METALS & NON METALS

CONTENTS

INTRODUCTION

There are 115 elements known till today. These elements can be classified into various groups on the basis of their properties. Metals are usually hard, malleable and ductile. They are sonorous and have metallic lustre. They are good conductors of heat and electricity.

Non-metallic are usually soft. They do not possess lustre. They are not malleable and ductile rather they are brittle. They are not good conductors of heat and electricity.

We shall study about metals and non-metals in detail.

IMPORTANT TERMS AND CONCEPTS

1. **Metals :** Those elements which can lose electrons easily and from positive ions. They are mostly solids, possess high density. They have high melting and boiling points. They have metallic lustre and they are sonorous, i.e., produce metallic sound. They are good conductors of heat and electricity. They are usually malleable and ductile, e.g., gold , silver, copper, tin, lead, iron, mercury, cobalt, nickel, aluminium, sodium, potassium are metals.

2. **Hardness :** Most of the metals are hard. If you try to cut them with knife, it will not be possible in most of the metals. Some of metals like sodium, potassium are soft metals and can be cut with knife.

3. **Malleability :** The ability of a metals due to which it can be beaten into sheets is called malleability. Iron, copper, zinc, aluminium, magnesium are available in the form of sheets. Aluminium, steel, copper, brass, bronze are used in making utensils.

   Brass and bronze are also used for making statues. Bronze is used for making medals. Aluminium and silver metals are converted in foils. Aluminium foils are used for packaing whereas silver foils are used in decorating sweets.

4. **Ductility :** It is the ability of metals due to which it can be drawn into wires. Copper, aluminium, iron can be drawn into wires. Silver, gold and platinum are highly ductile metals. 1 gram of gold can be drawn into 2 km long wire.

5. **Effect of Tapping (Sonorous) :** When metals are struck with hard substance, they produce sound, i.e., they are sonorous. Brass and bronze are highly sonorous. They are used in making bells and gongs.

6. **Electrical Conductance :** It is the property due to which electric current can pass through the metal. It is due to presence of free electrons or mobile electrons, e.g., copper, silver, gold, aluminium are good conductors of electricity. Silver is best conductor of electricity followed by copper, gold, aluminium and tungsten. Mercury and lead have low electrical conductivity due to high resistance.

7. **Thermal Conductivity :** It is the property due to which metals can conduct heat e.g., copper, silver, aluminium, gold and iron are good conductors of heat.

8. **Metallic Lustre :** Most of metals have shiny surface i.e., they show metallic lustre e.g., Au, Ag, Pt are lustrous.

9. **Exceptions if classification of metals and non-metals is done on the basis of physical properties :**
   - (i) All metals are hard except sodium, potassium, lithium, They can be cut even with knife. Osmium (Os) is hardest metal. Lithium is lightest metal.
   - (ii) All metals are solids except mercury. Cesium, francium, germanium and gallium are low melting solids. Gallium becomes liquid if kept on palm. But gallium has very high boiling point which makes it useful for high temperature thermometers.
   - (iii) Iodine is a non-metals but has metallic lustre. Diamond (an allotrope of carbon) is highly lustrous.
   - (iv) Non-metals have low melting and boiling points but diamond, graphite, boron and silicon have high melting and boiling points.
   - (v) Metals have high melting and boiling points. Tungsten has highest melting point whereas sodium, potassium have low melting and boiling points.
   - (vi) Non-metals are bad conductors of heat and electricity, but graphite is a non-metals which is good conductor of heat and electricity.
   - (vii) Metals generally have high density e.g., Cu, Cr, Ag, Au, Pt, Os, Ir, Ti, W, etc. Sodium, potassium, and lithium have density less than 1 g cm^{-3}. Diamond is non-metals but still has high density.
   - (viii) Generally, metals have 1 to 3 valence electrons. Hydrogen has one valence electron but still it is non-metal. Helium has 2 valence electrons but still it is non-metals.
   - (ix) Metals usually have a bright metallic lustre and are silvery white in appearance, but gold is yellow in colour, copper is reddish brown.
   - (x) Non-metals generally do not from alloys but carbon is alloyed with iron to form steel.

   Thus metals and non-metals cannot be classified on the basis of physical properties only. Chemical properties are more suitable to classify them.

10. **Anodising :** The process of forming oxide layer on the surface of metal is called anodising, e.g., Aluminium forms an oxide layer on its surface when exposed to air. It is non-penetrating layer which protects it from corrosion. The layer can be made more thick with the help of anodising. It is a process in which cleaned aluminum metals is taken as anode. Dilute H_2SO_4 acts as electrolyte. When electric current is passed, O_2 gas is liberated which reacts with aluminium to form a layer of aluminium oxide.

    The oxide layer can be made coloured like red, blue, etc., so as to make it more attractive for making decorative articles.

    Metal burns in presence of oxygen to from metal oxide. Let us perform and experiment.

11. **Solubility of Oxides in water :** Some metal oxides are soluble in water. Soluble base are called alkalies e.g.,

$$K_2O + H_2O \longrightarrow \underset{\text{Potassium hydroxide}}{2KOH}$$

$$\underset{\text{Sodium oxide}}{Na_2O} + H_2O \longrightarrow \underset{\text{Sodium hydroxide}}{2NaOH}$$

$$\underset{\text{Calcium oxide}}{CaO} + H_2O \longrightarrow \underset{\text{Calcium hydroxide}}{Ca(OH)_2}$$

$$\underset{\text{Magnesium oxide}}{MgO} + H_2O(\text{hot}) \longrightarrow \underset{\text{Magnesium hydroxide}}{MgO(H)_2}$$

12. **Alkalies :** Those base, which are soluble in water are called alkalies. For example, sodium hydroxide (NaOH), potassium hydroxide (KOH), calcium hydroxide $Ca(OH)_2$ etc. They turn phenolphthanlein pink and turn red litmus blue.

13. **Basic Oxides :** Those oxide, which react with acids or acidic or acidic oxides to form salt and water are called basic oxides. Metals react with oxygen to form metallic oxides which are

generally basic in nature e.g., Na_2O (sodium oxide.) CaO (calcium oxide), K_2O (potassium oxide), MgO (magnesium oxide) are basic oxides.

$$MgO + HCl \longrightarrow MgCl_2 + H_2O$$

14. **Acidic oxides :** Those oxides, which react with bases or basic oxides to form salt and water, are called acidic oxides. Non-metals react with oxygen to form non-metallic oxides which are generally acidic in nature, e.g., CO_2 (carbon dioxide), SO_2 (sulphur dioxide), SO_3 (sulphur trioxide), P_2O_5 (phosphorus pentoxide), SiO_2 (silicon dioxide) are acidic oxides.

$$CO_2 + 2NaOH \rightarrow Na_2CO_3 + H_2O$$

15. **Amphotoric oxide :** These oxides are both acidic as well as basic in nature. They react both with acids as well as bases to form salt and water. Some metals form amphoteric oxides. ZnO (zinc oxide) and Al_2O_3 (aluminium oxide) are amphoteric oxides.

$$ZnO + 2HCl \rightarrow ZnCl_2 + H_2O$$

$$\underset{\text{Zinc oxide}}{ZnO} + \underset{\text{Sodium hydroxide}}{2NaOH} \rightarrow \underset{\text{Sodium zincate}}{Na_2ZnO_2} + H_2O$$

16. **Neutral Oxides :** These oxides are neither acidic nor basic in nature. They neither react with acids nor with bases. Some non-metals form neutral oxides. Carbon monoxide (CO), nitrogen oxide (NO), Nitrous oxide (N_2O) are examples of neutral oxides.

