atom accepts 1 more electron to form a chloride ion, then this chloride ion has 8 electrons in the outermost shell. So we put 8 dots and write $\left[:\underset{\cdot\cdot}{\ddot{\text{Cl}}}:\right]^-$ for a chloride ion. The electron-dot structures for other non-metal elements and their anions can be written in a similar way as shown in the table on the next page.

Some Common Non-metal Elements that form Negative Ions (or Anions)

Non-metal element	*Symbol*	*Atomic number*	*Electronic configuration* K L M	*No. of outermost electrons*	*Electron-dot structure*	*Ion formed*
1. Fluorine	F	9	2, 7	7	$\cdot\ddot{\underset{\cdot\cdot}{F}}:$	$[:\ddot{\underset{\cdot\cdot}{F}}:]^-$ Fluoride ion, F^-
2. Chlorine	Cl	17	2, 8, 7	7	$\cdot\ddot{\underset{\cdot\cdot}{Cl}}:$	$[:\ddot{\underset{\cdot\cdot}{Cl}}:]^-$ Chloride ion, Cl^-
3. Oxygen	O	8	2, 6	6	$:\ddot{O}:$	$[:\ddot{\underset{\cdot\cdot}{O}}:]^{2-}$ Oxide ion, O^{2-}
4. Sulphur	S	16	2, 8, 6	6	$:\ddot{S}:$	$[:\ddot{\underset{\cdot\cdot}{S}}:]^{2-}$ Sulphide ion, S^{2-}
5. Nitrogen	N	7	2, 5	5	$\cdot\ddot{N}:$	$[:\ddot{\underset{\cdot\cdot}{N}}:]^{3-}$ Nitride ion, N^{3-}

Types of Chemical Bonds

There are two types of chemical bonds :

(*i*) Ionic bond, and

(*ii*) Covalent bond.

Ionic bonds are formed by the transfer of electrons from one atom to another whereas **covalent bonds are formed by the sharing of electrons between two atoms.** When a chemical bond is formed between the atoms, then both the combining atoms acquire the stable, inert gas electron configuration. We will now discuss the ionic bond and covalent bond in detail, one by one. Please note that **ionic bond is also called electrovalent bond.** The name electrovalent bond is derived from the fact that there are electrical charges on the atoms involved in the bond formation.

Figure 48. Common salt is an ionic compound containing ionic bonds.

Figure 49. Cane sugar is a covalent compound containing covalent bonds.

IONIC BOND

The chemical bond formed by the transfer of electrons from one atom to another is known as an ionic bond. The transfer of electrons takes place in such a way that the ions formed have the stable electron arrangement of an inert gas. The ionic bond is called so because it is a chemical bond between oppositely charged ions. Before we give examples to understand the formation of ionic bonds, we should know what type of elements form ionic bonds. This is discussed below.

An ionic bond is formed when one of the atoms can donate electrons to achieve the inert gas electron configuration, and the other atom needs electrons to achieve the inert gas electron configuration. Now, the metal atoms have usually 1, 2 or 3 electrons in their outermost shells which they can donate to form stable positive ions. On the other hand, non-metal atoms have usually 5, 6 or 7 electrons in their outermost shells, so they need electrons to form stable negative ions. Thus, **when a metal reacts with a non-metal,**

transfer of electrons takes place from metal atoms to the non-metal atoms, and an ionic bond is formed. For example, sodium is a metal and chlorine is a non-metal, so when sodium reacts with chlorine to form sodium chloride, transfer of electrons takes place from sodium atoms to chlorine atoms, and an ionic bond is formed. It is obvious that **the ionic bonds are formed between metals and non-metals.**

In the formation of an ionic bond between a metal and a non-metal, the metal atom donates one or more electrons to the non-metal atom. By losing electrons, the metal atom forms a positively charged ion (cation). The non-metal atom accepts electrons (donated by the metal atom) and forms a negatively charged ion (anion). The positive ions and negative ions attract one another. **The strong force of attraction developed between the oppositely charged ions is known as an ionic bond**. The compounds containing ionic bonds are called ionic compounds. **Ionic compounds are made up of ions**. We will now describe the formation of some ionic compounds such as sodium chloride, magnesium chloride and magnesium oxide, etc.

1. Formation of Sodium Chloride

Sodium is a metal whereas chlorine is a non-metal. **Sodium metal reacts with chlorine to form an ionic compound, sodium chloride.** We will now explain how sodium chloride is formed and what changes take place in the electron arrangements of sodium and chlorine atoms in the formation of this compound.

The atomic number of sodium is 11, so its electronic configuration is 2, 8, 1. Sodium atom has only 1 electron in its outermost shell. So, the sodium atom donates 1 electron (to a chlorine atom) and forms a sodium ion, Na^+

$$\underset{\substack{\text{Sodium atom}\\ 2,\,8,\,1}}{Na\cdot} \quad - \quad \underset{\text{Electron}}{e^-} \quad \longrightarrow \quad \underset{\substack{\text{Sodium ion}\\ 2,\,8}}{Na^+}$$

The atomic number of chlorine is 17, so its electronic configuration is 2, 8, 7. Chlorine atom has 7 electrons in its outermost shell and needs 1 more electron to achieve the stable, 8-electron inert gas configuration. So, a chlorine atom takes 1 electron (from the sodium atom) and forms a negatively charged chloride ion, Cl^-

$$\underset{\substack{\text{Chlorine atom}\\ 2,\,8,\,7}}{\cdot\ddot{\underset{\cdot\cdot}{Cl}}\!:} \quad + \quad \underset{\text{Electron}}{e^-} \quad \longrightarrow \quad \underset{\substack{\text{Chloride ion}\\ 2,\,8,\,8}}{\left[:\ddot{\underset{\cdot\cdot}{Cl}}:\right]^-}$$

When sodium reacts with chlorine, it transfers its 1 outermost electron to the chlorine atom. By losing 1 electron, sodium atom forms a sodium ion (Na^+) and by gaining 1 electron, the chlorine atom forms a chloride ion (Cl^-). This is shown below :

Electron transfer

$$\underset{\substack{\text{Sodium atom}\\ 2,\,8,\,1}}{Na\cdot} \quad + \quad \underset{\substack{\text{Chlorine atom}\\ 2,\,8,\,7}}{\cdot\ddot{\underset{\cdot\cdot}{Cl}}\!:} \quad \longrightarrow \quad \underbrace{\underset{\substack{\text{Sodium ion}\\ 2,\,8}}{Na^+} \quad \underset{\substack{\text{Chloride ion}\\ 2,\,8,\,8}}{\left[:\ddot{\underset{\cdot\cdot}{Cl}}:\right]^-}}_{\text{Sodium chloride}} \quad \text{or} \quad NaCl$$

Sodium ions have positive charge whereas chloride ions have negative charge. Due to opposite charges, sodium ions and chloride ions are held together by the electrostatic force of attraction to form sodium chloride, $Na^+ Cl^-$ or NaCl.

In sodium chloride compound, the electronic configuration of sodium ion (2, 8) resembles that of inert gas neon, and the electronic configuration of chloride ion (2, 8, 8) resembles that of another inert gas argon. Due to this, the sodium chloride compound is very stable.

The formation of sodium chloride can be shown more clearly with the help of a diagram shown in Figure 50. It is obvious from this diagram that in the formation of sodium chloride compound, one electron

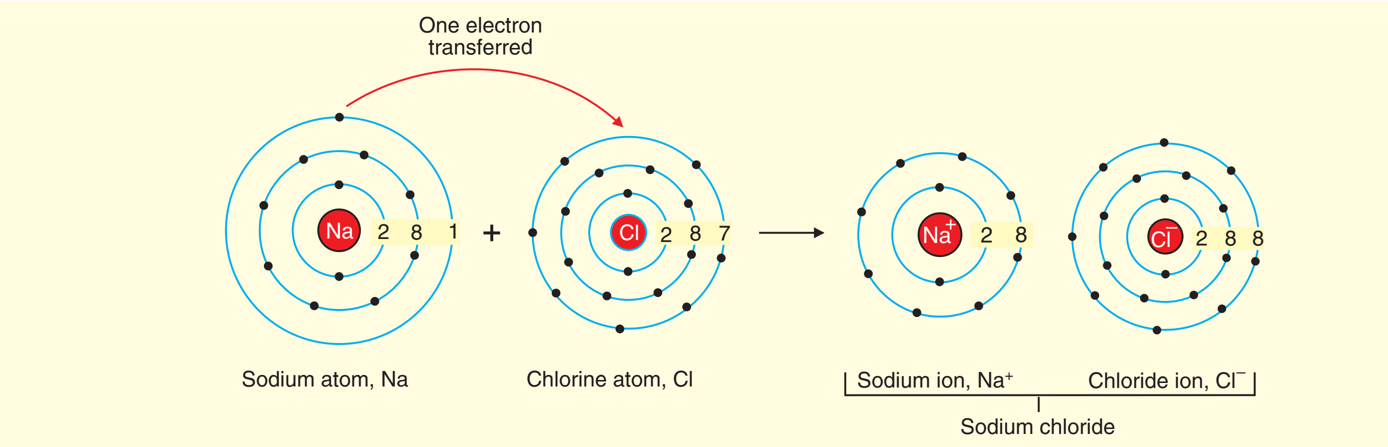


Figure 50. Diagram to show the formation of sodium chloride.

is transferred from each sodium atom to each chlorine atom resulting in the formation of oppositely charged sodium ions and chloride ions. Thus, **sodium chloride is an ionic compound and contains ionic bonds**. It should be noted that **in the formation of ionic bonds, the reacting atoms achieve the inert gas electron configuration by the transfer of electrons.** By convention, the formulae of ionic compounds are written with the positive ion first. Another point to be noted is that the charges on the ions of an ionic compound are usually not written in the formula. For example, sodium chloride is written as NaCl and not as Na^+Cl^-. Please note that sodium chloride does not consist of molecules like NaCl or Na^+Cl^- made up of one sodium ion and one chloride ion. Sodium chloride consists of a large aggregate of an equal number of sodium ions, Na^+, and chloride ions, Cl^-, so the actual formula of sodium chloride should be $(Na^+)_n(Cl^-)_n$ or $(Na^+Cl^-)_n$, where n is a very large number. NaCl is the simplest formula of sodium chloride and not its actual formula.

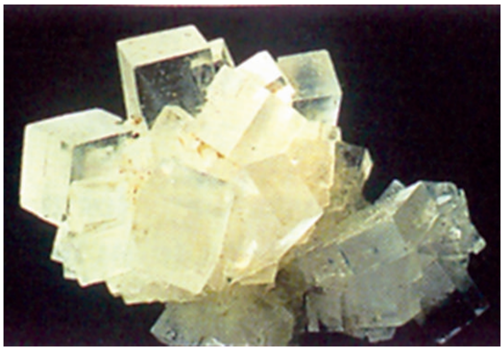


Figure 51. These are crystals of the ionic compound 'sodium chloride' made up of a large number of sodium ions and chloride ions held together.

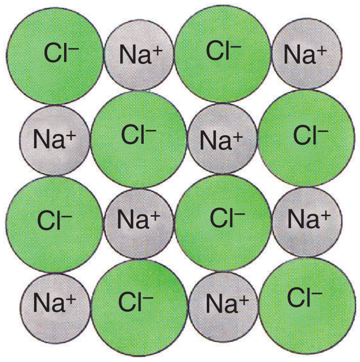


Figure 52. This is how sodium ions and chloride ions are arranged in a sodium chloride crystal.

The formation of potassium chloride (KCl) is similar to the formation of sodium chloride which has been discussed above. Knowing that the atomic number of potassium is 19 and that of chlorine is 17, explain the formation of potassium chloride yourself. The electron-dot representation for the formation of potassium chloride is given below :

Electron transfer

$$K\cdot \;+\; \cdot\underset{\cdot\cdot}{\overset{\cdot\cdot}{Cl}}: \;\longrightarrow\; K^+ \quad \left[:\underset{\cdot\cdot}{\overset{\cdot\cdot}{Cl}}:\right]^- \quad \text{or} \quad KCl$$

Potassium atom	Chlorine atom	Potassium ion	Chloride ion
2, 8, 8, 1	2, 8, 7	2, 8, 8	2, 8, 8

$$\underbrace{K^+ \;(2, 8, 8) \quad Cl^- \;(2, 8, 8)}_{\text{Potassium chloride}}$$

Please note that in potassium chloride, the electronic configurations of both, potassium ion as well as chloride ion (2, 8, 8), resemble that of inert gas argon. Another point to be noted is that **the formation of fluorides, bromides and iodides of alkali metals is similar to the formation of chlorides.** This is because like chlorine, the other halogens, fluorine, bromine and iodine, also have 7 electrons each in their outermost shells. We will now discuss the formation of another ionic compound, magnesium chloride.

2. Formation of Magnesium Chloride

Magnesium is a metal whereas chlorine is a non-metal. **Magnesium reacts with chlorine to form an ionic compound magnesium chloride.** We will now explain how magnesium chloride is formed and what changes take place in the electronic configurations of magnesium and chlorine atoms in the formation of this compound.

The atomic number of magnesium is 12, so its electronic configuration is 2, 8, 2. It has 2 valence electrons. A magnesium atom donates its 2 valence electrons (to two chlorine atoms) and forms a stable magnesium ion, Mg^{2+}

$$\underset{\substack{\text{Magnesium atom}\\ 2,\ 8,\ 2}}{Mg{:}} \quad - \quad \underset{\text{Electrons}}{2e^-} \quad \longrightarrow \quad \underset{\substack{\text{Magnesium ion}\\ 2,\ 8}}{Mg^{2+}}$$

The atomic number of chlorine is 17, and its electronic configuration is 2, 8, 7. Chlorine atom has 7 valence electrons, so it requires only 1 more electron to complete its octet. Since one magnesium atom donates 2 electrons, so two chlorine atoms take these two electrons and form two chloride ions :

$$\underset{\substack{\text{Two chlorine atoms}\\ 2\ (2,\ 8,\ 7)}}{2\ \cdot\ddot{\underset{..}{Cl}}{:}} \quad + \quad \underset{\text{Electrons}}{2e^-} \quad \longrightarrow \quad \underset{\substack{\text{Two chloride ions}\\ 2\ (2,\ 8,\ 8)}}{2\left[{:}\ddot{\underset{..}{Cl}}{:}\right]^-}$$

When magnesium reacts with chlorine, the magnesium atom transfers its two outermost electrons to two chlorine atoms. By losing 2 electrons, the magnesium atom forms a magnesium ion (Mg^{2+}), and by gaining 2 electrons, the two chlorine atoms form two chloride ions ($2Cl^-$). This is shown below :

Electron transfer

$$\underset{\substack{\text{Magnesium}\\ \text{atom}\\ 2,\ 8,\ 2}}{Mg{:}} \quad + \quad \underset{\substack{\text{Two chlorine}\\ \text{atoms}\\ 2\ (2,\ 8,\ 7)}}{\begin{matrix}\cdot\ddot{\underset{..}{Cl}}{:}\\ \cdot\ddot{\underset{..}{Cl}}{:}\end{matrix}} \quad \longrightarrow \quad \underbrace{\underset{\substack{\text{Magnesium}\\ \text{ion}\\ 2,\ 8}}{Mg^{2+}} \quad \underset{\substack{\text{Two chloride}\\ \text{ions}\\ 2\ (2,\ 8,\ 8)}}{\begin{matrix}\left[{:}\ddot{\underset{..}{Cl}}{:}\right]^-\\ \left[{:}\ddot{\underset{..}{Cl}}{:}\right]^-\end{matrix}}}_{\text{Magnesium chloride}} \quad \text{or} \quad Mg^{2+}\,2\left[{:}\ddot{\underset{..}{Cl}}{:}\right]^- \quad \text{or} \quad MgCl_2$$

Electron transfer

The positively charged magnesium ions and negatively charged chloride ions are held together by electrostatic force of attraction to form magnesium chloride compound.

We can see from the above equation that a magnesium ion, Mg^{2+}, has 2 units of positive charge whereas a chloride ion, Cl^-, has only 1 unit of negative charge. So, one magnesium ion, Mg^{2+}, combines with two chloride ions, $2Cl^-$, to form magnesium chloride compound $Mg^{2+}2Cl^-$ or $MgCl_2$. Thus, for each magnesium ion, there are two chloride ions and the formula of magnesium chloride becomes $MgCl_2$.

Figure 53. Magnesium chloride ($MgCl_2$) is an ionic compound containing ionic bonds which is made up of magnesium ions (Mg^{2+}) and chloride ions (Cl^-) held together by electrostatic force of attraction.

Please note that in magnesium chloride compound, the electron arrangement of magnesium ion (2, 8) resembles that of a neon atom whereas the electron arrangement of each chloride ion (2, 8, 8) resembles that of an argon atom. This makes the magnesium chloride compound very stable. Another point to be noted is that the magnesium ion and chloride ions have opposite charges, so they attract one another. The force of attraction between magnesium ion and chloride ions is very strong. It is called an ionic bond. Thus, **magnesium chloride contains ionic bonds**.

The formation of calcium chloride ($CaCl_2$) is similar to the formation of magnesium chloride which has been discussed above. Knowing that the atomic number of calcium is 20 and that of chlorine is 17, explain the formation of calcium chloride yourself. The electron-dot representation for the formation of calcium chloride is given below :

$Ca{:}$	$+$	$2\ {\cdot}\ddot{\underset{\cdot\cdot}{Cl}}{:}$	$\longrightarrow$	Ca^{2+}	$2\left[{:}\ddot{\underset{\cdot\cdot}{Cl}}{:}\right]^{-}$	or	$CaCl_2$
Calcium atom		Two chlorine atoms		Calcium ion	Two chloride ions		
2, 8, 8, 2		2 (2, 8, 7)		2, 8, 8	2 (2, 8, 8)		
				Calcium chloride			

The positively charged calcium ions and negatively charged chloride ions are held together by electrostatic force of attraction. So, **the chemical bond present in calcium chloride ($CaCl_2$) is an ionic bond.**

3. Formation of Magnesium Oxide

Magnesium is a metal whereas oxygen is a non-metal. **Magnesium metal burns in oxygen to form an ionic compound magnesium oxide.** We will now explain how magnesium oxide is formed and what changes take place in the electronic configurations of magnesium and oxygen atoms during the formation of this compound.

The atomic number of magnesium is 12, so its electronic configuration is 2, 8, 2. We see that the magnesium atom has 2 electrons in its outermost shell. So, the magnesium atom donates its 2 outermost electrons (to an oxygen atom) and forms a stable magnesium ion, Mg^{2+}, having the electron arrangement of a neon atom :

$Mg{:}$	$-$	$2e^-$	$\longrightarrow$	Mg^{2+}
Magnesium atom		Electrons		Magnesium ion
2, 8, 2				2, 8

The atomic number of oxygen is 8, so its electronic configuration is 2, 6. We find that oxygen atom has 6 electrons in its outermost shell so it requires 2 more electrons to achieve the stable, 8-electron structure of an inert gas. Thus, an oxygen atom accepts 2 electrons (donated by a magnesium atom) and forms a stable oxide ion, O^{2-}, having the electron arrangement of a neon atom :

$\ddot{\underset{\cdot\cdot}{O}}{:}$	$+$	$2e^-$	$\longrightarrow$	$\left[{:}\ddot{\underset{\cdot\cdot}{O}}{:}\right]^{2-}$
Oxygen atom		Electrons		Oxide ion
2, 6				2, 8

When magnesium reacts with oxygen, the magnesium atom transfers its two outermost electrons to an oxygen atom. By losing 2 electrons, the magnesium atom forms a magnesium ion (Mg^{2+}), and by gaining 2 electrons, the oxygen atom forms an oxide ion (O^{2-}). This is shown below :

Transfer of electrons

$Mg{:}$	$+$	$\ddot{\underset{\cdot\cdot}{O}}{:}$	$\longrightarrow$	Mg^{2+}	$\left[{:}\ddot{\underset{\cdot\cdot}{O}}{:}\right]^{2-}$	or	MgO
Magnesium atom		Oxygen atom		Magnesium ion	Oxide ion		
2, 8, 2		2, 6		2, 8	2, 8		
				Magnesium oxide			

We find that the magnesium ion has 2 units of positive charge whereas the oxide ion has 2 units of negative charge. The oppositely charged magnesium ions, Mg^{2+}, and oxide ions, O^{2-}, are held together by a strong force of electrostatic attraction to form magnesium oxide compound $Mg^{2+}O^{2-}$ or MgO. Thus, **magnesium oxide contains ionic bonds.**

Please note that in magnesium oxide compound, the electronic configurations of magnesium ion as well as the oxide ion (2, 8), resemble the electronic configuration of the inert gas neon.

Calcium reacts with oxygen to form an ionic compound calcium oxide, CaO. **The formation of calcium oxide is similar to the formation of magnesium oxide** which has been discussed above. Knowing that the atomic number of calcium is 20 and that of oxygen is 8, please explain the formation of calcium oxide compound yourself. The electron-dot representation of the reaction for the formation of calcium oxide is given below :

$$\mathrm{Ca\!:} \;+\; \mathrm{\ddot{\underset{\cdot\cdot}{O}}\!:} \;\longrightarrow\; \mathrm{Ca^{2+}} \quad \left[\mathrm{:\ddot{\underset{\cdot\cdot}{O}}:}\right]^{2-} \quad \text{or} \quad \mathrm{CaO}$$

Calcium atom	Oxygen atom	Calcium ion	Oxide ion
2, 8, 8, 2	2, 6	2, 8, 8	2, 8
		$\underbrace{\qquad\qquad\qquad}_{\text{Calcium oxide}}$	

Please note that in the calcium oxide compound, the electronic configuration of a calcium ion, Ca^{2+}, is 2, 8, 8 which is the same as that of inert gas argon. But the electronic configuration of the oxide ion, O^{2-}, is 2, 8 which resembles that of another inert gas neon.

Ionic Compounds

The compounds containing *ionic bonds* are known as *ionic compounds*. They are formed by the transfer of electrons from one atom to another. The ionic compounds are made up of positively charged ions (cations) and negatively charged ions (anions). That is, **the ionic compounds consist of ions and not molecules.** Some of the common ionic compounds and the ions of which they are made, are given below. Please note that *ionic compounds* are also known as *electrovalent compounds*.

Some Ionic Compounds (or Electrovalent Compounds)

Name	*Formula*	*Ions present*
1. Sodium chloride	$NaCl$	Na^+ and Cl^-
2. Potassium chloride	KCl	K^+ and Cl^-
3. Ammonium chloride	NH_4Cl	NH_4^+ and Cl^-
4. Magnesium chloride	$MgCl_2$	Mg^{2+} and Cl^-
5. Calcium chloride	$CaCl_2$	Ca^{2+} and Cl^-
6. Sodium oxide	Na_2O	Na^+ and O^{2-}
7. Magnesium oxide	MgO	Mg^{2+} and O^{2-}
8. Calcium oxide	CaO	Ca^{2+} and O^{2-}
9. Aluminium oxide	Al_2O_3	Al^{3+} and O^{2-}
10. Sodium hydroxide	$NaOH$	Na^+ and OH^-
11. Copper sulphate	$CuSO_4$	Cu^{2+} and SO_4^{2-}
12. Calcium nitrate	$Ca(NO_3)_2$	Ca^{2+} and NO_3^-

Please note that all the above ionic compounds are made up of a metal and a non-metal (except ammonium chloride which is an ionic compound made up of only non-metals). So, **whenever we see a compound made up of a metal and a non-metal, we should at once say that it is an ionic compound and contains ionic bonds.** We will now answer a question based on the formation of an ionic compound.

