METALS AND NON METALS

On the basis of similarities and difference among 115 chemical element, all the elements can be divided into two main groups: metals and non-metals.

Topic : Metals & Their Physical Properties

Definition 1: Those elements which are malleable, ductile, sonours and good conductor of heat and electricity are called metals.

Definition 2: Those elements which form positive ions by losing electrons from their outermost shell are called metals. For example sodium loses one electron from their outermost shell to form sodium ions.

$$\underset{2,\ 8,\ 1}{Na - 1e^-} \longrightarrow \underset{2,\ 8}{Na^+}$$

Definition 3: Those elements which have 1 or 2 or 3 electrons in their outer most shell (except hydrogen and helium) are called metals. All the metals are solid, except **mercury**, which is a liquid metal at room temperature.

☞ **Gallium** and **Cesium** are two metals which melts in our hands at room temperature.

Occurance of metsals in the earth crust: The most abundant metal in the earth's crust is aluminium. It is 7% of earth's crust. The second most abundant metal in the earth crust's is iron, it is about 4% the earth crust. The major metals in the earth crust in the decreasing order of their abundance are. Aluminium, Iron, calcium sodium potassium and magnesium.

$$A\ell > Fe > Ca > Na > K > Mg$$

Physicsal Properties of Metals

1. **Metals are malleable:** Metals are malleable that means they can be beaten into thin sheet when hammered, this property of metal is called **mallelability**

   ☞ **Gold and silver** are best malleable metals due to this both are used to make ornaments and Silver foils are also used for decorating sweets.

   ☞ Aluminium metal also good melleable metal due to this aluminium foils are used for packing food items like Biscuits, Chocolets and medicines, cigrats, milk bottle cap etc.

   ☞ Iron is also quite malleable due to this Iron sheets are used to make boxes, buckets, drums and water tank etc.

2. **Metals are sonorous,:** Metals are sonorous i.e., they make sound when hit with an object. This property of metal is called **sonority** due to this they are used for making bells and strings of musical instruments like sitar and violin.

Activity – 3.1

Aim: To show that metals are malleable and sonorous.

Material required: Take pieces of iron, zinc, lead , copper, hammer, block of iron.

Procedure:

Step 1: Take pieces of iron, zinc, lead and copper.

Step 2: Place any one metal on a block of iron.

Step 3: Strike the metal four or five times with a hammer.

Step 4: Note the observation.

Observation: Metals produce sound when struck with hard substance. They can be beaten into sheets when hammered hard for long time.

Conclusion: Metals are sonorous and most of them can be beaten into sheet. This property is called malleability.

2. **Metals are ductile:** Metals are ductile that means they can be drawn into wires. This property of metal is called **ductility**.

- Gold and silver are most ductile metal, i.e., (1gm of gold can be drawn into 2km long wire). Due to high ductility both are used to make ornaments.
- Copper and aluminium are also very ductile metals due to this both are used to make electric wires.
- Iron, magnesium and tungsten metal are also quite ductile due to this iron wires are used for making wire gauzes. Magnesium wires are used in science experiment and tungsten wires are used for making the filament of electric bulbs.

Activity – 3.2

Aim: To show that metals are ductile i.e., they can be drawn into wires.

Material required: Iron, copper, aluminium, lead, sodium, potassium.

Procedure:

Step 1: Observe various metals.

Step 2: Note down which of the metals can be drawn into wires and those cannot be drawn into wires.

Observation: Iron, copper, aluminium can be drawn into wires and lead, sodium, potassium cannot be drawn into wires.

Conclusion: Most of the metals are ductile except lead, sodium and potassium.

3. **Metals are good conductors of heat:** Metals are good conductor of heat i.e., they allow heat to pass through them easily. Thir property of metals is called **thermal conductivity**.
   - **Silver** is the best conductor of heat.
   - Copper and aluminium metals are also very good conductor of heat , due to this cooking utensils and water boilers are made up of copper and aluminium metals.
   - **Lead and Mercury** are the poor conductor of heat but **lead** is the poorest conductor of heat.

Activity – 3.3

Aim: To show that metals are good conductor of heat.

Material required: Aluminium or copper wire, a pin, wax, stand, burner, etc.

Procedure:

Step 1: Take an aluminium or copper wire.

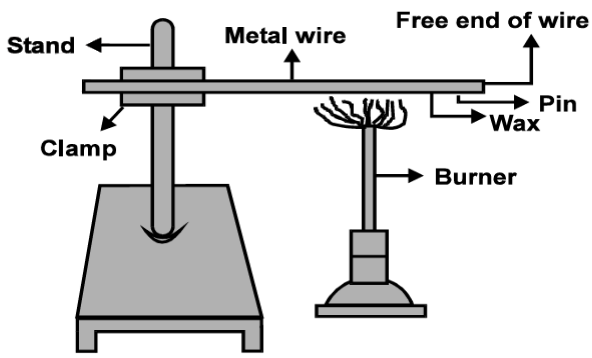
Step 2: Clamp this wire on a stand.

Step 3: Fix a pin to the free end of the wire using wax.

Step 4: Heat the wire with a burner near the place where it is clamped.

Step 5: Note the observation.

Observation: The wax melts and pin falls down.



Conclusion: Metals are good conductors of heat.

4. **Metals are good conductors of electricity:** Metals are good conductor of electricity that is they allow electricity to pass through them easily. This property of metal is called **electrical conductivity.**
   - **Silver** is the best conductor of electricity. Copper is next and is followed by gold, aluminium and tungsten.
   - Electric wires are made of copper and aluminium because they are good conductor of electricity.

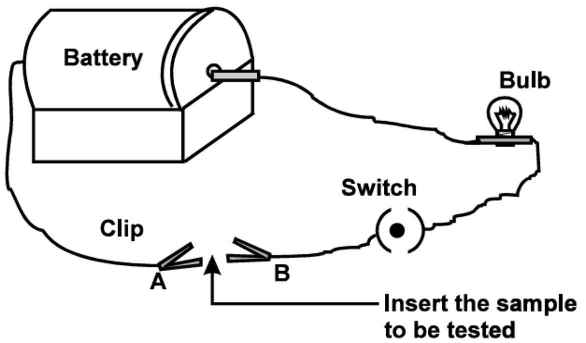
Activity – 3.4

Aim: To show that metals aer good conductors of electricity.

Material required: Au, Ag, Mg, Fe, Zn, Cu etc. battery, bulb, clips, switch, connecting wires.

Procecudure:

Step 1: Set up an electric circuit as shown in the figure.



Step 2: Place the metal to be tested in the circuit between terminals A and B.

Step 3: Note the observations.

Observation: Bulb glows in case of all the metals tested.

Conclusion: Metals are good conductor of electricity.

5. **Metals are lustrous:** Metals are lustrous (or shiny) i.e., they can be polished,
   - ☞ Gold, silver and copper are shiny metals this is because they are used in making jewellery and decorative items.
   - ☞ **Note:** When metals kept in air for long time. They lose their brightness and acquire dull appearance due to formation of thin layer of oxide, carbonate or sulphide on their surface.

Activity – 3.5

Aim: To show that metals are lustrous.

Materials required: Iron, copper, magnesium and aluminium.

Procedure:

Step 1: Take samples of iron, copper aluminium and magnesium.

Step 2: Clean the surface of each sample by rubbing them with sand paper.

Step 3: Note the appearance (observation).

Observation s:

Iron: Greyish black metal.**Copper:** Reddish brown.

Magnesium: Silvery white **Aluminium:** Silvery white.

Conclusion: Metals taken as samples have shining surface or lustrous.

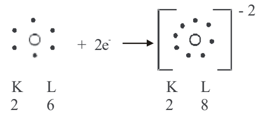
6. **Metals are hard:** Metals are generally hard except sodium and potassium which can be cut by knife.
7. **Metals are solids at room temperature:** Metals are solid at room temperature except mercury which is liquid.
8. **Metals have high melting and boiling point** : Metals have high melting point & boiling point except sodium, potassium, gallium and cesium.
9. **Metals have high densities:** Metals have densities excepts sodium, potassium, cesium and gallium. For Ex. The density iron is 7.8 g/ cm^3 and density of sodium is 0.97 g/cm^3.
10. **Metals are grey in colour:** Metals are generally grey in colour except copper and gold. Copper is reddish brown colour and gold is yellow colour.

☞ Some metals such as titanium (Ti), chromium (Cr), manganese (Mn), and zirconium (Zr), are called strategic metals because these metals and their alloys are used in atomic energy, space science projects, jet engines, etc. They are also used in making high grade steel.

Topic : Non-Metals & Their Physical Properties

Definition 1: Those elements which are normalleable, nor-ductile, non-sonorous and bad conductor of heat and electricity but they are brittle are called non-metals.

Definition 2: Non metals are those elements which form negative ions (anion) by gaining electron from another atom. For example oxygen atom gain 2 electrons from another atom to form oxide ion.



Definition 3: The elements having 4 or 5 or 6 or 7 or 8 electrons in their outer most shell are called non-metals (except hydrogen and helium)

Occurance of Non-metal in the earth crust: The most abudant non metal in the earth crust's is oxygen (50%). The second most abundant metal in the earth crust is silicon (26%). The major non-metals in the earth crust inthe decreasing order are oxygen, silicon, phosphorus and sulphur. Non-metals are the major constituents of earth, air and oceans. Two non-metals oxygen and nitrogen are the main constituents of air; two non metals oxygen and silicon are the main constituents of earth and two non-metals hydrogen and oxygen are the main constituents of aceans (in the form of water).

<u>Physicsal Properties of Non Metals</u>

1. **Non - metals are brittle:** Non - metals are brittle i.e., they can be broken into small pieces on hammering.
2. **Non - metals are bad conductor of heat & electricity:** Non - metals are bad conductors of heat & electricity, this is because they have no free electrons. Graphite is the only non - metal which is a good conductor of electricity, due to this graphite is used for making electrodes.
3. **Non-metals are not lustrous** (not shiny). Non - metals are not lustrous except Iodine.
4. **Non-metals are generally soft:** Non - metals are generally soft except diamond, infact diamond is the hardest substance in the nature.
5. **Non-metals are not strong:** Non - metals are not strong i.e., they are easily broken.
6. **Non-metals are solid, liquid or gases at the room temperature.** Bromine is only one non-metal which is liquid at the room temperature.
7. **Non-metals have comparatively low melting points and boiling points** except diamond, graphite and boron.
8. **Non-metals have low densities.** Non-metals have low densities. except diamond and graphite i.e., non-metals are light substances.
9. **Non-metals are non-sonorous.** i.e., they do not produce sound when hit with an object.
10. **Non - metals are different in colours,** for example sulphur is yellow, phosphorus is white, yellow or red, graphite is black, chlorine is yellowish green, bromine is reddish brown while hydrogen, oxygen and nitrogen are colourless.