17. **Electrovalent Compounds :** The compounds in which metal loses electrons and non-metal gains electrons are called electrovalent compounds, e.g., NaCl (sodium chloride), KCl (potassium chloride), etc.

$$\underset{\text{Sodium}}{2Na} + \underset{\text{Chlorine}}{Cl_2} \rightarrow \underset{\text{Sodium chloride}}{2NaCl}$$

18. **Hydrides :** When metals react with hydrogen, the compounds formed are called hydrides. e.g., sodium hydride (NaH), calcium hydride (CaH_2). In these compounds, metals lose electrons whereas hydrogen gains electron.

$$\underset{\text{Sodium}}{2Na} + \underset{\text{Hydrogen}}{H_2} \rightarrow \underset{\text{Sodium hydride}}{2NaH}$$

19. **Reaction of metal oxides with Acids :** Metal oxides react with acid to form salt and water because most of the metal oxides are basic or amphoteric in nature. Amphoteric oxides react with both acids as well as bases.

$$Na_2O + 2HCl(dil) \rightarrow 2NaCl + H_2O$$

$$CaO(s) + H_2SO_4(dil) \rightarrow CaSO_4 + H_2O$$

$$MgO(s) + 2HCl(dil) \rightarrow MgCl_2 + H_2O$$

$$CuO(s) + 2HCl(dil) \rightarrow CuCl_2 + H_2O$$

$$CuO(s) + H_2SO_4(dil) \rightarrow CuSO_4 + H_2O$$

$$Fe_2O_3(s) + 6HCl(dil) \rightarrow 2FeCl_3 + 3H_2O$$

$$SnO_2 + 4HCl \rightarrow SnCl_4 + 2H_2O$$

$$PbO + 2HCl \rightarrow PbCl_2 + H_2O$$

$$PbO + H_2SO_4 \rightarrow PbSO_4 + H_2O$$

$$AlP_3 + 6HCl \rightarrow 2AlCl_3 + 3H_2O$$

$$\underset{\text{Aluminium oxide}}{Al_2O_3} + \underset{\text{Sodium hydroxide}}{2NaOH} \rightarrow \underset{\text{Sodium aluminate}}{2NaAlO_2} + \underset{\text{Water}}{H_2O}$$

$$\underset{\text{Stannic oxide}}{SnO_2} + \underset{\text{Sodium hydroxide}}{2NaOH} \rightarrow \underset{\text{Sodium stannate}}{Na_2SnO_3} + \underset{\text{Water}}{H_2O}$$

$$HgO + 2HCl \rightarrow HgCl_2 + H_2O$$

$$HgO + H_2SO_4 \rightarrow HgSO_4 + H_2O$$

$$ZnO + 2HCl \rightarrow ZnCl_2 + H_2O$$

$$ZnO + H_2SO_4 \rightarrow ZnSO_4 + H_2O$$

$$ZnO + 2KOH \rightarrow \underset{\text{Potassium zincate}}{K_2ZnO_2} + H_2O$$

20. **Reactivity of Metals :** All the metals do not react with the same rate. Some react very fast, some react moderately whereas others react very slowly e.g., sodium, potassium react with oxygen

at room temperature vigorously to form oxide. They can catch fire in presence of moist air. These metals are kept in kerosene oil or benzene so as to protect them from formation of oxide and hydroxide in open air.

At room temperature, metals like Al, Zn, Cu, Mg, Sn, Pb form oxide layer on their surface and become dull. This oxide layer makes aluminium passive and does not allow it to react further with H_2O, O_2 and even conc. HNO_3. Copper is less reactive and forms black coloured oxide and gives green coloured flame with blue tip in burner. Magnesium burns with dazzling light forming MgO. Silver, gold and platinum do not react with oxygen. Mercury forms red coloured oxide, HgO.

Reactivity Series of Metals :

The series of metals in decreasing order of reactivity is called reactivity or activity series of metals. The metals at the top are most reactive whereas metals at the bottom are less reactive. The following is activity series of metals. The metals above hydrogen are more reactive than hydrogen. They can displace hydrogen from dilute acids and water. Metals below hydrogen are less reactive than hydrogen and cannot displace hydrogen from dilute acids and water.

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

21. **Reaction of Metals with Water :** Some metals like Na, K, Ca react with cold water vigorously to form hydroxide and liberate hydrogen gas. Some metals like Mg, Zn, Al react with hot water to form oxides and hydroxides and hydrogen gas. Some metals like Fe reacts with steam to form Fe_3O_4 and $H_2(g)$. Some metals like Cu, Ag, Au, Hg, Pb and Pt do not react with water at all because they are less reactive than hydrogen. Let us perform the following experiment.

22. **Reaction of Metals with Acids :** Metals react with dilute acids to from salt and hydrogen gas. The metal replaces hydrogen of the acid to from salt. Let us perform experiment to show the reaction of metals with acids.

23. **Aqua Regia :** It is mixture of conc. HCl and conc. HNO_3 (nitric acid) in the ratio of 3 : 1. It can dissolve gold and platinum. Gold and platinum do not react with conc. HCl. They do not react even with conc. HNO_3. They dissolve in aqua regia. Aqua regia is a strong oxidising agent due to formation of NOCl (Nitrosyl chloride) and chlorine produced by reaction of two acids. Aqua regia (Latin word for royal water) is a highly corrosive and fuming liquid. Therefore, it should be kept away from eyes and skin.

24. **Reaction of metals with solution of other metal salts :** More reactive metal can displace less reactive metal from their salt solution. These reactions are called displacement reactions. Let is carry out these reaction experimentally.

25. **Reason for difference in Reactivity of Metals :**

    We have observed with the help of experiments that some metals are less reactive whereas other metals are more reactive. Let us find out the reason why are some metals more reactive whereas other metals are less reactive.

    Metals are those elements which can lose electrons easily. The reactivity of metals will depend upon how easily metal can lose electron and form positively charged ion. Metals have 1 to

3 electrons whereas non-metals have 4 to 8 electrons. hydrogen and Helium have one and two electrons respectively but still they are non-metals because they cannot lose electrons easily. Boron has three valence electrons but still it is non-metals.

26. Electronic Configuration of elements and classification as Metals and Non-metals :

S.No.	Element	Symbol	Atomic number	Electronic configuration K L M N O	Valence electrons	Metal/Non-metal
1.	**Hydrogen**	**H**	**1**	**1**	**1**	**Non-metal**
2.	**Helium**	**He**	**2**	**2**	**2**	**Non-metal**
3.	**Neon**	**Ne**	**10**	**2, 8**	**8**	**Non-metal(noble gas)**
4.	**Boron**	**B**	**5**	**2, 3**	**3**	**Non-metal**
5.	**Carbon**	**C**	**6**	**2, 4**	**4**	**Non-metal**
6.	**Nitrogen**	**N**	**7**	**2, 5**	**5**	**Non-metal**
7.	**Oxygen**	**O**	**8**	**2, 6**	**6**	**Non-metal**
8.	**Fluorine**	**F**	**9**	**2, 7**	**7**	**Non-metal**
9.	**Sodium**	**Na**	**11**	**2, 8, 1**	**1**	**Metal**
10.	**Magnesium**	**Mg**	**12**	**2, 8, 2**	**2**	**Metal**
11.	**Aluminium**	**Al**	**13**	**2, 8, 3**	**3**	**Metal**
12.	**Potassium**	**K**	**19**	**2, 8, 8, 1**	**1**	**Metal**
13.	**Calcium**	**Ca**	**20**	**2, 8, 8, 2**	**2**	**Metal**

27. Reason for Metals for Losing Electrons : In the electronic configuration of elements, you have observed that noble gases have 8 electrons in their outermost shell and they are quite stable except helium which has two valence electrons. It is also quite stable.

