• Metals and Nonmetals -

• On the basis of their properties, all the elements can be divided into two main groups.

• Metals • Nonmetals

• Malleable - means which can be beaten with hammer to form thin sheets.

• Ductile - means which can be stretched to form thin wires.

• Brittle - means which breaks into pieces on hammering or stretching.

• METALS -

• Metals are the element that conduct heat and electricity, and are malleable and ductile.

• Metals are lustrous (shiny), hard, strong, heavy and sonorous.

• Example - Iron, Aluminium, copper, silver, Gold, Platinium, zinc, calcium etc.

• All the metals are solid, except mercury which is a liquid metal

• Metals are the elements (except hydrogen) which form positive ions by losing electrons (or donating electrons).

• Ex- Aluminium (AL) is a metal which form +vely charged aluminium ion (AL^{+3}) by losing electrons.

• Why metals are known as electropositive elements?

• Bcoz they form positive ions by losing electrons.

• The most abundant metal in earth crust is Aluminium, constitues 7% of earth.

• The second most abundant metal in earth crust is iron, constitue 4% of earth crust.

• NON-METALS -

• Non metal are the element that do not conduct heat and electricity and are neither malleable nor ductile. and they are brittle.

• It is not lustrous (not shiny), they have dull appearance.

• They are generally soft and not strong they are light and non-sonorous

• Example- carbon, silicon, Phosphorus, Sulphur, Hydrogen, oxygen, Jodine etc.

• The two allotropic forms of carbon element - Diamond and graphite

• 22 nonmetals are present, in which 10 are non metal - solids, 1 are liquid and other all 11 are gases.

• Bromine is a liquid nonmetal at room temperature.

• Non metals are the elements which form -ve ions by gaining electrons

- Non metals are known as electronegative elements because they can form -ve ions by gaining electrons.
- Exception - H^+ (Hydrogen) is the only non-metal element which loses electrons to form +ve ions, H^+ ions.
- CARBON is one of the most important non metal because all life on this earth is based on carbon compound.
- OXYGEN is equally important for existence of life bcoz it is imp. for breathing to maintain life. necessary for the combustion of fuels.
- NITROGEN is an inert gaseous non metal whose presence in air reduces the rate of combustion.
- Most abundant non metal in earth crust - OXYGEN - 50% present.
- 2nd most abundant non metal in earth crust - SILICON - 26% present

• PHYSICAL PROPERITIES OF METALS -

- Metals are mallaeble, that is metal beaten into thin sheets -
- Malleability - The property which allows the metal to be hammered into thin sheets. It is important characteristic of metal.
- Ex- Gold and silver metals are best malleable metals.
- Ex- Aluminium and copper metal are also highly malleable metals.
- foils are used for decorating sweets.
- Aluminium foils are used for packing food items like biscuit, cigarette and medicine. they all turn into thin sheets bcoz they are mallaeble

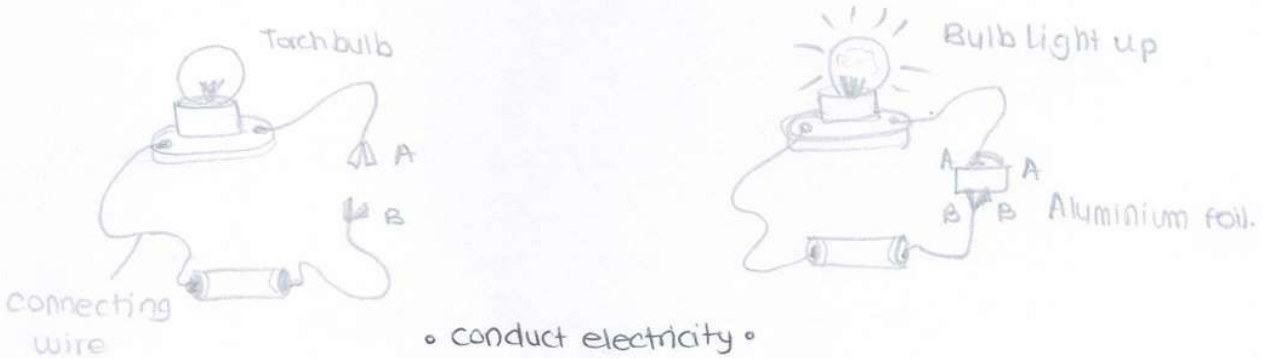
- Metals are ductile, that is metals turns into thin wires -

Ductility - The property which allows the metal to be drawn into thin wire

- Gold is the most ductile metal and silver are among the best ductile metal.
- Copper and aluminium metals are very ductile and drawn into thin copper and Aluminium wires.
- Magnesium metals are used in experiment in the laboratory and Tungsten metal are used for making the filament of electric bulbs.

- Metals are good conductor of heat -
- Metal allows heat to pass through them easily.
- Experiment -
- Silver metal is the best conductor of heat. It has highest thermal conductivity
- copper and aluminium metal are also very good conductor of heat.
- The cooking utensil and water boilers are usually made of copper or aluminium metals bcoz they are good conductor of heat.
- Poorest conductor of heat - Lead, Mercury

- Metals are good conductor of electricity -
- Metals allows electricity to pass through them easily.
- Demonstrated as follows -
- we take a dry cell, a torch is fitted in a holder and some connecting wire with crocodile clips and connected them to make an electric circuit.
- There is gap between the ends of the crocodile clips A and B so no current flows in incomplete circuit and hence does not glow (bulb).
- Lets us now insert a piece of aluminium foil between the ends of crocodile clips A and B.
- We will now insert a piece of aluminium foil between the end of clip the buld light up at once.
- The Aluminium foil allow electric current to pass through it.
- This conducts the electricity, so we conclude that the aluminium metal is good conductor of electricity.
- Silver metal is the best conductr of electricity and copper metal is the best conductor of electricity.
- Electrical conductivity is the characteristic of property of metals.



- conduct electricity -

5. Metal are lustrous or shiny and can be polished -
- lustrous means having a shining surface, the shiny apperance of the substance is called 'metallic lustre'.
- The shiny apperance of metals makes them useful in making jewelery and decoration pieces.
- The metal lose their shine or brightness on keeping in air for long time due to formation of a thin layer of oxide, carbonate and sulphide on their surface.

6. Metals are generally hard -
- Most of the metals like iron, copper and aluminium are very hard they cannot cut with knife.
- Sodium and potassium are soft metals which can be easily cut with knife. (Exception)

- Metals are strong -
- Metal can hold large weight without snapping (without breaking).
- Exception - Except sodium and potassium metal which are not strong
- iron metal used in construction of bridges, buildings and railway line.

- Metals are solids at room temperature -
- Metals a. iron, b. copper, c. silver, d. gold are solids at room temp. only one metal, Mercury is in liquid state at room temperature.

- Metals have high melting points and boiling point -
- Iron metal has high melting point of 1535°C, this means that solid iron melts and turns into liquid on heating to a high temp.
- Copper metal has high melting point of 1083°C.
- Exceptions - Sodium and potassium metal has low melting point of 98°C and 64°C.
- Gallium and cesium metal also has low melting point - 30°C and 28°C.

- Metal are sonorous, that is metal makes sound when hit an object.
- Sonorous means capable of producing a deep or ringing sound.
- The property of metals of being sonorous is called sonority.
- It is due to the property of sonorousness that metals are used for making bells and strings of musical instruement like sitar and violin

- Metals usually have a silver or grey colour -
- Copper has red-dish brown colour whereas gold has yellow colour.

- PHYSICAL PROPERTIES OF NON-METAL -
- Non metal are neither malleable nor ductile. Non metal are brittle -
- Non metal are not malleable, they cannot beaten into thin sheets. nor drawn into wires.
- Ex - Sulphur and phosphorus are beaten with hammer, they break into pieces, they do not form thin sheet.
- Carbon is a solid non-metal which is brittle, the brittleness is property of non metal.

- Non metal do not conduct heat and electricity -
- Bcoz, they have no electrons.
- Carbon in form of graphite is the only non-metal which is good conductor of electricity, it is used for making electrode.
- Sulphur do not conduct heat and electricity.

- Non metals are not lustrous. They are dull -
- Non metal do not have lustre, which means that non metal do not have a shining surface. and solid non-metal have a dull apperance
- Iodine is a non-metal having lustrous apperance. it has shining surface.

- Non metals are generally soft -
- Non metal are quite soft. sulphur and phosphorus are solid non-metal, which are soft. (Exception).
- only Diamond is very hard and it is the hardest natural substance.
- Non metal are not strong. They are easily broken -
- Graphite is a non-metal which is not strong. It has low strength so, when a large weight is placed on graphite sheet.
- Non metal may be solid, liquid or gases at room temp.
- sulphur and phosphorus - solid non-metal
- Bromine - a liquid non-metal
- Hydrogen, oxygen - Nitrogen and chlorine - gaseous non metal,
- Non metal have low melting points and boiling points -
- Melting point of sulphur - 115°c quite low.
- Melting point of Diamond - more than 3500°c, very high.
- Non metals are non-sonorous. They do not produce sound when hit with an objects.
- Non metals have many different colours - sulphur - yellow phosphorus - white or red, graphite - black, chlorine - yellowish green, oxygen and hydrogen - colourless.
- CHEMICAL PROPERTIES OF METAL -
- Reaction of metals with oxygen -
- When metal are burnt in air, they react with oxygen of air to form metal oxides -
  - Metal + oxygen ⟶ Metal oxide (Basic oxide)
- Metal oxides, being basic turns red litmus solution blue.
- The vigour of reaction with oxygen depends on the chemical reactivity of metal -
- Sodium metal reacts with the oxygen of air at room temperature to form a basic oxide called sodium oxide.
  - $4Na(s) + O_2(g) \longrightarrow 2Na_2O(s)$
    sodium oxygen sodium oxide
- Why potassium and sodium metal are stored under Kerosene oil to?
- The potassium and sodium metal are so reactive that they react vigorously with oxygen (of air).
- They catch fire and start burning when kept in air.
- So, it is kept in Kerosene to prevent their reaction with the O_2, moisture and CO_2 of air.

- Most of the metal oxides are insoluble in water. But some of the metal oxides dissolve in water to form alkalis.
- Sodium oxide and potassium oxide are the two metal oxide dissolve in water which are soluble in water.
- Sodium oxide is a basic oxide which react with water to form an alkali called sodium hydroxide.

$$Na_2O(s) + H_2O(l) \longrightarrow 2NaOH(aq)$$

Sodium oxide (Basic oxide) + water → sodium hydroxide (An alkali)

- Due to formation of NaOH alkali, a solution of sodium oxide in water turns red litmus to blue.
- Potassium oxide is also a basic oxide which reacts with water to form an alkali called potassium hydroxide.

$$K_2O(s) + H_2O(l) \longrightarrow 2KOH(aq)$$

Potassium oxide (Basic oxide) + water → Potassium hydroxide (An alkali)

- Turns red litmus to blue.
- Aluminium metal burns in air, on heating to form aluminium oxide-

$$4Al(s) + 3O_2(g) \longrightarrow 2Al_2O_3(s)$$

Aluminium (metal) + oxygen → Aluminium oxide (Amphoteric oxide)

- Aluminium is less reactive than magnesium.
- **Amphoteric oxide-**
- Those metal oxides which shows basic as well as acidic behaviour are called Amphoteric oxide.
- Aluminium metal and zinc metal form amphoteric oxide, aluminium oxide and zinc oxide are amphoteric in nature.
- React with both acids as well as bases to form salt and water.

a. **Aluminium oxide reacts with HCl to form aluminium chloride and water-**

$$Al_2O_3(s) + 6HCl(aq) \longrightarrow 2AlCl_3(aq) + 3H_2O(l)$$

Aluminium oxide + Hydrochloric acid (Acid) → Aluminium chloride (Salt) + water

- In this reaction, aluminium oxide behave as basic oxide

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

• $Al_2O_3(s) + 2NaOH(aq) \longrightarrow 2NaAlO_2(aq) + H_2O(l)$

Aluminium oxide — sodium hydroxide (Base) — Sodium aluminate (salt) — water

• In this reaction aluminium oxide behave as an acidic oxide.

• zinc metal burns in air only on strong heating to form zinc oxide -

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

zinc — oxygen — zinc oxide

• zinc is less reactive than aluminium.

• zinc oxide reacts with hydrochloric acid to form zinc chloride and water -

a. • $ZnO(s) + 2HCl \longrightarrow ZnCl_2 + H_2O$

zincoxide — Hydrochloric acid (Acid.) — zincchloride (salt) — water

• In this reaction, zinc oxide behaves a basic oxide.

b. zinc oxide reacts with sodium hydroxide to form sodium zincate and water -

• $ZnO + 2NaOH \longrightarrow Na_2ZnO_2 + H_2O$

• zinc oxide — • sodium hydroxide (Base) — sodium zincate (salt) — water

• In this reaction, zinc oxide behaves as an acidic oxide.

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

$3Fe(s) + 2O_2(g) \longrightarrow Fe_3O_4(s)$

Iron — oxygen — Iron (II, III) oxide

• The reaction between iron with oxygen takes place readily than that of zinc, so iron is less reactive than zinc.

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

• Metal react with water to form a metal hydroxide and hydrogen gas.

a. when a metal react with water, then produce metal hydroxide and hydrogen gas.

