CHAPTER-3

METALS AND NONMETALS

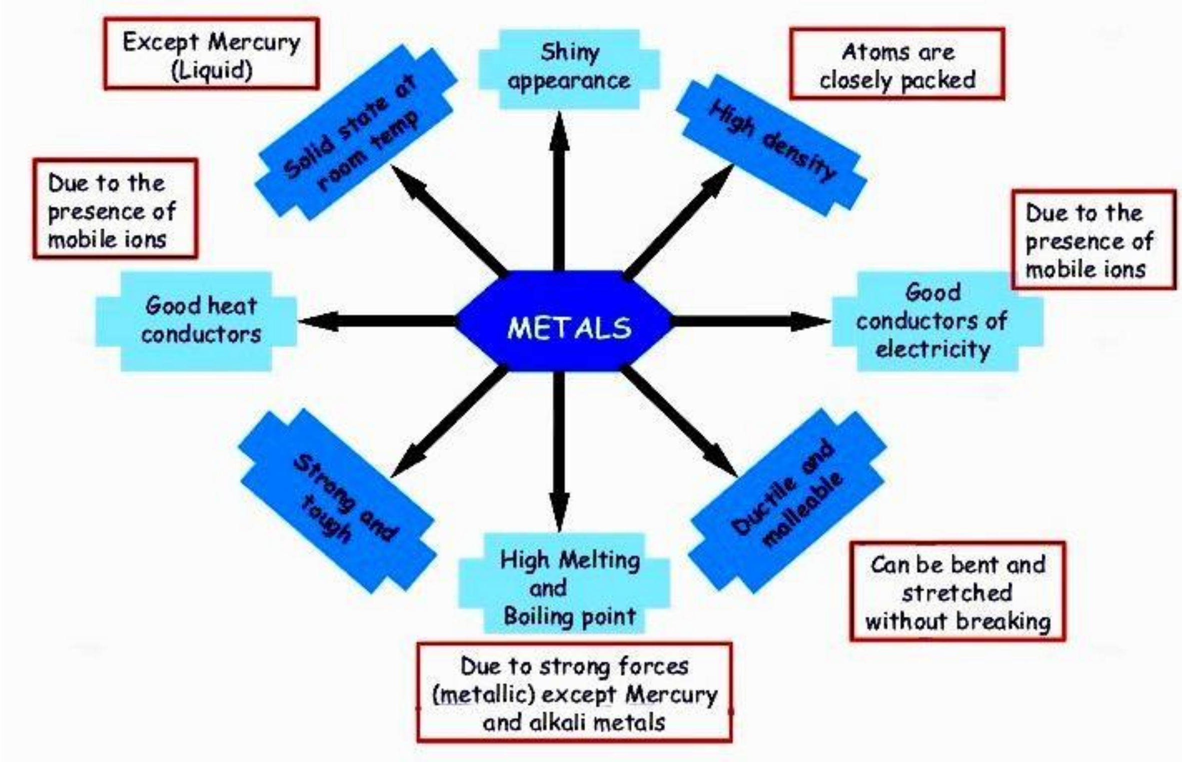
STUDY NOTES :

Metals occupy the bulk of the periodic table, while non-metallic elements can only be found on the right-hand-side of the Periodic Table . A diagonal line, drawn from boron (B) to polonium (Po), separates the metals from the nonmetals. Most elements on this line are metalloids, sometimes called semiconductors. This is because these elements exhibit electrical properties intermediate to both, conductors and insulators. Elements to the lower left of this division - line are called metals, while elements to the upper right of the division - line are called non-metals.

On the basis of their general physical and chemical properties, every element in the periodic table can be termed either a metal or a nonmetal.

PHYSICAL PROPERTIES OF METALS:

- **Physical state** - Metals are solids at room temperature e.g. sodium, aluminium, potassium, magnesium. There are exceptions to this. Mercury and gallium are metals but they are in liquid state at room temperature.
- **Luster** – Metals have a shining surface called luster when freshly prepared. They have a quality of reflecting light from their surface and they can be polished e.g. metals like gold, silver, copper show this property.
- **Malleability** - Metals can be beaten into thin sheets. This property is called malleability. Due to this property, metals can be rolled into sheets e.g. aluminium, copper, zinc can be beaten into sheets.
- **Ductility** - Metals can be drawn into thin wires. This property is called ductility. For example, 100 grams of silver can be drawn into a thin wire about 200 meters long.
- **Hardness** – Metals are generally hard e.g. iron, cobalt, nickel. There are few exceptions to this. Sodium and potassium are soft and they can be cut with a knife.
- **Sound:** Metals produce ringing sound, so, metals are called sonorous. Sound of metals is also known as metallic sound. This is the cause that metal wires are used in making musical instruments.
- **Conduction** – Generally, metals are good conductors of heat and electricity because they have free electrons. Silver and copper are the two best conductors . Relatively, lead and bismuth are poor conductors of heat and electricity.
- **Density** - Metals generally have high density and they are heavy. Iridium and osmium have the highest densities while lithium has the lowest density.
- **Melting and boiling point** – Metals usually have high melting point and boiling point. For example, iron, cobalt and nickel have high melting and boiling point. Tungsten has the highest melting point. There are some exceptions to this. For example , most of the alkali metals have low melting and boiling point.
- **Strength:** Most of the metals are strong and have high tensile strength. Because of this big structures are made using metals, such as copper and iron.
- **Color:** Most of the metals are grey in color. But gold and copper are exceptions.



CHEMICAL PROPERTIES OF METALS

REACTION WITH OXYGEN:

Most of the metals form respective metal oxides when react with oxygen.

Metal + Oxygen → Metal oxide

Examples:

- Reaction of magnesium metal with oxygen: Magnesium metal gives magnesium oxide when reacts with oxygen. Magnesium burnt with dazzling light in air and produces lot of heat.

  $2Mg + O_2 \rightarrow 2MgO$
- Reaction of aluminium metal with oxygen: Aluminium metal does not react with oxygen at room temperature but it gives aluminium oxide when burnt in air.

  $4Al + 3O_2 \rightarrow 2Al_2O_3$
- Reaction of potassium with oxygen: Potassium metal forms potassium oxide when reacts with oxygen.

  $4K + O_2 \rightarrow 2K_2O$
- Reaction of sodium with oxygen: Sodium metal forms sodium oxide when reacts with oxygen.

  $4Na + O_2 \rightarrow 2Na_2O$

  Lithium, potassium, sodium, etc. are known as alkali metals. Alkali metals react vigorously with oxygen.
- Reaction of Iron metal with oxygen: Iron does not react with oxygen at room temperature. But when iron is heated strongly in air, it gives iron oxide.

  $3Fe + 2O_2 \rightarrow Fe_3O_4$

  Iron fillings give sparkle in flame when burnt.
- Reaction of copper metal with oxygen: Copper does not react with oxygen at room temperature but when burnt in air, it gives copper oxide.

  $2Cu + O_2 \rightarrow 2CuO$
- Reaction of zinc metal with oxygen: Zinc does not react with oxygen at room temperature.

But it gives zinc oxide when heated strongly in air.

$2Zn + O_2 \rightarrow 2ZnO$

REACTION OF METALS WITH WATER:

Metals form respective metal hydroxide and hydrogen gas when react with water.

Metal + Water → Metal hydroxide + Hydrogen

Most of the metals do not react with water. However, alkali metals react vigorously with water.

Examples:

- Reaction of sodium metal with water: Sodium metal forms sodium hydroxide and liberates hydrogen gas along with lot of heat when reacts with water.

