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BOOKS

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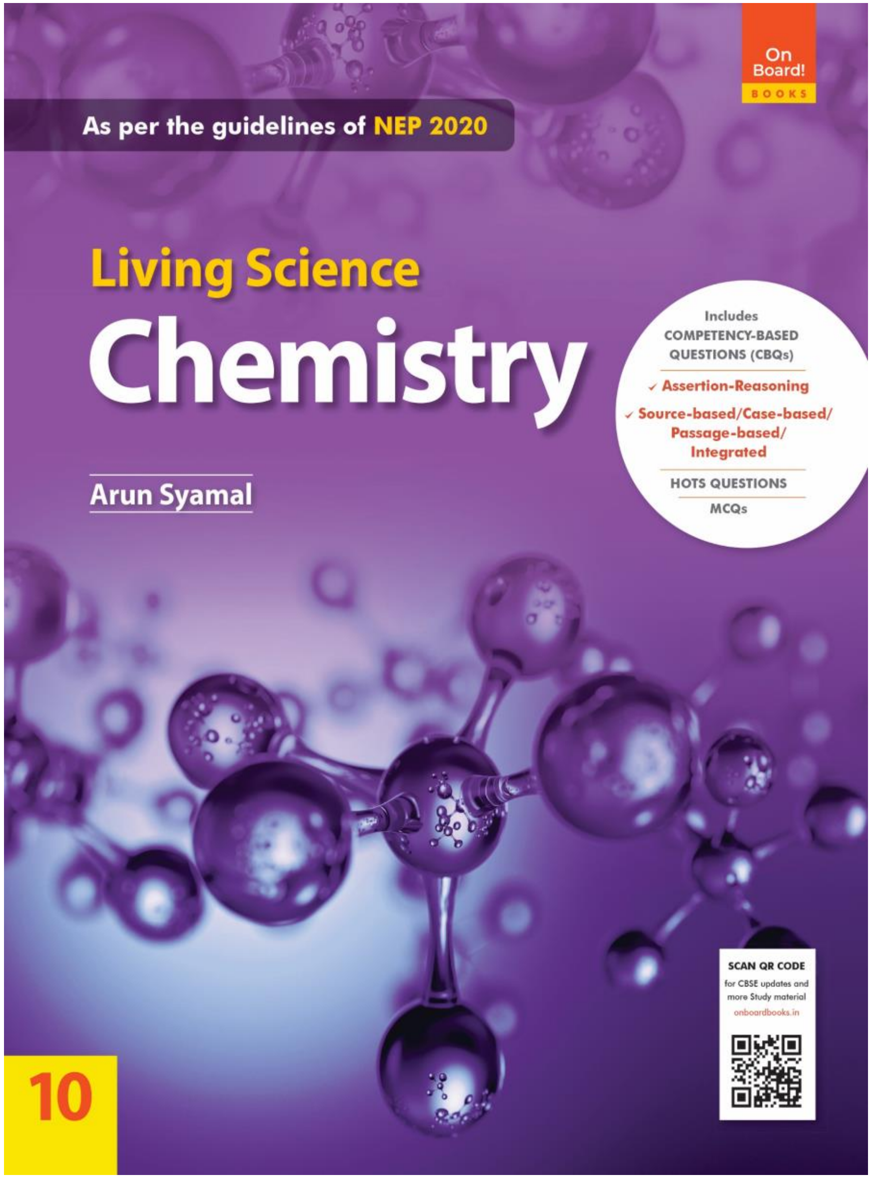


CBSE LIVING SCIENCE CHEMISTRY

CLASS 10

Chapter 3

Metals and Non-metals

Learning Objectives

- Properties of metals and non-metals
- Elementary idea about bonding in metals
- Brief discussion of basic metallurgical processes

The elements are classified into metals, non-metals and metalloids (semi-metals) on the basis of their chemical properties.

There are only 20 non-metals and out of these 7 are solids, 12 are gases and only 1 is liquid (bromine) at room temperature. Since non-metals exist as solids, liquids or gases, they exhibit a wide range of physical properties in comparison to metals.

Some elements exhibit the properties of metals as well as of non-metals and these elements are called metalloids or semi-metals.

PROPERTIES OF METALS

Physical properties of metals

The important physical properties of metals are discussed below:
1. **Metallic lustre:** All metals in their pure state, have a lustre called metallic lustre. All metals except copper and gold have a shining silvery white appearance.

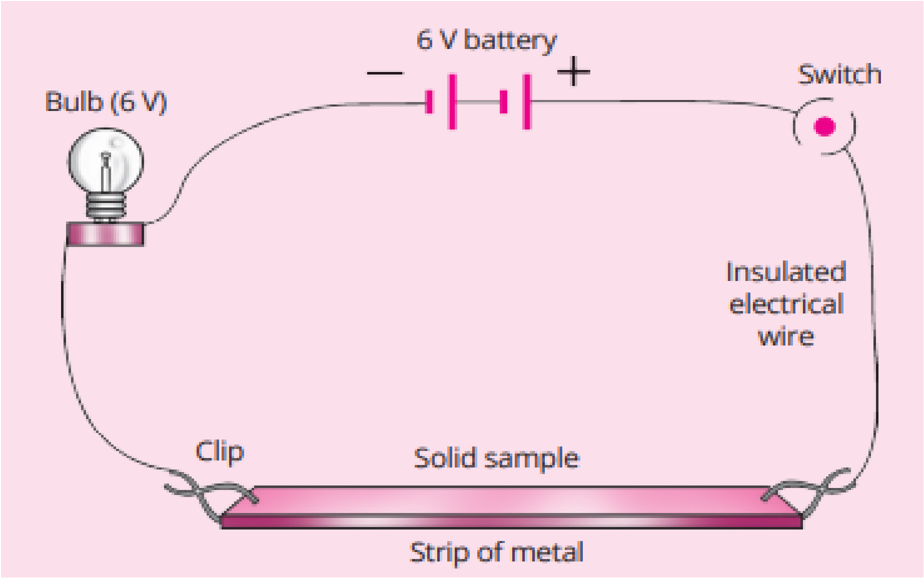
2. **Hardness:** Metals are generally hard except alkali metals such as sodium, potassium, etc. The hardness varies from metal to metal.

3. **Malleability and ductility:** Metals are malleable and ductile. Malleability is the property by which metals can be beaten into thin sheets with a hammer without breaking. Ductility is the property by which metals can be drawn into thin wires. Gold is the most malleable metal and silver is the most ductile metal.

4. **Melting point and boiling point:** Generally, metals have high melting and boiling points except mercury and alkali metals which have comparatively lower melting and boiling points.

5. **Sonority:** Metals are sonorous. They produce sharp ringing sound when hit by an object.

6. **Thermal and electrical conductivity:** Metals are good conductors of heat and electricity. Silver is the best conductor of heat and electricity. Mercury and lead are comparatively poor conductors of heat.



Circuit for finding electrical conductance of metal sample

7. **Tensile strength:** Metals possess high tensile strength. They can resist stretching without breaking.

Some exceptions in the physical properties of metals are mentioned below:

- Metals are usually hard but alkali metals such as lithium, sodium, potassium, etc., are very soft and they can be cut easily with a knife. The alkali metals exhibit low melting points and low densities.
- All metals are solid at room temperature except mercury which is liquid at room temperature. Metals have high melting points but gallium and caesium exhibit very low melting points. They can melt when kept on palm.

Chemical properties of metals

Electropositive character

All metals are electropositive in nature. They have the tendency to lose their valence electrons easily to form cations.

$$M(g) \rightarrow M^{+}(g) + e^{-}$$

We know that the loss of electrons is called oxidation. Since the metals lose electrons easily, they are oxidised easily and hence they are good reducing agents.

Activity series of metals

The arrangement of metals in the order of their decreasing reactivity is called the activity series of metals. Although hydrogen is not a metal, it is included in the activity series of metals because similar to metals, hydrogen loses electron to form the positive ion H^+.