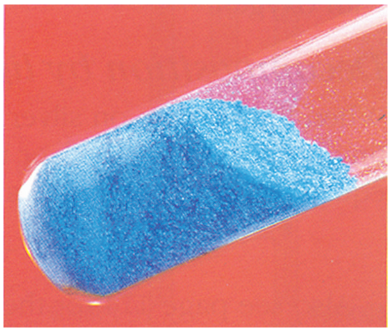


Figure 54. Copper sulphate ($CuSO_4$) is an ionic compound made up of copper ions (Cu^{2+}) and sulphate ions (SO_4^{2-}).

Sample Problem. (*i*) Write the electron-dot structures for sodium and oxygen.

(*ii*) Show the formation of sodium oxide (Na_2O) by the transfer of electrons.

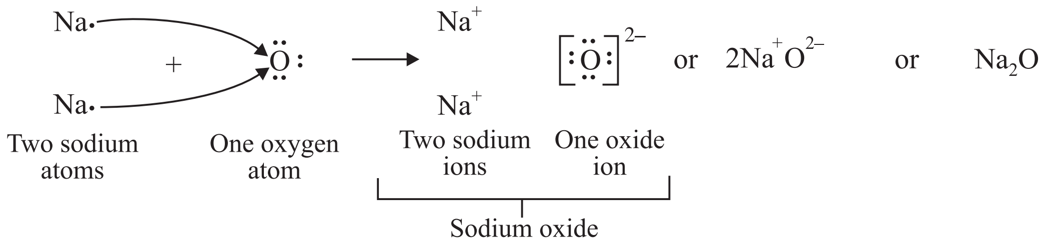
(*iii*) What are the ions present in this compound ?

(NCERT Book Question)

Solution. (*i*) Sodium atom has 1 electron in its outermost shell, so the electron dot structure of sodium is Na· (1 dot on the symbol Na). Oxygen atom has 6 electrons in its outermost shell, so the electron-dot structure of oxygen is :Ö: (6 dots on the symbol O).

(*ii*) The formation of sodium oxide (Na_2O) can be explained as follows : A sodium atom has 1 outermost electron to donate but an oxygen atom requires 2 electrons to achieve the 8-electron structure. So, two sodium atoms will combine with one oxygen atom to form sodium oxide compound.

In the formation of sodium oxide, two sodium atoms transfer their 2 outermost electrons to an oxygen atom. By losing 2 electrons, the two sodium atoms form two sodium ions ($2Na^+$). And by gaining 2 electrons, the oxygen atom forms an oxide ion (O^{2-}) :



The oppositely charged sodium ions and oxide ion are held together by strong electrostatic forces of attraction to form the ionic sodium oxide compound $2Na^+O^{2-}$ or Na_2O.

(*iii*) The ions present in sodium oxide compound (Na_2O) are : sodium ions ($2Na^+$) and oxide ion (O^{2-}).

COVALENT BOND

The chemical bond formed by the sharing of electrons between two atoms is known as a covalent bond. *The sharing of electrons takes place in such a way that each atom in the resulting molecule gets the stable electron arrangement of an inert gas.* It should be noted that the atoms share only their *outermost electrons* in the formation of covalent bonds. Before we give examples to understand the formation of covalent bonds, we should know what type of elements form covalent bonds. This is discussed below.

A covalent bond is formed when both the reacting atoms need electrons to achieve the inert gas electron arrangement. Now, the non-metals have usually 5, 6 or 7 electrons in the outermost shells of their atoms (except carbon which has 4 and hydrogen which has just 1 electron in the outermost shell). So, all the non-metal atoms need electrons to achieve the inert gas structure. They get these electrons by mutual sharing. Thus, **whenever a non-metal combines with another non-metal, sharing of electrons takes place between their atoms and a covalent bond is formed.** For example, hydrogen is a non-metal and chlorine is also a non-metal, so when hydrogen combines with chlorine to form hydrogen chloride, HCl, sharing of electrons takes place between hydrogen and chlorine atoms and a covalent bond is formed. It should be noted that **a covalent bond can also be formed between two atoms of the same non-metal.** For example, two chlorine atoms combine together by the sharing of electrons to form a chlorine molecule, Cl_2, and a covalent bond is formed between the two chlorine atoms. From this we conclude that **the bond formed between the atoms of the same element is a covalent bond.**

In the formation of a covalent bond between two non-metals, each non-metal atom shares one or more electrons with the other non-metal atom. **The shared electrons are counted with both the atoms due to which each atom in the resulting molecule gets an inert gas electron arrangement of 8 electrons (or 2 electrons) in the outermost shell.** The shared electron pair constitutes the covalent bond.

Covalent bonds are of three types :

(*i*) Single covalent bond
(*ii*) Double covalent bond
(*iii*) Triple covalent bond

We will now discuss the formation of these three types of covalent bonds in detail. Let us take the single bond first.

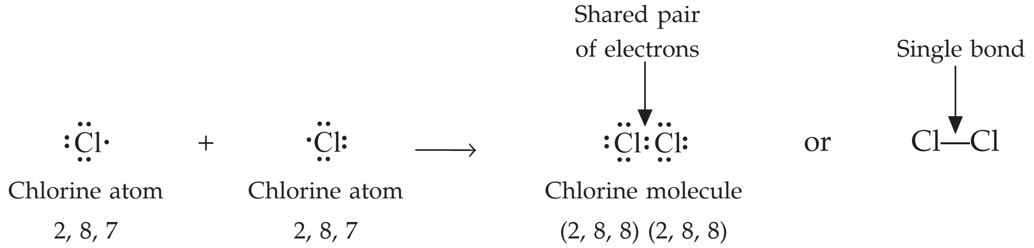
SINGLE BOND

A single covalent bond consists of one pair of shared electrons. In other words, **a single bond is formed by the sharing of one pair of electrons between two atoms.** Now, one pair of electrons means 2 electrons, so we can also say that **a single covalent bond is formed by the sharing of 2 electrons between the atoms, each atom contributing one electron for sharing.** *The shared electron pair is always drawn between the two atoms.* For example, a hydrogen molecule H_2, contains a single covalent bond and it is written as H **:** H, the two dots drawn between the hydrogen atoms represent a pair of shared electrons which constitutes the single bond. A single covalent bond is denoted by putting a short line (—) between the two atoms. So, a hydrogen molecule can also be written as H—H. **The short line between the two hydrogen atoms represents a single covalent bond consisting of two shared electrons, one from each hydrogen atom.** A chlorine molecule, hydrogen chloride, methane, carbon tetrachloride, water and ammonia, all contain single covalent bonds. Let us discuss the formation of these molecules in detail.

1. Formation of a Chlorine Molecule, Cl_2

A chlorine atom is very reactive and cannot exist free because it does not have the stable electron arrangement of an inert gas. Chlorine gas, therefore, does not consist of single atoms, it consists of more stable Cl_2 molecules. **Each molecule of chlorine contains two chlorine atoms joined by a single covalent bond.** We will now explain the formation of a chlorine molecule on the basis of electronic theory of valency.

The atomic number of chlorine is 17, so its electronic configuration is 2, 8, 7. Chlorine atom has 7 electrons in its outermost shell and needs 1 more electron to complete its octet and become stable. It gets this electron by sharing with another chlorine atom. So, **two chlorine atoms share one electron each to form a chlorine molecule :**



Because the two chlorine atoms share electrons, there is a strong force of attraction between them which holds them together. This force is called a covalent bond. The bonded chlorine atoms thus form a chlorine molecule. Since the two chlorine atoms share one pair of electrons, the bond between them is called a single covalent bond or just a single bond.

The two shared electrons are counted with both the chlorine atoms for the purpose of determining the inert gas configuration. For example, in the chlorine molecule, each chlorine atom has now 8 outermost electrons (7 its own and 1 shared from other atom). In fact, **each chlorine atom in the chlorine molecule has the electronic configuration 2, 8, 8 resembling its nearest inert gas argon.** Since the chlorine atoms in a chlorine molecule have inert gas electron arrangements, therefore, a chlorine molecule is more stable than two separate chlorine atoms.

The formation of a chlorine molecule by the sharing of electrons between two chlorine atoms can also be shown by means of a diagram. Since the atoms share only their outermost electrons with one another, therefore, only the outermost electrons of each atom are shown in the diagram. For example, the combination of two chlorine atoms by the sharing of electrons to form a covalent chlorine molecule can be shown by the diagram given on the next page.

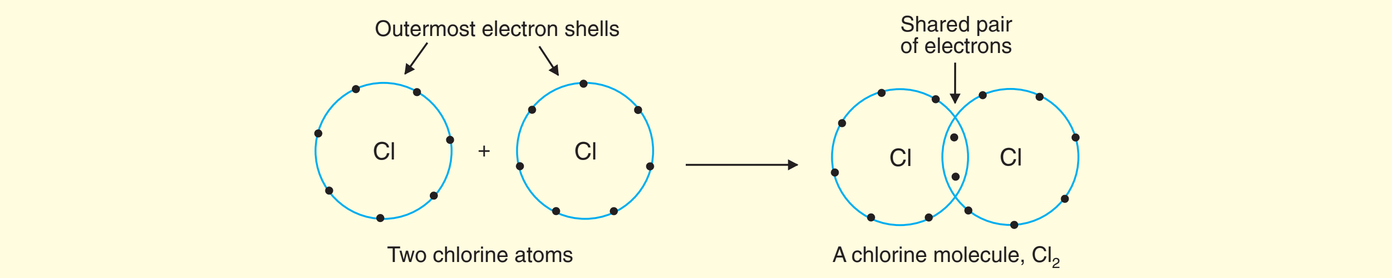


Figure 55. Diagram to show the formation of a chlorine molecule.

In the above example, two atoms of the same element, chlorine, combine to form a molecule containing a covalent bond. In general, **whenever two atoms of the same element combine to form a molecule, a covalent bond is formed.** Please remember that **in the formation of a covalent bond (or a covalent compound), the reacting atoms achieve the inert gas electron arrangement by the sharing of electrons.** There is no transfer of electrons in the formation of a covalent bond.

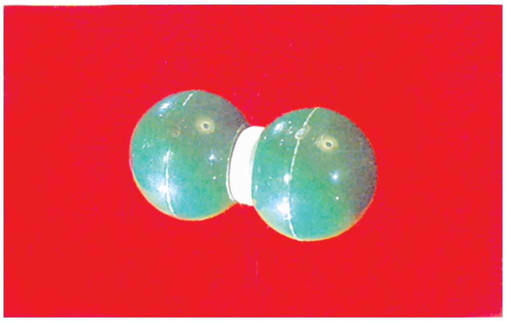


Figure 56. Model of a chlorine molecule. The two green balls represent two chlorine atoms joined together.

2. Formation of a Hydrogen Molecule, H_2

A hydrogen atom is very reactive and cannot exist free because it does not have the stable, inert gas electron arrangement. So, hydrogen gas does not consist of single atoms, it consists of more stable H_2 molecules. Each molecule of hydrogen gas has two hydrogen atoms joined by a covalent bond. We will now explain the formation of a hydrogen molecule in detail.

The atomic number of hydrogen is 1, so its electronic configuration is $\underset{1}{K}$. Hydrogen atom has only 1 electron in the outermost shell (which is K shell), and this is not a stable arrangement of electrons. A stable arrangement is to have 2 electrons in the K shell because then the helium gas electron structure will be achieved. Thus, a hydrogen atom needs 1 more electron to become stable. It gets this electron by sharing with another hydrogen atom. So, **two hydrogen atoms share one electron each to form a hydrogen molecule :**

$$\underset{\substack{\text{Hydrogen atom} \\ K \\ 1}}{H\cdot} \quad + \quad \underset{\substack{\text{Hydrogen atom} \\ K \\ 1}}{\cdot H} \quad \longrightarrow \quad \underset{\substack{\text{Hydrogen molecule} \\ \binom{K}{2}\binom{K}{2}}}{\overset{\substack{\text{Shared pair} \\ \text{of electrons} \\ \downarrow}}{H:H}} \quad \text{or} \quad \overset{\substack{\text{Single bond} \\ \downarrow}}{H{-}H}$$

In the hydrogen molecule, each hydrogen atom is supposed to have 2 electrons in its outermost shell, K shell (1 its own and 1 shared). So, each hydrogen atom in the hydrogen molecule has the stable electron arrangement like that of inert gas helium (which has 2 electrons in its outermost K shell). This makes a hydrogen molecule very stable. Please note that **hydrogen is one of the elements which cannot achieve the 8-electron configuration (octet configuration) in its outermost shell during the bond formation.** This is because the outermost shell of a hydrogen atom is the first shell or K shell which can accommodate a maximum of 2 electrons only (and not 8 electrons). Thus, hydrogen atoms can achieve only the helium gas electron configuration of having 2 electrons in the outermost K shell. **Lithium is another element which cannot acquire the 8-electron configuration (octet configuration) during bond formation.**

The formation of a hydrogen molecule from two hydrogen atoms can be shown by the diagram given on the next page.

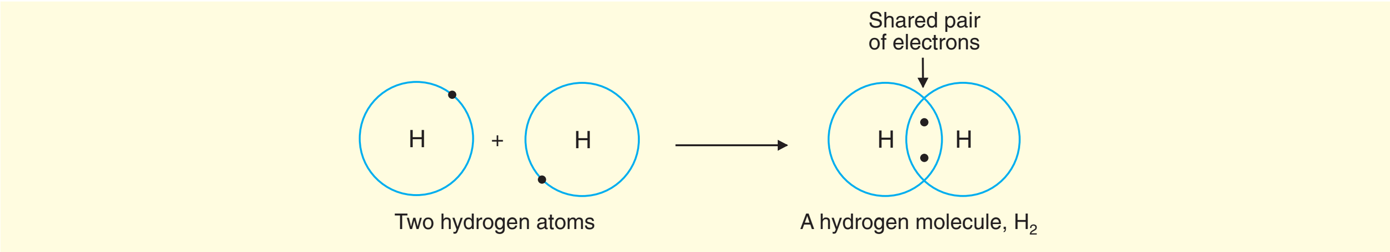


Figure 57. Diagram to show the formation of a hydrogen molecule.

It is clear from the above diagram that when the two reacting hydrogen atoms come close enough, their shells overlap and then their electrons get shared to form a hydrogen molecule. Hydrogen gas is made up of hydrogen molecules and for this reason it is called a molecular substance. Its formula is H_2. Hydrogen gas is called diatomic because it has 2 atoms in each molecule.

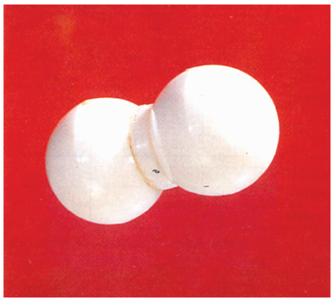


Figure 58. Model of a hydrogen molecule. The two white balls represent two hydrogen atoms.

3. Formation of a Hydrogen Chloride Molecule, HCl

Hydrogen atom has 1 valence electron, so it needs 1 more electron to get 2-electron helium gas electron structure and become stable. Chlorine atom has 7 valence electrons, so it also needs 1 more electron to achieve the 8-electron structure and become stable. Since both hydrogen atom and chlorine atom need 1 electron each, they will become stable by sharing 1 electron with each other. So, **hydrogen atom and chlorine atom share one electron each and form a hydrogen chloride molecule :**

				Shared electron pair		Single bond
H·	+	·Cl:	⟶	H :Cl:	or	H—Cl
Hydrogen atom		Chlorine atom		Hydrogen chloride molecule		
1		2, 8, 7		(2) (2, 8, 8)		

In the hydrogen chloride molecule, the hydrogen atom has 2 electrons in its outermost K shell (1 its own and 1 shared), so it resembles inert gas helium in electron arrangement. The chlorine atom in hydrogen chloride molecule has 8 electrons in its outermost shell (7 its own and 1 shared), and it resembles inert gas argon in electron arrangement (of 2, 8, 8). **Hydrogen chloride gas is a covalent compound containing a covalent bond.**

The combination of a hydrogen atom and a chlorine atom to form a hydrogen chloride molecule can be shown by means of a diagram as follows :

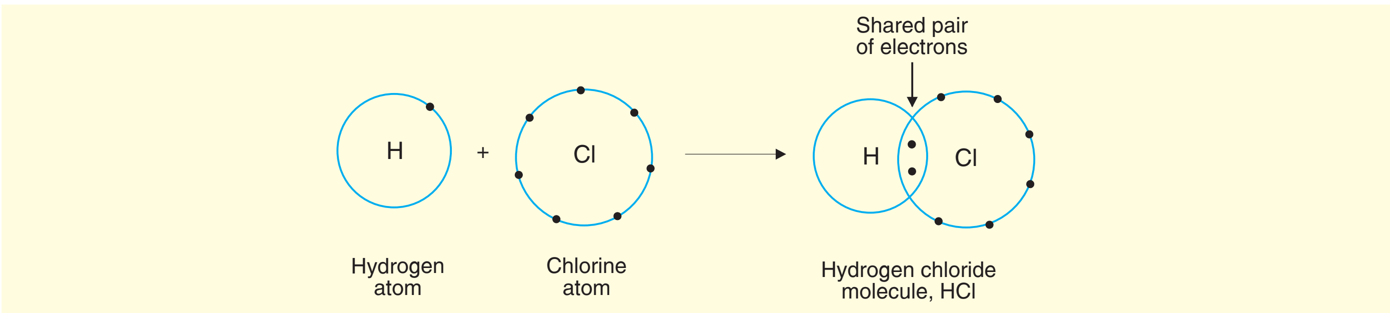
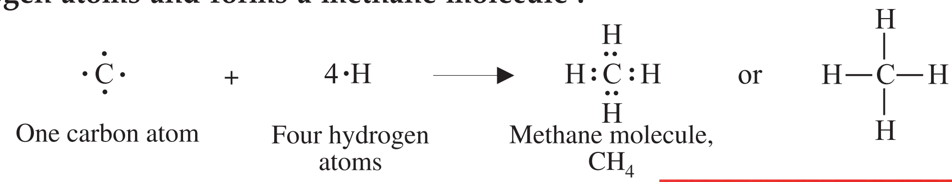


Figure 59. Diagram to show the formation of a hydrogen chloride molecule.

4. Formation of a Methane Molecule, CH_4

Methane is a covalent compound containing covalent bonds. We will now explain the formation of a methane molecule on the basis of electronic theory of valency.

The atomic number of carbon is 6, so its electronic configuration is $\underset{2,\ 4}{\text{K L}}$. Carbon has 4 valence electrons so it needs 4 more electrons to complete the 8-electron structure and become stable. The atomic number of hydrogen is 1, so its electronic configuration is $\underset{1}{\text{K}}$. Hydrogen atom has 1 electron in its K shell and it needs 1 more electron to complete the 2-electron, helium structure. **The carbon atom shares its 4 valence electrons with four hydrogen atoms and forms a methane molecule :**



In the methane molecule both carbon atom as well as the hydrogen atoms have stable inert gas electron arrangements. The carbon atom in methane has 8 electrons in its outermost shell (4 its own and 4 shared), and it resembles inert gas neon in electron arrangement. Each hydrogen atom in methane has 2 electrons in its K shell and resembles helium gas in electron arrangement. There are four 'carbon-hydrogen' single bonds in methane. Each single bond consists of one pair of shared electrons. So, a methane molecule has four pairs of shared electrons.

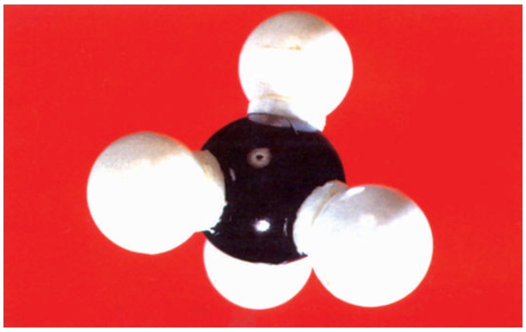
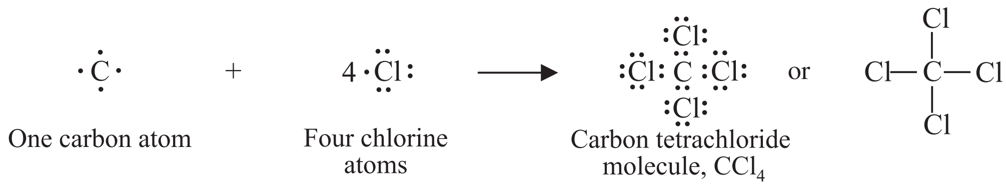


Figure 60. Model of a methane molecule (CH_4). The black ball represents carbon atom whereas white balls represent hydrogen atoms.

5. Formation of a Carbon Tetrachloride Molecule, CCl_4

Carbon tetrachloride, CCl_4, is a covalent compound containing single covalent bonds. The formation of a carbon tetrachloride molecule from carbon and chlorine can be explained as follows.

Carbon atom has 4 valence electrons, so it needs 4 more electrons to complete the 8-electron configuration of inert gas. Chlorine atom has 7 valence electrons, so it needs 1 more electron to achieve the eight-electron structure. **The carbon atom shares its four valence electrons with four chlorine atoms to form carbon tetrachloride molecule :**

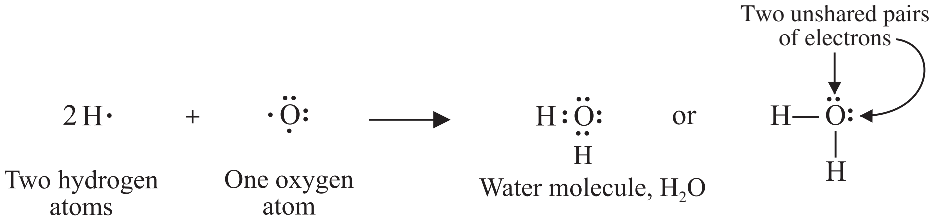


All the atoms in the carbon tetrachloride molecule have a stable 8-electron inert gas configuration in their outermost shells. For example, C atom in CCl_4 has 8 electrons in its valence shell (4 its own and 4 shared). The electronic configuration of C atom in CCl_4 resembles its nearest inert gas neon. Each Cl atom in CCl_4 has also 8 electrons in its outermost shell (7 its own and 1 shared). The electronic configuration of each Cl atom in CCl_4 resembles its nearest inert gas argon. Please note that carbon tetrachloride (CCl_4) is also known as tetrachloromethane.

6. Formation of a Water Molecule, H_2O

Water is a covalent compound consisting of hydrogen and oxygen. It contains single covalent bonds. The formation of a water molecule from hydrogen and oxygen can be explained as follows :

The hydrogen atom has only 1 electron in its outermost K shell, so it needs 1 more electron to achieve the stable, 2-electron arrangement of the inert gas helium. The oxygen atom has 6 electrons in its outermost shell, and it needs 2 more electrons to complete the stable, 8-electron arrangement of inert gas neon. So, **one atom of oxygen shares its two electrons with two hydrogen atoms to form a water molecule :**



In the water molecule, H_2O, each H atom has the electron arrangement of a helium atom whereas the O atom has an electron arrangement of a neon atom. Please note that the central oxygen atom in the water molecule has two pairs of unshared electrons which have not been utilised in the formation of bonds.