  $Na + H_2O \rightarrow NaOH + H_2$
- Reaction of aluminium metal with water: Reaction of aluminium metal with cold water is too slow to come into notice. But when steam is passed over aluminium metal; aluminium oxide and hydrogen gas are produced.

  $2Al + 3H_2O \rightarrow Al_2O_3 + 2H_2$
- Reaction of zinc metal with water: Zinc metal produces zinc oxide and hydrogen gas when steam is passed over it. Zinc does not react with cold water.

  $Zn + H_2O \rightarrow ZnO + H_2$
- Reaction of Iron with water: Reaction of iron with cold water is very slow and come into notice after a long time. Iron forms rust (iron oxide) when reacts with moisture present in atmosphere.

  Iron oxide and hydrogen gas are formed by passing of steam over iron metal.

  $3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2$
- Reaction of potassium metal with water: Potassium metal forms potassium hydroxide and liberates hydrogen gas along with lot of heat when reacts with water.

  $K + H_2O \rightarrow KOH + H_2$
- Reaction of calcium metal with water: Calcium forms calcium hydroxide along with hydrogen gas and heat when reacts with water.

  $Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$
- Reaction of magnesium metal with water: Magnesium metal reacts with water slowly and forms magnesium hydroxide and hydrogen gas.

  $Mg + 2H_2O \rightarrow Mg(OH)_2 + H_2$
- When steam is passed over magnesium metal, magnesium oxide and hydrogen gas are formed.

REACTION OF METALS WITH DILUTE ACID:

Metals form respective salts when react with dilute acid.

Metal + dil. acid → Metal salt + Hydrogen

Examples:

- Reaction of aluminium with dilute hydrochloric acid: Aluminium chloride and hydrogen gas are formed.

  $2Al + 6HCl \rightarrow 2AlCl_3 + 3H_2$
- Reaction of zinc with dilute sulphuric acid: Zinc sulphate and hydrogen gas are formed when zinc reacts with dilute sulphuric acid. This method is used in laboratory to produce hydrogen gas.

$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$

- Reaction of sodium metal with dilute acid: Sodium metal gives sodium chloride and hydrogen gas when react with dilute hydrochloric acid.

$2Na + 2HCl \rightarrow 2NaCl + H_2$

- Reaction of potassium with dilute sulphuric acid: Potassium sulphate and hydrogen gas are formed when potassium reacts with dilute sulphuric acid.

$2K + H_2SO_4 \rightarrow K_2SO_4 + H_2$

- Reaction of magnesium metal with dilute hydrochloric acid: Magnesium chloride and hydrogen gas are formed when magnesium reacts with dilute hydrochloric acid.

$Mg + 2HCl \rightarrow MgCl_2 + H_2$

Copper, gold and silver are known as noble metals. These do not react with water or dilute acids.

WEIGHT: (Light) → (Heavy)

Metal		Reaction with AIR	Reaction with WATER	Reaction with ACIDS
Potassium	K	Burn vigorously to form metal oxides	React with cold **water H_2O (l)** to form $H_{2\,(g)}$ and (metal)$OH_{(aq)}$	Strong reaction with **diluted acid (aq)** to form $\mathbf{H_{2\,(g)}}$. Metal replaces H in compound to form a salt.
Sodium	Na			
Calcium	Ca	Burn with decreasing vigour down the series to form metal oxides		
Magnesium	Mg		Only reacts with **steam $H_2O(g)$** to form $H_{2\,(g)}$ and metal oxide	
Aluminium	Al			
Zinc	Zn			
Iron	Fe			
Lead	Pb	React slowly (when heated) to form an oxide layer	No reaction	React with **concentrated acid (l)**. Metal replaces H to make a salt. Some of the acid decomposes into $\mathbf{NO_{2\,(g)}}$ and $\mathbf{H_2O}_{\,(l)}$.
Copper	Cu			
Mercury	Hg			
Silver	Ag	No reaction		No reaction
Gold	Au			

METAL OXIDES: CHEMICAL PROPERTIES

Metal oxides are basic in nature. Aqueous solution of metal oxides turns red litmus blue.

REACTION OF METAL OXIDES WITH WATER:

Most of the metal oxides are insoluble in water. Alkali metal oxides are soluble in water. Alkali metal oxides give strong base when dissolved in water.

Examples:

- Reaction of sodium oxide with water: Sodium oxide gives sodium hydroxide when reacts with water.

  $Na_2O + H_2O \rightarrow 2NaOH$

- Reaction of magnesium oxide with water: Magnesium oxide gives magnesium hydroxide with water.

  $MgO + H_2O \rightarrow Mg(OH)_2$

- Reaction of potassium oxide with water: Potassium oxide gives potassium hydroxide when reacts with water.

  $K_2O + H_2O \rightarrow 2KOH$

- Reaction of zinc oxide and aluminium oxide: Aluminium oxide and zinc oxide are insoluble in water. Aluminium oxide and zinc oxide are amphoteric in nature. An amphoteric substance shows both acidic and basic character. It reacts with base like acid and reacts with acid like a base.

  When zinc oxide reacts with sodium hydroxide, it behaves like an acid. In this reaction, sodium zicate and water are formed.

  $ZnO + 2NaOH \rightarrow Na_2ZnO_2 + H_2O$

- Zinc oxide behaves like a base when reacts with acid. Zinc oxide gives zinc chloride and water on reaction with hydrochloric acid.

  $ZnO + 2HCl \rightarrow ZnCl_2 + H_2O$

- In similar way aluminium oxide behaves like a base when reacts with an acid and behaves like an acid when reacts with a base.
- Aluminium oxide gives sodium aluminate along with water when reacts with sodium hydroxide.

  $Al_2O_3 + 2NaOH \rightarrow 2NaAlO_2 + H_2O$

- Aluminium oxide gives aluminium chloride along with water when it reacts with hydrochloric acid.

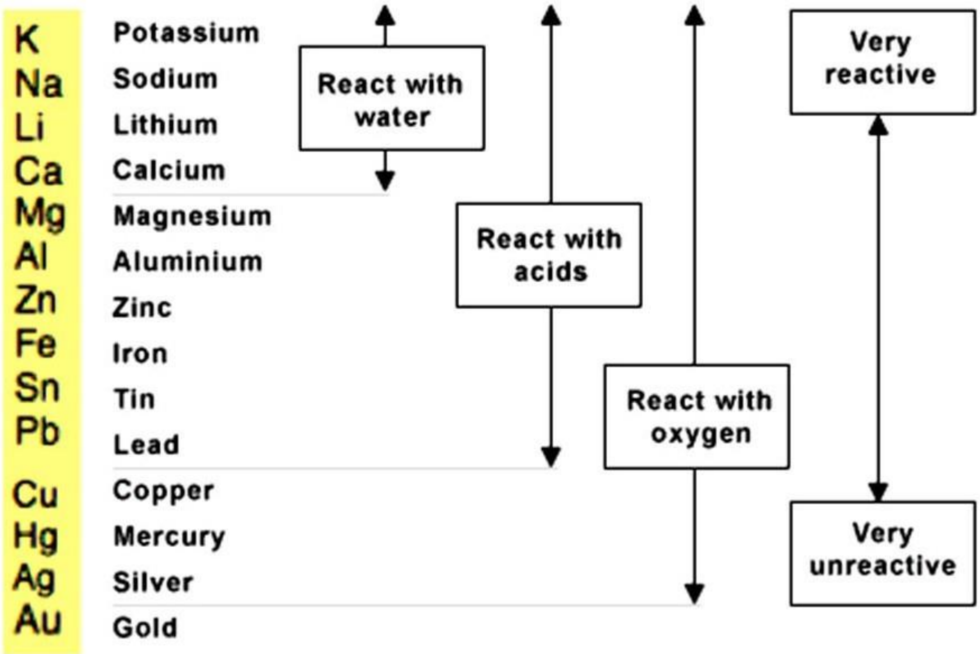
  $Al_2O_3 + 6HCl \rightarrow 2AlCl_3 + 3H_2O$

REACTIVITY SERIES OF METALS

A series of metallic elements arranged in the increasing or decreasing order of their reactivity is called a reactivity series of metals.
In the reactivity series, copper, gold, and silver are at the bottom and hence least reactive. These metals are known as noble metals.

