	INDIAN SCHOOL AL WADI AL KABIR	
Class: X	**DEPARTMENT OF SCIENCE 2023 - 24**	DATE OF COMPLETION: 10.10.2023
HANDOUTS	**CHAPTER- METALS AND NON-METALS**	
CLASS & SEC:	**NAME OF THE STUDENT:**	**ROLL NO.**

CHAPTER-3-METALS AND NON-METALS

PHYSICAL PROPERTIES OF METALS:-

- Metals are hard solids.
  Exception:-Sodium, potassium and lithium are soft metals. Mercury is a liquid metal.
- Metals have metallic lustre.(shining appearance on the new cut surface of metals)
- Metals are malleable. (Metals can be beaten into thin sheets.)
  Gold and silver are the most malleable metals.
- Metals are ductile. (Metals can be drawn into thin wires.)
  Gold is the most ductile metal.
- Metals are sonorous. (Metals produce a ringing sound on striking a hard surface.)
- Metals are good conductors of heat and electricity.
  The best conductors are silver and copper. Lead and mercury are poor conductors of heat.
- Metals have high melting and boiling points.
  Exception:-Gallium and caesium are metals with low melting points.

PHYSICAL PROPERTIES OF NON-METALS:-

- Non-metals are either solids or gases.
  Bromine is a liquid non-metal.
- Non-metals do not have metallic lustre.
  Exception: - Iodine is a non-metal which has metallic lustre.
- Non-metals are non-malleable, non-ductile and non-sonorous.
- Non-metals are bad conductors of heat and electricity.
  Exception: - Graphite is a good conductor of electricity.
- Non-metals have low melting and boiling points.
  (Carbon is a non-metal that can exist in different forms. Diamond, an allotrope of carbon is the hardest natural substance known and has a very high melting and boiling point. Graphite another allotrope of carbon, is a conductor of electricity.)

CHEMICAL PROPERTIES OF METALS: -

1. **REACTION OF METALS WITH OXYGEN**

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

- Metal oxides are basic in nature.

Eg:- $2Cu + O_2 \rightarrow 2CuO$

$4Al + 3O_2 \rightarrow 2Al_2O_3$
(Aluminium oxide)

Some metal oxides show both acidic as well as basic properties. Such metal oxides are known as **amphoteric oxides.** Amphoteric oxides react with both acids as well as bases to produce salt and water.(Eg:- Aluminium oxide and Zinc oxide)

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

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

Different metals show different reactivities towards oxygen.

- Metals like sodium and potassium react vigorously with oxygen and catch fire. So to prevent accidental fires and to protect these metals, they are stored in kerosene.
- Metals like Magnesium, Aluminium, Zinc, lead etc. are covered with a thin layer of oxide. This protective oxide layer prevents the metal from further oxidation.
- Iron does not burn on heating. Iron filings burn vigorously at high temperatures.
- Copper does not burn, but the hot metal is coated with a black coloured layer of copper oxide.
- Least reactive metals like gold, silver etc. do not react with oxygen even at high temperatures.

REACTION OF METALS WITH WATER:-

$$\boxed{\begin{array}{c}\text{Metal + water} \rightarrow \text{metal oxide + hydrogen} \\ \text{Metal oxide + water} \rightarrow \text{Metal hydroxide.}\end{array}}$$

- Metals like sodium and potassium react vigorously even with cold water. The reaction is exothermic and the heat evolved is sufficient for hydrogen gas to catch fire.

$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$ + heat energy

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$ + heat energy

- Calcium reacts less violently with water. The heat evolved is not sufficient for hydrogen to catch fire.

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

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

- Magnesium reacts with hot water. Magnesium starts floating because bubbles of hydrogen gas stick to its surface.
  $Mg + 2H_2O \rightarrow Mg(OH)_2 + H_2$
- Metals like aluminium, iron and zinc react with steam to form metal oxide and hydrogen gas.
  $2Al(s) + 3H_2O(g) \rightarrow Al_2O_3(s) + 3H_2(g)$
  $3Fe(s) + 4H_2O(g) \rightarrow Fe_3O_4(s) + 4H_2(g)$
- Metals like lead, copper, silver and gold do not react with water at all.