Metal + water $\longrightarrow$ Metal hydroxide + Hydrogen

b. Metal reacts with steam, then produces products such as metal oxide and hydrogen-

- Metal + Steam ⟶ Metal oxide + Hydrogen

- Forexample-
- Potassium reacts violently with cold water to form potassium hydroxide and hydrogen gas -

- $2K(s) + 2H_2O(l) \longrightarrow 2KOH(aq) + H_2(g) + Heat$

- Potassium
- water
- Potassium hydroxide
- Hydrogen

- It is highly exothermic due to which the hydrogen gas formed during the reaction catches fire immediately.
- Sodium reacts vigorously with cold water forming sodium hydroxide and hydrogen gas.

- $2Na(s) + 2H_2O(l) \longrightarrow 2NaOH + H_2 + Heat$

- Sodium
- water (cold)
- Sodium hydroxide
- Hydrogen

- It is highly exothermic due to which the hydrogen gas formed during the reaction catches fire., bcoz sodium is also a very reactive metal.
- calcium reacts with cold water to form calcium hydroxide and hydrogen gas -

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

- Calcium
- water (cold)
- Calcium Hydroxide
- Hydrogen

- The piece of calcium metal starts floating in water bcoz of the bubble the hydrogen gas formed during the reaction stick to surface.
- The reaction of calcium metal with water is less violent.
- calcium is less reactive than sodium.

- Metals like aluminium, zinc and iron do not react with either cold water or hot water. they react with steam to form a metal oxide and H_2.
- Aluminium reacts with steam to form aluminium oxide and hydrogen gas-

- $2Al(s) + 3H_2O(g) \longrightarrow Al_2O_3(s) + 3H_2(g)$

- Aluminium
- steam
- Aluminium oxide
- Hydrogen

Aluminium metal does not react with water under ordinary condition bcoz of the presence of thin layer of aluminium oxide.

- zinc reacts with steam to form zinc oxide and hydrogen-

$Zn(s) + H_2O(g) \longrightarrow ZnO(s) + H_2(g)$

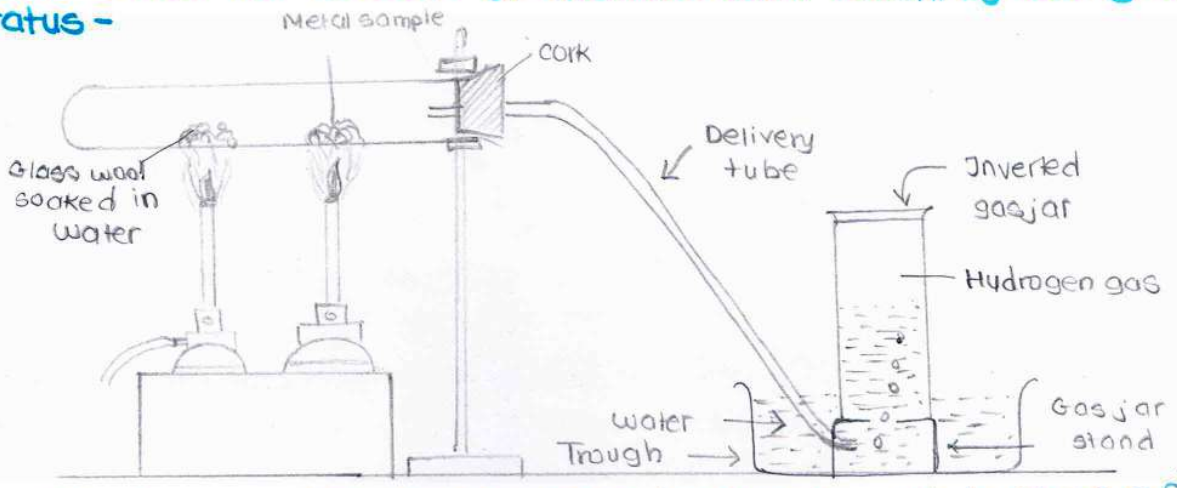
zinc — steam — zinc oxide — Hydrogen

- Red-hot iron reacts with steam to form iron(II, III) oxide and Hydrogen.

- $3Fe_{(s)} + 4H_2O_{(g)} \longrightarrow Fe_3O_{4(s)} + 4H_{2(g)}$

  Iron — Steam — iron(II, III) oxide — Hydrogen

- We can study the reaction of metals with steam by using the apparatus -



- A lump of glass wool soaked in water is placed at bottom of a tube The water in glass wool will form steam on heating.
- The sample of metal is placed in the middle of the horizontally kept boiling tube. The boiling tube containing water, soaked glass wool and metal sample is arranged in apparatus.
- To start the experiment, the metal sample is heated by using a burner when the metal gets hot, then the glass wool is heated by using other burner.
- The water present in glass wool forms steam on heating. this steam then passes over the hot metal. The metal react with steam to form the metal oxide and H_2 gas.
- The H_2 gas come out of the boiling tube and it is collected over water when a lighted match stick is applied to gas collected jar, the gas burns with a pop sound. indicating that it is hydrogen. The metal oxide formed remains behind in the boiling tube.
- This experiment is performed by taking magnesium, aluminium, zinc and iron as metal sample.
- It is found that the reaction of steam with magnesium is most vigorous followed by mx with aluminium and zinc, but mx with iron is slow.
- Magnesium is very reactive whereas the iron is least reactive.
- Decreasing order - $Mg > Al > Zn > Fe$
- Metals like lead, silver and gold do not react with water or even steam.
- Only those metals displace hydrogen from water which are hydrogen in the reactivity series.

- Reaction of metals with Dilute Acids -
- Metals usually displace hydrogen from dilute acids. only less reactive metals like copper, silver and gold do not displace hydrogen from dilute acids.

Metal + Dilute acid ⟶ Metal salt + Hydrogen

- All metals, do not react with dilute acids. the vigour of rx of metal with dilute acids depend on chemical reactivity.
- Metal react with dil HCl to give metal chloride and hydrogen gas -
- Sodium metal react violently with dil. HCl to form sodium chloride and hydrogen gas.

$2Na(s) + 2HCl(aq) \longrightarrow 2NaCl(aq) + H_2(g)$

Sodium — Hydrochloric acid — Sodium chloride — Hydrogen

- The reaction shows that sodium metal is very reactive.
- Zinc reacts with dil. HCl acid to give zinc chloride and hydrogen gas -

$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$

zinc — Hydrochloric acid — zinc chloride — Hydrogen.

- Zinc is less than reactive aluminium
- Note - Silver and gold also do not react with dilute acids.
- How metals displaces hydrogen from dilute reacts - **(2M | 3M)**
- All those metals which are more reactive than hydrogen, that is those metal which lose electrons more easily than hydrogen displace hydrogen from dilute acids to produce hydrogen gas.
- The metals like copper and silver which are less reactive than hydrogen do not displace hydrogen from dilute acids. bcoz they do not give out electrons required for reduction of hydrogen ions present in acid.
- All the metal which are above hydrogen in the activity series, displace hydrogen from dilute acids.
- Those metals which are below hydrogen in activity series, do not displace hydrogen from dilute acids.
- Metals reacts with dil sulphuric acide to give metal sulphate and H_2 gas -

$2Na(s) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + H_2(g)$

Sodium — sulphuric acid — sodium sulphate — Hydrogen

$Cu + H_2SO_4 \longrightarrow CuSO_4 + H_2$ No reaction

copper — sulphuric acid

$Fe + H_2SO_4 \longrightarrow FeSO_4 + H_2$

Iron — sulphuric acid — Iron(II) sulphate — Hydrogen.

- When a metal reacts with dil. nitric acid, then hydrogen gas is not evolve.
- Explanation - Nitric acid is oxidising agent so, as soon as hydrogen gas is formed in the reaction between a metal and dil. nitric acid, the nitric acid oxidises this hydrogen to water.
- When nitric acid oxidises hydrogen to water, then nitric acid itself reduced to any of the nitrogen oxide.

- Dil. nitric acids, however reacts with magnesium and manganese metals to evolve hydrogen gas - this is bcoz the dil. nitric acid is a weak oxidising agent which is not able to oxidise hydrogen to water.

a. Magnesium reacts with very dil. nitric acid to form magnesium nitrate and hydrogen gas.

- $Mg(s) + 2HNO_3(aq) \longrightarrow Mg(NO_3)_2(aq) + H_2(g)$

Magnesium nitric acid magnesium nitrate Hydrogen.

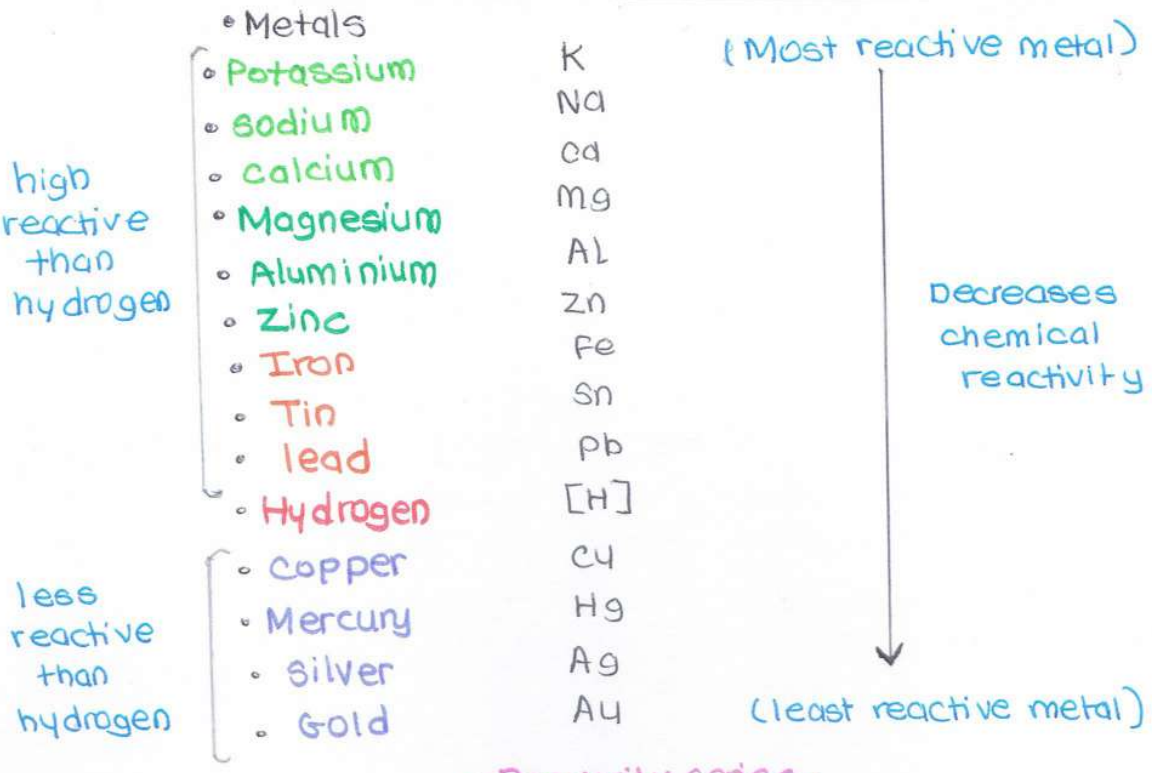
b. Manganese reacts with dil. nitric acid to form managanese nitrate and hydrogen gas.

$Mn(s) + 2HNO_3(aq) \longrightarrow Mn(NO_3)_2(aq) + H_2(g)$

Manganese dil nitric acid Manganese nitrate hydrogen

- **Aqua - Regia -**
- Aqua - regia is freshly prepared mixture of 1 part of conc. nitric acid and 3 part of conc. HCl.
- Ratio - conc. HNO_3 : conc HCl - 1:3, it is a highly corrosive fuming liquid.
- Aqua regia can dissolve all metals.
- Aqua regia can dissolve even gold and platinum metals.

- **The reactivity series of metal -**
- Some metals are chemically very reactive whereas others are less reactive or unreactive.
- The arrangement of metals in a vertical column in order of decreasing reactivites is called reactivity of series of metals.
- The reactivity series, the most reactive metal is placed at the top whereas the least reactive metal is placed at bottom.
- As we come down in the series, the chemical reactivity of metal decreases.
- Most reactive metal - Potassium
- least reactive metal - Gold
- Hydrogen is also placed in reactivity series, as it is not a metal. it is usually places as it loses electrons and forms +ve ions.



• Reactivity series •

Q. Why some metals are more reactive and others less reactive

• When metal reacts, they lose electrons to form positive ions hence it will be a reactive metal.

• When metal reacts, they lose electrons less readily to form positive ions, it will react slowly. such metal is less reactive.

• Ex- Sodium, Potassium

Q. Why some metals are more reactive than hydrogen?

• Those metal which lose electrons more readily than hydrogen are said to be more reactive than hydrogen.

• The metal above hydrogen are lose electrons more readily than H_2 and hence they are more reactive

Q. why some metals are less reactive than hydrogen?

• Those metal which lose electrons less readily than hydrogen are said to be less reactive. than hydrogen.

• All metal below hydrogen are in reactivity series, lose electrons less readily than hydrogen, they are less reactive than hydrogen.