Activity series	Reactivity	Electropositive character	Reducing power
K	Decreases ↓	Decreases ↓	Decreases ↓
Na			
Ca			
Mg			
Al			
Zn			
Fe			
Sn			
Pb			
H			
Cu			
Hg			
Ag			
Au			
Pt			

Properties of metals based on the activity series

Chemical reactions of metals

1. **Reaction with oxygen:** Metals combine with oxygen to form metal oxides. Their reaction with oxygen can be explained in terms of activity series.

Potassium and sodium react vigorously with oxygen to form oxide (K_2O, Na_2O) and they can catch fire if they are kept in open. Hence, they are kept immersed in kerosene oil in order to protect them from oxidation and to prevent accidental fires. Calcium reacts with oxygen slowly to form CaO. Metals such as Mg, Al, Zn, Fe, Sn, Pb and Cu react with oxygen when heated. Mercury reacts with oxygen reversibly while silver and gold do not react with oxygen even at high temperatures.

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

$$\underset{\text{Sodium}}{4Na(s)} + \underset{\text{Oxygen}}{O_2(g)} \xrightarrow{\text{room temperature}} \underset{\text{Sodium oxide}}{2Na_2O(s)}$$

$$\underset{\text{Copper}}{2Cu(s)} + \underset{\text{Oxygen}}{O_2(g)} \xrightarrow{\text{heat}} \underset{\substack{\text{Copper(II) oxide} \\ \text{(black)}}}{2CuO(s)}$$

$$\underset{\text{Aluminium}}{4Al(s)} + \underset{\text{Oxygen}}{3O_2(g)} \xrightarrow{25^\circ C} \underset{\text{Aluminium oxide}}{2Al_2O_3}$$

$$\underset{\text{Calcium}}{2Ca(s)} + \underset{\text{Oxygen}}{O_2(g)} \xrightarrow{\text{heat}} \underset{\text{Calcium oxide}}{2CaO(s)}$$

When the metals such as aluminium, magnesium, lead, zinc, etc., are exposed to air at ordinary temperature, a thin layer of oxide is formed on their surfaces due to oxidation. This protective layer of oxide protects the metal from further oxidation.

The oxides of metals are either basic or amphoteric in nature. For example, Na_2O, K_2O, CaO, MgO, etc., are basic oxides. Most of the metal oxides are insoluble in water but some basic oxides such as Na_2O, K_2O, CaO, etc., dissolve in water to form alkalis. Their aqueous solutions turn red litmus blue.

$$\underset{\text{Sodium oxide}}{Na_2O} + \underset{\text{Water}}{H_2O} \rightarrow \underset{\text{Sodium hydroxide}}{2NaOH}$$

$$\underset{\text{Potassium oxide}}{K_2O} + \underset{\text{Water}}{H_2O} \rightarrow \underset{\text{Potassium hydroxide}}{2KOH}$$

A basic oxide reacts with acids to form a salt and water. For example,

$$\underset{\text{Calcium oxide}}{CaO} + \underset{\text{Hydrochloric acid}}{2HCl} \rightarrow \underset{\text{Calcium chloride}}{CaCl_2} + H_2O$$

The metal oxides such as aluminium oxide (Al_2O_3) and zinc oxide (ZnO), which react with both acids and bases to form salts and water, are called amphoteric oxides. An amphoteric oxide is both basic and acidic in nature. It reacts with both acid and base to produce salt and water. For example, acidic and basic character of aluminium oxide can be described as:

$$\underset{\text{Aluminium oxide}}{Al_2O_3} + \underset{\text{Sodium hydroxide}}{2NaOH} \rightarrow \underset{\text{Sodium aluminate}}{2NaAlO_2} + H_2O$$

$$\underset{\text{Aluminium oxide}}{Al_2O_3} + \underset{\text{Hydrochloric acid}}{6HCl} \rightarrow \underset{\text{Aluminium chloride}}{2AlCl_3} + 3H_2O$$

2. **Reaction with water:** Metals react with water to form metal oxides and hydrogen gas. The metal oxides which are soluble in water react with water to form the corresponding metal hydroxides.

Metal + Water → Metal oxide(*aq*) + Hydrogen(*g*)

Metal oxide + Water → Metal hydroxide

Potassium and sodium react vigorously with cold water displacing hydrogen. The reaction is highly exothermic and the evolved hydrogen gas (H_2) catches fire immediately. Calcium reacts less vigorously with cold water. The heat evolved is not sufficient for hydrogen gas to catch fire.

The bubbles of H_2 gas formed stick to the surface of metal and it starts floating on the surface of water.
Magnesium does not react with cold water. It reacts only with hot water forming $Mg(OH)_2$ and H_2. It also starts floating on the surface of water due to the bubbles of H_2 gas sticking on the surface of water.

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

$$\underset{\text{Calcium}}{Ca(s)} + \underset{\text{Cold water}}{2H_2O(l)} \rightarrow \underset{\substack{\text{Calcium}\\ \text{hydroxide}}}{Ca(OH)_2(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

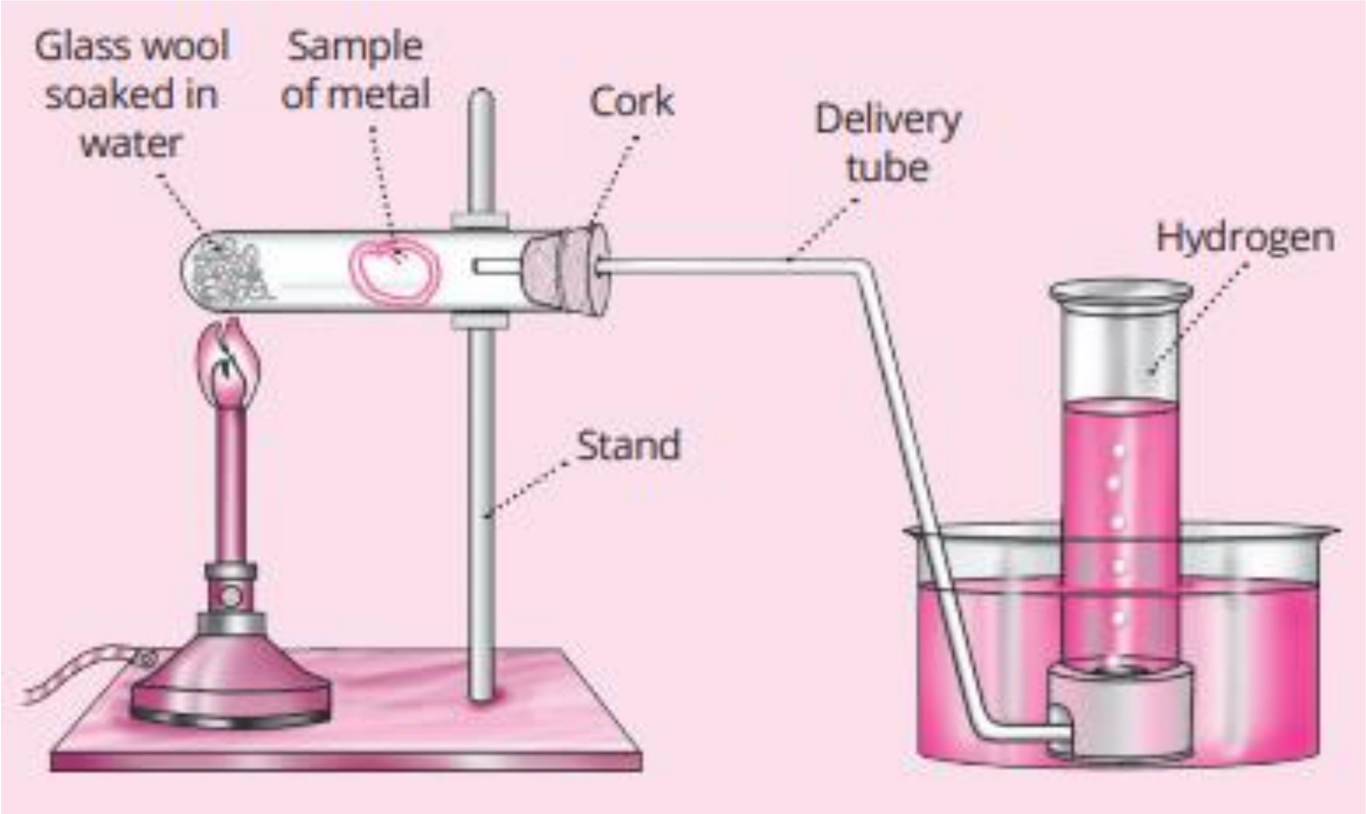
$$\underset{\text{Magnesium}}{Mg(s)} + \underset{\text{Hot water}}{2H_2O(l)} \rightarrow \underset{\substack{\text{Magnesium}\\ \text{hydroxide}}}{Mg(OH)_2(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

$$\underset{\text{Aluminium}}{2Al(s)} + \underset{\text{Steam}}{3H_2O(g)} \rightarrow \underset{\substack{\text{Aluminium}\\ \text{oxide}}}{Al_2O_3(s)} + \underset{\text{Hydrogen}}{3H_2(g)}$$

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

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

Only the metals which are placed above hydrogen in the activity series displace hydrogen from water or steam. The reactions of metals with water are redox reactions.



