Metals and Non-metals

Physical Properties of Metals and Non-Metals

Do you know how many elements are there in our periodic table?

There are 118 elements in the modern periodic table. These elements can be broadly classified as metals and non-metals depending on their properties.

Elements that lose electrons to form compounds are called **metals** whereas elements that gain electrons to form compounds are called **non-metals**. Elements such as Si, Ge, As, Sb and Te show the characteristic properties of both metals and non-metals. They are called **semi-metals** or **metalloids**. Here, we will discuss metals and non-metals along with their physical properties in detail.

Metals

These elements are electropositive and contain less than or equal to three electrons in their valence shell. Metals such as aluminium, copper, and iron are widely used around us. Metals are used for the construction of bridges, automobiles, airplanes, ships, trains, etc. We will now discuss the physical properties of metals.

Physical properties of metals:

1. Metallic Lustre: The surface of most metals is shiny. The lustre associated with metals is known as **metallic lustre**. For example, iron, copper, gold, and silver are very shiny. Metals such as gold and silver are very lustrous. Therefore, they are used for making jewellery.

> Silver is used for making mirrors because of its excellent shine and reflective nature.

Do you know that metals like gold, silver, platinum, paladium and rhodium are known as **noble metals.** They occur in the elemental state in nature.

Some metals do not look very lustrous. This is because they either lose their lustre or their lustre gets reduced when exposed to air for a long time. This happens due to the formation of a layer of oxide, carbonate, and sulphide on their surface. If a metal surface is rubbed with sand paper, then this layer gets removed and the shiny surface of the metal can be seen. The layer formed in some cases is stable and sticks on the surface of the metal, but in other cases, it is unstable and falls off (as in the case of rusting of iron).

2. Hardness: Metals are generally hard in nature. However, this hardness varies from metal to metal. Most metals such as iron, aluminium, etc. are very hard and cannot be cut with a knife whereas some metals such as sodium and potassium are very soft and can be cut using a knife.

3. Malleability: Metals are malleable. Most metals such as iron, copper, silver, and gold can be hammered without breaking to form thin sheets. Aluminium, and silver are highly malleable metals and are often used for making foils, which are extensively used in the decoration of sweets, packing of food items, etc.

4. Ductility: Most metals are ductile, which means that they can be drawn into thin wires without breaking. For example, iron, copper, silver, and gold can be drawn into thin wires without breaking. For this reason, copper and aluminium are extensively used for making electrical wires.

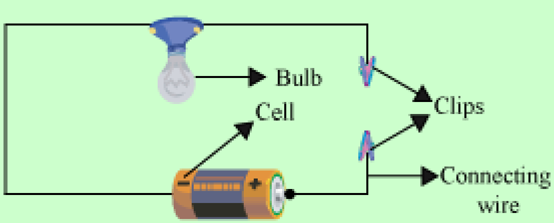
Gold and silver are the most malleable and ductile metals. Hence, they are extensively used in jewellery.

5. Conduction of heat: Metals are generally good conductors of heat. This means that if one end of a metal rod is heated for some time, then the entire rod becomes hot. For example, aluminium, copper, and silver are good conductors of heat. Hence, copper and aluminium are generally used for making vessels. The following activity can be performed to explain that metals can conduct heat.

6. Conduction of electricity: Metals are good conductors of electricity i.e., they allow an electric current to pass through them easily. Silver, copper, and aluminium are the best conductors of electricity. For this reason, most electric wires are made of copper and aluminium. However, using silver for making electric wires is not cost effective. The following activity can be performed to explain that metals can conduct electricity.

Activity:

Take two electric wires and attach two clips to each wire (as shown in the given figure). Then, take a bulb fitted in a holder and connect it to a battery with the help of electric wires. Now, take pieces of iron, copper, and aluminium and place them one by one between the clips.



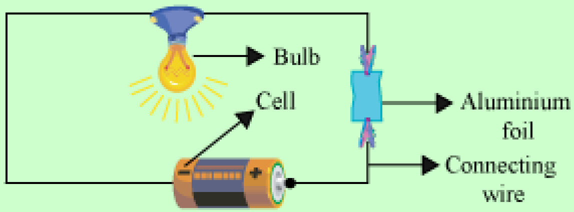


Figure 2: Conduction of electricity in metals

It will be observed that the bulb glows every time when the metal pieces are used to connect the two clips. This shows that metals are good conductors of electricity.

7. **Melting and boiling points:** Melting and boiling points of metals are usually high.

8. **Physical state:** All metals exist as solids at room temperature except mercury, which exists as a liquid.

9. **Sonority:** Metals such as iron and copper produce a sound on being struck. Hence, metals are said to be sonorous.

Non-metals

Many elements in the periodic table do not behave like metals. These elements are known as **non-metals**. These elements gain electrons to form compounds. These are electronegative and contain more than three electrons in their valence shell. Carbon, sulphur, iodine, oxygen, etc. are some examples of non-metals. Non-metals exist in all three physical states i.e., as solids, liquids, and gases. Bromine is the only non-metal, which exists as a liquid.

Physical properties of non-metals:

1. **Lustre:** Non-metals do not have a shiny surface. However, iodine is an exception, which has a very shiny surface.

2. **Hardness:** Non-metals generally exist as solids, liquids, or gases. Non-metals that exist in a solid state are very soft. For example, sulphur, which exists in solid state, is quite soft. Similarly, carbon, in the form of graphite, is quite soft. However, diamond, another allotrope of carbon, is very hard. It is in fact the hardest known natural substance.

3. **Malleability and ductility:** Non-metals that exist in solid states are not very strong. They are brittle and break when pressure is applied on them. Therefore, non-metals are neither malleable nor ductile.

4. Conduction of heat and electricity: Non-metals are poor conductors of heat and electricity. Examples include sulphur and phosphorus. However, there is an exception. Graphite, an allotrope of carbon, is a good conductor of electricity.

5. Physical state: Non-metals exist in all three physical states at room temperature. Non-metals such as carbon, sulphur, and phosphorus exist in solid states while oxygen, chlorine, and nitrogen exist in gaseous states. Bromine is the only non-metal that exists in a liquid state.

6. Melting and boiling points: Melting and boiling points of non-metals are quite low. For example, the melting point of phosphorus is 44.2°C. However, diamond, an allotrope of carbon, is the only non-metallic substance that has a very high melting and boiling point. The melting point of diamond is more than 3500°C.

7. Sonority: Non-metals are not sonorous.

The given table summarizes the properties of metals and non-metals.

Metals	**Non-metals**
Metals are very hard and strong.	Solid non-metals are soft and can be easily broken.
Metals have a shiny lustre.	Non-metals are not shiny and have a dull appearance.
Metals are sonorous.	Non-metals are not sonorous.
Metals are malleable and ductile.	Non-metals are neither malleable nor ductile.
Metals are good conductors of heat and electricity.	Non-metals are poor conductors of heat and electricity.