• Note- We should remember the reactivity series of metals to decide whether a particular displacement rnx will take place or not.

• **Reaction of metal with salt solutions**

• When a more reactive metal is put in the salt-solution of a less reactive metal, then the more reactive metal displaces the less reactive metal from its salt solution.

• A more reactive metal displaces less reactive metal from its solution.

• **Salt sol. of metal B + Metal A → Salt sol. of A + Metal B**

- The reaction of zinc with copper sulphate solution -
  - $CuSO_{4(aq)} + Zn_{(s)} \longrightarrow ZnSO_{4(aq)} + Cu_{(s)}$

copper sulphate (Blue) | zinc (silvery white) | zinc sulphate (colourless) | copper (Red brown)

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

- **Reaction of copper with silver nitrate solution -**
- When a strip of copper metal is kept immersed in silver nitrate for some time, the solution graudually becomes blue and shining greyish white deposit of silver metal form on copper strip.
  - $2AgNO_{3(aq)} + Cu_{(s)} \longrightarrow Cu(NO_3)_{2(aq)} + 2Ag_{(s)}$

silver nitrate (colourless) | copper (Red brown) | copper nitrate (Blue) | silver (Greyish white)

- This is because silver is less reactive than copper and hence cannot displace copper from copper nitrate solution.

- **Reaction of metals with chlorine -**
- Metals react with chlorine to form ionic chlorides because they can give electrons to chlorine atoms to form ions.
- Metal chlorides have high melting and boiling point. metal chlorides are usually solid and conduct electricity in the solution in molten state.
- Metal chloride are weak electrodytes. and are non volatile.
- Sodium is a metal. so sodium reacts with chlorine to form an ionic chloride.
  - $2Na_{(s)} + Cl_{2(g)} \longrightarrow 2NaCl_{(s)}$

Sodium | chlorine | sodium chloride

- This is an ionic compound or electrovalent compound containing sodium ion Na^+ and chloride ions (Cl^-).
- calcium is a metal which react with chlorine to form an ionic chloride called calcium chloride -
  - $Ca_{(s)} + Cl_{2(g)} \longrightarrow CaCl_{2(s)}$

calcium | chlorine | calcium chloride

- Magnesium on heating with chlorine form magnesium chloride which is an ionic chloride -

$Mg_{(s)} + Cl_{2(g)} \longrightarrow MgCl_{2(s)}$

Magnesium | chlorine | magnesium chloride

- **All these metal chlorides are ionic compounds.**

5. Reaction of metals with hydrogen-

- Metals generally do not react with hydrogen bcoz metals form compounds by losing electrons and hydrogen also form compound by losing electron.
- Most of the metals do not combine with hydrogen. only a few reactive metals like sodium, potassium and magnesium react with hydrogen to form metal hydrides.
- In a metal hydride, the hydrogen present in the form of a negative ion (anion) called hydride ions H^-.
- when hydrogen gas is passed over heated sodium, then sodium hydride is formed:

$$2Na(s) + H_2(g) \longrightarrow 2NaH(s)$$

Sodium hydrogen sodium hydride

- NaH is an ionic compound containing sodium ions (Na^+) and hydride ions, H^-. When hydrogen gas is passed over heated potassium, then (KH) is formed. Potassium hydride is also an ionic hydride.

- CHEMICAL PROPERTIES OF NONMETAL-

- **Reaction of non metal with oxygen-**
- Non metal reacts with oxygen to form acidic oxides or neutral oxides.
- carbon forms an acidic oxide CO_2 and hydrogen form a neutral oxide. H_2O.
- The non-metal oxides are covalent in nature which are formed by sharing of electrons. The acidic oxides of non-metals dissolve in water to form acids.
- The acidic oxides of non metal turns blue litmus solution to red.
- Carbon is non-metal. when carbon burns in air it reacts with the oxygen of air to form an acidic oxide called carbon dioxide.

$$C + O_2 \longrightarrow CO_2 \quad \text{(Acidic oxide)}$$

carbon oxygen carbondioxide

The acidic oxide, carbon dioxide, dissolve in water to form an acid called carbonic acid.

$$CO_2(g) + H_2O(l) \longrightarrow H_2CO_3(aq)$$ Acidic in nature

carbondioxide water carbonic acid.

- Turns blue litmus to red.

- sulphur is a nonmetal. when sulphur is burned in air, it reacts with the oxygen of air to form an acidic oxide called sulphur dioxide.

$$S(s) + O_2(g) \longrightarrow SO_2(g)$$

sulphur oxygen sulphurdioxide

• $SO_2(g) + H_2O(l) \longrightarrow H_2SO_3(aq)$

• sulphur dioxide • water • sulphurous acid

• Turns blue litmus to red •

• This shows that it is acidic in nature.

∴ The non metal oxides like CO_2, SO_2 turns blue litmus solution red, showing that they are acidic in nature these acidic oxides are called acid anhydrides.

• Phosphorus is also an non-metal which reacts with oxygen to form acidic oxide, phosphorous pentoxide (P_2O_5)

• $P + O_2 \longrightarrow P_2O_5$

phosphorus oxygen phosphorous pentoxide

• $4P + 5O_2 \longrightarrow 2P_2O_5$

• Non metal oxides which are neutral, being acidic or basic –
Ex- carbon monoxide, water, nitrogen monoxide, dinitrogen monoxide.

• These oxides do not turn blue litmus to red litmus. or red to blue litmus

• Carbon is a nonmetal. when carbon burns in an supply of oxygen then it forms and neutral oxide called carbon monoxide-

• $2C(s) + O_2(g) \longrightarrow 2CO(g)$ (neutral)

• carbon • oxygen • carbon monoxide

• Does not produce an acid with water.

• Hydrogen is an non metal. when hydrogen combines with the oxygen of air, then it forms an neutral oxide.

• $2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$

• Hydrogen • oxygen • water

• Water is actually a hydrogen oxide, Formed by sharing of electrons so they are covalent compounds. they not contain any oxide ion.

• **Reaction of non metal with water –**
Non metals do not react with water to evolve hydrogen gas bcoz non-metal cannot give electrons to reduce the hydrogen ions of water into hydrogen gas.

• Reaction of non metal with salt solutions –

• A more reactive nonmetal displaces a less reactive non metal from its salt solution.

• when chlorine is passed through a solution of sodium bromide, then sodium chloride and bromine.

• $2NaBr(aq) + Cl_2(g) \longrightarrow 2NaCl(aq) + Br_2(aq)$

sodium bromide chlorine sodium chloride Bromine

- Reaction of non metal with chlorine -
- Non metals reacts with chlorine to form covalent compounds which is chloride which are non electrolytes do not conduct electricity
- Non metal chloride are usually liquids or gases.
- Hydrogen reacts with chlorine to form a covalent chloride called hydrogen chloride -
  - $H_2 + Cl_2 \longrightarrow 2HCl$ (covalent chloride)
    Hydrogen chlorine Hydrochloric acid
- **Phosphorous reacts with chlorine to form a covalent chloride called phosphorous chloride**
  - $P_4(s) + 6Cl_2(g) \longrightarrow 4PCl_3(l)$ (covalent chloride)
    Phosphorus chlorine Phosphorous trichloride
- Non metals form covalent chloride because they cannot give electrons to chlorine atoms to form chloride ions.

- Reaction of nonmetal with Hydrogen -
- Non metals react with hydrogen to form covalent hydrides. The hydrides are formed by the sharing of electrons that is non-metal hydrides are formed by covalent bonding.
- Sulphur is a nonmetal which combines com with hydrogen to form a covalent hydride called hydrogen sulphide
  - $H_2(g) + S(l) \longrightarrow H_2S(g)$ (covalent hydride)
    Hydrogen sulphur Hydrogen sulphide
- The H_2S gas has a characteristics smell resembeling the smell of rotten eggs.
- Non metals form covalent hydrides because non metal atoms cannot give electrons to hydrogen atoms to form hydride ions.
- Non metal hydrides are liquids or gases. do not conduct ions, they do not conduct electricity and are stable compounds.

- Reaction of non-metal with Dilute Acids -
- Non metal do not react with dilute acids, they do not displace hydrogen from acids because in order to displace hydrogen ions of an acid and convert them into hydrogen gas, electrons should be supplied to the hydrogen of the acid. It's (nonmetal) being itself an acepptor of electrons, cannot give electrons to the hydrogen ions of the acid to reduce them to H_2 gas.
- If non metals like carbon, sulphur or phosphrous are put into test tube containing dil. H_2SO_4, then no hydrogen gas is evolve.

Uses of metal

- copper and aluminium metals are used to make wire to carry electric current
- Iron is used as a catalyst in the prep. of ammonia gas by Haber's process
- Iron, copper and Aluminium metals are used to make household utensils
- zinc is used as galvanizing iron to protect it from rusting
- Lead metal is used in making car battery
- Mercury used in thermometer
- zirconium metal is used for making bullet proof alloy steels.
- silver and gold used in making jewellery

Uses of Nonmetal

- Carbon is used for making the electrodes of electrolytic cells and dry cells.
- Hydrogen is used as hydrogenation of vegetable oil to make vegetable ghee
- Nitrogen used to preserve food materials
- Hydrogen is used in the manufacture of ammonia
- Nitrogen is used in the manufacture of ammonia and fertilisers
- Liquid hydrogen used as a rocket fuel
- Sulphur is used for manufacturing sulphuric acid
- sulphur is used in the vulcanisation of rubber
- sulphur is used as a fungicide and in making gunpowder.

- How do metal and nonmetal react -
- When metal react with nonmetals, they form ionic compounds.
- When non-metal react with other non-metal, they form covalent compound
- **Chemical bond** - The force which links the atoms in a molecule is called a chemical bond.
- Noble gases are also called inert gases (bcoz they are chemically very inert or unreactive).

- **Inertness of Noble Gases** -
- There are some elements in group 18 of the periodic table which do not combine with other elements - Helium, Argon, Neon, krypton, Xenon and Radon they are known as noble elements or gases because they are unreactive and do not react with other elements.
- Inert gases do not form chemical bonds. The noble gases are chemically unreactive, bcoz of the stable arrangement of atoms which do not allow the outermost electrons to take part in chemical reaction.

- Electronic configuration of Noble Gas -

• Noble gas	• Symbol	• Atomic no.	• Electronic config.	• No. of e^- in outermost shell.
1. Helium	He	2	2	2
2. Neon	Ne	10	2,8	8
3. Argon	Ar	18	2,8,8	8
4. Krypton	Kr	36	2,8,18,8	8
5. Xenon	Xe	54	2,8,18,18,8	8
6. Radon	Rn	86	2,8,18,32,18,8	8

- Here only one has - Helium - 2 electrons in outermost shell and all others have 8 electrons in outermost shells of their atoms.
- Outermost shell electrons are more stable arrangement of electrons. Atoms with 8 electrons are more stable and unreactive.
- Note - Though 8 electrons in the outermost shell always impart stability to an atom, but 2 electrons in the outermost shell impart stability only when the outermost shell is the first shell and no other shells are present in the atom.
- "8 electrons" in the outermost shell of an atom known as "octet of electrons"
- "2 electrons" in the outermost shell, outermost K shell is known as "duplet of electrons.
- The usual number of electrons in the outermost shell of the atom of a noble gas is 8. only in the case of one noble gas helium, the no. of outermost electrons is 2.

• we can explain the reactivity of elements as a tendency of their atoms to achieve a completely filled outermost shell or valence shell. and becomes stable.

• **Cause of chemical Bonding -**

• Everything in this world wants to be stable. stability means having the electron arrangement and become more stable.

• Atoms form chemical bonds to achieve stability by acquiring the inert gas electron configuration.

• The atoms get 8 electron in valence shell are stable

• The atoms less than 8 electron in valence shell are unstable

• All atoms have tendency to achieve the inert gas electron arrangement of 8 electrons or 2 electrons in the outermost shell and become more stable.

• An atom can achieve the inert gas electron arrangement in 3 ways-

• By losing one or more electrons.

• By gaining one or more electrons.

• By sharing one or more electrons.

• **Valence shell -** The outermost electron shell of an atom is known as valence shell.

• **Valence electrons -** The outermost electrons are also known as valence electrons

• **IONS -**

• An ion is an electrically charged atom.

Ex- sodium ions Na^+, chloride ions Cl^-, oxide ion O^{2-}

• An ion is formed by loss or gain of electrons by an atom, so it contains an unequal number of electrons and protons.

• **2 types -** • cation • Anion

• CATION - A positively charged ion known as cation

• Ex- sodium ions Na^+, magnesium ions Mg^{+2} are cation bcoz they are "+vely" charged.

• A cation is formed by the loss of one or more electrons by an atom.

• For example- sodium atom loses 1 electron to form a sodium ion Na^+.

$$Na - e^- \longrightarrow Na^+$$

Sodium — electron — sodium ion

• cation contain less electron than protons.

• It is formed by removal of electrons from an atom.

• cation is for med by loss of one or more electrons by an atoms.

• The ions of all metal element are called cation.

• cation

- A negatively charged ions is known as anion.
- Forexample- chloride ion Cl^- and oxide O^{2-} are anion bcoz they are negatively charged ions'.
- An anion is formed by gaining of one or more electrons by an atom.
- Ex- A chlorine atoms gains (accepts) 1 electron to form an chloride ion.

$$Cl + e^- \longrightarrow Cl^-$$

chlorine atom Electron Chloride ion.