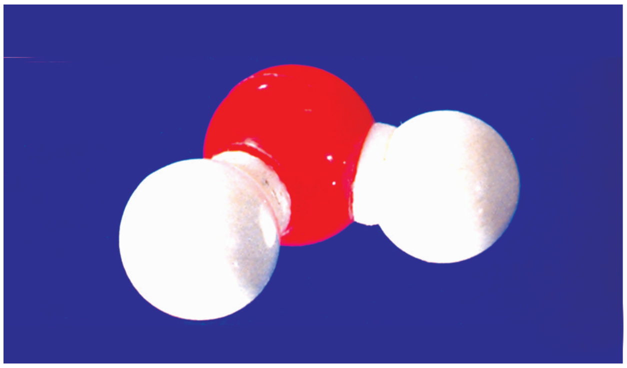


Figure 61. Model of a water molecule. The white balls represent hydrogen atoms whereas the red ball represents oxygen atom.

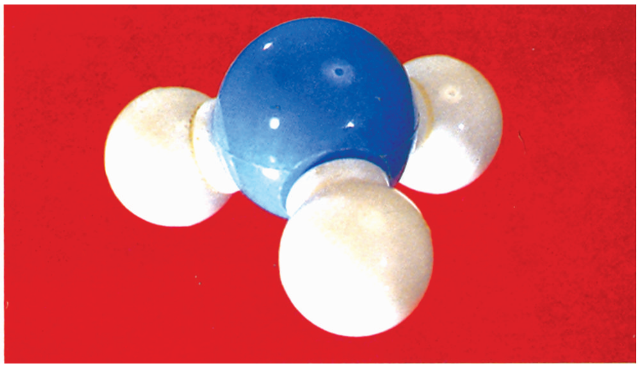
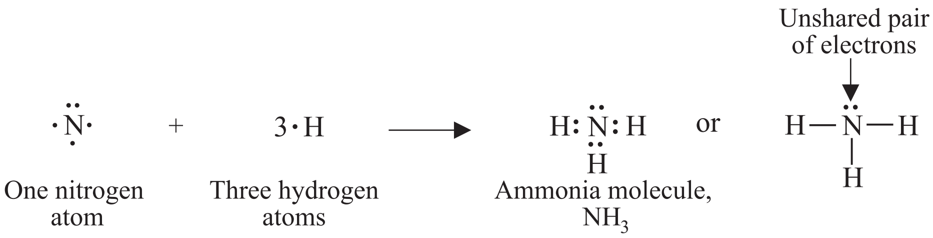


Figure 62. Model of an ammonia molecule. The blue ball represents nitrogen atom whereas white balls represent hydrogen atoms.

7. Formation of Ammonia Molecule, NH_3

Nitrogen combines with hydrogen to form a covalent compound ammonia having covalent bonds in it. The formation of ammonia from nitrogen and hydrogen atoms can be explained as follows.

The nitrogen atom has 5 valence electrons, so it needs 3 more electrons to complete the 8 electrons in the valence shell and become stable. The hydrogen atom has 1 valence electron in the K shell, so it needs 1 more electron to complete 2 electrons in its K valence shell and become stable. The nitrogen and hydrogen atoms get these electrons by sharing with one another. So, **one atom of nitrogen shares its three valence electrons with three hydrogen atoms and forms the ammonia molecule :**



In the ammonia molecule, NH_3, each hydrogen atom attains the electron arrangement of inert gas helium, and the nitrogen atom achieves the electron arrangement of its nearest inert gas neon. Please note that the nitrogen atom in the ammonia molecule has an unshared pair of electrons on it. This pair of electrons has not been utilised in chemical bonding.

DOUBLE BOND

A double covalent bond consists of two pairs of shared electrons. In other words, **a double bond is formed by the sharing of two pairs of electrons between two atoms.** Since two pairs of electrons means 4 electrons, we can also say that **a double covalent bond is formed by the sharing of four electrons between two atoms, each atom contributing two electrons for sharing.** A double bond is actually a combination of two single bonds, so it is represented by putting two short lines (═) between the two atoms. For example, oxygen molecule, O_2, contains a double bond between two atoms and it can be written as O═O. Carbon dioxide and ethene also contain double bonds. We will now take some examples to understand the formation of double bonds.

1. Formation of Oxygen Molecule, O_2

Oxygen atom is very reactive and cannot exist free because it does not have the stable, inert gas electron arrangement in its valence shell. Oxygen gas, therefore, does not consist of single atoms O, it consists of more stable O_2 molecules. The formation of an oxygen molecule from two atoms of oxygen can be explained on the basis of electronic theory of valency as follows :

The atomic number of oxygen is 8, so its electronic configuration is 2, 6. Thus, an oxygen atom has 6 electrons in its outermost shell. Since an oxygen atom has 6 electrons in its outermost shell, therefore, it requires 2 more electrons to achieve the stable, 8-electron inert gas configuration. The oxygen atom gets these electrons by sharing its two electrons with the two electrons of another oxygen atom. So, **two oxygen atoms share two electrons each and form a stable oxygen molecule :**

Two pairs of electrons are shared ↓ Double bond ↓

:Ö: + :Ö: ⟶ :Ö::Ö: or O═O

Two oxygen atoms Oxygen molecule, O_2

Since the oxygen atoms share two pairs of electrons, the bond between them is called a double covalent bond or just a double bond. Thus, **in the oxygen molecule, the two oxygen atoms are held together by a double bond.** Please note that a double bond is stronger than a single bond. In the oxygen molecule, each oxygen atom has 8 electrons in its outermost shell (6 of its own and 2 shared from the other atom), therefore, an oxygen molecule is more stable than the two separate oxygen atoms. Please note that each oxygen atom in oxygen molecule resembles its nearest inert gas neon in electronic configuration.

2. Formation of Carbon Dioxide Molecule, CO_2

Carbon dioxide is a covalent compound made up of carbon and oxygen elements and it contains covalent bonds in it. Knowing that a carbon atom has 4 valence electrons and an oxygen atom has 6 valence electrons, the formation of carbon dioxide molecule can be explained as follows :

Carbon atom has 4 valence electrons, so it needs 4 more electrons to achieve the eight-electron inert gas configuration and become stable. Oxygen atom has 6 valence electrons and it needs 2 more electrons to achieve the eight-electron configuration and become stable. So, **one carbon atom shares its four electrons with two oxygen atoms and forms a carbon dioxide molecule :**

:C: + 2 :Ö: ⟶ :Ö::C::Ö: or O═C═O

One carbon atom Two oxygen atoms Carbon dioxide molecule, CO_2

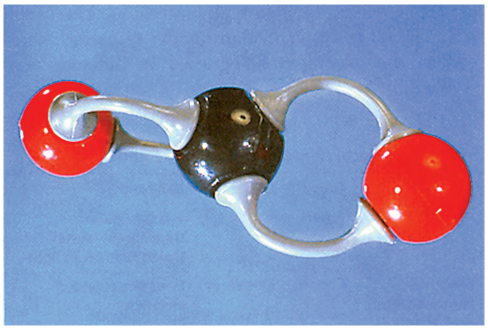


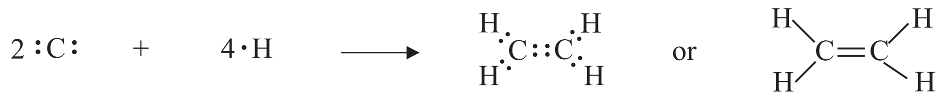
Figure 63. Model of a carbon dioxide molecule. Please note the double bonds between carbon and oxygen atoms.

Please note that **there are two double bonds in a carbon dioxide molecule.** The carbon atom is in the middle of the molecule and the two oxygen atoms are held to it by means of two double bonds, one on each side of the carbon atom. Another point to be noted is that in the carbon dioxide molecule, the carbon atom as well as the two oxygen atoms have attained the electron arrangement of their nearest inert gas neon.

3. Formation of Ethene Molecule, C_2H_4

Ethene is a covalent compound made up of two carbon atoms

and four hydrogen atoms and its formula is C_2H_4. In the formation of ethene molecule, the two carbon atoms share two electrons each to form a double bond among themselves. The remaining four electrons of the two carbon atoms are shared with four hydrogen atoms to form four carbon-hydrogen single bonds. The formation of ethene molecule can be represented as :



Two carbon atoms Four hydrogen atoms Ethene molecule, C_2H_4

It is obvious that in the ethene molecule, the two carbon atoms are joined together by a *double bond* but the hydrogen atoms are joined to the carbon atoms by *single bonds*. Thus, **in ethene molecule we have one carbon-carbon double bond and four carbon-hydrogen single bonds.** So, ethene is a covalent compound which contains single bonds as well as a double bond. Please note that in the ethene molecule, each C atom has achieved an octet of electrons in its valence shell and resembles inert gas neon in its electron arrangement, whereas each H atom has achieved a duplet of electrons in its K valence shell and resembles inert gas helium in its electron arrangement. Please note that the common name of *ethene* is *ethylene*. We will study ethene in Chapter 4 on carbon and its compounds.

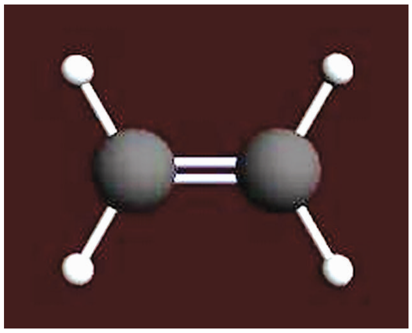


Figure 64. Model of an ethene molecule (C_2H_4). The black balls represent carbon atoms whereas white balls represent hydrogen atoms.

TRIPLE BOND

A triple covalent bond consists of three pairs of shared electrons. In other words, **a triple bond is formed by the sharing of three pairs of electrons between two atoms.** Since three pairs of electrons are equal to six electrons, we can also say that **a triple bond is formed by the sharing of six electrons between two atoms, each atom contributing three electrons for sharing.** A triple bond is actually a combination of three single bonds, so it is represented by putting three short lines ($\equiv$) between the two atoms. Nitrogen molecule, N_2, contains a triple bond, so it can be written as $N \equiv N$. Ethyne molecule also contains a triple bond. We will now explain the formation of a triple bond by taking some examples.

1. Formation of a Nitrogen Molecule, N_2

A nitrogen atom is very reactive and cannot exist free because it does not have the stable electron arrangement of an inert gas. Nitrogen gas, therefore, does not consist of single atoms, it consists of more stable N_2 molecules. The formation of a nitrogen molecule from two nitrogen atoms can be explained as follows :

The atomic number of nitrogen is 7, so its electronic configuration is 2, 5. This means that a nitrogen atom has 5 electrons in its outermost shell. Since a nitrogen atom has 5 electrons in its outermost shell, it needs 3 more electrons to achieve the 8-electron structure of an inert gas and become stable. So, **two nitrogen atoms combine together by sharing 3 electrons each to form a molecule of nitrogen gas :**

Three pairs of electrons are shared ↓ Triple bond ↓

$$:\dot{N}\vdots \;+\; \vdots\dot{N}: \longrightarrow :N\vdots\vdots N: \quad \text{or} \quad :N\equiv N:$$

Two nitrogen atoms Nitrogen molecule, N_2

Since the nitrogen atoms share three pairs of electrons among themselves, the bond between them is called a triple covalent bond or just a triple bond. Thus, **in the nitrogen gas molecule, the two nitrogen atoms are held together by a triple bond**. In the nitrogen molecule, each nitrogen atom has 8 electrons in the outermost shell (5 of its own and 3 shared), so the nitrogen molecule is more stable than two separate nitrogen atoms. Each nitrogen atom in the nitrogen molecule resembles its nearest inert gas neon in electron arrangement. Nitrogen gas, N_2, is diatomic and there is a triple covalent bond between the two atoms of a nitrogen molecule.

2. Formation of Ethyne Molecule, C_2H_2

Ethyne is a covalent compound made up of two carbon atoms and two hydrogen atoms and its formula is C_2H_2. In the formation of an ethyne molecule, the two carbon atoms share three electrons each to form a triple bond among themselves. The remaining two electrons of the two carbon atoms are shared with two hydrogen atoms to form two carbon-hydrogen single bonds. The formation of an ethyne molecule can be represented as follows :

$$\underset{\text{Two carbon atoms}}{2\cdot\dot{\underset{\cdot}{C}}\vdots} \quad + \quad \underset{\text{Two hydrogen atoms}}{2\cdot H} \longrightarrow \underset{\text{Ethyne molecule, } C_2H_2}{H:C\vdots\vdots C:H} \quad \text{or} \quad H-C\equiv C-H$$

It is obvious that in the ethyne molecule, the two carbon atoms are joined together by a *triple bond* but the hydrogen atoms are joined to the carbon atoms by *single bonds.* Thus, in ethyne molecule we have one carbon-carbon triple bond and two carbon-hydrogen single bonds. So, ethyne is a covalent compound which contains single bonds as well as a triple bond. Please note that in the ethyne molecule, each C atom has achieved an octet of electrons in its valence shell, whereas each H atom has achieved a duplet of electrons in its K valence shell, which are very stable arrangements. The common name of *ethyne* is *acetylene.* We will study ethyne in Chapter 4 on carbon and its compounds.

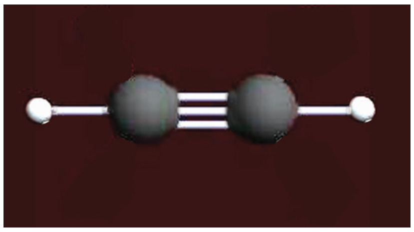


Figure 65. Model of an ethyne molecule (C_2H_2). The black balls represent carbon atoms whereas white balls represent hydrogen atoms.

Covalent Compounds

The compounds containing covalent bonds are known as covalent compounds. Covalent compounds are formed by the sharing of electrons between atoms. The covalent compounds are made up of molecules, so they are also known as molecular compounds. Some of the common covalent compounds and the elements of which they are made, are given below :

Some Covalent Compounds

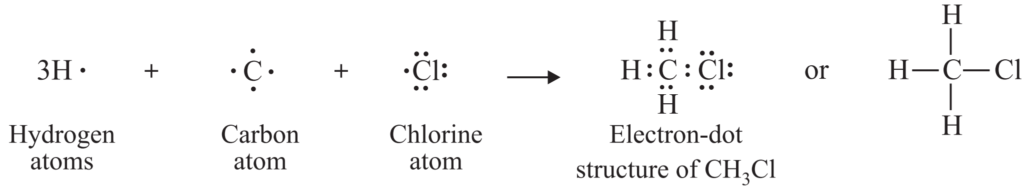
Name	*Formula*	*Elements present*
1. Methane	CH_4	C and H
2. Ethane	C_2H_6	C and H
3. Ethene	C_2H_4	C and H
4. Ethyne	C_2H_2	C and H
5. Water	H_2O	H and O
6. Ammonia	NH_3	N and H
7. Alcohol (Ethanol)	C_2H_5OH	C, H and O
8. Hydrogen chloride gas	HCl	H and Cl
9. Hydrogen sulphide gas	H_2S	H and S
10. Carbon dioxide	CO_2	C and O
11. Carbon disulphide	CS_2	C and S
12. Carbon tetrachloride	CCl_4	C and Cl
13. Glucose	$C_6H_{12}O_6$	C, H and O
14. Cane sugar	$C_{12}H_{22}O_{11}$	C, H and O
15. Urea	$CO(NH_2)_2$	C, O, N and H

Please note that all the above covalent compounds are made up of two (or more) non-metals. So, **whenever we see a compound made up of two (or more) non-metals, we should at once say that it is a covalent compound and contains covalent bonds.** Apart from the above compounds, the elements fluorine, chlorine, bromine, iodine, hydrogen, oxygen and nitrogen also consist of covalent molecules F_2, Cl_2, Br_2, I_2, H_2, O_2, and N_2 respectively.We will now answer some questions based on covalent bonds.

Figure 66. Water (H_2O) is a liquid covalent compound having covalent bonds between hydrogen and oxygen atoms.

Sample Problem 1. Explain the nature of the covalent bond using the bond formation in CH_3Cl. **(NCERT Book Question)**

Solution. CH_3Cl is methyl chloride (or chloromethane). It is made up of one carbon atom, three hydrogen atoms and one chlorine atom. Carbon atom has 4 outermost electrons (or valence electrons), each hydrogen atom has 1 outermost electron, and chlorine atom has 7 valence electrons. Carbon atom shares its 4 valence electrons with three hydrogen atoms and one chlorine atom to form CH_3Cl as shown below :

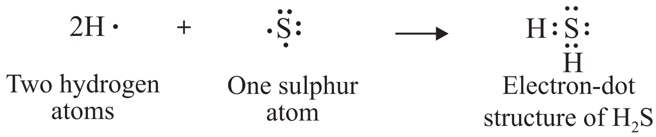


We can see from the above electron-dot structure of CH_3Cl that there are four pairs of shared electrons between carbon and other atoms. Each pair of shared electrons constitutes one single covalent bond. So, CH_3Cl has four single covalent bonds. Please note that each atom in CH_3Cl has a noble gas electron arrangement (of 2 or 8 electrons in the outermost shell).

Sample Problem 2. Draw the electron-dot structures for :

(*a*) H_2S (*b*) F_2 **(NCERT Book Question)**

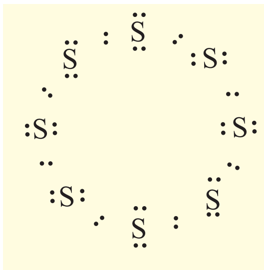
Solution. (*a*) H_2S is hydrogen sulphide. It is made up of two hydrogen atoms and one sulphur atom. Each hydrogen atom has 1 valence electron whereas a sulphur atom has 6 valence electrons. The sulphur atom shares its two electrons with two hydrogen atoms to form hydrogen sulphide as shown below :



(*b*) F_2 is fluorine molecule. Each fluorine atom has 7 valence electrons. Two fluorine atoms share 1 electron each to form a fluorine molecule as shown below :

Sample Problem 3. 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). **(NCERT Book Question)**

Solution. A sulphur atom has 6 outermost electrons. Eight sulphur atoms combine by sharing two electrons among themselves to form a ring type sulphur molecule, S_8 (shown alongside).



Electron-dot structure of sulphur molecule, S_8.

PROPERTIES OF IONIC COMPOUNDS

The important properties of ionic compounds are as follows :

1. Ionic compounds are usually crystalline solids. For example, sodium chloride is a crystalline solid. **The ionic compounds are solids because their oppositely charged ions attract one another strongly and form a regular crystal structure**. The crystals of ionic compounds are hard and brittle.

2. Ionic compounds have high melting points and high boiling points. For example, sodium chloride has a high melting point of 800°C and a high boiling point of 1413°C. *The ionic compounds are made up of positive and negative ions. There is a strong force of attraction between the oppositely charged ions, so a lot of heat energy is required to break this force of attraction and melt or boil the ionic compound. Due to this, ionic compounds have high melting points and high boiling points.* If a substance has high melting point and high boiling point, then we can say that it is an ionic compound and contains ionic bonds.

Figure 67. Copper sulphate is a blue coloured ionic compound. So, it dissolves in water to form a blue coloured copper sulphate solution.

3. Ionic compounds are usually soluble in water but insoluble in organic solvents (like ether, acetone, alcohol, benzene, kerosene, carbon disulphide and carbon tetrachloride). For example, sodium chloride is soluble in water but insoluble in organic solvents like ether, benzene or kerosene oil. Similarly, copper sulphate is an ionic compound which is readily soluble in water (see Figure 67). *The ionic compounds dissolve in water because water has a high dielectric constant due to which it weakens the attraction between the ions.* The organic liquids like ether, benzene or kerosene oil cannot do so.

4. Ionic compounds conduct electricity when dissolved in water or when melted. This means that ionic compounds are electrolytes. **Ionic compounds conduct electricity because they contain charged particles called ions.** Although solid ionic compounds are made up of ions but they do not conduct electric current in the solid state. This is due to the fact that in the solid ionic compound, the ions are held together in fixed positions by strong electrostatic forces and cannot move freely. So, solid ionic compounds are non conductors of electricity. **When we dissolve the ionic solid in water or melt it, the crystal structure is broken down and ions become free to move and conduct electricity. Thus, an aqueous solution of an ionic compound (or a molten ionic compound) conducts electricity because there are plenty of free ions in the solution which are able to conduct electric current.** This point will become more clear from the following example.

Though solid sodium chloride is made up of ions but it does not conduct electricity. This is due to the fact that the sodium ions and chloride ions are held together in fixed positions in the sodium chloride crystal and cannot move freely. When sodium chloride is dissolved in water or melted, it becomes a good conductor of electricity. On dissolving in water or on melting, the sodium chloride crystal is broken up, sodium ions, Na^+, and chloride ions, Cl^-, become free to move and conduct electricity.

We will now describe **experiments to demonstrate some of the properties of ionic compounds**. We will take sodium chloride as the ionic compound in these experiments.

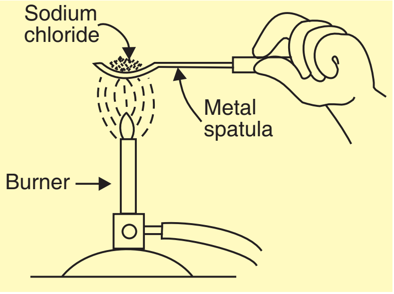


Figure 68. Sodium chloride being heated on a spatula.

(*i*) The property of ionic compounds that they have high melting points can be shown as follows : Take a small amount of sodium chloride on a metal spatula (having an insulated handle). Heat it directly over the flame of a burner (as shown in Figure 68). We will see that sodium chloride does not melt easily. Sodium chloride melts (and becomes a liquid) only on strong heating. This shows that sodium chloride (which is an ionic compound) has a high melting point.

(*ii*) The property of ionic compounds that they are soluble in water but insoluble in organic solvents can be shown as follows : Take some water in a test-tube and add a pinch of sodium chloride to it. Shake the test-tube. We will see that sodium chloride dissolves in water. Thus, sodium chloride (which is an ionic compound) is soluble in water. Let us now take an organic solvent called ether in another test-tube and add a pinch of sodium chloride to it. Shake the test-tube. We will find that sodium chloride does not dissolve in ether. It remains at the bottom of the test-tube as such. Thus, sodium chloride (which is an ionic compound) is insoluble in an organic solvent ether.

(*iii*) The property of ionic compounds that they conduct electricity when dissolved in water can be shown as follows : Fill a beaker half with water and dissolve some sodium chloride in it. Two carbon rods or electrodes (made of graphite) are placed in the sodium chloride solution in the beaker. An electric circuit is then set up by including a battery, a bulb and a switch (see Figure 69). Let us now press the switch. On pressing the switch, the bulb lights up at once. This means that the sodium chloride solution taken in the beaker allows the electric current to pass through it. In other words, the sodium chloride solution conducts electricity. Since sodium chloride is an ionic compound, in general we can say that ionic compounds conduct electricity when dissolved in water.