Action of steam on a metal

3. **Reaction with acids:** Metals which are above hydrogen in the activity series displace hydrogen from dilute acids (HCl or H_2SO_4). For example, potassium and sodium react vigorously with dilute acids. Calcium and magnesium react less vigorously with dilute acids than sodium. Aluminium reacts with dilute acids less rapidly than magnesium. Copper, mercury and silver do not react with dilute acids at all.

$$\underset{\text{Potassium}}{2K(s)} + \underset{\substack{\text{Hydrochloric}\\\text{acid}}}{2HCl(aq)} \rightarrow \underset{\substack{\text{Potassium}\\\text{chloride}}}{2KCl(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

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

$$\underset{\text{Aluminium}}{2Al(s)} + \underset{\substack{\text{Sulphuric}\\\text{acid}}}{3H_2SO_4(aq)} \rightarrow \underset{\substack{\text{Aluminium}\\\text{sulphate}}}{Al_2(SO_4)_3(aq)} + \underset{\text{Hydrogen}}{3H_2(g)}$$

$$\underset{\text{Zinc}}{Zn(s)} + \underset{\substack{\text{Sulphuric}\\\text{acid}}}{H_2SO_4(aq)} \rightarrow \underset{\substack{\text{Zinc}\\\text{sulphate}}}{ZnSO_4(aq)} + \underset{\text{Hydrogen}}{H_2(g)}$$

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

The rate of reaction of some metals with dil. Hcl increases in the order:

$$Cu < Fe < Zn < Al < Mg < Ca < Na$$

Thus, a more electropositive metal displaces hydrogen from dilute acids more readily.

4. **Reaction with dilute nitric acid:** Only magnesium(Mg) and manganese(Mn) react with very dilute nitric acid to produce metal nitrate and hydrogen gas.

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

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

When other metals react with nitric acid, hydrogen gas is not evolved. Since nitric acid is a strong oxidizing agent and hydrogen is a reducing agent, nitric acid oxidises the hydrogen produced to water and itself undergoes reduction to give any of the nitrogen oxides (N_2O, NO, NO_2) and ammonia (NH_3), etc.

5. **Reactions of metals with other metal salt solutions (displacement reactions):** The more reactive metals displace less reactive metals from their salt solutions. For example, when a piece of granulated zinc is added to a blue solution of copper(II) sulphate, the blue solution fades gradually due to the formation of zinc(II)sulphate and red copper metal is deposited on granulated zinc. Since zinc is a more reactive metal, it displaces copper from a solution of copper(II)sulphate. This means, if metal A displaces metal B from its solution, it is more reactive than B.

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

Some examples of metal displacement reactions are as follows:

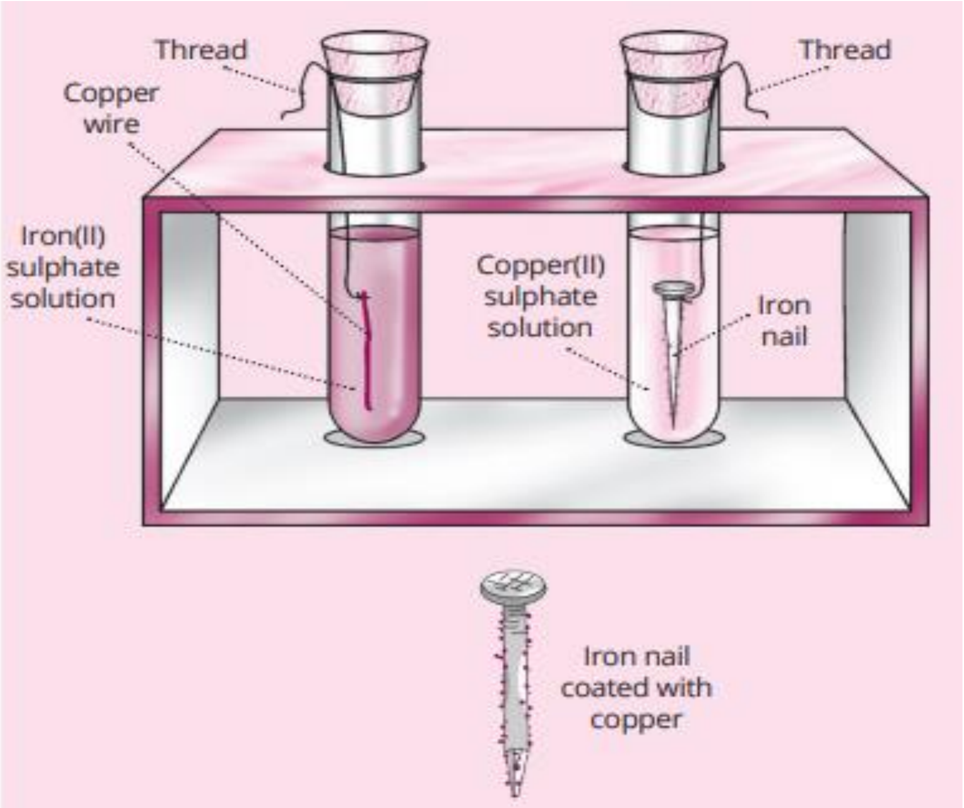
$$\underset{\substack{\text{Copper(II)}\\\text{sulphate}\\\text{(blue)}}}{CuSO_4(aq)} + \underset{\text{Zinc}}{Zn(s)} \rightarrow \underset{\substack{\text{Zinc(II)}\\\text{sulphate}\\\text{(colourless)}}}{ZnSO_4(aq)} + \underset{\substack{\text{Copper}\\\text{(red)}}}{Cu(s)}$$

$$\underset{\substack{\text{Silver nitrate}\\\text{(colourless)}}}{2AgNO_3(aq)} + \underset{\substack{\text{Copper}\\\text{(red)}}}{Cu(s)} \rightarrow \underset{\substack{\text{Copper(II) nitrate}\\\text{(blue)}}}{Cu(NO_3)_2(aq)} + \underset{\substack{\text{Silver}\\\text{(white)}}}{2Ag(s)}$$

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

Activity

Take a clean copper wire and a clean iron nail. Take a test tube and add to it a dilute solution of iron(II) sulphate. Put the copper wire in the solution of iron(II) sulphate in the test tube as shown in the figure and note the time. Record your observation after twenty minutes. Take another test tube and add to it a dilute solution of copper(II) sulphate. Put the iron nail in the test tube and note the time. Record your observation after twenty minutes. It is observed that in the second test tube a reaction has occurred. The iron nail gets coated with a layer of red copper metal. It is also observed that in the first test tube no reaction has occurred. The copper wire does not get coated with iron metal. In this reaction, a more reactive metal, iron displaces a less reactive metal, copper from its compound copper(II) sulphate.



Reactions of metals with salt solutions of other metals

The chemical equation for the reaction is as follows:

$CuSO_4(aq) + Fe(s) \rightarrow FeSO_4(aq) + Cu(s)$

PROPERTIES OF NON-METALS

Physical properties of non-metals

The important physical properties of non-metals are discussed here:

1. **Hardness:** Non-metals are generally soft and brittle except diamond, an allotrope of carbon, which is the hardest natural substance known.

2. **Malleability and ductility:** Non-metals are neither malleable nor ductile. Solid non-metals are brittle.