- An anion contains more electrons than normal atom
- An anion containg more electrons than protons.
- The ions of all the nonmetal elements are anions (except H ion and ammonium ions)

- **FORMATION OF IONS - CATIONS**

If an element has 1, 2, 3 electrons in the outermost shell of its atoms, then it loses these elements to achieve the inert gas e^- arrangement of eight valence electrons and forms positively charged ion.

- Formation of sodium ions, Na^+-
- Atomic No - 11, so atom contains 11 electrons.
- The electronic configuration of sodium is **K-2 L-8 M-1.**
- Sodium atoms has $1e^-$ in its outermost shell. it is not stable arrangement bcoz it does not have $8\ e^-$ in the outermost shell.
- In order to become more stable a sodium atom donates its outer 1 atom to get stable.
- The M shell, the e^- are removed and L shell becomes an outermost shell. By losing $1e^-$, the sodium atom gets 1 unit of +ve charge and become a sodium ion, Na^+.

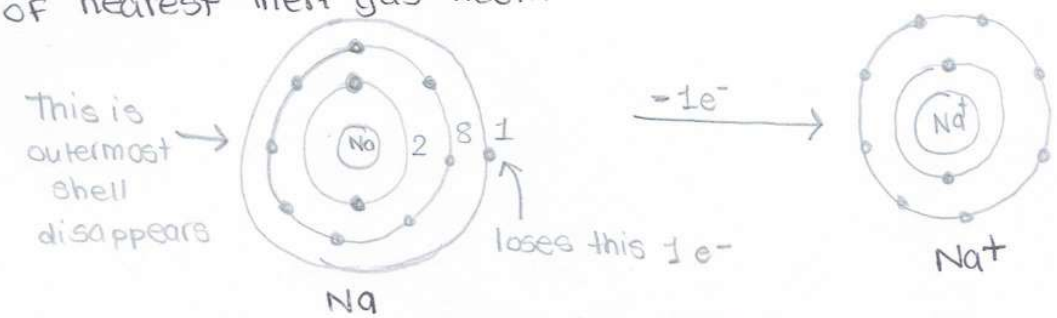
- anion

$$Na - e^- \longrightarrow Na^+$$

Sodium atom Sodium ion.

Electronic confi -	K L M	K L (same as neon)
	2, 8, 1	2, 8
	unstable	stable

- The electronic configuration of the sodium ion is the same as that of nearest inert gas neon.



• The formation of potassium ion (K^+) is similar to the formation of sodium ion because like sodium atom, the potassium atom has 1 electron in its outermost shell.
• The atomic no. of K is 19. Electronic configuration is K L M N
The noble gas nearest to potassium is argon having 2 8 8 1
18 and electronic configuration of K L M.
2 8 8

• **Formation of an Aluminium Ion, Al^{+3} -**
The atomic number of Aluminium is 13, so its electronic configuration K L M. The Al^{+3} atom has 3 electrons in its outermost shell which 2 8 3 it denates to some other atom and Al^{+3} having 3 units of positive charges.

$$\underset{\text{Aluminium}}{\bullet\ Al} - \underset{\text{electron}}{3e^-} \longrightarrow \underset{\text{Aluminium ion}}{Al^{3+}}$$

• K L M
2 8 3

K L
2, 8

• The electronic confiugration of an aluminium ion is the same as that of its nearest noble gas neon.
• Aluminium ion has 13 protons but only $10e^-$, that is it has 3 protons more than electrons. since an Al^{+3} has 3protons more than electrons, it has 3 units of positive charge.

• **Electron - Dot - Representation-**
• The valence electrons in an atom are represented by putting (•) on the symbol of the element, one dot for each valence electron.
• Sodium atoms has 1 valence electron in its outermost shell, so we put the one dot with symbol of sodium and write Na•. sodium atom loses this 1 electron to form a sodium ion. so we write - Na^+
• Sodium atom loses this $1e^-$ to form sodium ion.
• Magnesium - atom has 2 valence electrons so we write. Mg:
• Aluminium - atom has 3 valence electrons so we write - •Al:
• Potassium - atom has 1 valence electron so we write - K•

• **FORMATION OF NEAGTIVE IONS (OR) ANIONS-**
• If an element has 5, 6, 7 electron in the outermost shell of its atom, then it gains electrons to achieve the stable, ineft gas electron configuration of 8 valence electrons and form -vely charged ion called anion.
• Non metal atoms have usually 5, 6 or 7 electrons in their outermost shell, so the non-metal atoms accept electrons to form negative ions or anions.
• Fluorine, chlorine, bromine, iodine, oxygen are non-metals which accepts electron to form negative ions.

• The element carbon having four electrons in its outermost shell, is also an nonmetal but it neither lose 4 e^- nor gets gain 4 e^- due to energy considerations, so a carbon atom does not form ions.

• Note - An atom having 7 electrons in its outermost shell accepts 1 more electron to form an anion having 1 units negative charge.

• An atom having 6 electrons in its outermost shell accepts 2 more e^-'s to form an anion having two unit negative charge.

• **Formation of chloride ion, Cl^-**

• Atomic no - 17 , Electronic configuration - K L M
2 8 7

• Chlorine atom has 7 electrons in the outermost shell. It needs 1 more electron to achieve the stable configuration of an inert gas.

• In order to become stable, atom accepts 1 electron from other atom.

• $Cl + e^- \longrightarrow Cl^-$

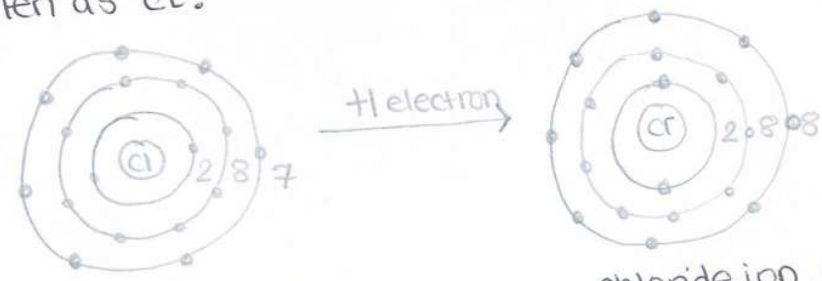
	Chlorine ion	Electron	Chloride ion
Electronic config -	K L M 2, 8, 7 unstable	• Gain one electron	K L M 2, 8, 8 stable

(same as Argon gas)

• It is stable than chlorine atom, it is the same as that its of nearest inert gas **argon**.

• A chlorine, the no. of protons and electrons in a chlorine atom is equal, therefore it is electrically neutral, having no overall charge.

• Due to 1 more electron than protons, a chloride ion has 1 unit -ve charge. written as Cl^-.



Cl, chlorine atom

Chloride ion, Cl^-

• **The formation of Fluoride ion (F^-) is similar to the formation of chloride ions** because like chlorine atom, a fluorine atom (F) also has 7 electron in the outermost shell.

• The atomic number of Fluorine is 9, and its electronic configuration is K L . The inert gas nearset to Fluorine is neon having an electron and 2, 7 atomic number of 10 and electronic configuration K L. The other halogen, bromine and iodine have also 7 valence electrons 2 8 each in their atoms and accepts 1 e^- each to form bromide ion (Br^-).

- **Formation of an oxide Ion O^{2-} -**
- Atomic no. of oxygen - 8, so its electronic configuration K L 2, 6
- The oxygen atom has 6 electrons in its outermost shell, it needs 2 more electrons to achieve the stable.
- By taking 2 electrons from some other atom, the oxygen atom form an oxide ion, O^{2-}, having 2 units of negative charge.

$$O + 2e^- \longrightarrow O^{2-}$$

- O: oxygen atom — K L: 2, 6 (unstable)
- $2e^-$: electrons
- O^{2-}: oxide ion — K L: 2, 8 (stable) (same as neon atom)

- since an oxide ions has 2 electrons more than protons, it has 2 units of negative charge (as written as O^{2-}).
- **The formation of a sulphuide ion (S^{2-}) is similar to the formation of an oxide ion** - because like an oxygen atom, a sulphur atom has 6 e^- in its outermost shell.
- The atomic number of sulphur is 16. and electronic configuration K L M. 2 8 6
- The atomic number of 18 and electronic configuration. K L M 2 8 8. (Argon)

- **Formation of Nitride ion, N^{3-}**
- Atomic no - nitrogen - 7 • Electronic configuration - K L 2 5. Nitrogen atom has 5 valence electrons, so it needs more 3 electrons to get stable. So by taking 3 electrons from other atom, a nitrogen atoms forms a nitride ion N^{3-} having 3 unit of negative charge.

$$N + 3e^- \longrightarrow N^{3-}$$

- N: Nitrogen — K L: 2, 5 — unstable
- $3e^-$: electron
- N^{3-}: Nitride ion — K L: 2, 8 — stable (same as neon)

- The electronic configuration of a nitride ion is same as that of neon atom. since a nitride ion contain 3 electrons more than protons it has 3 units of negative charge.
- **The formation of phosphide ion (P^{3-}) is similar to the formation of nitride ion.** because like a nitrogen atom, a phosphorus atom has 5 electrons in the outermost shell.
- Atomic number is 15 and electronic configuration is K L M. 2 8 5

• Electron Dot representation –

Non-metal	Symbol	Atomic no.	K L M	No. of outermost e⁻	Electron dot struc.	• Ion formed
• Fluorine	F	9	2, 7	7		
• chlorine	Cl	17	2, 8, 7	7		
• oxygen	O	8	2, 6	6		
• sulphur	S	16	2, 8, 6	6		
• Nitrogen	N	7	2, 5	5		

• TYPES OF CHEMICAL BOND •

• There are two types –

• Ionic Bond • covalent bond.

• **Ionic Bond** – formed by transfer of electrons from one atom to another. also called as **electrovalent bond.**

• **Covalent Bond** – formed by the sharing of electrons between two atoms.

• Ionic Bond –

• The chemical bond formed by the transfer of electrons from one atom to another atom known as ionic bond.

• Transfer of electron takes place in such way that the ions formed have stable electron arrangement of inert gas.

• An ionic bond is formed when one of the atoms can donate electron to achieve inert gas configuration.

• When a metal reacts with non metal, transfer of electrons takes place from metal atoms to the non metal atoms and ionic bond is formed. Ionic compound made up of ions.

• The ions of metal and nonmetal attract each other, the strong force of attraction develop between the oppositely charged ions are known as Ionic bond.

• By losing electrons, the metal atoms forms a positively charged ion (cation). The non metal atoms forms or accepts electrons from a negatively charged ion (anion).

• Ions formed the stable chemical bond between oppositely charged ions.

• Ionic bonds are formed between metals and non metals.

- **Formation of Sodium chloride -**
- Sodium is a metal whereas chlorine is a non metal.
- Sodium metals reacts with chlorine to form ionic compound, sodium chloride.
- The atomic number of sodium is 11, so electronic configuration 2,8,1. sodium atom only has 1 electron in outermost shell.

- $Na^+ \longleftarrow Na\cdot - e^-$

Sodium ion — Sodium atom 2,8,1 — electron

- The atomic number of chlorine is 17, so its electronic configuration is 2,8,7. chlorine atoms has 7 electrons in outer most shell and needs 1 electron to get stable.

- $\cdot\ddot{Cl}: + e^- \longrightarrow [:\ddot{Cl}:]^-$

Chlorine 2,8,7 — electron — Chloride ion

- When sodium react with chlorine, it transfers its 1 outermost electron to the chlorine atom.
- By losing 1 electron, sodium atoms forms a sodium ion (Na^+) and by gaining 1 electron, the chlorine atoms forms a chloride ion. (Cl^-).

Sodium — Chlorine — Sodium chloride

- $Na\cdot + \cdot\ddot{Cl}: \longrightarrow Na^+ [:\ddot{Cl}:]^-$ or NaCl

Electron transfer

- Positive charge
- Negative charge

Na^+ ion + $[:\ddot{Cl}:]^-$ ion = Sodium chloride

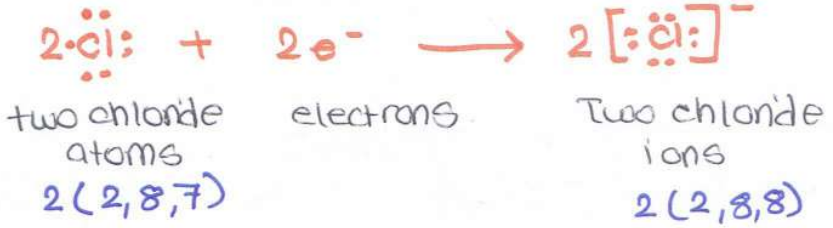
- They held together by electro static force of attraction.
- Sodium ion resemble to inert gas neon. and chloride ion resemble inert gas argon. Due to this sodium chloride is very stable.
- Note - In formation of ionic bonds, the reacting atoms achieve the inert gas electron configuration by transfer of electrons.
- Note - The charges are not written in the formula - Ex - NaCl is correct, not the Na^+Cl^-.
- **Formation of magnesium chloride -**
- Atomic no. 7 of magnesium is 12, so electronic configuration is 2,8,2, it has two valence electrons - Mg^{+2}

- $Mg - 2e^- \longrightarrow Mg^{+2}$

Magnesium — Electron — Magnesium oxide

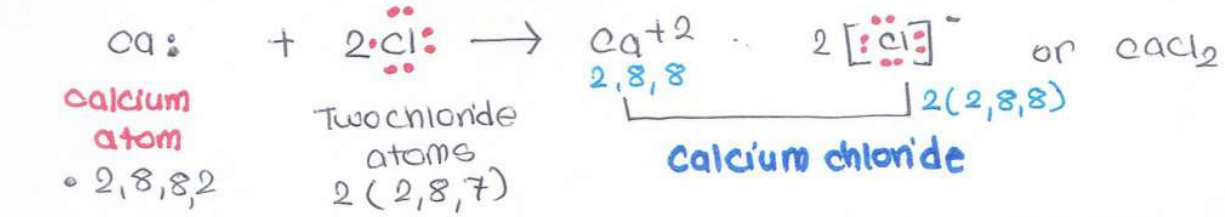
- 2,8,2

- The atomic no. of chlorine is 17, so electronic configuration is 2,8,7, it has 7 valence electrons. So it requires only 1 electron to complete its octet.