Reactions of Metals and Non-Metals with Oxygen and Water

Metals such as aluminium, copper, and iron are widely used around us. Metals are used for the construction of bridges, automobiles, airplanes, ships, trains etc.

We have earlier studied about the physical properties of metals. Now, let us try to learn about their chemical properties. Here, we will study about the reaction of metal with oxygen, water, and acids.

You must have observed that when a piece of iron is kept in the open for some time, it gets covered with a brownish substance. This brownish substance is called **rust** and the process is called **rusting**. Rust is formed when iron reacts with oxygen (present in air) to form iron oxide. Also, a ribbon of magnesium burns in air to form magnesium oxide. These reactions represent reactions of metals with oxygen. Hence, metals react with oxygen to produce metals oxides.

Metals react with oxygen to produce metal oxides which are basic in nature. These oxides thus turn red litmus paper blue, but have no effect on blue litmus paper.

Sulphur (S) is a non-metal. It reacts with oxygen to produce sulphur dioxide (SO_2), which is an acidic oxide. Sulphur dioxide then reacts with water to produce sulphurous acid (H_2SO_3), which changes blue litmus to red. The chemical equations involved in the reaction can be represented as:

$$\underset{\text{Sulphur}}{S} + \underset{\text{Oxygen}}{O_2} \rightarrow \underset{\text{Sulphur dioxide}}{SO_2}$$

Non-metals react with oxygen to produce their oxides, which are generally acidic in nature.

We will now study the reaction of metals and non-metals with water.

While some metals react very vigorously with water, others react very slowly. However, there are some metals which do not react with water at all. For example, sodium metal reacts vigorously with water and iron reacts slowly with water.

Metals react with water to produce hydrogen gas and metal hydroxides. These metal hydroxides are basic in nature. However, non-metals usually do not react with water.

Do You Know:

- Sodium and potassium are very reactive metals. They react vigorously with oxygen and water to produce a lot of heat. Hence, to prevent their reaction with air and water, they are stored under kerosene.

- Non-metals react very vigorously with air, but generally do not react with water. Phosphorus is a very reactive non-metal, which catches fire when exposed to air. Hence, phosphorus is stored under water to prevent contact between phosphorus and air.

1. **Reaction of metals with oxygen**

   On heating, magnesium burns with a dazzling white flame to form magnesium oxide. Similarly, when aluminium is heated, it reacts with oxygen present in the air to form aluminium oxide.

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

Almost all metals combine with oxygen to form metal oxides. The general reaction for the process is:

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

All metals are not equally reactive. Therefore, the reactivity of metals with oxygen also varies. Some metals such as sodium react with oxygen at room temperature. Metals such as magnesium do not react with oxygen at room temperature and require heating. On the other hand, metals such as zinc do not react with oxygen easily and require very strong heating. Silver and gold do not react with oxygen even at high temperatures.

All metal oxides are basic in nature and turn red litmus paper blue. These basic oxides react with acids to form salt and water. However, the oxides of aluminium and zinc show the properties of both acids and bases. **Chemicals that show both acidic and basic properties are said to be amphoteric in nature**. Hence, aluminium oxide and zinc oxide are amphoteric oxides. They react with both acids and bases to give their respective salts and water.

Almost all metal oxides are insoluble in water. However, the oxides of sodium and potassium dissolve in water to form hydroxides.

$$\underset{\text{Sodium oxide}}{Na_2O(s)} + \underset{\text{Water}}{H_2O(l)} \rightarrow \underset{\text{Sodium hydroxide}}{2\,NaOH(aq)}$$

$$\underset{\text{Potassium oxide}}{K_2O(s)} + \underset{\text{Water}}{H_2O(l)} \rightarrow \underset{\text{Potassium hydroxide}}{2\,KOH(aq)}$$

2.

Reaction of metals with water

Do you know that sodium reacts explosively with cold water? The reaction results in the formation of their respective hydroxides and hydrogen gas. The reaction is so violent and exothermic that the evolved hydrogen catches fire. These metals give hydroxides with water as their oxides are soluble in water.

$$\underset{\text{Sodium}}{2Na(s)} + \underset{\text{Water}}{2H_2O(l)} \rightarrow \underset{\text{Sodium hydroxide}}{2NaOH(aq)} + \underset{\text{Hydrogen}}{H_2(g)} + \text{Heat}$$

On the other hand, metals such as iron do not react with cold water or hot water. However, they react with steam to give their respective oxides and hydrogen gas.

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

Thus, metals react with water to form metal oxides and hydrogen gas. Some metal oxides are soluble in water. These metal oxides form hydroxides by reacting with one or more water molecules. The general reaction for the process is given as:

$$\text{Metal} + \text{Water} \rightarrow \text{Metal oxide} + \text{Hydrogen}$$
$$\text{Metal oxide} + \text{Water} \rightarrow \text{Metal hydroxide} \quad (\text{if metal oxide is soluble in water})$$

The vigour with which a metal reacts with water differs from metal to metal. Some metals react with cold water, others with hot water, while some react only with steam. There are also metals that do not even react with steam. For example, silver and gold do not react with water at all.

Reaction of non-metals with hydrogen.

Non-metals react with hydrogen under specific conditions to form their corresponding compounds containing hydrogen. Few examples are given below:

$O_2 + 2H_2 \rightarrow 2H_2O$ (Water)

$S + H_2 \rightarrow H_2S$ (Hydrogen sulphide)

$N_2 + 3H_2 \rightarrow 2NH_3$ (Ammonia)

$Cl_2 + H_2 \rightarrow 2HCl$ (Hydrogen chloride)

Unlike metals, non-metals do not react with water or dilute acids.

Reaction of Metals and Non-Metals with Acids and Bases

You know that the substances which turn blue litmus paper to red are called **acids,** and the substances which turn red litmus paper to blue are called **bases**. **Do you how these substances react with metals and non-metals?**

Let us study how metals and non-metals react with acids.

Therefore, it can be concluded that metals react with acids to release hydrogen gas, which burns with a ‘pop’ sound. On the other hand, non-metals do not react with acids.

Some Interesting Facts:

- Hydrogen gas is colourless and odourless. It has no effect on moist litmus paper. It burns with a characteristic 'pop' sound when a flame is introduced.

- Copper is a less reactive metal. It does not react with dilute hydrochloric acid, even on heating.

Thus, metals react with bases to produce hydrogen gas. However, not all the metals react with bases to produce hydrogen gas. The reactions of non-metals with bases are complex.