Importance of Non metals: Although the number of non metals are very less as compared to metals, yet they play an important role in our daily life. For example,

1. Carbon is one of the most important non metal because all the life on this earth is made of carbon compounds.For example, Carbohydrates, Proteins, Oils & Fats, Vitamins, Enzymes etc which are made up of carbon compounds, are essential for development, growth and maintainance of living organisms. Carbon, in the form of graphite, is also used as electrodes in electrolytic cells and dry cells.
2. Another important non metal is hydrogen. It is present in almost all the compounds of carbon which are required for growth and develpoment of fat. it is used in hydrogenation of vegetable oils to make vegetal Ghee. It is also used in the manufacture of ammonia which is used as a refrigerant and also for making fertilizers.
3. Oxygen is very important for existence of life and combustion process.
4. Nitrogen is very important non metal. Its presence in air reduces the rate of combustion. Nitrogen is mainly used in the manufacture of ammonia, nitric acid, fertilizers, tri nitro tolune (used as explosive).

1. Reaction of metals with oxygen of air:

- Metals react with oxygen to form metal oxides. Metal oxides are basic in nature, which turn red litmus solution blue.

Concept:

$$\text{Metal} + \text{Oxygen} \longrightarrow \underset{\text{Basic Oxide}}{\text{Metal oxide}}$$

Ex.1 Sodium, Potassium metals reacts with the oxygen of air at room temperature to form a basic oxide called sodium oxide and potassium oxide.

(a) $$\underset{\substack{\text{Sodium}\\ \text{(metal)}}}{4Na(s)} + \underset{\substack{\text{Oxygen}\\ \text{(from air)}}}{O_2(g)} \longrightarrow \underset{\substack{\text{Sodium oxide}\\ \text{(basic oxide)}}}{2Na_2O(s)}$$

(b) $$\underset{\substack{\text{Potassium}\\ \text{(Metal)}}}{4K} + \underset{\substack{\text{Oxygen}\\ \text{(from air)}}}{O_2} \longrightarrow \underset{\substack{\text{Potassium oxide}\\ \text{Basic Oxide}}}{2K_2O}$$

☞: Potassium and sodium metal are stored under kerosene oil to prevent their reaction with the oxygen, moisture and carbon dioxide of air.

Note: Most of the metal oxides are insoluble in water. But some of the metal oxides soluble in water to form alkalies. **For example:** Sodium oxide and potasium oxide are soluble in water and form Alkalies called sodium Hydroxide and Potassium Hydroxide.

$$\underset{\substack{\text{Sodium Oxide}\\ \text{(basic oxide)}}}{Na_2O(s)} + \underset{\text{Water}}{H_2O(\ell)} \longrightarrow \underset{\substack{\text{Sodium hydroxide}\\ \text{(an alkali)}}}{2NaOH(aq)}$$

$$\underset{\substack{\text{Potassium oxide}\\ \text{(basic oxide)}}}{K_2O(s)} + \underset{\text{Water}}{H_2O(\ell)} \longrightarrow \underset{\substack{\text{Potassium}\\ \text{Hydroxide}}}{2KOH\ (aq)}$$

Activity – 3.6

Aim: To test the nature of oxides formed by metals

Material required: Magnesium ribbon, red litmus paper, water, test tube.

Procedure:

Step 1: Take a magnesium ribbon and sulphur powder.

Step 2: Burn the magnesium ribbon in the presence of oxygen of air.

Step 3: Collect the ashes formed and dissolve them in water.

Step 4: Add red litmus paper into it.

Step 5: Note your observation.

Observations: The oxide formed by metal turns red litmus to blue.

Chemical reactions:

$$2Mg(s) + O_2 \xrightarrow{\text{burning}} 2MgO(s)$$

$$\underset{\text{(Basic oxide)}}{MgO(s)} + H_2O(l) \longrightarrow Mg(OH)_2(aq)$$

Conclusion: Most of metallic oxides are basic in nature.

Ex. 2 Magnesium metal does not react with oxygen at room temperature. But on heating magnesium metal burns in air and give intense heat and light to form a basic oxide called magnesium oxide (which is a white powder):

$$\underset{\substack{\text{Magnesium}\\ \text{(metal)}}}{2Mg(s)} + \underset{\substack{\text{Oxygen}\\ \text{(from air)}}}{O_2(g)} \longrightarrow \underset{\substack{\text{Magnesium oxide}\\ \text{(basic oxide)}}}{2MgO}$$

$$\underset{\substack{\text{Magnesium}\\ \text{oxide}}}{MgO(s)} + \underset{\text{Water}}{H_2O(\ell)} \longrightarrow \underset{\substack{\text{Magnesium Hydroxide}\\ \text{(a base)}}}{Mg(OH)_2(aq)}$$

Ex. 3 Aluminium metal burns in air, on heating to form aluminium oxide:

$$\underset{\substack{\text{Alu minium}\\ \text{(metal)}}}{4Al(s)} + \underset{\substack{\text{Oxygen}\\ \text{(from air)}}}{3O_2(g)} \longrightarrow \underset{\substack{\text{Aluminium oxide}\\ \text{(amphoteric oxide)}}}{2Al_2O_3(s)}$$

Ex. 4 Zinc metal burns in air only on strong heating to form zinc oxide:

$$\underset{\text{Zinc}}{2Zn(s)} + \underset{\text{Oxygen}}{O_2(g)} \longrightarrow \underset{\text{Zinc Oxide}}{2ZnO(s)}$$

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

Ex. 5 Iron metal does not burn in air even on heating. It reacts with the oxygen of air on very strong heating to form iron (II, III) oxides:

$$\underset{\text{Iron}}{3Fe(s)} + \underset{\text{Oxygen}}{2O_2(g)} \longrightarrow \underset{\substack{\text{Iron (II, III)}\\ \text{Oxide}}}{Fe_3O_4(s)}$$

Ex. 6 Copper metal also does not burn in air even on strong heating. Copper reacts with the oxygen of air on prolonged heating to form a black substance copper (II) oxide:

$$\underset{\text{Copper}}{2Cu(s)} + \underset{\text{Oxygen}}{O_2(g)} \longrightarrow \underset{\text{Copper(II) oxide}}{2CuO(s)} \text{ (Black)}$$

☞ Gold, silver mercury and platinum do not react with oxygen of air even at a high temperature.

Amphoteric Oxides

Those metal oxides which shows basic as well as acidic behaviour are known as amphoteric oxide. Aluminium oxide and zinc oxides are the examples of amphoteric oxide.

a. Case of Aluminium Oxide (Al_2O_3):

(i) Aluminium oxide reacts with Di--Hydrochloric acid to form aluminium chloride and water.

$$\underset{\text{Aluminium Oxide}}{Al_2O_3(s)} + \underset{\text{Hydrochloric acid}}{6HCl(aq)} \rightarrow \underset{\text{Aluminium chloride (salt)}}{2AlCl_3(aq)} + \underset{\text{Water}}{3H_2O(\ell)}$$

In the above reation Al_2O_3 acts as a base.

(ii) Aluminium oxide reacts with sodium hydroxide to form sodium aluminate and water.

$$\underset{\text{Aluminium oxide}}{Al_2O_3(s)} + \underset{\text{Sodium hydroxide (base)}}{2NaOH(aq)} \rightarrow \underset{\text{Sodium aluminate (salt)}}{2NaAlO_2(aq)} + \underset{\text{Water}}{H_2O(\ell)}$$

In the above reation Al_2O_3 acts as a acid.

b. Case of zinc oxide (ZnO:

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

$$\underset{\text{Zinc oxide}}{ZnO(s)} + \underset{\text{Hydrochloric acid (Acid)}}{2HCl(aq)} \longrightarrow \underset{\text{Zinc chloride (Salt)}}{ZnCl_2(aq)} + \underset{\text{Water}}{H_2O(1)}$$

In the above Reaction ZnO acts as a base.

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

$$\underset{\text{Zinc oxide}}{ZnO(s)} + \underset{\text{Sodium hydroxide (Base)}}{2NaOH(aq)} \longrightarrow \underset{\text{Sodium zincate (Salt)}}{Na_2ZnO_2(aq)} + \underset{\text{Water}}{H_2O(1)}$$

In the above Reaction ZnO acts as a acid.

2. Reaction of metals with water:

- Metal react with water to form a metal hydroxide (or metal oxide) and hydrogen gas.

Concept 1: Metals react with cold or hot water to form metal hydroxide and hydrogen gas.

(a) Metal + water (cold or hod) → Metal hydroxide + Hydrogen gas

Concept 2: Metals react with steam to form metal oxide and hydrogen gas.

(b) Metal + Steam → Metal oxide + Hydrogen gas

Ex. 1 Potassium, Sodium reacts violently with cold water to form potassium hydroxide and sodium hydroxide and hydrogen gas.

(a) $2K + 2H_2O \longrightarrow 2KOH + H_2 + \text{Heat}$

(b) $2Na + 2H_2O \longrightarrow 2NaOH + H_2 + \text{Heat}$

Ex. 2 Calcium reacts less violently with cold water to form calcium hydroxide and hydrogen gas:

$$Ca + 2H_2O \longrightarrow Ca(OH)_2 + H_2$$

Ex. 3 Magnesium metal does not react with cold water. Magnesium reacts with hot water to form magnesium hydroxide and hydrogen gas:

$$Mg + 2H_2O \rightarrow Mg(OH)_2 + H_2$$

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

$$Mg + 2H_2O \rightarrow MgO + H_2$$

Ex. 4 Aluminium reacts with steam to form aluminium oxide and hydrogen gas:

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

Ex. 5 Zinc reacts with steam to form zinc oxide and hydrogen gas.

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

Ex. 6 Red-hot iron reacts with steam to form iron (II, III) oxide and hydrogen gas:

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

Ex.7 Metals like copper, silver, gold, mercury and platinum do not react with even steam

$Cu + H_2O \rightarrow$ No reaction

☞ Above metals do not react with water because they can not reduce hydrogen ion present in water to the hydrogen atom because they are less reactive then hydrogen.