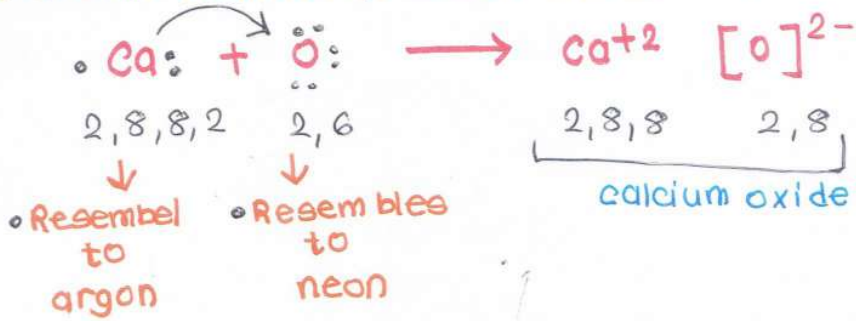


- since magnesium atoms has donates 2 electron, the two chlorine atoms take these two electrons and form two chloride ions.
- When magnesium reacts with chlorine, the magnesium atom transfers its two outermost electrons to two chlorine atoms-

- It is also held by electrostatic force of attraction to form compound. so formula becomes - $MgCl_2$. It contains ionic bonds. the force of attraction is very strong.
- The formation of calcium chloride ($CaCl_2$) is similar to formation of magnesium chloride.
- The atomic no. of calcium is 20 and chlorine is 17



- They held together by electrostatic force of attraction, $CaCl_2$ is an ionic bond.
- Calcium reacts with oxygen to form ionic compound calcium oxide. The atomic number is 20 of calcium and oxygen is 8.



Ionic compounds-

- The compounds containing ionic bonds are known as ionic compound.
- They are made up of positively charged ions (cation) and negatively ions (anion). Also known as electrovalent compound.
- The ionic compounds consist of ions and not molecule

Name	Formula	Ions present
• Sodium chloride	NaCl	$Na^+ - Cl^-$
• Potassium chloride	KCl	$K^+ - Cl^-$
• Calcium nitrate	$Ca(NO_3)_2$	$Ca^{+2} - NO_3^-$
• copper sulphate	$CuSO_4$	$Cu^{+2} - SO_4^{2-}$
• Ammonium chloride	NH_4Cl	$NH_4^+ - Cl^-$
• Sodium oxide	Na_2O	$Na^+ - O^{2-}$
• Calcium chloride	$CaCl_2$	$Ca^{+2} - Cl^-$
• Magnesium chloride	$MgCl_2$	$Mg^{+2} - Cl^-$
• Sodium hydroxide	NaOH	$Na^+ - OH^-$
• Aluminium oxide.	Al_2O_3	$Al^{+3} - O^{2-}$

- Only Ammonium chloride which is an ionic compound made up of only non metals. (Exception)

• Covalent Bond-

- The chemical Bond formed by sharing of electrons between two atoms.
- The sharing of electrons takes place in a such a way that each atom in resulting molecule gets stable.
- Atoms share only their outermost electrons in formation of covalent bonds. they get stable by mutual sharing.
- Whenever a non-metal combines with another non-metal, sharing of e⁻'s takes place between their atoms. (e⁻'s - electrons)
- A covalent bond can also be formed between two atoms of the same non metal. Ex- Between two chlorine atoms to form Cl_2.
- Covalent bonds are of three types-
- single covalent bond • Double covalent bond • Triple covalent bond

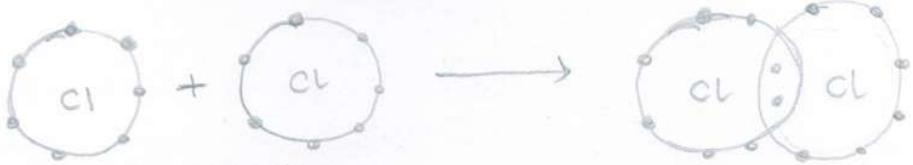
• SINGLE COVALENT BOND -

- A single bond is formed by sharing of one pair of electrons between two atoms.
- A shared electron pair is always drawn between the atoms.
- Ex- A hydrogen molecule H_2 contains a single covalent bond and written as- H:H. It is denoted by putting the a short line (—) between two atoms.
- H—H, it represented as. This short line represents as a single bond between the two hydrogen atoms.

- Formation of chlorine molecule -
- chlorine atom is very reactive and cannot exist free because it does not have the stable electron arrangement of an inert gas.
- Each molecule of chlorine contains two chlorine atoms joined by a single covalent bond.
- The atomic number of chlorine is 17, so electronic configuration is 2,8,7. It has 7 electrons in outermost shell. and needs 1 more electron to get stable.
- Two chlorine atoms share one electrons each form a chlorine molecule

chlorine atom 2,8,7 | chlorine atom 2,8,7 | chlorine molecule (2,8,8)

- Resemble to inert gas Argon
- Because the two chlorine atoms share electrons there is a strong force of attraction between them which hold them together.
- The two chlorine atoms share one pair of electrons, the bond between them is called single bond.



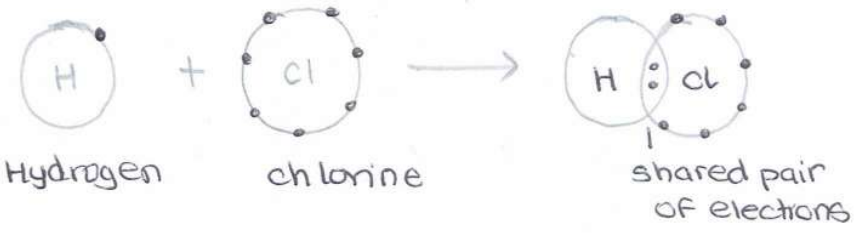
- Two chlorine atoms
- Cl_2 – shared electron

- Whenever two atoms of the same element combine to form a molecule a covalent bond in formed.

- **Formation of a Hydrogen chloride, HCl -**
- Hydrogen atom has 1 valence electron, so it needs 1 more electron to get 2 electron helium gas electrons and become stable.
- chlorine atom has 7 valence electrons, so it also needs 1 more electron to achieve the stable configuration.

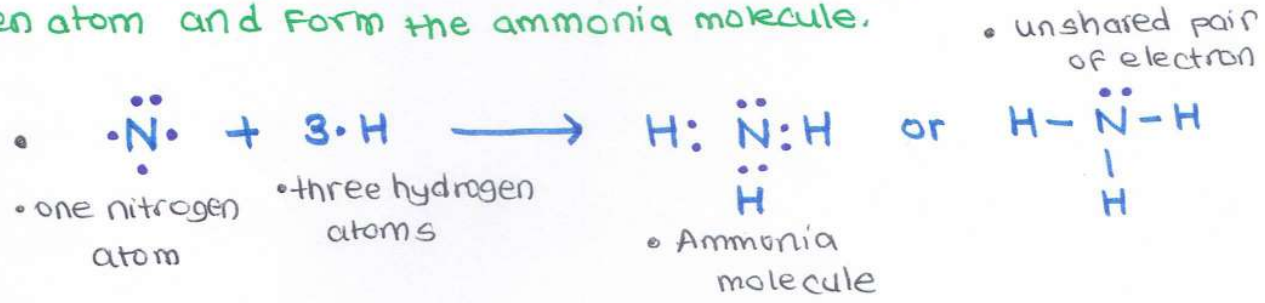
1 | 2,8,7 | Shared electron pair (2) | (2,8,8)

- Hydrogen gas resembels to Helium.
- chlorine gas resembels to Argon.
- Hydrogen chloride gas is a covalent compound containing covalent bond.



Hydrogen chlorine shared pair of electrons

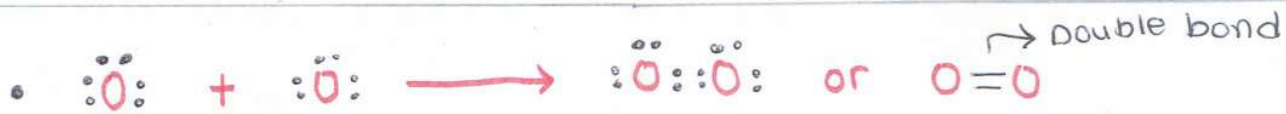
- Formation of Ammonia - NH_3
- Nitrogen combines with hydrogen to form covalent compound ammonia having covalent bond in it.
- Nitrogen has 5 valence electrons, so it needs 3 electrons to complete the octet and become stable.
- Hydrogen has 1 valence electrons, so it needs one more electrons to complele 2 electrons in its K shell and becomes stable.
- one atom of nitrogen shares its three valence electrons with three hydrogen atom and form the ammonia molecule.



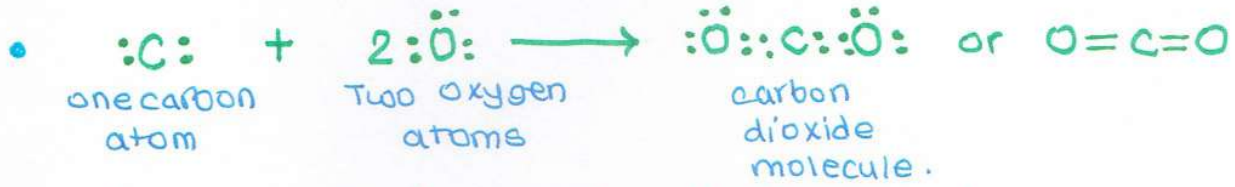
- one nitrogen atom
- three hydrogen atoms
- Ammonia molecule
- unshared pair of electron

- Hydrogen resemble to Helium (inert gas)
- Nitrogen resemble to neon (inert gas)
- By combination of nitrogen and hydrogen, the ammonia molecule is formed.

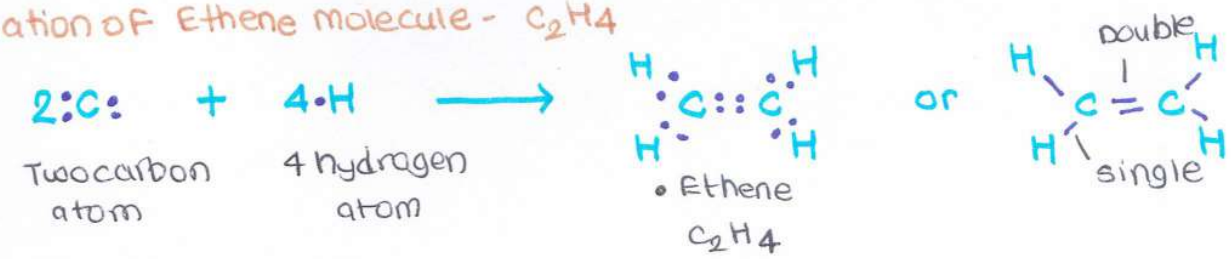
- **DOUBLE BOND -**
- A double covalent bond consist of two pairs of shared electrons.
- Bond is formed by sharing of two pairs of electrons between two atoms. It is represented by putting two short lines (=) between atom.
- A double covalent bond is formed by the sharing of four electrons between two atoms, each contributing two electrons for sharing.
- Ex- Oxygen molecule O_2 contains a double bond between two atoms and it is written as - O=O.
- Formation of oxygen molecule, O_2
- Oxygen atom is very reactive and cannot exist free bcoz it does not have the stable, inert gas electron arrangement in its valence shell.
- If does not consist of single bond, it is more stable O_2 molecule.
- Oxygen atomic number is 8, its electronic configurgation is 2,6. It has 6 valence electrons it reauire more 2 electrons to get fully stable.
- Two oxygen atoms share two electrons each and form a stable oxygen molecule.



- The oxygen molecule are held together by a double bond.
- Double bond is stronger than single bond
- oxygen atom resemble as - Neon inert gas configuration.
- **Formation of carbon dioxide - CO_2**
- Carbon dioxide is a covalent compound made up of carbon atom and oxygen element and it contains covalent bonds.
- Carbon has 4 valence electrons and oxygen atom has 6 valence electrons.
- Carbon needs more 4 electrons to get stable and oxygen need more 2 electrons to get stable inert gas configuration.
- One carbon atom shares its four electrons with two oxygen atoms and forms a carbon dioxide molecule.