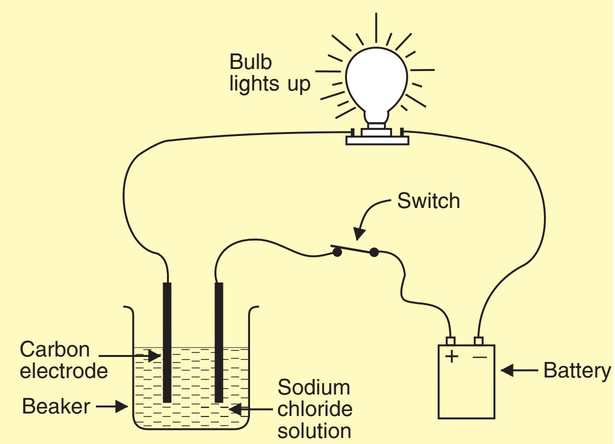


Figure 69. Sodium chloride solution conducts electricity. It is an ionic compound.

PROPERTIES OF COVALENT COMPOUNDS

The important properties of covalent compounds are as follows :

1. Covalent compounds are usually liquids or gases. Only some of them are solids. For example, alcohol, ether, benzene, carbon disulphide, carbon tetrachloride and bromine are liquids; methane, ethane, ethene, ethyne, and chlorine are gases. Glucose, cane sugar, urea, naphthalene and iodine are, however, solid covalent compounds. The covalent compounds are usually liquids or gases due to the weak force of attraction between their molecules.

2. Covalent compounds have usually low melting points and low boiling points. For example, naphthalene has a low melting point of 80°C and carbon tetrachloride has a low boiling point of 77°C. *Covalent compounds are made up of electrically neutral molecules. So, the force of attraction between the molecules of a covalent compound is very weak. Only a small amount of heat energy is required to break these weak molecular forces, due to which covalent compounds have low melting points and low boiling points.* Please note that some of the covalent solids like diamond and graphite have, however, very high melting points and boiling points.

Figure 70. These are naphthalene balls. Naphthalene is a covalent compound. It has a low melting point of 80°C. Naphthalene is also insoluble in water.

3. Covalent compounds are usually insoluble in water but they are soluble in organic solvents. For example, naphthalene is insoluble in water but dissolves in organic solvents like ether. Some of the covalent compounds like glucose, sugar and urea, etc., are, however, soluble in water. The polar covalent compounds like hydrogen chloride and ammonia are also soluble in water.

4. Covalent compounds do not conduct electricity. This means that covalent compounds are non-electrolytes. **Covalent compounds do not conduct electricity because they do not contain ions.** For example, covalent compounds like glucose, cane sugar, urea, alcohol and carbon tetrachloride, etc., do not conduct electricity (because they do not contain ions). Some polar covalent compounds like hydrogen chloride gas, however, conduct electricity when dissolved in water. This is due to the fact that hydrogen chloride chemically reacts with water to form hydrochloric acid containing ions.

We will now describe **experiments to demonstrate some of the properties of covalent compounds**. We will use naphthalene and sugar as the covalent compounds in these experiments.

(*i*) Take a small amount of naphthalene on a metal spatula. Heat it directly over the flame of a burner. We will see that naphthalene melts easily and turns into a liquid. This means that naphthalene (which is a covalent compound) has a low melting point.

(*ii*) Take some water in a test-tube and add a little of naphthalene to it. Shake the test-tube. We will see that naphthalene does not dissolve in water. Thus, naphthalene (which is a covalent compound) is insoluble in water. Let us now take an organic solvent ether in another test-tube and add some naphthalene to it. Shake the test-tube. We will see that naphthalene dissolves in ether. Thus, naphthalene (which is a covalent compound) is soluble in an organic solvent ether.

(*iii*) Set up the apparatus as shown in Figure 69 on page 63 but take sugar solution in the beaker (in place of sodium chloride solution). On pressing the switch, the bulb does not light up. This shows that sugar solution does not conduct electricity. Since sugar is a covalent compound, in general we can say that covalent compounds do not conduct electricity when dissolved in water. (Please note that we have not taken naphthalene as the covalent compound in this case because it does not dissolve in water).

How to Distinguish between Ionic Compounds and Covalent Compounds

The ionic compounds can be distinguished from covalent compounds by making use of the differences in their melting points, boiling points, solubility in water, and solubility in organic solvents. For example :

1. (*a*) If a compound has *high* melting point and boiling point, then it will be an *ionic* compound.
   (*b*) If a compound has comparatively *low* melting point and boiling point, then it will be a *covalent* compound.
2. (*a*) If a compound is *soluble* in water but *insoluble* in organic solvents, it will be an ionic compound.
   (*b*) If a compound is *insoluble* in water but *soluble* in organic solvents, it will be a covalent compound. (Some of the covalent compounds are, however, soluble in water).

The best test to distinguish between ionic compounds and covalent compounds is the electrical conductivity test. Because :

(*i*) If a compound ***conducts electricity*** (in the solution form or molten state), it will be an ***ionic*** compound.

(*ii*) If a compound ***does not conduct electricity*** (in the solution form or molten state or liquid form), then it will be a ***covalent*** compound.

Before we end this discussion, we would like to give the major points of difference between ionic compounds and covalent compounds in the tabular form.

Differences between Ionic Compounds and Covalent Compounds

Ionic compounds	*Covalent compounds*
1. Ionic compounds are usually crystalline solids.	1. Covalent compounds are usually liquids or gases. Only some of them are solids.
2. Ionic compounds have high melting points and boiling points. That is, ionic compounds are non-volatile.	2. Covalent compounds have usually low melting points and boiling points. That is, covalent compounds are usually volatile.
3. Ionic compounds conduct electricity when dissolved in water or melted.	3. Covalent compounds do not conduct electricity.
4. Ionic compounds are usually soluble in water.	4. Covalent compounds are usually insoluble in water (except, glucose, sugar, urea, etc.).
5. Ionic compounds are insoluble in organic solvents (like alcohol, ether, acetone, etc.).	5. Covalent compounds are soluble in organic solvents.

Let us solve some problems now.

Sample Problem 1. In the formation of the compound AB, atoms of A lost one electron each while atoms of B gained one electron each. What is the nature of bond in AB ? Predict the two properties of AB.

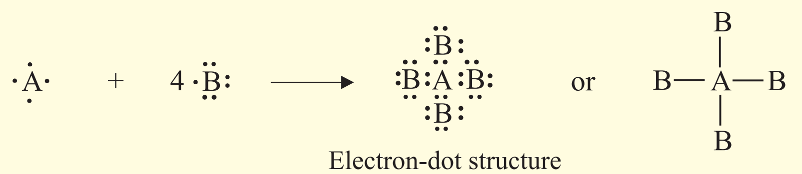
Solution. Here, the atoms of A lose electrons whereas the atoms of B gain electrons. This means that there is a transfer of electrons from atoms of A to atoms of B. Now, the bond formed by the transfer of electrons is called ionic bond. So, the nature of bond in the compound AB is ionic. The two properties of the ionic compound AB will be : (*i*) It will be soluble in water, and (*ii*) it will conduct electricity when dissolved in water or melted.

We know that in the formation of sodium chloride, NaCl, atoms of Na lose one electron each while atoms of Cl gain one electron each. So, the above problem is similar to the formation of sodium chloride.

Sample Problem 2. An element 'A' has 4 electrons in the outermost shell of its atom and combines with another element 'B' having 7 electrons in the outermost shell of its atom. The compound formed does not conduct electricity. What is the nature of the chemical bond in the compound ? Give the electron-dot structure of its molecule.

Solution. The atom of A has 4 valence electrons so it needs 4 more electrons to achieve the stable, 8-electron configuration in the outermost shell. The atom of B has 7 valence electrons, so it needs 1 more electron to complete the 8-electron structure. Since both the reacting atoms need electrons to achieve the inert gas electron arrangements, they will combine by the sharing of electrons and form covalent bonds. Thus, the nature of chemical bond present in the compound is "covalent bond". The presence of covalent bonds in the compound is confirmed by the fact that the compound does not conduct electricity (only ionic compounds containing ionic bonds conduct electricity). We will now give the electron-dot structure of a molecule of the compound formed.

We have been given that the atom A has 4 valence electrons whereas atom B has 7 valence electrons in it. Now, one atom of A shares its four electrons with four atoms of B to form the covalent molecule AB_4 as shown below :



Electron-dot structure of AB_4 molecule

We know that a carbon atom has 4 electrons in its outermost shell, so the element A in the above problem may be carbon. A chlorine atom has 7 electrons in its outermost shell, so the element B may be chlorine. Thus, the compound AB_4 may be carbon tetrachloride, CCl_4. From this discussion we conclude that the above problem is similar to the formation of carbon tetrachloride from carbon and chlorine.

Sample Problem 3. Give the formulae of the chlorides of the elements A and B having atomic numbers of 6 and 11 respectively. Will the properties of the two chlorides be similar or different ? Explain.

Solution. (*i*) The atomic number of element A is 6, so its electronic configuration is 2, 4. Now, an atom of A has 4 electrons in its outermost shell and requires 4 more electrons to achieve the 8-electron configuration and become stable. Thus, the valency of element A will be 4. We know that the valency of chlorine is 1. So, one atom of element A will share its four electrons with 4 atoms of Cl to form a covalent compound having the formula ACl_4.

(*ii*) The atomic number of element B is 11, so its electronic configuration is 2, 8, 1. Now, an atom of B has only 1 electron in its outermost shell which it can give (to a Cl atom), and form a cation B^+. And by gaining 1 electron, the chlorine atom (Cl) forms an anion Cl^-. Now, the ions B^+ and Cl^- combine to give an ionic compound having the formula BCl.

From the above discussion we conclude that the chloride of element A is a covalent compound ACl_4 whereas the chloride of element B is an ionic compound BCl. So, the properties of the two chlorides will be different. *(Please note that the chloride ACl_4 is actually carbon tetrachloride, CCl_4, whereas the chloride BCl is actually sodium chloride, NaCl).*

Before we go further and describe the extraction of metals, **please answer the following questions :**

Very Short Answer Type Questions

1. What is the name of the chemical bond formed :
(*a*) by the sharing of electrons between two atoms ?
(*b*) by the transfer of electrons from one atom to another ?

2. Name a carbon containing molecule which has two double bonds.

3. What would be the electron-dot structure of carbon dioxide which has the formula CO_2 ?

4. What type of chemical bond is formed between :
(*a*) potassium and bromine ?
(*b*) carbon and bromine ?

5. (*a*) What do we call those particles which have more or less electrons than the normal atoms ?
(*b*) What do we call those particles which have more electrons than the normal atoms ?
(*c*) What do we call those particles which have less electrons than the normal atoms ?

6. (*a*) The atomic number of sodium is 11. What is the number of electrons in Na^+ ?
(*b*) The atomic number of chlorine is 17. What is the number of electrons in Cl^- ?

7. The atomic number of an element X is 8 and that of element Y is 12. Write down the symbols of the ions you would expect to be formed from their atoms.

8. (*a*) Write down the electronic configuration of (*i*) magnesium atom, and (*ii*) magnesium ion.
(At. No. of Mg = 12)
(*b*) Write down the electronic configuration of (*i*) sulphur atom, and (*ii*) sulphide ion.
(At. No. of S = 16)

9. What type of chemical bonds are present in a solid compound which has a high melting point, does not conduct electricity in the solid state but becomes a good conductor in the molten state ?

10. State whether the following statement is true or false :
The aqueous solution of an ionic compound conducts electricity because there are plenty of free electrons in the solution.

11. What type of bonds are present in hydrogen chloride and oxygen ?

12. Write the electron-dot structures for the following molecules :
(*i*) NaCl (*ii*) Cl_2

13. What type of bonds are present in water molecule ? Draw the electron-dot structure of water (H_2O).

14. What type of bonds are present in methane (CH_4) and sodium chloride (NaCl) ?

15. State one major difference between covalent and ionic bonds and give one example each of covalent and ionic compounds.

16. What type of bonds are present in the following molecules ? Draw their electron-dot structures.
(*i*) H_2 (*ii*) CH_4 (*iii*) Cl_2 (*iv*) O_2

17. Which inert gas electron configuration do the Cl atoms in Cl_2 molecule resemble ? What is this electron configuration ?

18. Which of the following compounds are ionic and which are covalent ?
Urea, Cane sugar, Hydrogen chloride, Sodium chloride, Ammonium chloride, Carbon tetrachloride, Ammonia, Alcohol, Magnesium chloride.

19. Give one example each of the following :
(*i*) A molecule containing a single covalent bond
(*ii*) A molecule containing a double covalent bond
(*iii*) A molecule containing a triple covalent bond
(*iv*) A compound containing an ionic bond

20. Fill in the blanks in the following sentences :

(*i*) Two atoms of the same element combine to form a molecule. The bond between them is known as .... bond.

(*ii*) Two chlorine atoms combine to form a molecule. The bond between them is known as ....

(*iii*) In forming oxygen molecule, ... electrons are shared by each atom of oxygen.

(*iv*) In forming N_2 molecule, ... electrons are shared by each atom of nitrogen.

(*v*) The number of single covalent bonds in C_2H_2 molecule are ....

(*vi*) Melting points and boiling points of ionic compounds are generally .... than those of covalent compounds.

Short Answer Type Questions

21. (*a*) What is a covalent bond ? What type of bond exists in (*i*) CCl_4, and (*ii*) $CaCl_2$?

(*b*) What is an ionic bond ? What type of bond is present in oxygen molecule ?

22. (*a*) What is an ion ? Explain with examples.

(*b*) What is the nature of charge on (*i*) a cation, and (*ii*) an anion ?

(*c*) Name the cation and anion present in $MgCl_2$. Also write their symbols.

23. (*a*) What type of chemical bond is present in chlorine molecule ? Explain your answer.

(*b*) Explain the formation of a chlorine molecule on the basis of electronic theory of valency.

24. (*a*) Giving one example each, state what are (*i*) ionic compounds, and (*ii*) covalent compounds.

(*b*) Compare the properties of ionic compounds and covalent compounds.

25. Explain why :

(*a*) covalent compounds have generally low melting points.

(*b*) ionic compounds have generally high melting points.

26. (*a*) Give two general properties of ionic compounds and two those of covalent compounds.

(*b*) State one test by which sodium chloride can be distinguished from sugar.

27. (*a*) Explain why, ionic compounds conduct electricity in solution whereas covalent compounds do not conduct electricity .

(*b*). Which of the following will conduct electricity and which not ?

$MgCl_2$, CCl_4, $NaCl$, CS_2, Na_2S

Give reasons for your choice.

28. (*a*) Name one ionic compound containing chlorine and one covalent compound containing chlorine.

(*b*) How will you find out which of the water soluble compound A or B is ionic ?

29. Explain why, a solution of cane sugar does not conduct electricity but a solution of common salt is a good conductor of electricity.

30. Give the formulae of the compounds that would be formed by the combination of the following pairs of elements :

(*a*) Mg and N_2 (*b*) Li and O_2

(*c*) Al and Cl_2 (*d*) K and H

31. (*a*) What are noble gases ? What is the characteristic of the electronic configuration of noble gases ?

(*b*) What is the cause of chemical bonding (or chemical combination) of atoms of elements ?

32. (*i*) Write electron-dot structures for magnesium and oxygen.

(*ii*) Show the formation of MgO by the transfer of electrons.

(*iii*) What are the ions present in this compound ?

33. Draw the electron-dot structure of a hydrogen chloride molecule :

(*i*) Which inert gas does the H atom in HCl resemble in electron arrangement ?

(*ii*) Which inert gas does the Cl atom in HCl resemble in electron arrangement ?

34. What type of bonding would you expect between the following pairs of elements ?

(*i*) Calcium and Oxygen

(*ii*) Carbon and Chlorine

(*iii*) Hydrogen and Chlorine

35. Describe how sodium and chlorine atoms are changed into ions when they react with each other to form sodium chloride, NaCl. What is the name given to this type of bonding ? (At. No of sodium = 11 ; At. No. of chlorine = 17)

36. What is the difference between a cation and an anion ? How are they formed ? Give the names and symbols of one cation and one anion.

37. Using electron-dot diagrams which show only the outermost shell electrons, show how a molecule of nitrogen, N_2, is formed from two nitrogen atoms. What name is given to this type of bonding ? (Atomic number of nitrogen is 7)

38. Draw the electron-dot structures of the following compounds and state the type of bonding in each case :
(*i*) CO_2 (*ii*) MgO (*iii*) H_2O (*iv*) HCl (*v*) $MgCl_2$

39. Using electron-dot diagrams which show only the outermost shell electrons, show how a molecule of oxygen, O_2, is formed from two oxygen atoms. What name is given to this type of bonding ? (At. No. of oxygen = 8)

40. Draw the electron-dot structures of the following compounds and state the type of bonding in each case :
(*i*) KCl (*ii*) NH_3 (*iii*) CaO (*iv*) N_2 (*v*) $CaCl_2$

41. Explain why, a salt which does not conduct electricity in the solid state becomes a good conductor in molten state.

Long Answer Type Questions

42. (*a*) Write down the electronic configuration of (*i*) sodium atom, and (*ii*) chlorine atom.
(*b*) How many electrons are there in the outermost shell of (*i*) a sodium atom, and (*ii*) a chlorine atom ?
(*c*) Show the formation of NaCl from sodium and chlorine atoms by the transfer of electron(s).
(*d*) Why has sodium chloride a high melting point ?
(*e*) Name the anode and the cathode used in the electrolytic refining of impure copper metal.

43. (*a*) Write the electron arrangement in (*i*) a magnesium atom, and (*ii*) an oxygen atom.
(*b*) How many electrons are there in the valence shell of (*i*) a magnesium atom, and (*ii*) an oxygen atom ?
(*c*) Show on a diagram the transfer of electrons between the atoms in the formation of MgO.
(*d*) Name the solvent in which ionic compounds are generally soluble.
(*e*) Why are aqueous solutions of ionic compounds able to conduct electricity ?

44. (*a*) What is the electronic configuration of (*i*) a sodium atom, and (*ii*) an oxygen atom ?
(*b*) What is the number of outermost electrons in (*i*) a sodium atom, and (*ii*) an oxygen atom ?
(*c*) Show the formation of Na_2O by the transfer of electrons between the combining atoms.
(*d*) Why are ionic compounds usually hard ?
(*e*) How is it that ionic compounds in the solid state do not conduct electricity but they do so when in molten state ?

45. (*a*) Write down the electron arrangement in (*i*) a magnesium atom, and (*ii*) a chlorine atom.
(*b*) How many electrons are there in the valence shell of (*i*) a magnesium atom, and (*ii*) a chlorine atom ?
(*c*) Show the formation of magnesium chloride from magnesium and chlorine by the transfer of electrons.
(*d*) State whether magnesium chloride will conduct electricity or not. Give reason for your answer.
(*e*) Why are covalent compounds generally poor conductors of electricity ?

Multiple Choice Questions (MCQs)

46. The atomic number of an element X is 19. The number of electrons in its ion X^+ will be :
(*a*) 18 (*b*) 19 (*c*) 20 (*d*) 21

47. The atomic number of an element Y is 17. The number of electrons in its ion Y^- will be :
(*a*) 17 (*b*) 18 (*c*) 19 (*d*) 20

48. The atomic numbers of four elements A, B, C and D are 6, 8, 10 and 12 respectively. The two elements which can react to form ionic bonds (or ionic compound) are :
(*a*) A and D (*b*) B and C (*c*) A and C (*d*) B and D

49. The atomic numbers of four elements P, Q, R and S are 6, 10, 12 and 17 respectively. Which two elements can combine to form a covalent compound ?
(*a*) P and R (*b*) Q and S (*c*) P and S (*d*) R and S

50. The solution of one of the following compounds will not conduct electricity. This compound is :
(*a*) NaCl (*b*) CCl_4 (*c*) $MgCl_2$ (*d*) $CaCl_2$

51. The electronic configurations of three elements X, Y and Z are :
X : 2 Y : 2, 8, 7 Z : 2, 8, 2
Which of the following is correct regarding these elements ?
(*a*) X is a metal (*b*) Y is a metal
(*c*) Z is a non-metal (*d*) Y is a non-metal and Z is a metal

52. Which one of the following property is generally not exhibited by ionic compounds ?
(*a*) solubility in water (*b*) electrical conductivity in solid state
(*c*) high melting and boiling points (*d*) electrical conductivity in molten state

53. The electrons present in the valence shell of a noble gas atom can be :
(*a*) 8 only (*b*) 2 only (*c*) 8 or 2 (*d*) 8 or 4

54. The atomic number of an element X is 16. The symbol of ion formed by an atom of this element will be :
(*a*) X^{2+} (*b*) X^{3+} (*c*) X^{2-} (*d*) X^{-}

55. The number of protons in the nucleus of one atom of an element Y is 5. The symbol of ion formed by an atom of this element will be :
(*a*) Y^{3-} (*b*) Y^{2+} (*c*) Y^{2-} (*d*) Y^{3+}

56. Out of KCl, HCl, CCl_4 and NaCl, the compounds which are not ionic are :
(*a*) KCl and HCl (*b*) HCl and CCl_4
(*c*) CCl_4 and NaCl (*d*) KCl and CCl_4

57. Element X reacts with element Y to form a compound Z. During the formation of compound Z, atoms of X lose one electron each whereas atoms of Y gain one electron each. Which of the following property is not shown by compound Z ?
(*a*) high melting point (*b*) low melting point
(*c*) occurrence as solid (*d*) conduction of electricity in molten state

58. One of the following compounds is not ionic in nature. This compound is :
(*a*) Lithium chloride (*b*) Ammonium chloride
(*c*) Calcium chloride (*d*) Carbon tetrachloride

59. The rechargeable battery used in a mobile phone hand set is usually :
(*a*) lead ion battery (*b*) sodium ion battery
(*c*) hydrogen ion battery (*d*) lithium ion battery

60. The number of protons in one atom of an element X is 8. What will be the number of electrons in its ion X^{2-} ?
(*a*) 8 (*b*) 9 (*c*) 10 (*d*) 11

61. If the number of protons in one atom of an element Y is 20, then the number of electrons in its ion Y^{2+} will be :
(*a*) 20 (*b*) 19 (*c*) 18 (*d*) 16

62. The noble gas having only two electrons in its valence shell is :
(*a*) Ar (*b*) Ne (*c*) He (*d*) Kr

63. A covalent molecule having a double bond between its atoms is :
(*a*) Hydrogen (*b*) Oxygen (*c*) water (*d*) ammonia

64. The molecules having triple bond in them are :
(*a*) oxygen and ethyne (*b*) carbon dioxide and ammonia
(*c*) methane and ethene (*d*) nitrogen and ethyne

65. One of the following contains a double bond as well as single bonds. This is :
(*a*) CO_2 (*b*) O_2 (*c*) C_2H_4 (*d*) C_2H_2

66. Which of the following has a triple bond as well as single bonds ?
(*a*) ethene (*b*) methane (*c*) ethyne (*d*) nitrogen

Questions Based on High Order Thinking Skills (HOTS)

67. Two non-metals combine with each other by the sharing of electrons to form a compound X.

(*a*) What type of chemical bond is present in X ?

(*b*) State whether X will have a high melting point or low melting point.

(*c*) Will it be a good conductor of electricity or not ?

(*d*) Will it dissolve in an organic solvent or not ?

68. A metal combines with a non-metal by the transfer of electrons to form a compound Y.

(*i*) State the type of bonds in Y.

(*ii*) What can you say about its melting point and boiling point ?

(*iii*) Will it be a good conductor of electricity ?

(*iv*) Will it dissolve in an organic solvent or not ?