3. Reaction of Metals with Dilute Acids:

Metals react with dilute acids to form metal salt and Hydrogen gas.

Concept:

Metal + Dilute acid → Metal Salt + Hydrogen

Ex.1 **Sodium and potassium metals** reacts violently with dilute hydrochloric acid to form sodium chloride and hydrogen gas.

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

☞ Calcium also react vigrously with acid but less than sodium and potassium

Ex.2 **Magnesium** reacts quite rapidly with dilute hydrochloric acid to form magnesium chloride and hydrogen gas:

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

Ex.3 **Aluminium metal** first reacts slowly with dilute hydrochloric acid due to the presence of layer of aluminium oxide. But when the layer gets dissolved in acid, then fresh aluminium react rapidly with dilute hydrochloric acid to form aluminmum chloride and hydrogen gas:

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

Ex.4 **Zinc** reacts slowly with dilute hydrochloric acid to form zinc chloride and hydrogen gas :

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

Ex.5 **Iron** reacts very solwly with dilute hydrochloric acid to give iron (II) chloride and hydrogen gas :

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

Ex.6 **Copper, silver, platinium, mercury and gold** do not react with dilute hydrochloric acid (or dilute sulphuric acid) at all.

$$\underset{\text{Copper}}{Cu(s)} + \underset{\text{Hydrochloric acid}}{HCl(aq)} \rightarrow \text{No reaction}$$

☞ **:** Copper, silver, mercury, platinum and gold do not react with dilute acid because they are less reactive then hydrogen.

• • • **Reaction of Metals with Nitric Acid:**

Since Nitric acid is strong oxidising agent when Nitric acid react with metal to form Hydrogen gas. Dilute Nitric acid Oxidises. This Hydrogen gas to water but very very dilute nitric acid, reacts with magnesium and maganese metals to form hydrogen gas.

Ex.7 **Magnesium** reacts with very dilute nitric acid to form magnesium nitrate and hydrogen gas:

$$\underset{\text{Magnesium}}{Mg} + \underset{\text{Nitric acid}}{2HNO_3} \rightarrow \underset{\text{Magneium Nitrate}}{Mg(NO_3)_2} + \underset{\text{Hydrogen gas}}{H_2}$$

Ex.8 **Maganese** reacts with very dilute nitric acid to form magnese nitrate and hydrogen gas:

$$\underset{\text{Manganese}}{Mn(s)} + \underset{\text{Nitric acid (very dilute)}}{2HNO_3(aq)} \longrightarrow \underset{\text{Manganese Nitrate}}{Mn(NO_3)_2} + \underset{\text{Hydrogen gas}}{H_2}$$

Aqua-Regia: Aqua-regia is a freshly prepared mixture of 1 part of concentrated nitric acid and 3 parts of concentrated hydrochloric acid. Aqua-regia can dissolve all metals, even gold and platinum metals. The latin name of aqua - regia is **royal water.**

Reactivity Series of Metals

The arrangement of metals in a vertical column according to decreasing. Their reactivities is called reactivity series of metals (or activity series of metals).

Reactivity Series of some metals

Nature of Reactivity	Name of Metal	Symbol	
Metals more reactive than hydrogen	Potassium	K	(Most reactive metal)
	Sodium	Na	REACTIVITY DECREASES
	Calcium	Ca	
	Aluminum	Al	
	Magnesium	Mg	
	Zinc	Zn	
	Iron	Fe	
	Nickel	Ni	
	Tin	Sn	
	Lead	Pb	
Metals less reactive than hydrogen	Copper	Cu	
	Mercury	Hg	
	Silver	Ag	
	Gold	Au	
	Platinum	Pt	Least reactive metal

- **Why some metals are more reactive and other less reactive:** The reactivity of metals depends upon how fast a metal loses electron from their outermost shell to form positive ions. If a metal atom losses electrons from their outer most shell rapidly than it is more reactive in comparison to other metals which loses electrons from their outer most shell slowly.

- **Why some metals are more reactive than hydrogen while others are less reactive:** Metals which can loose electron/electrons more easily than hydrogen are more reactive than hydrogen. Hence, they can displace hydrogen from the acid. Metals which can loose electrons less radily than hydrogen are less reactive than hydrogen. Hence they cannot displace hydrogen from the acid.

4. Reaction of metals with salt solutions:

Highly reactive metals displaces less reactive metals from their salt solutions.

Concept:

Salt solution of less reactive metal + more reactive metal $\rightarrow$ salt solution of more reactive metal + less reactive metal.

Ex.1 When a strip of zinc metal is placed in copper sulphate solution, then blue colour of copper sulphate becomes colourless due to formation zinc sulphate solution, and red-brown copper metal is deposited on the zinc strip and zinc strip becomes reddish brown due to formation not larger of copper metal.

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

Ex.2 Reaction of iron with Copper Sulphate: When a strip of iron metal is placed in copper sulphate solution, then the blue colour of copper sulphate solution turns in to dirty green due to formation of a ferrous sulphate and Iron red become reddish brown due to formation of a layer of copper metal on it.

$$\underset{\text{Blue colour}}{CuSO_4(aq)} + \underset{\text{Iron}}{Fe(s)} \rightarrow \underset{\text{dirtygreen}}{FeSO_4(aq)} + \underset{\text{(raddishbrown)}}{Cu(s)}$$

Note: If a strip of copper metal is placed in iron (II) sulphate solution, then no reaction occurs. Because copper is less reactive than iron.

Ex.3 Reaction of copper metal with silver Nitrate Solution; when a strip of copper metal is placed in a silver Nitrate Solution. The solution becomes gradually blue due to formation of copper Nitrate $Cu(NO_3)_2$ solution and copper strips becomes greyish white due to formation of layer of a silver metal on the copper metal.

$$\underset{\text{Silver Nitrate (colourless)}}{2AgNO_3} + \underset{\text{(Copper)}}{Cu} \rightarrow \underset{\text{Copper Nitrate (Bluecolour)}}{Cu(NO_3)_2} + \underset{\text{(greyish white)}}{2Ag}$$

5. Reaction of metals with Chlorine:

Metals react with chlorine to form ionic metal chlorides.

Concept:

Metal + chlorine $\rightarrow$ ionic metal chloride

Ex. 1 Sodium and potassium metals readily reacts with chlorine to form an ionic chloride called sodium chloride and potassium chloride.

$$\underset{\text{Sodium}}{2Na(s)} + \underset{\text{Chlorine}}{Cl_2(g)} \rightarrow \underset{\text{Sodiumchloride}}{2NaCl(s)}$$

$$\underset{\text{Potassium}}{2K(s)}\underset{\text{Chlorine}}{Cl_2(q)} \longrightarrow \underset{\text{Potassium Chloride}}{2KCl(s)}$$

Ex. 2 Magnesium reacts with chlorine on heating to form magnesium chloride.

$$\underset{\text{Magnesium}}{Mg(s)} + \underset{\text{Chlorine}}{Cl_2(g)} \rightarrow \underset{\text{MagnesiumChloride}}{MgCl_2(g)}$$

Ex. 3 Aluminium reacts with chlorine, on heating, to form aluminium chloride:

$$\underset{\text{Aluminium}}{2Al(s)} + \underset{\text{Chlorine}}{3Cl_2(g)} \rightarrow \underset{\text{Aluminium Chloride}}{2AlCl_3(s)}$$

Ex. 4 Zinc combines directly with chlorine to form zinc chloride:

$$\underset{\text{Zinc}}{Zn(s)} + \underset{\text{Chlorine}}{Cl_2(g)} \rightarrow \underset{\text{Zinc chloride}}{ZnCl_2(s)}$$

Ex. 5 Iron react with chlorine, on heating to form iron (III) chloride:

$$\underset{\text{Iron}}{2Fe(s)} + \underset{\text{Chlorine}}{3Cl_2(g)} \rightarrow \underset{\text{Iron Chloride}}{2FeCl_3(s)}$$

Ex. 6 Copper reacts with chlorine to form copper (II) chloride:

$$\underset{\text{Copper}}{Cu(s)} + \underset{\text{Chlorine}}{Cl_2(g)} \rightarrow \underset{\text{Copper (II) chloride}}{CuCl_2(s)}$$

6. Reaction of metals with Hydrogen:

Most of the metals do not combine with hydrogen. Only a few reactive metals like sodium, potassium, calcium and magnesium react with hydrogen to form ionic metal Hydrides.

Concept:

Metal + Hydrogen → ionic metal hydride

Ex. 1 When hydrogen gas is passed over heated sodium, then sodium hydride is formed:

$$\underset{\text{Sodium}}{2Na(s)} + \underset{\text{Hydrogen}}{H_2(g)} \rightarrow \underset{\text{Sodium hydride}}{2NaH(s)}$$

Ex. 2 When hydrogen gas is passed over heated lithium, then lithium hydride is formed

$$\underset{\text{(Lithium)}}{2Li} + \underset{\text{(Hydrogen)}}{H_2} \longrightarrow \underset{\text{(Lithium Hydride)}}{2LiH}$$

Ex. 3 When hydrogen gas is passed over heated potassium, then potassium hydride is formed

$$\underset{\text{(Potassium)}}{2K} + \underset{\text{(Hydrogen)}}{H_2} \longrightarrow \underset{\text{(Potassium Hydride)}}{2KH}$$

Ex. 4 When hydrogen gas is passed over heated calcium, then calcium hydride is formed:

$$\underset{\text{Calcium}}{Ca(s)} + \underset{\text{Hydrogen}}{H_2(g)} \rightarrow \underset{\text{Calcium hydride}}{CaH_2(s)}$$

Ex.5 When hydrogen gas is passed over heated Aluminium, then Aluminium Hydride is formed:

$$\underset{\text{Aluminium}}{Al} + \underset{\text{Hydrogen}}{H_2} \longrightarrow \underset{\text{Aluminium Hydride}}{AlH_3}$$

Topic : Chemical Properties of Non-Metal

1. Reaction of Non-metals with oxygen:

Non-metal react with oxygen of air to form acidic oxide or neutral oxide. Acidic oxide tunrs blue litmus solution to red and nutral oxide do not have any effects on any type litmus solution. Non-metal oxide are called acidic oxide. This is because when they dissolve in water to form acid.