- There are two double bonds in carbon dioxide molecule.
- Both atoms resembles to - Neon inert gas configuration.
- Formation of Ethene molecule - C_2H_4



- Ethene molecule have one carbon-carbon double bond and four carbon hydrogen single bond.

TRIPLET BOND

- A triple covalent bond consist of three pairs of shared electrons. A triple bond formed by sharing of three pairs of electrons between two atoms.
- Three pairs of electrons equal to six electrons. Triple atom is represent by putting (≡) between two atoms.
- Nitrogen - $N \equiv N$ $H-C \equiv C-H$

- **Formation of nitrogen molecule-**
- A nitrogen atom is very reactive and cannot exist free because it does not have the stable electron arrangement of inert gas.
- It does not consist single atoms, it consist of more stable N_2 molecule.
- Atomic number of nitrogen is 7 so electronic configuration 2,5.
- It has 5 valence electrons in its outermost shell, it needs 3 more electrons to achieve the stable configuration.
- Two nitrogen atoms combine together by sharing 3 electrons each to form a molecule of nitrogen gas.

$$:\!\dot{N}\!: + :\!\dot{N}\!: \longrightarrow :\!N\!::\!N\!: \text{ or } :N \equiv N:$$

Two nitrogen atoms — Nitrogen molecule — Triple bond.

- In the nitrogen gas molecule, the two nitrogen held by triple bond.
- It resembles to inert gas molecule- Neon.
- **Covalent compounds-**
- The compounds containing covalent bonds are called covalent compounds. covalent compound are formed by the sharing of electrons between them.
- The covalent compounds are made up of molecules, so they are also known as Molecular compounds.

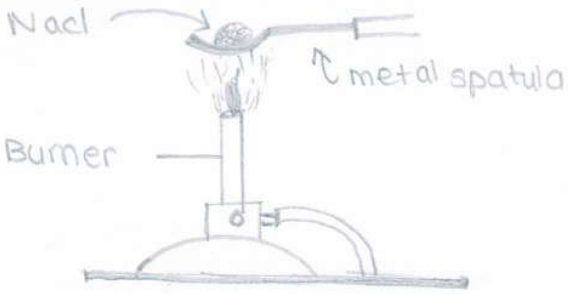
Name	Formula	Elements present
Methane	CH_4	C and H
Ethane	C_2H_5	C and H
Ethene	C_2 4	C and H
Ethyne	C_2H_2	C and H
water	H_2O	H and O
Ammonia	NH_3	N and H
Alcohol	C_2H_5OH	C, H, and O
HCl gas	HCl	H and Cl
H_2S gas	H_2S	H and S
Urea	$CO(NH_2)_2$	C, O, N and H
Glucose	$C_6H_{12}O_6$	C, H and O

- Note- Whenever we see a compound made up of two nonmetals we can say it is covalent compound.

PROPERTIES OF IONIC COMPOUNDS-

- Ionic compounds are usually crystalline solid because their oppositely charged ions attract one another strongly and form a regular crystal structure. The crystals of ionic compound are hard and brittle.
- Ionic compounds are having high melting and high boiling point- For example- Nacl has high melting point of 800°C and high boiling point 1413°C.
- There is a strong force of attraction between the oppositely charged ions, so a lot of heat energy is required to break the force of attraction between the bonds and melt and boil the ionic compound.
- **Ionic compound are usually soluble in water but insoluble in organic solvent.** Nacl is soluble in water but insoluble in solvent like ether, benzene or Kerosene. bcoz water has high dielectric constant due to which it weakens the attraction between the ions.
- **Ionic compound conduct electricity when dissolved in water or when melted.** bcoz they contain charged particles called ions. It is conduct electricity because there are plently of free ions in the solution which are able to conduct electric current.

- **EXPERIMENTS-**
- The property of ionic compounds that they have high melting point-
- Take a small amount of Nacl on a metal spatula. Heat it directly over the burner.
- we can see that Nacl does not melt easily. It melts on strong heating. This shows that sodium chloride has high melting point.

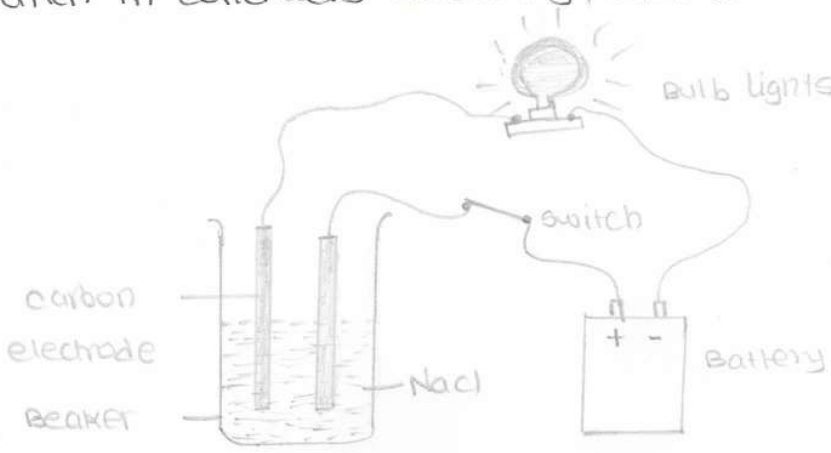


- The property of ionic compounds that they are soluble in water, but insoluble in other solvents-
- Take a water in a test tube and add pinch of Nacl to it. we can see Nacl get dissolve in water, but when a pinch of Nacl is put in other test tube with ether, cannot dissolve in ether. It remains same at bottom of test tube. Nacl is insoluble in organic solvent ether.

• The property of ionic compounds that they con ducts electricity when dissolved in water -

• Fill a beaker with Nacl and dissolve some Nacl with water. Two carbon rods are placed in Nacl solution. A electric circuit is set then on pressing the switch, the bulb light up at once.

• Nacl taken in conducts electricity, when dissolve in water.



• PROPERTIES OF COVALENT COMPOUNDS -

• covalent compounds are usually liquids or gases. only some of them are solids.

• solids - Glucose, cane sugar, urea

• liquids - Alcohol, ether, benzene, carbon tetrachloride

• Gases - Methane, chlorine, ethane.

• covalent compounds are usually has low melting and boiling point -

• Naphthalene has low melting point of 80°C. and ccl4 has low boiling point of 77°C.

• Covalent compounds are made up of electrically neutral molecules. The force of attraction between the molecules are very weak.

• covalent compounds are usually insoluble in water but they are soluble in solvent.

• Naphthalene insoluble in water but dissolves inorganic solvent like ether. some compounds like glucose, sugar, and urea. soluble in water.

• covalent compounds do not conduct electricity this means that they are non electrolytes. covalent compounds do not conduct electricity because they do not contain ions.

• sugar, urea, alcohol and ccl4 - do not conduct electricity.

• Hcl chemically reacts with water to form Hcl containing ions.

Ionic compound	Covalent compound
• Ionic compound are usually crystalline solids.	• covalent compound are usually gas or liquids, only some are solids.
• They are usually soluble in water	• They are usually insoluble in water
• They conduct electricity when dissolve in water.	• They do not conduct electricity in water.
• They are insoluble in organic solvent	• They are soluble in organic solvent.
• They have high melting point and boiling point.	• They have low melting point and boiling point.

Occurrence of metals-

- Earth is main and major source of minerals, ores and some metals.
- Most of the metals are quite reactive and hence they do not occurs as free element in nature.
- only few less reactive metals are found in 'free state' as metal bcoz of their low chemical reactivity.
- When a metal is found in free state or element, it is said to occure in 'native state'. EX- copper, silver, gold and platinium metals.
- copper and silver metals occurs in free state as well as in the combined state (in form of compounds).
- The metals which are high up in reactivity, so reactive that they never found in nature as free elements. (Potassium, sodium) they find in combined state.
- All the metals which are placed above copper in the reactivity series are found in nature only in the form of their compounds.
- Minerals - The natural materials in which the metals or compounds are found in earth are called minerals.
- Ores - Minerals from which the metals can be extracted profitably are called ores. An ore contains a good percentage of metal

"All the ores are minerals, but all the minerals are not ores"

Metal	Ore name	Name of compo. in ore	formula of ore
• sodium	• Rock salt	• sodium chloride	NaCl
• Aluminium	• Bauxite	• Aluminium oxide	$Al_2O_3 . 2H_2O$
• Maganese	• Pyrolusite	• Manganese dioxide	MnO_2
• zinc	1. calamine	• zinc carbonate	$ZnCO_3$
	2. zinc blende	• zinc sulphide	ZnS
• Mercury	• cinnabar	• Mercury(II) sulphide	HgS

• Iron	• Haematite	• Iron (III) oxide	Fe_2O_3
• copper	1. cuprite	• copper(I) oxide	Cu_2O
	2. copper glance	• copper (I) sulphide	Cu_2S

• Extraction of metals-

• After the mining of the ore from the ground, it converted into pure metal. To obtain a metal from its ore is called the extraction of metal.

• Metallurgy - The various processes involved in the extraction of metal from their ores, and refining is called metallurgy.

• Note- No single process can be used for the extraction of all metal.

• Three major steps-

• concentration of ore.

• conversion of concentrated ore into metal.

• Refining of impure metal.

• concentration of ore -

• ore is an impure compound of metal containing a large amount of sand and rocky material.

• The unwanted impurities like sand, rocky material, limestone, mica present in an ore are called Gangue.

• Before extraction, the metal from an ore, it is necessary to remove these impurities. The methods used for removing gangue from an ore depend on some difference in physical properties of the ore and gangue.

• concentration of concentrated ore into metal-

• All the metals into following three categories-

• Metals of high reactivity or highly reactive metals.

• Metals of medium reactivity or moderately reactive metals.

• Metals of low reactivity or less reactive metals.

• Manganese metal (mn) lies just above zinc (zn) in reactivity series

• Manganese is obtained by the reduction of its oxide with aluminium powder and not carbon. this is because carbon is less reactive than manganese

• carbon can reduce the oxides of zinc and all other metals below zinc to form metals.

• Copper can extracted from by the reduction of oxide with carbon as well as by heating its sulphide ores.

• Carbon which is an non metal, is more reactive than zinc (Zn). so it placed just above Zn in reactivity series
• Tin metal (Sn) is more reactive than lead (Pb) so it is placed above the Pb in reactivity series.
• The extraction of a metal from its concentrated ore is essentially a process of reduction of metal compound present in the ore.

• Extraction of Highly reactive metals-

• Metals	• Method of extraction
K Na Ca Mg Al	Electrolysis of molten chloride or oxide
Zn Fe Pb Cu	Reduction of oxide with carbon
Cu Hg	Heating sulphide in air
Ag Au Pt	Found in native state

• The highly reactive metal such as potassium, sodium, chloride, mangnesium and Aluminium placed in high up in the reactivity series in upper part.
• Highly reactive metals cannot be extracted by reducing their oxides with carbon.
• The highly reactive metals are extracted by the electrolytic reduction of their molten chloride or oxides. Electrolytic reduction is brought about by passing electric current through the molten state.
• It is takes place by extraction of oxide and chlorides by the electrolysis.
• During electrolysis, the negatively charged electrode (cathode) acts as a powerful reducing agent by supplying electrons to reduce the metal ions into metal.
• During the electrolysis of molten salts, the metals are always produced at the cathode (negative electrodes) fact that the metal ion are always (+vely charged and get attracted to the (-)vely charged electrode (cathode) when electricity is passed through it

- The metals extracted by electrolysis methods are very pure. They do not contain any impurities.
- When a molten metal chloride is electrolysed by passing electric current, then pure metal is produced at the cathode (negative electrode) and chlorine gas is formed at anode (+ve electrode).
- When a molten metal oxide is electrolysed by passing electric current, then pure metal is produce at cathode (negative electrode) and oxygen gas is formed at anode (+ve electrode).

- **Extraction of Sodium metal-**
- Sodium metal is extracted by the electrolytic reduction of molten sodium chloride. When electric current is passed through molten sodium chloride, it decomposes to form sodium metal and chlorine gas.

$$2NaCl(l) \xrightarrow{\text{Electrolysis}} 2Na(s) + Cl_2(g)$$

Sodium chloride — Sodium — Chlorine gas

- **Explanation-** During the electrolysis of molten sodium chloride the rx takes place at the two electrodes.
- **At cathode-** Positive sodium ion (Na^+) are attracted to the cathode. The sodium take electrons from the cathode and get reduced to form sodium atoms.

$$2Na^+ + 2e^- \longrightarrow 2Na$$

Sodium ion — electrons — Sodium atoms

- Produce at cathode (negative electrode).
- **At anode-** Negative chloride ion (Cl^-) are attracted to anode. The chloride ion give electron to the anode and get oxidised to form chlorine gas.

$$2Cl^- - 2e^- \longrightarrow Cl_2$$

Chloride ion — electrons — Chlorine gas.