The most active metal, potassium, is at the top of the list and the least reactive metal, gold, is at the bottom of the list. Although hydrogen is a non-metal it is included in the activity series due to the fact that it behaves like a metal in most chemical reactions i.e., the hydrogen ion has a positive charge [H^+] like other metals.

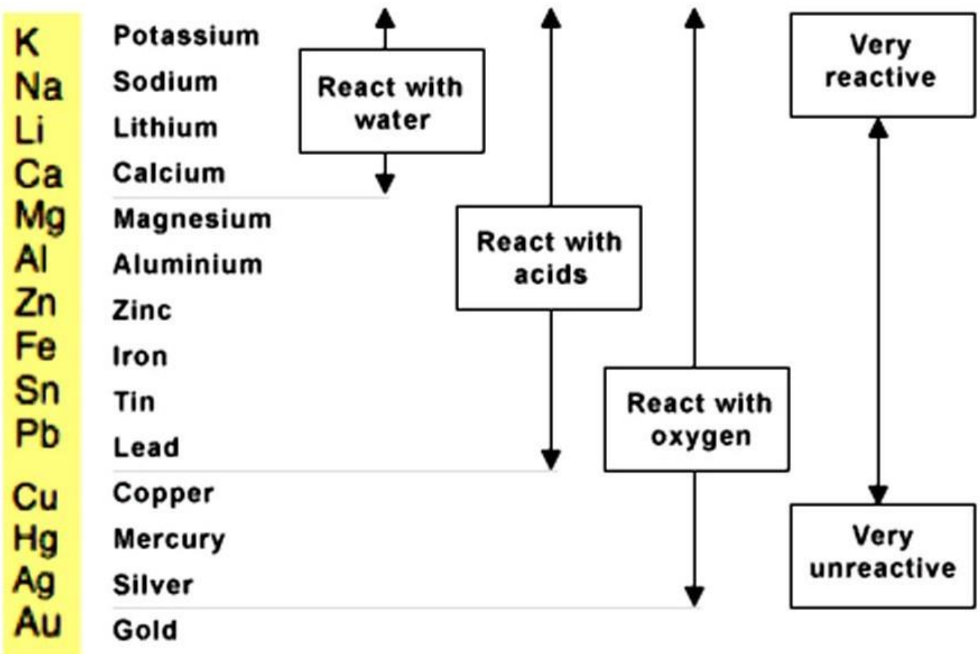
Reactivity Series of Metals



Following points become evident from the activity series of metals.

- The higher the metal in the series, the more reactive it is i.e., its reaction is fast and more exothermic.
- This also implies that the reverse reaction becomes more difficult i.e., the more reactive a metal, the more difficult it is to extract it from its ore. The metal is also more susceptible to corrosion with oxygen and water.

Reactivity Series of Metals



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- This also implies that the reverse reaction becomes more difficult i.e., the more reactive a metal, the more difficult it is to extract it from its ore. The metal is also more susceptible to corrosion with oxygen and water.
- The reactivity series can be established by observation of the reaction of metals with water, oxygen or acids.
- Within the general reactivity or activity series, there are some periodic table trends:
  a) Down Group 1(I) the "Alkali Metals", the activity increases Cs > Rb > K > Na > Li.
  b) Down Group 2(II) the activity increases e.g., Ca > Mg.
  c) In the same period, the Group 1 metal is more reactive than the group II metal and the group II metal is more reactive than the Group III metal and all three are more reactive than the "Transition Metals". e.g., Na > Mg > Al (in Period 3) and K > Ca > Ga > Fe/Cu / Zn etc. (in Period 4)

REACTION OF METALS WITH SOLUTION OF OTHER METAL SALTS:

Reaction of metals with solution of other metal salt is displacement reaction. In this reaction more reactive metal displace the less reactive metal from its salt.

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

Examples:

Iron displaces copper from copper sulphate solution.

$Fe + CuSO_4 \rightarrow FeSO_4 + Cu$

Similarly, aluminium and zinc displace copper from the solution of copper sulphate.

$2Al + 3CuSO_4 \rightarrow Al_2(SO_4)_3 + 3Cu$

$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$

In all the above examples, iron, aluminium and zinc are more reactive than copper. That's why they displace copper from its salt solution.

When copper is dipped in the solution of silver nitrate, it displaces silver and forms copper nitrate.

$Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$

In this reaction copper is more reactive than silver and hence displace silver from silver nitrate solution forming copper nitrate.

Silver metal does not react with copper sulphate solution. Because silver is less reactive than copper and not able to displace copper from its salt solution.

$Ag + CuSO_4 \rightarrow$ No reaction

Similarly, when gold is dipped in the solution of copper nitrate, no reaction takes place. Because copper is more reactive than gold.

$Au + CuSO_4 \rightarrow$ No reaction

In similar way no reaction takes place when copper is dipped in the solution of aluminium nitrate. Because copper is less reactive than aluminium.

$Al(NO_3)_3 + Cu \rightarrow$ No reaction

PHYSICAL PROPERTIES OF NON-METALS

- **Physical state –** Non-metals can exist in solid or liquid or gaseous state at room temperature. . For example, carbon, sulphur, phosphorus, iodine are in solid state, bromine is in liquid state while oxygen, nitrogen, chlorine are in gaseous state at room temperature.

- **Luster** – Non-metals do not have luster. They do not reflect light from their surface. (exception – diamond and iodine) Non-metals have dull appearance. For example, sulphur, phosphorus and carbon show this property.

- **Malleability** - Non-metals are non-malleable. If solids, they are brittle i.e. they break or shatter on hammering. For example, coal, sulphur, phosphorus are brittle.

- **Ductility** – Non-metals can not be drawn into thin wires. So they are not ductile.

- **Hardness** – Non-metals are usually not hard. They are soft. For example, coal, sulphur and phosphorus are soft. Diamond is exception to this. It is the hardest substance known.

- **Sonority:** Non-metals are not sonorous, i.e. they do not produce a typical sound no being hit.

- **Conduction** - Non- metals are usually poor conductors of heat and electricity. However, carbon in the form of gas carbon and graphite is exception to this. These forms of carbon are good conductors of electricity.

- **Density –** Non- metals which are gases have low density. Solid non-metals have low to moderate density. They are medium light. For example, sulphur, phosphorus and boron have densities 1.82, 2.07 and 2.34 respectively. . However, diamond has high density which is about 3.5.

- **Melting and boiling point –** Non-metals usually have low melting and boiling points. For example, phosphorus, sulphur, and iodine have melting points 440, 1150 and 1140 C respectively and boiling points 2800 , 4450 and 1840C respectively. . However, carbon, silicon and boron possess very high melting and boiling points.

- **Tensile strength** – Non-metals have low tensile strength i.e. they have no tenacity.

- **Color:** Non-metals are of many colors.