Metals react with hydrochloric acid in the similar fashion as they do with sulphuric acid. Sodium reacts very vigorously with hydrochloric acid to form a salt, and hydrogen gas is evolved in the reaction.

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

Magnesium reacts vigorously with hydrochloric acid, but not as vigorously as sodium and potassium.

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

Zinc and iron also react with dilute hydrochloric acid to give zinc chloride and iron (II) chloride respectively. These reactions are comparatively less vigorous than the reaction of hydrochloric acid with aluminium metal.

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

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

Thus, it can be concluded that metals react with acids to give a salt and hydrogen gas. The general equation for the process can be represented as:

$$\text{Metal} + \text{acid} \rightarrow \text{Salt} + \text{Hydrogen}$$

However, all metals do not react with dilute hydrochloric and sulphuric acids. Also, hydrogen gas is not evolved when a metal reacts with nitric acid. This is because nitric acid acts as an oxidizing agent and oxidizes hydrogen gas produced in the reaction to form water. At the same time, nitric acid itself gets reduced to form nitrogen oxides such as nitrous oxide (N_2O), nitric oxide (NO), and nitrogen dioxide (NO_2). However, there are some metals such as magnesium, which react with very dilute nitric acid to evolve hydrogen gas.

$$\underset{\text{Magnesium}}{Mg(s)} + \underset{\text{Nitric acid}}{2HNO_3(aq)} \rightarrow \underset{\text{Magnesium nitrate}}{Mg(NO_3)_2(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

Metals such as gold and silver, which are very less reactive, do not react with acids. The only acid that dissolves gold is *aqua regia*. *Aqua regia* is the Latin name for 'holy water' or 'royal water'. It is called so because it is the only liquid that dissolves gold. It is prepared by mixing three parts of concentrated hydrochloric acid and one part of concentrated nitric acid. It is a highly corrosive and fuming solution having yellow or red colour. It can also dissolve platinum metal.

- **Reaction of metals with bases**

When metals react with base they forms hydrogen.

Corrosion: Causes and Prevention

You must have observed that when metals such as iron, silver, and copper are exposed to air for some time, they lose their shine. For example, iron, when exposed to moist air for a long period of time, acquires a coating of a brown-flaky substance.

This is because metals react with moisture and the different gases present in the air. **The reaction of metals with moisture and gases present in the air is known as corrosion.** Rusting of iron is the most common example of corrosion.

DO YOU KNOW?

Rust is a general term given to iron oxides, which are formed when iron reacts with oxygen in the presence of moisture. Rust primarily consists of hydrated iron (III) oxides, $Fe_2O_3 .nH_2O$. The number of water molecules in rust is variable. Hence, they are represented by *n.*

Other examples of corrosion:

1. You must have observed that ornaments made of silver lose their shine after some time. This is because silver reacts with sulphur present in the air to form silver sulphide, which forms a layer over its surface.

2. Copper reacts with carbon dioxide to form copper carbonate, which is greenish in appearance. This is the reason why a copper article loses its shiny brown surface when exposed to air.

DO YOU KNOW?

Corrosion of aluminium metal is extremely slow. This is because of aluminium oxide, which is formed when aluminium reacts with oxygen, is very stable and forms a protective coating or layer on the surface of the metal. This prevents the oxidation of the remaining metal.

Corrosion can drastically reduce the quality and strength of metals. The higher a metal lies in the reactivity series, more readily it is corroded. Here, we will study about the rusting of iron and the conditions necessary for the same. Let us find out about the conditions necessary for rusting of iron to take place with the help of the following activity.

Therefore, we can say that both air and water are required for rusting to take place.

Effect of Corrosion on Other Metals based on the Reactivity Series:

1. Reactive alkali metals react with oxygen, water and carbon dioxide present in air to form oxide, hydroxide and carbonate, respectively. Hence, they are kept immersed in kerosene oil to prevent the corrosion.
2. Aluminium and magnesium when exposed to air form a white layer of the oxide on their surface.
3. Iron forms hydrated ferric oxide (rust) on exposure to moisture in the air.
4. Lead forms a white deposit of lead hydroxide and lead carbonate called basic lead carbonate on coming in contact with moist air.
5. Copper forms a green deposit of copper hydroxide and copper carbonate called basic copper carbonate on exposure to moist air.
6. Siver forms a black coating of silver sulphide on its surface on coming in contact with hydrogen sulphide present in the air. This phenomenon is known as tarnishing of silver.

Factors Affecting Corrosion

Besides oxygen and moisture in the air, there are other factors that enhance the corrosion of metals. These are:

- Reactive nature of metal: Highly reactive metals corrode easily.
- Presence of dissolved salts: They act as electrolyte and increase the rate of corrosion.
- Presence of pollutants: They increase the rate of corrosion.
- Presence of less reactive metal: If a less reactive metal is present, it will make the more reactive metal susceptible to corrosion.

Every year our world suffers a huge monetary loss owing to the process of rusting, which causes harm to articles made of iron. Attempts were made to prevent rusting. Here are some ways that can prevent rusting or corrosion.

We now know that both air and water are required for rusting to take place. Thus, rusting can be prevented by cutting off the contact of iron articles with air or water or both. There are different methods by which rusting can be checked:

- Rusting can be prevented by electroplating, painting, oiling, and greasing of iron articles. In fact, paints and grease should be applied regularly to prevent rusting.
- Rusting can also be prevented by applying a layer of a metal such as chromium or zinc on the surface of iron articles. **The process of depositing zinc on iron is called galvanization.**
- Rusting can also be prevented by connecting the iron object with a more reactive metal like zinc with the help of a wire. **The process of connecting iron with a more reactive metal through a wire is called cathode protection**.
- Alloying can also be used to prevent rusting or corrosion.

Do you know what alloys are?

An alloy is a homogeneous mixture of two or more elements, at least one of which is a metal. An alloy of a metal is made by first melting the metal and then, adding and dissolving the element with which it is to be alloyed. This is done in a molten state so that an even distribution of elements can take place. Usually, the resulting substance has properties different from those of its components.

Alloy	**Components**	**Properties**	**Uses**
Stainless steel	Iron, nickel and chromium	Does not get affected by the action of air, water and alkali.	In preparation of utensils, blades and surgical instruments.

Brass	Copper and zinc	Malleable, strong, corrosion resistant and can be easily shaped.	In preparation of cooking utensils, parts of machines and instruments.
Bronze	Copper and tin	Stronger and more corrosion resistant.	In preparation of statues, coins and medals.

Do you know?

Pure gold is known as 24 carat gold. In India, the gold that is generally used to make ornaments is 22 parts of pure gold alloyed with 2 parts of either silver or copper.

Do you know that alloying is a good method for improving the properties of metals?