Concept:

Non-metal + oxygen → acidic oxide or neutral oxide

Ex. 1 When carbon burns in air it reacts with the oxygen of air to form on acidic oxide called carbon dioxide:

$$\underset{\text{(non–metal)}}{C(s)} + \underset{\text{(from air)}}{O_2(g)} \rightarrow \underset{\text{(acidic oxide)}}{CO_2(g)}$$

When, carbon dioxide dissolve in water to form an acid called carbonic acid:

$$\underset{\text{Carbondioxide}}{CO_2(g)} + \underset{\text{Water}}{H_2O(\ell)} \rightarrow \underset{\text{Carbonic acid}}{H_2CO_3(aq)}$$

Activity – 3.7

Aim: To test the nature of oxides formed by non-metals.

Materials required: Magnesium ribbon, sulphur powder, blue litmus paper, water and test tube.

Procedure:

Step 1: Burn sulphur powder in the presence of oxygen of air

Step 2: Place the test tube over the burning sulphur to collect the fumes produced.

Step 3: Add some water to the test tube and shake well of the test tube.

Step 4: Add blue litmus paper into test tube.

Step 5: Note your observations.

Observations: The oxide formed by non-metal turns blue litmus to red.

Chemical reactions:

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

$$\underset{\text{(Acidic oxide)}}{SO_2} + H_2O \longrightarrow \underset{\text{(Sulphurous Acid)}}{H_2SO_3}$$

Conclusion: Most of non-metallic oxides are acidic in nature.

Ex. 2 When sulphur is burn in air, it reacts with the oxygen of air to form an acidic oxide called sulphur dioxide.

$$\underset{\substack{\text{Sulphur}\\\text{(non–metal)}}}{S(s)} + \underset{\substack{\text{Oxygen}\\\text{(from air)}}}{O_2} \rightarrow \underset{\substack{\text{Sulphur dioxide}\\\text{(acidic oxide)}}}{SO_2(g)}$$

When, sulphur dioxide, dissolves in water to form an acid called sulphurous acid:

$$\underset{\substack{\text{Sulphur dioxide}\\\text{(acidic oxide)}}}{SO_2(g)} + \underset{\text{Water}}{H_2(\ell)} \rightarrow \underset{\substack{\text{Sulphurous acid}\\\text{(an acid)}}}{H_2SO_3(aq)}$$

Neutral oxide: Those non-metal oxide which have no effects on any types of litmus solution carbon monoxide (CO), Water (H_2O), Nitric oxide (NO) and Nitrious oxide (N_2O) are exampe of Neutra oxide.

2. **Reaction of Non-metals with water:**

Non metals do not react with water at any condition

Concept:

Non metal + Water $\rightarrow$ No reaction

Reason: Non metals do not react with water, this is because non metals are electron acceptor. Hence they cannot give electrons to hydrogen ions present in water to reduce hydrogen atoms. So, they do not react with water.

3. **Reaction of Non-metals with Dilute Acids:**

Non metal do not react with acids at any condition

Concept:

Non metal + dil. acid $\rightarrow$ No reaction

Reason: Non-metals do not react with dilute acids to produce hydrogen gas. This is because non-metal cannot give electron to hydrogen ions present in dilute acid to reduce hydrogen atom. This is because non-metals are electron acceptor.

4. **Reaction of Non-metals with salt solution**

A more reactive non-metal displaces a less reactive non-metal form their salt solution.

Concept:

Salt solution of less reactive non - metal + more reactive non-metal $\rightarrow$ salt solution of more reactive non-metal + less reactive non-metal

Ex. 1 When chlorine gas is passed through sodium bromide solution then chlorine displace bromine from sodium bromide solution to form sodium chloride solution and bromine gas.

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

5. **Reaction of Non-metals with chlorine:**

Non-metals react with chlorine to form covalent non metal chlorides.

Concept:

Non-metal + chlorine $\rightarrow$ co-valent non-metal chloride

Ex.1 Hydrogen reacts with chloride to form a covalent hydride called hydrogen chloride.

$$\underset{\substack{\text{Hydrogen}\\\text{(non–metal)}}}{H_2(g)} + \underset{\text{Chlorine}}{Cl_2(g)} \rightarrow \underset{\substack{\text{Hydrogen Chloride}\\\text{(Covalent chloride)}}}{2HCl(g)}$$

Ex. 2 Phosphorus reacts with chlorine to form a covalent chloride called phosphorus trichloride:

$$\underset{\substack{\text{Phosphorus}\\\text{(non–metal)}}}{P_4(s)} + \underset{\text{Chlorine}}{6C\ell_2(g)} \rightarrow \underset{\substack{\text{Phosphorus+ trichloride}\\\text{(covalent chloride)}}}{4PCl_3(\ell)}$$

6. **Reaction of Non-metals with Hydrogen:**

Non-metals react with hydrogen to form covalent non-metal hydrides.

Concept:

Non-metals + hydrogen $\rightarrow$ Co-valent non-metal hydride.

Ex. 1 Sulphur is combines with hydrogen to form a covalent hydrides called hydrogen sulphide, H_2S.

$$\underset{\text{Hydrogen}}{H_2(g)} + \underset{\substack{\text{Sulphur}\\\text{(non–metal)}}}{S(\ell)} + \underset{\substack{\text{Hydrogen Sulphide}\\\text{(covalent hydride)}}}{H_2S(g)}$$

Ex. 2 Nitrogen combinds with hydrogen in the prsence of iron catalyst to form ammonia.

$$\underset{\substack{\text{Nitrogen}\\\text{(non–metal)}}}{N_2(g)} + \underset{\text{Hydrogen}}{3H_2(g)} \rightarrow \underset{\substack{\text{Ammonia}\\\text{(covalent hydride)}}}{2NH_3(g)}$$

Note:

Difference between Metals and Non Metals on the basis of Physical Properties

Metals	Non - Metals
1. Metals are malleable and ductile.	1. Non - Metals are brittle.
2. Metals are good conductors of heat	2. Non - Metals are bad conductors of Heat.
3. Metals are good conductor of electricity.	3. Non - Metals are bad conductor of electricity except graphite.
4. Metals are lustrous.	4. Non - Metals are non lustours except Iodine.
5. Metals are sonorous.	5. Non - Metals are non sonorous

Difference between Metals and Non Metals on the basis of Chemical Properties

Metals	Non - Metals
1. Metals react with oxygen to form basic oxide	2. Non - Metals react with oxygen to form acidic or neutral oxide.
2. Metals react with water to form hydrogen gas.	2. Non - Metals do not react with water.
3. Metals react with Dil acid to form H2 gas.	3. Non - Metals do not react with Dil Acid.
4. Metals react with chlorine to form Ionic metal chloride.	4. Non - Metals react with chlorine to form covalent non - metal chloride.
5. Metals react with hydrogen to form ionic metal hydride.	5. Non - Metals react with hydrogen to form covalent non - metal hydride.

Uses of Metals	Uses of Non-metals
1. Copper and aluminium metals are used to make wires to carry electric. This is because copper and aluminium have very low electrical resistance and hence very good conductors of electricity.	1. Hydrogen is used in the hydrogenation of vegetable oils to make vegetable ghee (or vanaspati ghee).
2. Iron, copper and aluminium metals are used to make household utensile and factory equipments.	2. Hydrogen is used in the manufacture of ammo nia (whose compounds are used as fertilizers).
3. Zinc is used for galvanizing iron to protect from rusting.	3. Liquid hydrogen is used as a rocket fuel.
4. Chromium and nickel metals are used for electroplating and in this manufacture of stainless steel.	4. Carbon (in the form of graphite) is used for making the electrodes of electrolytic cells and dry cells.
5. The aluminium foils are used in packaging of medicines, cigarettes and good materials.	5. Nitrogen is used in the manufacture of ammonia, nitric acid and fertilizers.
6. Silver and gold metals are used to make jewellery. The thin foils made of silver and gold are used to decorate sweets.	6. Due to its inertness, nitrogen is used to preserve food materials.
7. The liquid metal 'mercury' is used in making thermometers.	7. Compounds of nitrogen like Tri Nitro Toluence (TNT) and nitroglycerine are used as explosives.
8. Sodium, titanium and zirconium metals are used in atomic energy (nuclear energy) and space science projects.	Sulphur is used for manufacutring sulphuric acid.
9. Zirconium metal is used in making bullet-proof alloy steels.	8. Sulphur is used as a fungicide and in making gun powder.

TOPIC : HOW DO METALS REACT WITH NON - METALS

Metals reac with non-metals to achieve the electronic configuration of nearest inert gas. Which are quite stable. When metals react with non metals, to form ionic compounds and when a non - metal react with another non - metal to form covalent compounds. The force which links the atoms in a molecule is called chemical bond. An atom can achieve the electronic configuration of the inert gas in three ways:

(1) by losing one or more electrons.

(2) by gaining one or more electrons.

(3) by sharing one or more electrons.

Note: The usual number of electrons in the shell of an atom of a noble gas is 8. Except Helium - it has only two electrons in their outermost shell.

Types of Chemical Bonds

There are two types of chemical bonds. These are:

(1) Ionic bond, and

(2) Covalent bond

(1) **Ionic Bond:** The chemical bond formed by the transfer of electrons from one atom to another atom is known as an ionic bond. The ionic bonds are formed between metals and non-metals. The compounds containing ionic bonds are called ionic compounds.

(2) **Covalent bond:** The chemical bond formed between two atoms by sharing of electrons is called covalent bond. Covalent bonds are formed between non-metals. The compounds containing covalent bonds are called covalent compounds.

1. **Formation of Sodium chloride. (NaCl)**

Step 1: The atomic number of sodium is 11, so its electronoic configuration is 2, 8, 1. So, the sodium atom loses 1 electron from their outermost shell and achieve the electronic configuration of innert gas neon and form sodium ion (Na^+).

$$\cdot Na \ (\text{Sodium Atom, } 2, 8, 1) - e^- \ (\text{electron}) \longrightarrow Na^+ \ (\text{Sodium ion, } 2, 8) \quad \text{(i)}$$

Step 2: The atomic number of clorine is 17, so its electronic configuration is 2, 8, 7. So chlorine atoms needs one electron to achieve the electronic configuration of innert gas Argon and form chloride ion (Cl^-).

$$:\dot{C}l: \ (\text{K L M: } 2\ 8\ 7) + e^- \longrightarrow [:\dot{C}l:]^- \ (\text{K L M: } 2\ 8\ 8) \quad \text{(ii)}$$

Adding (i) and (ii) equation:

$$Na\cdot \ (\text{Sodium atom, } 2,8,1) + \cdot\ddot{C}l: \ (\text{Clorine atoms, } 2,8,7) \longrightarrow Na^+ \ (\text{Sodium ion, } 2,8) \ [:\ddot{C}l:]^- \ (\text{Chloride ion, } 2,8,8)$$

due to opposite charge both sodium ion and chloride ion (Cl^-) combine to form sodium chloride (NaCl) molecule.