Physical Properties	Metals	Non-Metals
Malleability and Ductilily	Metals are malleable. They can be beaten into thin sheets. They are also ductile and can be drawn into wire (except a few metals like Na, K etc.)	Non-metals are neither malleable nor ductile. For e.g. coal, (carbon) and sulphur
Metallic Lusture	All the metals show metallic lusture.	They do not show any metallic lusture.
Hardness	Metals are generally hard	Non-metals are soft in comparison to metals
Physical state	They exist in solid and liquid states	Non-metals exist in solid, liquid and gaseous states.
Sonorous	Metals are sonorous and produce characteristic metallic sound when struck (e.g school bell)	They are non sonorous
Density	High density	Low density
Electrical conductivity	Good conductor of electricity	Bad conductor of electricity

CHEMICAL PROPERTIES OF NON-METALS

REACTION OF NON-METALS WITH OXYGEN:

Non-metals form respective oxide when react with oxygen.

Non-metal + Oxygen → Non-metal oxide

Examples:

- When carbon reacts with oxygen, carbon dioxide is formed along with production of heat.

  $C + O_2 \rightarrow CO_2 + Heat$

- When carbon is burnt in insufficient supply of air, it forms carbon monoxide. Carbon monoxide is a toxic substance. Inhaling of carbon monoxide may prove fatal.

$2C + O_2 \rightarrow 2CO + Heat$

- Sulphur gives sulphur dioxide when react with oxygen. Sulphur caught fire when exposed to air.

$S + O_2 \rightarrow SO_2$

- When hydrogen reacts with oxygen it gives water.

  $2H_2 + O_2 \rightarrow 2H_2O$

NON-METAL OXIDE: Non-metal oxides are acidic in nature. Solution of non-metal oxides turns blue litmus d.

Examples:

- Carbon dioxide gives carbonic acid when dissolved in water.

  $CO_2 + H_2O \rightarrow H_2CO_3$
- Sulphur dioxide gives sulphurous acid when dissolved in water.

  $SO_2 + H_2O \rightarrow H_2SO_3$
- Sulphur dioxide gives sulphur trioxide when it reacts with oxygen.

  $2SO_2 + O_2 \rightarrow 2SO_3$
- Sulphur trioxide gives sulphuric acid when dissolved in water.

  $SO_3 + H_2O \rightarrow H_2SO_4$

REACTION OF NON-METAL WITH CHLORINE:

Non metals give respective chloride when they react with chlorine gas.

Non-metal + Chlorine → Non-metal chloride

Examples:

- Hydrogen gives hydrogen chloride and phosphorous gives phosphorous trichloride when react with chlorine.

$H_2 + Cl_2 \rightarrow 2HCl$ $P_4 + 6Cl_2 \rightarrow 4PCl_3$

REACTION OF METAL AND NON-METAL

Many metals form ionic bonds when they react with non-metals. Compounds so formed are known as ionic compounds.

Ions: Positive or negative charged atoms are known as ions. Ions are formed because of loss or gain of electrons. Atoms form ion to obtain electronic configuration of nearest noble gas, this means to obtain stable configuration.

Positive ion: A positive ion is formed because of loss of electrons by an atom. Following are some examples of positive ions.

Examples:

- Sodium forms sodium ion because of loss of one electron. Because of loss of one electron; one positive charge comes over sodium.

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

- Similarly; potassium gets one positive charge by loss of one electron.

  $K \rightarrow K^+ + e^-$
- Magnesium forms positive ion because of loss of two electrons. Two positive charges come over magnesium because of loss of two electrons.

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

- Similarly calcium gets two positive charges over it by loss of two electrons.

  $Ca \rightarrow Ca^{++} + 2e^-$

Negative ion: A negative ion is formed because of gain of electron. Some examples are given below.

Examples:

- Chlorine gains one electron in order to achieve stable configuration. After loss of one electron chlorine gets one negative charge over it forming chlorine ion.

$Cl + e^- \rightarrow Cl^-$

- Similarly, fluorine gets one negative charge over it by gain of one electron forming chloride ion; in order to achieve stable configuration.

$F + e^- \rightarrow F^-$

- Oxygen gets two negative charge over it by gain of two electrons forming oxide ion; in order to obtain stable configuration.

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

USES OF METALS

Metals find number of applications. Some of them are given below.

- Zinc metal is used for galvanizing iron , in anti corrosion material, in medicinal fields and in alloys.
- Iron is used as a construction material in bridges, houses, ships etc. Iron, in the form of steel is used for making domestic utensils.
- Tin is used for soldering, for preparing foils, for metal coatings to prevent chemical action and corrosion, for panel lighting etc.
- Lead is used in making water pipes, in pigments, batteries, in alloys etc.
- Titanium finds extensive use in aircraft industries
- Pure metals, which display zero resistance to electrical currents, are called superconductors. Hg, Nb are examples of superconductors. They become superconductors below a critical temperature of 4.2 K and 9.2 K respectively. Superconductors have many applications in research and industry.
- Almost all metals including Zr, Ti find wide applications in atomic and space programmes and experiments.
- Mercury is used in thermometers.
- Silver, gold and platinum are precious metals and they are used in making ornaments.
- Radioactive metals like uranium and plutonium are used in nuclear power plants to produce atomic energy via nuclear fission.

USES OF NON - METALS

Non - metals find number of applications. Some of them are given below.

- Sulphur is used in making compounds like sulpha drugs, sulphuric acid, in matches, in gun powder , for vulcanization of rubber etc.
- Boron, in the form of compound borax, is used in making skin ointments.
- Phosphorus is used in making crackers.
- Oxygen is used for respiration.
- Chlorine, in the form of bleaching powder, is used for purification of water.
- Carbon is used as a fuel, as electrodes (graphite), as a reducing agent in metallurgy.
- Oxygen, hydrogen and nitrogen are used by all living things, they are the 'building blocks' of life.
- Iodine is used to prevent thyroid problems.
- Bromine is used in the preparation of dyes.
- Some compounds of fluorine (such as sodium fluoride, stannous fluoride) are added to toothpastes to prevent dental decays or formation of cavities.

<u>INTEXT QUESTIONS PAGE NO. 46</u>

Question 1: Why is sodium kept immersed in kerosene oil?

Answer : Sodium and potassium are very reactive metals and combine explosively with air as well as water. Hence, they catch fire if kept in open. Therefore, to prevent accidental fires and accidents, sodium is stored in kerosene oil.

Question 2: Write equations for the reactions of

(i) iron with steam

(ii)calcium and potassium with water

Answers:

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

Iron Steam Iron (II,III) oxide Hydrogen

(ii) $Ca_{(s)} + 2H_2O_{(l)} \longrightarrow Ca(OH)_{2(aq)} + H_{2(g)}$ + Heat

$2K_{(s)} + 2H_2O_{(l)} \longrightarrow 2KOH_{(aq)} + H_{2(g)}$ + Heat

Calcium/ Potassium Water Calcium Hydroxide/ Potassium hydroxide Hydrogen

Question 3: Samples of four metals A, B, C and D were taken and added to the following solution one by one. The results obtained have been tabulated as follows.

Metal	Iron(II) sulphate	Copper(II) sulphate	Zinc sulphate	Silver nitrate
A	No reaction	Displacement		
B	Displacement		No reaction	
C	No reaction	No reaction	No reaction	Displacement
D	No reaction	No reaction	No reaction	No reaction

Use the Table above to answer the following questions about metals A, B, C and D.

(i) Which is the most reactive metal?

(ii) What would you observe if B is added to a solution of copper (II) sulphate?

(iii) Arrange the metals A, B, C and D in the order of decreasing reactivity.