Properties of metals can be improved by combining them with other elements i.e., by alloying. Alloying can also be used to prevent rusting. Pure iron is not very hard and stretches when heated. However, when it is mixed with a small amount of carbon, it becomes very hard. This is known as steel. Even though steel is hard, it does rust. Stainless steel is obtained when nickel and chromium are added to iron. **Stainless steel contains iron as the primary constituent, but it does not rust at all.** Thus, by adding different elements, the properties of iron can be changed.

The iron pillar near the Qutub Minar in New Delhi was made around 400 B.C. It is 8 m tall and weighs around 6 tonnes (6000 kg). The workers who made it knew that pure iron would rust after some time and devised a method to prevent the pillar from rusting. They painted the surface of the pillar using a mixture of salts, followed by heating and quenching (rapid cooling). This finishing treatment resulted in the formation of a thin layer of magnetic oxide (Fe_3O_4) on the surface of the pillar and prevented the iron present in the pillar from rusting.

Even though corrosion and rusting causes much damage, but sometimes this phenomenon has an **advantage**. Let us understand with the help of an example. Aluminium and zinc articles when exposed to air form a white deposit of their respective oxides on their surface. These oxides stick to the surface of the metal and are impervious in nature. So in a way, this oxide prevents the next layer of metal from getting corroded. This is the reason why objects made from aluminium and zinc do not corrode easily.

Reaction of Metals with Solution of Salts of other Metals; And The Reactivity Series

Categorisation of Metals

Based on the positioning of metals in the periodic table, their characteristics are given below:

	Alkali Metals	Alkaline Metals
Elements	Li, Na, K	Mg, Ca, Sr
Position	Placed in IA group	Placed in IIA group
Occurence	Do not occur in free state	Do not occur in free state
Nature	Soft Low melting and boiling point	Hard Greyish white in colour
Bonding	Salts of alkali metals form ionic compounds (except for some lithium salts)	Salts of alkaline metals form ionic compounds (except for beryllium)
Action of air	React rapidly Reactivity increases down the group	Less reactive Reactivity increases down the group
Action of water	$2M + 2H_2O \rightarrow\rightarrow 2MOH + H_2$	$M + 2H_2O \rightarrow\rightarrow M(OH)_2 + H_2$
Action of acids	$2 M + 2 HCl \dashrightarrow 2 MCl + H_2$	$M + 2HCl \rightarrow MCl_2 + H_2$
Ionisation energy	Low ionisation energy	Low ionisation energy (but higher than alkali metals)
Colour of the flame	Impart characteristics colour to the flame Crimson red – Lithium (Li) Golden yellow – Sodium (Na) Pale violet – Potassium (K)	Impart colour to the flame Calcium – Brick red Strontium – Crimson Barium – Apple green Mg and Be do not impart colour because the electrons are too strongly bound to be excited.
Obtained	By electrolysis of their molten salts	By electrolysis of their molten salts

Some other categories of metals:

Transition Metals:

- Elements like Fe, Zn etc. are transition metals. They are placed in the middle of the periodic table.
- Have high melting and boiling point
- Good conductors of heat and electricity
- Show variable valencies

Inner transition metals:

- Elements like La, Ce are inner transition metals. They are placed at the bottom of the periodic table.
- Heavy metals with high melting and boiling point
- Good conductor of heat and electricity
- Show variable valencies

In the reaction of metals with air, water, and acids, we observed that some metals react very vigorously, some others react rather slowly, and some do not react at all.

What can you conclude from the given information? Are all elements equally reactive?

As different metals react with the same chemicals in different ways, the reactivity of metals cannot be similar.

If the reactivity of metals is different, then how can we determine the reactivity of two metals?

Displacement reactions help us for this. Actually, some metals are more reactive than others. Metals that are more reactive can displace the less reactive metals from their salts in a solution or molten form. The general equation for such reactions is given as:

$$\text{Metal A} + \text{Salt solution of metal B} \rightarrow \text{Salt solution of metal A} + \text{Metal B}$$

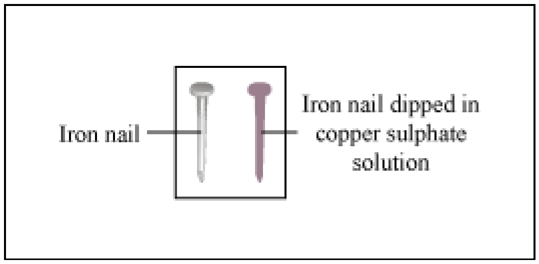
Such reactions are called **displacement reactions.** In displacement reactions, a more reactive metal replaces a less reactive metal from the latter's salt.

For example, iron can replace copper from copper sulphate solution, but copper cannot replace iron from iron (II) sulphate solution.

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

$$\underset{\text{Copper}}{Cu(s)} + \underset{\text{Iron (II) sulphate}}{FeSO_4(aq)} \rightarrow \text{No reaction}$$

When an iron nail is kept in a copper sulphate solution, the intensity of copper sulphate solution decreases, and the iron nail gets covered with copper.

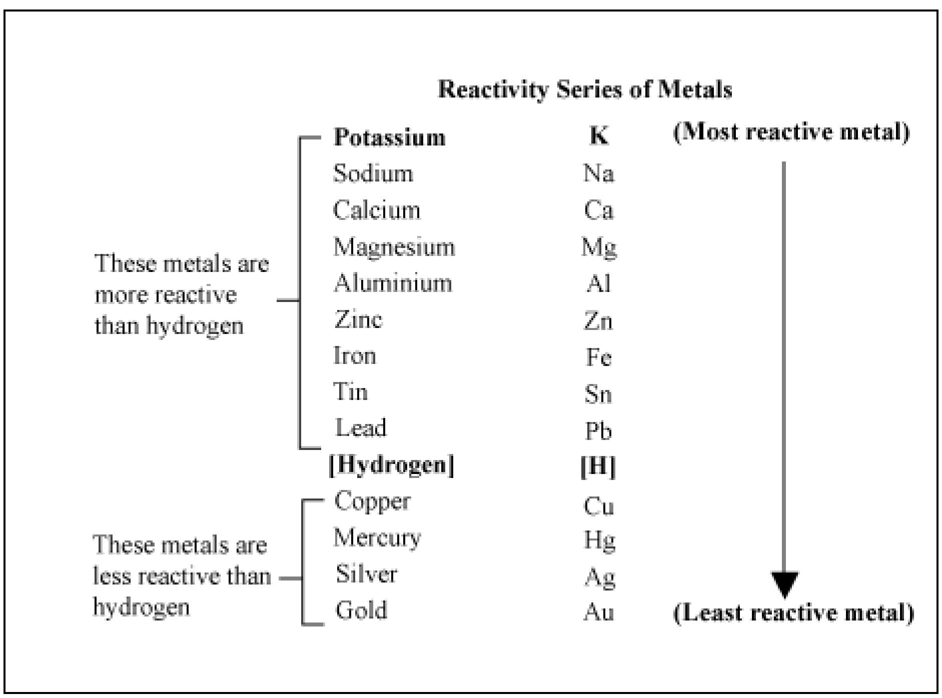


This means that iron is more reactive than copper as it can replace copper from copper sulphate solution.