It means all metals will try to lose electrons to acquire nearest noble gas configuration, e.g.,

$$\underset{2,1}{\mathrm{Li}} \longrightarrow \underset{2}{\mathrm{Li}^+} + e^-$$

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

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

$$\underset{2,8,18,8,1}{\mathrm{Rb}} \longrightarrow \underset{2,8,18,8}{\mathrm{Rb}^+} + e^-$$

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

$$\underset{2,8,8,2}{\mathrm{Ca}} \longrightarrow \underset{2,8,8}{\mathrm{Ca}^{2+}} + 2e^-$$

$$\underset{2,8,3}{\mathrm{Al}} \longrightarrow \underset{2,8}{\mathrm{Al}^{3+}} + 3e^-$$

28. Reason for Non-metals for Gaining Electrons : Non-metals can gain electrons to form negative ions so as to acquire stable noble gas configuration, i.e., 8 electrons in its outermost orbit.

$$\underset{2,8,7}{\mathrm{Cl}^{3+}} + e^- \longrightarrow \underset{2,8,8}{\mathrm{Cl}^-}$$

$$\underset{2,5}{\mathrm{N}} + 3e^- \longrightarrow \underset{2,8}{\mathrm{N}^{3-}}$$

Non-metals gain electrons to complete their octet.

29. Octet : A stable group of eight electrons in the outermost orbit of the atom.

30. Ionic bond : The bond which is formed by loss and gain of electrons is called ionic or electrovalent bond.

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

$\underset{2,8,7}{Cl} + e^- \longrightarrow \underset{2,8,8}{Cl^-}$

$(Na^+)(:\underset{\cdot\cdot}{\ddot{Cl}}:^-$ is electron dot diagram.

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

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

$(Mg^{2+})\ (:\underset{\cdot\cdot}{\ddot{Cl}}:^-)_2$ is electron dot diagram.

$\underset{2,8,3}{Al} \longrightarrow Al^{3+} + 3e^-$

$3F + 3e^- \longrightarrow 3F^-$

$(Al^{3+})\ (:\underset{\cdot\cdot}{\ddot{F}}:^-)_3$ is electron dot diagram.

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

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

$(Na^+)_2(:\underset{\cdot\cdot}{\ddot{F}}:^{2-})$ is electron dot diagram.

31. Properties of Ionic Compounds :

(i) Ionic compounds are solids, e.g., NaCl, KCl, $CuSO_4$, K_2SO_4, $NaNO_3$ are solids.

(ii) Ionic compounds are somewhat hard and brittle. It is due to strong force of attraction between them.

(iii) Ionic compounds have closed packed structures, e.g., NaCl has face centred cubic structure.

(iv) Ionic compounds have high melting and boiling points. It is because of strong force of attraction between oppositely charged ions, therefore, high energy is required to break the metallic bonds between ions.

(v) Ionic or electro valent compounds are soluble in water because they form ions in aqueous solution.

(vi) Ionic compounds are insoluble in organic solvents like alcohol, acetone, ether, chloroform, carbon disulphide etc. because they do not form ions in organic solvents.

(vii)Ionic compounds conduct electricity in molten state and in aqueous solution because ions carry current. The movement of ions take place towards oppositely charged electrode in electric field.

(viii)They do not conduct electricity in solid state because ions are not free to move.

32. Conclusion : It shows that ionic compounds conduct electricity in aqueous solution. Corrosion. It is a process in which metal reacts with substance present in atmosphere to form surface compounds, e.g., silver metal turns black due to formation of Ag_2S. Copper forms a greenish layer of $CuCO_3.Cu(OH)_2$ on its surface. Iron forms reddish brown coating of hydrated ferric oxide, $Fe_2O_3{\cdot}xH_2O$.

Distinctions between metals and Nonmetals

(A) Distinctions based on physical properties

Metals	Non-metals
1. Metals have lustre called metallic lustre.	Non-metals do not possess any metallic luster. **Exceptions :** Iodine and graphite possess metallic lustre
2. Metals are generally electropositive	Nonmetals are generally electronegative . **Exception :** Hydrogen is a non-metal but it is electropositive.
3. Metals are generally good conductors of heat and electricity.	Non-metals are generally bad conductor of heat and electricity. Exception : Graphite is a non-metals but it is a good conductor of electricity. Hydrogen, though a nonmetal, is a good conductor of heat.
4. Metals are usually solids at ordinary temperature. **Exception :** Mercury is a metal but it is a liquid at ordinary temperature.	Nonmetals are gases or solids at ordinary temperature. **Exception :** Bromine is a non-metal but it is a liquid at ordinary temperature.

5. Metals are malleable and ductile, have high density, and reflect light. **Exception :** Sodium and potassium are metals but their densities are less than that of water	Non-metals are not malleable and ductile, have low density, and do not reflect light. **Exception :** Plastic sulphur is ductile.
6. Metals when hammered emit a characteristic sound called metallic sound	Nonmetals do not emit metallic sound.

(B) Distinctions based on physical properties

Metals	Non-metals
1. Metals form positive ions. Example: Na^+, K^+, Ca^{2+}, Mg^{2+}, Fe^{2+}, etc.	Nonmetals form negative ions. Example: Cl^-, S^{2-}, N^{3-}, etc.
2. Oxides of metals are basic in nature, i.e., the oxides of metals react with water to give bases or alkalis. $Na_2O + H_2O \longrightarrow 2NaOH$ $CaO + H_2O \longrightarrow Ca(OH)_2$	Oxides of nonmetals are acidic in nature, i.e., the oxides of nonmetals react with water to give acids. $CO_2 + H_2O \rightarrow \underset{\text{carbonic acid}}{H_2CO_3}$ $SO_2 + H_2O \rightarrow \underset{\text{sulphurous acid}}{H_2SO_3}$ $SO_3 + H_2O \rightarrow \underset{\text{sulphuric acid}}{H_2SO_4}$
3. Metals dissolve in dilute acids to produce hydrogen gas. $Zn + H_2SO_4 \rightarrow ZnSO4 + H_2$ Exceptions Some metals do not produce hydrogen when treated with an acid under ordinary circumstances.	Nonmetals generally do not dissolve in dilute acids.
4. Metals in general do not combine with hydrogen. **Exception :** Some metals (Na, Ca, Li, Be, etc.) combine with hydrogen to form non-volatile unstable hydrides.	Nonmetals combine with hydrogen to form stable compounds.
5. Metallic chlorides are generally not hydrolyzed by water, or are only partially hydrolyzed. $NaCl + H_2O \rightarrow$ No hydrolysis $AlCl_3 + 3H_2O \rightarrow Al(OH)_3 + 3HCl$	Chlorides of nonmetals are usually hydrolyzed by water. $PCl_3 + 3H_2O \rightarrow 3HCl + H_3PO_3$ $SiCl_4 + 4H_2O \rightarrow Si(OH)_4 + 4HCl$

➤ OCCURRENCE OF METALS IN NATURE

Metals occur in nature in free state or in combined state. A metal is said to occur native or free when it is found in nature in the metallic state. For example, gold may be found in nature as metal. This is because gold when left exposed to air practically does not undergo any change. It is not reacted upon by moisture, oxygen and carbon dioxide of the air. Thus, those metals which remain unaffected by moisture, oxygen and carbon dioxide of the air can occur native or free. In other words, the unreactive metals occur in nature in free state because of their low reactivity towards chemical reagents. Another example of an unreactive metal is silver.

The reactive metals, i.e., the metals which react with moisture, oxygen, carbon dioxide or other chemical reagents, are not found in nature in free state, but in combined state in the form of compounds.

Metals usually occur in combination with nonmetallic elements. The native occurrence is comparatively rare.

Occurrence of metals in nature

In free state
(gold, platinum, silver,
mercury etc.)

In combined state
(sodium, calcium, potassium,
aluminium, lead, copper etc)

Minerals and Ores :

Metal-bearing substances, found in the earths crust, are called minerals. In other words, the solid compounds of metals occurring in nature are called minerals. For example, NaCl, KCl, $CaCO_3$, $MgCO_3$, ZnS, Cu_2S, Fe_2S_3 etc., which are found in nature are minerals.

Some minerals and their occurrence in India are given below.