2. **Formation Potassium Chloride (KCl)**

Step - 1: The atomic number of potassium is 19. So its electronic configuration is 2, 8, 8, 1. So one atom of potassium losses one electron from their outer most shell and achieve the electronic configuration of innert gas Argon and form potassium ion (K^+).

$$K\cdot \ (\text{K L M N: } 2\ 8\ 8\ 1) - e^- \longrightarrow K^+ \ (\text{K L M: } 2\ 8\ 8) \quad \text{(i)}$$

Step - 2: The atomic number of chlorine is 17. So, its electronic configuration is 2, 8, 7. So, one atom of chlorine needs only one electron to achieve the electronic configuration of innert gas Argon and form chloride ion (Cl^-)

$$:\dot{C}l: \ (\text{Chlorine atom, K L M: } 2\ 8\ 7) + e^- \longrightarrow [:\dot{C}l:] \ (\text{Chloride ion, K L M: } 2\ 8\ 8) \quad \text{(ii)}$$

Adding equation (i) and (ii), we get

$$K \ (\text{K L M N: } 2\ 8\ 8\ 1) + Cl: \ \longrightarrow K^+ \ (\text{K L M: } 2\ 8\ 8) \ [:\dot{C}l:] \ (\text{K L M: } 2\ 8\ 8)$$

KCl (Potassium Chloride)

due to opposite Charge both Potasion ion and chloride ion combine to form potassium chloride (KCl) molecule.

3. **Formation of Magnesium Chloride ($MgCl_2$)**

Step 1: The atomic number of magnesium is 12, so its electronic configuration is 2, 8, 2. So, one atom of magnesium loses two electrons from their outermost shell and active the electronic configuration of inert gas neon and form magnesium ion (Mg^{2+}).

$$\underset{\substack{\text{Magnesium Atom}\\2,8,2}}{Mg:} - 2e^- \longrightarrow \underset{\substack{\text{Magnesium ion}\\2,8}}{Mg^{2+}} \qquad (i)$$

Step 2: The atomic number of chlorine is 17. So, its electronic configuration is 2, 8, 7. Each atom of chlorine needs only one electron to achieve the electronic configuration of innert gas. Since one magnesium atom loses 2 electrons, so two chlorine atoms takes these two electrons and form two chloride ions (Cl^-).

$$\underset{\substack{\text{Two chlorine}\\\text{atoms}\\2(2,8,7)}}{2.\ddot{C}l:} + \underset{\text{Electrons}}{2e^-} \longrightarrow \underset{\substack{\text{Two Chloride}\\\text{ions}\\2(2,8,8)}}{2[:\ddot{C}l:]^-} \qquad (ii)$$

Adding (i) and (ii) equation

$$\underset{\substack{K\,L\,M\\2\,8\,2}}{Mg:} + 2\underset{\substack{K\,L\,M\\2\,8\,7}}{.\ddot{C}l:} \longrightarrow \underset{\substack{K\,L\\2\,8}}{Mg^{2+}} \;\underset{\substack{\text{Magnesium chloride}\\MgCl_2}}{2[:\ddot{C}l:]^-}$$

due to opposite charge both magnesium ion and chloride ion combine to form magnesium chloride ($MgCl_2$) molecule.

4. **Formation of Calcium Chloride ($CaCl_2$)**

Step - 1: The atomic number of calcium is 20, So, its electronic configuration is 2, 8, 8, 2. So one atom of calcium loses two electrons from their outermost shell to achieve the electronic configuration of inert gas Argon and form calcium ion (Ca^{2+}).

$$\underset{\substack{K\,L\,M\,N\\2\,8\,8\,2}}{Ca:} - 2e^- \longrightarrow \underset{\substack{K\,L\,M\\2\,8\,8}}{Ca^{2+}} \qquad (i)$$

Step - 2: The atomic number of Cl is 17. So its electronic configuration is 2, 8, 7. So each atom of chlorine needs any one electron to achieve the electronic configuration of innert gas Argon. Since calcium atom losses two electrons. So two chlorine atoms takes these electrons and form two chloride ions

$$2\underset{\substack{K\,L\,M\\2\,7\,8}}{\dot{C}l:} + 2e^- \longrightarrow 2\underset{\substack{K\,L\,M\\2\,8\,8}}{[:\ddot{C}l:]^-} \qquad (ii)$$

Adding equation (i) and (ii)

$$\underset{\substack{K\,L\,M\,N\\2\,7\,8\,2}}{Ca:} + 2\dot{C}l: \longrightarrow \underset{\substack{K\,L\,M\\2\,8\,8}}{Ca^{2+}} \quad \underset{\substack{\text{(Calcium chloride)}\\CaCl_2}}{2[:\ddot{C}l:]}$$

5. **Formation of calcium oxide (CaO)**

Step 1: The atomic number of Ca is 20. So, its electronic configuration is 2, 8, 8, 2. So, one atom of calcium loses two electrons from their outermost shell to achieve the electronic configuration of inert gas Argon.

$$\underset{\substack{K\,L\,M\,N\\2,8,8,2}}{Ca:} - \underset{\substack{\text{Loses two}\\\text{electrons}}}{2e^-} \longrightarrow \underset{\substack{K\,L\,M\\2,8,8}}{Ca^{2+}} \qquad (i)$$

Step 2: The atomic number of oxygen is 8. So, its electronic configuration is 2, 6. Hence one atom of oxygen gain to electrons to achieve the electronic configuration of innert gas Neon.

$$\underset{\substack{K\,L\\2,6}}{:\dot{O}:} + \underset{\substack{\text{gain two}\\\text{electrons}}}{2e^-} \longrightarrow \underset{\substack{K\,L\\2,8}}{[:\ddot{O}:]^{-2}} \qquad (ii)$$

Adding (i) and (ii) equation, we get

$$\underset{\substack{K\,L\,M\,N\\2,8,8,2}}{Ca:} + \underset{\substack{K\,L\\2,6}}{:\dot{O}:} \longrightarrow \underset{\substack{K\,L\,M\\2,8,8}}{Ca^{2+}} \quad \underset{\substack{\text{CaO (Calcium Oxide)}\\K\,L\\2,8}}{[:\ddot{O}:]^{-2}}$$

due to opposite charge both calcium ion and oxide ion combine to form calcium oxide (CaO) molecules.

6. **Formation of magnesium oxide (MgO)**

Step1: The atomic number of Mg is 12. So its electronic configuration is 2, 8, 2. So, one atom of Mg loses two electrons from their outermost shell to achieve the electronic configuration of inert gas Neon and form magnesium ion (Mg^{2+}) molecule.

$$\underset{\substack{\text{Magnesium atom}\\\text{ion K L M}\\2,8,2}}{Mg:} - 2e^- \longrightarrow \underset{\substack{\text{Magnesium}\\K\,L\\2,8}}{Mg^{2+}} \qquad (i)$$

Step 2: The atomic number of oxygen is 8. So, its electronic configuration is 2, 6. Hence one atom of oxygen needs only two electrons to achieve the electronic configuration of inert gas Neon and form oxide ion (O^{-2}).

$$\underset{\substack{K\ L\\2,\ 6}}{:\ddot{O}:} + 2e^- \longrightarrow \underset{\substack{\text{Oxide ion}\\K\ L\\2,\ 8}}{[:\ddot{O}:]^{-2}} \quad \text{(ii)}$$

Adding (i) and (ii) equation, we get

$$\underset{\substack{K\,L\,M\\2,8,2}}{Mg:} + \underset{\substack{K\,L\\2,6}}{:\dot{O}:} \longrightarrow \underset{\substack{K\,L\\2,8}}{Mg^{2+}} \underset{\substack{\text{MgO}\\\text{Magnesium}}}{[:\ddot{O}:]^{-2}}$$

Due to opposite charge both magnesium ion and oxide ion combine to form magnesium oxide (MgO) molecules.

7. **Formation of sodium oxide (Na_2O) T.B.Q**

Step 1: The atomic number of oxygen atom is 8. So, its electronic configuration is 2, 6. So, one atom of oxygen needs only two electrons to achieve the electronic configuration of inert gas neon and form oxide ion (O^{-2})

$$\underset{\substack{\text{Oxygen atom}\\K\ L\\2\ \ 6}}{:\ddot{O}:} + 2e^- \longrightarrow \underset{\substack{K\ L\\2\ 8}}{[:\ddot{O}:]} \quad \text{(i)}$$

Step 2: The atomic number of sodium is 11, so its electronic configuration is 2, 8, 1. Since, one atom of oxygen need two electrons. So, two atoms of sodium loses two electrons from their outermost shell to form two sodium ions and achieve the electronic configuration of inert gas Neon.

$$\underset{\substack{\text{Sodium atom}\\K\ L\ M\\2\ 8\ 1}}{2Na\cdot} - 2e^- \longrightarrow \underset{\substack{K\ L\\2\ 8}}{2Na^+} \quad \text{(ii)}$$

Adding (i) and (ii) equation, we get

$$\underset{\substack{K\,L\,M\\2\ 8\ 1}}{2Na\cdot} + \underset{\substack{K\ L\\2\ 6}}{\ddot{O}:} \longrightarrow \underset{\substack{K\ L\\2\ 6}}{2Na^+} \underset{Na_2O}{[:\ddot{O}:]^{-2}}$$

Due to opposed charge both sodium ion and oxide ion combine to form sodium oxide (Na_2O) molecules.