The reactivity of metals can be determined by observing their reactions with salt solutions of other metals. When the reactivity of a metal is determined, it can be arranged in an increasing or decreasing order of their reactivities.

The series in which various metals are arranged in the order of their decreasing reactivity is called a Reactivity series.

This series is prepared by performing displacement reactions between various metals and their salt solutions. The reactivity series is given as follows:



Reactivity series of metals

In the reactivity series, metals present above a particular metal are more reactive than that metal, while the metals present below the particular metal are less reactive than it.

We know that all metals lose electrons to form positive ions. The tendency to lose an electron can be related to the reactivity of metals. If a metal can lose electrons easily, then it is very reactive. On the other hand, if the tendency of a metal to lose electrons is less, then it is less reactive.

In general, we say that the metals present above hydrogen are more reactive than it, and displace it from acids to liberate hydrogen gas. However, the metals present below hydrogen are less reactive than it, and cannot displace it from acids to liberate hydrogen gas.

Metals such as sodium and potassium (that lie above hydrogen) readily react with dilute acids to evolve hydrogen gas, whereas metals such as copper, gold, and silver (that lie below hydrogen) do not react with dilute acids.

Main Features of Reactivity Series

- Metals are arranged in the decreasing order of their electropositive character.
- Metals at the top have greater reducing power. This power decreases on moving down the series.
- Metals at the top show greater tendency to get oxidised.
- Metals above hydrogen in the reactivity series liberate hydrogen gas from mineral acids.
- Metals at the top displace metals lower in the series from the aqueous solution of their salts.
- Metal oxides above Al, cannot be reduced by common reducing agents, the reverse is true for metal oxides below Al.

Let us now see the action of heat on some metallic compounds like oxides, hydroxide, carbonates and nitrates.

Oxides:

Metal	**K, Na, Ca, Mg, Al**	**Zn**	**Fe, Pb, Cu**	**Hg, Ag**
Action of Heat on metal oxide	Stable to heat Reduced by electrolysis	Reduced by coke only	Reduced by C, CO, H_2, NH_3	Decompose on heating to give metal and oxygen

Hydroxides:

Metal	**K, Na**	**Ca, Mg, Al, Zn, Fe, Pb, Cu**	**Hg, Ag**

Action of Heat on metal oxide	Stable to heat Soluble in water	Decompose on heating to give metal oxide and water vapour	Decompose on heating to give metal, oxygen and water vapour

Carbonates:

Metal	**K, Na**	**Ca, Mg, Al, Zn, Fe, Pb, Cu**	**Hg, Ag**
Action of Heat on metal oxide	Stable to heat Soluble in water	Decompose on heating to give metal oxide and carbon dioxide	Decompose on heating to give metal, oxygen and carbon dioxide

Nitrates:

Metal	**K, Na**	**Ca, Mg, Al, Zn, Fe, Pb, Cu**	**Hg, Ag**
Action of Heat on metal oxide	Decompose on heating to give metal nitrite and oxygen	Decompose on heating to give metal oxide, nitrogen dioxide and oxygen	Decompose on heating to give metal, nitrogen dioxide and oxygen

Elementary Idea of Chemical Bonding

Chemical Bonding

Elements are rarely capable of free existence. In a compound, atoms of different elements are held together by bonds. The types of bonds present in a compound are largely responsible for its physical and chemical properties. The different bonds can be classified as **strong** and **weak**.

Why do elements undergo bond formation?

Elements are made of atoms, which comprise of protons, electrons, and neutrons. The protons and the neutrons reside in the nucleus and the electrons revolve around in definite paths called **orbits**. The electrons present in the last shell are called valence electrons. These electrons are responsible for all the chemical reactions of that element.

Every element has a tendency to attain a stable outer octet. To do so, it either gains or loses or shares its electrons; and in this process, it forms the bonds.

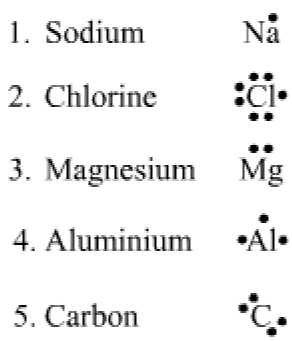
Types of strong bonds:

- Ionic or electrovalent bond
- Covalent bond
- Metallic bond

Types of weak bonds:

- Bonds formed due to van der Waal's interaction
- Hydrogen bond

This representation of elements with valence electrons as dots around elements is referred to as **Electron Dot structures** for elements. The electron dot structure of some of the elements are:



Formation of Ionic Compounds and Their Properties

We know that common salt is an important dietary mineral essential for animal life. Common salt is chemically known as sodium chloride. The chemical formula of sodium chloride is NaCl. It suggests that it is made up of sodium, which is a reactive metal, and chlorine, which is a non-metal.

Do you know that sodium chloride does not exist as molecules, but aggregates as oppositely charged ions?

An ion is a charged species, which can be negatively charged or positively charged. A negatively charged species is called an 'anion' and a positively charged species is called a 'cation'.

Sodium chloride (NaCl) is formed by the combination of sodium (Na^+) and chloride (Cl^-) ions. Sodium and chloride ions are oppositely charged. Hence, they are held by a strong electrostatic force of attraction in sodium chloride compound. **But why do they react or combine with each other?** This can be explained by considering the formation of sodium chloride.

This representation of elements with valence electrons as dots around elements is referred to as **Electron Dot structures** for elements.

Do you know what type of a compound sodium chloride is? Sodium chloride is an ionic compound.

Ionic compounds:

These are compounds that are formed by the transfer of electrons. In other words, these are compounds that are made up of ions.

The bonding in such compounds is called **ionic bonding or electrovalent bonding**. This type of bonding is also known as **electrostatic bonding** as the forces that hold the ions together are electrostatic in nature. The transfer of electrons always takes place from a metal to a non-metal. Thus, metals and non-metals combine with each other to attain a noble gas configuration.

We know that inert (noble) gases are very stable and almost unreactive. This is because of their stable electronic configuration in which their valence shell is complete. Hence, they do not take part in the formation of ionic compounds. The given table lists some elements with their electronic configurations.