69. The electronic configurations of three elements X, Y and Z are as follows :

X	2, 4
Y	2, 7
Z	2, 1

(*a*) Which two elements will combine to form an ionic compound ?

(*b*) Which two elements will react to form a covalent compound ?

Give reasons for your choice.

70. An element A has 4 valence electrons in its atom whereas element B has only one valence electron in its atom. The compound formed by A and B does not conduct electricity. What is the nature of chemical bond in the compound formed ? Give its electron-dot structure.

71. In the formation of a compound XY_2 atom X gives one electron to each Y atom. What is the nature of bond in XY_2 ? Give two properties of XY_2.

72. An element 'A' has two electrons in the outermost shell of its atom and combines with an element 'B' having seven electrons in the outermost shell, forming the compound AB_2. The compound when dissolved in water conducts electric current. Giving reasons, state the nature of chemical bond in the compound.

73. The electronic configurations of two elements A and B are given below :

A	2, 6
B	2, 8, 1

(*a*) What type of chemical bond is formed between the two atoms of A ?

(*b*) What type of chemical bond will be formed between the atoms of A and B ?

74. Four elements A, B, C and D have the following electron arrangements in their atoms :

A	2, 8, 6
B	2, 8, 8
C	2, 8, 8, 1
D	2, 7

(*a*) What type of bond is formed when element C combines with element D ?

(*b*) Which element is an inert gas ?

(*c*) What will be the formula of the compound between A and C ?

75. An element X of atomic number 12 combines with an element Y of atomic number 17 to form a compound XY_2. State the nature of chemical bond in XY_2 and show how the electron configurations of X and Y change in the formation of this compound.

76. The electronic configurations of three elements A, B and C are as follows :

A	2, 8, 1
B	2, 8, 7
C	2, 4

(*a*) Which of these elements is a metal ?

(*b*) Which of these elements are non-metals ?

(*c*) Which two elements will combine to form an ionic bond ?

(*d*) Which two elements will combine to form a covalent bond ?

(*e*) Which element will form an anion of valency 1 ?

77. The electronic configurations of four particles A, B, C and D are given below :

A	2, 8, 8
B	2, 8, 2
C	2, 6
D	2, 8

Which electronic configuration represents :

(*i*) magnesium atom ? (*ii*) oxygen atom ?

(*iii*) sodium ion ? (*iv*) chloride ion ?

78. The atomic number of an element X is 12.

(*a*) What must an atom of X do to attain the nearest inert gas electron configuration ?

(*b*) Which inert gas is nearest to X ?

79. The atomic number of an element Y is 16.

(*a*) What must an atom of Y do to achieve the nearest inert gas electron arrangement ?

(*b*) Which inert gas is nearest to Y ?

80. You can buy solid air-freshners in shops. Do you think these substances are ionic or covalent ? Why ?

81. Give the formulae of the chlorides of the elements X and Y having atomic numbers of 3 and 6 respectively. Will the properties of the two chlorides be similar or different ? Explain your answer.

ANSWERS

1. (*a*) Covalent bond (*b*) Ionic bond **2.** Carbon dioxide, CO_2 **4.** (*a*) Ionic bond (*b*) Covalent bond **5.** (*a*) Ions (*b*) Anions (*c*) Cations **6.** (*a*) 10 (*b*) 18 **7.** X^{2-} ; Y^{2+} **8.** (*a*) (*i*) 2,8, 2 (*ii*) 2, 8 (*b*) (*i*) 2, 8, 6 (*ii*) 2, 8, 8 **9.** Ionic bonds **10.** False (It should be 'ions' in place of 'electrons') **11.** Covalent bonds **13.** Covalent bonds **14.** Methane : Covalent bonds ; Sodium chloride : Ionic bonds **17.** Argon ; 2, 8, 8 **18.** Ionic compounds : Sodium chloride , Ammonium chloride, Magnesium chloride ; Covalent compounds : Urea, Cane sugar, Hydrogen chloride, Carbon tetrachloride, Ammonia, Alcohol **19.** (*i*) Hydrogen (*ii*) Oxygen (*iii*) Nitrogen (*iv*) Sodium chloride **20.** (*i*) covalent (*ii*) covalent (*iii*) two (*iv*) three (*v*) two (*vi*) higher **21.** (*a*) (*i*) Covalent bonds (*ii*) Ionic bonds (*b*) Covalent bond **22.** (*b*) (*i*) Positive charge (*ii*) Negative charge (*c*) Magnesium ion, Mg^{2+} ; Chloride ions, $2Cl^-$ **23.** (*a*) Covalent bond **26.** (*b*) An aqueous solution of sodium chloride conducts electricity but a sugar solution does not conduct electricity **27.** (*a*) Ionic compounds are made up of electrically charged ions but covalent compounds are made up of electrically neutral molecules (*b*) Conduct electricity : $MgCl_2$, NaCl, Na_2S (Ionic compounds) ; Do not conduct electricity : CCl_4, CS_2 (Covalent compounds) **28.** (*a*) Ionic compound : Sodium chloride, NaCl ; Covalent compound : Carbon tetrachloride, CCl_4 (*b*) Out of A and B, the compound whose aqueous solution conducts electricity will be an ionic compound **30.** (*a*) Mg_3N_2 (*b*) Li_2O (*c*) $AlCl_3$ (*d*) KH **33.** (*i*) Helium (*ii*) Argon **34.** (*i*) Ionic bonding (*ii*) Covalent bonding (*iii*) Covalent bonding **35.** Ionic bonding **37.** Covalent bonding **42.** (*e*) Anode : Thick block of impure copper metal ; Cathode : Thin strip of pure copper metal **43.** (*d*) Water **45.** (*d*) Magnesium chloride will conduct electricity because it is an ionic compound **46.** (*a*) **47.** (*b*) **48.** (*d*) **49.** (*c*) **50.** (*b*) **51.** (*d*) **52.** (*b*) **53.** (*c*) **54.** (*c*) **55.** (*d*) **56.** (*b*) **57.** (*b*) **58.** (*d*) **59.** (*d*) **60.** (*c*) **61.** (*c*) **62.** (*c*) **63.** (*b*) **64.** (*d*) **65.** (*c*) **66.** (*c*) **67.** (*a*) Covalent bond (*b*) Low melting point (*c*) No (*d*) Yes **68.** (*i*) Ionic bond (*ii*) High melting point and boiling point (*iii*) Yes (*iv*) No **69.** (*a*) Y and Z (*b*) X and Y **70.** Covalent bond, $\begin{array}{c} \mathrm{B} \\ \mathrm{B}:\underset{..}{\overset{..}{\mathrm{A}}}:\mathrm{B} \\ \mathrm{B} \end{array}$ **71.** Ionic bond **72.** Ionic bond **73.** (*a*) Covalent bond (*b*) Ionic bond **74.** (*a*) Ionic bond (*b*) B (*c*) C_2A **75.** Ionic bond; The electronic configuration of X changes from 2, 8, 2 to 2, 8 ; The electronic configuration of Y changes from 2, 8, 7 to 2, 8, 8 **76.** (*a*) A (*b*) B and C (*c*) A and B (*d*) B and C (*e*) B **77.** (*i*) B (*ii*) C (*iii*) D (*iv*) A **78.** (*a*) Lose 2 electrons (*b*) Neon **79.** (*a*) Accept 2 electrons (*b*) Argon **80.** Solid air-freshners are covalent compounds because they are volatile **81.** Formula of chloride of element X is XCl ; Formula of chloride of element Y is YCl_4 ; The properties of two chlorides will be different because XCl is an ionic chloride whereas YCl_4 is a covalent chloride.

OCCURRENCE OF METALS

The earth's crust is the major source of metals. Sea-water also contains salts of metals like sodium chloride, magnesium chloride, etc. **Most of the metals are quite reactive and hence they do not occur as free elements in nature.** So, most of the metals are found in the form of their compounds (with other elements) called 'combined state'. The compounds of metals found in nature are their oxides, carbonates, sulphides and chlorides, etc. In these compounds, the metals are present in the form of positive ions (or cations). **Only a few less reactive metals (like copper, silver, gold and platinum) are found in the 'free state' as metals** (because of their low chemical reactivity). When a metal is found as *free element,* it is said to occur in '*native state*'. So, we can also say that copper, silver, gold and platinum metals occur in native state. Please note that **copper and silver metals occur in free state (native state) as well as in the combined state (in the form of compounds)**.

Figure 71. Some metals such as gold are very unreactive due to which they can be found in their metallic form in the Earth's crust. On a small scale, gold is extracted by panning (which involves washing the gold containing gravel in a pan to separate out gold).

We have already studied the reactivity series of metals. We can relate the occurrence of metals to the reactivity series of metals as follows : The metals which are high up in the reactivity series (like potassium, sodium, calcium, magnesium and aluminium) are so reactive that they are never found in nature as free elements. They are always found in combined state. The metals placed in the middle of reactivity series (like zinc, iron and lead) are moderately reactive metals which are also found in the combined state. In fact, **all the metals which are placed above copper in the reactivity series are found in nature only in the form of their compounds.** The metals which are quite low in the reactivity series (such as copper, silver, gold and platinum) are the least reactive or unreactive and hence found in free state as metals. Copper and silver metals are found in free state only to a small extent. They are mainly found in the combined state as their sulphides or oxides.

Minerals and Ores

The natural materials in which the metals or their compounds are found in earth are called minerals. Some minerals may contain a large percentage of metal whereas others may contain only a small percentage of the metal. Some minerals may not contain any objectionable impurities whereas others may contain objectionable impurities which hamper the extraction of metals. Thus, all the minerals cannot be used to extract metals. **Those minerals from which the metals can be extracted conveniently and profitably are called ores.** An ore contains a good percentage of metal and there are no objectionable impurities in it. Thus, **all the ores are minerals, but all the minerals are not ores.** Some of the common ores are given in the following table :

Figure 72. This is an iron ore called haematite. It contains iron (III) oxide, Fe_2O_3.

Metal (to be extracted)	*Name of ore*	*Name of compound in ore*	*Formula of ore*
1. Sodium	Rock salt	Sodium chloride	NaCl
2. Aluminium	Bauxite	Aluminium oxide	$Al_2O_3.2H_2O$
3. Manganese	Pyrolusite	Manganese dioxide	MnO_2
4. Zinc	(*i*) Calamine	Zinc carbonate	$ZnCO_3$
	(*ii*) Zinc blende	Zinc sulphide	ZnS

5. Iron	Haematite	Iron (III) oxide	Fe_2O_3
6. Copper	(*i*) Cuprite	Copper (I) oxide	Cu_2O
	(*ii*) Copper glance	Copper (I) sulphide	Cu_2S
7. Mercury	Cinnabar	Mercury (II) sulphide	HgS

We can see from the above table that the ores of many metals are oxides. The ores of many metals are oxides because oxygen is a very reactive element and very abundant on the earth.

EXTRACTION OF METALS

An ore contains a metal in the form of its compound with other elements. So, after the mining of the ore from the ground, it must be converted into pure metal. **To obtain a metal from its ore is called the extraction of metal.** The ores are converted into free metals by a number of steps which depend on the type of the ore used, nature of the impurities present and reactivity of the metal to be extracted. **The various processes involved in the extraction of metals from their ores, and refining are known as metallurgy.** Please note that no single process can be used for the extraction of all the metals. The process to be used varies from metal to metal. The three major steps involved in the extraction of a metal from its ore are :

(*i*) Concentration of ore (or Enrichment of ore),

(*ii*) Conversion of concentrated ore into metal, and

(*iii*) Refining (purification) of impure metal.

We will now describe all these steps in detail, one by one. Let us start with the concentration of ore.

1. Concentration of Ore (or Enrichment of Ore)

Ore is an impure compound of a metal containing a large amount of sand and rocky material. **The unwanted impurities like sand, rocky material, earthy particles, limestone, mica, etc., present in an ore are called gangue.** Before extracting the metal from an ore, it is necessary to remove these impurities (or gangue). *The methods used for removing gangue from ore depend on some difference in the physical properties or chemical properties of the ore and gangue.* By removing the gangue, we get a concentrated ore containing a much higher percentage of the metal. We will discuss the various methods of ore concentration in higher classes. Please note that the concentration of ore is also known as enrichment of ore.

2. Conversion of Concentrated Ore into Metal

For the purpose of extracting metals from the concentrated ores, we can group all the metals into following three categories :

(*i*) Metals of high reactivity (or Highly reactive metals)

(*ii*) Metals of medium reactivity (or Moderately reactive metals)

(*iii*) Metals of low reactivity (or Less reactive metals)

Different methods are used for extracting metals belonging to the above three categories. This is shown in Figure 73. Manganese metal (Mn) lies just above zinc (Zn) in the reactivity series (but it has not been shown in Figure 73). Manganese metal is obtained by the reduction of its oxide with aluminium powder and not carbon. This is because carbon is less reactive than manganese. Please note that carbon (C), which is a non-metal, is more reactive than zinc and it can be placed just above Zn in the reactivity series. So, **carbon can reduce the oxides of zinc and of all other metals below zinc to form metals.** Another point to be noted is that tin metal (Sn) is more reactive than lead (Pb), so its place is just above Pb in the reactivity series. A yet another point to be noted is that copper can be extracted by the reduction of its oxide with carbon as well as by heating its sulphide ore in air.

Metal	*Method of extraction*
K Na Ca Mg Al	Electrolysis of molten chloride or oxide
Zn Fe Pb Cu	Reduction of oxide with carbon
Cu Hg	Heating sulphide in air (Reduction by heat alone)
Ag Au Pt	Found in native state (as metals)

Figure 73. The method of extraction of a metal from its concentrated ore depends on its chemical reactivity.

The extraction of a metal from its concentrated ore is essentially a process of reduction of the metal compound present in the ore. The method of reduction to be used depends on the reactivity of the metal to be extracted. This will become clear from the following discussion.

Extraction of Highly Reactive Metals

The highly reactive metals such as potassium, sodium, calcium, magnesium and aluminium are placed high up in the reactivity series in its upper part. So, the extraction of highly reactive metals means the extraction of metals which are towards the top of the reactivity series. The oxides of highly reactive metals (like potassium, sodium, calcium, magnesium and aluminium) are very stable and cannot be reduced by the most common reducing agent 'carbon' to obtain free metals. This is because these metals have more affinity (more attraction) for oxygen than carbon. So, carbon is unable to remove oxygen from these metal oxides and hence cannot convert them into free metals. Thus, ***the highly reactive metals cannot be extracted by reducing their oxides with carbon.***

The highly reactive metals are extracted by the *electrolytic reduction* of their molten chlorides or oxides. Electrolytic reduction is brought about by passing electric current through the molten salt. This process is called electrolysis (which means splitting by electricity). So, we can also say that : **The highly reactive metals (which are placed high up in the reactivity series) are extracted by the electrolysis of their molten chlorides or oxides.** During electrolysis, the negatively charged electrode (cathode) acts as a powerful reducing agent by supplying electrons to reduce the metal ions into metal. ***During the electrolysis (or electrolytic reduction) of molten salts, the metals are always produced at the cathode (negative electrode).*** This is due to the fact that metal ions are always positively charged and get attracted to the negatively charged electrode (cathode) when electricity is passed through the molten metal salt (Molten salt means melted salt. Salts are melted by strong heating). The metals extracted by electrolysis method are very pure. They do not contain any impurities.

(*i*) When a molten metal chloride is electrolysed by passing electric current, then pure metal is produced at the cathode (negative electrode) and chlorine gas is formed at the anode (positive electrode).

(*ii*) When a molten metal oxide is electrolysed by passing electric current, then pure metal is produced at the cathode (negative electrode) whereas oxygen gas is formed at the anode (positive electrode).

Figure 74. This is a sodium ore called 'rock salt'. It contains sodium in the form of sodium chloride, NaCl.

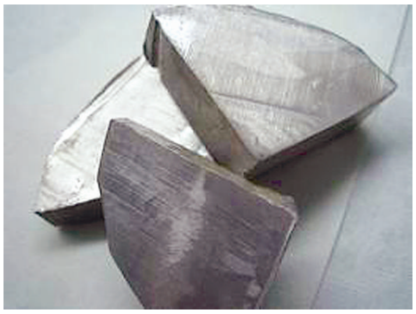


Figure 75. This sodium metal has been produced by the electrolysis of molten sodium chloride (obtained from rock salt).

The highly reactive metals potassium, sodium, calcium, and magnesium are extracted by the electrolysis of their molten chlorides whereas **aluminium metal is extracted by the electrolysis of its molten oxide.** We will now describe the extraction of two very reactive metals, sodium and aluminium, as examples.

Extraction of Sodium Metal. Sodium metal is extracted by the electrolytic reduction (or electrolysis) of molten sodium chloride. When electric current is passed through molten sodium chloride, it decomposes to form sodium metal and chlorine gas :

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

The formation of sodium and chlorine by the electrolysis of molten sodium chloride can be explained as follows : Molten sodium chloride (NaCl) contains free sodium ions (Na^+) and free chloride ions (Cl^-). During the electrolysis of molten sodium chloride, the following reactions take place at the two electrodes :

(*i*) The positive sodium ions (Na^+) are attracted to the cathode (negative electrode). The sodium ions take electrons from the cathode and get reduced to form sodium atoms (or sodium metal) :

At cathode :

$$\underset{\substack{\text{Sodium ions}\\\text{(From molten NaCl)}}}{2Na^+} + \underset{\substack{\text{Electrons}\\\text{(From cathode)}}}{2e^-} \longrightarrow \underset{\substack{\text{Sodium atoms}\\\text{(Sodium metal)}}}{2Na}$$

Thus, sodium metal is produced at the cathode (negative electrode).

(*ii*) The negative chloride ions (Cl^-) are attracted to the anode (positive electrode). The chloride ions give electrons to the anode and get oxidised to form chlorine gas :

At anode :

$$\underset{\substack{\text{Chloride ions}\\\text{(From molten NaCl)}}}{2Cl^-} - \underset{\substack{\text{Electrons}\\\text{(Given to anode)}}}{2e^-} \longrightarrow \underset{\text{Chlorine gas}}{Cl_2}$$

Thus, chlorine gas is formed at the anode (positive electrode).

Please note that **we cannot use an aqueous solution of sodium chloride to obtain sodium metal.** This is because if we electrolyse an aqueous solution of sodium chloride, then as soon as sodium metal is produced at cathode it will react with water present in the aqueous solution to form sodium hydroxide. So, electrolysis of an aqueous sodium chloride solution will produce sodium hydroxide and not sodium metal.

Please note that just like sodium metal, **potassium metal** is produced by the electrolysis of molten potassium chloride (KCl); **calcium metal** is obtained by the electrolysis of molten calcium chloride ($CaCl_2$); and **magnesium metal** is extracted by the electrolysis of molten magnesium chloride ($MgCl_2$).

Extraction of Aluminium Metal. Aluminium metal is extracted by the electrolytic reduction (or electrolysis) of molten aluminium oxide. When electric current is passed through molten aluminium oxide, it decomposes to form aluminium metal and oxygen gas :

$$\underset{\substack{\text{Aluminium oxide}\\\text{(Molten)}}}{2Al_2O_3\ (l)} \xrightarrow{\text{Electrolysis}} \underset{\text{Aluminium metal}}{4Al\ (l)} + \underset{\text{Oxygen}}{3O_2\ (g)}$$

The formation of aluminium and oxygen by the electrolysis of molten aluminium oxide can be explained as follows : Molten aluminium oxide (Al_2O_3) contains free aluminium ions (Al^{3+}) and free oxide ions (O^{2-}). During the electrolysis of molten aluminium oxide, the following reactions take place at the two electrodes :

(*i*) The positively charged aluminium ions (Al^{3+}) are attracted to the cathode (negative electrode). The aluminium ions accept electrons from the cathode and get reduced to form aluminium atoms (or aluminium metal) :

At cathode :

$$\underset{\substack{\text{Aluminium ion}\\\text{(From molten }Al_2O_3)}}{Al^{3+}} + \underset{\substack{\text{Electrons}\\\text{(From cathode)}}}{3e^-} \longrightarrow \underset{\substack{\text{Aluminium atom}\\\text{(Aluminium metal)}}}{Al}$$

Thus, aluminium metal is formed at the cathode.

(*ii*) The negatively charged oxide ions (O^{2-}) are attracted to the anode (positive electrode). The oxide ions give electrons to the anode and get oxidised to form oxygen gas :

At anode :

$$\underset{\substack{\text{Oxide ions}\\\text{(From molten }Al_2O_3)}}{2O^{2-}} - \underset{\substack{\text{Electrons}\\\text{(Given to anode)}}}{4e^-} \longrightarrow \underset{\text{Oxygen gas}}{O_2}$$

Thus, oxygen gas is produced at the anode.

(*a*) This is aluminium ore called bauxite. It contains hydrated aluminium oxide, $Al_2O_3.2H_2O$

(*b*) This is purified aluminium oxide (which has been obtained from bauxite ore)

(*c*) This is the electrolytic tank in which molten aluminium oxide is electrolysed to obtain aluminium metal

(*d*) Aluminium metal

Figure 76. Extraction of aluminium metal.

Extraction of Moderately Reactive Metals

The moderately reactive metals such as zinc, iron, tin and lead, etc., are placed in the middle of the reactivity series. So, the extraction of moderately reactive metals means the extraction of metals which are in the middle of reactivity series. **The moderately reactive metals which are in the middle of reactivity series are extracted by the reduction of their oxides with carbon, aluminium, sodium or calcium.** Some of the moderately reactive metals occur in nature as oxides but others occur as their carbonate or sulphide ores. Now, **it is easier to obtain metals from their oxides (by reduction) than from carbonates or sulphides.** So, before reduction can be done, the ore must be converted into metal oxide which can then be reduced. **The concentrated ores can be converted into metal oxide by the process of calcination or roasting.** The method to be used depends on the nature of the ore. A *carbonate ore* is converted into oxide by *calcination* whereas a *sulphide ore* is converted into oxide by *roasting.*

(*i*) **Calcination is the process in which a carbonate ore is heated strongly in the absence of air to convert it into metal oxide.** For example, zinc occurs as zinc carbonate in calamine ore, $ZnCO_3$. So, in order to extract zinc metal from zinc carbonate, this zinc carbonate should be first converted into zinc oxide. This is done by calcination. Thus, when calamine ore (zinc carbonate) is heated strongly in the absence of air, that is, when calamine is calcined, it decomposes to form zinc oxide and carbon dioxide :

$$\underset{\substack{\text{Zinc carbonate}\\\text{(Calamine ore)}}}{ZnCO_3\,(s)} \xrightarrow{\text{Calcination}} \underset{\text{Zinc oxide}}{ZnO\,(s)} + \underset{\text{Carbon dioxide}}{CO_2\,(g)}$$

Thus, *calcination converts zinc carbonate into zinc oxide.*

(*ii*) **Roasting is the process in which a sulphide ore is strongly heated in the presence of air to convert it into metal oxide.** For example, zinc occurs as sulphide in zinc blende ore, ZnS. So, in order to extract zinc metal from zinc sulphide, this zinc sulphide has to be converted into zinc oxide first. This is done by roasting. When zinc blende ore (zinc sulphide) is strongly heated in air (roasted), it forms zinc oxide and sulphur dioxide :

$$\underset{\substack{\text{Zinc sulphide}\\\text{(Zinc blende ore)}}}{2ZnS\,(s)} + \underset{\substack{\text{Oxygen}\\\text{(From air)}}}{3O_2\,(g)} \xrightarrow{\text{Roasting}} \underset{\text{Zinc oxide}}{2ZnO\,(s)} + \underset{\text{Sulphur dioxide}}{2SO_2\,(g)}$$

Thus, *roasting converts zinc sulphide into zinc oxide.*

The metal oxides (obtained by calcination or roasting of ores) are converted to the free metal by using reducing agents like carbon, aluminium, sodium or calcium. The reducing agent used depends on the *chemical reactivity* of the metal to be extracted.