8. **Formation Potassium oxide (K_2O)**

Step-1: The atomic number of potassium is 19. So, its electronic configuration is 2, 8, 8, 1. So, each atom of Potassium loses one electron from their outermost shell to achieve the electronic configuration of inert gas Argon. Since one atom of oxygen need two electrons. So two atom of potassium loses two electrons and form two potassium ion (K^+).

$$\underset{\substack{K\,L\,M\,N\\2\ 8\ 8\ 1}}{2K\cdot} - 2e^- \longrightarrow \underset{\substack{K\,L\,M\\2\ 8\ 8}}{2K^+}$$

Step-2: The atomic number of oxygen is 8. So its electronic configuration is 2, 6. So one atom of oxygen needs two electron to achieve the electronic configuration of Inert gas Neon and form oxide ion (O^{-2}).

$$\underset{\substack{K\ L\\2\ 6}}{\dot{O}\cdot} + 2e^- \longrightarrow \underset{\substack{K\ L\\2\ 8}}{[:\ddot{O}:]^{-2}}$$

Adding (i) and (ii) equation, we get

$$\underset{\substack{K\,L\,M\,N\\2\ 8\ 8\ 1}}{2K} + \underset{\substack{K\ L\\2\ 6}}{\cdot\ddot{O}\cdot} \longrightarrow \underset{\substack{K\,L\,M\\2\ 8\ 8}}{2K^+} \underset{\substack{K\ L\\2\ 8}}{[:\ddot{O}:]^{-2}}$$

K_2O (Potassium oxide)

Properties of Ionic Compounds

1. **Ionic compounds are usually crystalline solids.** This is because the oppositely charged ions attract one another strongly and form a regular crystal structure.
2. **Ionic compounds have high melting and boiling points.** This is because they are made up opposite ions. There is a strong force of attraction between the oppositely charged ions, so a lot of heat energy is required to break this strong force of attraction, due to this ionic compounds have high melting point and boiling points.
3. **Ionic compounds are usually soluble in water but insoluble in organic solvents:** This is because water has a high dielectric constant due to which it weakens the force of attraction between the ions.

4. **Ionic compounds conduct electricity when dissolved in water or taken in molten:** This is because in aqueous solution or in molten state. They provides free ions due to this they conduct electricity in aqueous solution or in molten state.

Properties of Covalent Compounds:

1. **Covalent compounds are usually liquids or gases.** This is because the force of attraction between the molecules of a covalent compounds is very weak.
2. **Covalent compounds have usually low melting and boiling points.** Thsi is because the force of attraction between the molecules of a covalent compounds is very weak.
3. **Covalent compounds are usually insoluble in water but they are soluble in organic solvents.**
4. **Covlent compounds do not conduct electricity.** This is because they do not contain ions.

How to distinguish between ionic compounds and covalent compounds:

(1) If a compound has high melting point and boiling point, then it will be an ionic compound otherwise it will be covalent compound.

(2) If a compound is soluble in water but insoluble in organic solvents, it will be an ionic compound otherwise it will be a covalent compound.

(3) If a compound conducts electricity (in the solution form and molten state), it will be an ionic compound otherwise it will be a covalent compound.

Difference b/w Ionic and Covalent Compounds

Ionic Compounds	Covalent Compounds
1. Ionic compounds are generally solids.	1. Covalent compounds are liquids or gas
2. Ionic compounds are soluble in inorganic solvent like water.	2. Covalent compounds are soluble in organic solvent.
3. Ionic compounds are electrolytes.	3. Covalent compounds are arenot electrolyte.
4. Ionic compounds have high melting points and boiling points.	4. Covalent compounds have low melting point and boiling point.

Notes:

TOPIC : EXISTENCE OF METALS, (ORES, MINERALS AND REDUCTION OF METALS)

The earth's crust is the major source of metals. Most of the metals are quite reactive and hence they do not occur in the free state in nautre. Only a few less reactive metals like copper, silver, gold and platinum are found in the 'free state' in the earth crust. **Copper and silver metals** occur in free state as well as in the combined state.

☞ All the metals which are placed above copper in the reactivity series are found in nature only in the form of combine state.

Minerals and Ores

Minerals: The naturally occurring substances in the earth crust in which metals or their compounds are found called minerals. **For examples:** Iron found in the earth crust in the form of Iron IIIrd oxide (Fe_2O_3) and Iron ((ii), (iii)) oxide (Fe_3O_4). Hence Fe_2O_3 and Fe_3O_4 both are the mineral of Iron metal.

Ores: Those minerals from which metals can be extracted conveniently and profitably are called ores. **For example:** Iron can be extracted easily and profitavely from Iron III oxide. So Fe_2O_3 is an ore of Iron metal.

Note: All the ores are minerals, but all the minerals are not ores.

Types of ores: There are four types of ore:

(i) Oxide ores (ii) Sulphide ores

(iii) Carbonate ores

(iv) Halides ores. Important ores and their formulae are given below:

Nature of ores	Metal	Name of Ore	Composition
Oxides Ores	Alum.	Bauxite	$Al_2O_3.2H_2O$
	Copper	Cuprite	Cu_2O
	Iron	Haematite	Fe_2O_3
Sulphide ores	Copper	Copper glance	Cu_2S
	Zinc	Zinc blende	ZnS
	Mercury	Cinnabar	HgS
Carbonate ores	Calcium	Limestone	$CaCO_3$
	Zinc	Calamine	$ZnCO_3$
Halide ores	Sodium	Rock salt	NaCl

Topic : Extraction of Metals

To obtain a metal from its ore is called the extraction of metal.

Metallurgy: The various steps involved in the extraction of metal from their ore, and refining of metal is called metallurgy. There are three major steps involved in the extraction of a metal from its ore these are:

1. Concentration of ore (or Enrichment of ore),
2. Conversion of concentrated ore into metals (or reduction of metals from concentrated ores)
3. Refining of impure metals or purification of metals .

1. **Concentration centre of ore:** The removal of unwanted impurities present in an ore is called concentration of ore. The unwnated impurities like sand, rocky materials, earthy particles, limestone, mica etc., present in an ore are called **metrix** (or gangue). The methods used for removing matrix from an ore depends on the physical and chemical properties of the nature of ores and nature of metrix. Concentration of ore is also known as **enrichment of ore**.
2. **Reduction of metal from concentrated ore** The reduction of metals from concentrated ore depends upon the reactivity of metals. On the basis of reactivity of metals all metals are divided into three categories:-
   (i) Highly reactive metals
   (ii) Moderate reactive metals
   (iii) Less reactive metals.

(i) **Highly reactive metals:** Highly reactive metals (like Na, K, Ca, Mg and Al) are reduced from their concentrated ores with the help of electric current which is called electrolytic reduction.

(ii) **Moderate reactive metas:** Moderate reactive metals (like Zn, Fe, Pb, Mn, Sn) are reduced from their concentrated ores by using chemicals.

(iii) **Less reactive metals:** Less reactive metals (like Cu, Ag, Hg) are reduced from their concentrated ores by heating alone.

Nature of Metals & their Reduction

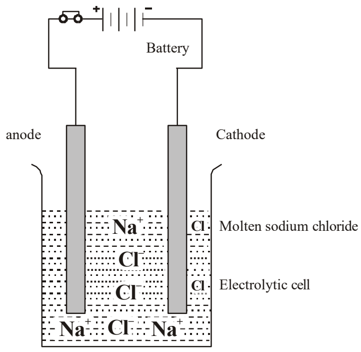
Nature of Metal	Name of Metal	Methodof extraction (Reduction)
(i) Highly reactive metals	K Na Ca Mg	Electrolysis of molten chloride or oxide
(ii) Moderately reactive metals	Al Zn Fe, Mn Pb, Sn	Reduction of oxide with carbon
(iii) less reactive metals	Cu Ag Hg	Heating in air
(iv) Very less reactive metals	Au Pt	Found in native state

(i) **Extraction of Highly Reactive Metals:** It is found that highly reactive metals (like Na, K, Ca, Mg and Al) are found in the earth crust in the form of their chloride and oxide form. Since the chloride and oxide of these metals are highly stable. So, they rquired large amount of energy to reduce from their chloride and oxide. So they are reduced only with the help of electric current. So highly reactive metals are reduced by the electrolysis of their molten chlorides or xides.

☞ <u>The reduction of metals from their ores with the help of electric current is called electrolytic reduction.</u> In the electrolydic reduction cathod act as reducing agent.

Ex.1 **Reduction of sodium metal:** When electric current is passes through molten sodium chloride, it decomposes to form sodium metal and chlorine gas:

$$\underset{\text{Sodium chloride (Molten)}}{2NaCl(l)} \xrightarrow{\text{Electrolysis}} \underset{\text{Sodium metal}}{2Na(s)} + \underset{\text{Chlorine gas}}{Cl_2(g)}$$



During the electrolysis of molten sodium chloride the following chemical reactions take place at cathode and anode.

Reaction at cathode: The positively charged sodium ions (Na^+) from molten sodium chloride are moved towards the cathode and they takes electrons from the cathode and gets reduced to form sodium atoms

$$\underset{\text{(from molten NaCl)}}{2Na^+} + 2e^- \xrightarrow{\text{Reduction}} \underset{\text{Sodium atom}}{2Na(metal)}$$

This sodium metal is formed at the cathode.

Reaction at anode: The negatively charged chloride ions (Cl^-) from molten sodium chloride are moved towards the anode and they give electrons to the anode and gets oxidised to from chlorine gas.

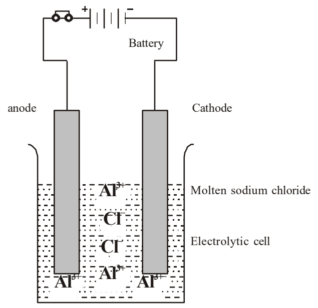
$$\underset{\text{Chlorides ion}}{2Cl^-} - 2e^- \xrightarrow{\text{Oxidation}} \underset{\text{Chlorine gas}}{Cl_2}$$

Important discussion: We do not use an aqueous solution of metal chloride or metal oxide in electroletic reduction. This is because as soon as metal is produced at cathode react with water present in aqueous solution to form metal hydroxide and hydrogen gas.

Ex.2 **Extraction of Aluminium Metal:** When electric current is passes through molten aluminium oxide. It decomposes to form aluminium metal and oxygen gas.

$$\underset{\text{Aluminium oxide (Molten)}}{2Al_2O_3(l)} \xrightarrow{\text{Electrolysis}} \underset{\text{Aluminium metal}}{4Al(l)} + \underset{\text{Oxygen gas}}{3O_2(g)}$$

During the electrolysis of molten aluminium oxide, the following chemical reactions take place at cathode and anode.