S.No.	Type of element	Element		Symbol	Atomic number	Electronic configuration K L M N	Number of valence electrons
1.	**Noble gases**	1. 2. 3.	Helium Neon Argon	He Ne Ar	2 10 18	2 2, 8 2, 8, 8	2 8 8
2.	**Metals**	1. 2. 3. 4. 5.	Sodium Potassium Magnesium Calcium Aluminium	Na K Mg Ca Al	11 19 12 20 13	2, 8, 1 2, 8, 8, 1 2, 8, 2 2, 8, 8, 2 2, 8, 3	1 1 2 2 3
3.	**Non-metals**	1. 2. 3. 4. 5. 6.	Nitrogen Phosphorus Oxygen Sulphur Fluorine Chlorine	N P O S F Cl	7 15 8 16 9 17	2, 5 2, 8, 5 2, 6 2, 8, 6 2, 7 2, 8, 7	5 5 6 6 7 7

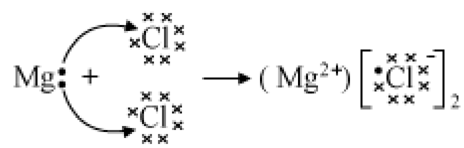
Let us now see the formation of magnesium chloride, which is also an ionic compound.

The atomic number of magnesium is 12. Thus, its electronic configuration is 2, 8, 2. Since it contains two more electrons than a stable noble gas configuration, it loses these two electrons to form Mg^{2+}. On the other hand, the atomic number of chlorine is 17. Thus, its electronic configuration is 2, 8, 7. It requires one more electron to complete its octet. For this, two chlorine atoms accept two electrons that were lost by Mg atom to form two chloride (Cl^-) ions. The chemical equations involved in the process are given below:

$$\underset{\substack{\text{Magnesium} \\ (2,\ 8,\ 2)}}{Mg} \rightarrow \underset{\substack{\text{Magnesium ion} \\ (2,\ 8)}}{Mg^{2+}} + \underset{\text{Electron}}{2e^-}$$

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

The reaction between magnesium and chlorine can be represented as follows:



On the similar basis, the formation of sodium chloride (NaCl) and calcium oxide (CaO) is depicted in the table below:

Compound	Formation
Sodium chloride (NaCl)	$\underset{\text{Sodium oxide}}{Na_2O(s)} + \underset{\text{Water}}{H_2O(l)} \rightarrow \underset{\text{Sodium hydroxide}}{2\ NaOH(aq)}$ $\underset{\text{Potassium oxide}}{K_2O(s)} + \underset{\text{Water}}{H_2O(l)} \rightarrow \underset{\text{Potassium hydroxide}}{2\ KOH(aq)}$
Calcium oxide (CaO)	$\underset{(2,8,8,2)}{Ca} \longrightarrow \underset{(2,8,8)}{Ca^{2+}} + 2e^-$ $\underset{(2,6)}{O} + 2e^- \longrightarrow \underset{(2,8)}{O^{2-}}$ $Ca^{2+} + O^{2-} \longrightarrow [Ca^{2+}][O^{2-}]$

Potassium oxide (K_2O) is also an ionic compound. It is made of two potassium atoms and one oxygen atom.

Can you draw the Electron Dot structure of potassium and oxygen atoms? Can you show the formation of potassium oxide?

Let us now try to find out the properties of ionic compounds by performing the following activities.

1) Take samples of sodium chloride, potassium iodide, and barium chloride and observe their physical state.

2) After that, take a small amount of a sample on a metal spatula and heat it directly on a flame. Observe what happens to the sample.

3) Now, try to dissolve each sample in water, petrol, and kerosene and observe the solubility of compounds.

4) Now, take a container and fill it with distilled water. Take two electrodes and place them in water. Then, connect the electrodes to a bulb and a battery through electric wires (as shown in **figure 1)**. When the switch is closed, the bulb will not glow as distilled water does not conduct electricity. Now, instead of distilled water, take a solution of an ionic compound and observe.

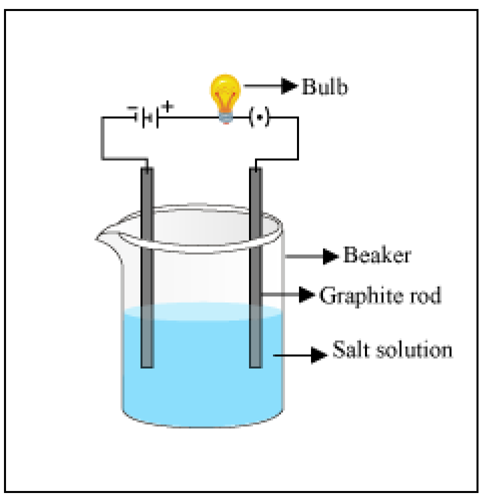


Figure 1: Conductivity of salt

We will observe that

- all compounds are solids
- all have high melting and boiling points
- all samples are soluble in water but insoluble in kerosene and petrol
- the solution of all samples can conduct electricity

When the switch is closed, the bulb starts glowing. This shows that solutions of ionic compounds conduct electricity.

Hence, we can summarize the properties of ionic compounds as follows:

Ionic compounds are hard and brittle crystalline solids: The electrostatic force holding the ions present in ionic compounds are very strong. Therefore, these compounds are quite hard, as they are made up of small crystals.

Ionic compounds have high melting and boiling points: A lot of energy is required to overcome the strong electrostatic force of attraction, which holds the ions present in ionic compounds together. Thus, these compounds have high melting and boiling points.

Table 1: Melting and boiling points of some ionic compounds

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

Ionic compounds dissolve only in polar solvents: Ionic compounds are polar in nature due to the presence of opposite charges. Therefore, these compounds dissolve only in polar solvents such as water. These compounds are insoluble in organic solvents such as kerosene, alcohol, and petrol.

Ionic compounds conduct electricity in a solution or molten state: Ionic compounds consist of small ions, which can conduct electricity.

Extraction and Refining of Metals

Elements on earth are found in different forms from the different parts of the earth:

Lithosphere: This part of earth is made up of sand, clay, stone and elements such as aluminium, copper, iron, calcium, sodium etc. which are found in the form of sulphides or oxides.

Hydrosphere: This part of the earth includes water of seas, lakes and ice of polar regions. In this part, many non-metals and metals are obtained in combined forms such as, chlorine, flourine, sodium, potassium, magnesium and calcium.

Atmosphere: The blanket of air around the earth is called atmosphere. Non-metallic gases like nitrogen, carbon dioxide and oxygen are present majorly.

We know that metals are highly reactive. Therefore, they do not occur in the free state. For example, sodium, potassium, etc. are never found in the free state and occur in nature in chemically combined forms known as **minerals.**

Elements or compounds, which occur naturally in the Earth's crust, are known as minerals. Most minerals found in the earth's crust contain metals. Some metals are also found in the oceans in the form of salts such as sodium chloride, magnesium chloride, etc.

