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- **All the materials around us are made up of chemical elements.**
- **There are more than 114 elements known at present.**

Classification of Elements :

Elements can be classified in many ways depending on similarities and differences in their properties.

One of the most useful classifications is to broadly classify elements into two main categories :

-----(i) Metals and ------------(ii) non-metals.

- Metals show many properties which are opposite to those of non metals.
- Metals are generally **hard solids except mercury which is a liquid**. The non metals are solids, liquids and gases.
- Of all the known elements, only 22 are non metals. Out of these, 10 non metals are solids, 1non metals is a liquid whereas the remaining11 non metals are gases.

---**Bromine is the only liquid non metal**.

- ***The periodic table is dominated by metals*** because of all the known elements about 80 % are metals.
  - --- **Iron** is the most commonly used metal.
  - --- **Aluminum** and **copper** are also commonly used metals in our household items.
  - --- **Gold** and **silver** are widely used for making jewelry.
  - --- **Silver** has a great ability to reflect light and is commonly used for making high quality mirrors.
  - --- **Copper** and **aluminum** are used in electrical gadgets are household wiring.
  - --- **Platinum** and **rhodium** are also being used in jewellery these days.

Non metals are also essential for life.

- --- **Oxygen** is used by plants and animals for their survival. It is also used for combustion reaction in homes, factories, airplanes and missiles.
- --- **Nitrogen** is widely used in its natural form and in its compounds.
- --- **chlorine** is used in water purification to kill germs.
- --- **Sulphur** is used in the manufacture of sulphuric acid, which is an important chemical and is regarded as king of chemicals.

Positions Of Metals and Non-Metals In The Periodic Table

- **Metals are placed on the left-hand side and in the centre of the periodic table. On the other hand, the non metals are placed on the right side of the periodic table.**
  Exception: Hydrogen (H) is an exception because it is a non metal but is placed on the left-hand side of the periodic table.

- Metals **and nonmetals are separated from each other in the periodic table by a zig zag line. The elements close to zig zag line show some properties of metals and some properties of non metals. These are called metalloids.** The common examples of metalloids are boron (B), silicon (Si), germanium (Ge) arsenic (As), antimony (Sb), tellurium (Te) and polonium (Po).

- Metals present at the extreme left are known as **light metals** while those present in the centre of the periodic table are called **heavy metals or transition metals.**

- Most metallic elements are on the **extreme left-hand side of the periodic table** whereas the most non-metallic elements are on the **extreme right-hand side of the periodic table**.
- **In general, the metallic character decreases on going from left to right side in the periodic table. Thus, The elements at the extreme left of the periodic table are most metallic and those on the right are least metallic or non metallic.**
  ---However, on going down the group, the metallic character increases. Thus, we can predict the metallic or non metallic character of the elements on the basic of their positions in the periodic table. For example,
  (i) sodium is more metallic than aluminum because sodium is on the left-hand side of aluminium.
  (ii)carbon is non metal while lead is metal because metallic character increases down the group.

Metals and their Characteristics

General properties of metals electronic configurations of metals

- The atoms of metals have 1 to 3 electrons in their outermost shells.

  For example, all the alkali metals have one electron in their outermost shells.

  Sodium, magnesium and aluminum are metals having 1, 2 and 3 electrons respectively in their valence shells as shown below :

Metal	At. No.	K	L	M
Sodium	11	2	8	1
Magnesium	12	2	8	2
Aluminium	13	2	8	3

Similarly, other elements have 1 to 3 electrons in their outermost shells. It may be noted that hydrogen and helium are excep tions because hydrogen is a non metal having only 1 electron in the outermost shell of its atom and helium is also a non metal having 2 electrons in the outermost shell.

Physical properties of metals

1. Metals are solids at room temperature

Metals such as iron, copper, aluminium, silver, gold, zinc, sodium, lead are solids at room temperature. In fact all metals are solids at room temperature. **Mercury is a liquid at room temperature.**

2. Metals are lustrous and can be polished

Most of the metals have metallic lustre and they can be polished.

The shining appearance of metals is also known as metallic lustre.

For example, if we take samples of iron, copper, aluminium and magnesium and clean their surfaces by rubbing them with a sand paper we observe that they have shiny surface. As we see, gold, silver and copper metals have their metallic lustre. Metals like aluminium and magnesium appear white. Gold is yellow in colour and copper is reddish in colour.

3 Metals are hard

Most of the metals hard . But all metals are not equally hard. If you take small pieces of some metals such as iron, copper, magnesium, aluminium, etc and try to cut these metals with a sharp knife, you will observe that the hardness of metals varies from metal to metal. The metals like iron, copper, aluminium etc. are quite hard. They cannot be cut with a knife. The most metals are hard and strong. Therefore, they can bear a heavy load on them.

- This property of metals is being used in the construction of buildings, bridges and heavy machines. However, sodium and potassium are common exceptions which are soft and can be easily cut with a knife.

4 Metals are malleable

Metals are generally malleable.

- ***The malleability means that the metals can be beaten with a hammer into very thin sheets without breaking.***

For example, place small pieces of some metals such as iron or copper on a block of iron and strike it 4 to 5 times with a hammer. We will find **that metal pieces become slightly larger**.

- This means that metals on being hammered can be beaten into thin sheets. This property of beating a metal into sheets is called malleability.
- All the metals can be beaten with a hammer to form very thin sheets.

" This property of beating a metal into sheets is called malleability".

- All the metals can be beaten with a hammer to form very thin sheets also called foils.

Example: Gold and silver are among the best malleable metals. They can be hammered into sheets which are much thinner than the paper. Aluminium and copper are also highly malleable metals. Because of this property, metals can be used for many purposes.

For example, gold and silver ornaments of different designs are prepared from small pieces of metals on hammering. Gold can be beaten into leaves as thin as 0.00002 mm.

5 Metals are ductile :

Ductility is also an important property of metals. It means that metals can be drawn into thin wires.

Gold and silver are the mot ductile metals. For example, **100 mg of silver can be drawn into a thin wire of about 200 meters long.**

Similarly, **we can draw a wire of about 2 kilometers from only one gram of gold.**

Copper and aluminium are also very ductile, and therefore, these can be drawn into thin wires which are used in electrical wiring.

Thus, we can say that metals are malleable and ductile.

- It is because of these properties of malleability and ductility that metals can be given different shapes to make various articles.

For example, silver foils are used for decorative purposes on sweets. Similarly, aluminium foils are used to wrap chocolates, cigarettes, medicines, food stuffs, etc. These are also used to seal bottles and containers.

6 Metals are good conductors of heat and electricity :

All metals are good conductor of heat.

The conduction of heat is called thermal conductivity. Silver is the best conductor of heat. Copper and aluminium are also good conductors of heat and therefore, they are used for making household utensils. Water boilers are generally made of copper and aluminium because they are good conductors of heat. Lead is the poorest conductor of heat. Mercury metal is also a poor conductor of heat.

Metals allow the current to pass through them. Thus, **metals are good conductors of electricity.**

- The electrical and thermal conductivities of metals are due to the presence of free electrons in them.
- Among all metals, silver is the best conductor of electricity. Copper and aluminium are the next best conductors of electricity. Since silver is expensive therefore, copper and aluminium are commonly used for making electric wires. However, metals like iron and mercury offer comparatively greater resistance to the flow of current and therefore, they have low electrical conductivities.

7 Metals have high densities

Most of the metals are heavy and have high densities. For example, the density of mercury metal is very high.

Exceptions: Sodium, potassium, magnesium and aluminium have low densities. Densities of metals are generally proportional to their atomic masses. The smaller the atomic mass and the greater the size of the metal atom, the smaller is its density.

Example : Magnesium, aluminium and titanium are used in construction purposes because they are light metals. They are commonly used to make transportation equipments.

8 Metals have high melting and boiling points

Most of the metals have high melting and boiling points.

- Tungsten has very high melting point and therefore, it is used whenever high temperatures are required.

For example, it is used in filaments of electric bulbs.

9 Metals are strong

Most of the metals are strong and they have high tensile strength.

10 Metals are sonorous

Most of the metals produce sound when they strike a hard surface or hard object.

The metals which produce sound on hitting a hard object or surface are called sonorous. That is why school bells are made up of metals.

NON- METALS and their physical properties

Non metals are very few in number as compared to metals.

- Among the 114 elements known these are only 22 non metals. Some common examples of nonmetals are carbon, Sulphur, iodine, oxygen, hydrogen, phosphorous, chlorine, etc.

These are either solids or gases except bromine which is a liquid.

Though nonmetals are small in number yet they represent an extremely important class of elements.

- Nonmetals display many properties opposite to those of metals.

1.Electronic Configurations of non metals

The atoms of non metals have usually 4 to 8 electrons in their outermost shells. For examples, carbon, nitrogen, oxygen, fluorine and neon are non metals. The electronic configurations of these non metals are :

Non-Metal	Atomic No.	Electronic configuration	
		K	L
Carbon	6	2	4
Nitrogen	7	2	5
Oxygen	8	2	6
Fluorine	9	2	7
Neon	10	2	8

It is clear that the non metals have usually 4 to 8 electrons in their outermost shells.

- However, there are two exceptions namely hydrogen and helium. These have one and two electrons respectively in their valence shells but are non metals.

Physical properties of non metals

1. NON- metals may be solid, liquid or gases at room temperature.

The non metals exist in all the three states. For examples :

Gaseous non metals : Hydrogen (H_2), oxygen (O_2), nitrogen (N_2), fluorine (F_2), chlorine (Cl_2), helium (He), neon (Ne), argon (Ar).

Liquid non metal : Bromine

Solid non metals : Carbon (C), Phosphorous (P_4), Sulphur (S_8), Iodine (I)

2. Non metals are non lustrous and cannot be polished

Most of the non metals are non lustrous and dull and these cannot be polished. Only graphite and iodine are lustrous non metals.

3. Non metals are generally soft.

Most of the non metals are soft, except diamond. Diamond is the hardest known substance .

4. Non metals are brittle.

Non metals are brittle and these break into pieces when hammered or stretched. For example, sulphur and phosphorous are brittle non metals. ■ Since non metals are not malleable, they cannot be beaten into sheets.

5. Non metals are not ductile.

Non metals are not ductile and, therefore, these cannot be drawn into thin wires.

6. Non metals are bad conductors of heat and electricity .

Non metals are generally bad conductors of heat and electricity. **This is due to fact that the non metals do not have free electrons.**

■ Exception. The allotropic form of carbon, graphite is a good conductor of heat and electricity like metals. Therefore, graphite is used for making electrodes.

7. Non metals have low densities.

Most of the non metals are light. For example, the density of Sulphur is 2 g cm^{-3} .

8. Non metals have generally low melting and boiling points.

Most of the non metals have low melting and boiling points ■ except **graphite** which has high melting point.

9. Non metals are not sonorous.

Non metals do not produce any sound when struck on a hard surface. Therefore, non metals are not sonorous.

Some Exceptions To General Trends

A comparison of physical properties of metals and non metals shows that we cannot group elements according to their physical properties alone. There are many exceptions to the general properties. **Some of the exceptions are** :

1. All metals **except mercury** are **solids at room temperature**.
2. **Metals in general have very high melting and boiling points**. However, **gallium** and **cesium** have very low melting points. Gallium has such a low melting point that it melts on our palm.
3. **Metals are generally hard,** but **sodium** and **potassium** are the examples of metals which are very soft. Therefore, unlike metals these can be easily cut with a knife. They have low densities and low melting points.
4. **Non metals do not have shiny lustre** but **iodine** is a non-metal and has luster.
5. Non metals are generally soft. But **carbon** is a non metals which can exist in different form called allotropes. **Diamond** is an allotrope of carbon and is very hard. It is the hardest natural substance known and has very high melting and boiling points. Another allotrope of carbon is graphite. It is good conductor of electricity unlike other non metals.

Chemical Properties of METALS

Metals and non metals react differently with other substances and form different products.

■ Most metals form basic oxides when dissolved in water. On the other hand, non metals form acidic oxides when dissolved in water.

1. Reaction of metals with oxygen

Magnesium burns in air with a brilliant white flame. some metals give characteristics colour of the flame.
For example, sodium burns with a golden yellow colour and copper burns with bluish green colour.

- **Metals react with air or oxygen to form oxides.**

Metal + Oxygen ⟶ Metal oxide

[a] when **magnesium** is heated in air, it combines with oxygen to form magnesium oxide.

$$2Mg + O_2 \longrightarrow 2MgO$$
Magnesium ⟶ Magnesium oxide

[b] **copper** reacts with oxygen, when heated in air to form copper (II) oxide, a black oxide.

$$Cu + O_2 \longrightarrow 2CuO$$
Copper ⟶ Copper (II) oxide

[c] **Aluminium** forms aluminium oxide

$$4Al + 3O_2 \longrightarrow 2Al_2O_3$$
Aluminium ⟶ Aluminium oxide

- **These oxides are basic in nature. When these oxides are dissolved in water, they give alkaline solutions.**

[d] **Sodium metal** reacts with oxygen of the air and forms sodium oxide.

$$4Na\,(s) + O_2\,(g) \longrightarrow 2Na_2O\,(s)$$
Sodium oxide

Sodium oxide reacts with water to form an alkali called sodium hydroxide. Therefore, sodium oxide is a basic oxide.

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

Due to the formation of sodium hydroxide, the solution of sodium oxide in water turns red litmus blue.

Similarly, potassium oxide dissolves in water to produce alkalies as :

$$4K\,(s) + O_2\,(g) \longrightarrow 2K_2O\,(s)$$
$$K_2O\,(s) + H_2O\,(l) \longrightarrow KOH\,(aq)$$

Reactivity of metals with oxygen

All metals do not react with oxygen with equal ease. ■ The reactivity of oxygen depends upon the metal. Some metals react with oxygen even at room temperature, some react on heating while still others react only on strong heating.

For example, (i) Metals like sodium, potassium and calcium react with oxygen even at room temperature to form their oxides.

$$4Na + O_2 \longrightarrow 2Na_2O$$
Sodium ⟶ Sodium oxide

$$4K + O_2 \longrightarrow 2K_2O$$
Potassium ⟶ Potassium oxide

$$2Ca + O_2 \longrightarrow 2CaO$$
Calcium ⟶ Calcium oxide

(ii) Metals like magnesium and zinc do not react with oxygen at room temperature. They burn in air only on strong heating to form corresponding oxides.

$$2Mg + O_2 \longrightarrow 2MgO$$
Magnesium ⟶ Magnesium oxide

$$2Zn + O_2 \longrightarrow 2ZnO$$
Zinc ⟶ Zinc Oxide

(iii) Metals like iron and copper do not burn in air even on strong heating. However, they react with oxygen only on prolonged heating. Iron filings burn vigorously when sprinkled in the flame of the burner.

$$3Fe + 2O_2 \xrightarrow{Heat} 3Fe_3O_4$$
Iron ⟶ Iron (II, III) oxide

Fe_3O_4 is a mixture of iron (II) oxide or ferrous oxide (FeO) and iron (III) oxide or ferric oxide (Fe_2O_3).