Metal	**Minerals**	**Places of occurrence in India**
1. Sodium	Tincal, borax	Ladakh (Kashmir)
2. Magnesium	Dolomite, magnesite	Tamil Nadu
3. Calcium	Gypsum	Rajasthan, Tamil Nadu, Jammu and Kashmir
4. Aluminium	Bauxite	UP, Maharashtra,'MP, Orissa
5. Copper	Chalcopyrites or copper pyrites, malachite	Jharkhand, Orissa, MP
6. Zinc	Zincblende	Rajasthan
7. Manganese	Pyrolusite	Maharashtra, Karnataka and Jharkhand
8. Iron	Haematite	Jharkhand and Karnataka

Ores :

The minerals from which metals can be obtained on a commercial scale are called ores. In other words, the minerals from which metals can be extracted profitably are called ores. Both bauxite

($Al_2O_3 \cdot 2H_2O$) and clay ($Al_2O_3 \cdot 2SiO_2 . 2H_2O$) are minerals of aluminium. However, it is bauxite that is chiefly used to obtain aluminium commercially. So, bauxite, not clay, is an ore of aluminium. Thus:

(i) All ores are minerals, but all minerals are not ores.

(ii) An are is rich in the amount of the metal. The amount of foreign materials or impurities is low in an ore.

<u>Different types of ores</u>

The different types of ores that are used in the extraction of metals are listed below.

1. **Oxides :** Copper, aluminium, zinc, tin, iron, etc., occur as oxides.
2. **Sulphides :** Silver, copper, zinc, mercury, lead, iron, etc., occur as sulphides.
3. **Carbonates :** Sodium, copper, calcium, magnesium, zinc, lead, iron, etc., occur as carbonates.
4. **Sulphates :** Sodium, calcium, magnesium, lead, etc., occur as sulphates.

5. **Halides :** Sodium, calcium, silver, etc., occur as halides.
6. **Phosphates :** Calcium occurs as phosphate.

➤ METALLURGY

The process of extracting metals from their ores and refining them for use is known as metallurgy. In other words, the process of obtaining a metal from its ores is called metallurgy of the metal.

Common terms used in discussing metallurgical operations :

1. **Charge :** The mixture of materials fed to a furnace for obtaining the metal is called charge.
2. **Dressing of the ore :** The removal of impurities associated with ore is called dressing or enrichment of the ore.
3. **Calcination :** The process of heating a metal-rich ore to a high temperature to convert the metal into its oxide, either in absence or insufficient supply of air is called calcination.
4. **Roasting :** The process of heating a finely ground ore to a high temperature in excess of air is called roasting. Roasting converts the metal present in the ore to its oxide.
5. **Flux :** A flux is a substance that is mixed with the fumance charge (calcined or roasted ore and coke) to remove the infusible impurities present in the ore.
6. **Slag :** Flux combines with the infusible impurities to convert them into a fusible substance called slag. Being light, slag floats over the molten metal and is removed from there.

   Impurities present in metal oxides may be acidic or basic. For acidic impurities, such as SiO_2 or P_2O_5, a basic flux (e.g., CaO) is added to the charge. If basic impurities such as MnO are present, silica is added to the charge.

Impurity	Flux	Slag
SiO_2 +	$CaO \longrightarrow$	$CaSiO_3$
P_2O_5 +	$3CaO \longrightarrow$	$Ca_3(PO_4)_2$
MnO +	$SiO_2 \longrightarrow$	$MnSiO_3$

7. **Gangue or matrix :** The ore mined from the earth's crust contains some unwanted substances or impurities, such as sand, rocky or clayey materials. These substances are called gangue or matrix. The gangue has to be removed before the process of extraction of metals starts.
8. **Smelting** The process of obtaining the metal by reducing its oxide ore with coke is known as smelting.

Principles of metallurgy

The extraction of a metal from its ore depends upon the reactivity of the metals.

1. Metals at the top of the activity series (K, Na, Ca, Mg, etc.) are highly reactive. They do not occur in the free state. They are extracted by the electrolysis of the molten ore.
2. Metals in the middle of the activity series (Zn, Fe, Pb, etc.) are moderately reactive. These are obtained by roasting and calcination of their sulphide or carbonate ore.
3. Metals at the bottom of the activity series (Au, Ag, Pt, Cu) being the least reactive are found in the free state. Copper and silver also occur as their sulphide or oxide ores. These are obtained by the process of roasting.

Metallurgical Operations :

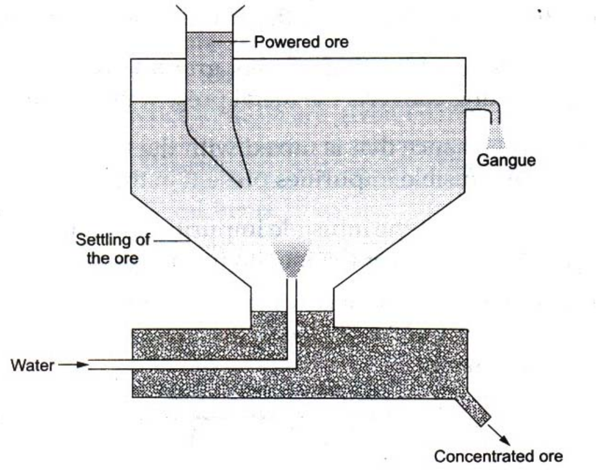
The various steps used in metallurgy are listed below.

1. Enrichment or dressing of the ore
2. Conversion of the enriched ore into the oxide of metal
3. Extraction of metal from the metal oxide
4. Refining or purification of the metal

Enrichment or dressing of an ore :

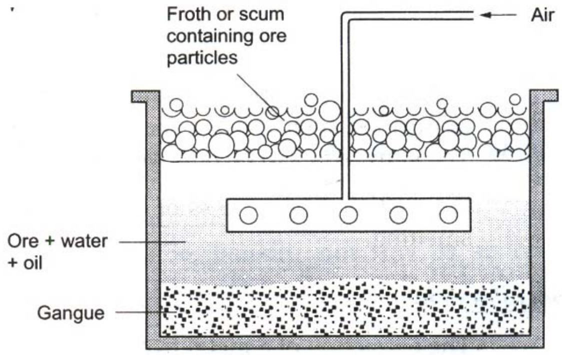
An ore mined from the earth's crust contains a number of impurities (gangue), which must be removed. The ore, free from gangue, then becomes suitable for subsequent treatment. Enrichment or dressing of an ore is carried out by the following methods.

(i) **Levigation :** The powdered ore is washed in a jet of water. The lighter, rocky and earthy impurities are washed away by water, while heavier ore particles are left behind to settle down at the bottom. This process is also called hydraulic washing.



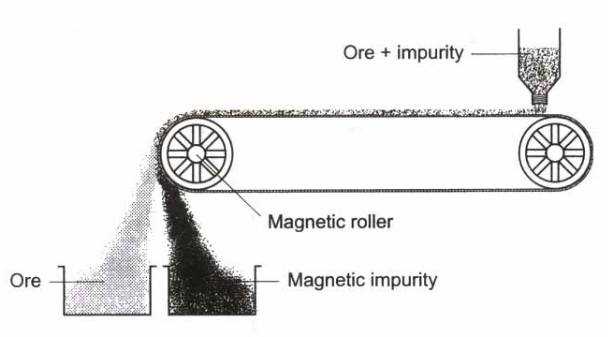
(ii) Froth floatation : Sulphide ores of copper, lead and zinc are generally concentrated by this method.

The finely powdered ore is mixed with water and a small amount of oil in a tank. Air is blown into the mixture. A froth or scum is produced at the surface. The ore particles are carried by the froth to the surface. The earthy impurities sink to the bottom. The froth along with the ore is removed. An acid is added to break up the froth. The concentrated ore is filtered and dried.



(iii) Liquation : This process is used to concentrate the ore whose melting point is lower than that of the impurities. Stibnite, an ore of antimony, is concentrated by this method. The impure ore is heated. The ore melts and flows along the surface. The impurities are left behind.