Answer: Explanation

A + FeSO4 → No reaction, i.e., A is less reactive than iron

A + CuSO4 →Displacement, i.e., A is more reactive than copperB

+ FeSO4 → Displacement, i.e., B is more reactive than iron

B + ZnSO4 →No reaction, i.e., B is less reactive than zincC

+ FeSO4 → No reaction, i.e., C is less reactive than iron

C + CuSO4 →No reaction, i.e., C is less reactive than copperC +

ZnSO4 →No reaction, i.e., C is less reactive than zinc

C + AgNO3 → Displacement, i.e., C is more reactive than silver

D + FeSO4/CuSO4/ZnSO4/AgNO3 →No reaction, i.e., D is less reactive than iron, copper, zinc, and silver

From the above equations, we obtain:

Zn — Most Reactive
B
Fe — Reactivity Series
A
Cu
C
Ag
D — Least Reactive

(i) B is the most reactive metal.

IONIC BONDS

Ionic bonding is the complete transfer of valence electron(s) between atoms. It is a type of chemical bond that generates two oppositely charged ions. In ionic bonds, the metal loses electrons to become a positively charged cation, whereas the nonmetal accepts those electrons to become a negatively charged anion. Ionic bonds require an electron donor, often a metal, and an electron acceptor, a nonmetal.

Ionic bonding is observed because metals have few electrons in their outer-most orbitals. By losing those electrons, these metals can achieve noble gas configuration and satisfy the octet rule. Similarly, nonmetals that have close to 8 electrons in their valence shells tend to readily accept electrons to achieve noble gas configuration. In ionic bonding, more than 1 electron can be donated or received to satisfy the octet rule. The charges on the anion and cation correspond to the number of electrons donated or received. In ionic bonds, the net charge of the compound must be zero.

FORMATION OF IONIC BOND:

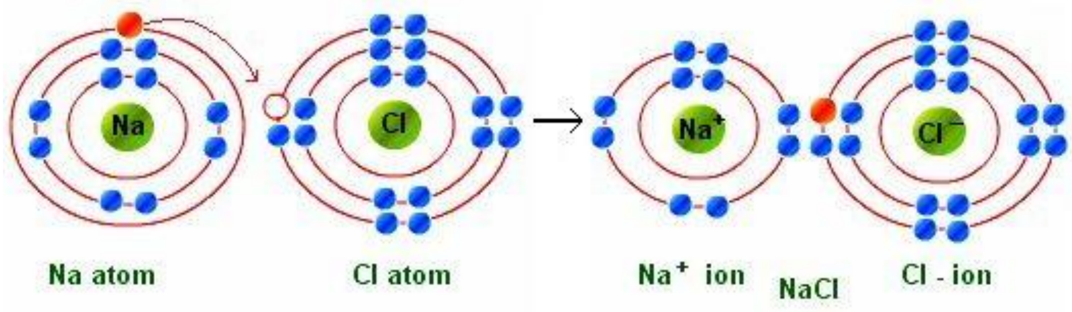
The positive ions (cations) and negative ions (anions) that are formed experience the electrostatic forces and get attracted to form chemical bond. As this bond is between charged particles known as ions, it is called *ionic bond*. Sometimes based on the forces being electrostatic, the bond is also called *the electrostatic bond*. As the valence concept has been explained in terms of electrons, it is also called the *electrovalent bond*.

Thus, we can define ionic bond as follows: The electrostatic attractive force that keeps cation and anion together to form a new electrically neutral entity is called *'ionic bond'*.

<u>EXAMPLES</u>

FORMATION OF SODIUM CHLORIDE (NaCl):

In sodium chloride; sodium is a metal (alkali metal) and chlorine is non-metal.
Atomic number of sodium = 11
Electronic configuration of sodium: 2, 8, 1
Number of electrons in outermost orbit = 1
Valence electrons = Electrons in outermost orbit = 1
Atomic number of chlorine = 17
Electronic configuration of chlorine: 2, 8, 7
Electrons in outermost orbit = 7
Therefore, valence electrons = 7



Sodium has one valence electron and chlorine has seven valence electrons. Sodium requires losing one electron to obtain stable configuration and chlorine requires gaining one electron in order to obtain stable electronic configuration. Thus, in order to obtain stable configuration sodium transfers one electron to chlorine.

After loss of one electron sodium gets one positive charge (+) and chlorine gets one negative charge after gain of one electron. Sodium chloride is formed because of transfer of electrons. Thus, ionic bond is formed between sodium and chlorine. Since, sodium chloride is formed because of ionic bond, thus it is called ionic compound. In similar way; potassium chloride (KCl) is formed.

FORMATION OF MAGNESIUM CHLORIDE ($MgCl_2$):

The atomic number of magnesium is 12

Electronic configuration of magnesium: 2, 8, 2

Number of electrons in outermost orbit = 2

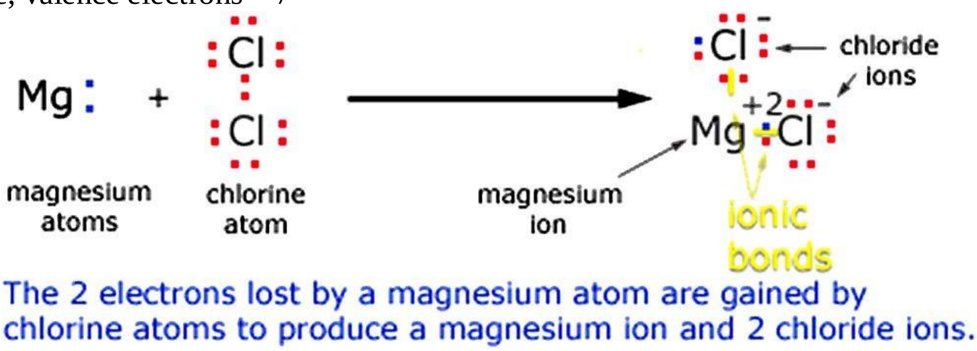
Valence electron = 2

Atomic number of chlorine = 17

Electronic configuration of chlorine: 2, 8, 7

Electrons in outermost orbit = 7

Therefore, valence electrons = 7



Magnesium loses two electrons in order to obtain stable electronic configuration. Each of the two chlorine atoms gains one electron lost by magnesium to obtain stable electronic configuration. The bonds so formed between magnesium and chlorine are ionic bonds and compound (magnesium chloride) is an ionic compound.

FORMATION OF CALCIUM CHLORIDE: ($CaCl_2$):

Atomic number of calcium is 20.

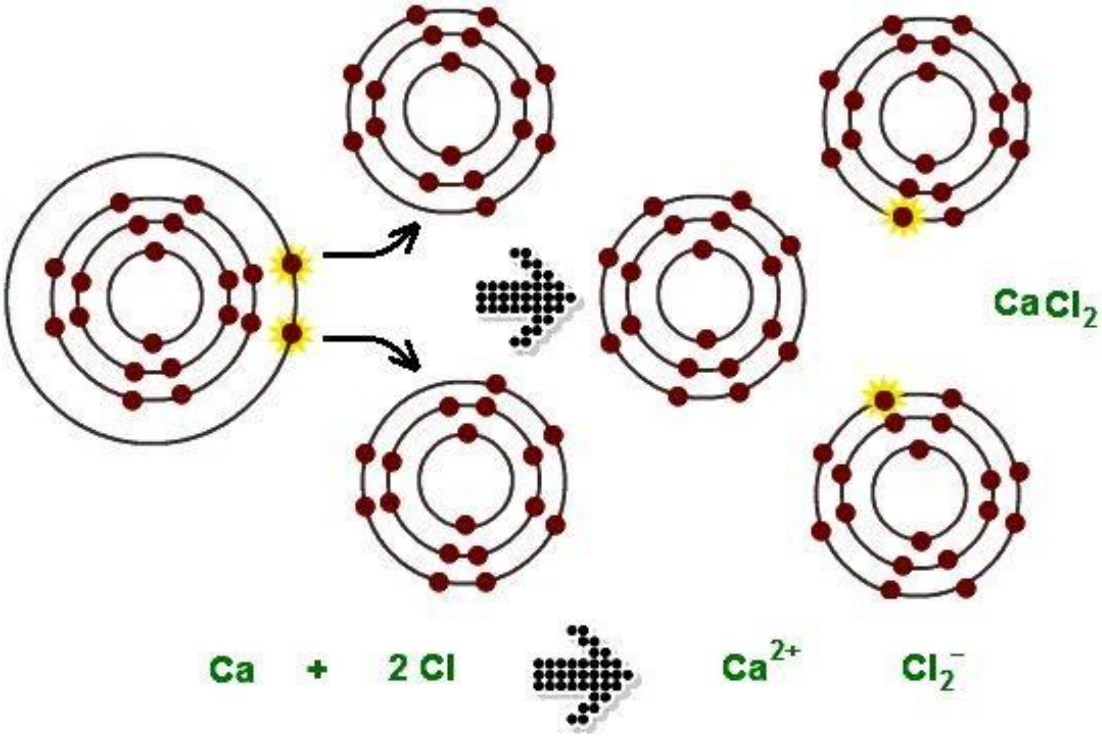
Electronic configuration of calcium: 2, 8, 8, 2