∴ it is written as iron (II, III) oxide.

$$2Cu + O_2 \xrightarrow{Heat} 2CuO$$
Copper (II) oxide

Copper has low reactivity with oxygen in comparison to iron.

--------From the above reactions of sodium, magnesium, zinc, iron and copper with oxygen, we observe that among these metals, sodium is the most reactive metal while copper is the least reactive metal.

The order of reactivity of these metals with oxygen is :

Na > Mg > Zn > Fe > Cu

Reactivity with oxygen decreases

- Metals like **silver** and **gold** ***do not react*** with oxygen **even at high temperatures**.
- At ordinary temperature, the surfaces of metals such as magnesium, aluminium, zinc, lead etc. are covered with a thin layer of oxide. This protective layer of oxide prevents the metal from further oxidation.

The reactivity of the metal to react with oxygen depends on the ease with which it can lose its valence electrons.

- ***The metal oxides are basic in nature.*** However, ***some metal oxides such as aluminium oxide*** (Al_2O_3), ***zinc oxide*** (ZnO), ***etc. show both acidic and basic behaviour. Such metal oxides are called amphoteric oxides***.

Such oxides react with both acids as well as bases to produce salt and water. For example, aluminium oxide reacts with acids and bases as :

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

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

- **Most metal oxides are insoluble in water. But some metal oxides dissolve in water to form alkali.** For example, sodium oxide, potassium oxide dissolve in water to form alkaline solutions.

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

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

2. Reaction of metals with water

-Metals react with water to form metal oxide or metal hydroxide and hydrogen.

Metal + Water ⟶ Metal oxide + Hydrogen

Metal oxide + Water ⟶ Metal hydroxide

- **The reactivity of metals towards water depends upon the nature of the metals. Some metals react even with cold water, some react with water only on heating while there are some metals which do not react even with steam. For example**,

(i) Sodium and potassium metals react **vigorously** with cold water to form **sodium hydroxide** and **hydrogen gas** is liberated.

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

$$\underset{\text{Potassium}}{2K\,(s)} + 2H_2O\,(l) \longrightarrow \underset{\text{Potassium hydroxide}}{2KOH\,(aq)} + H_2\,(g)$$

This reaction is so violent and exothermic that the hydrogen gas evolved catches fire.

(ii) Calcium reacts with cold water to form calcium hydroxide and hydrogen gas. The reaction is less violent.

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

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

(iii) **Magnesium reacts very slowly with cold water but reacts rapidly with hot boiling water forming magnesium oxide and hydrogen.**

$$\underset{\text{Magnesium}}{Mg\,(s)} + \underset{\text{Boiling water}}{H_2O\,(l)} \longrightarrow \underset{\text{Magnesium oxide}}{MgO\,(s)} + H_2\,(g)$$

It is also starts floating due to the bubbles of hydrogen gas sticking to its surface.

(iv) Metals like zinc and Aluminium do not react either with cold or hot water. **But they react only with steam to form metal oxide and hydrogen**.

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

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

(v) Iron metal does not react with water under ordinary conditions. **The reaction occurs only when steam is passes over red-hot iron and the products are iron (II, III) oxide and hydrogen.**

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

(vi) **Metals like copper, silver and gold do not react with water even under strong conditions**.

Conclusion: Metals react with water to form oxides or hydroxides and hydrogen gas is evolved. Out of sodium, magnesium, zinc, iron and copper, sodium is the most reactive while copper is the least reactive towards water. The order of reactivities of these metals with water is :

Na > Mg > Zn > Fe > Cu

Reactivity with water decreases

3. Reaction of metals with dilute acids

Many metals react with dilute acids and liberate hydrogen gas. Only less reactive metals such as copper, silver, gold, etc. do not liberate hydrogen from dilute acids. These metals are labelled as less reactive.

The reactions of metals with dilute hydrochloric acid and dilute sulphuric acid are similar. With dil. HC*l*, they give metal chlorides and hydrogen whereas with dil.H_2SO_4, they give metal sulphates and hydrogen.

(i) Sodium, magnesium and calcium react violently with dilute hydrochloric acid or dilute sulphuric acid liberating hydrogen gas and corresponding metal salt.

$$\underset{\text{Sodium}}{2Na\,(s)} + 2HCl\,(aq) \longrightarrow \underset{\text{Sodium chloride}}{2NaCl\,(aq)} + H_2\,(g)$$

$$\underset{\text{Sodium}}{2Na\,(s)} + H_2SO_4 \longrightarrow \underset{\text{Sodium Sulphate}}{Na_2SO_4\,(aq)} + H_2$$

Similarly,

$$\underset{\text{Magnesium}}{2Mg\,(s)} + 2HCl\,(aq) \longrightarrow \underset{\text{Magnesium chloride}}{MgCl_2\,(aq)} + H_2\,(g)$$

$$\underset{\text{Magnesium}}{Mg\,(s)} + H_2SO_4\,(aq) \longrightarrow \underset{\text{Magnesium sulphate}}{MgSO_4\,(aq)} + H_2\,(g)$$

Though both sodium and magnesium react violently with dil. acids, the reaction with magnesium is less violent than that of sodium. Therefore, sodium is more reactive than magnesium.

(ii) Metals like zinc and aluminium react with dil. HC*l* or dil. H_2SO_4 but the reaction is less rapid than that of magnesium.

$$\underset{\text{Zinc}}{Zn\,(s)} + 2HCl\,(aq) \longrightarrow \underset{\text{Zinc chloride}}{ZnCl_2\,(aq)} + H_2\,(g)$$

$$\underset{\text{Zinc}}{Zn\,(s)} + 2H_2SO_4\,(aq) \longrightarrow \underset{\text{Zinc sulphate}}{ZnSO_4\,(aq)} + 2H_2\,(g)$$

Similarly,

$$\underset{\text{Aluminium}}{2Al\,(s)} + 6HCl\,(aq) \longrightarrow \underset{\text{Aluminium chloride}}{2AlCl_3\,(aq)} + 3H_2\,(g)$$

$$\underset{\text{Aluminium}}{2Al\,(s)} + 3H_2SO_4\,(aq) \longrightarrow \underset{\text{Aluminium sulphate}}{Al_2(SO_4)_3\,(aq)} + 3H_2\,(g)$$

(iii) Iron reacts slowly with dilute HC*l* or dil. H_2SO_4 and therefore, it is less reactive than zinc and aluminium.

$$Fe\,(s) + 2HCl\,(aq) \longrightarrow \underset{\text{Ferrous chloride}}{FeCl_2\,(aq)} + H_2\,(g)$$

$$Fe\,(s) + H_2SO_4\,(aq) \longrightarrow \underset{\text{Ferrous sulphate}}{FeSO_4\,(aq)} + H_2\,(g)$$

(iv) Copper does not react with dil. HC*l* or dil. H_2SO_4 at all.

$$Cu\,(s) + HCl\,(aq) \longrightarrow \text{No reaction}$$

$$Cu\,(s) + H_2SO_4\,(aq) \longrightarrow \text{No reaction}$$

Therefore, copper is even less reactive than iron.

Conclusion: metals react with dil. HC*l* or dil. H_2SO_4 liberating hydrogen gas. Among the metals, sodium, magnesium, zinc, aluminium, iron and copper, sodium is most reactive metal, while copper is least reactive metal.

The order of reactivity of metals with dilute acids is :

Na > Mg > A*l* = Zn > Fe > Cu **Reactivity with dilute acids decreases**

4. Reaction of metals with chlorine

Metals react with chlorine to form metal chlorides. They are solids having high melting and boiling points. They also conduct electricity in their molten state or when dissolved in water. The reactivity of metals with chlorine is also different. Sodium, magnesium, calcium, zinc, etc. react with chlorine readily while metals like iron, copper, etc. react with chlorine on heating.

$$2Na + Cl_2 \longrightarrow 2NaCl$$

Sodium chloride

$$Ca + Cl_2 \longrightarrow CaCl_2$$

Calcium chloride

$$Zn + Cl_2 \longrightarrow ZnCl_2$$

Zinc chloride

$$Mg + Cl_2 \xrightarrow{\text{Heat}} MgCl_2$$

Magnesium chloride

$$2Fe + 3Cl_2 \xrightarrow{\text{Heat}} 2FeCl_3$$

Ferric chloride

Cu + Cl_2 —Heat→ $CuCl_2$
Cupric chloride

5. Reaction of metals with hydrogen

All metals do not combine with hydrogen. Only a few metals such as sodium, potassium, magnesium, calcium, etc. react with hydrogen to form hydrides. For example,

$2Na + H_2 \rightarrow 2NaH$ (Sodium hydride)

$2K + H_2 \rightarrow 2KH$ (Potassium chloride)

$Ca + H_2 \rightarrow CaH_2$ (Calcium hydride)

$Mg + H_2 \rightarrow MgH_2$ (Magnesium hydride)

- Calcium hydride is commercially known as hydrolith.

6. Reaction of metals with solutions of other metals

When a more reactive metal is placed in a salt solution of less reactive metal, then the more reactive metal displaces the less reactive metal from its salt solution. For example, take a solution of copper sulphate and put a strip of metal in th solution.

It is observed that the blue color of copper sulphate fades gradually and copper metal is deposited on the zinc strip.

This means that the reaction occurs :

$Zn (s) + CuSO_4 (aq) \rightarrow ZnSO_4 (aq) + Cu (s)$
Copper sulphate, Zinc sulphate

Zinc displaces copper from its solution. However, if we take zinc sulphate solution and put a strip of copper metal in this solution, no reaction occurs.

$ZnSO_4 (aq) + Cu (s) \rightarrow$ No reaction
Zinc sulphate

- Copper cannot displace zinc metal from its solution.

Conclusion : zinc is more reactive than copper.

However, if we put gold or platinum strip in the copper sulphate solution, then copper is not displaced by gold or platinum. **Thus, gold and platinum are less reactive than copper.**

- **Reaction of iron with copper sulphate solution**. When a strip of iron metal or iron filings are placed in copper Sulphate solution then blue colour of copper sulphate fades gradually and **red brown precipitate of copper** is formed.

$CuSO_4(aq) + Fe (s) \rightarrow FeSO_4 (aq) + Cu (s)$
Copper Sulphate, Ferrous sulphate, (Brown ppt.)

However, if a strip of copper metal is placed in ferrous sulphate solution, no reaction occurs. This means that iron an displace copper from its solution but copper cannot displace iron from its solution. **In other words, iron is more reactive than copper.**

- **Reaction of copper with silver nitrate solution**. When a strip of copper metal is placed in a solution of silver nitrate, the solution becomes gradually blue and a shining coating of silver metal gets deposited on the copper strip.

$2AgNO_3 (aq) + Cu (s) \rightarrow Cu (NO_3)_2 + 2Ag (aq)$
Silver nitrate, Copper nitrate, Silver

However, if we place silver wire in a copper sulphate solution no reaction occurs. This means that copper can displace silver from its solution, but silver cannot displace copper from its solution. Thus, **copper is more reactive than silver.**

Reactivity Series of Metals

Some metals are chemically very reactive while others are less reactive or do not react at all.

On the basis of reactivity of different metals with oxygen, water and acids as well as displacement reactions, the metals have Been arranged in the decreasing order of their reactivities.

"The arrangement of metals in order of decreasing reactivities is called reactivity series or activity series of metals".

- The activity series of some common metals ,the most reactive metal is placed at the top whereas the least reactive metal is placed at the bottom. As we go down the series the chemical reactivity of metals decreases.

Reactions for Different Reactivities

In the activity series of metals, **the basis of reactivity is the tendency of metals to lose electrons.**

---**If a metal can lose electrons easily to form positive ions, it will react readily with other substances. Therefore, it will be reactive metal.** On the other hand,

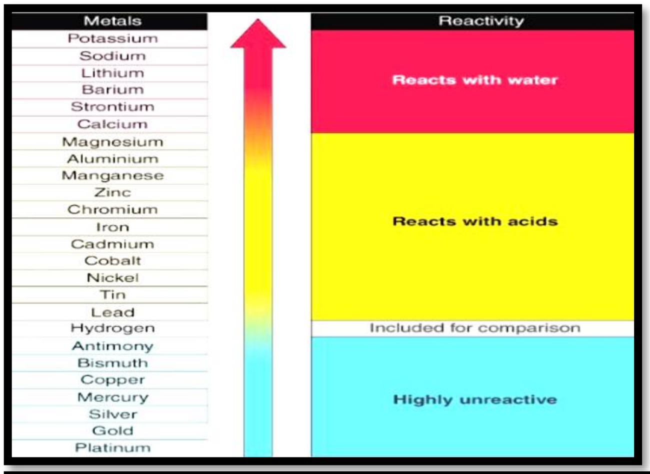
---**if a metal loses electrons less rapidly to form a positive ion, it will react slowly with the other substances.** Therefore, such a metal will be less reactive.