(iv) Magnetic separation : This method is used when the magnetic properties of the ore and the impurities are different. For example, tinstone, an ore of tin, contains wolfram as an impurity that is magnetic. To remove this impurity, the ore is finely powdered to make the magnetic and the nonmagnetic particles distinctly separate. The powdered tinstone is spread on a belt moving over electromagnetic rollers in figure. The wolfram, being magnetic, is attracted and gets collected in the pot near the magnet. Tinstone falls away from the magnet.



(v) Leaching or chemical separation : In this method, the powdered ore is treated with a suitable solvent. The ore dissolves in it while the impurities remain undissolved. For example, the bauxite ore contains Fe_2O_3, SiO_2, etc., as impurities. The ore is powdered and treated with a solution of sodium hydroxide. $A1_2O_3$ and SiO_2 present in the ore dissolve, forming sodium aluminate and sodium silicate respectively. The impurities are left behind undissolved. The impurities are filtered off. The filtrate containing sodium aluminate and sodium silicate is stirred with some freshly prepared aluminium hydroxide for several hours. Sodium aluminate undergoes hydrolysis producing aluminium hydroxide as precipitate. The addition of aluminium hydroxide accelerates the precipitation of hydroxide. Soluble sodium silicate remains in solution. The precipitate, when filtered, washed, dried and ignited, gives pure alumina (Al_2O_3).

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

$$NaAlO_2 + 2H_2O \rightarrow Al(OH) + NaOH$$

$$2Al(OH) \rightarrow Al_2O_3 + 3H_2O$$

Conversion of the enriched ore into the oxide of metal

It is easier to obtain metals from their oxides than from their carbonates or sulphides. Hence, the concentrated ore is converted into the oxide of metal which is then reduced to metal. This conversion to oxide is done by the process of calcination or roasting. In this process the ore is heated very strongly in the absence of air, keeping the temperature below its melting point so that volatile impurities are driven off.

EXAMPLES :

(i) Oxide ores are calcined to remove moisture and other volatile impurities.

$Al_2O_3 . 2H_2O \rightarrow Al_2O_3 + 2H_2O$

(ii) Carbonate ores are calcined to expel carbon dioxide.

$CaCO_3 \rightarrow CaO + CO_2$

$CaC0_3 \cdot MgCO_3 \rightarrow CaO + MgO + 2CO_2$

$\underset{\text{calamine ore}}{ZnCO_3} \rightarrow ZnO + CO_2$

$CuCO_3 \cdot Cu(OH) \rightarrow 2CuO + H_2O + CO_2$

Sulphide ores are usually converted to oxides by roasting. The process involves heating the ore at a temperature below its fusion point, but always in the presence of air so that it may be oxidized.

Arsenic and similar other elements present in free state or combined state are also oxidized to volatile oxides.

(i) Zinc blende (ZnS) is roasted in air to convert it into zinc oxide.

$2ZnS + 3O_2 \rightarrow 2ZnO + 2SO_2$

(ii) Galena (PbS) is converted into litharge (PbO) by roasting.

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

(iii) Cinnabar (HgS) is roasted to convert it directly into mercury (Hg).

$HgS + O_2 \rightarrow Hg + SO_2$

(iv) Iron pyrite (FeS_2) is converted into ferric oxide (Fe203) by roasting.

$4FeS_2 + 11O_2 \rightarrow 2Fe_2O_3 + 8SO_2$

Thus, both calcination and roasting produce oxide of the metal. However, there are a few points of difference between the two processes.

Calcination	Roasting
1. The ore is heated in the absence of air.	The ore is heated in the presence of air.
2. It is used for oxide or carbonate ores.	It is used for sulphide ores.

Chloride ores remains unchanged by calcination or roasting.

Extraction of metal from metal oxide :

A metal oxide thus produced is then reduced into metal. For this, the method used depends upon the reactivity of the metal being extracted. The following methods are used.

(i) Reduction by heat alone Metals occupying lower positions in the activity series can be obtained by heating their oxides.

$$2HgS + 3O_2 \rightarrow 2HgO + 2SO_2$$

$$2HgO \rightarrow 2Hg + O_2$$

(ii) Chemical reduction Metals in the middle of the activity series (Fe, Zn, Ni, Sn, etc.) cannot be obtained by heating their compounds alone. They require to be heated with a reducing agent, usually carbon (coke). When a metal oxide is heated with carbon, it is reduced to free metal.

$$MO + \underset{\text{(M=metal)}}{C \rightarrow M} + CO$$

The reduction of metal oxides with carbon is known as smelting. The impurities are removed as slag

EXAMPLES :

(i) When zinc oxide is heated with carbon, zinc metal is obtained.

$$ZnO + C \rightarrow Zn + CO$$

(ii) When stannic oxide is heated with carbon, tin metal is produced.

$$SnO_2 + 2C \rightarrow Sn + 2CO$$

(iii) Ferric oxide (Fe_2O_3) is reduced to iron by heating with coke in a blast furnace.

$$Fe_2O_3 + 3C \rightarrow 2Fe + 3CO$$

Reduction with aluminium (thermit process or alumino-thermic process)

Some metal oxides cannot be reduced satisfactorily by carbon. For them, aluminium, a more reactive metal, is used. The process is called thermic process or alumino-thermic process.

EXAMPLES :

(i) Manganese dioxide is reduced to manganese by heating with aluminium.

$$3MnO_2 + 4Al \rightarrow 3Mn + 2Al_2O_3$$

(ii) Ferric oxide (FeP3) is reduced by aluminium to free iron.

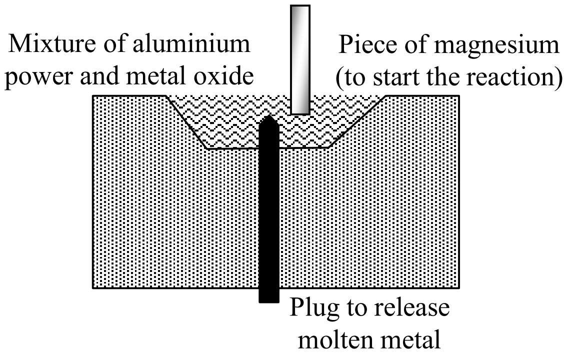
$$Fe_2O_3 + 2Al \rightarrow 2Fe + Al_2O_3$$

(iii) Chromium sesquioxide is reduced by aluminium to chromium metal.

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

In the thermit process, aluminium powder is mixed with metal oxide. A piece of magnesium is set alight to start the reaction. The aluminium reduces the oxide to free metal.

In case of iron oxide, iron is obtained in the molten state. (The mixture of iron oxide and aluminium powder is called thermite). The molten iron may be allowed to trickle down to weld two iron objects together. Cracked machine parts, railway tracks, etc., are joined by this method.



(iii) Electrolytic reduction : The reactive metals (high up in the activity series) cannot be produced by any of the above methods. They are obtained by electrolytic reduction of their molten oxides or chlorides. During electrolysis, the cathode supplies electrons to metal ions for their reduction to the metal.

EXAMPLES :

(i) Sodium metal is obtained by the electrolysis of molten sodium chloride.

$$\begin{array}{c}
2NaCl \xrightarrow{\text{fused}} 2Na^+ + 2Cl^- \\
2Cl \xrightarrow{\qquad} Cl_2 + 2e(\text{at anode}) \\
2Na^+ \xrightarrow{\qquad} 2Na(\text{at anode}) \\
\hline
2NaCl \xrightarrow{\text{electrolysis}} 2Na + Cl_2
\end{array}$$

(ii) Magnesium metal is obtained by the electrolysis of molten magnesium.

$$\begin{array}{c}
MgCl_2 \xrightarrow{\text{fused}} Mg^{2+} + 2Cl^- \\
2Cl^- \xrightarrow{\qquad} Cl_2 + 2e(\text{at anode}) \\
Mg^{2+} + 2e \xrightarrow{\qquad} Mg(\text{at cathode}) \\
\hline
MgCl_2 \xrightarrow{\text{electrolysis}} Mg + Cl_2
\end{array}$$

(iii) Aluminium oxide (Al_2O_3) is reduced to aluminium by the electrolysis of molten aluminium oxide.

$$Al_2O_3 \rightarrow 2Al^{3+} + 3O^{2-}$$

The aluminium ions present in aluminium oxide go to the cathode and are reduced there to aluminium atoms.