3. **Thermal and electrical conductivity:** Non-metals are poor conductors of heat and electricity. This is due to the absence of free electrons for electrical conduction in non-metals. There is one exception in this point. Graphite, an allotrope of carbon, is a conductor of electricity.

4. **Sonority:** Non-metals are non-sonorous and do not produce sound when hit by an object.

5. **Lustre:** Non-metals are not lustrous and they have dull appearance except iodine and diamond which are lustrous.

6. **Melting and boiling points:** Generally non-metals have low melting and boiling points except boron, carbon and silicon, which have a very high melting and boiling points.

Electronegative character of non-metals

Except hydrogen and noble gases, non-metals have a strong tendency to accept electrons to form anions having the stable electronic configuration of the nearest noble gas. This means, they exhibit electronegative character in nature. Thus, they have the capacity to remove electrons from other elements. Since removal of electrons is termed as oxidation, the non-metals act as oxidising agents.

HOW DO METALS REACT WITH NON-METALS?

All elements possess a certain number of valence electrons in their valence shell and the valence electrons dictate the chemical reactions of elements.

During the formation of a compound, all elements tend to acquire the electronic configuration of the nearest noble gas which has a completely filled valence shell. For example, an atom of potassium has a tendency to lose one electron so as to acquire the electronic configuration of argon having a stable octet. But an atom of chlorine needs one electron to attain the stable electronic configuration of argon. Hence, one atom of potassium must combine with one atom of chlorine to form potassium chloride. After losing an electron, the potassium atom acquires a unit positive charge because its nucleus has 19 protons and there are total 18 electrons in its *K*, *L* and *M* shells. Thus, a potassium cation (K^+) is formed.

After accepting an electron from potassium atom, the chlorine atom acquires a unit negative charge because its nucleus has 17 protons and there are total 18 electrons in its *K*, *L* and *M* shells. Thus, a chloride anion (Cl^-) is formed. Since potassium and chloride ions are oppositely charged, they are held by strong electrostatic forces of attraction to exist as potassium chloride. A compound formed by the transfer of electron(s) from a metal to a non-metal is called **ionic compound** and the bond formed between them is called **ionic bond**. The ionic compounds exist as aggregates of a very large number of oppositely charged ions.

The formation of potassium chloride from its constituent atoms is shown below:

$$\underset{2,8,8,1}{\mathrm{K}} \longrightarrow \underset{\substack{2,8,8\\ \text{Potassium cation}}}{\mathrm{K}^+} + e^-$$

$$\underset{2,8,7}{\mathrm{Cl}} + e^- \longrightarrow \underset{\substack{2,8,8\\ \text{Chloride anion}}}{\mathrm{Cl}^-}$$

$$\dot{\mathrm{K}} + \cdot\ddot{\underset{\cdot\cdot}{\mathrm{Cl}}}: \longrightarrow \mathrm{K}^+ + :\ddot{\underset{\cdot\cdot}{\mathrm{Cl}}}:^-$$

	K	Cl	K^+	Cl^-
	Atom	Atom	Ion	Ion
Number of electrons	19	17	18	18
Electronic configuration	2, 8, 8, 1	2, 8, 7	2, 8, 8	2, 8, 8

Formation of potassium chloride

Electronic configuration of some metals, non-metals and noble gases

Type of element	Element	Atomic number	Number of electrons in shells			
			K	*L*	*M*	*N*
Metals	Lithium	3	2	1		
	Sodium	11	2	8	1	
	Magnesium	12	2	8	2	
	Aluminium	13	2	8	3	
	Potassium	19	2	8	8	1
	Calcium	20	2	8	8	2
Non-metals	Boron	5	2	3		
	Carbon	6	2	4		
	Nitrogen	7	2	5		
	Oxygen	8	2	6		
	Fluorine	9	2	7		
	Phosphorus	15	2	8	5	
	Sulphur	16	2	8	6	
	Chlorine	17	2	8	7	
Noble gases	Helium	2	2			
	Neon	10	2	8		
	Argon	18	2	8	8	
	Krypton	36	2	8	18	8

Properties of ionic compounds

The properties of ionic compounds are described below.

1. **Physical nature**

The ionic compounds are solids. Their hardness is due to the presence of strong electrostatic forces of attraction between the oppositely charged ions. They are also brittle and their crystals break into pieces on applying pressure.

2. **Melting and boiling points**

The ionic compounds exhibit high melting and boiling points since a large amount of energy is needed to break the strong inter-ionic electrostatic forces of attraction.

Ionic compound	Melting point (K)	Boiling point (K)
$LiCl$	887	1600
$NaCl$	1074	1686
KCl	1049	1773
$MgCl_2$	981	1685
$CaCl_2$	1045	1900
CaO	2850	3120
$BaCl_2$	1235	1833

Melting point and boiling point of some ionic compounds

3. **Solubility**
The ionic compounds are highly soluble in polar solvents such as water and insoluble in non-polar solvents such as benzene, chloroform, carbon tetrachloride, petrol, kerosene, etc.

4. **Electrical Conductivity**
The ionic compounds conduct electricity in aqueous solution and in molten state. In the molten state, the heat breaks the inter-ionic electrostatic forces of attraction and as a result the ions are able to move freely and they conduct electricity. Ionic compounds do not conduct electricity in the solid state as the movement of ions in solids is not possible due to their rigid structure.

5. **Aggregates**
The ionic compounds do not exist as discrete molecules in the crystal lattice but exist as aggregates of a very large number of oppositely charged ions.

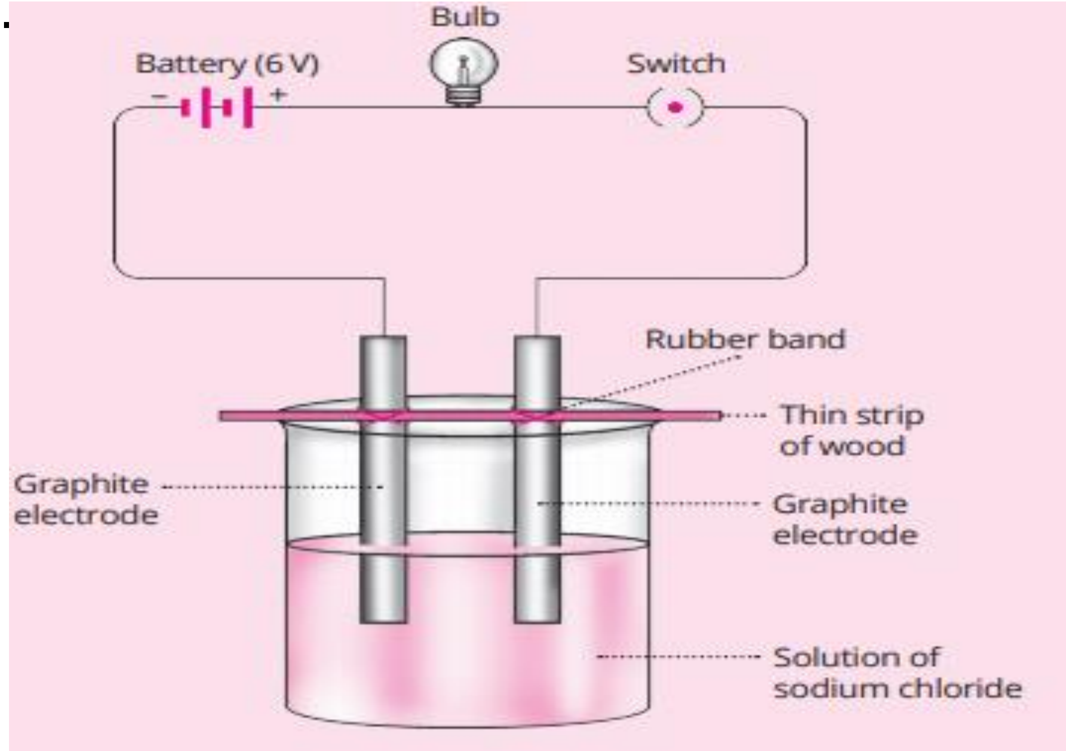
Examples of ionic compounds
The examples of ionic compounds are given below:
1. Chlorides, fluorides, nitrates, oxides and sulphides of sodium/potassium.
2. Chlorides, fluorides, nitrates and oxides of magnesium/calcium/

Activity

Testing the properties of ionic compounds
Take about 2 g each of sodium chloride, potassium iodide and barium chloride in three different test tubes. Record the physical state of these salts. Take a small amount of sodium chloride on a stainless steel spatula and heat directly on a Bunsen flame as shown in figure. Repeat heating of other two salts one by one. Record the colour of flame observed in each case. Record whether these salts melt on heating.