- The, chlorine gas is formed at anode (positive electrode).
- Potassium metal- is produced of molten potassium chloride
- calcium metal- is obtained by electrolysis of molten calcium chloride
- magnesium metal- is obtained by electrolysis of molten chloride

- **Extraction of Aluminium metal-**
- Aluminium metal is extracted by the electrolytic reduction of molten aluminium oxide. When electric current pass through molten aluminium oxide it decomposes to form aluminium metal and oxygen gas.

$$2Al_2O_3(l) \xrightarrow{\text{Electrolysis}} 4Al(l) + 3O_2(g)$$

Aluminium oxide — Aluminium metal — oxygen

Explanation- Molten Aluminium oxide (Al_2O_3) contains free aluminium ions (Al^{+3}) and free oxide ions (O^{2-}).

- **At cathode-** The positively charged aluminium ions are attracted to the cathode (negative electrode).

$$Al^{3+} + 3e^- \longrightarrow Al$$

Aluminium ion — Electrons (From cathode) — Aluminium atom

- Aluminium atom form at cathode.
- **At anode-** The negatively charged oxide ions are attracted to the anode (positive electrode).

$$2O^{2-} - 4e^- \longrightarrow O_2$$

Oxide ions — Electrons (at anode) — oxygen gas

- The oxygen gas form at anode.
- **Extraction of moderately reactive metals-** Placed in middle of the reactivity series.
- The reduction of their oxides with carbon, aluminium, sodium or calcium.
- It is easier to obtain metal from their oxides than from carbonate and sulphide.
- The concentrated ores can be converted into metal oxide by the process of calcination or roasting.
- A carbonate ore is converted into oxide by calcination.
- A sulphide ore is converted into oxide by roasting
- **Calcination -** is the process in which a carbonate ore is heated strongly in the absence of air to convert it into metal oxide.
- For example- Zinc occurs as zinc carbonate in calamine ore $ZnCO_3$.
- Zinc carbonate should be first converted into zinc oxide. This is done by calcination.
- When calamine ore is heated strongly in the absence of air, when calamine is calcined, it decomposes to form zinc oxide and carbon dioxide-

$$ZnCO_3(s) \xrightarrow{\text{calcination}} ZnO(s) + CO_2(g)$$

Zinc carbonate — calcination — zinc oxide — carbon dioxide

Thus, calcination converts zinc carbonate into zinc oxide.

- **Roasting -** It is the process in which a sulphide ore is strongly heated in presence of air to convert it into metal oxide.

- Zinc occurs as sulphides ore in zinc blend ore ZnS.
- When zinc blend ore is strongly heated in air, it forms zinc oxide and sulphur dioxide-
  - $2ZnS(s) + 3O_2(g) \xrightarrow{\text{Roasting}} 2ZnO(s) + 2SO_2(g)$
  - zinc sulphide (zinc blend) + oxygen → zinc oxide + sulphur dioxide
- Roasting converts zinc sulphide into zinc oxide
- **Reduction of metal oxide with carbon-**
- The oxide of less reactive metals like iron, nickel, tin and lead are usually reduced by using carbon as the reducing agent.
- In the reduction by carbon, the metal oxide is mixed with carbon and heated in a furnace.
- carbon reduces the metal oxide to free metal.
- Zinc metal is extracted by the reduction of its oxide with carbon (or coke). When zinc oxide heated with carbon, zinc metal is produce-
  - $ZnO(s) + C(s) \longrightarrow Zn(s) + CO(g)$
  - zinc oxide + carbon (Reducing agent) → zinc metal + carbon monoxide
- Carbon is a cheap reducing agent, but it contaminates the metal.
- **Iron metal is extracted from its oxide ore 'Haematite'** (Fe_2O_3) by reduction with carbon
- **Tin and lead metal are also extracted by the reduction of their oxides with carbon.**
- Even the less reactive metal copper is extracted by the reduction of its oxide with carbon.
- **Thermite reaction-** The reduction of metal oxide to form metal by using aluminium powder as a reducing agent is called a thermite reaction. It is a exothermic reaction in which a large amount of heat is evolved.
- The amount of heat evolved is so large that metals are produce in molten state.
- The property of the reduction by aluminium is made use of in thermite welding for joining the broken pieces of heavy iron object like railway track or machine parts.
- Aluminium reduces iron oxide to produce iron metal with the evolution of lot of heat. Due to this heated, iron metal is produce in molten state. used for welding process is called thermite welding
  - $Fe_2O_3(s) + 2Al(s) \longrightarrow 2Fe(l) + Al_2O_3(s) + Heat$
  - Iron (III) oxide + Aluminium powder → Iron metal + Aluminium oxide

- Extraction of less reactive metals-
- The less reactive metals such as mercury and copper are quite low in the reactivity series.
- The less reactive metals which are quite low in the activity series are extracted by the reduction of their oxides by heat alone.
- Mercury and copper metals are extracted by the reduction of their oxides by heat alone.

- Extraction of Mercury-
- Mercury is a less reactive metal which is quite low in the activity series. Mercury metal can be extracted by heating its sulphide ore in air.
- Mercury metal is produced from the sulphide ore called cinnabar. HgS which is actually mercury (II) sulphide.
- Extraction of mercury from cinnabar ore involves two steps

a. The concentrated mercury (II) sulphide ore is roasted in air when mercury (II) oxide is formed.

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

Mercury (II) sulphide (cinnabar ore) | oxygen | Mercury (II) oxide | sulphur dioxide

b. when this mercury (II) oxide is heated over about 300°C, it decomposes to form mercury metal.

- $2HgO(s) \xrightarrow{\Delta} 2Hg(l) + O_2(g)$

Mercury (II) oxide | Mercury metal | oxygen

- Mercury (II) sulphide is also called mercuric sulphide and mercury (II) oxide is also known as mercuric oxide.

- Extraction of copper-
- copper is a less reactive metal which is quite low in the reactivity series. copper metal can be extracted just by heating its sulphide ore in air.

copper metal - copper glance - Cu_2S which is copper (I) sulphide.

- Extraction of copper from copper glance involve two steps-

a. The concentrated copper (I) sulphide ore (copper glance) is roasted in air when a part of copper (I) sulphide is oxidised to copper (I) oxide.

- $2Cu_2S + 3O_2 \xrightarrow{\text{Roasting}} 2Cu_2O + 2SO_2$

- Copper glance ore | copper (I) sulphide | oxygen | copper (I) oxide | sulphur dioxide

b. When a good amount of copper (I) sulphide has been converted into copper (I) oxide, then the supply of air for roasting is stopped. In the absence of air, copper (I) oxide formed above reacts with copper (I) sulphide to form copper metal and sulphur dioxide -

- $2Cu_2O(s) + Cu_2S(s) \longrightarrow 6Cu(s) + SO_2(g)$

copper (I) oxide — copper (I) sulphide (unoxidised) ore. — copper metal — sulphur dioxide

- They need a reducing agent for their reduction to the metals.
- **Refining of metals-**
- The process of purifying impure metals is called refining of metals. Thus refining of metals means purification of metals.
- The most important and most widely used method for refining impure metals is **Electrolytic Refining.**
- It means refining by electrolysis, many metals like tin, lead and chromium are refined electrolytically.
- **The refining of an impure metal by electrolysis -**

a. The thick block of the impure metal is made anode. connected to the positive terminal of battery.

b. The thin strip of the pure metal is made **cathode**. connected to the negative terminal of battery.

c. A water soluble salt is taken as electrolytes.

- **On passing electric current, impure metal dissolve from the anode and goes on the electrolytic solution.** and pure metal from the electrolyte deposit at the cathode.
- The soluble impurities present in the impure metal go into the solution whereas insoluble impurities settle at the bottom of the anode as "anode mud."

- **Electrolytic Refining of copper -**
- The apparatus consist of an electrolytic tank containing acidified copper sulphate solution as an electrolyte. ($CuSO_4$ with dil H_2SO_4)
- A thick block of impure copper metal is made anode (connect with the +ve terminal of battery).
- A thin strip of pure copper metal is made cathode (connected to the −ve terminal of battery).
- On passing electric current impure copper from the anode dissolves and goes into copper sulphate solution. and pure copper from copper sulphate solution deposit on cathode.
- The soluble impurities goes into solution, Insoluble collect below the anode as anode mud.

Explanation - $CuSO_4$ solution contains copper ions, Cu^{+2} and sulphate ions SO_4^{2-}. on passing the electric current through copper sulphate.

- The positively charged copper ions, Cu^{+2} from the copper sulphate solution go to the negative electrode cathode and by taking electrons from the cathode, get reduced to copper atoms.

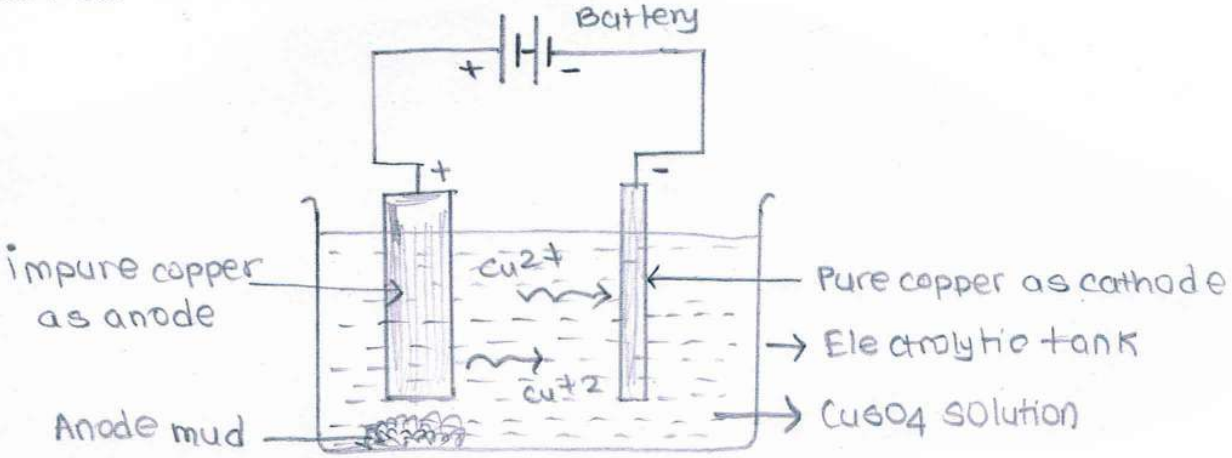
- At cathode- $Cu^{+2} + 2e^- \longrightarrow Cu$

  Copper atom (from electrolyte) | Electrons (from cathode) | Copper atom (Deposit on cathode)

- These copper atoms get deposited on the cathode giving pure copper metal.
- Copper atoms of the impure anode lose two electrons each to anode and form copper ions Cu^{+2}, which go into the electrolyte solution.
- At anode- $Cu - 2e^- \longrightarrow Cu^{+2}$

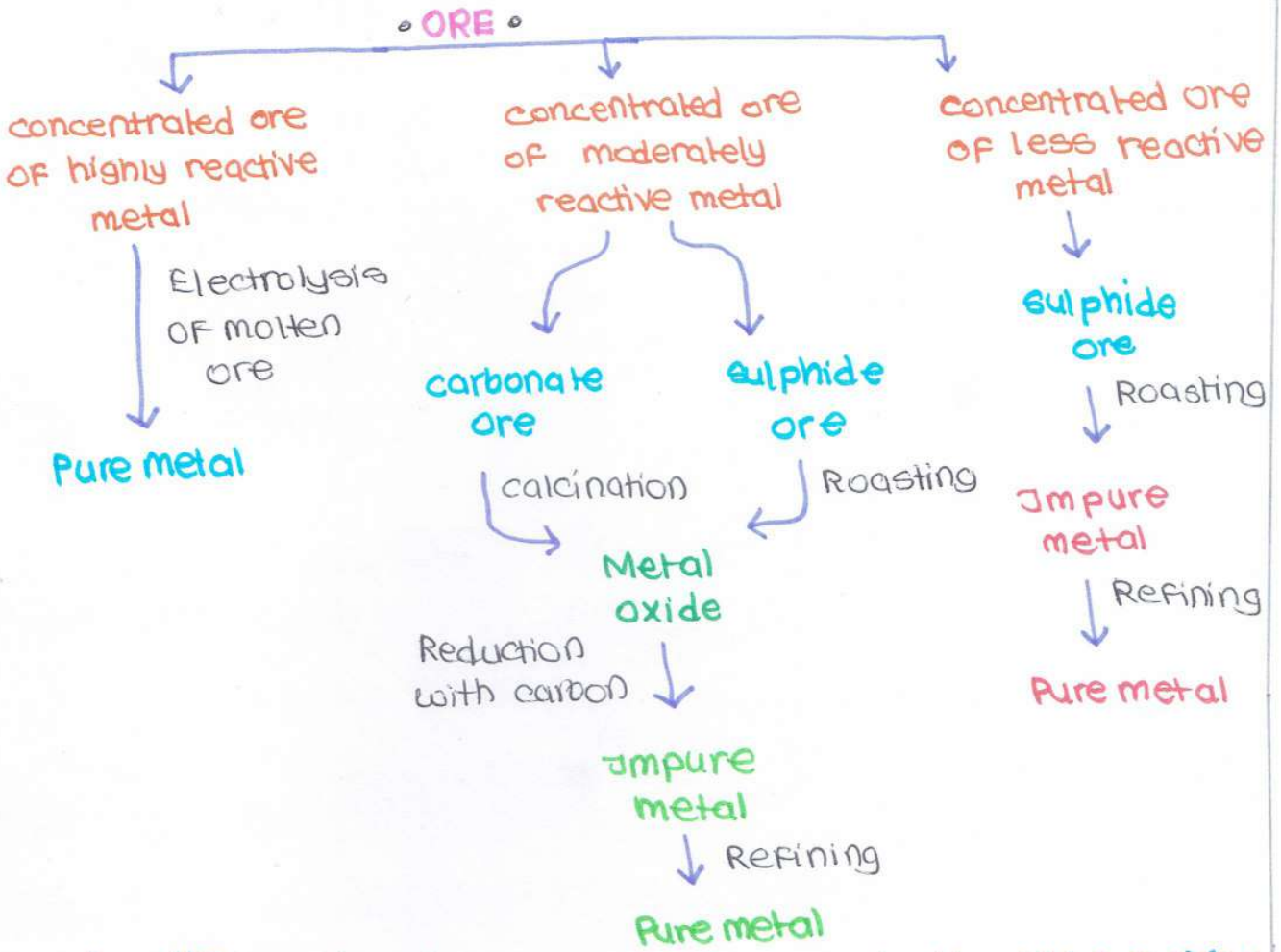
  Copper atom (from impure anode) | Electrons (Give to) anode | copper ion (goes into) electrolyte

- copper sulphate solution at the cathode and put into the solution at the anode.
- As process goes on, impure anode becomes thinner and thinner whereas pure cathode becomes thicker and thicker. Thus copper is obtained at the cathode.