Note : During electrolytic reduction of the molten salts, the metals are always liberated at the cathode.

(iv) Some specific methods Silver and gold are obtained by treating the ore with a solution of sodium cyanide. Sodium argentocyanide (in case of silver) or sodium aurocyanide (in case of gold) is obtained in the solution. On adding zinc dust to the solution, silver or gold is precipitated.

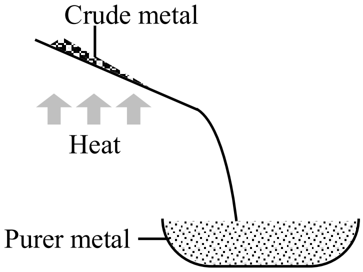
$$2Na[Ag(CN)_2] + Zn \rightarrow Na_2[Zn(CN)_4] + 2Ag$$

$2Na[Au(CN)_2] + Zn \rightarrow Na[Zn(CN)_4] + 2Au$

Refining of Metals :

The metal obtained from the ore is not pure. It contains various substances as impurities. The process of removing these impurities is called refining of the metal. Some of the methods generally applied for refining metals are discussed below.

1. **Liquation :** This process is used to separate metals of low melting points (e.g., tin and lead) from the metals of high melting points.

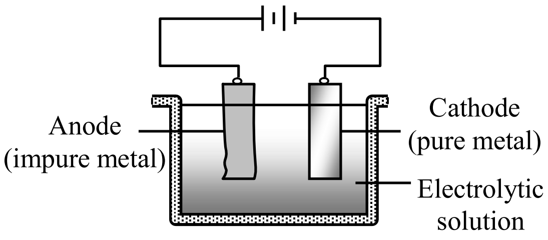


In this process, a sloping hearth is used. The hearth is kept at a temperature a little above the melting point of the metal. The impure metal is placed at the top of the hearth. The metal melts and flows down the hearth. The infusible impurities are left behind. This method is used in the purification of tin.

2. **Cupellation :** This method is used to purify silver, containing lead as an impurity. The impure silver is heated in the presence of air in a vessel made of bone-ash. This vessel is called cupel. Lead is oxidized to lead monoxide. Most of the lead monoxide is carried away in the blast of air. The remaining portion of the lead monoxide melts and is absorbed by the bone-ash. Pure silver is left behind.

3. **Poling :**Copper is purified by this method. The molten impure copper (called blister copper) is stirred thoroughly with poles of green wood. The gases escaping from the poles reduce the oxide of metal to the metal. The surface of the molten copper is kept covered with powdered charcoal so that copper may not be reoxidized in contact with air.

4. **Electrolytic refining :** This method is widely used for purification of metals. Several metals such as aluminium, copper, tin, lead, gold, zinc and chromium are purified by this method. The impure metal is made the anode while a strip of pure metal acts as the cathode. A solution of the salt of the metal acts as the electrolyte.



On passing electric current through the solution, pure metal gets deposited on the cathode. The more reactive impurities present in the metal to be purified go into solution and remain there. The less reactive impurities fall to the bottom of the electrolytic cell.

Ultra-pure Metals :

In the present age of technological advancement, metals of high purity are required for special purposes. For example, pure germanium is needed for semiconductor devices. Uranium of high-grade purity is used as fuel in nuclear reactors.

Two special techniques have been devised to prepare metals of very very high purity.

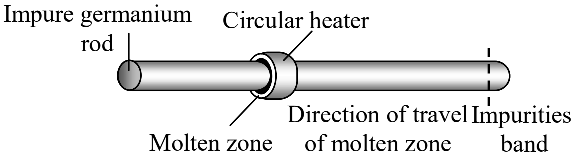
1. **Van Arkel method :** This method was developed by van Arkel to obtain ultra-pure metals. It is based on the thermal decomposition of metal compounds. It is used for obtaining pure titanium which is used in space technology. The impure titanium metal is converted into titanium tetra-iodide.

   The air in the barrel used in this process is removed to create a high vacuum. An iodine bulb is broken. Titanium metal is heated which reacts with iodine to form gaseous titanium tetra-iodide.

   $$Ti + 2I_2 \rightarrow TiI_4$$

   The impurities do not react with iodine. The vapour of titanium tetra-iodide is passed over a heated tungsten filament (1674 K). Titanium tetra-iodide gets decomposed into titanium and iodine. Pure titanium is deposited upon the filament and can be removed. The regenerated iodine can be reused to react with more titanium. The process is repeated.

2. **Zone refining method :** This method is capable of producing metals of high purity. Germanium, which is used in semiconductor devices, is purified by this method. In this method, advantage is taken of the fact that impure molten metal, when allowed to cool, deposits crystals of pure metal.



An impure germanium rod is provided with a circular heater. The heater is slowly moved along the metal rod. A band of the rod melts. As the heater moves away, the metal crystallizes out of the melt. The impurities are swept along the molten zone. Finally, the impurities reach the other end of the rod, and are removed.

SOME COMMON METALS

IRON

Symbol	Fe	Atomic number	26

Electronic Configuration of Iron :

The atomic number of iron is 26. This means that an atom of iron contains 26 electrons in its shells. The electronic configuration of iron is shown below.

	K	L	M	N
Fe(26)	2	8	14	2

Thus, an atom of iron contains two electrons in its outermost shell.

Occurrence of Iron

Iron is second to aluminium in terms of abundance in the earth's crust. It makes up 4.7% of the earth's crust. Free iron has been found in most meteorites.

Iron is a reactive metal. So it does not occur free in nature. In combined state, it occurs as oxide, sulphide, carbonate, etc. The important ores of iron are:

(i) Haematite, Fe_2O_3

(ii) Magnetite, Fe_3O_4

(iii) Limonite, $2Fe_2O_3 . 3H_2O$

(iv) Siderite, $FeCO_3$

(v) Iron pyrites, FeS_2

The most important ore of iron is haematite, which is used most commonly in the extraction of iron. The pyrite ore (FeS_2) is not used for the extraction of iron because of its high sulphur content.

Iron in India :

Iron metal has great economic importance. The world output of iron exceeds two hundred million tonnes per annum. In 2002-03, India's total production of iron reached almost 97 million tonnes. Besides, India has a vast deposit of iron ore: about 12,318 million tonnes of haematite and 5,396 million tonnes of magnetite. Most of these deposits are located in Jharkhand, Orissa, Chhattisgarh, Tamil Nadu, Karnataka and Maharashtra. The important iron and steel plants are located at Bhillai, Bokaro, Jamshedpur, Rourkela, Durgapur, Asansol and Bhadravati.

Extraction fo ron from Haematite :

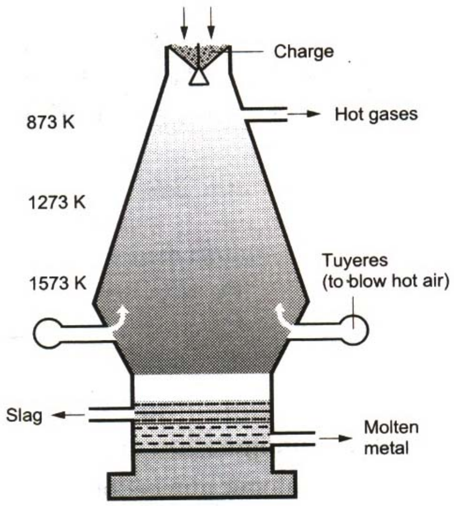
1. **Dressing of the ore :**

The big lumps of the ore are broken into small pieces and then washed with water to remove clay, sand and other adhering impurities. The ore thus becomes ready for treatment in the blast furnace.