There are some minerals which contain a large amount of a particular metal and from them, metals can be extracted profitably (using practically possible techniques).

The minerals from which metals can be extracted commercially are known as ores.

The process by which a pure metal is obtained from its ore is known as extraction.

Do you know that metals are classified into three groups on the basis of their reactivity series? The three groups are as follows:

1. Metals of low reactivity

2. Metals of medium reactivity

3. Metals of high reactivity

Reactivity series

The reactivity series is a list of metals arranged in the order of their decreasing reactivity (as shown in **Figure 1**).

The metals at the bottom of the reactivity series are less reactive and they often exist in nature in the free state. For example, gold, silver, and platinum are less reactive metals.

The metals at the top of the reactivity series are very reactive. Hence, they never occur in nature in the free state. For example, sodium, potassium, calcium, etc. are highly reactive but do not occur in nature in the free state.

The metals in the middle of the reactivity series are moderately reactive. For example, zinc, iron, and lead are moderately reactive.

Now, let us study how these metals are extracted. A number of steps are involved in the extraction of metals. They can be summarized in the form of a figure as follows.

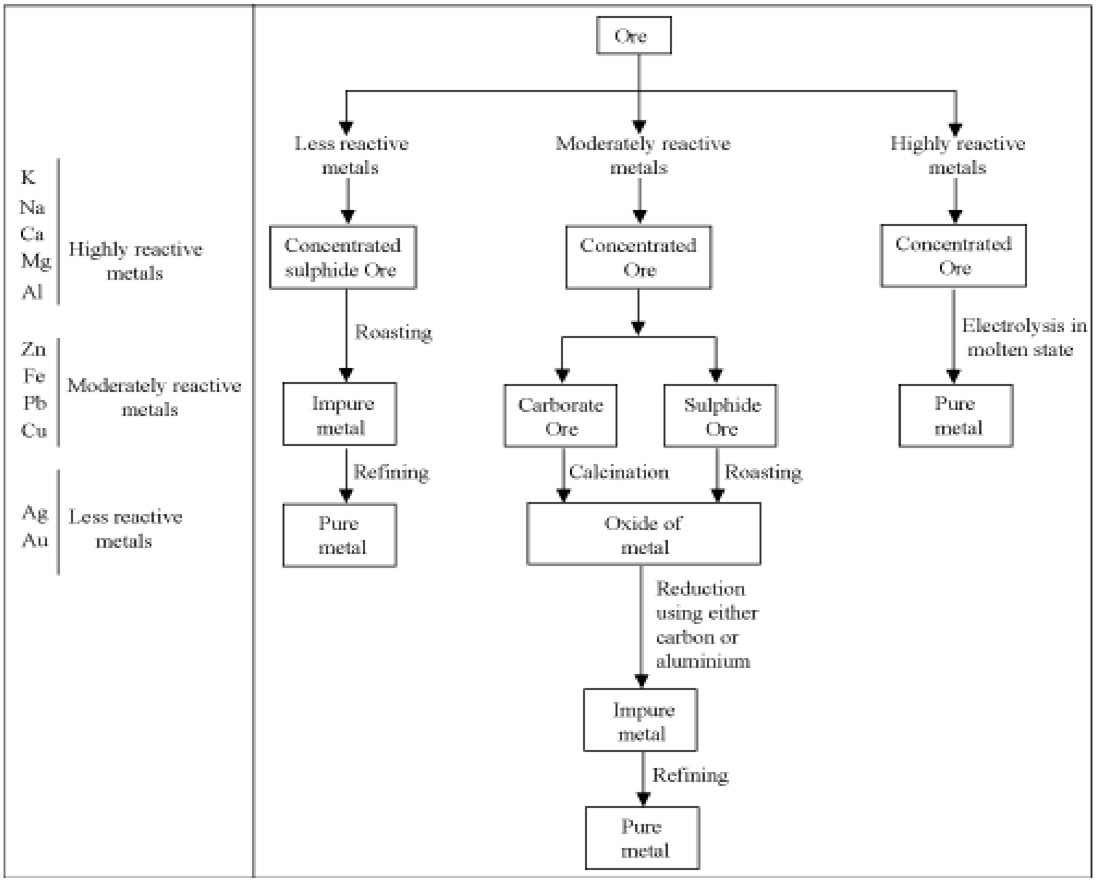


Figure 1: Reactivity series and steps involved in the extraction of metals from ores.

An ore is found in nature along with a large number of impurities such as sand, silt, soil, and gravel. The impurities are called **gangue**.

Hence, before the process of extracting a metal from its ore begins, these impurities, called gangue, have to be removed. When the gangue is removed from the ore, the **enriched or concentrated ore** is obtained.

Thus, the process of removing gangue from the ore is known as the **enrichment of ore**. The process that is employed to separate the gangue from the ore depends upon the physical and chemical properties of both the gangue and the ore.

Now, let us see how metals at the bottom of the reactivity series can be extracted. **These metals are unreactive.** The oxides of these metals can be reduced easily on heating as follows:

Step I: In the first step, cinnabar (an ore of mercury, HgS) is converted into mercury (II) oxide. Sulphur dioxide is also released in the process.

$$\underset{\text{Mercury (II) sulphide}}{2HgS(s)} + \underset{\text{Oxygen}}{3O_2(g)} \xrightarrow{\Delta} \underset{\text{Mercury (II) oxide}}{2HgO(s)} + \underset{\text{Sulphur dioxide}}{2SO_2(g)\uparrow}$$

Step II: In the second step, mercury (II) oxide obtained by heating the cinnabar is reduced to mercury metal on further heating.

$$\underset{\text{Mercury (II) oxide}}{2HgO(s)} \xrightarrow{\Delta} \underset{\text{Mercury}}{2Hg(l)} + \underset{\text{Oxygen}}{O_2(g)}$$

Concentration (or Dressing or Benefaction) of ores

- Removal of unwanted materials such as sand and clay from ores

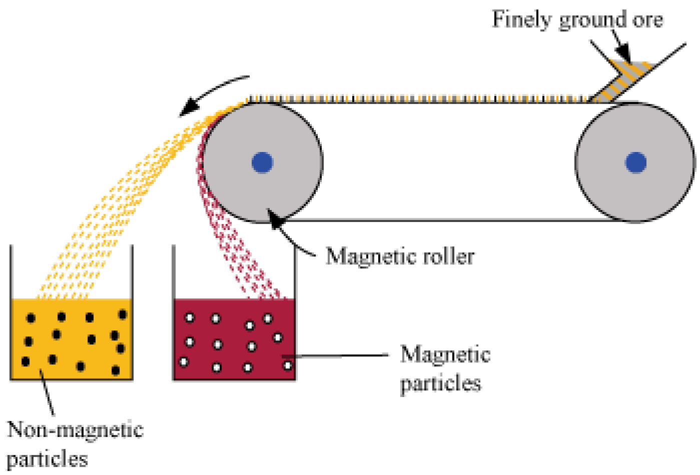
Some Important Procedures

Hydraulic washing

- Based on gravity differences between the ore and the gangue particles
- In one such process, the lighter gangue particles are washed away by a stream of water, leaving behind the heavier ore.