Reaction at Cathode: The positively charged aluminium ions (Al^{3+}) are moved towards the cathode and they takes electrons from the cathode and get reduced to form aluminium atoms.

$$\underset{\text{(from molten } Al_2O_3)}{4Al^{3+}} + \underset{\text{(from cathode)}}{12e^-} \rightarrow \underset{\text{(aluminium metal)}}{4Al}$$

Thus, aluminium metal is formed at the cathode.

Reaction at Anode: The negatively charged oxide ions (O^{2-}) are moved towards the anode and they gives electrons to the anode and gets oxidised to form oxygen gas:

$$\underset{\text{from molten } (Al_2O_3)}{6O^{2-}} - \underset{\text{(give to anode)}}{12e^-} \xrightarrow{\text{oxidation}} \underset{\text{Oxygen Gas}}{3O_2}$$

Note: In the electrolytic reduction we get pure metal, further refining of metal is not required.

Extraction of Moderately Reactive Metals: The moderately reactive metals (like Fe, Ni, Sn, Mn and Pb) are extracted from their oxides with the help of chemicals. These chemical may be carbon or aluminium. Some moderately reactive metal occurs in nature in the form of oxides (like Fe_2O_3). MnO_2) but others are occure in the form of carbonates or sulphides. It is easier to reduced metals from their oxides in comparison of carbonates or sulphides. So before reduction metals. The carbonates and sulphides ores must be converted in the form of oxides. This is done by roasting or calcination.

(i) **Calcination:** Calcination is the process in which a carbonate ore is heated strongly in the absence of oxygen of air to convert it into metal oxides.
Example, when calamine (or zinc carbonate) is heated strongly in the absence of oxygen of air, it decomposes to form zinc oxide and carbon dioxide:

$$\underset{\text{(calamine)}}{ZnCO_3(s)} \xrightarrow{\text{calcination}} \underset{\text{Zinc oxide}}{ZnO(s)} + \underset{\text{Carbon dioxide}}{CO_2(g)}$$

(ii) **Roasting:** Roasting is the process in which a sulphide ore is strongly heated in the presence of oxygen of air to convert it into metal oxide.
Example, When zinc blende (or zinc sulphide) is heated strongly in the presence of oxygen of air, it decomposes to form zinc oxide and sulphur dioxide:

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

Differences b/w Roasting and Calcination

	Roasting	Calcination
1.	Roasting process is used for sulphide ore.	Calcination process is used for carbonate ore.
2.	In roasting SO_2 gas is evolved.	In Calcination CO_2 gas is evolved.
3.	Roasting process is done in the presence oxygen of air.	Calcination process is done in the absence of oxygen of air.

Reduction of metals with carbon:

(i) **Reduction of zinc metal from calamin ($ZnCO_3$):** The reduction Zinc metal from calamine ore ($ZnCO_3$) completed in two steps.
Step-1: In the first step calamine is heated strongly in the absense of oxygen of air, it decompasses to form Zinc oxide and carbon dioxide.

$$ZnCO_3 \xrightarrow{\text{Calcination}} ZnO + CO_2$$

Step-2: In second step zinc oxide is heated with carbon zinc metal is produced.

$$ZnO + C \xrightarrow{\text{Reductions}} Zn + CO$$

(ii) **Reduction of zinc metal from zinc blende-(ZnS):** The reduce of zinc metal from zinc blend is complete in two steps:

Step-1: In first step Zinc blend (ZnS) is heated strongly in the presence of oxygen of air, it decomposes to form Zinc oxide and sulphur dioxide.

$$ZnS + O_2 \xrightarrow{\text{Roasting}} ZnO + SO_2$$

Step-2: In second step zinc oxide is heated with carbon, zinc metal is produced.

$$ZnO + C \longrightarrow Zn + CO$$

Note: Iron, Tin and lead metals are also reduced from their oxides with the help of carbon. Even less reactive metal like copper also can be reduced from its oxide by using as carbon reducing agent.

Reduction of metal with aluminium powder: If we required high grade metal then we reduced moderate reactive metal from their oxide with the help of aluminium powder(thermite powder).

Ex.1 **Reduction of manganese metal:** When manganese dioxide is heated strongly with aluminium powder, then magnese metal is obtained and aluminium metal itself oxidised to form aluminium oxide.

$$\underset{\substack{\text{Manganese}\\\text{dioxide}}}{3MnO_2(s)} + \underset{\substack{\text{Aluminium powder}\\\text{(reducing agent)}}}{4Al(s)} \longrightarrow \underset{\text{Manganese}}{3Mn(l)} + \underset{\substack{\text{Aluminium}\\\text{Oxide}}}{2Al_2O_3(s)} + \text{Heat}$$

Ex.2 **Reduction of Iron metal:** In thermite welding when Haemetite ore (Fe_2O_3) is heated with aluminium powder then Iron metal is obtained and aluminium itself oxidised metal to form aluminium oxide.

$$\underset{\text{Haemetite}}{Fe_2O_3} + 2Al \longrightarrow \underset{\text{(molten)}}{2Fe} + Al_2O_3$$

Thermite Welding: <u>The welding done with the help of aluminium powder is called thermite welding.</u> In thermite welding Iron metal is obtained in the form of molten state. Which is poured between the broken iron pieces to join them. This process is also called **alumino thermy.**

Thermite reaction: <u>The reduction of metal from their oxide ore with the help of aluminium powder is called thermite reaction.</u>

(iii) Extraction of Less Reactive Metals:

Less reactive metal is also occur in the form of sulphide ores carbonate ores. So, before reduction of metal. They also converted in the form of oxide. The less reactive metals are extracted by by heated alone from their oxides.

Ex.1 **Extraction of Mercury metal (HgS):** Mercury metal extracted from cinnabar ore (HgS). The extraction of mercury metal from cinnabar ore (HgS) is completed in two steps.

Step-1: In the first step the concentrated mercury (II) Sulphide ore (cinnabar ore) is strongly heated in the presence of oxygen of air to form mercury (II) oxide and sulphur dioxide gas.

$$2HgS + 3O_2 \xrightarrow{\text{roasting}} 2HgO + 2SO_2$$

Step-2: In the second step mercury (II) oxide is heated strongly it decomposes to form mercury metal.

$$\underset{\substack{\text{Mercury (II)}\\\text{Oxide}}}{2HgO(s)} \xrightarrow[\text{(reduction)}]{\text{heat}} \underset{\substack{\text{Mercury}\\\text{metal}}}{2Hg(s)} + \underset{\text{Oxygen}}{O_2(g)}$$

Ex.2 **Extraction of copper metal (Cu_2S):** Copper metal is extracted from copper glance (Cu_2S). The extraction of copper metal from copper glance (Cu_2s) is completed in two steps.

Step-1: The concentrated copper (I) sulphide ore (or copper glance) is heated strongly the presence of oxygen of air when a part of copper (I) sulphide has been converted into copper (I) oxide, then supply of air is stopped for roasting

$$\underset{\substack{\text{Copper}\\\text{glance}}}{2Cu_2S} + 3O_2 \xrightarrow{\text{roasting}} \underset{\substack{\text{Copper(I)}\\\text{oxide}}}{2Cu_2O} + 2SO_2$$

Step-2: In the absence of oxygen of air, copper (I) oxide formed reacts with the remaining copper (I) sulphide to form copper metal and sulphur dioxide

$$\underset{\text{Copper(s) oxide}}{2Cu_2O} + \underset{\text{Copper glance}}{Cu_2S} \xrightarrow{\text{heat}} \underset{\text{Copper Metal}}{6Cu} + SO_2$$

Refining of metals

The process of purification of impure metals is called refining of metals. The most widely used method for refining of impure metals is electrolytic refining.

Electrolytic Refining: The purification of metals with the help of electric current is called electrolytic refining. Many metals like copper, zinc, tin, lead, chromium, nickel, silver and gold are refined electrolytically. For the refining of an impure metal by electrolysis:

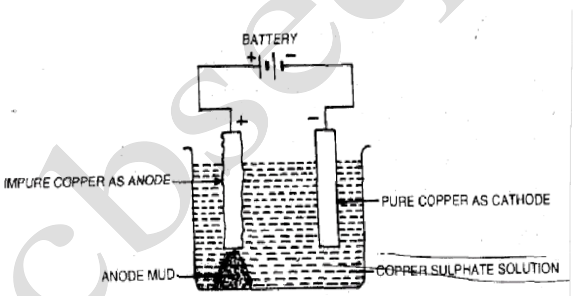
(a) A thick block of the impure metal is made anode.

(b) A thin strip of the pure metal is made cathode.

(c) A water soluble salt of the metals to be refined is taken as electrolyte.

Electrolytic Refining of Copper: In the refining of copper metal an acidified aqueous solution of copper sulphate taken as electroyle. A thick block of impure copper metal is made anode and a thin strip of pure copper metal is made cathode.

When electric current is passed through aquesous solution of copper sulphate solution then following chemical reactions occurs at cathode and anode.



Reaction at cathode:

The positively charged copper ions, (Cu^{2+}) are moved towards the cathode and they takes electrons from the cathode, gets reduce to copper atoms.

$$\underset{\substack{\text{Copper ion}\\ \text{(from electrolyte)}}}{Cu^{2+}} + \underset{\substack{\text{Electrons}\\ \text{(from cathode)}}}{2e^-} \xrightarrow{\text{reduction}} \underset{\substack{\text{Copper atom}\\ \text{(deposits on cathode)}}}{Cu}$$

Reaction at Anode:

Copper atoms of impure anode lose two electrons and form copper ions, (Cu^{2+}), which go into the solution.

$$\underset{\substack{\text{(from}\\ \text{impure}\\ \text{anode)}}}{Cu} - \underset{\substack{\text{(give to}\\ \text{anode)}}}{2e^-} \xrightarrow{\text{oxidation}} \underset{\substack{\text{(goes into}\\ \text{electrolyte)}}}{Cu^{2+}}$$

As the process goes on, impure anode rod becomes thinner and thinner. Whereas pure cathode rod becomes thicker and thicker in this way pure copper is obtained at the cathode and impurities like gold and silver present in the impure copper rod are collected below the anode which is called **anode mud.** Thus, the electrolytic refining of metals serves two purposes these are:

(i) It refines the metal concerned.

(ii) It enables to recover other valuable metals like gold and silver.

Note: In electrolytic reduction. The metals are extracted from their ore with the help of electric current. But in the electrolytic refining metals are purified with the help of electric current.