Number of electrons in outermost orbit = 2

Valence electron = 2

Valence electrons of chlorine = 7

Calcium loses two electrons in order to achieve stable electronic configuration. Each of the two chlorine atoms on the other hand gains one electron losing from calcium to get stability. By losing of two electrons calcium gets two positive charges over it. Each of the chlorine atoms gets one positive charge over it.



The bonds formed in the calcium chloride are ionic bonds and compound (calcium chloride) is an ionic compound. In similar way; Barium chloride is formed.

Formation of Calcium oxide (CaO):

Valence electron = 2

Atomic number of oxygen is 8

Electronic configuration of oxygen is: 2, 6

Number of electrons in outermost orbit = 6

Valence electron = 6

Calcium loses two electrons and gets two positive charges over it in order to get stability. Oxygen gains two electrons; lost by calcium and thus gets two negative charges over it.

Bond formed between calcium oxide is ionic bond. Calcium oxide is an ionic compound. In similar way; magnesium oxide is formed.

PROPERTIES OF IONIC COMPOUND:

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

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

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

- ***Conduction of Electricity*:** The conduction of electricity through a solution involves the movement of charged particles. A solution of an ionic compound in water contains ions, which move to the opposite electrodes when electricity is passed through the solution. Ionic compounds in the solid state do not conduct electricity because movement of ions in the solid is not possible due to their rigid structure. But ionic compounds conduct electricity in the molten state. This is possible in the molten state since the elecrostatic forces of attraction between the oppositely charged ions are overcome due to the heat. Thus, the ionsmove freely and conduct

OCCURENCE AND EXTRACTION OF METALS

Metals occur in nature in free as well as combined form. Metals having low reactivity show little affinity for air, moisture, carbon dioxide or other non-metals present in nature. Such metals may remain in elemental or native (free) state in nature. Such metals are called "noble metals" as they show the least chemical reactivity. For example gold, silver, mercury and platinum occur in free state.

On the other hand, most of the metals are active and combine with air, moisture, carbon dioxide and non-metals like oxygen, sulphur, halogens, etc. to form their compounds, like oxides, sulphides, carbonates, halides and silicates. i.e., they occur in nature in a combined state.

A naturally occurring material in which a metal or its compound occurs is called a *mineral*. A mineral from which a metal can be extracted economically is called an *ore*.

An ore is that mineral in which a metal is present in appreciable quantities and from which the metal can be extracted economically.

Metals found at the bottom of reactivity series are least reactive and they are often found in nature in free-state; such as gold, silver, copper, etc. Copper and silver are also found in the form of sulphide and oxide ores.

Metals found in the middle of reactivity series, such as Zn, Fe, Pb, etc. are usually found in the form of oxides, sulphides or carbonates.

METHODS USED TO ENRICH THE ORE

Hand picking

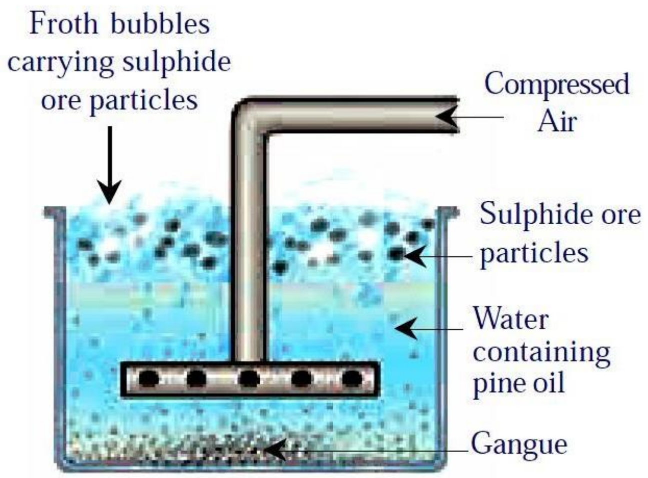
If the ore particles and the impurities are different in one of the properties like colour, size etc. Then using that property the ore particles are handpicked separating them from other impurities.

Washing

Ore particles are crushed and kept on a slopy surface. They are washed with controlled flow of water. Less densive impurities are carried away by water flow, leaving the more densive ore particles behind.

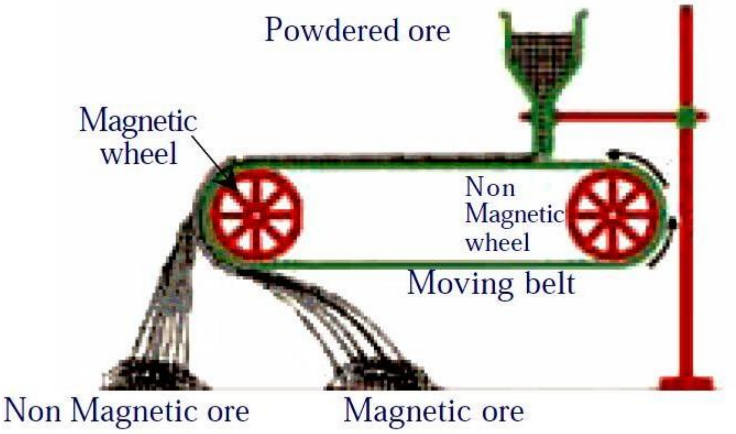
Froth flotaton

This method is mainly useful for sulphide ores which have no wetting property whereas the impurities get wetted. The ore with impurities is finely powdered and kept in water taken in a flotation cell. Air under pressure is blown to produce froth in water. Froth so produced, takes the ore particles to the surface whereas impurities settle at the bottom. Froth is separated and washed to get ore particles.



Magnetic separation

If the ore or impurity, one of them is magnetic and the other non-magnetic they are separated using electromagnets.



EXTRACTION OF CRUDE METAL FROM THE ORE:

After concentration and dressing of ore that obtained earth, we get a concentrated or enriched ore. To extract metal from this enriched ore it is converted into metallic oxide by reduction reaction. Then this metallic oxide further reduced to get a metal with certain impurities.