Heating of sodium chloride on a stainless steel spatula

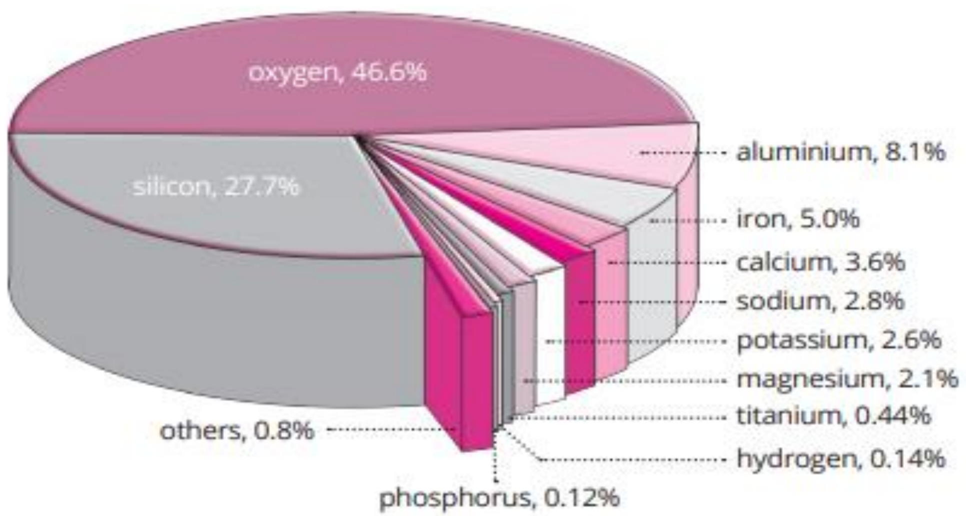
Now, take a small amount of each salt in different test tubes and try to dissolve a small amount of each salt in water, kerosene and petrol separately in different test tubes. Make an electrical circuit using electrical wire, battery (6 V), bulb and switch as shown in figure. Dissolve about 1 g of sodium chloride in 250 smL water in a 500 mL beaker. Place two graphite electrodes in the beaker. Turn on the switch. Record your observation. It is observed that the bulb glows. This indicates that an aqueous solution of sodium chloride conducts electricity. Repeat the electrical conductivity test taking the solutions of the other two salts in other beakers. Record your inference regarding the nature of these salts.



Testing the electrical conductivity of an aqueous solution of sodium chloride

OCCURRENCE OF METALS

The major source of metals is the earth's crust. Most of these metals occur in the earth's crust in the form of their salts. Metals are also present in seawater in the form of metal salts such as $NaCl$, KCl, $MgCl_2$, etc. The relative abundance of the elements in the earth's crust is shown in the pie chart below.



The relative abundance of elements in the earth's crust

Oxygen is the most abundant element in the earth's crust. The next most abundant being silicon followed by aluminium.

Minerals and ores

The natural materials found inside the earth's crust containing elements or compounds are called minerals. The minerals are usually associated with earthy materials called gangue. All minerals are not suitable for the extraction of metals economically. The naturally occurring minerals from which metals can be extracted profitably are called ores. Thus, all ores are minerals but all minerals are not ores. The metal compounds which are found in the form of minerals, carbonates or sulphides and ores in the earth's crust have low solubility in water. The names of some common ores are given in the table below.

Type of ore	Example
Native	Cu, Ag, Au, Hg, Bi, Sb, Pt, Pd
Oxides	$Al_2O_3{\cdot}2H_2O$, Cu_2O, Fe_2O_3, Fe_3O_4, SnO_2, MnO_2, ZnO
Carbonates	$CaCO_3$, $CaCO_3{\cdot}MgCO_3$, $BaCO_3$, $SrCO_3$, $FeCO_3$, $PbCO_3$, $ZnCO_3$, $CuCO_3{\cdot}Cu(OH)_2$, $MnCO_3$
Sulphides	Cu_2S, PbS, ZnS, HgS, FeS_2, CaS, NiS, Ag_2S
Halides	$NaCl$, KCl, $MgCl_2{\cdot}6H_2O$, $AgCl$, CaF_2, Na_3AlF_6
Sulphates	$CaSO_4{\cdot}2H_2O$, $BaSO_4$, $SrSO_4$, $PbSO_4$, $CuSO_4{\cdot}Cu(OH)_2$, $MgSO_4{\cdot}7H_2O$

Extraction of metals from their ores

Metals are classified into the following three categories on the basis of their reactivity:

1. Metals of low reactivity.
2. Metals of moderate reactivity.
3. Metals of high reactivity.

An element which is less reactive in the activity series often occurs in native (free) state in the earth’s crust. For example, Cu, Ag, Au, Pt and Hg. The metals which occur at the top of the activity series never occur in the free state as they are very reactive metals. For example, Na, K, Ca, Mg and Al. The metals such as Zn, Fe, Sn and Pb occurring in the middle of the activity series are moderately reactive and occur in the earth’s crust usually as oxides, carbonates or sulphides. The techniques used for the extraction of metals belonging to the above three categories are given below.

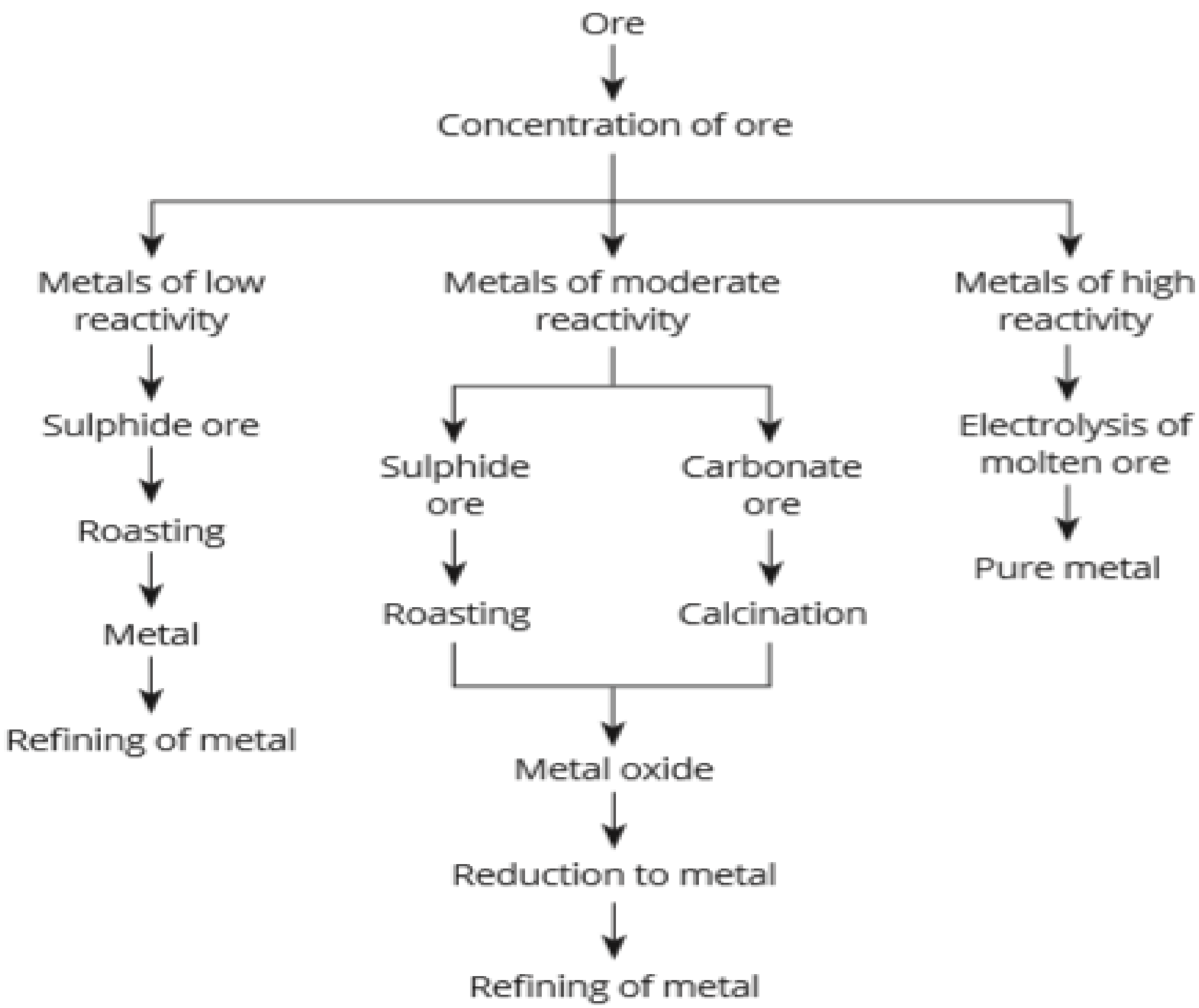
Highly reactive metals	$\begin{bmatrix} K \\ Na \\ Ca \\ Mg \\ Al \end{bmatrix}$	**By electrolysis**
Moderately reactive metals	$\begin{bmatrix} Zn \\ Fe \\ Sn \\ Pb \end{bmatrix}$	**By reduction of oxide ore with carbon**