2. **Smelting in the blast furnace :**

The concentrated ore is mixed with coke and limestone. The mixture is charged at the top of a blast furnace. The following reactions occur in the furnace.

(i) As the charge comes down to the 873 K region, the iron oxide is reduced by the ascending carbon monoxide gas produced by the burning of coke.



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

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

The iron thus obtained is called sponge iron.

(ii) At the 1273 K region, the silica is converted to slag.

$$CaCO_3 \rightarrow CaO + CO_2$$

$$CaO + SiO_2 \rightarrow \underset{\text{slag}}{CaSiO_3}$$

(iii) At the 1573 K region, sponge iron melts and dissolves carbon, phosphorus, silica, etc. The slag also fuses. The molten mass collects at the base of the furnace. The slag floats over it. The molten iron is taken out as required. This iron is called pig iron.

Function of limestone : Limestone is decomposed to give quicklime.

$$CaCO_3 \rightarrow CaO + CO_2$$

Quicklime combines with impurities like sand to form a molten slag (calcium silicate).

$$CaO + SiO_2 \rightarrow CaSiO_3$$

The slag floats on the surface of molten iron. It is taken out through a hole from time to time.

The formation of calcium silicate as slag not only removes unwanted silica but also keeps iron away from being oxidized.

Varieties of Iron :

1. **Pig iron and cast iron :** The iron produced in the blast furnace is pig iron. It contains a comparatively high percentage of carbon due to which it is hard and brittle. It also contains phosphorus, silicon and manganese as impurities. Pig iron is melted, mixed with steel scrap and allowed to cool in moulds to give cast iron. Cast iron is impure iron, and is hard and brittle.

2. **Wrought iron :** It is almost a pure form of iron. It contains only 0.12% to 0.25% carbon. It melts at a higher temperature (1773 K) than that at which cast iron melts. Wrought iron is obtained by melting cast iron on a hearth lined with ferric oxide (Fe_2O_3). The impurities such as carbon, phosphorus, silicon and manganese are oxidized by Fe_2O_3.

   $$Fe_2O_3 + 3C \rightarrow 2Fe + 3CO$$

   Wrought iron is soft, grey and tough. It is malleable and ductile. Hence, it can be drawn into sheets and stretched into wires. It is used in making chains, wire, anchors and cores of electromagnets.

3. **Steel :** It is an alloy of iron and carbon. It contains about 0.15 to 1.7% of carbon. There are different types of steel.

   (a) Mild steel : It contains less than 0.3% carbon. It is also called soft iron. Mild steel is used for making sheets and wires.

   (b) Hard steel : It contains higher percentage (0.7-1.7%) of carbon. It is used in making tools and instruments.

   (c) Alloy steels : Alloy steels are prepared by adding small amounts of nickel, cobalt, chromium, tungsten, molybdenum, manganese and silicon to steel. Alloy steels are used extensively in making rock-crushing machinery, helmets, armour plate, cutlery, springs, etc.

   (d) Medium steel : It contains 0.3-D.7% carbon. It is hard and is used in making rails, bridges, etc.

	C	Si	S	P	Mn
Cast iron	2-4.5%	0.7-3.5%	0.7-3.5%	0.05-1.5%	0.5-1.0%
Wrought iron	0.12-0.25%	0.03-2.0%	0.02-0.2%	0.04-0.07%	0.1-0.4%
Mild steel	0.15%	0.03%	0.05%	0.05%	0.50%

Tempering :

The hardness and elasticity of steel can be controlled by heat treatment. The steel is heated to a temperature below redness. It is then cooled slowly. The process is called tempering of steel. It is used to bring the steel to a suitable state of hardness and elasticity.

Annealing of steel :

Hard steel can be softened by heating it to a high temperature and then allowing it to cool down slowly. This process is called annealing.

Quenching of steel :

Hard steel is heated to a high temperature. It is then suddenly cooled by plunging into oil or water. Steel becomes as hard and brittle as glass. Steel produced in this way is known as quenched steel and the process of making such steel is known as quenching or hardening of steel.

Properties of Iron

Physical properties :

Pure iron has a grey colour. It is malleable and ductile. It is a good conductor of heat and electricity. It melts at 1808 K and boils at 3023 K. It has a density of 7.9×10^3 kg cm^{-3}.

Chemical properties :

1. **Valency :** Iron shows variable valency: 2 and 3. It forms divalent ion (Fe^{2+}) as well as trivalent ion (Fe^{3+}). The compounds in which iron shows divalency are known as ferrous compounds, whereas the compounds in which iron shows trivalency are known as ferric compounds. For example, in $FeCl_2$ the valency of iron is 2. So, it is called ferrous chloride. In $FeCl_3$, the valency of iron is 3. Hence, it is called ferric chloride.

2. **Action of air** : In the presence of moist air and carbon dioxide, iron gets covered with a thin deposit of rust. The rust consists of hydrated ferric oxide ($2Fe_2O_3{\cdot}3H_2O$).

3. **Action of water :** Red hot iron decomposes steam, forming ferrosoferric oxide and evolving hydrogen gas.

$$\underset{\text{iron}}{2Fe} + \underset{\text{steam}}{4H_2O} \rightarrow \underset{\text{ferrosoferric oxide}}{Fe_3O_4} + \underset{\text{hydrogen}}{4H_2}$$

4. **Action of acids** : Iron lies above hydrogen in the activity series of metals. So, it can displace hydrogen from dilute hydrochloric acid or dilute sulphuric acid. The corresponding ferrous salts are also produced.

(i) With hydrochloric acid

(a) Iron dissolves in dilute hydrochloric acid, forming ferrous chloride with the evolution of hydrogen gas.

$$Fe + 2HCl \rightarrow FeCl_2 + H_2$$

(b) Concentrated hydrochloric acid also produces hydrogen with iron.

(ii) With sulphuric acid

(a) Iron dissolves in dilute sulphuric acid, forming ferrous sulphate. Hydrogen gas is evolved in the reaction.

$$Fe + H_2SO_4 \rightarrow FeSO_4 + H_2$$

(b) Iron reacts with concentrated sulphuric acid to form ferrous sulphate with the evolution of sulphur dioxide. Hydrogen gas does not evolve in this reaction.

$$Fe + 2H_2SO_4 \rightarrow FeSO_4 + SO_2 + 2H_2O$$

Some ferric sulphate is also formed due to the oxidation of $FeSO_4$ by concentrated H_2SO_4,

$$2FeSO_4 + 2H_2SO_4 \rightarrow Fe_2(SO_4) + 2H_2O + SO_2$$

(iii) With nitric acid :

(a) Iron reacts with dilute nitric acid to form ferrous nitrate and ammonium nitrate.

$$4Fe + 10HNO_3 \rightarrow 4Fe(NO_3)_2 + NH_4NO_3 + 3H_2O$$

(b) With concentrated nitric acid, iron is rendered passive due to the formation of insoluble ferrosoferric oxide (Fe_3O_4) on the surface of iron.

5. **Action of halogens** : Halogens combine with heated iron, forming the halides of iron. For example, chlorine combines with heated iron to form ferric chloride.

$$2Fe + 3Cl_2 \rightarrow 2FeCl_3$$

6. **Action with sulphur :** When iron filings are heated with sulphur, iron sulphide is produced.

$$Fe + S \rightarrow FeS$$

7. **Displacement of less electropositive metal :** When an iron piece is dipped into a solution of copper sulphate, copper is displaced from the salt and gets deposited on the surface of iron. This is because copper is less electropositive than iron.

$$Fe + CuSO_4 \rightarrow Cu + FeSO_4$$

Tests to distinguish between ferrous and ferric salts

(i) When a ferrous salt solution is treated with a solution of sodium hydroxide, a greenish precipitate of ferrous hydroxide is obtained.

$$FeSO_4 + 2NaOH \rightarrow Fe(OH) + Na_2SO_4$$

When a ferric salt solution is treated with a solution of sodium hydroxide, a brown precipitate of ferric hydroxide is produced.

$$Fe_2(SO_4)_3 + 6NaOH \rightarrow 2Fe(OH) + 3Na_2SO_4$$

(ii) The ferrous salts are generally green-coloured, whereas the ferric salts are brown in colour.