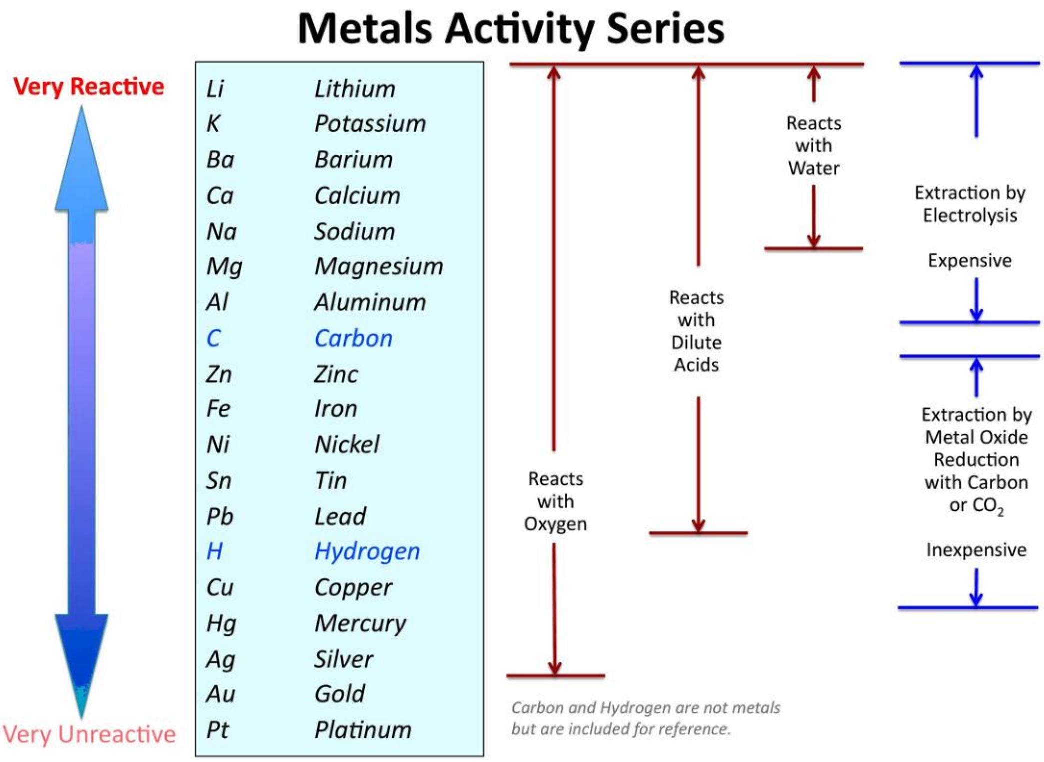
Extraction of the metal from its ores depends on the reactivity of the metal.

Arrange the metal in decreasing order of their reactivity is known as *activity series*. The classification of the metals on the basis of their reactivity:

EXTRACTION OF METALS AT THE TOP OF THE ACTIVITY SERIES:

(K, Na, Ca, Mg and Al). Simple chemical reduction methods like heating with C, CO etc to reduce the ores of these metals are not feasible. The temperature required for the reduction is too high and more expensive. To make the process economical, electrolysis methods are to be adopted. Again the electrolysis of their aqueous solutions also is not feasible because water in the solution would be discharged at the cathode in preference to the metal ions.

The only method viable is to extract these metals by electrolysis of their fused compounds. For example to extract Na from NaCl, fused NaCl is electrolysed with steal cathode (-) and graphite anode (+). The metal (Na) will be deposited at cathode and chloride liberated at the anode. At Cathode $2Na^{+} + 2e^{-}$ ’! $2Na$; and At Anode $2Cl^{-}$ ’! $Cl_2 + 2e^{-}$



Metals	Action of Oxygen	Reaction with cold water	Reaction with steam	Reaction with dilute strong Acids	Reaction with chlorine on heating
K	Form Na_2O, K_2O in limited supply of O_2 but form peroxides in excess of O_2	K to Mg displace H_2 from coldwater with decreasing reactivity {K violently but Mg very slowly}	K to Fe displace H_2 with steam without decreasing reactivity. {K very violently but Fe very slowly}	K to Pb displace H_2 from dilute strong acids with decreasing reactivity. {K-explosively, Mg-very vigorously, Fe-steadily, Pb-very slowly}	All metals react with Chlorine on heating toform their respective Chlorides but with decreasing reactivity from top to bottom. This is understood from the heat evolved when the metal reacts with one mole of Chlorine gas to form Chloride. KCl, NaCl, $CaCl_2$, $MgCl_2$, Al_2Cl_3, $ZnCl_2$, $FeCl_3$, $PbCl_2$, $CuCl_2$, $HgCl_2$, AgCl, $PtCl_3$ and $AuCl_3$ are formed
Na					
Ca	Burn with decreasing vigour to form oxides CaO, MgO, Al_2O_3, ZnO, Fe_2O_3				
Mg					
Al		From Al to Au do not displace H_2 from cold water			
Zn					
Fe					
Pb	Don't burn, but only form a surface layer of oxide PbO, CuO, HgO		From Pb to Au donot displace H_2 from steam		
Cu				Cu to Au do not displace H_2 from dilute strong acids	
Hg					
Ag	Don't burn or oxidise even on the surface				
Pt					
Au					

B) EXTRACTION OF METALS IN THE MIDDLE OF THE ACTIVITY SERIES:

(Zinc, iron, tin, lead and copper): The ore of these metals are generally present as Sulphides or Carbonates in native. Therefore prior to reduction of ores of these metals, they must be converted into metal oxides. **Sulphide** ores are converted into oxides by heating them strongly in excess of air. This process is known as *roasting*. Generally the sulphide ores are roasted to convert them into oxides before reducing them to metal.

Eg: $2PbS + 3O_2 \rightarrow 2PbO + 2SO_2$

The metal oxides are then reduced to the corresponding metal by using suitable reducing agent such as carbon

(i) **Reduction of metal oxides with carbon**: The oxides are reduced by coke in a closed furnace which gives the metal and carbon monoxide (CO).

Eg: $PbO + C \rightarrow Pb + CO$ at 1400^0C

(ii) **Reduction of oxide ores with CO**. Eg: $Fe_2O_3 + 3CO \rightarrow 2Fe + 3O_2$ in blast furnace

(iii) **Auto (self) reduction of sulphide ores:** In the extraction of Cu from its sulphide ore, the ore is subjected partial roasting in air to give its oxide.

$2Cu_2S + 3O_2 \rightarrow 2Cu2O + 2SO2$

When the supply of air is stopped and the temperature is raised. Therest of the sulphide reacts with oxide and forms the metal and SO2.$2Cu_2O + Cu_2S \rightarrow 6Cu + 2SO_2$

(iv) Reduction of ores (compounds) by more reactive metals.

Eg: $TiCl_4 + 2Mg \rightarrow Ti + 2MgCl_2$ at $850^0C$$TiCl_4 + 4Na \rightarrow Ti + 4NaCl$ at 850^0C

Thermite process: When highly reactive metals such as sodium, calcium, aluminium etc., are used as reducing agents, they displace metals of lower reactivity from the compound. These displacement reactions are highly exothermic. The amount of heat evolved is so large that the metals produced in molten state. This type of reaction is used in thermite process. The reaction of Iron (III) oxide (Fe2O3), with aluminium is used to join railings of railway tracks or cracked machine parts. This reaction is known as the **thermite reaction**.

$2Al + Fe_2O_3 \rightarrow Al_2O_3 + 2Fe + Heat$

$2Al + Cr_2O_3 \rightarrow Al_2O_3 + 2Cr + Heat$

C) EXTRACTION OF METALS AT THE BOTTOM OF THE ACTIVITY SERIES (AG, HG ETC)

Metals at bottom of the activity series are often found in free state. They reactivity with other atoms is very low. The oxides of these metals can be reduced to metals by heat alone and sometimes by displacement from their aqueous solutions.