TOPIC : CORROSION

The eating up of metals by the action of air, moisture or a chemical (such as an acid) on their surface is called corrosion. Corrosion of iron is a continuous process which ultimately eats up the whole iron object. The corrosion of iron is called rusting.

Rusting of Iron: When an iron object is left in damp air (or moist air) for a considerable time, it gets covered with a red-brown flaky substance called rust.

$$\underset{(Iron)}{4Fe} + \underset{dempair}{xH_2O + 3O_2} \longrightarrow \underset{\substack{\text{rust}\\ \text{(red, brown colour)}}}{2Fe_2O_3 \,.\, xH_2O}$$

Conditions for Rusting: These are two conditions for rusting of Iron. These are:

1. Presence of air (or oxygen)
2. Presence of water (or moisture)

☞ Damp air alone supplies both the things, air and water, required for the rusting of iron.

ACTIVITY – 3.8

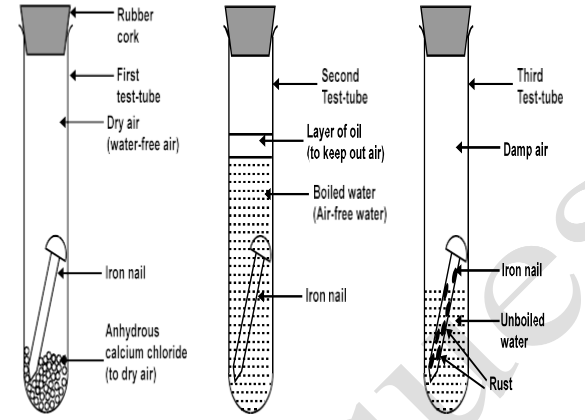
Aim: To show that Rusting of iron requires both, Air and Water

Materials Required: Three test tube, iron nail, anhydrous calcium chloride, boiled water and oil.

Step 1: In the first test-tube containing iron nail, put some anhydrous calcium chloride and close its mouth with a tight cork. The anhydrous calcium chloride absorbs water from the damp air present in the test-tube and make it dry. This test tube is kept aside for a weak.

Step 2: In the second test-tube containing iron nail, poured boiled distilled water. A layer of oil is put over boiled water in the test-tube to prevent the outside air from mixing with boiled water. The mouth of this test-tube is closed with a cork and it is kept aside for a weak.

Step 3: In the third test-tube containing iron nail, poured unboiled water so that about two-third of the nail is immersed in water and and the rest is above the water. The mouth of this test tube is closed with a cork and is kept aside for a weak.



(a) Only air : No rusting (b) Only water : No rusting (c) Air and water together : Rusting takes place

Observations:

(1) No rust is seen on the surface of iron nail in the first test-tube. This result shows that rusting of iron does not take place in air alone.

(2) No rust is seen on the surface of iron nail kept in second test-tube. This result shows that rusting of iron does not take place in water alone.

(iii) Red-brown rust is seen on the surface of iron nail in the third test tube. This result shows that rusting of iron takes place in the presence of both air and water together.

Prevention of Rusting of Iron: Rusting of Iron can be prevented by following methods:

1. Rusting of iron can be prevented by painting.
2. Rusting of iron can be prevented by applying greases or oil on the surface of objects.
3. **Rusting of iron can be prevented by galvanization.** (The process of depositing a thin layer of zinc metal on an iron object is called galvanization). It is done by dipping an iron object in molten zinc metal.

   Note: Zinc is quite reactive metal so it react with the oxygen of air to form a very thin coating of zinc oxide. The galvanized iron object remains uneffected against rusting. Even if a break occurs in the zinc layer.
4. Rusting of iron can be prevented by tin-plating and chromium-plating by electro plating process.
5. Rusting of iron can be prevented by alloying it to make stainless steel.

1. Corrosion of Aluminium Metal: When aluminium objects are exposed to air for long time then, a thin layer of aluminium oxide is formed on the surface of aluminium object which protects aluminium metal from further corrosion.

Note: In the case of aluminium metal corrosion is beneficial. A common metal which is highly resistant to corrosion is aluminium.

Anodising: The proces in which a thick layer of aluminium oxide is depositing on an aluminium object by making it anode during the electrolysis of dilute sulphuric acid.

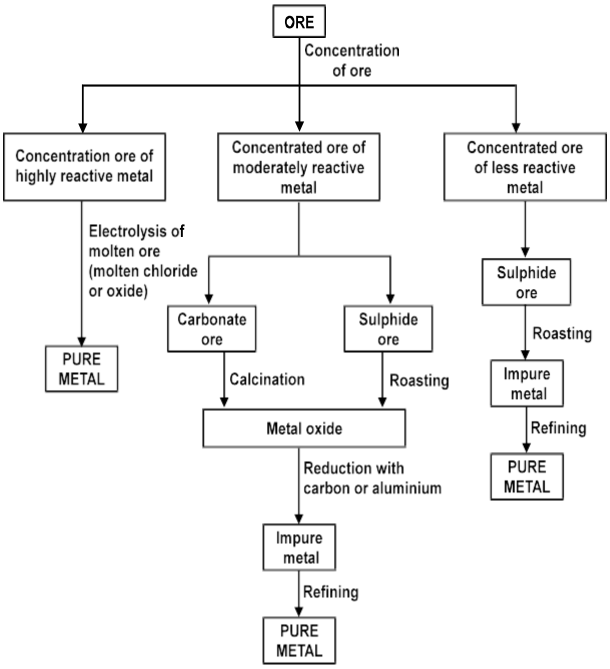
2. Corrosion of Copper Metal: When a copper object exposed in damp air for a consirable time, then a green coating of basic copper carbonate ($CuCO_3.Cu(OH)_2$) is depositing on the surface of the copper object.

$$\underset{\text{(Copper)}}{Cu} + \underset{\text{Dampair}}{\boxed{CO_2 + H_2O + O_2}} \rightarrow \underset{\text{Basic Copper Carbonate (Green)}}{CuCO_3.Cu(OH)_2}$$

3. Corrosion of Silver: When silver ornaments are exposed to the atmosphere for long time then they react with hydrogen sulphide gas (H_2S) present in the atmosphere to form a black layer of silver sulphide.

$$\underset{\text{(Silver metal)}}{Ag} + \underset{\text{Hydrogen sulphide}}{H_2S} \longrightarrow \underset{\left(\text{Silver sulphide Black}\right)}{Ag_2S} + H_2(\text{gas})$$

Note: Gold and platinium metal have high resistance to corrossion and they have bright shiny surface. Due to this both are used to make jewellery.



TOPIC : ALLOYS

Definition: Alloys are the homogeneous mixture of two or more than two metals or metals and non-metals. *For example*, brass is an alloy of two metals : copper and zinc, whereas steel is an alloy of Iron (metal) and carbon (non-metal).

Preparation of Alloy: An alloy is prepared by mixing various metals in molten state in required proportions, and then cooling their mixture to the room temperature.

Note: The alloys of a metals and non-metals are prepared by first melting the metal and then dissolving the non-metal in it followed by cooling to the room temperature.

Objectives of alloys: The objective of alloys are given below:

1. Alloys are stronger than the metals from which they are made.
2. Alloys are harder than the constituent metals
3. Alloys have high resistant to corrosion.
4. Alloys have low melting points & boiling point than the constituent metals.
5. Alloys have low electrical conductivity than pure metals.

Important alloys and their uses: Some of the common alloys are Duralumin or Duralium, Magnalim, Steel, Stainless steel, Bras, Bronze, Solder and Amalgam.

(i) **Brass:** Brass is an alloy of **Copper** and **Zinc** (Cu + Zn). It contains 80% copper and 20% zinc. Brass is more malleable and more strong than pure copper. Brass is used for making cooking utensils, Screws, nuts, bolts, wires, tubes, scientific instruments (like microscope) ornaments, vessels of flowers and fancy lamps etc.

(ii) **Bronze:** Bronze is an alloy of **Copper and Tin** (Cu + Sn). It contains 90% copper and 10% tin. Bronze is very tough and highly resistant to corrosion. It is used for making statues, coins, medals, cooking utensils and ship's propellers.

(iii) **Solder:** Solder is an alloy of **lead** and **tin** (Pb + Sn). It contains 50% lead and 50% tin. Solder is used for welding of electrical wires together.

(iv) **Amalgam:** An alloy of mercury metal with one or more other metals is known as amalgam. A solution of sodium metal in liquid mercury metal is called sodium amalgam. An amalgam consisting of **mercury, silver, tin** and **zinc** (Hg + Ag + Sn + Zn) is used by dentists for fillings in teeth.

(v) **Alloys of Gold:** The purity of gold is expressed in terms of **'carats'**. Pure gold (known as 24 carat gold) is very soft due to which it is not suitable for making jewellery. Gold is alloyed with a small amount of silver or copper to make it hard.

SOME COMMON ALLOYS, THEIR CONSTITUENTS, PROPERTIES AND USES

Name of Metal	Name of Alloys	Constituents	Properties	Uses
Aluminium	*1. Duralium or Duralium	Al, Cu, Mg, Mn	Its is Light and strong and resistance to corrosion	It is used to make aircraft bodies and parts of aircraft space satellite, kitchen-ware, bodies of ships.
.	2. Magnalium	Al, Mg	It is Very light and hard	It is used to make Balance beam, light Instrument and Pressure Cookers etc.
Iron	1. Steel	Fe,C	It is Hard, tough and strong	It is used to make Nails, screws, girders, bridges, railway lines, vehicles, ships, construction of building.
	2. Stainless Steel	Fe, Ni, Cr	It is Hard and does not rust.	It is used to make Cooking utensils, knives, scissors, ornaments tools, surgical instruments and equipment for food processing.
Copper	1. Brass	Cu, Zn	It is Malleable, strong resists to corrosion, can be easily cut	It is used to make cooking utensils, screw, nuts, bolts, wires, tubes, scientific instruments and fancy lamp etc.
	2. Bronze	Cu, Zn, Ni	It is Very strong and highly Resistance to corrosion	It is used to make Statues, coins, medals,cooking utensils and ships propellers etc.
Mercury	1. Amalgam	Hg,Ag,Sn,Zn	Does not react with acid	It is used to make as a filler in teeth.
Lead	1. Solder	Sn Pb	It has low melting point	It is used for welding of electrical wires together.
Gold		Au,Cu, or Ag	It has good casting and does not corrode.	It is used to make ornaments.