(i) Reduction of Metal Oxide With Carbon. The oxides of comparatively less reactive metals like zinc, iron, nickel, tin, lead and copper, are usually reduced by using carbon as the reducing agent. In the reduction by carbon, the metal oxide is mixed with carbon (in the form of coke) and heated in a furnace. Carbon reduces the metal oxide to free metal. Here is an example.

Zinc metal is extracted by the reduction of its oxide with carbon (or coke). Thus, when zinc oxide is heated with carbon, zinc metal is produced :

$$\underset{\text{Zinc oxide}}{ZnO\,(s)} + \underset{\substack{\text{Carbon}\\ \text{(Reducing agent)}}}{C\,(s)} \longrightarrow \underset{\text{Zinc metal}}{Zn\,(s)} + \underset{\text{Carbon monoxide}}{CO\,(g)}$$

Carbon is a cheap reducing agent but it contaminates the metal.

(*a*) Calamine ore (*b*) Zinc blende ore (*c*) Zinc oxide (*d*) Carbon (as Coke) (*e*) Zinc metal

Figure 77. Extraction of zinc metal.

Iron metal is extracted from its oxide ore 'haematite' (Fe_2O_3) by reduction with *carbon* (in the form of coke). **Tin and lead metals** are also extracted by the reduction of their oxides with *carbon.* Even the less reactive metal **copper** is extracted by the reduction of its oxide with *carbon.*

(ii) Reduction of Metal Oxide With Aluminium. A more reactive metal like aluminium can also be used as a reducing agent in the extraction of metals from their oxides. Aluminium is used as a reducing agent in those cases where the metal oxide is of a comparatively more reactive metal than zinc, etc., which cannot be satisfactorily reduced by carbon. This is because *a more reactive metal (like aluminium) can displace a comparatively less reactive metal from its metal oxide to give free metal.* Thus, displacement reactions can also be used to reduce certain metal oxides into free metals. For example, the oxides of manganese and chromium metals are not satisfactorily reduced by carbon. So, **manganese and chromium metals are extracted by the reduction of their oxides with aluminium powder.** Aluminium powder reduces the metal oxide to metal and is itself oxidised to aluminium oxide. We will now give the example of extraction of manganese metal by using aluminium as the reducing agent.

Manganese metal is extracted by the reduction of its oxide with aluminium powder as the reducing agent. Thus, when manganese dioxide is heated with aluminium powder, then manganese metal is produced :

$$\underset{\substack{\text{Manganese}\\ \text{dioxide}}}{3MnO_2\,(s)} + \underset{\substack{\text{Aluminium powder}\\ \text{(Reducing agent)}}}{4Al\,(s)} \longrightarrow \underset{\substack{\text{Manganese}\\ \text{metal}}}{3Mn\,(l)} + \underset{\substack{\text{Aluminium}\\ \text{oxide}}}{2Al_2O_3\,(s)} + \text{Heat}$$

This is a displacement reaction between MnO_2 and Al (which is also an oxidation and reduction reaction). This example illustrates the use of a displacement reaction in the extraction of metals. **The reduction of manganese dioxide with aluminium is a highly *exothermic* reaction.** A lot of *heat* is evolved during the reduction of manganese dioxide with aluminium powder because of which the manganese metal produced is in the molten state (or liquid state). Please note that aluminium is an expensive reducing agent as compared to carbon (coke). From the above discussion we conclude that the two reducing agents which are commonly used in the extraction of metals are (*i*) carbon (in the form of coke), and (*ii*) aluminium powder. The use of sodium and calcium metals as reducing agents in the extraction of metals will be discussed in higher classes.

Thermite Reaction. The reduction of a metal oxide to form metal by using aluminium powder as a reducing agent is called a **thermite reaction** (or thermite process). The reactions of metal oxides with aluminium powder to produce metals are highly exothermic in which a large amount of heat is evolved. In fact, the amount of heat evolved is so large that the metals are produced in the molten state. This property of the reduction by aluminium is made use of in thermite welding for joining the broken pieces of heavy iron objects like girders, railway tracks or cracked machine parts. This is done as follows :

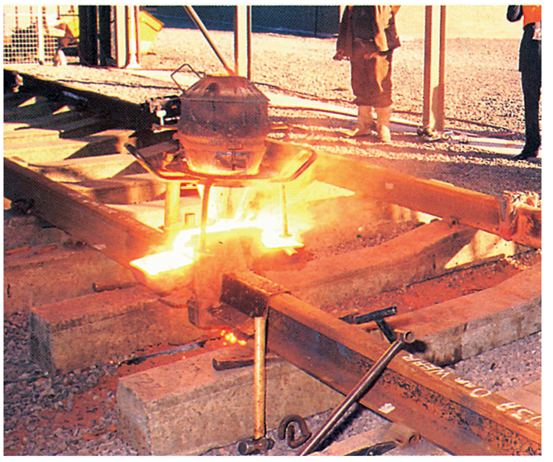


Figure 78. This picture shows thermite welding of broken rail track. The molten iron formed from the reaction between Fe_2O_3 and Al is run into a mould around the rails to be welded (or joined). When the molten iron has cooled, the mould is removed, and excess iron trimmed off.

A mixture of iron (III) oxide and aluminium powder is ignited with a burning magnesium ribbon. Aluminium reduces iron oxide to produce iron metal with the evolution of lot of heat. Due to this heat, **iron metal is produced in the molten state.**

$$\underset{\text{Iron (III) oxide}}{Fe_2O_3\ (s)} + \underset{\substack{\text{Aluminium powder}\\ \text{(Reducing agent)}}}{2Al\ (s)} \longrightarrow \underset{\substack{\text{Iron metal}\\ \text{(Molten state)}}}{2Fe\ (l)} + \underset{\text{Aluminium oxide}}{Al_2O_3\ (s)} + \text{Heat}$$

The molten iron is then poured between the broken iron pieces to weld them (to join them). This process is called aluminothermy or thermite welding. Thus, thermite welding makes use of the reducing property of aluminium.

Extraction of Less Reactive Metals

The less reactive metals such as mercury and copper, etc., are placed quite low in the reactivity series. So, the extraction of less reactive metals means the extraction of metals which are quite low in the reactivity series. **The less reactive metals which are quite low in the activity series are extracted by the reduction of their oxides by heat alone.** For example, mercury and copper are less reactive metals which are placed quite low in the reactivity series. So, mercury and copper metals are extracted by the reduction of their oxides by heat alone. This is described below.

(i) Extraction of Mercury. Mercury is a less reactive metal which is quite low in the activity series. **Mercury metal can be extracted just by heating its sulphide ore in air.** This happens as follows.

Mercury metal is produced from the sulphide ore called cinnabar, HgS, which is actually mercury (II) sulphide. The extraction of mercury from cinnabar ore involves the following two steps :

(*a*) The concentrated mercury (II) sulphide ore (cinnabar ore) is roasted in air when mercury (II) oxide is formed :

Figure 79. This is a mercury ore called cinnabar. It contains mercury (II) sulphide, HgS. Mercury metal can be obtained by simply heating this cinnabar ore in air.

Figure 80. This is mercury metal (which has been obtained from cinnabar ore). Please note that mercury metal is a liquid.

$$\underset{\substack{\text{Mercury (II) sulphide} \\ \text{(Cinnabar ore)}}}{2HgS\ (s)} + \underset{\substack{\text{Oxygen} \\ \text{(From air)}}}{3O_2\ (g)} \xrightarrow{\text{Roasting}} \underset{\text{Mercury (II) oxide}}{2HgO\ (s)} + \underset{\text{Sulphur dioxide}}{2SO_2\ (g)}$$

(*b*) When this mercury (II) oxide is heated to about 300°C, it decomposes (gets reduced) to form mercury metal :

$$\underset{\substack{\text{Mercury (II) oxide} \\ \text{(Formed above)}}}{2HgO\ (s)} \xrightarrow[\text{(Reduction)}]{\text{Heat}} \underset{\text{Mercury metal}}{2Hg\ (l)} + \underset{\text{Oxygen}}{O_2\ (g)}$$

Thus, here mercury metal has been produced by the reduction of mercury (II) oxide by heat alone. Please note that mercury (II) sulphide is also called mercuric sulphide and mercury (II) oxide is also known as mercuric oxide.

(ii) Extraction of Copper. Copper is a less reactive metal which is quite low in the reactivity series. **Copper metal can be extracted just by heating its sulphide ore in air.** This happens as follows.

One of the ores from which copper metal is produced is copper glance, Cu_2S, which is actually copper (I) sulphide. The extraction of copper from copper glance ore involves the following two steps :

(*a*) The concentrated copper (I) sulphide ore (copper glance) is roasted in air when a part of copper (I) sulphide is oxidised to copper (I) oxide :

$$\underset{\substack{\text{Copper (I) sulphide} \\ \text{(Copper glance ore)}}}{2Cu_2S\ (s)} + \underset{\substack{\text{Oxygen} \\ \text{(From air)}}}{3O_2\ (g)} \xrightarrow{\text{Roasting}} \underset{\text{Copper (I) oxide}}{2Cu_2O\ (s)} + \underset{\text{Sulphur dioxide}}{2SO_2\ (g)}$$

(*b*) When a good amount of copper (I) sulphide has been converted into copper (I) oxide, then the supply of air for roasting is stopped. In the absence of air, copper (I) oxide formed above reacts with the remaining copper (I) sulphide to form copper metal and sulphur dioxide :

$$\underset{\substack{\text{Copper (I) oxide} \\ \text{(Formed above)}}}{2Cu_2O\ (s)} + \underset{\substack{\text{Copper (I) sulphide} \\ \text{(From unoxidised ore)}}}{Cu_2S\ (s)} \xrightarrow{\text{Heat}} \underset{\text{Copper metal}}{6Cu\ (s)} + \underset{\text{Sulphur dioxide}}{SO_2\ (g)}$$

The oxides of moderately reactive metals like chromium, manganese, zinc, iron, tin and lead, etc., which occur in the middle of reactivity series, cannot be reduced by heating alone. They need a reducing agent (such as carbon or aluminium) for their reduction to metals. This has already been discussed.

Figure 81. This is the copper ore called copper glance. It contains copper (I) sulphide, Cu_2S.

Figure 82. These copper pipes have been made from copper metal extracted from copper glance ore.

4. Refining of Metals

The metals prepared by the various reduction processes usually contain some impurities, so they are impure. **The process of purifying impure metals is called refining of metals.** Thus, refining of metals means purification of metals. The method to be used for refining an impure metal depends on the nature of metal as well as on the nature of impurities present in it. Different refining methods are used for different metals. **The most important and most widely used method for refining impure metals is electrolytic refining.** This is described below.

Electrolytic Refining. Electrolytic refining means refining by electrolysis. Many metals like copper, zinc, tin, lead, chromium, nickel, silver and gold are refined electrolytically.

For the refining of an impure metal by electrolysis :

(*a*) **A thick block of the impure metal is made anode** (It is connected to the positive terminal of the battery).

(*b*) **A thin strip of the pure metal is made cathode** (It is connected to the negative terminal of the battery).

(*c*) **A water soluble salt (of the metal to be refined) is taken as electrolyte.**

On passing electric current, impure metal dissolves from the anode and goes into the electrolyte solution. And pure metal from the electrolyte deposits on the cathode. The soluble impurities present in the impure metal go into the solution whereas the insoluble impurities settle down at the bottom of the anode as 'anode mud'. We will now take an example to make the electrolytic refining of metals more clear. Let us describe the refining of copper metal by this method.

Electrolytic Refining of Copper. The apparatus used for the electrolytic refining of copper has been shown in Figure 83. The apparatus consists of an electrolytic tank containing **acidified copper sulphate solution as electrolyte** (The copper sulphate solution is acidified with dilute sulphuric acid). **A thick block of impure copper metal is made anode** (it is connected to the +ve terminal of the battery), and **a thin strip of pure copper metal is made cathode** (it is connected to the –ve terminal of the battery).

On passing electric current, impure copper from the anode dissolves and goes into copper sulphate solution, and pure copper from the copper sulphate solution deposits on cathode. Thus, **pure copper metal is produced on the cathode.** The soluble impurities go into the solution whereas insoluble impurities collect below the anode as anode mud (see Figure 83).

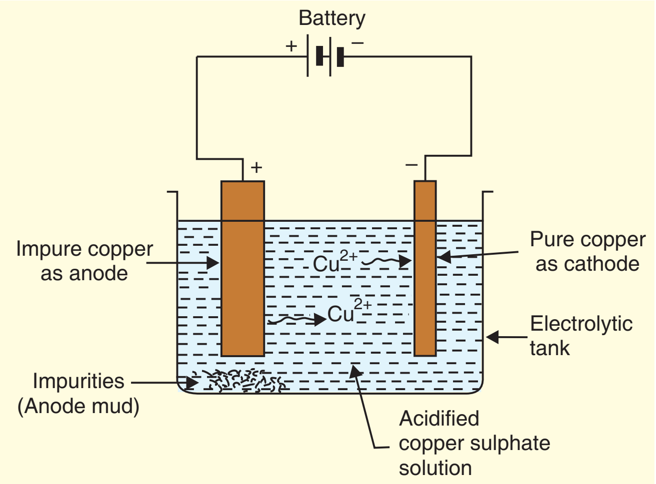


Figure 83. Experimental set up for the electrolytic refining of copper.

Explanation. Copper sulphate solution ($CuSO_4$ solution) contains copper ions, Cu^{2+}, and sulphate ions, SO_4^{2-}. On passing the electric current through copper sulphate solution, the following reactions take place at the two electrodes :

(*i*) The positively charged copper ions, Cu^{2+}, from the copper sulphate solution go to the negative electrode (cathode) and by taking electrons from the cathode, get reduced to copper atoms :

***At cathode* :**

Cu^{2+}	+	$2e^-$	$\longrightarrow$	Cu
Copper ion		Electrons		Copper atom
(From electrolyte)		(From cathode)		(Deposits on cathode)

These copper atoms get deposited on cathode giving pure copper metal.

(*ii*) Copper atoms of the impure anode lose two electrons each to anode and form copper ions, Cu^{2+}, which go into the electrolyte solution (this requires less energy than the discharge of SO_4^{2-} ions) :

***At anode* :**

Cu	–	$2e^-$	$\longrightarrow$	Cu^{2+}
Copper atom		Electrons		Copper ion
(From impure anode)		(Given to anode)		(Goes into electrolyte)

In this way **copper ions are taken from the copper sulphate solution at the cathode and put into the solution at the anode**. As the process goes on, impure anode becomes thinner and thinner whereas pure cathode becomes thicker and thicker. Thus, pure copper is obtained at the cathode.

We will now discuss **what happens to the metallic impurities present in the impure copper (crude copper) which is being refined.** The metallic impurities present in impure copper can either be *more reactive* or *less reactive*. Now, the more reactive metals like iron present in impure copper, pass into the electrolyte solution and remain there. On the other hand, the less reactive metals like gold and silver present in the impure copper, collect at the bottom of electrolytic cell below the anode in the form of anode mud. **Gold and silver metals can be recovered from the anode mud.** Thus, the electrolytic refining of metals serves two purposes :

(*i*) It refines (purifies) the metal concerned.

(*ii*) It enables to recover other valuable metals (like gold and silver) present as impurities in the metal being refined.

It is clear from the above discussion on the extraction of metals that several steps are involved in the production of pure metals from their naturally occurring ores. **A summary of the various steps involved in the extraction of pure metals from their ores is given below.**

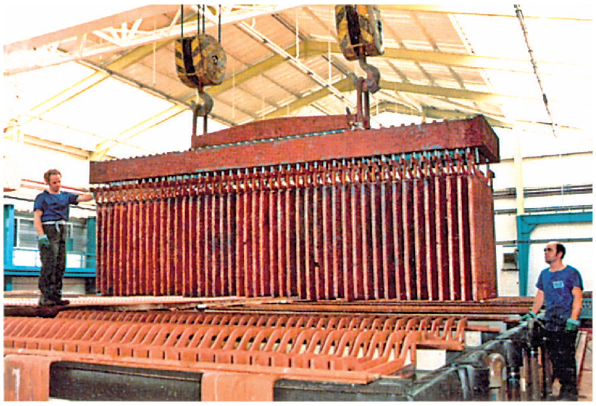
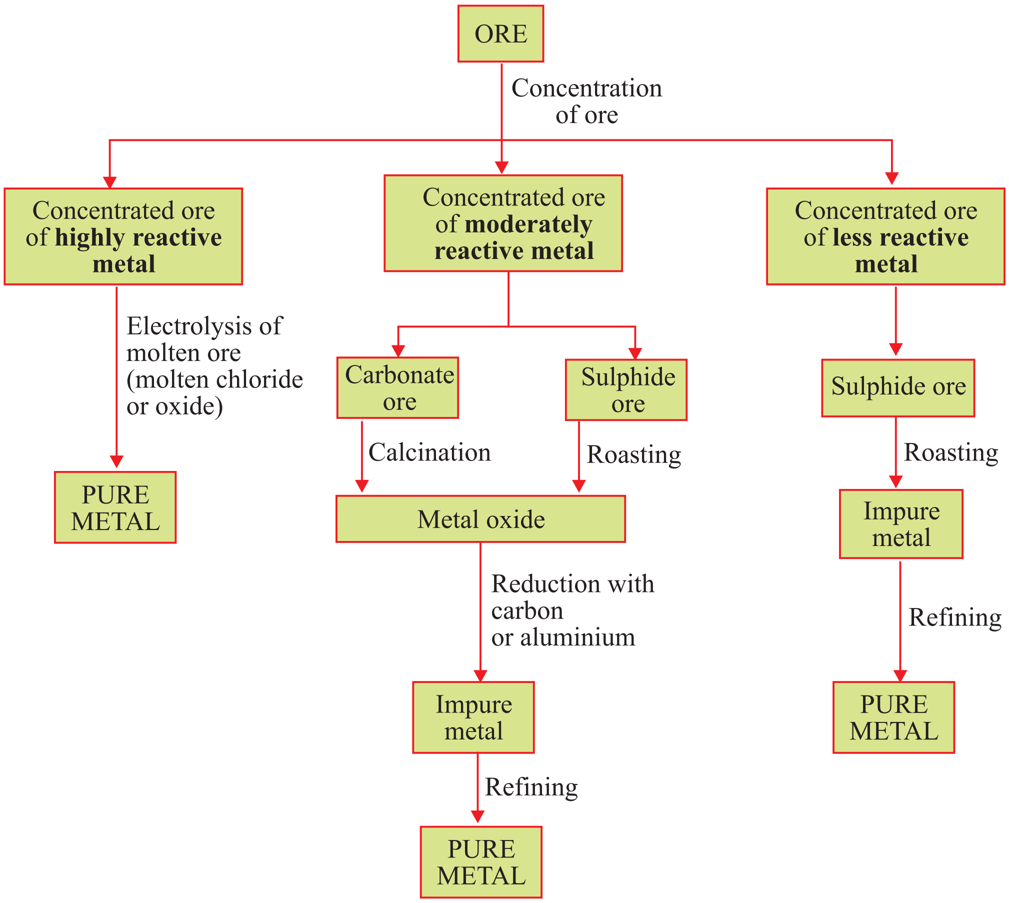


Figure 84. This picture shows impure copper anodes being transferred to an electrolytic tank for refining (or purification).



Please note that if an ore gives *carbon dioxide* on heating or on treatment with a dilute acid, it will be a *carbonate ore.* On the other hand, if an ore gives *sulphur dioxide* on heating in air, then it will be a *sulphide ore* or if an ore gives *hydrogen sulphide gas* (H_2S gas) on treatment with a dilute acid, then also it will be a *sulphide ore.*

Let us solve one problem now.

Sample Problem. An ore gives carbon dioxide on treatment with a dilute acid. What steps will you take to convert such a concentrated ore into free metal ?

Solution. Whenever a metal carbonate reacts with a dilute acid, carbon dioxide is formed. Since this ore gives carbon dioxide on treatment with a dilute acid, so it is a carbonate ore. A carbonate ore can be converted into free metal in two steps : Calcination and Reduction.

(*i*) The carbonate ore is strongly heated in the absence of air (calcined) to get the metal oxide :

$$\underset{\text{(Carbonate ore)}}{\text{Metal carbonate}} \xrightarrow{\text{Calcination}} \text{Metal oxide} + \text{Carbon dioxide}$$

(*ii*) The metal oxide is reduced with carbon to get free metal :

$$\text{Metal oxide} + \text{Carbon} \xrightarrow{\text{Reduction}} \text{Metal} + \text{Carbon monoxide}$$

CORROSION

If a metal is reactive, its surface may be attacked slowly by the air and water (moisture) in the atmosphere. The metal reacts with the oxygen of air and water vapour of air forming compounds on its surface. The formation of these compounds tarnishes the metal, that is, it makes the surface of metal appear dull. The compounds formed on the surface of metal are usually porous and gradually fall off from the surface of metal, and then the metal underneath is attacked by air and water. This process goes on and on. In this way, the action of air and water gradually eats up the whole metal. At some places (especially in industrial areas) there are some acidic gases in the air which mix with rain water to form chemicals such as acids. These acids also attack the surface of metals and eat them up slowly. We can now define corrosion as follows.

The eating up of metals by the action of air, moisture or a chemical (such as an acid) on their surface is called corrosion. Most of the metals corrode when they are kept exposed to damp air (or moist air). For example, iron metal corrodes when kept in damp air for a considerable time. When an iron object is kept in damp air for a considerable time, then a red-brown substance called 'rust' is formed on its surface. Rust is soft and porous, and it gradually falls off from the surface of iron object, and then the iron below starts corroding. Thus, **corrosion of iron is a continuous process which ultimately eats up the whole iron object.** The corrosion of metals is a highly undesirable process. A large amount of metals is lost every year because of corrosion. In general, the more reactive a metal is, the more readily it corrodes.

Figure 85. The corrosion of iron is called rusting. Rusting eats up iron objects gradually and makes them useless. This picture shows a rusted iron gate post.

The corrosion of iron is called rusting. While other metals are said to 'corrode', iron metal is said to 'rust'. In fact, most of the examples of corrosion which we come across in our daily life are due to the rusting of iron. The rusting of iron is actually the most troublesome and damaging form of corrosion. We will now discuss the rusting of iron and its prevention in detail.

Rusting of Iron

When an iron object is left in damp air (or water) for a considerable time, it gets covered with a red-brown flaky substance called rust. This is called rusting of iron. During the rusting of iron, iron metal combines with the oxygen of air in the presence of water to form hydrated iron (III) oxide, $Fe_2O_3.xH_2O$. This hydrated iron (III) oxide is called rust. So, ***rust is mainly hydrated iron (III) oxide,*** $\boldsymbol{Fe_2O_3.xH_2O}$ (the number of molecules of water x varies, it is not fixed). Rust is red-brown in colour. We have all seen iron nails, screws, pipes, and railings covered with red-brown rust here and there. It is not only the iron which rusts, even the steel rusts on being exposed to damp air (or on being kept in water). But steel rusts less readily than iron. We will now describe the conditions which are necessary for the rusting of iron.