Less reactive metals	[Cu, Hg]	**By self-reduction of sulphide ore (by heating alone)**
Least reactive metals	[Ag, Au, Pt]	**Occur in nature in the native state**

The technology employed for the extraction of metals economically on a large-scale from their ores and refining them for use is called metallurgy. The process of extraction of a metal depends on the nature of the ore, impurities present in the ore and physical and chemical properties of the metal to be extracted. The following common steps are applicable for metallurgy of most metals:

1. Concentration of ore.
2. Conversion of concentrated ore into metal oxide.
3. Extraction of metal by the reduction of metal oxide to metal.
4. Refining of impure metals.

A flow chart indicating the steps for the extraction of metals from ores is given as follows.



Steps involved in the extraction of metals from ores

Concentration or enrichment of ore

The ores are generally associated with unwanted foreign materials such as clay, sand and granite. The impurities present in an ore are called gangue. In the first step of processing of an ore, it is essential to remove the gangue. The process of removal of unwanted material (gangue) from an ore is called concentration or enrichment. Some of the methods commonly used for the concentration of ore are gravity separation, froth floatation and electromagnetic separation.

Conversion of concentrated ore into metal oxide

The ore is subjected to any one of the following processes:

1. **Calcination:** The carbonate ores are subjected to calcination. Calcination is the process in which the concentrated ore is heated strongly below its melting point in the absence or limited supply of air. It is used for carbonate ores.

$$\underset{\text{Limestone}}{CaCO_3(s)} \xrightarrow{\text{heat}} \underset{\text{Calcium oxide}}{CaO(s)} + CO_2(g)$$

$$ZnCO_3(s) \xrightarrow{\text{heat}} \underset{\text{Zinc oxide}}{ZnO(s)} + CO_2(g)$$

2. **Roasting:** The sulphide ores are subjected to roasting. Roasting is the process in which the concentrated ore is heated strongly below its melting point in the presence of excess of air. It is used for sulphide ores such as ZnS, HgS, etc.

$$\underset{\substack{\text{Zinc}\\\text{blende}}}{2ZnS(s)} + \underset{\text{Oxygen}}{3O_2(g)} \xrightarrow{\text{heat}} \underset{\substack{\text{Zinc}\\\text{oxide}}}{2ZnO(s)} + \underset{\substack{\text{Sulphur}\\\text{dioxide}}}{2SO_2(g)}$$

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

Calcination and roasting are generally applicable for the extraction of metals (Fe, Zn, Pb, etc.) which lie in the middle of the activity series and which occur in the nature in the form of their oxide, sulphide or carbonate ores.

Extraction of the metal by reduction of metal oxide

The calcined or roasted ore is converted into the free metal by simple heating or chemical reduction or electrolytic reduction. In the chemical reduction method, the choice of reducing agent depends upon the chemical reactivity of the metal.

The oxides of moderately reactive metals such as zinc, iron, lead and tin can be reduced by carbon or carbon monoxide. The oxides of very reactive metals such as alkali metals can only be reduced by the electrolytic method. The various reduction processes are discussed under the following categories:

1. Self-reduction (reduction by heat alone)
2. Chemical reduction
3. Displacement method
4. Electrolytic reduction

Self-reduction

The oxides of metals (Cu, Hg, Ag, Pt, Au) which occur near the bottom of the activity series can be reduced to free metals by the action of heat alone. No reducing agent is required. Mercury occurs in nature as the sulphide ore, cinnabar (HgS).

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Self-reduction

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When concentrated cinnabar is roasted in the presence of excess of air, mercury(II) sulphide is converted into mercury(II) oxide which is thermally unstable and decomposes to form mercury.

$$\underset{\text{Cinnabar}}{2HgS(s)} + \underset{\text{Oxygen}}{3O_2(g)} \xrightarrow[\text{(roasting)}]{\text{heat}} \underset{\substack{\text{Mercury(II)}\\ \text{oxide}}}{2HgO(s)} + \underset{\substack{\text{Sulphur}\\ \text{dioxide}}}{2SO_2(g)}$$

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

The equation for the overall reaction occurring during roasting of cinnabar is obtained by adding the above two equations.

$$\underset{\text{Cinnabar}}{HgS(s)} + \underset{\text{Oxygen}}{O_2(g)} \xrightarrow{\text{heat}} \underset{\text{Mercury}}{Hg(l)} + \underset{\substack{\text{Sulphur}\\ \text{dioxide}}}{SO_2(g)}$$

Similarly, copper is extracted from its sulphide ore copper glance (Cu_2S) by just heating in air.

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

$$2Cu_2O(s) + Cu_2S(s) \xrightarrow{\text{heat}} 6Cu(s) + SO_2(g)$$

Chemical reduction
Coke or carbon monoxide reduces the oxides of moderately reactive metals such as zinc, iron, tin, lead and copper, which occur at the middle of the activity series, to the corresponding metals. For example,

$$\underset{\text{Zinc oxide}}{ZnO(s)} + \underset{\text{Coke}}{C(s)} \xrightarrow{\text{heat}} \underset{\text{Zinc}}{Zn(s)} + \underset{\text{Carbon monoxide}}{CO(g)}$$

Aluminothermic reduction
The carbon impurity is difficult to remove from many metals (Mn, Cr, Fe), the carbon reduction method is unsuitable for the metallurgy of these metals in pure form. In such cases, a reactive electropositive metal aluminium is used as a reducing agent. This is because a more electropositive element can displace a less electropositive element from its oxides to give free metals. Thus, displacement reaction can also be used to reduce metal oxides into free metals.

When aluminium is used as a reducing agent the process is called aluminothermic reduction. The reduction of metal oxide using aluminium as the reducing agent is called aluminothermy. These reactions are highly exothermic. The amount of heat liberated is so high that the metal is obtained in molten state.

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

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

The above exothermic reaction of iron(III) oxide with aluminium is known as thermit reaction. The molten iron produced in this process is used for welding red hot steel. It is used for joining the broken pieces of heavy steel objects such as railway tracks or cracked machine parts.

Displacement method: A more reactive metal can displace a less reactive metal from its salt solution. This technique is used in cases where the leaching method has been used for concentration of an ore. An important example is the extraction of copper from copper glance (Cu2S). When copper glance is exposed to air and water, it is oxidised to copper(II) sulphate. Copper is precipitated from copper(II) sulphate solution by adding a more electropositive metal such as iron.

Electrolytic reduction

The carbon reduction method is unsuitable for the metallurgy of highly reactive metals (K, Ca, Na, Mg and Al). These metals cannot be obtained by reduction of their oxide ores by carbon or aluminium.