For example, **alkali metals such as sodium and potassium lose electrons very readily to form alkali metal ions, therefore, they are very reactive.**

- **Hydrogen is not a metal but even then, in has been placed in the reactivity series.** This is due to the fact that

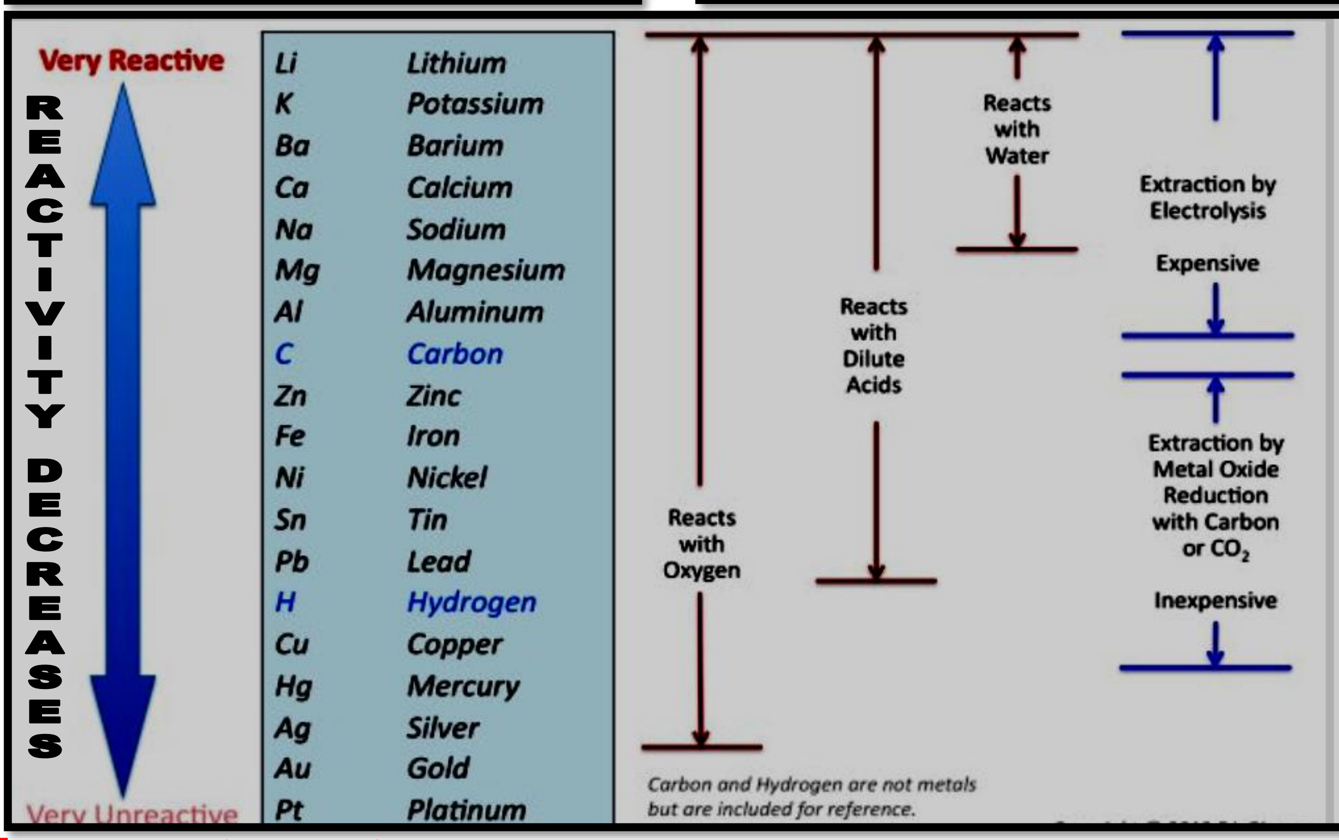
----(i) Hydrogen can also lose electron and form positive ion, H^+ .

----(ii) Hydrogen has also been included in the series to compare the reactivities of metals with respect to it.



How to remember the Reactivity Series?

Please	Potassium	Most reactive
Stop	Sodium	
Calling	Calcium	
Me	Magnesium	
A	Aluminium	
Careless	(Carbon)	
Zebra	Zinc	
Instead	Iron	
Try	Tin	
Learning	Lead	
How	(Hydrogen)	
Copper	Copper	
Saves	Silver	
Gold	Gold	Least reactive



Displacement of Hydrogen from Acids by Metals

All metals above hydrogen in the reactivity series like zinc, magnesium, nickel can liberate hydrogen from acids like HCl and H_2SO_4.

These metals have greater tendency to lose electrons than hydrogen. Therefore, the H^+ ions in the acids will accept electrons and give hydrogen gas as :

$$\underset{\text{Metal}}{M} \longrightarrow M^+ (aq) + e^-$$

H^+ (aq) + e^- ⟶ H
(From acid)
H + H ⟶ H_2

The metals which are below hydrogen in the reactivity series like copper, silver, gold cannot liberate hydrogen from acids like HCl, H_2SO_4. etc. These metals have lesser tendency to lose electrons than hydrogen. Therefore, they cannot lose electrons to H^+ ions.

Reactivity Series and Displacement Reactions

In general, a more reactive metal can displace the less reactive metal from its solution. For example, zinc displaces copper from its solution.

Zn (s) + $CuSO_4$ (aq) ⟶ $ZnSO_4$ (aq) + Cu (s)

This displacement reaction occurs because zinc is more reactive than copper and can readily lose electrons. These electrons are accepted by copper ions and from copper metal, which gets deposited on the zinc strip. Since the Cu^{2+} ions which give blue colour of the solution fades. As a result, the solution becomes colorless.

Zn (s) ⟶ Zn^{2+} + $2e^-$
Cu^{2+} (aq) + $2e^-$ ⟶ Cu (s)

Or Zn (s) + Cu^{2+} (aq) ⟶ Zn^{2+} (aq) + Cu (s)
(Blue colour) (Colourless)

or $CuSO_4$ (aq) + Zn (s) ⟶ $ZnSO_4$ (aq) + Cu (s)

■ However, copper cannot displace zinc from $ZnSO_4$.

Cu + $ZnSO_4$ ⟶ **No reaction** This is due to the fact that copper is less reactive than zinc. Therefore, copper cannot lose electrons in preference to zinc and the above reaction cannot occur.

USEFULNESS OF ACTIVITY SERIES

- (i) The metal which is higher in the activity series is more reactive than the other. Potassium is most reactive and platinum is least reactive.
- (ii) The metals which have been placed above hydrogen are more reactive than hydrogen and these can displace hydrogen from its compounds like water and acids to liberate hydrogen gas.
- (iii) The metals which are placed below hydrogen are less reactive than hydrogen and these cannot displace hydrogen from its compounds like water and acids.
- (iv) A more reactive metal can displace the less reactive metal from its solution.
- (v) Metals at the top of the series are very reactive and, therefore, they do not occur free in nature. The metals at the bottom of the series are least reactive and, therefore, they normally occur free in nature. For example, gold, the last element of the series is found almost as free element.

2NaBr (aq) + Cl_2 ⟶ 2NaCl(aq) + Br_2
Sodium bromide

However, bromine cannot displace chlorine from its salt solution.

2NaCl (aq) + Br_2 ⟶ No reaction . Therefore, chlorine is more reactive non metal than bromine.

Chemical Properties of Non-Metals

Non metals have usually 4 to 8 electrons in their outermost shells. They have the tendency to accept electrons to complete their octets. By accepting the electrons, they from negatively charged ions and, therefore, they are **electro negative elements.** For example, nitrogen, oxygen and fluorine can accept 3, 2 and 1 electrons respectively to complete their octets as :

N + $3e^-$ ⟶ N^{3-} gains 3 electrons
O + $2e^-$ ⟶ O^{2-} gains 2 electrons
F + e^- ⟶ F^- gains 1 electrons

Let us compare some of the reactions of non metals with those of metals. These are :

1. Reaction of non metals with oxygen

Non metals react with oxygen to form ***acidic or neutral oxides*** . These oxides are covalent in nature and are formed by sharing of electrons. The acidic oxides dissolve in water to give acids.

Acidic oxides

The oxides of **carbon, sulphur, phosphorous, etc, are acidic** and, therefore, they turn blue litmus solution red. For example,

(i) Carbon reacts with oxygen of air to form carbon dioxide.

C (s) + O_2 (g) ⟶ CO_2
Carbon Carbon dioxide

Carbon dioxide dissolves in water to form an acid called carbonic acid.

CO_2 (g) + H_2O (*l*) ⟶ H_2SO_3 (aq)
Carbonic acid

(ii) Sulphur burns in air to give Sulphur dioxide.

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

Sulphur dioxide dissolves in water to from an acid called Sulphurous acid.

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

Sulphur also forms an oxide, sulphur trioxide which dissolve in water to give sulphuric acid.

$$2S (s) + 3O_2 (q) \longrightarrow 2SO_3 (g)$$
$$SO_3 (g) + H_2O (l) \longrightarrow H_2SO_4 (aq)$$
Sulphuric acid

(iii) When phosphorous is burnt in air, it reacts with oxygen of air to form phosphorous pentaoxide (P_2O_5). This is also acidic oxide and dissolves in water to give an acid, phosphoric acid.

$$P_4 (s) + 5O_2 (g) \longrightarrow 2P_2O_5 (s)$$
Phosphorous pentoxide

$$2P_2O_5 (s) + 6H_2O (l) \longrightarrow 4H_3PO_4 (aq)$$
Phosphorous acid

These acidic oxides of non metals, neutralize bases to from salt and water.

For example, $SO_2 + 2NaOH \longrightarrow Na_2SO_3 + H_2O$
Sodium sulphite

Neutral oxides

Some oxides of non metals are neutral. For example, carbon monoxide (CO), nitric oxide (NO), nitrous oxide (N_2O), water (H_2O) etc. For example,

$$2C (s) + O_2 (g) \longrightarrow 2CO (g)$$
Carbon monoxide

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

These oxides do not turn blue litmus solution red.

2. Reaction of non metals with water

Non metals do not react with water or steam to give hydrogen gas. This is because nonmetal cannot give electrons to reduce the hydrogen ions of water into hydrogen gas.

3. Reaction with acids

Non metals do not react with dilute acids and, therefore, hydrogen gas is not liberated when non metals are treated with dilute acids. Therefore, non metals do not displace hydrogen from dilute acids. For example, carbon, sulphur or phosphorous do not react with dilute acids such as dil. HC*l* or dil. H_2SO_4 to produce hydrogen gas. We have seen that hydrogen can only be displaced from dilute acids if electrons are supplied of H^+ ions of the acid.

$$H_2SO_4(aq) \longrightarrow 2H^+ (aq) + SO_4^{2-} (aq)$$
$$2H^+ (aq) + 2e^- \longrightarrow H_2 (g)$$

But the non metals are electron acceptors and, therefore, they cannot give electron to H^+ ions of an acid. Hence hydrogen gas is not liberated.

4. Reaction with hydrogen

Non metals react with hydrogen under different conditions to form corresponding hydrides. For example, H_2O, H_2S, NH_3, HCl, CH_4 etc., are common hydrides of oxygen, Sulphur, nitrogen, chlorine and carbon respectively.

$$2H_2 (g) + O_2 (g) \xrightarrow{\text{Electric spark}} 2H_2O (l)$$
$$H_2 (g) + S (s) \longrightarrow H_2S (g)$$
$$N_2 (g) + 3H_2 (g) \xrightarrow[\text{400 – 600 atm}]{\text{775 K}} 2 NH_3 (g)$$
$$H_2 (g) + Cl_2 (g) \xrightarrow{\text{Diffused sunlight}} 2 HCl (g)$$

5. Reaction with chlorine

Non metals react with chlorine to form covalent chlorides such as HC*l*, PC*l*₃, CC*l*₄ etc. For example,

Non metals react with chlorine under different conditions to form corresponding chlorides. For example,

$$H_2 (g) + Cl_2 \xrightarrow{\text{Diffused sunlight}} 2HCl (g)$$
Hydrogen chloride

$$P_4 (s) + 6Cl_2 (g) \longrightarrow 4PCl_3 (l)$$

Phosphorous trichloride

6. Reaction with salt solutions

A more reactive non metal displaces a less reactive non metal from its salt solution. For example, when chlorine is passed through a solution of sodium bromide, then bromine is liberated.

Distinction between Metals and Non metals

Although there is no sharp line of distinction between metals and non metals yet there are some distinctive differences. The main points of difference between metals and non metals are summarized below :

Property	Metals	Non metals
Electronic structure	Metals have 1 to 3 electrons in the outermost shell of their atoms.	Non metals have 4 to 8 electrons in the outer most shell of their atoms (except hydrogen and helium which have 1 to 2 electrons respectively in their outer most shells).
Physical properties		
1. State of existence	Metals are mostly solids at room temperature except mercury which is a liquid metal.	Non metals exist in all the three states of matter i.e., solids, liquids and gases.
2. Lustre	Metals are lustrous and can be polished.	They are usually non lustrous except diamond which is a very hard substance.
3. Hardness	Metals are generally hard.	Non metals are comparatively soft .
4. Density	Metals have usually high densities except alkali and alkaline earth metals which are light.	Non metals usually have low densities.
5. Conductivity	**Metals** are good conductors of heat and electricity.	Non metals are poor conductors of heat and electricity. Graphite is an exception because it is a good conductor of electricity.
6. Malleability and ductility	Metals are usually malleable and ductile.	They are usually brittle.
Chemical properties		
7. Nature of oxides	Metals form basic oxides; some are amphoteric also.	Nonmetal's form acidic or neutral oxides.
8. Displacement of hydrogen from acids	Metals displace hydrogen from acids and form salts.	Nonmetals do not displace hydrogen from acids.
9. Reaction with chlorine	Metals react with Cl_2 to form electrovalent chlorides.	Nonmetals react with Cl_2 to form covalent chlorides.
10. Reaction with hydrogen	with hydrogen, only a few metals combine to Form electrovalent hydrides.	with hydrogen, non metals form many stable hydrides which are covalent.
11. Electropositive or electronegative character	Metals are electropositive in character.	Non metals are electronegative.
12. Oxidising and reducing agent	Metals act as reducing agents.	Non metals act as oxidizing agents.

Metals and Non metals react :

Metals have 1 to 3 electrons in their outermost shells while non metals have 4 to 8 electrons in their outermost shells. The atoms containing 8 electrons in their outermost shells are noble gases helium which has only two electrons in its outermost first K shell. The electronic configurations of some metals, nonmetals and noble gases are given in Table.

Electronic configuration of some elements.

Type of element	Element	Atomic number	Number of electrons in shells K	L	M	N	No. of electrons in outermost shell
Metals	Sodium (Na)	11	2	8	1		1
	Magnesium (Mg)	12	2	8	2		2
	Aluminium (Al)	13	2	8	3		3
	Potassium (K)	19	2	8	8	1	1
	Calcium (Ca)	20	2	8	8	2	2
Non metals	Nitrogen (N)	7	2	5			5
	Oxygen (O)	8	2	6			6
	Fluorine (F)	9	2	7			7
	Phosphorous (P)	15	2	8	5		5
	Sulphur (S)	16	2	8	6		6
	Chlorine (Cl)	17	2	8	7		7
Noble gases	Helium (He)	2	2				2
	Neon (Ne)	10	2	8			8
	Argon (Ar)	18	2	8	8		8
	Krypton (Kr)	36	2	8	18	8	8

◘The **noble gases** are very stable. It is clear from the table that **except for helium**, ***all other noble gases have eight electrons*** in their outermost shell. Helium, on the other hand has only two electrons in its first shell because the first shell cannot have more two electrons. This means that the noble gases have completely filled outer most shell. Since the noble gases do not take part in bonding, this means that eight electrons in outermost shell represent a highly stable electronic configuration.