- The less reactive metals like gold and silver present in the impure copper collect at the bottom of the electrolytic cells below the anode in the form of anode mud.
- Gold and silver metals can be recovered from the anode mud.
- The electrolytic refining of metal serve two purpose -
- It refines the metal concerned.

• It enable to recover the other valuable metals present as impurities in the metal being refined.



• CO_2 on heating on treatment of a dilute acid, it will be a carbonate ore.
• SO_2 on heating in air, then it will be sulphide ore.

• **CORROSION-**
• The eating up of metals by the action of air, moisture or a chemical on their surface is called corrosion.
• Most of the metals corrode when they are exposed to damp air.
• Rust - It is soft and porous which gradually falls off from the surface of iron object and then the iron below starts corroding.
• Corrosion of iron is continuous process which ultimately eats up the whole iron object.
• The corrosion of iron is called rusting while other metals are said to 'rust'.
• Rusting of iron-
• When an iron object is left in damp air for a considerable time it gets covered with a red-brown flaky substance called rust. this is called rusting of iron.

• During the rusting of iron, iron metal combines with the oxygen of air in the presence of water to form hydrated iron(III) oxide. $Fe_2O_3 \cdot xH_2O$. this hydrated iron(III) oxide is called rust.

• Rust is mainly hydrated iron(III) oxide $Fe_2O_3 \cdot xH_2O$

• Rust is brown-red in colour.

• Steel rusts less than iron.

• condition necessary for rusting of iron-

• Rusting of iron, needs both air and water, two conditions are necessary for rusting of iron.-

1. Presence of air 2. Presence of water

• The damp air alone supplies both the things air and water required for the rusting of iron.

• The ordinary water alone also supplies both the things, air and water, needed for rusting.

• Experiment to show that rusting of iron requires both air and water-

• In the first test tube containing iron nail, we put anhydrous $CaCl_2$ and close its mouth with cork. The anhydrous $CaCl_2$ is added to absorb water from the damp air present in the test tube and make it dry.

• The iron nail in the first test tube is kept in dry air (having no water vapour in it. The test tube kept aside for one week.

• In the second test tube containing iron nail, we put boiled distilled water. Boiled water does not contain any dissolved air in it. A layer of oil is put over boiled water. In this way, the iron nail in the second test tube is kept in air-free boiled water. The mouth of this test tube is closed with a cork and kept aside for one week.

• In the third test tube containing iron nail, we put unboiled water so that about two-third of nail is immersed in water and the rest is above the water, exposed to damp air. in this way, the iron nail in the third test tube has been placed in air and water together. the mouth of cork is closed and kept aside for about one week.

conclusions-

• Test tube 1- No rust is seen on the surface of iron kept in dry air. This tells us rusting of iron does not takes place alone in air.

• Test tube 2- No rust is seen on the surface of nail kept in air free, boiled water in the second test tube. this tell us that rusting of iron does not take place in water alone.

Test tube 3- Red-brown rust is seen on the surface of iron nail kept in presence of both air and water together in third test tube . this tell us that rusting of iron takes place in presence of both air and water together.

Hence, in above experiment shows that for the rusting of iron to take place, both air and water is essential. this means that the rusting of iron objects can be prevented if damp air is not allowed to come in contact with iron objects.

- **Corrosion of Aluminium-**
- when a shining aluminium vessels is exposed to moist air, the oxygen of air reacts with aluminium to form a thin dull layer of aluminium oxide. all over the vessels.
- **Aluminium is more reactive than iron** so fresh aluminium metal begins to corrode quickly when it comes in contact with moist air.
- The action of moist air on aluminium metals forms a thin layer of aluminium oxide all over the aluminium metals
- The aluminium oxide layer is very tough and prevents the metal underneath from corrosion which will be further can be done.
- This means that the sometimes corrosion is useful.
- **The common metal which is highly resistant to corrosion is Aluminium.**
- The tough layer of aluminium oxide over aluminium metal, this can done by a process called 'anodising' it can be done thicker by electrolysis process.
- In this Electrolysis process-
- **Aluminium object made an- anode (+ve electrode) in an electrolytic**
- **tank in which dil H_2SO_4 is electrolysed.**
- During electrolysis of dil. H_2SO_4 acid, oxygen gas is liberated at anode and react with aluminium object and form a thickcer layer of Aluminium oxide on the surface.
- **Anodising-** It is the process of forming thick layer of Aluminium oxide on aluminium object by making it anode during electrolysis of dil H_2SO_4 acid.

- **Corrosion of copper-**
- copper oxide layer formation takes place on copper objects by which they lose there shine. after some day.
- copper reacts with air in considerable manner, then copper reacts with carbondioxide and water slowly to form a "**green coating of basic $CuCO_3$ on surface**"

- The green coating of basic copper carbonate is a mixture of of copper carbonate and copper hydroxide. $CuCO_3 \cdot Cu(OH)_2$
- Since copper metal is low in the reactivity series, therefore corrosion of copper is very slow.
- The corroded copper vessels can cleaned with dilute acid solution.
- The acid solution dissolves green coloured basic copper carbonate present on the corroded copper vessels and looks them shinky. Red brown again.

- **Corrosion of Silver-**
- When a shining metal objects loses its shine they become dull we can say that it has been tarnished.
- **When silver objects kept in air they graudally be come black colour.**
- Explanation- silver is very highly unreactive metal it does not react with oxygen of air easily. But air contains a little of sulphur compounds such as hydrogen sulphide gas (H_2S)
- The silver objects combine slowly with H_2S to form a black coating of silver sulphide. (Ag_2S)
- It graudally turns black due to formation of thin silver sulphide layer on their surface by the action of H_2S gas present in air.
- Silver metal used for making silver coins, jewellery and silver-ware because its bright shiny surface and resistance to corrosion.

- **Corrosion of Gold and Platinum-**
- Gold metal does not corrode when exposed in air. because it is very unreactive metal which remains unaffected by air water and other gases.
- Gold is used for making jewelery because of its bright shiny surface and high resistance to corrosion.
- Note- Gold dissolves only in aqua-regia solution.
- Platinum is highly resistant to corrosion
- Platinum also dissolves in aqua regia solution. it is a white metal with a silvery shine.
- Platinum used to make jewellery bcoz of its bright shiny surface and high resistance to corrosion.
- Titanium is also very resistant to corrosion. (Note)

•ALLOYS-

• An alloy is a homogenous mixture of two or more metals.

◦ For example- • Brass is an alloy of two metals- copper and zinc.

• Steel is an alloy of a metal and non metals - iron and carbon.

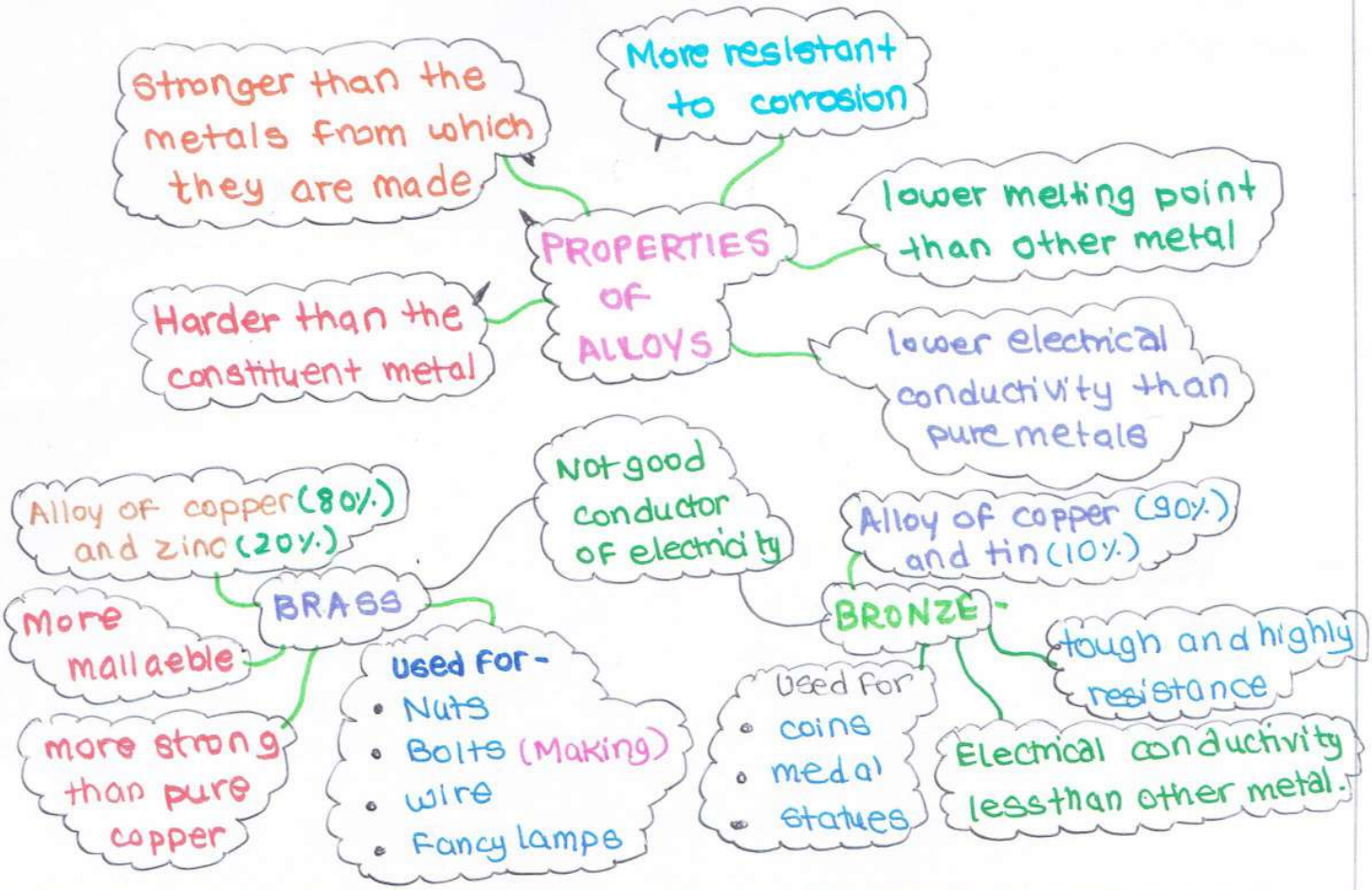
• An alloy is prepared by mixing the various metals in molten state in required proportions and then cooling their mixture to the room temperature.

• **Duralumin** - An alloy of aluminium with copper, mangesium and manganese is light as well as strong.

◦ it is used in making aircraft bodies, space satellite and pressure cookers. (**Aluminium metal is light but not strong**).

• **Magnalium** - An alloy of aluminium with magnesium, is light as well as hard. (**Aluminium metal is light but not hard**)
It is used to make balance beams and light instruement.

• **stainless steel** - Alloy of chromium and nickel with iron we get an alloy stainless steel. (which is strong, tough and does not rust at all.
EX- cooking utensils, knives, scissors, tools and ornamental pieces. used for making surgical instruement.



- Alloy of gold - The purity of gold in terms of carats.
- Pure gold is said to be of 24 carats
- Pure gold is very soft due to which it is not suitable for making jewellery
- Gold alloyed with a small amount of silver or copper to make it hard. The harder alloy of gold make it suitable for making ornaments.
- In india, gold ornaments usually made up of 22 carat gold. It means that 22 parts of pure gold is alloyed with 2 parts of either silver or copper.

Q. How rusting of Qutab minar in Delhi has been prevented-

- The rusting of iron has been prevented because of formation of thin film of magnetic oxide to form on its surface. (Fe_3O_4) as the result of treatment given to the pillar, painting with the mixture of many salts, then heating and quenching. bcoz of the good treatment it is the wonder of ancient indian metallurgy. bcoz of good knowledge of alloys.