REACTION OF METALS WITH ACIDS:-

Metal +dilute acid → salt + hydrogen gas

Eg:- $Zn + 2HCl \rightarrow ZnCl_2 + H_2$

- Hydrogen gas is not evolved when a metal reacts with nitric acid. Nitric acid is a strong oxidising agent. It oxidises the hydrogen produced to water and itself gets reduced to any of the nitrogen oxides(N_2O, NO_2, NO)

- Metals like magnesium and manganese react with very dilute nitric acid to release hydrogen gas.

REACTION OF METALS WITH SOLUTIONS OF OTHER METAL SALTS-(DISPLACEMENT REACTION):-

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

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

Eg:- $Fe + CuSO_4 \rightarrow FeSO_4 + Cu$

THE REACTIVITY SERIES:-

Reactivity series is the arrangement of metals in the decreasing order of reactivity.

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

REACTION OF METALS WITH NON-METALS-(FORMATION OF IONIC COMPOUNDS):-

When **metals react** with **non-metals**, electrons are transferred from the **metal** atoms to the **non-metal** atoms, **forming ions**. (the metal atoms give electrons to the non-metal atoms. The **metal atoms become positive ions** and the **non-metal atoms become negative ions**.)
There is a strong electrostatic force of attraction between these oppositely charged ions – this is called an **ionic bond**. The resulting **compound** is called an **ionic compound**.

ie, the compounds formed by the transfer of electrons from a metal to a non-metal are called ionic compounds or electrovalent compounds.

Formation of NaCl:-

Atomic number of sodium-11, electronic configuration- K L M
2 8 1
Sodium loses one electron from its outermost shell and forms Na^+.
Atomic number of Chlorine is 17, electronic configuration- K L M
2 8 7
Chlorine has seven electrons in its outermost shell and it requires one more electron to complete its octet. It accepts one electron and forms Cl^- .

Sodium and chloride ions, being oppositely charged, attract each other and are held by strong electrostatic force of attraction to exist as Sodium chloride (NaCl)

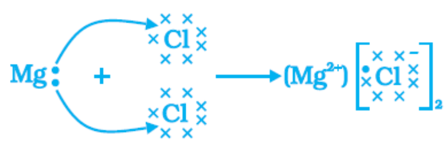
$$\underset{2,8,1}{Na} \rightarrow \underset{\substack{2,8 \\ \text{(Sodium cation)}}}{Na^{+}} + e^{-}$$

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

Formation of $MgCl_2$:-

$$\underset{2,8,2}{Mg} \longrightarrow \underset{\substack{2,8 \\ \text{(Magnesium cation)}}}{Mg^{2+}} + 2e^{-}$$

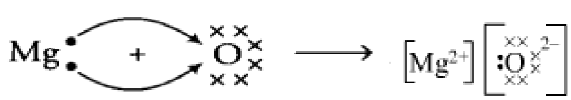
$$\underset{2,8,7}{Cl} + e^{-} \longrightarrow \underset{\substack{2,8,8 \\ \text{(Chloride anion)}}}{Cl^{-}}$$



Formation of MgO:-

$$\underset{[2,8,2]}{Mg} \longrightarrow \underset{[2,8]}{Mg^{2+}} + 2e^{-}$$

$$\underset{[2,6]}{O} + 2e^{-} \longrightarrow \underset{[2,8]}{O^{2-}}$$



PROPERTIES OF IONIC COMPOUNDS:-

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

2. **Melting and boiling points:-** Ionic compounds have high melting and boiling points. This is because a considerable amount of energy is required to break the strong inter-ionic attraction.

3. **Solubility:** - Ionic compounds are generally soluble in water and insoluble in solvents such as kerosene, petrol etc.

4. **Conduction of electricity:** - Ionic compounds in the solid state do not conduct electricity because movement of ions in the solid is not possible due to their rigid structure. But ionic compounds conduct electricity in the molten state and in the solution form as ions are free to move to conduct electricity.