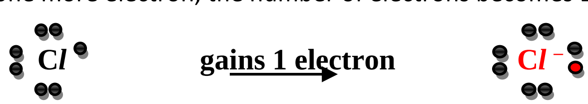
Due to this **stable configuration**, the noble gases have neither any tendency to lose nor gain electrons. ***Therefore, they remain as such and exist as monoatomic.***

▫ IONS {CATIOS & ANIONS}

◘ A sodium atom has one electron in its outermost shell. If it loses the electron from its outermost M shell than its L shell becomes the outermost shell which has also stable octet like noble gases. The nucleus of sodium atom still has 11 protons but the number of electrons has become 10. Therefore, **it becomes positively charged sodium ion or cation.**

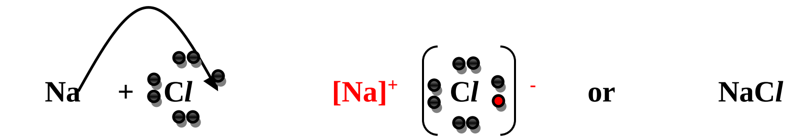
$$\underset{\substack{\text{Sodium atom} \\ (2,\,8,\,1)}}{Na} \longrightarrow \underset{\substack{\text{Sodium ion} \\ (2,\,8)}}{Na^+}$$

◘ let us consider a chlorine atom it has seven electrons in its outermost shell and it requires one electron more to complete its octet. Therefore, if sodium and chlorine were to react, the electron lost by sodium could be taken up by chlorine atom. After gaining an electron, the chlorine atom gets a unit negative charge. This is because the nucleus of chlorine has 17 protons and 17 electrons. When it gets one more electron, the number of electrons becomes 18. ***This makes chlorine ion, Cl^- as negatively charged***.



Chlorine atom (2, 8, 7) **Chlorine ion** 2, 8, 8

◘ In this case, both sodium and chlorine acquires complete octets. Now sodium and chlorine ions being oppositely charged ions attract each other and are held by strong electrostatic forces of attraction of exist as sodium chloride (NaCl). In other words, we may say that Na^+ and Cl^- ions are held together by electrovalent or ionic bond. This may be represented as :



(2, 8, 1) (2, 8, 7) (2, 8) (2, 8, 8)

The compound formed is called **ionic compound** or **electrovalent compound**. Thus, we can say that all atoms other than noble gases have less than eight electrons in their outermost shells. Therefore, they combine will each other or with other atoms to get stable electronic arrangement of noble gases. Thus, the tendency of atoms of various elements to get stable electronic **configuration of eight electrons in their valence shells is the cause of formation of chemical bond.** This principle of attaining maximum of eight electrons is also called **octet rule .**

Formation of more ionic compounds.

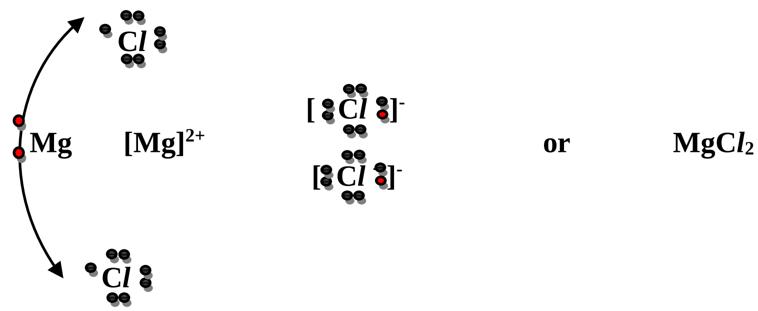
(i) Formation of magnesium chloride ($MgCl_2$). The electronic configuration of magnesium and chlorine atoms are :

Mg (At. no. = 12) : 2, 8, 2

Cl (At. no. = 17) : 2, 8, 7

Magnesium atoms has two electrons in its valence shell. It has a tendency to lose both of its electrons to attain the electronic configuration of the nearest noble gas. On the other hand, chlorine has only one electron less than the nearest noble gas electronic configuration. Therefore, magnesium loses its both the valence electrons to two chlorine atoms, each of which is in the need of one electron. This result in the formation of magnesium ion carrying two positive charges, Mg^{2+} and two chloride ions each having one negative charge, Cl^-. The Mg^{2+} ion has the electronic configuration 2,8 and each Cl^- ion has the electronic configuration 2, 8, 8. The Mg^{2+} ion and two Cl^- ions form bonds to give $MgCl_2$.

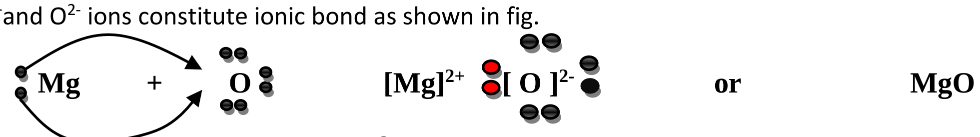
In this case, the valency of magnesium is **two** and that of chlorine is **one.**



(ii) Formation of magnesium oxide MgO. The electronic configurations of magnesium and oxygen are :

Mg (At. No = 12) : 2, 8, 2   O (At. No. = 8) : 2, 6

It is clear that magnesium can lose two electrons to form mg^{2+} ion, which has stable noble gas configuration (2, 8, 8). On the other hand, oxygen atom has two electrons less than the nearest noble gas configuration, so it can also acquire noble gas configuration by gaining two electrons. As a result, oxygen changes to O^{2-} ion. The electrostatics force of attraction between $Mg^{2+}$and O^{2-} ions constitute ionic bond as shown in fig.



Mg —loses $2e^-$→ $[Mg]^{2+}$
2, 8, 2   2, 8

◙In this case, the electro valency of both magnesium and oxygen is two.

◙ Structures of some common ionic compounds.

Sodium chloride	**Na** (2, 8, 1)	+	**Cl** (2, 7)	→	**$Na^+ [Cl]^-$ or NaCl**
Magnesium fluoride	**Mg** (2, 8, 2)	+	**2F** (2, 7)	→	**Mg^{2+} 2$[F]^-$ or MgF_2**
Calcium oxide	**Ca** (2, 8, 8, 2)	+	**O** (2, 6)	→	**$Ca^{2+} [O]^{2-}$ or CaO**
Aluminium oxide	**Al** (2, 8, 8, 1)	+	**O** (2, 6)	→	**$2K^+ [O]^{2-}$ or K_2O**
Magnesium chloride	**Mg** (2, 8, 2)	+	**2Cl** (2, 8, 7)	→	**Mg^{2+} 2 $[Cl]^-$ or $MgCl_2$**

◙ **The reactivity of metals to form metal oxide**. For example, the tendency of magnesium to form magnesium oxide can be easily explained. Magnesium atom gives up two electrons to from Mg^{2+} ion. Oxygen atom accepts two electrons to form O^{2-} ion. These ions are held together by electrostatic forces forming magnesium oxide .

◘CHARACTERISTICS PROPERTIES OF IONIC COMPOUNDS ARE :

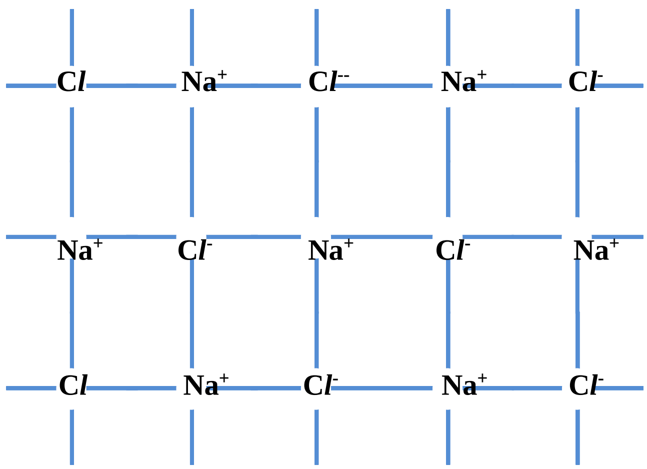
1.**Ionic compounds consist of ions**. All ionic compound consist of positively and negatively charged ions and not molecules. For example, sodium chloride consists of Na^+ and Cl^- ions, magnesium fluorine consists of Mg^{2+} and F^- ions and so on.

2.**Ionic compounds are solids**. In general, in ionic compounds, the ions are held together by strong electrostatic forces of attraction. Hence, the ionic compounds are solids and relatively hard in which in which the ions have regular close packed structure. These compounds are generally brittle and break into pieces when subjected to pressure or stress.

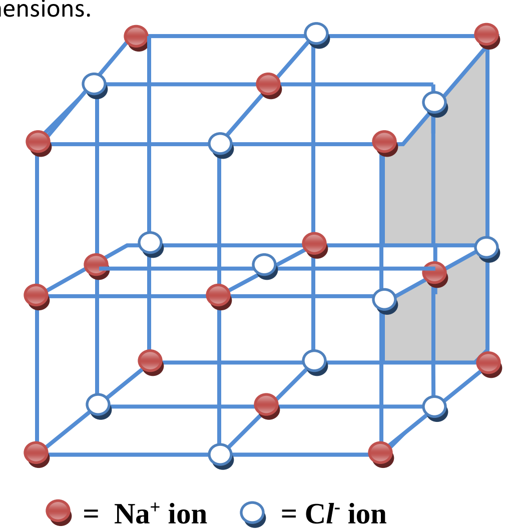
3. **The ionic compounds, in general, are crystalline in nature.** X- ray studies have shown that ionic compounds do not exist as simple single molecules as Na^+Cl^-. This is due to the fact that the forces of attraction are not restricted to single unit such as Na^+ and Cl^- but due to uniform electric field around an ion, each ion is attracted to a large number of other ions. For example, one Na^+ ion will not attract only one Cl^- ion but it can attract as many negative charges as it can. Similarly, the Cl^- ion will attract several Na^+ ions. As a result, there is a regular arrangement of these ions in three dimensions as shown in fig. Such a regular arrangement is called **lattice.**

Example :(a) ***Formation of regular arrangement of sodium chloride.***

As can be seen from the fig, the oppositely charged ions are arranged in a regular manner. It has been seen that the arrangement of sodium chloride lattice is such that each Na^+ ion is surrounded by six Cl^- ions and each Cl^- ion is surrounded by six Na^+ ions . This compound has equal number of oppositely charged ions, so it will have no charge. But what is its formula $Na^+ Cl^-$ or $(Na^+)_{100} (Cl^-)_{100}$ or $(Na^+)_n (Cl^-)_n$? In fact, sodium chloride has an extended lattice in three dimensions.



(a) Formation of regular arrangement of sodium chloride.



(b) Sodium chloride structure .

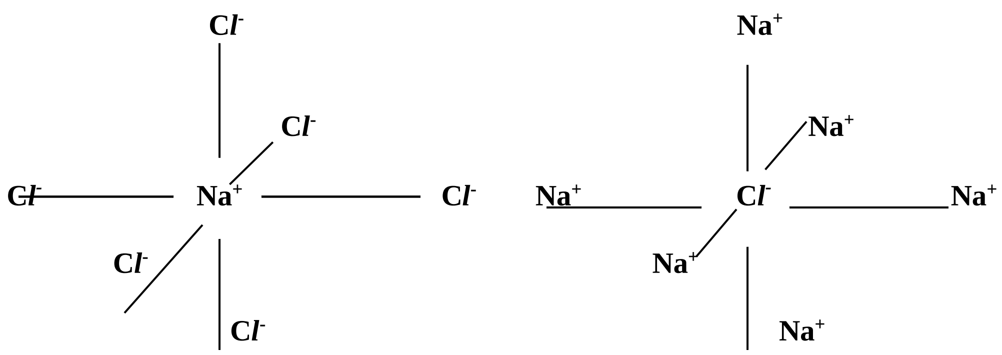
4. **Ionic compounds have high melting and boiling points.** The compounds have high melting and boiling points. This 15
is because of strong electrostatic forces of attraction between the oppositely charged ions. Therefore, large amount of energy is needed to break these bonds and hence melting and boiling points are high. Their densities are also high.

The melting and boiling points of some common ionic compounds are given below :

Ionic Compound	Melting point (K)	Boiling point (K)
LiC*l*	878	1570
NaC*l*	1074	1738
KBr	1007	1708
KI	953	1600
CaO	2845	3123
$MgCl_2$	987	1685
$CaCl_2$	1055	1870

5. **Ionic compounds are generally soluble in water and other polar solvents.** Ionic compounds are generally soluble in polar solvents like water and insoluble in non polar solvents like benzene, ether, alcohol, etc.

6. Ionic compounds conduct electricity when dissolved in water or melted. Ionic compounds do not conduct electricity in the solid state. This is because they have very rigid structures. But when these are melted or dissolved in water, the ions become free and therefore; they conduct electricity due to mobility of ions. Thus, the ionic compounds are good conductors of electricity in the molten or dissolved state.



Arrangement around Na^+ and Cl^- ions. Each Na^+ is surrounded by six Cl^- ions and each Cl^- ion is surrounded by six

◙ An aqueous solution of sodium chloride conducts electricity. Similarly, it can be observed that aqueous solution of other ionic compounds such as magnesium chloride, calcium carbonate, potassium iodide, copper sulphate, etc. conduct electricity.

Occurrence of Metals

◙ Earth is a main source of metals.
◙ Sea water also contains some soluble salts of metals like sodium chloride, magnesium chloride, etc.
◙ All metals are present in the earth's crust either in the free state or in the form of their compounds.
◙ Aluminium is the most abundant metal in the earth's crust. The second most abundant metal is iron and third one is calcium.

Native and Combined States of Metals

Metals occur in the crust of earth in two states : *native state and combined state.*

◙. Native state or free state

A metal is said to occur in a free or a native state when it is found in the crust of the earth in the elementary or uncombined form.

- The metals which are very uncreative are found in the free state.
- These have no tendency to react with oxygen and not attacked by moisture, carbon dioxide of air or other non metals. Silver copper, gold and platinum are some examples of such metals.

◙ Relative abundance of metals in earth

Metal	% by weight
Aluminum	7.5
Iron	4.7
Calcium	3.4
Sodium	2.6
Potassium	2.5
Magnesium	1.9
Titanium	0.6

................All other metals are present in smaller amounts.