Conditions Necessary for the Rusting of Iron

Rusting of iron (or corrosion of iron) needs both, *air* and *water.* Thus, two conditions are necessary for the rusting of iron to take place :

1. Presence of air (or oxygen)

2. Presence of water (or moisture)

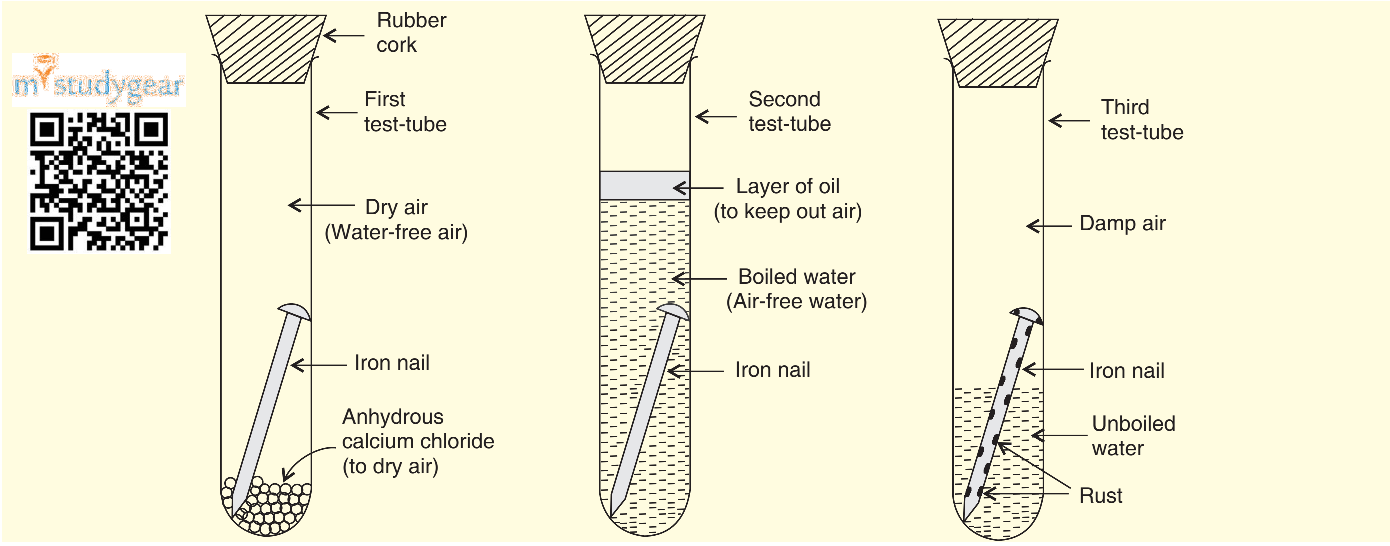
We know that iron rusts when placed in damp air (moist air) or when placed in water. Now, damp air (or moist air) also contains water vapour. Thus, **damp air alone supplies both the things, air and water, required for the rusting of iron.** Again, ordinary water has always some air dissolved in it. So, **ordinary water alone also supplies both the things, air and water, needed for rusting.** We will now describe an experiment to show that *air and water together* are necessary for the rusting of iron.

Experiment to Show that Rusting of Iron Requires Both, Air and Water

We take three test-tubes and put one clean iron nail in each of the three test-tubes :

1. In the first test-tube containing iron nail, we put some anhydrous calcium chloride and close its mouth with a tight cork [see Figure 86(*a*)]. The anhydrous calcium chloride is added to absorb water (or moisture) from the damp air present in the test-tube and make it dry. In this way, **the iron nail in the first test-tube is kept in dry air (having no water vapour in it).** This test-tube is kept aside for about one week.

2. In the second test-tube containing iron nail, we put boiled distilled water [see Figure 86(*b*)]. Boiled water does not contain any dissolved air (or oxygen) in it (This is because the process of boiling removes



(*a*) Only air : No rusting (*b*) Only water : No rusting (*c*) Air and water together : Rusting takes place

Figure 86.

all the dissolved air from it). A layer of oil is put over boiled water in the test-tube to prevent the outside air from mixing with boiled water. In this way, **the iron nail in the second test-tube is kept in air-free, boiled water.** The mouth of this test-tube is closed with a cork and it is kept aside for about one week.

3. In the third test-tube containing an iron nail, we put unboiled water so that about two-thirds of the nail is immersed in water and the rest is above the water, exposed to damp air [see Figure 86(*c*)]. In this way, **the iron nail in the third test-tube has been placed in air and water together.** The mouth of this test-tube is closed with a cork and it is also kept aside for about one week.

After one week, we observe the iron nails kept in all the three test-tubes, one by one. We find that :

(*i*) No rust is seen on the surface of iron nail kept in dry air (water-free air) in the first test-tube [see Figure 86(*a*)]. This tells us that *rusting of iron does not take place in air alone.*

(*ii*) No rust is seen on the surface of iron nail kept in air-free, boiled water in the second test-tube [see Figure 86(*b*)]. This tells us that *rusting of iron does not take place in water alone.*

(*iii*) Red-brown rust is seen on the surface of iron nail kept in the presence of both air and water together in the third test-tube [see Figure 86(*c*)]. This tells us that *rusting of iron takes place in the presence of both air and water together.*

The above experiment shows that for the rusting of iron to take place, both air (oxygen) and water (moisture) are essential. This means that **the rusting of iron objects can be prevented if damp air is not allowed to come in contact with iron objects.** We will now discuss how the rusting of iron objects can be prevented by various methods.

Prevention of Rusting

The wasting of iron objects due to rusting causes a big loss to the country's economy, so it must be prevented. Several methods are used to protect the iron objects from rusting (or corrosion). *Most of the methods involve coating the iron object with 'something' to keep out air and water (which cause rusting).* The various common methods of preventing the rusting of iron (or corrosion of iron) are given below :

1. Rusting of iron can be prevented by painting. The most common method of preventing the rusting of iron (or corrosion of iron) is to coat its surface with a paint. When a coat of paint is applied to the surface of an iron object, then air and moisture cannot come in contact with the iron object and hence no rusting takes place. The iron articles such as window grills, railings, steel furniture, iron pipes, iron bridges, railway coaches, ships, and bodies of cars, buses and trucks, etc., are all painted to protect them from rusting [see Figure 87(*a*)].

2. Rusting of iron can be prevented by applying grease or oil. When some grease or oil is applied to the surface of an iron object, then air and moisture cannot come in contact with it and hence rusting is prevented. For example, the tools and machine parts made of iron and steel are smeared with grease or oil to prevent their rusting [see Figure 87(*b*)].

(*a*) The iron pipe is being painted to prevent rusting

(*b*) Oil is being applied to bicycle chain and gear wheel to prevent rusting

(*c*) This bucket is made of iron sheet coated with zinc (galvanised iron sheet) to prevent rusting

Figure 87. Some methods of preventing rusting of iron and steel objects.

3. Rusting of iron can be prevented by galvanisation. The process of depositing a thin layer of zinc metal on iron objects is called galvanisation. Galvanisation is done by dipping an iron object in molten zinc metal. A thin layer of zinc metal is then formed all over the iron object. This thin layer of zinc metal on the surface of iron objects protects them from rusting because zinc metal does not corrode on exposure to damp air. The iron sheets used for making buckets, drums, dust-bins and sheds (roofs) are galvanised to prevent their rusting [see Figure 87 (*c*)]. The iron pipes used for water supply are also galvanised to prevent rusting.

We will now explain **how a more reactive metal zinc is able to protect iron from rusting.** Zinc is a quite reactive metal. The action of air on zinc metal forms a very thin coating of zinc oxide all over it. This zinc oxide coating is hard and impervious to air and hence prevents the further corrosion of zinc metal (because air is not able to pass through this hard zinc oxide coating). So, when a layer of zinc metal is deposited on an iron object, then the zinc oxide coating formed on its surface protects the zinc metal of zinc layer as well as the iron below it. Please note that **the galvanised iron object remains protected against rusting even if a break occurs in the zinc layer.** This is because zinc is more easily oxidised than iron. So, when zinc layer on the surface of galvanised iron object is broken, then zinc continues to corrode but iron object does not corrode or rust.

4. Rusting of iron can be prevented by tin-plating and chromium-plating. Tin and chromium metals are resistant to corrosion. So, when a thin layer of tin metal (or chromium metal) is deposited on iron and steel objects by electroplating, then the iron and steel objects are protected from rusting. For example, tiffin-boxes made of steel are nickel-plated from inside and outside to protect them from rusting. Tin is used for plating tiffin-boxes because it is non-poisonous and hence does not contaminate the food kept in them. Chromium-plating is done on taps, bicycle handle bars and car bumpers made of iron and steel to protect them from rusting and give them a shiny appearance (see Figure 88).

Figure 88. This tap made of iron (or steel) has been chromium-plated to prevent its rusting and also to give it a shining appearance.

5. Rusting of iron can be prevented by alloying it to make stainless steel. When iron is alloyed with chromium and nickel, then stainless steel is obtained. Stainless steel does not rust at all. Cooking utensils, knives, scissors and surgical instruments, etc., are made of stainless steel and do not rust at all. But stainless steel is too expensive to be used in large amounts. Please note that in the 'stainless steel formation' method of rust prevention, the iron is not coated with anything.

Corrosion of Aluminium

It is a common observation that aluminium vessels lose their shine and become dull very soon after use. This is due to the corrosion of aluminium metal when exposed to moist air. When a shining aluminium vessel is exposed to moist air, the oxygen of air reacts with aluminium to form a thin, dull layer of aluminium oxide all over the vessel. **Due to the formation of a dull layer of aluminium oxide on exposure to moist air, the aluminium vessel loses its shine very soon after use.**

Aluminium metal is more reactive than iron. So, fresh aluminium metal begins to corrode quickly when it comes in contact with moist air. The action of moist air on aluminium metal forms a thin layer of aluminium oxide all over the aluminium metal. This aluminium oxide layer is very tough and prevents the metal underneath from further corrosion (because moist air is not able to pass through this aluminium oxide layer). In this way, **a thin aluminium oxide layer formed on the surface of aluminium objects protects them from further corrosion.** This means that *sometimes corrosion is useful.* Because a newly cut piece of aluminium metal corrodes quickly to form a strong layer of aluminium oxide on its surface which then protects the aluminium piece from further corrosion. Please note that the aluminium articles (like aluminium vessels) are not attacked by air and water due to the presence of protective oxide layer, and hence not easily corroded. Thus, **a common metal which is highly resistant to corrosion is aluminium.**

We have just said that the formation of a thin aluminium oxide layer on the surface of aluminium objects on exposure to moist air, protects the aluminium objects from further corrosion. If the aluminium oxide layer on the surface of aluminium objects could somehow be made thicker, then the aluminium objects would be protected from corrosion even more effectively. This can be done by a process called 'anodising'.

The layer of aluminium oxide on the surface of aluminium objects can be made thicker by electrolysis (to give them even more protection from corrosion). This process is called anodising. In this process, the aluminium object is made an anode (positive electrode) in an electrolytic tank in which dilute sulphuric acid is electrolysed. During the electrolysis of dilute sulphuric acid, oxygen gas is liberated at the anode and reacts with the aluminium object to form a thicker layer of aluminium oxide on its surface. This thicker and more uniform aluminium oxide layer protects the aluminium object from corrosion very effectively. Thus, **anodising is a process of forming a thick layer of aluminium oxide on an aluminium object by making it anode during the electrolysis of dilute sulphuric acid.** The aluminium objects like pressure cookers, cooking utensils, saucepans, and window frames, etc., are anodised to protect them from corrosion (see Figure 89). The aluminium oxide layer can also be dyed to give the objects attractive colours.

Figure 89. These pictures show two objects made of anodised aluminium to prevent their corrosion.

Corrosion of Copper

The copper objects lose their shine after some time due to the formation of a copper oxide layer on them. When a copper object remains in damp air for a considerable time, then copper reacts slowly with the carbon dioxide and water of air to form a green coating of basic copper carbonate on the surface of the object. The formation of this green coating on the surface of a copper object corrodes it (see Figure 90). Please note that the green coating of basic copper carbonate is a mixture of copper carbonate and copper hydroxide, $CuCO_3.Cu(OH)_2$. Since copper metal is low in the reactivity series, therefore, the corrosion of copper metal is very, very slow. The corroded copper vessels can be cleaned with dilute acid solution. The acid solution dissolves green coloured basic copper carbonate present on the corroded copper vessels and makes them look shiny, red-brown again.

Figure 90. Copper metal corrodes in air to form a green substance called basic copper carbonate. The copper dome over this building is green because copper metal has reacted slowly with the carbon dioxide and moisture (water) of air to form a green coating of basic copper carbonate.

Corrosion of Silver

When a shining metal object loses its shine and becomes dull, we say that it has been tarnished. When silver objects are kept in air, they get tarnished and gradually turn black. This can be explained as follows : Silver is a highly unreactive metal so it does not react with the oxygen of air easily. But air usually contains a little of sulphur compounds such as hydrogen sulphide gas (H_2S). So, the silver objects combine slowly

Figure 91. The two silver coins on the left side in this picture are freshly made so they are bright and shiny. On the other hand, the two silver coins on the right side are very old so they have been tarnished (or corroded) by the action of air and hence turned black.

with the hydrogen sulphide gas present in air to form a black coating of silver sulphide (Ag_2S). The shining silver objects become tarnished due to the formation of silver sulphide coating on their surface. Thus, **silver ornaments (and other silver articles) gradually turn black due to the formation of a thin silver**

sulphide layer on their surface by the action of hydrogen sulphide gas present in air (see Figure 91). Silver is a bright, shiny metal which is chemically quite unreactive. Silver metal loses its shine and becomes dull (or tarnished) very slowly. Thus, silver metal is fairly resistant to corrosion. **Silver metal is used to make silver coins, jewellery and silverware (such as silver utensils and decorative articles) because of its bright shiny surface and resistance to corrosion.**

The Case of Gold and Platinum

Gold is a yellow, shining metal. Gold metal does not corrode when exposed to atmosphere. Gold does not corrode because it is a highly unreactive metal which remains unaffected by air, water vapour and other gases in the atmosphere. Gold does not tarnish and retains its lustre (*chamak*) for years. **Since gold does not corrode, therefore, gold ornaments look new even after several years of use.** We can now say that : **Gold is used to make jewellery because of its bright shiny surface and high resistance to corrosion.** Please note that though gold is highly resistant to corrosion but the shine of gold ornaments decreases with time and they become somewhat dull. Such gold ornaments are polished by jewellers to make them glitter again. Another point to be noted is that gold dissolves only in aqua-regia solution.

Figure 92. Gold metal is used in leaf form on 'Golden Temple' as it is highly unreactive and does not corrode on exposure to air. The thin gold leaf coating not only adds long-lasting beauty, it also protects the marble structure beneath from corrosion.

Platinum is another metal which is highly resistant to corrosion. Platinum also dissolves only in aqua-regia. Platinum is a white metal with a silvery shine. **Platinum is used to make jewellery because of its bright shiny surface and high resistance to corrosion.** A yet another metal which is very resistant to corrosion is titanium. We can now say that the metals which do not corrode easily are silver, gold, platinum and titanium. Let us solve some problems now.

Sample Problem 1. Which of the following methods is suitable for preventing an iron frying pan from rusting ?

(*a*) applying grease (*b*) applying paint (*c*) applying a coat of zinc (*d*) all of the above

(NCERT Book Question)

Solution. The most suitable method for preventing an iron frying pan from rusting is : (*c*) applying a coat of zinc (which is called galvanisation). Please note that we cannot apply grease because it will spoil the food to be cooked in frying pan. We can also not apply paint because it will gradually come out when frying pan is heated on a gas stove during the cooking of food.

Sample Problem 2. Food cans are coated with tin and not 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.

(NCERT Book Question)

Solution. (c) zinc is more reactive than tin.

Sample Problem 3. You must have seen tarnished copper vessels being cleaned with lemon (or tamarind juice). Explain why, these sour substances are effective in cleaning these vessels . **(NCERT Book Question)**

Solution. The sour substances such as lemon (or tamarind juice) contain acids. These acids dissolve the coating of copper oxide or basic copper carbonate present on the surface of tarnished copper vessels and makes them shining red-brown again.

Sample Problem 4. A woman gave old and dull gold bangles to a goldsmith for polishing to restore their glitter. The goldsmith dipped the gold bangles in a particular solution. The bangles sparkled like new but their weight was reduced drastically. Can you guess the solution used by the dishonest goldsmith ?
(**NCERT Book Question**)

Solution. The dishonest goldsmith dipped the gold bangles in aqua-regia solution (which contains 1 part of concentrated nitric acid and 3 parts of concentrated hydrochloric acid, by volume). Aqua-regia dissolved a considerable amount of gold from gold bangles and hence reduced their weight drastically. The dishonest goldsmith can recover the dissolved gold from aqua-regia by a suitable treatment.

ALLOYS

The various properties of a metal like malleability, ductility, strength, hardness, resistance to corrosion, appearance, etc., can be improved by mixing other metals with it. This mixture of two or more metals is called an alloy. For example, **aluminium metal is light but not strong, but an alloy of aluminium with copper, magnesium and manganese (called duralumin) is light as well as strong.** Since duralumin is light and yet strong, it is used for making the aircraft bodies and parts, space satellites, and kitchen-ware like pressure cookers, etc. Similarly, **aluminium metal is light but not hard, but an alloy of aluminium with magnesium (called magnalium) is light as well as hard.** Since magnalium alloy is light and yet very hard, it is used to make balance beams and light instruments. Alloys have properties which are different from the constituent metals. In fact, it is possible to make alloys having required properties. In some alloys, however, non-metals like carbon are also present. This will become more clear from the following example.

We know that iron is the most widely used metal. But it is never used in the pure form. This is because pure iron is very soft and stretches easily when hot. But **when a small amount of carbon (varying from about 0.1 per cent to 1.5 per cent) is mixed with iron, we get an alloy called steel. This alloy of iron called steel is hard and strong. It also rusts less readily than pure iron.** The strength and other properties of steel vary with the percentage of carbon present in it. Being very hard, tough and strong, steel is used for making nails, screws, girders, bridges and railway lines, etc. It is also used for the construction of buildings, vehicles and ships. And **when iron metal is alloyed with other metals such as chromium and nickel, we get an alloy called stainless steel which is strong, tough and does not rust at all.** Since stainless steel resists corrosion, it is used for making cooking utensils, knives, scissors, tools and ornamental pieces. Stainless steel is also used for making surgical instruments and equipment for food processing industry and dairy industry. We can now define an alloy as follows :

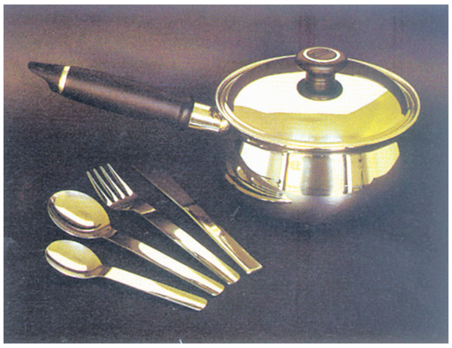
An alloy is a homogeneous mixture of two or more metals (or a metal and small amounts of non-metals). For example, *brass* is an alloy of two metals : *copper* and *zinc*, whereas *steel* is an alloy of a metal and a small amount of a non-metal : *iron* and *carbon*, **An alloy is prepared by mixing the various metals in molten state in required proportions, and then cooling their mixture to the room temperature.** The alloy of a metal and a non-metal can be prepared by first melting the metal and then dissolving the non-metal in it, followed by cooling to the room temperature.

Figure 93. To make an alloy, the molten metals are mixed, then allowed to cool.

Each alloy has certain useful properties. **The properties of an alloy are different from the properties of the constituent metals (from which it is made).** In general :

1. Alloys are stronger than the metals from which they are made.
2. Alloys are harder than the constituent metals.
3. Alloys are more resistant to corrosion.
4. Alloys have lower melting points than the constituent metals.
5. Alloys have lower electrical conductivity than pure metals.

Some of the common alloys are : Duralumin or Duralium, Magnalium, Steel, Stainless steel, Brass, Bronze, Solder and Amalgams. Duralumin and magnalium are the alloys of aluminium; steel and stainless steel are the alloys of iron; brass and bronze are the alloys of copper; solder is an alloy of lead and tin; whereas amalgams are the alloys of mercury. We have already discussed duralumin, magnalium, steel and stainless steel briefly. We will now discuss brass, bronze, solder and amalgams.



(*a*) Stainless steel alloy is used for making household utensils

(*b*) Light weight aluminium alloys are used in making aircrafts

(c) Different coloured alloys are used to make coins

Figure 94. Some of the uses of metal 'alloys'.

(*i*) **Brass.** Brass is an alloy of Copper and Zinc (Cu and Zn). It contains 80% copper and 20% zinc. Brass is more malleable and more strong than pure copper. Its colour is also more golden. Brass is used for making cooking utensils, screws, nuts, bolts, wires, tubes, scientific instruments like microscopes and ornaments. Brass is also used for making vessels like flower vases and fittings like that of fancy lamps.

(*ii*) **Bronze.** Bronze is an alloy of Copper and Tin (Cu and Sn). It contains 90% copper and 10% tin. Bronze is very tough and highly resistant to corrosion. It is used for making statues, coins, medals, cooking utensils and ship's propellers.

The electrical conductivity of an alloy is less than that of pure metals. For example, brass (an alloy of copper and zinc) and bronze (an alloy of copper and tin) are not good conductors of electricity but pure copper is an excellent conductor of electricity and used for making electrical circuits.

(*iii*) **Solder.** Solder is an alloy of lead and tin (Pb and Sn). It contains 50% lead and 50% tin. The melting point of an alloy is less than that of pure metals. Solder is an alloy which has a low melting point. So, it is used for soldering (or welding) electrical wires together.

(*iv*) **Amalgam.** An alloy of mercury metal with one or more other metals is known as an amalgam. A solution of sodium metal in liquid mercury metal is called sodium amalgam. An amalgam consisting of mercury, silver, tin and zinc is used by dentists for fillings in teeth.

(*v*) **Alloys of Gold.** The purity of gold is expressed in terms of 'carats'. **Pure gold is said to be of 24 carats.** Pure gold (known as 24 carat gold) is very soft due to which it is not suitable for making jewellery. Gold is alloyed with a small amount of silver or copper to make it hard. This harder alloy of gold is more suitable for making ornaments (because it becomes easier to work with it). In India, gold ornaments are usually made of 22 carat gold. It means that 22 parts pure gold is alloyed with 2 parts of either silver or copper for making ornaments. Thus, 22 carat gold is an alloy of gold with silver or copper.