This is because these highly reactive metals have more affinity for oxygen than for carbon or aluminium. The oxides of these highly reactive metals are very stable. So carbon is unable to remove oxygen from these metal oxides and hence cannot convert them into free metals. These metals are obtained by the electrolysis of their molten chlorides or oxides.

The process of obtaining a metal by the electrolysis of molten salt is called electrometallurgy. For example, sodium chloride is electrolysed using iron as cathode and graphite anod

$$Na^+Cl^- \text{ (molten)} \rightarrow Na^+ + Cl^-$$

At cathode:

$$\underset{\text{Sodium ion}}{2Na^+} + \underset{\text{Electron}}{2e^-} \rightarrow \underset{\text{Sodium metal}}{2Na}$$

At anode:

$$\underset{\text{Chloride ion}}{2Cl^-} \rightarrow \underset{\text{Chlorine gas}}{Cl_2(g)} + \underset{\text{Electron}}{2e^-}$$

Refining of impure metals

The metals extracted by the reduction operations contain impurities such as carbon, silicon, phosphorus, other metals, dissolved oxides, sulphides, etc. Hence, the impure metals require refining. The method which is commonly used for refining of metals is electrolytic refining.

Electrolytic refining

This method is used in the refining of copper, silver, gold, zinc, tin, lead, aluminium, nickel, chromium, etc. The thick plates of impure metal are made anode and a thin plate of pure form of the same metal is made cathode in an electrolytic cell.

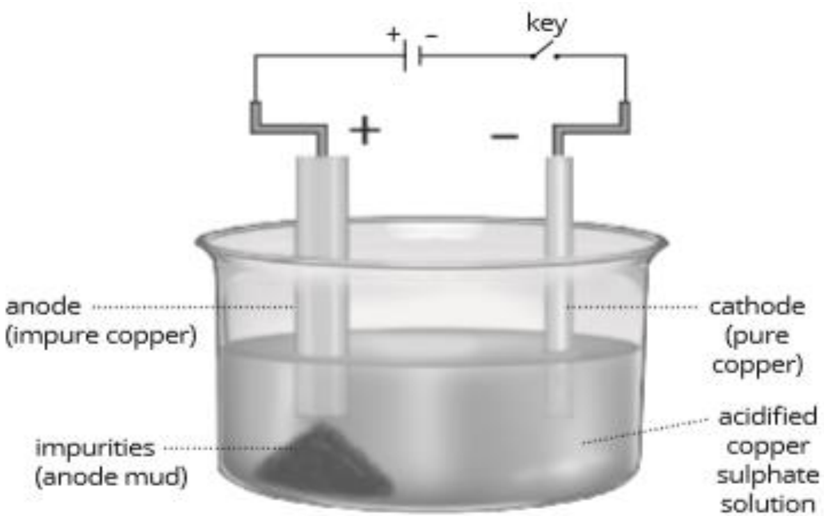
The electrolyte consists of an aqueous solution of a salt of the metal. When an electric current is passed, the anode undergoes dissolution while the pure metal is deposited on the cathode. The insoluble impurities settle at the bottom of the electrolytic cell and are recovered as valuable anode mud. The soluble impurities remain in the solution. Highly pure metal is produced by this method. For example, in the electrorefining of copper, anode is made of a block of impure copper containing traces of iron, silver and gold, cathode is made of a thin plate of pure copper and the electrolyte used is a solution of copper(II) sulphate containing a small amount of dilute H_2SO_4.

On passing electric current, the following reactions occur:

$$\textbf{At cathode: } Cu^{2+} + 2e^{-} \rightarrow Cu(s)$$

$$\textbf{At anode: } Cu(s) \rightarrow Cu^{2+} + 2e^{-}$$

Copper dissolves from the anode and goes into the electrolyte and an equivalent amount of copper from the electrolyte gets deposited on the cathode. The insoluble impurities containing less reactive metals such as silver and gold fall to the bottom of the electrolytic cell in the form of anode mud. The more reactive metals such as iron dissolve in the solution and remain in the solution.



Refining of a metal by electrolysis

CORROSION

The phenomenon of deterioration or destruction of a metal when exposed to air, water or any other substance around it in the atmosphere, is called corrosion. Corrosion occurs at the surface of a metal. It is a spontaneous irreversible process. The examples of corrosion are:

1. When a piece of iron or any iron article is exposed to moist air for a long time, a layer of reddish brown flaky and porous substance is formed on the surface of iron. This substance is called rust. Rust is hydrated iron(III) oxide ($Fe_2O_3 \cdot xH_2O$).

2. When silver articles are exposed to air, they become black after some time because they react with hydrogen sulphide gas present in air to form a coating of black silver sulphide.

$$\underset{\text{Silver}}{2Ag(s)} + H_2S(g) \rightarrow \underset{\substack{\text{Silver sulphide}\\ \text{(black)}}}{Ag_2S(s)} + H_2(g)$$

3. When a piece of copper or any copper utensil is exposed to moist air containing carbon dioxide, it loses its red shine after some time due to the formation of a green coating of basic copper(II) carbonate [$CuCO_3 \cdot Cu(OH)_2$].

4. When a piece of aluminium or any aluminium article is exposed to air, a thin protective layer of white aluminium oxide ($Al_2O_3 \cdot 2H_2O$) is formed on its surface after some time. Although there occurs initial damage to aluminium due to the formation of aluminium oxide but in the long run it is an advantage since the thin layer of aluminium oxide protects the aluminium underneath from further damage. Due to this, aluminium is used to make kitchen utensils, although aluminium is a very reactive metal.

Due to corrosion, most iron and steel structures get damaged. We have to spend a large amount of money for prevention of their damage due to corrosion or for replacement of iron and steel structures.

Prevention of corrosion

We have learned that the conditions necessary for the occurrence of rusting of iron are air and moisture. Hence, the methods of prevention of corrosion of a metal must aim at protecting it from air and moisture. The following methods are commonly used for prevention of corrosion:

1. **Protection by coating with oil or grease:** Corrosion is prevented by coating the material with oil or grease.
2. **Protection by covering with paint, plastic, rubber or ceramic:** Corrosion is prevented by covering the material with a coat of paint, plastic, rubber or ceramic.

3. **Prevention by galvanization:** Corrosion is prevented by coating iron objects with a thin layer of zinc. This process of coating iron objects with a thin layer of zinc is called galvanization. Zinc is a more reactive metal than iron. It reacts with air in preference to iron when exposed to moist air. On exposure to moist air, zinc being more electropositive than iron, gets oxidized protecting iron from oxidation. An invisible layer of basic zinc carbonate [$ZnCO_3 \cdot Zn(OH)_2$] is formed, which prevents iron objects from corrosion. Galvanization is done by dipping the iron or steel article in molten zinc and then taking it out. The galvanized articles are protected against corrosion even if the zinc coating is broken. When zinc coating gets a crack, zinc continues to get corroded in preference to iron since zinc is more reactive than iron.

4. **Prevention by electroplating:** Corrosion is prevented by electroplating iron with a more resistant metal such as chromium or nickel.

5. **Prevention by anodising:** Corrosion is prevented by anodizing aluminium articles by coating with a thick protective layer of aluminium oxide.

6. **Protection by alloying with metals:** Alloying is a very good method of improving the properties of a metal. For example, iron is protected from corrosion by alloying it with chromium, nickel, vanadium and tungsten. The rust proof stainless steel is an alloy of Fe, Cr, Ni (Fe, 74%; Cr, 18%; Ni, 8%) and C, and is used for making knives, cutlery, utensils, sinks, etc.

Formation of alloys

A homogeneous mixture of a metal with other metals or a metal and a non-metal is called alloy.

Alloys are extensively used in industries because most alloys have high melting and boiling points, are resistant to corrosion and possess high mechanical strength. Due to their higher melting and boiling points and stability towards oxidation, alloys are preferentially used in electrical heating devices instead of pure metals.

Alloys exhibit electrical conductivities and melting points which are less than those of the constituent metals. For example, the electrical conductivity of brass, an alloy of copper (60–80%) and zinc (20–40%) , is less than that of zinc and copper, whereas copper is used for making electrical circuits.