◙ The most abundant element in the earth's crust is oxygen. The second most abundant element is silicon and aluminium is the third most abundant element. It is the most abundant metal in the earth's crust because oxygen and silicon are non metals.

◙. Combined state

A metal is said to occur in a combined state if it is found in nature in the form of its compounds.

The metal which have a tendency to react with moisture, oxygen, sulphur, halogens, etc. occur in the crust of the earth in the form of their compounds such as: oxides, sulphides, halides, silicates, carbonates, nitrates, phosphates, etc. For example, sodium, potassium, calcium, aluminium, magnesium, etc. are very reactive metals and therefore, these are never found in the free state. The metals in the middle of the activity series are moderately reactive.

They are formed in the earth crust mainly as oxides, sulphides or carbonates. In fact, most of the metals are found in the combined form in the earth's crust. You will find that that the ores of many metals are oxides. This is because oxygen is very reactive element and is very abundant on the earth.
Copper and silver are metals which occur in the free state as well as in the combined state.

◘ MINERALS AND ORES

The natural substances in which metals or their compounds occur either in native state or combined state are called minerals

Ex: Bauxite and clay.

◙ The minerals are not pure and contain different types of other impurities. The impurities associated with minerals are collectively known as **gangue** or **matrix**.

◙ The metals are extracted from their minerals. But metals cannot be extracted from all the minerals conveniently and profitably.

The mineral from which the metal can be conveniently and profitably extracted, is called an ore.

For example, aluminum occurs in the earth's crust in the form of two minerals, bauxite and clay. Out of these two, aluminium can be conveniently and profitably extracted from clay by some easy and cheap method. Therefore, **the ore of aluminium is bauxite.** Similarly, the minerals of copper are copper glance, cuprite and copper pyrites. But copper can be conveniently extracted from copper pyrites. Therefore, **ore of copper is copper pyrites.**

◙ *All ores are minerals but all minerals are not ores*.

◘Types of Ores

The most common ores of metals are *oxides, sulphides, carbonates, sulphates, halides*, etc. In general, very uncreative metals occur in elemental form or free state.

Ores of some metals.

Nature of ore	Metal		Composition
◙Oxide ores	Aluminium	Bauxite	$Al_2O_3.2H_2O$
	Copper	Cuprite	Cu_2O
	Iron	Magnetite	Fe_3O_4
		Haematite	Fe_2O_3
◙ Sulphide ores	Copper	Copper pyrites	$CuFeS_2$
		Copper glance	Cu_2S
	Zinc	Zinc blende	ZnS
	Lead	Galena	PbS
	Mercury	Cinnabar	HgS
◙ Carbonate ores	Calcium	Limestone	$CaCO_3$
	Zinc	Calamine	$ZnCO_3$
◙ Halide ores	Sodium	Rock salt	$NaCl$
	Magnesium	Carnallite	$KCl\ MgCl_2.6H_2O$
	Calcium	Fluorspar	CaF_2
	Silver	Horn silver	$AgCl$
◙ Sulphate ores	Calcium	Gypsum	$CaSO_4.2H_2O$
	Magnesium	Epsom salt	$MgSO_4.7H_2O$
	Barium	Barytes	$BaSO_4$
	Lead	Anglesite	$PbSO_4$

◙ **Metals which are only slightly reactive occur as sulphides.**

◙ **Reactive metals occur as oxides.**

◙ **Most reactive metals occur as salts as carbonates, sulphates, halides etc.**

METALLURGY

"The process of extracting pure metals from their ores and then refining them use is called metallurgy".

The process of metallurgy involves extraction of metals from their ores and then refining them for use. The ores generally, contain unwanted impurities such as sand, stone, earthy particles, limestone, mica, etc. These are called gangue or matrix. The process of metallurgy depends upon the nature of the ore, nature of the metal and the types of impurities present. Therefore, there is not a single method for the extraction of all metals. However, most of the metals can be extracted by a general procedure which involves the following steps :

Various Steps involved in Metallurgical Processes

Important steps are involved in the extraction of metals from their ores :

1. *Crusting and grinding of the ore.*
2. *Concentration of the ore or enrichment of the ore.*
3. *Extraction of metal from the impure metal.*
4. *Refining or purification of the impure metal.*

Crushing and Grinding of the ore.

Most of the ores in nature occur as big rocks. They are broken to small pieces with the help of crushers. These pieces are then reduced to fine powder with the help of crushers. These pieces are then reduced to fine powder with the help of a ball mill or a stamp mill.

Concentration of Ore or Enrichment of ore.

The ores are usually found mixed up with a large number of nonmetallic impurities of sand and rocky materials known as gangue and matrix. These unwanted impurities have to removed before extracting the metals.

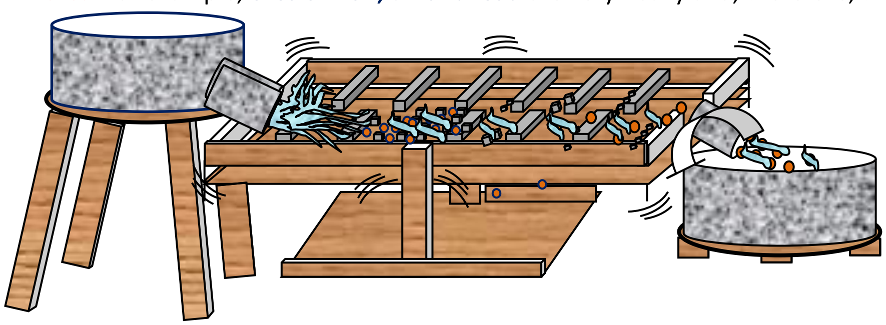
The process of removal of unwanted impurities from the ore is called ore concentration or ore enrichment.

Before the ore is subjected to metallurgical processes for the extraction of metal from the ore, it is essential to concentrate the ore. The finely ground ore is concentrated by any of the following processes depending **upon the nature of the ore and impurities present in it.**

(a) Washing with water - Hydraulic washing

This method *is* ***based upon the difference in the densities of the ore particles and the impurities***.

The gangue particles are generally lighter as compared to ore particles. The crushed and powdered ore is taken in large wooden tables with small obstacles. A stream of water is passed over the shaking table. The lighter impurities are washed away with the running stream of water while the heavier ore particles are left behind. This method of concentration is usually applicable to oxide ores. For example, **ores of iron, tin and lead** are very heavy and, therefore, they are concentrated by this method.

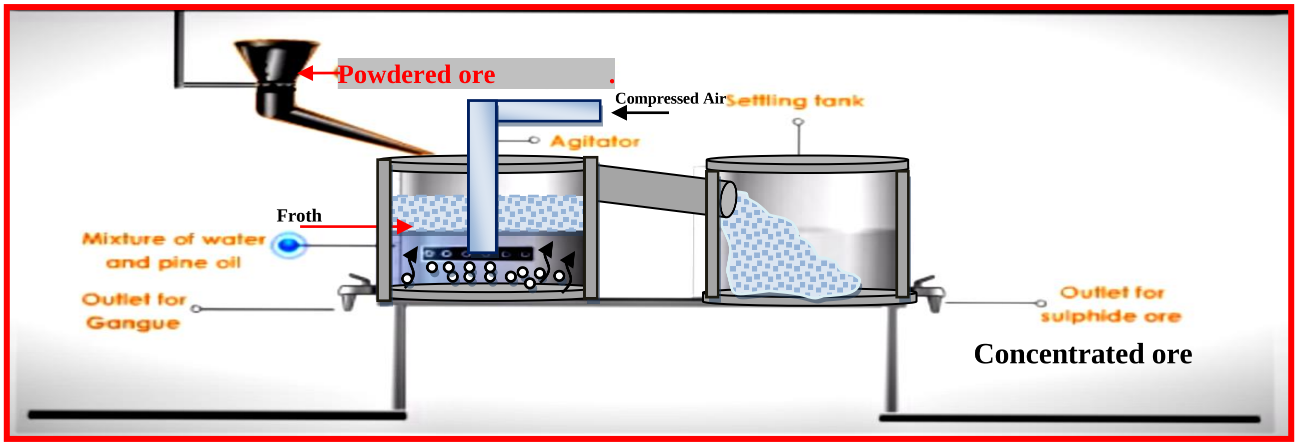


(b) Froth floatation process

This method is based on the principle of difference in the wetting properties of the ore and gangue particles with water and oil.

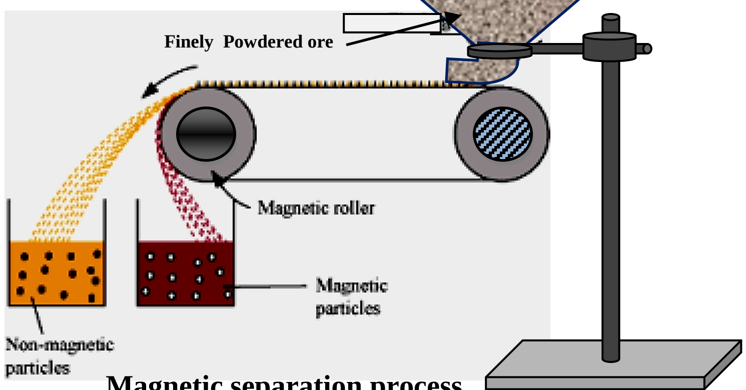
It is used for extraction of those metals in which the ore particles are preferentially wetted by oil and gangue by water. For example, **this method is commonly used for sulphide ores.** In this method, the powdered ore is mixed with water in a large tank to form a slurry. Then some oil is added to it. The sulphide ores are preferentially wetted by the pine oil while the gangue particles are wetted by water. The water is agitated by blowing air violently, when a froth is formed. The froth carries the lighter ore particles along with it to the surface. The heavier impurities are left behind in water and these settle to the bottom. Since the ore particles float with the froth at the surface, this process is called froth floatation process. The froth at the surface is transferred into another tank. The forth is broken by adding some acid and ore particles are separated by filtration and dried.

For example, the froth floatation process is commonly used for the sulphides ores of copper, zinc, lead, etc.



(c) Magnetic separation

This method depends upon the difference in the magnetic properties of the ores and gangue. The ores which are attracted by a magnet can be separated from the non-magnetic impurities with the help of magnetic separation method. For example, this method is used for the concentration of **Haematite, an ore of iron**. It consists of a leather belt moving over two rollers, one of which is magnetic in nature. The powdered ore is dropped over the magnetic portion of the ore is attracted by the magnetic roller and falls nearer to the roller while the non-magnetic impurities fall farther off.



Magnetic separation process.

(d) Chemical methods

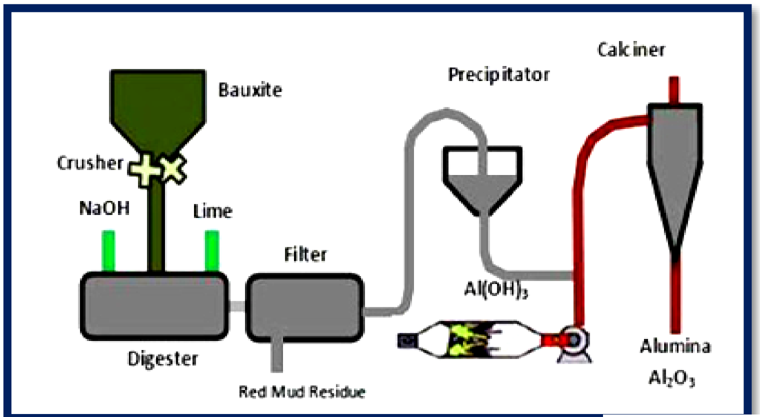
The chemical method used for concentration of the ore is called **leaching.**

- ***It is based on the difference in some chemical property of the metal and the impurities.***

In this method, the powdered ore is treated with certain chemical reagents in which the ore is soluble but the impurities are not soluble. The impurities left undissolved are removed by filtration.

For example, **bauxite** ore is aluminium is concentrated by this method. Bauxite, $Al_2O_3.2H_2O$ is an impure form of aluminium oxide. The main impurities present in it are (III) oxide and sand. The iron (III) oxide gives it a brown rd colour. Baeyer's method is used to obtain pure aluminium oxide from bauxite ore.

- ***This process of chemical separation of aluminium by chemical method*** is known s ***Baeyer's process***.
- ***Baeyer's process*** involves the following steps :



(i) The finely powdered ore is treated with hot sodium hydroxide solution. Sodium hydroxide reacts with aluminium oxide present in bauxite ore to form sodium aluminate which is soluble in water.

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

Aluminium oxide — **Sodium aluminate**

Iron (III) oxide present in bauxite ore does not dissolve in sodium hydroxide solution. It is, therefore, separated by filtration. Silica present in bauxite ore reacts with sodium hydroxide to form water soluble sodium silicate.

(ii) The filtrate is then stirred with small amount of freshly prepared aluminium hydroxide.

The Aluminium hydroxide is added to induce the precipitation of Aluminium hydroxide. It acts as a seeding agent and helps in quick precipitation. Under these conditions, sodium Aluminate gets hydrolysed to form a precipitate of aluminium hydroxide, $Al(OH)_3$.

$$NaAlO_2(aq) + 2H_2O(l) \longrightarrow Al(OH)_3(s) + NaOH(aq)$$

Sodium aluminate — **Aluminium hydroxide** — the impurity of silica remains dissolved as sodium silicate in the solution.

(iii) The precipitate of aluminium hydroxide, thus, formed is separated by filtration. It is dried heated to get pure aluminium oxide. Aluminium oxide is also called *alumina*.

$$2Al(OH)_3 \xrightarrow{Heat} Al_2O_3 + 3H_2O$$

Aluminium hydroxide — **Aluminium oxide**

3. Extraction of the Metal from the Concentrated Ore

The metal is extracted from the concentrated ore by the following steps :

(a) Conversion of the concentrated ore into its oxide. The production of metal from the concentrated ore mainly involves **reduction process.** But it is easier to reduce metal oxide than metal **sulphides and carbonates**. Therefore, before reduction can be done; the metal sulphides or carbonates must be converted into metal oxides. This can be usually done by two processes as **calcinations and roasting** process.

The method depends upon the nature of the ore.

Ex:A carbonate ore is converted into oxide by calcinations while a sulphide ore is converted into oxide by roasting

(b) Conversion of oxide to metal be reduction process.

(a) Conversion of Concentrated Ore into Metal Oxide

Calcination: It is the process of heating the concentrated ore in the absence or air.