(i) When cinnabar (HgS) whch is an ore of mercury, heated in air, it is first converted into (HgO) then reduced to mercury on further heating.

Eg: $2HgS + 3O_2 \rightarrow 2HgO + 2SO_2 \rightarrow 2HgO \rightarrow 2Hg + O_2$ on heating

(ii) Displacement from aqueous solutions:

Eg: $Ag_2S + 4CN^- \rightarrow 2[Ag(CN)_2]^- + S^{2-}$

$2[Ag(CN)_2]^-(aq) + Zn(s) \rightarrow [Zn(CN)_4]^{2-}(aq) + 2Ag(s)$

Here Ag_2S is dissolved in say KCN solution to get dicyanoargentate (I) ions. Fromthese ions Ag is precipitated by treating with Zn dust powder.

PURIFICATION OF THE CRUDE METAL:

The metal obtained by the reduction of the ore is usually contaminated with impurities like unchanged ore, other metals present in the ore and non metals from the anions in the ore.

For example, the (blister) copper obtained from its sulphide ore is a compound of copper iron pyrites ($CuFeS_2$). It contains some copper sulphide, iron and sulphur. It is purified by suitable methods including electrolysis. The process of obtaining the pure metal from the impure metal is called refining of the metal. Refining of the metal involves several types of processes. Some refining methods are given below:

a) Distillation
b) Poling
c) Liquation
d) Electrolysis etc.

The process that has to be adopted for purification of a given metal depends on the nature of the metal and its impurities?

a) **Distillation:** This method is very useful for purification of low boiling metals like zinc and mercury whether contain high boiling metals as impurities. The extracted metal in the molten state is distilled to obtain the pure metal as distillate.
b) **Poling:** The molten metal is stirred with logs (poles) of green wood. The impurities are removed either as gases or they get oxidized and form scum (slag) over the surface of the molten metal. Blister copper is purified by this method. The reducing gases, evolved from the wood, prevent the oxidation of copper.
c) **Liquation:** In this method a low melting metal like tin can be made to flow on a slopy surface to separate it from high melting impurities.
d) **Electrolytic refining:** In this method, the impure metal is made to act as anode. A strip of the same metal in pure form is used as cathode. They are put in a suitable electrolytic bath containing soluble salt of the same metal. The required metal gets deposited on the cathode in the pure form. The metal, constituting the impurity, goes as the anode mud.

The reactions are:

Anode: $M \rightarrow M^{n+} + ne^-$

Cathode: $M^{n+} + ne^- \rightarrow M$ (M = pure metal)

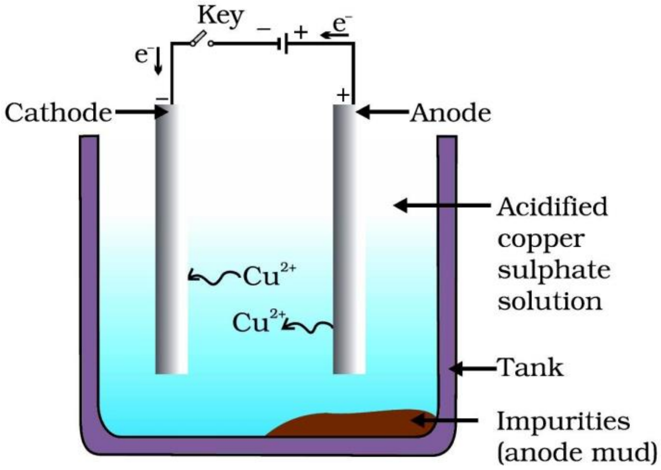
Where n = 1,2,3, ...

We use this electrolytic method to refine copper.

For this an impure copper is taken as anode and pure copper strips are taken as cathode. The electrolyte is a acidified solution of copper sulphate. As a result of electrolysis copper in pure form is transferred from the anode to the cathode.

Anode: $Cu \rightarrow Cu^{2+} + 2e^-$

Cathode: $Cu^{2+} + 2e^- \rightarrow Cu$



The suitable impurities go into the solution, where as insoluble impurities from the blister copper deposited at the bottom of anode as anode mud which contains antimony.

Selenium, tellurium, silver, gold and platinum; recovery of these elements may meet the cost of refining.

Zinc may also be refined this way.

CORROSION

Most of the metals keep on reacting with the atmospheric air. This leads to formation of a layer over the metal. In the long run, the underlying layers of the metal keep on getting lost due to conversion into oxides or sulphides or carbonate, etc. As a result, the metal gets eaten up. This process is called corrosion.

Rusting of Iron: Rusting of iron is the most common form of corrosion. When iron articles; like gate, grill, fencing, etc. come in contact with moisture present in air, the upper layer of iron turns into iron oxide. Iron oxide is brown-red in color and is known as rust. This phenomenon is called rusting of iron.

If rusting is not prevented in time, the whole iron article would turn into iron oxide. This is alsoknown as corrosion of iron. Rusting of iron gives huge loss every year.

Prevention of Rusting: For rusting, iron must come in contact with oxygen and water. Rusting is prevented by preventing the reaction between atmospheric moisture and the iron article. This can be done by painting, greasing, galvanization, electroplating, etc.

METHODS OF PREVENTION OF CORROSION

There are various methods of preventing corrosion and rusting of iron. Our main concern is to know the various methods to prevent the rusting of iron because iron is a strategic metal as it plays a very important role in the development of a nation. Some of the important methods of prevention of corrosion are as follows:

a) Painting

This is a common method of preventing iron from rusting. You might have observed that your parents paint iron gate in the garden and iron grills in your house. This painting prevents rusting by providing a coating over iron objects.

b) Oiling and greasing

To put a layer of oil and grease on the iron objects also prevents them from rusting. Iron parts of various machines and vehicles are oiled and greased to prevent rusting and to minimize friction.

c) Galvanization

Metals found at the top of the reactivity series are never found in free-state as they are veryreactive, e.g. K, Na, Ca, Mg and Al, etc.
Many metals are found in the form of oxides because oxygen is abundant in nature and is veryreactive.

TABLE : SOME IMPORTANT

Type of Ore	Metals (Common Ores)
Native Metals	Gold (Au), silver (Ag)
Oxide ores	Iron (Haematite, Fe_2O_3); Aluminium (Bauxite, $Al_2O_3 . 2H_2O$); Tin (Cassiterite, SnO_2); Copper (Cuprite, Cu_2O); Zinc (Zincite, ZnO); Titanium (Ilmenite, $FeTiO_3$, Rutile, TiO_2)
Sulphide ores	Zinc (Zinc blende, ZnS); Lead (Galena, PbS); Copper (Copper glance, Cu_2S); Silver (Silver glance or Argentite, Ag_2S); Iron (Iron pyrites, FeS_2)
Carbonate ores	Iron (Siferite, $FeCO_3$); Zinc (Calamine, $ZnCO_3$) , Lead (Cerrusite, $PbCO_3$)
Sulphate ores	Lead (Anglesite, $PbSO_4$)
Halide ores	Silver (Horn silver, AgCl); Sodium (Common salt or Rock salt, NaCl); Aluminium (Cryolite, Na_3AlF_6)
Silicate ores	Zinc (Hemimorphite, $2ZnO.SiO_2.H_2O$)

- The metal
- a mirror is reflected back among its ownpath.

	•			•	
	•			•	
Mirror:					

[METALS AND NONMETALS] |CHEMISTRY| STUDY NOTES

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[METALS AND NONMETALS] | CHEMISTRY| STUDY NOTES

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$Or, f = \frac{R}{2}$ $Or, R = 2f$

the mirror.