Magnetic separation

- Based on differences in magnetic properties between the ore and the gangue particles
- Magnetic field is applied to separate magnetically attractive particles from magnetically non-attractive particles.
- Schematic diagram of magnetic separation is as follows:



Froth-Floatation method

- Applied to remove gangue from sulphide ores

- Mineral and gangue particles are separated by first wetting the mineral particles with oil, and gangue particles with water, and then the mineral particles are carried out by forming froth.

  Sometimes, depressants are used for separating two sulphide ores by selectively preventing one ore from forming froth. For example, NaCN is used as a depressant for separating two sulphide ores, ZnS and PbS.

Leaching

- It is applied for extracting metals like gold, silver and aluminium.
- It is based on the difference in the solubility of the ore and gangue in a particular type of solution.
- In this method, the ore is soaked in a solution in which the ore dissolves whereas the gangue particles remain insoluble.
- So, the gangue particles and the ore can be separated easily.
- For example: Aluminium is extracted from its ore, bauxite, by soaking bauxite in an aqueous solution of NaOH or Na_2CO_3.

DO YOU KNOW?

Metals such as mercury and copper, which lie quite low in the reactivity series, exist in nature as sulphides. These ores when heated react with oxygen present in the air and get converted into oxides. When these oxides are further heated, pure metals are obtained. The process in which a sulphide ore is heated in the presence of air is known as **roasting.**

Metals of medium reactivity i.e., metals that are present in the middle of the reactivity series such as zinc, iron, lead, and manganese are quite reactive and exist in nature as oxides, sulphides, and carbonates. These metals are extracted from their ores by first converting ores to oxides and then by the reduction of these oxides, mostly using carbon.

There are two methods by which ores are converted into their respective oxides:

1. **Roasting**

2. **Calcination**

Roasting: It is used to convert sulphide ores into oxides. Roasting involves strong heating of iron ore in the **presence** of excess air. For example, copper sulphide in copper glance ore is converted into copper (I) oxide by heating it in the presence of oxygen.

$$\underset{\text{Copper (I) sulphide}}{2Cu_2S(s)} + \underset{\text{Oxygen}}{3O_2(g)} \xrightarrow{\text{Roasting}} \underset{\text{Copper (I) oxide}}{2Cu_2O(s)} + \underset{\text{Sulphur dioxide}}{2SO_2(g)\uparrow}$$

Calcination: It is used to convert carbonate ores into oxides. Calcination involves strong heating of the ore in the **absence** of air. For example, calamine ore, which is chemically zinc carbonate, is converted into zinc oxide by heating it in the absence of air.

$$\underset{\text{Calamine ore}}{2ZnCO_3(s)} \xrightarrow{\text{Calcination}} \underset{\text{Zinc oxide}}{2ZnO(s)} + \underset{\text{Carbon dioxide}}{2CO_2(g)\uparrow}$$

After obtaining metal oxides from the ores, reduction of these metal oxides is done to obtain pure metals. Mostly, carbon in the form of coke is used for this.

However, **the oxides of metals, which are present relatively higher in the reactivity series such as manganese, cannot be reduced with coke**. To reduce these oxides of metals, more reactive metals than manganese such as sodium, calcium, and aluminium are used. For example, iron is also a very reactive metal and cannot be reduced using carbon. Hence, it is reduced using aluminium metal. The reaction is highly exothermic. The heat evolved is so large that the metals are produced in the molten form. This reaction is known as thermite reaction and is used to join railway tracks or cracked machine parts.

$$\underset{\text{Iron (III) oxide}}{Fe_2O_3(s)} + \underset{\text{Aluminium}}{2Al(s)} \rightarrow \underset{\text{Iron}}{2Fe(l)} + \underset{\text{Aluminium oxide}}{Al_2O_3(s)} + \text{Heat}$$

Similarly, manganese cannot be reduced using carbon. Hence, it is also reduced by aluminium metal.

$$\underset{\text{Manganese dioxide}}{3MnO_2(s)} + \underset{\text{Aluminium}}{4Al(s)} \rightarrow \underset{\text{Manganese}}{3Mn(l)} + \underset{\text{Aluminium oxide}}{2Al_2O_3(s)} + \text{Heat}$$

Metals present at the top of the series such as sodium, potassium, calcium, manganese, and aluminium are very reactive**.** These metals cannot be reduced using coke as their affinity for oxygen is much more than that of carbon. Therefore, these metals are reduced by passing an electric current through their molten salts. This process is known as **electrolytic reduction.**

For example, sodium metal is extracted from sodium chloride. To extract the metal, electrolytic reduction of molten sodium chloride is carried out. When an electric current is passed, sodium ions which have positive charge move towards the cathode and get deposited over it after accepting electrons. The chloride ions have a negative charge and move towards the anode, lose their extra electrons, and escape out of the solution as chlorine gas.

Reaction at the cathode (negative electrode): $Na^+ + e^- \rightarrow Na$

Reaction at the anode (positive electrode): $2Cl^- \rightarrow Cl_2 + 2e^-$

$$\text{Net reaction: } \underset{\substack{\text{Sodium chloride}\\ \text{(Molten)}}}{2NaCl(l)} \xrightarrow{\text{Electrolytic reduction}} \underset{\text{Sodium}}{2Na(s)} + \underset{\text{chlorine}}{Cl_2(g)}$$

Do you know that the metals obtained by various reduction processes, except electrolytic reduction, contain many impurities and require purification? How are these metals purified?

The method that is most commonly used to purify metals is electrolytic refining. Many metals such as copper, zinc, gold, etc. are refined electronically.

Refining:
This method is used for the metals with low melting points i.e, which melt easily. A furnace with a slope in it, temperature is kept slightly higher than the melting point of the metal. When the impure metal is passed through the furnace, the pure metal is melted there and collected in the vessel. However, the melting points of the impurities is higher than the the metal so that they can be found solid on the slope.

Zone Refining:
This method works on the principle of fractional distillation and trace impurities are

removed from the metal using this method. The impurities remain more soluble in molten form which upon cooling the molten metal, decreases the solubility of impurities and separates in the from crystals. Semi-metals such as boron, silicon are refined by this method.