The calcination process is used for the following changes :

- (i) to convert carbonates ores into metal oxide.
- (ii) to remove water from the hydrates ores.
- (iii) to remove volatile impurities from the ore.

For example,

(i) Zinc occurs as zinc carbonate in calamine ore. The ore is calcined i.e., heated strongly in the absence of air to convert it to zinc oxide. During calcination, carbon dioxide is expelled.

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

Zinc carbonate

Similarly, in case of carbonate ore of iron, siderite, $FeCO_3$, calcination converts the carbonate to oxide.

$$FeCO_3(s) \xrightarrow{Calcination} FeO(s) + CO_2(g)$$

Siderite — **Iron (II) oxide**

Roasting : *It* is the process of heating the concentrated or strongly in the presence of excess air.

This process is used for converting sulphide ores to metal oxide. In this process, the following changes take place :

- (i) the sulphide ores undergo oxidation to their oxides.
- (ii) moisture is removed
- (iii) volatile impurities are removed

For example,

(i) Zinc occurs as sulphide in zinc blende. It is strongly heated in excess of air when it forms zinc oxide and sulphur dioxide gas is expelled.

$$2ZnS + 3O_2 \xrightarrow{Roasting} 2ZnO(s) + 2SO_2(g)$$

Zinc sulphate — **(From air)** — **Zinc oxide**

(ii) Iron occurs as sulphide in iron pyrites ore. During roasting it gets oxidized to ferric oxide.

$$4FeS_2 (s) + 11O_2 (g) \xrightarrow{Roasting} 2Fe_2O_3 (s) + 8 SO_2 (g)$$

Iron pyrites — Ferric oxide

Moreover, any ferrous compound present in the ore is oxidized to ferric .

$$4FeO(s) + 3O_2 (g) \longrightarrow 2Fe_2O_3 (s)$$

Ferrous oxide — Ferric oxide

If the ferrous oxide is not converted to ferric oxide, it would combine with impurities present and form slag.

Differences between Calcination and Roasting

Both calcination and roasting processes convert concentrated ore to metal oxide. The two processes differ in the respects as:

Calcination	Roasting
1. In calcinations, the ore is heated in the absence of air.	1. In roasting, the ore is heated in the presence of air.
2. It is used for carbonate and oxide ores. $ZnCO_3 \longrightarrow ZnO + CO_2$	2. It is generally used for sulphide ores. $2ZnS(s) + 3O_2 (g) \longrightarrow 2ZnO(s) + 2SO_2 (g)$
3. Moisture and organic impurities are removed.	3. Volatile impurities are removed as oxides SO_2, P_2O_5, As_2O_3, etc.

(b) Conversion of Metal Oxide into Metal

The metal oxide formed after calcination or roasting is converted into metal by reduction. The method used for reduction of metal oxide depends upon the nature and chemical reactivity of metal.

The metals can be grouped into the following three categories on the basis of their reactivity :

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

1. Extracting metals low in activity series : Reduction by Heating in Air :

Metals low in the reactivity series are very unreactive. They can be obtained from their oxides by simply heating in air. For example, mercury is obtained from its cinnabar ore by this method. The method involves the steps :

(i) The concentrated mercuric sulphide is roasting in air when mercuric oxide is formed.

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

Mercuric oxide — Mercury metal

Similarly, copper which is found as copper sulphide can be obtained from its ore by heating in air alone.

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

$$2Cu_2 O + Cu_2 S \xrightarrow{Heat} 6Cu (s) + SO_2 (g)$$

II. Extracting Metals in the middle of the activity series :

Chemical Reduction

The metals in the middle of the reactivity series, such as iron, zinc, lead, copper, etc. are moderately reactive. These are usually present as sulphides or carbonates. Therefore, before reduction the metal sulphides are carbonates must be converted to oxides. This is done by roasting and calcinations as described earlier. The oxides of these metals cannot be reduced by heating alone. Therefore, these metal oxides are reduced to free metal by using chemical agents like carbon, aluminium, sodium or calcium.

(i) Reduction with carbon

The oxides of moderately reactive metals like zinc, copper, nickel, tin, lead, etc. can be reduced by using carbon as reducing agent. In this process, the metal oxide is mixed with coke and heated in a furnace. Carbon reduces the metal oxide to free metal. For example, when zinc oxide is heated with carbon, zinc metal is produced.

$$ZnO (s) + C \xrightarrow{Heat} Zn (s) + CO (g)$$

Zinc oxide — Carbon — Zinc metal — Carbon monoxide

Similarly, iron and lead are obtained from their oxides by heating with carbon.

$$Fe_2O_3 (s) + 3C (s) \longrightarrow 2Fe (s) + 3CO (g)$$

Ferric oxide — Carbon monoxide

Carbon monoxide further reacts with ferric oxide to form iron :

$$Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$$

$$\underset{\text{Lead oxide}}{PbO(s)} + C \longrightarrow \underset{\text{Lead metal}}{Pb} + CO$$

Coke is very commonly used as a reducing agent because it is cheap. It is used in the reduction of oxides of copper, iron, tin etc. However, coke cannot be used for the reduction of oxides of more reactive metals like sodium, potassium, calcium, magnesium, Aluminium, Manganese, etc.

◙ ***One disadvantage of using carbon as reducing agent is that small traces of carbon are added to metal as impurity. Therefore, it contaminates the metals.***

(ii) Reduction with carbon monoxide

Metals can be obtained from oxides by reduction with carbon monoxide in the furnace. For example, iron is obtained from ferric oxide by heating with carbon monoxide.

$$Fe_2O_3(s) + 3CO(g) \xrightarrow{\text{Heat}} 2Fe(s) + 3CO_2(g)$$

(iii) Reduction with Aluminum

Certain metal oxide are reduced by aluminium to metals. **This method is known as Aluminothermy or thermite process.** For example, chromium, manganese, titanium, vanadium metals are obtained by the reduction of their oxides with aluminium powder For example, manganese dioxide is heated with aluminium powder. For example, manganese dioxide is heated with aluminium powder and manganese is obtained

$$\underset{\text{Manganese dioxide}}{3MnO_2(s)} + 4Al(s) \xrightarrow{\text{Heat}} \underset{\text{Manganese}}{3Mn(s)} + 2Al_2O_3$$

Similarly, chromium is obtained by heating chromium oxide with aluminium powder.

$$\underset{\text{Chromium oxide}}{Cr_2O_3} + 2Al \xrightarrow{\text{Heat}} \underset{\text{Chromium}}{2Cr} + Al_2O_3$$

◙ Aluminium is an expensive metal and, therefore, it is not used to reduce metals which are less expensive than aluminium.

III. Extraction of metals high up in activity series : Reduction by electrolysis of electrolytic reduction

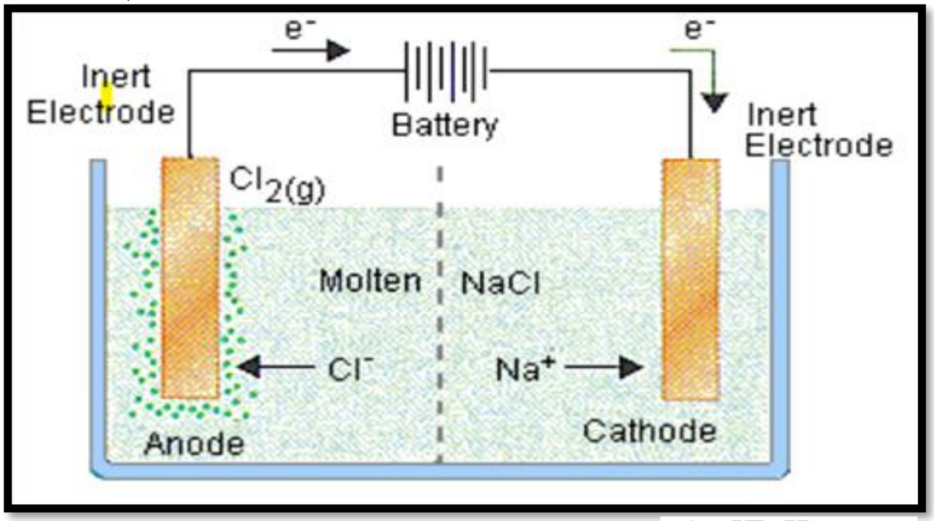
The oxides of active metals are very stable and cannot be reduced by carbon or aluminium. For example, carbon cannot reduce the oxides of sodium, magnesium, calcium, aluminium, etc. to the respective metal. This is because these metals have more affinity for oxygen than carbon. Therefore, carbon cannot eliminate oxygen from these metal oxides. These metals are commonly extracted by the electrolysis of their fused salts using suitable electrodes. This is also called **electrolytic reduction** meaning reduction by electrolysis. ***The process of extraction of metals by electrolysis process*** is called **electrometallurgy.**

For example, the reactive metals like sodium or potassium are prepared by the electrolytic reduction of their molten metal chlorides. The metals are deposited at cathode whereas chlorine gas is liberated at anode. The chemical reactions are :

At cathode : $Na^+ + e^- \longrightarrow Na$

At anode : $2Cl^- \longrightarrow Cl_2 + 2e^-$

or $2NaCl \xrightarrow{\text{Electrolysis}} 2Na + Cl_2$



Similarly, aluminium is obtained by electrolytic reduction of aluminium oxide. Aluminium oxide is very stable and aluminium cannot be prepared by reduction with carbon. It is prepared by the electrolysis of molten alumina. In this process, pure alumina is dissolved in molten cryolite in an iron tank linked with carbon. During electrolysis, the aluminium ions, Al^{3+} are reduced at cathode to form aluminium.

$$\underset{\text{Aluminium ion}}{Al^{3+}} + \underset{\text{Electrons}}{3e^-} \xrightarrow{\text{Heat}} \underset{\text{Aluminium}}{Al}$$

◙ ***During electrolysis reduction of molten salts, the metals are always obtained at the cathode.***

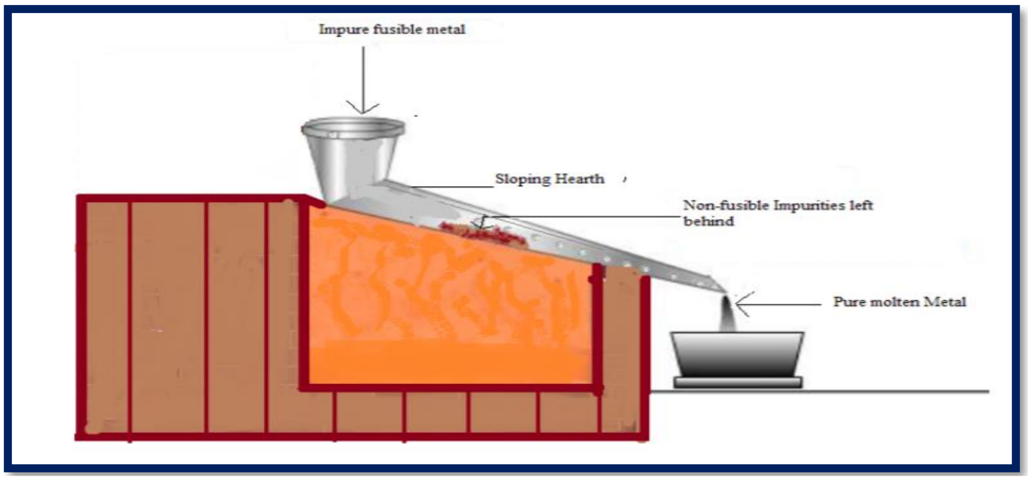
4. Purification of Refining of Metals

The metal obtained by any of the above methods is usually impure and is known as **crude** metal. *The process of purifying the crude metal is called* **refining. The method of refining depends upon the nature of the metal and the impurities present**. Some of the common methods of refining are:

(a) Liquation

This method is used for refining the metals having low melting points, such as tin, lead, bismuth, etc. ***This is based on the principle that the metal to be refined is easily fusible but the impurities do not fuse easily.***

In this method, the impure metal is placed on the sloping hearth of the furnace and is gently heated in an inert atmosphere of carbon monoxide. The hearth is maintained at a temperature slightly above the melting point of the metal. The metal melts and flows down to the bottom of sloping hearth. The solid impurities, whose melting point is higher than the melting point of the metal are left behind on the hearth. The pure metal is collected at the bottom of the sloping hearth in a receiver.



Liquation separation is based on differences in the melting temperatures and densities of alloy constituents and on the low level of mutual solubility of the constituents. For example, when molten crude lead is cooled, copper crystals (dross) separate out at established temperatures and, because of their low density, float to the surface and can be removed. This method is used to remove Cu, Ag, Au, and Bi from crude lead, to remove Fe, Cu, and Pb from crude zinc, and to refine tin and other metals.

(b) Distillation

This method is used for the purification of volatile metals such as mercury and zinc. In this method, the impure metal is heated strongly in a vessel. The pure metal distils over and its vapours are condensed separately in a receiver to get pure metal. The non volatile impurities are left behind in the retort.