The Iron Pillar at Delhi

The iron pillar near Qutab Minar in Delhi is made up of wrought iron (which is a low-carbon steel). This iron pillar was made around 400 BC by the Indian iron workers. Though wrought iron rusts slowly with time but the Indian iron workers had developed a process which prevented the wrought iron pillar from rusting even after thousands of years ! (see Figure 95). The rusting has been prevented because of the formation of a thin film of magnetic oxide of iron (Fe_3O_4) on the surface as a result of finishing treatment given to the pillar, painting it with a mixture of different salts, then heating and quenching (rapid cooling). The iron pillar is 8 metres high and 6000 kg (6 tonnes) in weight. This iron pillar stands in good condition more than 2000 years after it was made. **The iron pillar at Delhi is a wonder of ancient Indian metallurgy.** It tells us that ancient Indians had good knowledge of metals and their alloys. We are now in a position to **answer the following questions :**

Figure 95. This iron pillar in Delhi was made more than 2000 years ago. It has not rusted at all. Truly, a wonder of ancient Indian metallurgy. *Jai Ho* !

Very Short Answer Type Questions

1. A zinc ore gave CO_2 on treatment with a dilute acid. Identify the ore and write its chemical formula.
2. What chemical process is used for obtaining a metal from its oxide ?
3. State two ways to prevent the rusting of iron.
4. What is meant by galvanisation ? Why is it done ?
5. Name the metal which is used for galvanising iron.
6. Explain why, iron sheets are coated with zinc.
7. Why do we apply paint on iron articles ?
8. Give reason for the following :
   Carbonate and sulphide ores are usually converted into oxides during the process of extraction of metals.
9. Name a reducing agent that may be used to obtain manganese from manganese dioxide.
10. Name an alloy of lead and tin.
11. Give the composition of an alloy called solder. State its one property and one use.
12. What is an amalgam ?
13. How many carats is pure gold ? Why is pure gold not suitable for making ornaments ?
14. Name one method for the refining of metals.
15. State two conditions for the rusting of iron.
16. In one method of rust prevention, the iron is not coated with anything. Which is this method ?
17. Name two alloys of iron. What elements are present in these alloys ?
18. Give reason for the following :
    Silver, gold and platinum are used to make jewellery.
19. Which metal becomes black in the presence of hydrogen sulphide gas in air ?
20. Name the gas in air which tarnishes silver articles slowly.
21. Silver metal does not combine easily with oxygen but silver jewellery tarnishes after some time. How ?
22. Write the composition of the alloy called bronze. Give two uses of bronze.
23. Why does a new aluminium vessel lose shine so soon after use ?
24. Why do gold ornaments look new even after several years of use ?
25. Name two metals which are highly resistant to corrosion.
26. Which property of 'solder' alloy makes it suitable for welding electrical wires ?
27. Explain why, carbon cannot reduce oxides of sodium or magnesium.
28. Why are the metals like Na, K, Ca and Mg never found in their free state in nature ?
29. Name one metal each which is extracted by :
    (*a*) reduction with carbon. (*b*) electrolytic reduction.
    (*c*) reduction with aluminium (*d*) reduction with heat alone.
30. Fill in the following blanks with suitable words :
    (*a*) The corrosion of iron is called ....
    (*b*) .... and .... are necessary for the rusting of iron.
    (*c*) The process of depositing a thin layer of zinc on iron articles is called ....

(*d*) Tiffin boxes are electroplated with .... but car bumpers are electroplated with .... to protect them from rusting.

(*e*) The corrosion of copper produces a .... coating of basic copper carbonate on its surface.

(*f*) Brass is an alloy of copper and ....

(*g*) Bronze is an alloy of copper and ....

(*h*) The non-metal present in steel is ....

(*i*) The alloy in which one of the metals is mercury is called an ....

(*j*) The electrical conductivity and melting point of an alloy is .... than that of pure metals.

(*k*) The rocky material found with ores is called.............

Short Answer Type Questions

31. How is manganese extracted from manganese dioxide, MnO_2 ? Explain with the help of an equation.

32. What is a thermite reaction ? Explain with the help of an equation. State one use of this reaction.

33. Which one of the methods given in column I is applied for the extraction of each of the metals given in column II :

Column I	*Column II*
Electrolytic reduction	Aluminium
Reduction with Carbon	Zinc
Reduction with Aluminium	Sodium
	Iron
	Manganese
	Tin

34. (*a*) Give reason why copper is used to make hot water tanks but steel (an alloy of iron) is not.

(*b*) Explain why, the surface of some metals acquires a dull appearance when exposed to air for a long time.

35. (*a*) Why does aluminium not corrode right through ?

(*b*) What is meant by 'anodising' ? Why is it done ?

36. (*a*) Why is an iron grill painted frequently ?

(*b*) Explain why, though aluminium is more reactive than iron, yet there is less corrosion of aluminium when both are exposed to air.

37. (*a*) Name the method by which aluminium metal is extracted.

(*b*) Give the name and chemical formula of one ore of copper.

(*c*) How is zinc extracted from its carbonate ore (calamine) ? Explain with equations.

38. (*a*) Name two metals which occur in nature in free state as well as in combined state.

(*b*) Name one ore of manganese. Which compound of manganese is present in this ore ? Also write its chemical formula.

(*c*) A zinc ore on heating in air forms sulphur dioxide. Describe briefly any two stages involved in the conversion of this concentrated ore into zinc metal.

39. How does the method used for extracting a metal from its ore depend on the metal's position in the reactivity series ? Explain with examples.

40. Explain giving one example, how highly reactive metals (which are high up in the reactivity series) are extracted.

41. Describe with one example, how moderately reactive metals (which are in the middle of reactivity series) are extracted.

42. How are the less reactive metals (which are quite low in the reactivity series) extracted ? Explain with the help of an example.

43. What is meant by refining of a metal ? Name the most widely used method for the refining of impure metals obtained by various reduction processes. Describe this method with the help of a labelled diagram by taking the example of any metal.

44. (*a*) Define the terms (*i*) mineral (*ii*) ore, and (*iii*) gangue.

(*b*) What is meant by the 'concentration of ore' ?

(*c*) Name one ore of copper (other than cuprite). Which compound of copper is present in this ore ? Also, write its chemical formula.

45. Explain how, a reduction reaction of aluminium can be used for welding cracked machine parts of iron. Write a chemical equation for the reaction involved.

46. (*a*) What is corrosion ?

(*b*) Name any two metals which do not corrode easily.

(*c*) What is the corrosion of iron known as ?

(*d*) Explain why, aluminium is a highly reactive metal, still it is used to make utensils for cooking.

47. What is meant by 'rusting of iron' ? With the help of labelled diagrams, describe an activity to find out the conditions under which iron rusts.

48. (*a*) What is an alloy ? How is an alloy made ?

(*b*) What elements are present in steel ? How are the properties of steel different from those of pure iron ?

(*c*) Give the constituents and one use of brass.

49. (*a*) Name two metals which resist corrosion due to the formation of a thin, hard and impervious layer of oxide on their surface.

(*b*) Name five methods of preventing rusting of iron.

(*c*) What are the constituents of stainless steel ? What are the special properties of stainless steel ?

50. (*a*) Name an alloy of copper. State its chemical composition and any one use.

(*b*) Explain why, when a copper object remains in damp air for a considerable time, a green coating is formed on its surface. What is this process known as ?

51. (*a*) How does the painting of an iron object prevent its rusting ?

(*b*) How does the electrical conductivity of copper alloys, brass and bronze, differ from that of pure copper ?

(*c*) What is meant by 22 carat gold ? Name the metals which are usually alloyed with gold to make it harder.

52. Explain giving equation, what happens when :

(*a*) $ZnCO_3$ is heated in the absence of air ?

(*b*) a mixture of Cu_2O and Cu_2S is heated ?

53. (*a*) For the reduction of a metal oxide, suggest a reducing agent other than carbon.

(*b*) Explain why, an aqueous solution of sodium chloride is not used for the electrolytic extraction of sodium metal.

54. How are metals refined by the electrolytic process ? Describe the electrolytic refining of copper with the help of a neat labelled diagram.

55. (*a*) Name the chemical compound which is electrolysed in molten state to obtain aluminium metal. Which gas is evolved during this process ?

(*b*) Name the chemical compound which is electrolysed in molten state to obtain sodium metal. Which gas is produced in this process ?

(*c*) Name the gas produced when calamine ore is calcined.

(*d*) Name the gas evolved when cinnabar ore is roasted.

56. (*a*) Name two metals which are found in nature mainly in the free state (as metallic elements).

(*b*) Name two metals which are always found in combined state.

(*c*) What iron compound is present in haematite ore ? Also write its chemical formula.

Long Answer Type Questions

57. (*a*) What is the difference between a mineral and an ore ?

(*b*) Which metal is extracted from cinnabar ore ?

(*c*) Name one ore of sodium. Name the sodium compound present in this ore and write its chemical formula.

(*d*) How is sodium metal extracted ? Explain with the help of equation of the reaction involved.

(*e*) Name three other metals which are extracted in a manner similar to sodium.

58. (*a*) Name the metal which is extracted from haematite ore.

(*b*) Name one ore of aluminium. Name the aluminium compound present in this ore and write its chemical formula.

(*c*) How is aluminium metal extracted ? Explain with the help of an equation.

(*d*) Name the electrode at which aluminium metal is produced.

(*e*) Which gas is produced during the extraction of aluminium ? At which electrode is this gas produced ?

59. (*a*) Which metal is extracted from bauxite ore ?
(*b*) Give the name of one ore of iron. Which iron compound is present in this ore ? Write its chemical formula.
(*c*) Describe the extraction of zinc metal from its sulphide ore (zinc blende). Write equations of the reactions involved.
(*d*) Explain why, the galvanised iron article is protected against rusting even if the zinc layer is broken.
(*e*) Name a common metal which is highly resistant to corrosion.

60. (*a*) Name the metal which is extracted from the ore called 'rock salt'.
(*b*) Name two ores of zinc. Write the names of the chemical compounds present in them and give their chemical formulae.
(*c*) Explain how, mercury is extracted from its sulphide ore (cinnabar). Give equations of the reactions involved.
(*d*) In the electrolytic refining of a metal M, what would you take as anode, cathode and electrolyte ?
(*e*) Name any five metals which are purified by electrolytic refining method.

61. (*a*) Which metal is extracted from calamine ore ?
(*b*) Name one ore of mercury. Which mercury compound is present in this ore ? Write its chemical formula.
(*c*) How is copper extracted from its sulphide ore (copper glance), Cu_2S ? Explain with equations of the reactions involved.
(*d*) What is an alloy ? Give two examples of alloys.
(*e*) How are the properties of an alloy different from those of the constitutent elements ?

Multiple Choice Questions (MCQs)

62. An ore of manganese metal is :
(*a*) bauxite (*b*) haematite (*c*) cuprite (*d*) pyrolusite

63. Which of the following is an iron ore ?
(*a*) cinnabar (*b*) calamine (*c*) haematite (*d*) rock salt

64. The metal which can be extracted from the bauxite ore is :
(*a*) Na (*b*) Mn (*c*) Al (*d*) Hg

65. The two metals which can be extracted just by heating their sulphides in air are :
(*a*) sodium and copper (*b*) copper and aluminium
(*c*) potassium and zinc (*d*) mercury and copper

66. A common metal which is highly resistant to corrosion is :
(*a*) iron (*b*) copper (*c*) aluminium (*d*) magnesium

67. An important ore of zinc metal is :
(*a*) calamine (*b*) cuprite (*c*) pyrolusite (*d*) haematite

68. The major ore of aluminium is known as :
(*a*) cinnabar (*b*) calamine (*c*) bauxite (*d*) pyrolusite

69. The two metals which are extracted by means of electrolytic reduction of their molten salts are :
(*a*) magnesium and manganese (*b*) iron and aluminium
(*c*) zinc and magnesium (*d*) magnesium and aluminium

70. In stainless steel alloy, iron metal is mixed with :
(*a*) Cu and Cr (*b*) Cr and Ni (*c*) Cr and Sn (*d*) Cu and Ni

71. If copper is kept exposed to damp air for a considerable time, it gets a green coating on its surface. This is due to the formation of :
(*a*) hydrated copper sulphate (*b*) copper oxide (*c*) basic copper carbonate (*d*) copper nitrate

72. Which of the following alloys contains mercury as one of the constituents ?
(*a*) stainless steel (*b*) solder (*c*) duralumin (*d*) zinc amalgam

73. Which of the following is an ore of mercury metal ?
(*a*) rock salt (*b*) cinnabar (*c*) calamine (*d*) haematite

74. Calamine ore can be used to extract one of the following metals. This metal is :
(*a*) copper (*b*) mercury (*c*) aluminium (*d*) zinc

75. Which of the following pair of metals exists in their native state in nature ?
(*a*) Ag and Hg (*b*) Ag and Zn (*c*) Au and Hg (*d*) Au and Ag

76. Which of the following reactants are used to carry out the thermite reaction required for welding the broken railway tracks ?
(*a*) Al_2O_3 + Fe (*b*) MnO_2 + Al (*c*) Fe_2O_3 + Al (*d*) Cu_2O + Fe

77. Which of the following alloys contains a non-metal as one of the constituents ?
(*a*) brass (*b*) amalgam (*c*) steel (*d*) bronze

78. During the refining of an impure metal by electrolysis, the pure metal is a deposited :
(*a*) at cathode (*b*) on the walls of electrolytic tank
(*c*) at anode (*d*) at the bottom of electrolytic tank

79. Which of the following metals can be obtained from haematite ore ?
(*a*) copper (*b*) sodium (*c*) zinc (*d*) iron

80. Brass is an alloy of :
(*a*) Cu and Sn (*b*) Cu and Pb (*c*) Pb and Sn (*d*) Zn and Cu

81. The metal which is always present in an amalgam is :
(*a*) iron (*b*) aluminium (*c*) mercury (*d*) magnesium

82. Manganese metal is extracted from manganese dioxide by a reduction process by making use of :
(*a*) carbon (*b*) hydrogen (*c*) electrolysis (*d*) aluminium

83. The metal which can be extracted simply by heating the cinnabar ore in air is :
(*a*) Zn (*b*) Cu (*c*) Al (*d*) Hg

84. During galvanisation, iron metal is given a thin coating of one of the following metals. This metal is :
(*a*) chromium (*b*) tin (*c*) zinc (*d*) copper

85. Which of the following metals are extracted by the electrolysis of their molten chlorides ?
(*a*) Na and Hg (*b*) Hg and Mg (*c*) Na and Mg (*d*) Cu and Fe

86. Rock salt is an ore of one of the following metals. This metal is :
(*a*) Mn (*b*) Na (*c*) Fe (*d*) Cu

87. The articles made of silver metal become dark on prolonged exposure to air. This is due to the formation of a layer of its :
(*a*) oxide (*b*) hydride (*c*) sulphide (*d*) carbonate

88. A sulphide ore is converted into metal oxide by the process of :
(*a*) carbonation (*b*) roasting (*c*) calcination (*d*) anodising

89. The metal which can be extracted from pyrolusite ore is :
(*a*) mercury (*b*) manganese (*c*) aluminium (*d*) magnesium

90. Calamine ore can be converted into zinc oxide by the process of :
(*a*) dehydration (*b*) roasting (*c*) calcination (*d*) sulphonation

91. Zinc blende ore can be converted into zinc oxide by the process of :
(*a*) roasting (*b*) hydrogenation (*c*) chlorination (*d*) calcination

Questions Based on High Order Thinking Skills (HOTS)

92. An element A which is a part of common salt and kept under kerosene reacts with another element B of atomic number 17 to give a product C. When an aqueous solution of product C is electrolysed then a compound D is formed and two gases are liberated.
(*a*) What are A and B ?
(*b*) Identify C and D.
(*c*) What will be the action of C on litmus solution ? Why ?
(*d*) State whether element B is a solid, liquid or gas at room temperature.
(*e*) Write formula of the compound formed when element B reacts with an element E having atomic number 5.

93. A metal which exists as a liquid at room temperature is obtained by heating its sulphide ore in the presence of air.
(*a*) Name the metal and write its chemical symbol.

(*b*) Write the name and formula of the sulphide ore.
(*c*) Give the equations of chemical reactions involved in the production of metal from its sulphide ore.
(*d*) Name a common device in which this metal is used.
(*e*) Can this metal displace copper from copper sulphate solution ? Why ?

94. No chemical reaction takes place when granules of a rusty-brown solid A are mixed with the powder of another solid B. However, when the mixture is heated, a reaction takes place between its components. One of the products C is a metal and settles down in the molten state while the other product D floats over it. It was observed that the reaction is highly exothermic.
(*a*) What could the solids A and B be ?
(*b*) What are the products C and D most likely to be ?
(*c*) Write the chemical equation for the reaction between A and B leading to the formation of C and D. Mention the physical states of all the reactants and products in this equation and indicate the heat change which takes place.
(*d*) What is the special name of such a reaction ? State one use of such a reaction.
(*e*) Name any two types of chemical reactions under which the above reaction can be classified.

95. In an electrolytic tank, aluminium metal is being extracted by the electrolysis of molten aluminium oxide using carbon electrodes. It is observed that one of the carbon electrodes is gradually burnt away and has to be replaced.
(*a*) Which carbon electrode (cathode or anode) is burnt away ?
(*b*) Why is this carbon electrode burnt away ?

96. A metal X which is resistant to corrosion is produced by the electrolysis of its molten oxide whereas another metal Y which is also resistant to corrosion is produced by the reduction of its oxide with carbon. Metal X can be used in powder form in thermite welding whereas metal Y is used in making cathodes of ordinary dry cells.
(*a*) Name the metals X and Y.
(*b*) Which of the two metals is more reactive : X or Y ?
(*c*) Name one ore or metal X. Also write its chemical formula.
(*d*) Name one ore of metal Y. Also write its chemical formula.
(*e*) Name one alloy of metal X and one alloy of metal Y.

97. When an object made of metal A is kept in air for a considerable time, it loses its shine and becomes almost black due to the formation of a layer of substance B. When an object made of another metal C is kept in damp air for a considerable time, it gets covered with a green layer of substance D. Metal A is the best conductor of electricity whereas metal C is the next best conductor of electricity.
(*a*) What is metal A ?
(*b*) What is metal C ?
(*c*) Name the substance B.
(*d*) Name the substance D.
(*e*) What type of chemical can be used to remove the green layer from metal C and clean it ? Why ?

98. Four metals P, Q, R and S are all obtained by the reduction of their oxides with carbon. Metal P is used to form a thin layer over the sheets of metal S to prevent its corrosion. Metal Q is used for electroplating tiffin boxes made of metal S whereas metal R is used in making car batteries. Metals Q and R form an alloy called solder. What are metals P, Q, R and S ? How have you arrived at this conclusion ?

99. A black metal oxide XO_2 is used as a catalyst in the preparation of oxygen gas from potassium chlorate. The oxide XO_2 is also used in ordinary dry cells. The metal oxide XO_2 cannot be reduced satisfactorily with carbon to form metal X.
(*a*) Name the metal X.
(*b*) Name the metal oxide XO_2
(*c*) Which reducing agent can be used to reduce XO_2 to obtain metal X ?
(*d*) Name another metal which can also be extracted by the reduction of its oxide with the above reducing agent.

100. Metals X and Y can be recovered from the anode mud left behind after the electrolytic refining of copper metal. The coins made of metal X look new even after several years of use but the coins made of metal Y lose their shine gradually and get blackened soon. When metal X is alloyed with a small amount of metal Y, it becomes hard and hence suitable for making ornaments. What are metals X and Y ? Also state the colour of metal X.

ANSWERS

1. Calamine, $ZnCO_3$ **2.** Reduction **5.** Zinc **9.** Aluminium **16.** Alloying iron with chromium and nickel to make stainless steel **19.** Silver **20.** Hydrogen sulphide **25.** Aluminium and Zinc **26.** Low melting point **29.** (*a*) Zinc (*b*) Sodium (*c*) Manganese (*d*) Mercury **30.** (*a*) rusting (*b*) air ; water (*c*) galvanisation (*d*) tin ; chromium (e) green (*f*) zinc (*g*) tin (*h*) carbon (*i*) amalgam (*j*) less (*k*) gangue **33.** Electrolytic reduction : Aluminium and Sodium ; Reduction with carbon : Zinc, Iron and Tin ; Reduction with aluminium : Manganese **34.** (*a*) Copper does not corrode easily in the presence of water but steel rusts in the presence of water. **36.** (*a*) To prevent its rusting **37.** (*a*) Electrolytic reduction **38.** (*a*) Copper and Silver **46.** (*b*) Gold and Platinum (*c*) Rusting **49.** (*a*) Aluminium and Zinc **53.** (*a*) Aluminium **55.** (*a*) Aluminium oxide ; Oxygen (*b*) Sodium chloride ; Chlorine (*c*) Carbon dioxide (*d*) Sulphur dioxide **56.** (*a*) Gold and Platinum (*b*) Sodium and Magnesium **57.** (*e*) Potassium, Calcium and Magnesium **58.** (*d*) Cathode (Negative electrode) (*e*) Oxygen gas ; At anode (Positive electrode) **59.** (*e*) Aluminium **62.** (*d*) **63.** (*c*) **64.** (*c*) **65.** (*d*) **66.** (*c*) **67.** (*a*) **68.** (*c*) **69.** (*d*) **70.** (*b*) **71.** (*c*) **72.** (*d*) **73.** (*b*) **74.** (*d*) **75.** (*d*) **76.** (*c*) **77.** (*c*) **78.** (*a*) **79.** (*d*) **80.** (*d*) **81.** (*c*) **82.** (*d*) **83.** (*d*) **84.** (*c*) **85.** (*c*) **86.** (*b*) **87.** (*c*) **88.** (*b*) **89.** (*b*) **90.** (*c*) **91.** (*a*) **92.** (*a*) A is sodium and B is chlorine (*b*) C is sodium chloride and D is sodium hydroxide (*c*) It will turn litmus solution blue ; Because it is a base (*d*) Gas (*e*) EB_3 **93.** (*a*) Mercury, Hg (*b*) Cinnabar, HgS (*c*) See page 180 (*d*) Thermometer (*e*) No; Because it is less reactive than copper **94.** (*a*) A is iron (III) oxide and B is aluminium (*b*) C is molten iron metal and D is aluminium oxide (c) See page 179 (*d*) Thermite reaction ; Welding of broken pieces of heavy iron objects like railway tracks, etc. (*e*) Displacement reactions and Oxidation-reduction reactions **95.** (*a*) Positively charged carbon electrode (Anode) (*b*) Because oxygen produced during the electrolysis of molten aluminium oxide reacts gradually with the carbon of carbon anode to form carbon dioxide gas **96.** (*a*) X is aluminium and Y is zinc (*b*) X is more reactive (*c*) Bauxite ; $Al_2O_3.2H_2O$ (*d*) Calamine, $ZnCO_3$ (*e*) Alloy of metal X : Duralium ; Alloy of metal Y : Brass **97.** (*a*) Silver (*b*) Copper (c) Silver sulphide (*d*) Basic copper carbonate (*e*) Dilute acid solution ; The acid solution dissolves green coloured basic copper carbonate present on the corroded copper object **98.** Metal P is zinc ; Metal Q is tin ; Metal R is lead ; Metal S is iron ; Metal P (zinc) is used to form a thin layer on metal S (iron) to prevent its corrosion ; Metal Q (tin) is used for electroplating tiffin boxes made of metal S (iron) ; Metal R (lead) is used in making car batteries ; Metals Q (tin) and R (lead) form an alloy called solder **99.** (*a*) Manganese (*b*) Manganese dioxide (*c*) Aluminium (*d*) Chromium **100.** Metal X is gold and Metal Y is silver ; The colour of metal X (gold) is yellow.