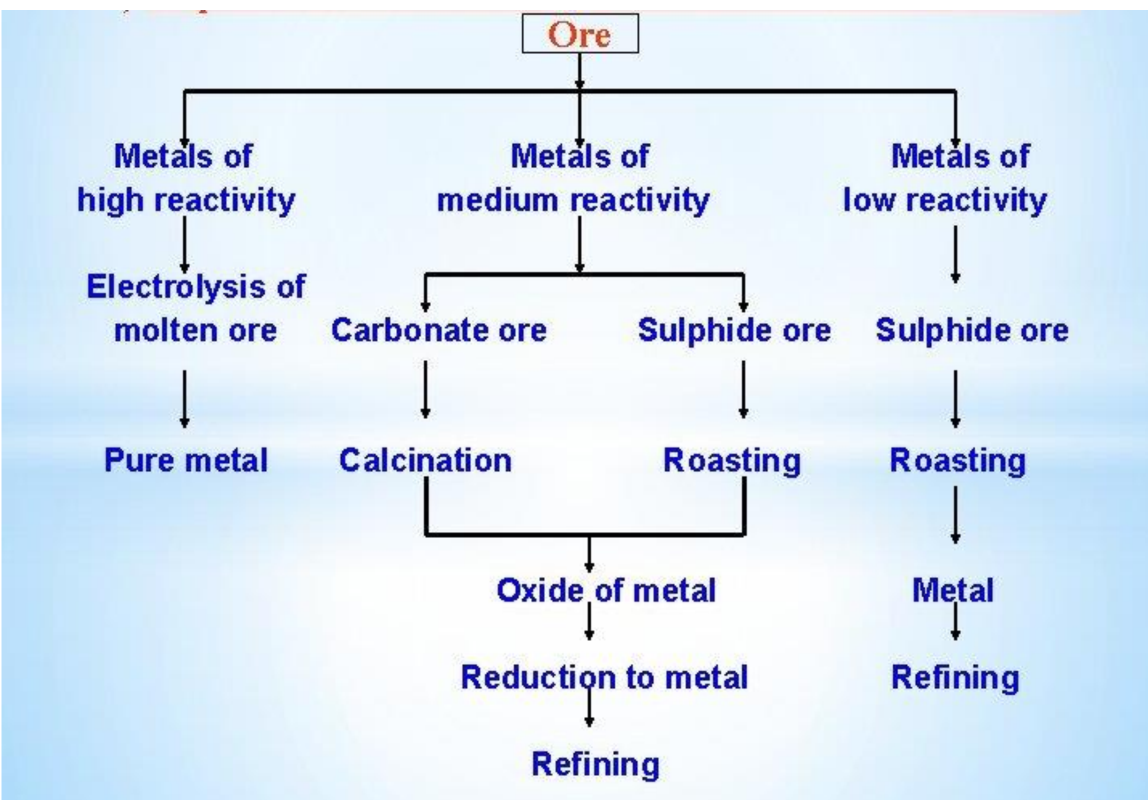
OCCURRENCE OF METALS

The elements or compounds which occur naturally in the earth's crust are known as minerals. The minerals from which we can profitably extract metals are known as ores.

EXTRACTION OF METALS

- Least reactive metals are present in nature in free state.
  Eg:- Gold, Silver, Platinum and Copper.
  Copper and silver are also present in the combined state as their sulphides or oxides.
- Moderately reactive metals are found mainly as oxides, sulphides or carbonates.
- Highly reactive metals like potassium, sodium, magnesium and aluminium are never found in nature in the free state.

Several steps are involved in the extraction of pure metal from ores. A summary of these steps is given in Fig:



ENRICHMENT OF ORES

Ores mined from the earth are usually contaminated with large amounts of impurities such as sand, soil etc. These impurities are called gangue. The removal of impurities from an ore is known as concentration of ore or enrichment of ore.

EXTRACTION OF LEAST REACTIVE METALS

These metals are very unreactive. The oxides of these metals can be reduced to metals by heating.
Eg: When we heat HgS (Cinnabar-an ore of mercury) it is first converted to HgO and then on heating again HgO reduces to Hg.

$$2HgS(s) + 3O_2(g) \xrightarrow{\text{Heat}} 2HgO(s) + 2SO_2(g)$$

$$2HgO(s) \xrightarrow{\text{Heat}} 2Hg(l) + O_2(g)$$

Similarly, copper can be obtained from Cu_2S.

$$2Cu_2S + 3O_2(g) \xrightarrow{\text{Heat}} 2Cu_2O(s) + 2SO_2(g)$$

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

EXTRACTION OF MODERATELY REACTIVE METALS

Moderately reactive metals are present in nature in the form of oxides, sulphides or carbonates. It is easy to obtain a metal from its metal oxide. Before reduction, the metal sulphides and carbonates must be converted into metal oxides.