(c) Oxidation method (oxidative refining)

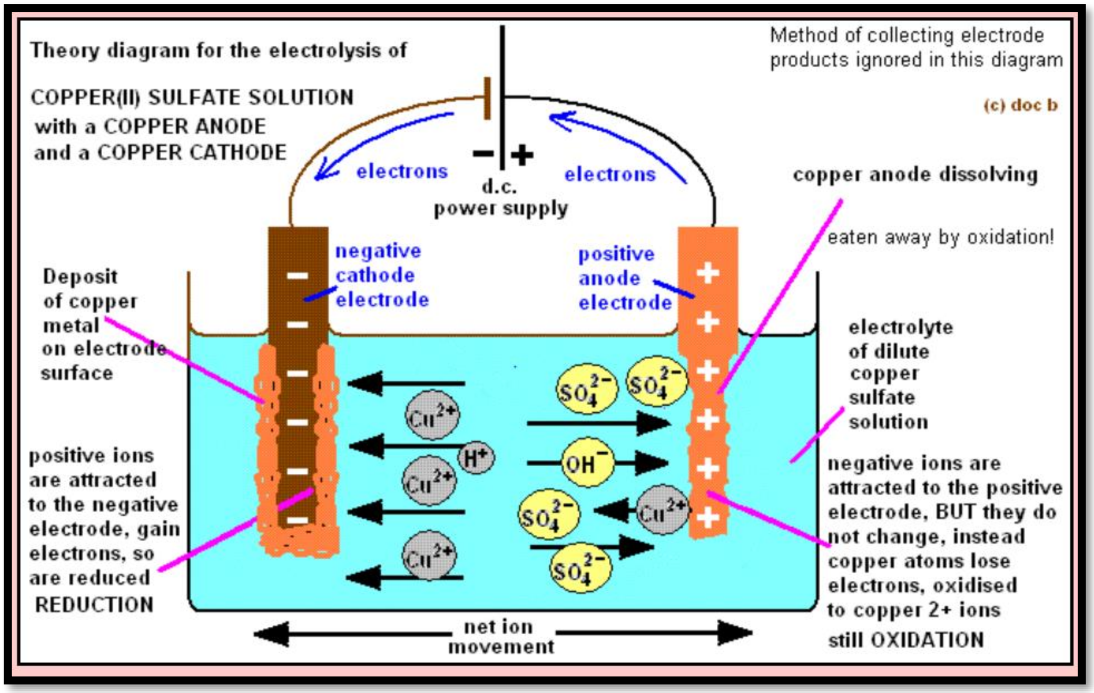
This method is used for the refining of metals in those cases in which the impurities have greater tendency to get oxidized than metal itself. In this method, air is passed through the impure molten metal. The oxygen of the air oxidizes the impurities to their oxides which are then removed. The pure metal is left behind. For example, impure iron is refined by oxidative refining method. Pig iron contains carbon, sulphur , phosphorous, silicon and manganese as impurities. When a blast of air is blown over molten pig iron these impurities are oxidized to their oxides and get removed.

The pure iron is left behind. Similarly, silver is refined by this method. The impure silver metal is fused in small boat shaped dishes made of bone ash called *cupels.* The cupels are heated in a suitable furnace by a blast of air blown over them. The lead is easily oxidized to lead monoxide and is carried away by the blast, while pure silver is left behind.

This method employed to purify silver containing lead as an impurity. The impure silver is heated in a shallow vessel made of bone-ash under a blast of air. The lead is easily oxidized to powdery lead monoxide. Most of it is carried away by the blast of air. The rest melts and is absorbed by the bone ash cupel. Pure silver is left behind. Silver itself is not oxidized under these condition

(d) Electrolysis refining

This is most general widely used method for the refining of impure metals. Many metals such as copper, zinc tin, nickel, silver, gold, etc. are refined electrolytically. **It is based upon the phenomenon of electrolysis**. In this method, the crude metal is cast into thick rods and are made as anodes, while the thin sheets of pure metal are made as cathodes. An aqueous solution of some salt of the metal is used as an electrolyte. On passing current through the electrolyte, the pure metal from the anode dissolves into the electrolyte . An equivalent amount of pure metal from the electrolyte is deposited on the cathode. The soluble impurities go in the solution whereas the insoluble impurities settle down at the bottom of the anode and are known as **anode mud.** In this way, the pure metal from anode goes into electrolyte and from electrolyte it goes to the cathode. Let us consider the example of refining of copper.



Electrolytic refining of copper

In the electrolysis refining of copper crude copper is made anode, a thin sheet of pure copper is made cathode. The electrolyte is a solution of copper sulpahte containing a small amount of dilute sulphuric acid. On passing the electric current copper dissolves from the anode into the electrolyte. An equivalent amount of copper from the electrolyte are deposited at the cathode in the form of pure metal.

<u>The reactions occur at the electrodes</u> :

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

At anode : $\underset{\text{Copper}}{Cu} - 2e^- \xrightarrow{\text{Oxidation}} \underset{\text{Copper ion}}{Cu^{2+}}$

Copper sulphate solution contains copper, Cu^{2+} and sulphate, SO^{2-}_4 ions. The positively charged copper go to the cathode and get reduced to copper by accepting the electrons.

At cathode : $\underset{\text{Copper ion}}{Cu^{2+}} + 2e^- \xrightarrow{\text{Reduction}} \underset{\text{Copper}}{Cu}$ The copper atoms get deposited on the cathode giving pure copper.

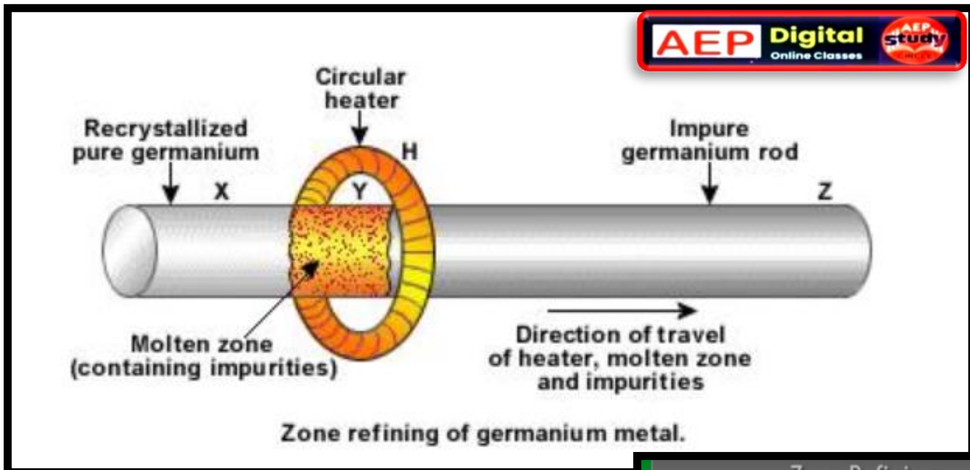
As the process continues, impure anode goes on dissolving in the solution and becomes thinner and thinner. At the same time, pure Copper gets deposited on the cathode. The impurities present in the crude copper either go into the solution or remain there. The less reactive metals like gold, silver present in the impure copper, collect at the bottom of the cell below the anode. This is called **anode mud.** Gold and silver metals if present in the impure metal can be recovered from the anode mud.

◙ The electrolytic refining of metals have two purposes :
(i) It refines or purifies the metal. (ii) It helps to recover some valuable metals if present as impurities in the crude metal.

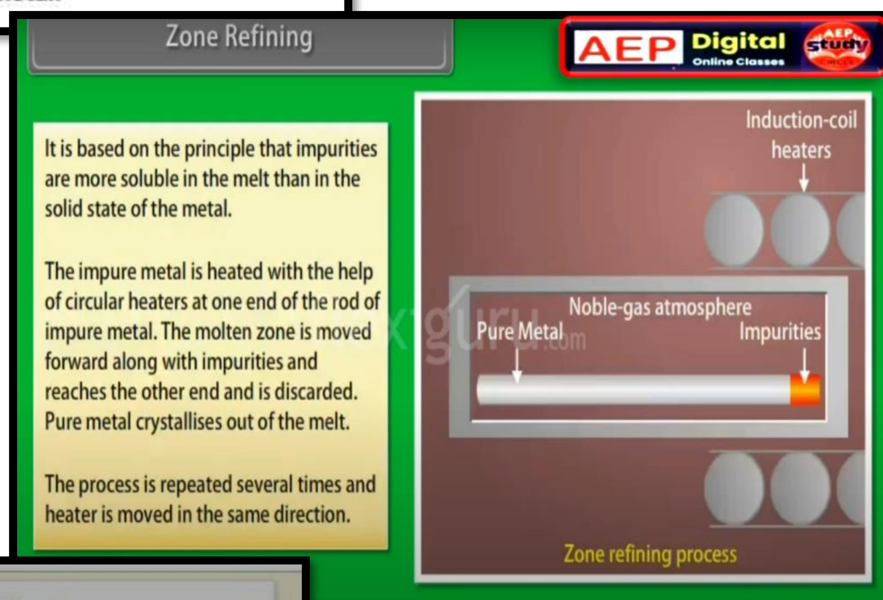
◘Methods for Obtaining Metals of very High Purity

These days, we require some metals in their purest form for certain specific applications. For example, extremely pure silicon and germanium are needed for semi conductors. Similarly, uranium to be used as fuel in nuclear reactors should be ultra pure. It has been calculated that uranium should not contain more than 1 part per million of carbon as impurity. The following two methods are used to get ultra pure metals :

(a) Zone refining [These methods are discussed in higher classes].



Zone refining of germanium metal.

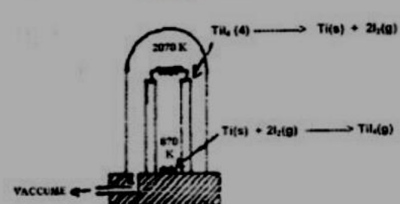


(b) Van Arkel method

Ultra pure metals are being prepared by the ***Van Arkel Method***. Crude metal is heated with a suitable substance so that the pure metal present in it may be converted into stable volatile compound leaving behind impurities. The compound so formed is then decomposed by heating to get the pure metal. Van Arkels method is used to purify crude titanium metal. It is heated with iodine to about 500K to form volatile compound. TiI_4 leaving behind the impurities .TiI_4 is further heated to 1700K when it decomposes to give pure titanium.

$$\underset{(Impure)}{Ti} + 2I_2 \xrightarrow{500K} TiI_4$$

$$TiI_4 \xrightarrow{1700K} \underset{(Pure\ Metal)}{Ti} + 2I_2$$



Steps of Extraction of Metals

Flow Chart: Extraction of Metals

Ore (Concentration of Ores)

- Metals of high reactivity → Electrolysis of molten ore → Pure metal
- Metals of medium reactivity
  - Carbonate ore → Calcination
  - Sulphide ore → Roasting
  - → oxide of metal → Reduction to metal → Purification
- Metals of least reactivity → Sulphide ore → Roasting → Metal → Purification

Elements

Metals

Physical Properties of Metals	Chemical Properties of Metals
Malleable	Highly Electropositive
Ductile	Metal+Water ⇩ Metal oxide+Hydrogen
Good conductor of Heat and Electricity	Metal+ Air ⇩ Metal oxides
Sonorous	Metal+Acid ⇩ Salt +Hydrogen
Generally solids	Reactive metals displace less reactive metals from their salt

Non-metals

Physical Properties of Non-metals	Chemical Properties of Non-metals
Non-Malleable	Generally Electronegative
Non-Ductile	Do not react with water
Poor conductor of Heat and Electricity	Form Ionic compounds with metals
Can be Solids, Liquids or Gases	

Ionic Compound

- **Metal:** Metal ⇩ Metal Ion + Electron
- ✚
- **Non- metal:** Non-metal+electron ⇩ Non metal Ion
- ➡
- **Ionic Compund:** Metal Ion+ Non-metal Ion ⇩ Ionic compound

Properties of Ionic compounds

- Solid But brittle
- High Melting and Boiling point
- High enthalpies of fusion and vaporisation
- They form crystals
- Readily soluble in polar solvents and insoluble in non-polar solvents
- Good conductor of electricty in molten and aqueous state

Concenteration of Ore

- Highly reactive metals → Electrolysis of Ore → Pure metal
- Medium reactive metals
  - Carbonate Ore → Calcination (Absence of air)
  - Sulphide Ore → Roasting (Presence of air)
  - → Oxide of metal → Reduction to metal → Purification of metal
- Low reactive metals → Sulphide Ore → Roasting → Metal → Refining

Corrosion

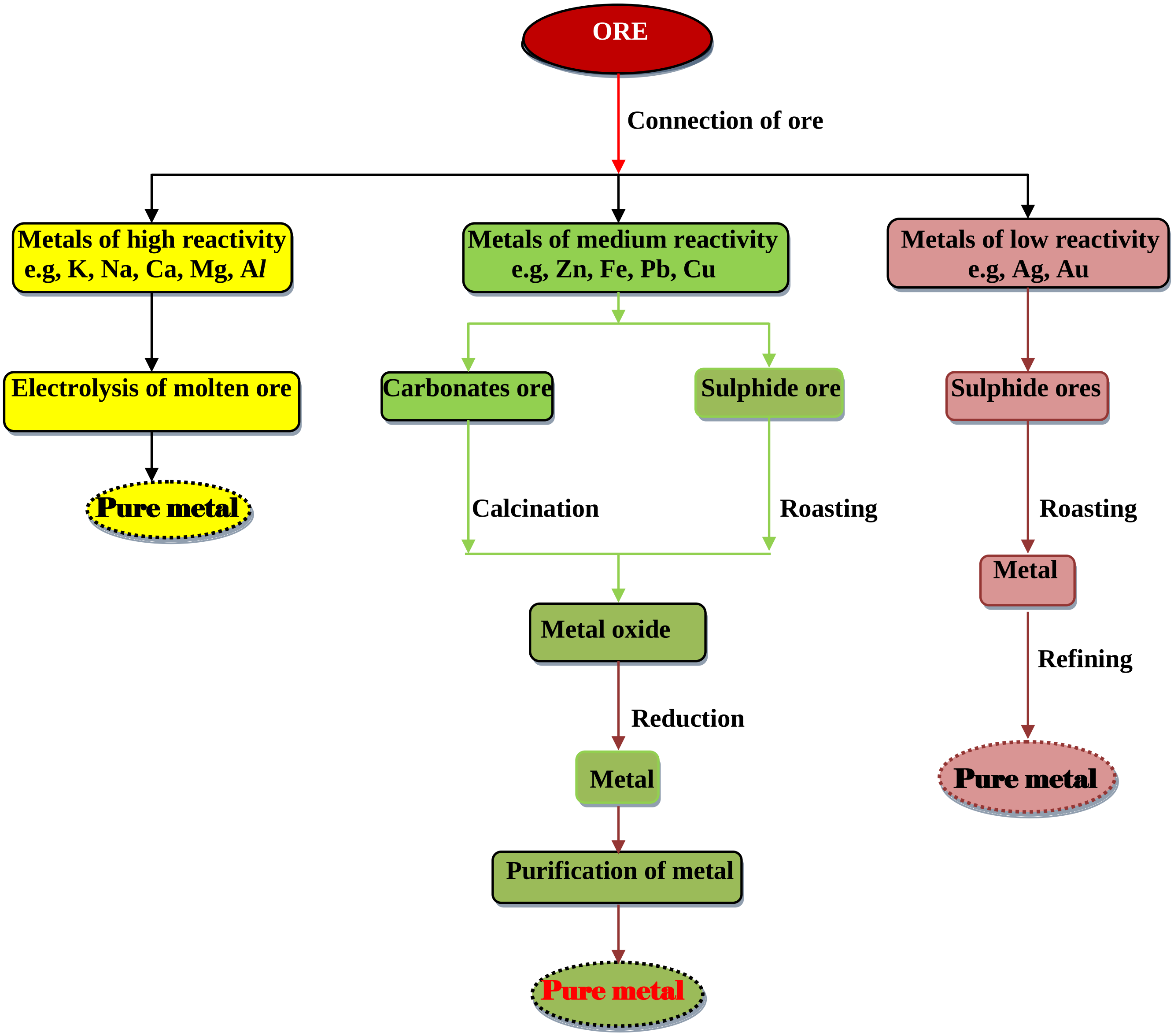
Metals on exposure to water / moisture / Acids tamish due to formation of oxides on their surface