Brass is used for making scientific instruments, telescopes, microscopes, etc. Other examples of few common alloys are: solder, an alloy of lead (Pb) 50% and tin (Sn) 50%, is used for soldering purposes, and welding two electrical wires together, it possesses a lower melting point than its constituents), bronze, an alloy of copper (Cu) 80%, zinc (Zn) 10% and tin (Sn) 10%, is used for making statues, machineries, coins and jewellery.

An alloy of mercury with another metal is called an amalgam. For example, sodium amalgam, zinc amalgam, etc.

Some common alloys, their constituents, properties and uses

Alloy	Constituents	Properties	Uses
Brass	Copper, zinc	Can be easily cast, strong, malleable, corrosion resistant	Making castings, sheets, tubes, valves, screws, nuts, bolts, utensils, school bells, cartridges, scientific instruments, telescopes, microscopes, barometers, decoration items
Bronze	Copper, tin, zinc	Very strong, highly corrosion resistant	Making coins, metal statues, machineries and ship's propellers
Steel	Iron, carbon	Very strong, hard, tough	Construction of bridges, vehicles, ships, as building material
Stainless steel	Iron, chromium, nickel	Strong, hard, resistant to corrosion (rust proof), acid proof	Making cutlery, utensils, ornamental pieces, surgical instruments
Duralumin	Aluminium, copper, magnesium, manganese	Strong, light, resistant to corrosion, ductile, easily castable	In pressure cooker, fluorescent tube caps, aircraft, automobile and ship parts
Magnalium	Aluminium, magnesium	Hard, very light (lighter than Al)	Light instruments, balance beams, in aircraft and automobile parts

SUMMARY

1. Out of 118 elements known, about 80% are metals. Number of metals = 91, number of non-metals = 20.
2. Elements are classified as metals, non-metals and semi-metals.
3. The semiconducting elements Si, Ge, As, Sb and Te are known as semi-metals.
4. Metals are hard except the alkali metals which are soft. Metals possess high tensile strength.
5. The lightest metal is lithium and the densest metal is osmium.
6. Metals exhibit lustre.
7. Metals are malleable and ductile.
8. Metals exhibit high melting and boiling points.
9. Metals are good conductors of heat and electricity.
10. Metals are electropositive in character.
11. **Alloy:** A homogeneous mixture of a metal with other metals or with a metal and a non-metal is called alloy.
12. Metals which appear above hydrogen in the activity series displace hydrogen from water (or steam) and dilute acids. Different metals have different reactivities with water and dilute acids. A more reactive metal displaces a less reactive metal from its salt solution.
13. Metal oxides are basic or amphoteric. The reactivity of metals with oxygen decreases on moving down the activity series.
14. Metals react with chlorine to form ionic metal chlorides.

15. Highly electropositive metals like Na, K, Ca, Mg, Sr, etc. react with hydrogen to form ionic metal hydrides.
16. Metals form positive ions by losing electrons to non-metals.
17. Metals occur in nature as free elements or in the form of their compounds.
18. **Mineral and ore:** The elements or compounds, which occur naturally in the earth's crust, are called minerals. A mineral from which a metal is extracted economically is called ore.
19. **Gangue:** The unwanted earthy material present in an ore is called gangue.
20. **Metallurgy:** The science and technology of metals and alloys employed for the extraction of metals economically on a large-scale from their ores and refining them is called metallurgy.
21. **Concentration of ore:** The process of removal of unwanted materials from an ore is called concentration of ore. The ores are concentrated by lavigation, froth floatation, magnetic separation and leaching.
22. **Calcination:** The process in which the concentrated ore is heated below its melting point in the absence or limited supply of air is called calcination.
23. **Roasting:** The process in which the concentrated ore is heated below its melting point in the presence of excess of air is called roasting.
24. The methods used for reduction of ore to the metal are: a. Reduction of metal oxides by heat alone, b. Reduction of metal oxides using carbon or carbon monoxide, c. Reduction of metal oxides using aluminium, d. Displacement of a less reactive metal by a more reactive metal from its salt solution, e. Reduction of fused salts by electrolysis.

25. **Refining of metals:** The process of removing impurities from the metals extracted from their ores is called refining of metals. Refining of metals is commonly carried out by electrorefining.
26. **Corrosion:** The phenomenon of deterioration or destruction of a metal when exposed to air, water or any other substance around it in the atmosphere is called corrosion.
27. **Rust:** When a piece of iron is exposed to moist air for a long time, a layer of reddish-brown flaky substance is formed on the surface of iron. This substance is called rust. Rust is hydrated iron(III) oxide ($Fe_2O_3 \cdot xH_2O$).
28. Non-metals exhibit properties opposite to that of metals. Unlike metals, they are neither malleable nor ductile. Unlike metals, they are poor conductors of heat and electricity.
29. Non-metals are soft and brittle.
30. Non-metals exhibit low melting and boiling points.
31. Non-metals are electronegative and form anions by accepting electrons when reacting with metals.
32. The oxides of non-metals are acidic or neutral.
33. Non-metals react with chlorine to form covalent and volatile chlorides.
34. Non-metals react with hydrogen to form covalent hydrides.

MIND MAP

METALS AND NON-METALS

METALS

Physical properties
- are solids except mercury
- are lustrous, malleable, sonorous and ductile
- generally possess high density
- high melting and boiling point
- good conductors of heat and electricity

Chemical properties

Reaction with oxygen
- Metal + $O_2 \rightarrow$ Metal oxide
- $4Na + O_2 \rightarrow 2Na_2O$

Reaction with water
- Metal + $H_2O \rightarrow$ Metal oxide + H_2
  $\downarrow H_2O$
  Metal hydroxide
- $Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$

Reaction with acids
- Metal + dilute acid $\rightarrow$ Salt + H_2
- $Mg + 2HCl \rightarrow MgCl_2 + H_2$

Reaction with Cl_2
- $2Na + Cl_2 \xrightarrow{\Delta} 2NaCl$

Reaction with H_2
- $2Na + H_2 \xrightarrow{\Delta} 2NaH$

Reaction with solutions of other metal salts (displacement)
- Metal A + Salt soln. of metal B $\rightarrow$ Salt soln. of metal A + metal B
- $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$

Activity series of metals

Activity series	Reactivity	Electropositive character	Reducing power
K	Decreases ↓	Decreases ↓	Decreases ↓
Na			
Ca			
Mg			
Al			
Zn			
Fe			
Sn			
Pb			
H			
Cu			
Hg			
Ag			
Au			
Pt			

NON-METALS

Physical properties
- are solids or gases except bromine (liquid)
- are not lustrous, malleable, ductile and sonorous
- have low density
- low melting and boiling point
- poor conductors of heat and electricity
- are soft and brittle
- exhibit allotropy

Chemical properties

Reaction with oxygen
- $C + O_2 \rightarrow CO_2$

Reaction with H_2
- $H_2 + Cl_2 \rightarrow 2HCl$

Reaction with Cl_2
- $P_4 + 6Cl_2 \rightarrow 4PCl_3$

MIND MAP

IONIC COMPOUNDS

- formed by the transfer of electrons from a metal to a non-metal
- **Formation of potassium chloride**

$$\underset{2,\,8,\,8,\,1}{K} \longrightarrow \underset{\substack{2,\,8,\,8\\ \text{Potassium cation}}}{K^+} + e^-$$

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

$$\dot{K} + \cdot\ddot{\underset{\cdot\cdot}{Cl}}: \longrightarrow K^+ + :\ddot{\underset{\cdot\cdot}{Cl}}:^-$$

CORROSION

- occurs when metals are exposed to oxygen, moisture, acids, etc.
- occurs at the surface of a metal
- example : $2Ag(s) + H_2S(g) \rightarrow Ag_2S(s) + H_2(g)$
- prevention of corrosion:
  - by galvanisation
  - by painting, oiling, chrome plating, anodising, greasing, etc.

EXTRACTION OF METALS

Ore → Concentration of ore

- Metals of low reactivity → Sulphide ore → Roasting → Metal → Refining of metal
- Metals of moderate reactivity
  - Sulphide ore → Roasting
  - Carbonate ore → Calcination
  - → Metal oxide → Reduction to metal → Refining of metal
- Metals of high reactivity → Electrolysis of molten ore → Pure metal