ROASTING

It is the process of conversion of sulphide ore to oxide ore by heating strongly in excess amount of air.

Eg:-

$$\underset{\text{zinc sunphide}}{2ZnS_{(S)}} + \underset{\text{oxygen}}{3O_{2(g)}} \xrightarrow{\textit{Roasting}} \underset{\text{Zinc Oxide}}{2ZnO} + \underset{\text{Sulphur dioxide}}{2SO_{2(g)}}$$

CALCINATION

It is the process of conversion of carbonate ore to oxide ore by heating strongly in limited air.
Eg:-

$$ZnCO_3 \xrightarrow{\Delta} ZnO + CO_2 \uparrow$$

The metal oxide can be reduced to metal by using a suitable reducing agent such as carbon.

$$ZnO(s) + C(s) \rightarrow Zn(s) + CO(g)$$

Instead of using carbon, displacement reactions can be used to obtain metal from metal oxide.

Eg:-

$$3MnO_2(s) + 4Al(s) \rightarrow 3Mn(l) + 2Al_2O_3(s) + Heat$$

These type of displacement reactions are highly exothermic. Large amount of heat energy is released and the metals are produced in the molten state.

The reaction of iron oxide (Fe_2O_3) with aluminium is used to join railway tracks and cracked machine parts. This reaction is known as Thermit reaction.

$$Fe_2O_3(s) + 2Al(s) \rightarrow 2Fe(l) + Al_2O_3(s) + Heat$$

EXTRACTION OF HIGHLY REACTIVE METALS

Highly reactive metals cannot be obtained from their compounds by heating with carbon. This is because these metals have more affinity for oxygen than carbon.

Highly reactive metals are obtained by electrolytic reduction.

Eg:- Sodium, Magnesium and Calcium are obtained by the electrolysis of their molten chlorides. The metals are deposited at cathode and chlorine gas is liberated at the anode.

At cathode $Na^+ + e^- \rightarrow Na$

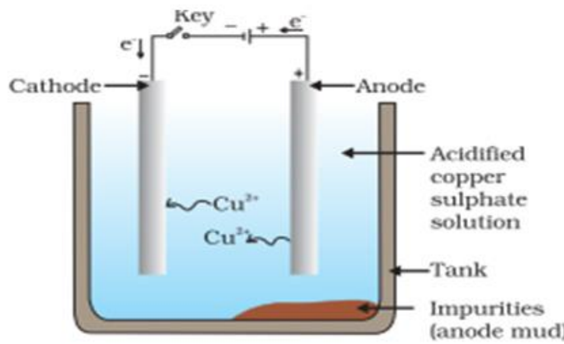
At anode $2Cl^- \rightarrow Cl_2 + 2e^-$

REFINING OF METALS

It is the process of purification of metals produced by various reduction processes. The most widely used refining process is electrolytic refining.

ELECTROLYTIC REFINING

In this process, the impure metal is made as the anode and a thin strip of pure metal is the cathode. A solution of the metal salt is used as an electrolyte. On passing current through the electrolyte, the pure metal from the anode dissolves into the electrolyte. An equivalent amount of pure metal from the electrolyte is deposited on the cathode. The soluble impurities go into the solution and the insoluble impurities settle down at the bottom of the anode and are known as anode mud.



CORROSION

- Silver articles become black when exposed to air. Silver reacts with Sulphur in the air to form silver sulphide.
- When copper articles are exposed to air, it reacts with moist CO_2 to form green-coloured copper carbonate.
- When iron articles are exposed to air it acquires a brown coating called rust.

PREVENTION OF CORROSION

- Rusting of iron can be prevented by painting, oiling, greasing, galvanizing, chrome-plating and alloying.
- Galvanisation is the method of protecting steel and iron from rusting by coating them with a thin layer of zinc.
- Alloying is a very good method to prevent corrosion. Alloys are homogeneous mixtures of two or more metals or a metal with non-metal.

Eg:-

Alloy	**Constituents**
Bronze	Cu, Sn
Brass	Cu, Zn
Steel	Fe, Ni, Cr, C
Solder	Sn, Pb

- ✓ If one of the metals is mercury, then the alloy is known as amalgam.
- ✓ The electrical conductivity and melting point of an alloy is less than that of pure metals.

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