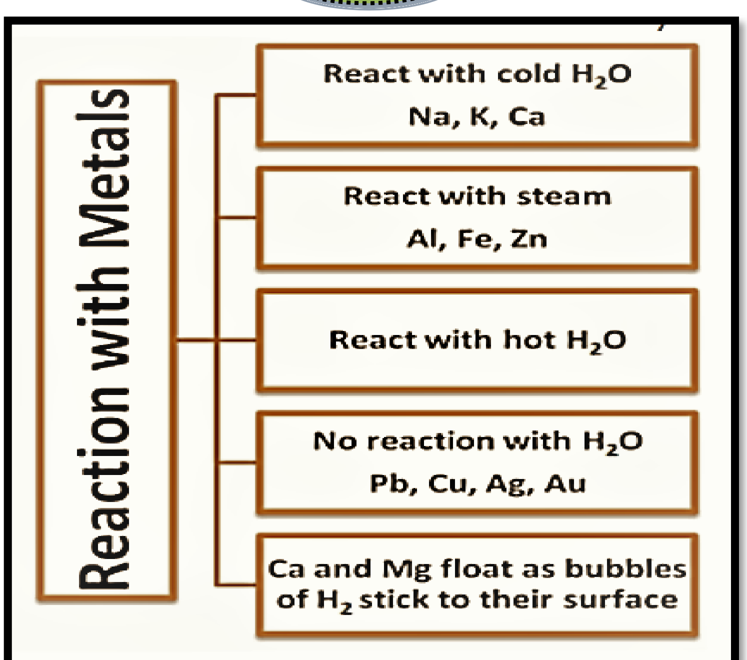
Prevention Of Corrosion

Alloying
Mixing Iron /steel with less reactive metals

Galvanization
Coating Iron / steel with zinc

Electroplating
Coating Iron by electrolytic deposition with chromium , silver or other metals


ORE
Connection of ore
Metals of high reactivity e.g, K, Na, Ca, Mg, Al
Metals of medium reactivity e.g, Zn, Fe, Pb, Cu
Metals of low reactivity e.g, Ag, Au
Electrolysis of molten ore
Pure metal
Carbonates ore
Sulphide ore
Calcination
Roasting
Metal oxide
Reduction
Metal
Purification of metal
Pure metal
Sulphide ores
Roasting
Metal
Refining
Pure metal


Reaction with Metals
React with cold H2O
Na, K, Ca
React with steam
Al, Fe, Zn
React with hot H2O
No reaction with H2O
Pb, Cu, Ag, Au
Ca and Mg float as bubbles of H2 stick to their surface

Corrosion of Metals

Surface of many metals is easily attacked when exposed to atmosphere. They react with air or water present in the environment and form undesirable compounds on their surfaces. These undesirable compounds are generally oxides. This process is called ***corrosion***. Almost all metals except the noble metals such as *gold, platinum and palladium* are attacked by the environment. In other words, almost all metals get corroded. For example,

- (i) When iron is exposed to moisture for a long time, its surface acquires a brown flaky substance called ***rust***.
- (ii)Copper reacts with moist carbon dioxide in the air and slowly loses its shiny brown surface and acquires a green coating of basic copper carbonate in moist air.
- (iii)Silver articles become black after sometime when exposed to air. This is because it reacts with sulphur in the air to form a coating of silver sulphide.
- (iv)Lead or stainless steel lose their luster due to corrosion. Thus,

Corrosion is a process of deterioration of a metal as a result of its reaction with air or water surrounding it.

The corrosion causes damage to buildings, bridges, ships and many other articles especially made of iron.

Rusting of iron

Iron corrodes readily when exposed to moisture and gets covered with a brown flaky substance called rust. This is also called **rusting of iron** . **Chemically, the rust is a hydrate iron (III) oxide, $Fe_2O_3.2H_2O$**

Resulting is an oxidation process in which iron metal is slowly oxidized by the action of air. Therefore, rusting of iron takes place under the following conditions :

- (i) **Presence of air (or oxygen)**
- (ii)**Presence of water (moisture).**

It has been observed that

- (a) The presence of impurities in the metal speed up the rusting process. Therefore, ***pure iron does not rust.***
- (b) The presence of electrolytes in water also speed up the process of rusting. Therefore, ***rusting of iron in water occurs quicker than in distilled in water.***
- (c) The position of metal in the electrochemical series determines the extent of corrosion. **More the reactivity of the metal, the more will be possibility of the metal getting corroded.**

Prevention of rusting

The rusting of iron can be prevented or decreased by:

- (i) Corrosion of metals can be prevented by **coating the metal surface with a thin layer of paint, varnish or grease.** For example, many vehicles such as cycles, motors, cars made from iron sheet are protected from rusting by paints.
- (ii) **Iron is protected from rusting by coating by coating it with a thin layer of another metal which is more reactive than iron.** This prevents the loss electrons in preference in iron. Therefore, the covering of metal is consumed with time but as long it is present on the surface of iron, the iron is not rusted. Zinc is commonly used for covering iron surfaces.

  The process of covering iron with zinc is called **galvanization.**

  Galvanized iron sheets maintain their shine. Iron is also coated with other metals such as tin known as **tin coating.**
- (iii) **The iron pipes which are in contact with water such as ground water pipes are protected from rusting by connecting these with more active metals** under the ground.
- (iv) **To decrease rusting of iron, certain antirust solutions are used.** For example, solutions of alkaline phosphates are used as **antirust solutions**.
- (v) **By alloying.** Some metals when alloyed with other metals become more resistance to corrosion. For example, when iron is alloyed with chromium and nickel form stainless steel. This is resistance to corrosion and does not rust at all.

Alloys and Amalgams

Alloying is a very good method for improving the properties of a metal. We can get the desired properties by this method.

Wonder of Ancient Indian Metallurgy

The iron pillar near Qutab Minar at Delhi was made around 400 BC by iron workers of India. They has developed a process which prevented wrought iron form rusting. The iron pillar is 8m high and has weight about 6000 kg. Is it not a wonder that the pillar still stands in a pristine condition more than 2000 years after it was manufactured.

Alloys and Amalgam

Iron is the most widely used metal. But it is never used in the pure form. This is because pure iron is very soft and stretches easily when hot. But when it is mixed with a small amount of carbon, it becomes hard and strong. The new form of iron is called steel. The strength and other properties of steel depend upon the percentage of carbon in it. Thus, ***an alloy is a homogeneous mixture of two or more metals or a metal and a nonmetal.***

Similarly, when iron is mixed with other metals such as nickel and chromium, we get an alloy called stainless steel. This is strong, hard and does not rust at all. Thus, if iron is mixed with some other substance, its properties are changed. In fact, the properties of any metal can be changed by mixing it with some other metal or non metal i.e., **by forming alloys.** We also come across brass articles. It is an alloy of copper and zinc. An alloy is generally prepared by first melting the main metal and then dissolving the other elements in it in a definite proportion. It is then cooled to room temperature so that the mixture gets solidified. It is then cooled to room temperature so that the mixture gets solidified. The solid mixture is an alloy.

Object of Alloy making

Alloys are generally prepared to have certain specific properties which are not possessed by the constituent metals. The main objects of alloy making are :

- (i) ***To increase resistance to corrosion.*** For example, stainless steel is prepared which has more resistance to corrosion than iron.
- (ii) ***To modify chemical reactivity.*** The chemical reactivity of sodium is increased by making an alloy with mercury which is known as sodium amalgam.
- (iii) ***To increase the hardness.*** Steel, an alloy of iron and carbon is harder than iron.
- (iv)***To increase tensile strength.*** Magnox (**alloy**), whose **name** is an abbreviation for "**magnesium** non-oxidizing", is 99% **magnesium** and 1% **aluminum**, is an alloy of magnesium and aluminium. It has greater tensile strength as compared to magnesium and aluminium.
- (v) ***To produce good casting***. Type metal is an alloy of lead, tin and mercury.
- (vi)***To lower the melting point.*** For example, solder is an alloy of lead and tin. It is having a low melting point and is used for welding electrical wires together.

Amalgam

- These are special class of alloys in which one of the constituent metals is mercury.

Thus, **amalgams *are homogeneous mixture of a metal and mercury***. For example, sodium amalgam contains sodium and mercury. These are formed by treating metals such as sodium, zinc, tin gold, etc. with mercury. Different amalgams are prepared depending upon their uses. For example,

- (i) Sodium amalgam is used to decrease the chemical reactivity of sodium metal. It is also used as a good reducing agent.
- (ii)Tin amalgam is used for silvering cheap mirrors.
- (iii)The process of amalgamation is used for the extraction of metals like gold or silver from their native ores.

Steel

Steel is an alloy of iron and carbon containing 0.1 to 1.5 % carbon. Thus, the constituents of steel are iron and carbon. Pure iron is not very hard and strong. It is brittle and, therefore, it cannot be used for structural purposes. Hence, most of the molten pig iron obtained from blast furnace is converted into steel. Steel is very hard, tough and strong. It is used for making rails, screws, girders, bridges, railway lines, etc. Steel can also used for the construction of buildings, vehicles, ships, etc.

Types of Iron

There are three commercial varieties of iron. They differ in their carbon content.

- 1. Cast iron or pig iron: 2.0 – 4.5 % carbon
- 2. Wrought iron : upto 0.5 % carbon
- 3. Steel : 0.1 – 1.5 % carbon
- Steel comes in between cast iron and wrought iron and shows intermediate properties.
- Wrought iron is the purest form of iron while cast iron is the least pure form of iron.
- Steel is manufactured form cast iron by reducing its carbon content form about 5 % to between 0.1 % to 1.5 % depending upon the quality of steel to be prepared.

Alloy Steels

Ordinary steel is affected by the action of air, acids, alkalies and other chemicals. However, when varying amounts of other elements are added to it, its properties are modified. These are called **alloy steels.** Thus, **steel obtained by the addition of some other elements such as chromium, vanadium, titanium, molybdenum, manganese, cobalt or nickel to carbon steel are called alloy steels.**

- The alloying elements modify the properties of steel. For example, stainless steel is an alloy of iron with chromium and nickel. It is hard and strong and does not rust on exposure to humid air. Since stainless steel resists rusting, it is used for making cooking utensils, knives, scissors, tools, surgical instruments and equipment for food processing industry and dairy industry.

Some common alloy steels and their important features are given in table.

Name	Percentage composition	Properties	Uses
1. Stainless steel	Fe = 73, Cr = 18, Ni = 8	Does not rust	Utensils, cycle and automobile parts, cutlery , razor blades.
2. Nickel steel	Fe = 96 – 98, Ni = 2 – 4	Hard, elastic and rust proof	Cables, automobile and aero plane parts, armor plates, gears.
3. Invar	Fe = 64, Ni = 36	Low expansion on heating	Meter scales, measuring instruments clock pendulums.
4. Chrome steel	Fe = 98, Cr = 1.5 – 2.0	High tensile strength	Cutting tools such as files, cutlery
5. Tungsten steel	Fe = 94, W = 5 and C	Hard, resistance to	High-speed cutting tools, springs.
6. Alnico	Fe = 60, Al = 12, Ni = 20, Co = 5	Strongly magnetic	Permanent powerful magnets.
7. Manganese steel	Fe = 86, Mn = 13 and C	Extremely hard, resistance	Rock crusher s, burglar proof safes, rail road tracks.

Alloys of Aluminium

Aluminium is very light metal but it is not strong. Therefore, it is alloyed with some strong metals to make it strong. The common alloys of aluminium are :

1. Duralumin. It is an alloy containing aluminium, copper and traces of magnesium and manganese. Its percentage composition is

Aluminium = 95 % , Copper = 4 %
Magnesium = 0.5 %, Manganese = 0.5%

It is stronger than pure Aluminium. Since duralumin is light and yet strong, it is used for making bodies of aircrafts, helicopters, jets, kitchenware like pressure cooker. Duralumin is resistance to sea water corrosion and, therefore, it is used for making bodies of ships. It is also known as **Duralumin**.

2 Magnalium. It is an alloy of aluminium and magnesium having the composition :

Aluminium = 95 % Magnesium = 5 % . It is very light and hard. It is more hard than pure aluminium. It is used for making light instruments, balance beams, pressure cookers, etc.

3. Alnico. It is an alloy containing aluminium, iron, nickel and cobalt. It is highly magnetic in nature and can be used for making powerful magnets.

Alloys of Copper

The important alloys of copper and brass and bronze.

1. Brass. It is an alloy of copper and zinc having the composition :

Copper = 80%, Zinc = 20%

Brass is more malleable and more strong than pure copper. It is used for making cooking utensils, condenser sheets, pipes, hardware, nuts, bolts, screws, wire, tubes, scientific instruments, springs, ornaments, etc.

2 Bronze. It is an alloy of copper and tin having the composition :

Copper = 90% Tin = 10%

Bronze is very tough and highly resistance to corrosion. It is used for making utensils, statues, cooling pipes, coins, hardware, etc.

Some other Common Alloys of copper

Gun metal : Cu = 90%
Sn = 10%

Used for making gun barrels, castings, bearings, etc.

Monel METAL is a group of nickel alloys, primarily composed of nickel (from 52 to 67%) and copper, with small amounts of iron, manganese, carbon, and silicon. (Alloys with copper contents 60% or more are called cupronickel.) Used for making bells, gangs, etc.

Constantan : Cu = 60%
Ni = 40%

Used for making electrical apparatus.

German silver :
Cu = 60%
Zn = 20%
Ni = 20% Used for making silverware, utensils and for electroplating.

It may be noted that German silver does not contain silver.

Alloying of Gold

Pure gold is very soft and cannot be used as such for jewellery. Therefore, it is generally alloyed with other metals commonly copper or silver to make it harder and modify its color. The purify of gold is expressed as carats. ***Pure gold is 24 carats.*** A 18 parts of gold in 24 parts by weight of alloy.

Therefore, percentage of gold in 18 carat gold

$$= \frac{18}{24} \times 100$$

$$= 75\ \%.$$