Rusting :

When iron is exposed to moist air, a reddish-brown coating of a mixture of ferric oxide (Fe_2O_3) and ferric hydroxide (($Fe(OH_3)$)) is deposited on the surface of the metal. This reddish-brown coating is known as rust, and this process is known as rusting. Thus, the slow conversion of iron into a mixture of Fe_2O_3 and $Fe(OH)$ by water and atmospheric oxygen is known as rusting.

Rusting of iron is an oxidation reaction that occurs due to the attack of water and oxygen. It has been found that rusting does not take place in air-free water. It also does not occur in presence of oxygen alone. Both water and oxygen are essential for rusting. Thus, the following conditions are necessary for rusting:

(i) Presence of oxygen or air

(ii) Presence of water or moisture

The process of rusting is continuous. The strength of iron decreases gradually and finally the metal is destroyed completely.

Prevention of Rusting :

Iron can be prevented from rusting by keeping it out of contact with air and water, and also by converting it into an alloy. This can be achieved in the following ways.

1. By covering the surface of iron with grease, paint, varnish, enamel, etc.
2. By galvanizing iron: a thin coating of zinc is deposited on the surface of the iron object. This is done by electroplating. Since zinc does not corrode on exposure to air, zinc metal prevents iron from rusting.
3. By coating the surface of the iron object with chromium, tin, nickel, or aluminium. These metals resist corrosion. Hence, they protect iron from rusting.
4. By converting it into an alloy with chromium and nickel. This alloy is called stainless steel.

Uses of Iron :

1. Iron is used in making household utensils and equipments.
2. Wrought iron and cast iron are largely used in the manufacture of locomotives, railway lines, springs, tubes, etc.
3. Iron finds wide application in house construction, e.g., in the reinforcement of roofs and other parts of buildings.

ALUMINIUM

Symbol	Al	Atomic number	13

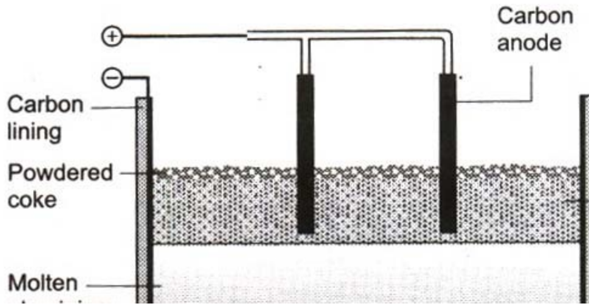
➤ HISTORY

Aluminium was first extracted in 1827 from aluminium chloride by treating it with sodium :

$AlCl_3 + 3Na \rightarrow Al + 3NaCl$. But the process was very expensive. Impressed by its properties, metallurgists all over the world attempted to develop a process for the commercial production of aluminium. But it remained a costly metal till 1886, when Heroult in France and Hall in USA succeeded in developing independently a process for the extraction of aluminium. Thus began large-scale production of aluminium in several parts of the world. Hall-Heroult process came into extensive use towards the end of the nineteenth century.

2. **Electrolytic reduction of alumina: Hall-Heroult process**

Alumina is mixed with cryolite and the mixture is melted in an iron cell.



Figure

Alumina melts at 2303 K. It is a bad conductor of electricity. But when mixed with cryolite and some calcium fluoride, the mixture becomes a good conductor of electricity, and melts at 1173-1223 K. The cryolite thus considerably reduces the' energy cost. The iron cell is lined inside with gas carbon, which serves as cathode. Carbon rods act as anode. The electrolyte, thus, contains Na+, Al^{3+}, F^{-} and O^{2-} ions.

On passing electric current, AP+ ions are discharged at the cathode and the 022ions, at the anode.

$Al^{3+} + 3e \rightarrow Al$ (at cathode)

$2O^{2-} \rightarrow O_2 + 4e$ (at anode)

Some quantity of oxygen formed in the reaction escapes and some reacts with the anode to form CO_2, Thus, carbon anode burns away due to its reaction with oxygen.

$C + O_2 \rightarrow CO_2$

Hence, the anode has to be replaced from time to time. This increases the cost of production of aluminium. Molten aluminium sinks to the bottom, and is taken out from there.

Alloy	Composition	Uses
1. Brass	Cu = 80%, Zn = 20%	Harder than pure Cu and Zn ; used for making utensils, cartridges, etc.
2. Bronze	Cu = 90%, Sn = 10%	For making statues, medals, ships, coins, machines, etc.
3. Solder (common)	Sn = 50%, Pb = 50%	For joining metals, soldering wires, electronic components, etc.
4. Duralumin	Al = 95.5%, Cu = 3% Mn = 1% Mg = 0.5%	In bodies of aircraft, kitchenware, automobile parts, etc
5. Babbit metal	Sn = 90%, Sb = 7%, Cu = 3%	In antifriction lining
6. German silver	Cu = 60%, Zn = 20%, Ni = 20%	For making utensils, ornaments, etc.
7. Gun metal	Cu = 60%, Sn = 10%	Gears, castings, etc.
8. Bell metal	Cu = 78%, Sn = 22%	Bells, gongs, etc.
9. Magnalium	Al = 90%, Mg = 10%	Balance beams, light instruments, etc.

10. Pewter	Sn = 75%, Pb = 25%	Cups, mugs, etc.
11. Type metal	Pb=82%, Sb = 15%, Sn = 3%	Casting type

Alloy steels			
Name	**Composition**	**Properties**	**Uses**
1. Manganese	Mn = 10 – 18%	Extremely hard, resistant to wear	Grinding machines, safes, etc.
2. Chrome-vanadium	Cr = 1-10%, V = 0.15%	Highly tensile, resistance to stress and torsin	Axle and other parts of automobiles
3. Nickel-chromium	Ni = 1-4%, Cr = 0.5-2%	High tensile strength, hard and highly elastic	Armour plates
4. 18–8	Cr = 18%, Ni = 8%	Resistance to corrosion	Cutlery, instrument
5. Alnico	Co = 5%	Highly magnetic	Powerful permanent magnet

Important Chemical Reactions :

1. $CaCO_3 \xrightarrow{\text{Heat}} CaO + CO_2$
2. $CaO + SiO_2 \longrightarrow CaSiO_3$
3. $C + O_2 \longrightarrow CO_2$
4. $4P + 5O_2 \longrightarrow 2P_2O_5$
5. $S + O_2 \longrightarrow SO_2$
6. $4Na + O_2 \longrightarrow 2Na_2O$
7. $4K + O_2 \longrightarrow 2K_2O$
8. $2Ca + O_2 \longrightarrow 2CaO$
9. $CO_2 + C \longrightarrow 2CO$
10. $Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$
11. $Al_2O_3 \cdot 2H_2O + 2NaOH \longrightarrow 2NaAlO_2 + 3H_2O$
12. $NaAlO_2 + HCl + H_2O \longrightarrow Al(OH)_3 + NaCl$
13. $2Al(OH)_3 \xrightarrow{\text{Heat}} Al_2O_3 + 3H_2O$
14. $2CuS + 3O_2 \longrightarrow 2CuO + 2SO_2$
15. $3MnO_4 + 4Al \longrightarrow 2Al_2O_3 + 3Mn$
16. $Fe_2O_3 + 2Al \longrightarrow Al_2O_3 + 3Mn$
17. $CuS + 2CuO \longrightarrow 3Cu + SO_2$
18. $Si + O_2 \longrightarrow SiO_2$
19. $SiO_2 + C \longrightarrow Si + CO_2$
20. $Si + \underset{\text{(Steam)}}{2H_2O} \longrightarrow SiO_2 + 2H_2$
21. $Ca(OH)_2 + Cl \longrightarrow \underset{\text{Bleaching power}}{CaOCl_2} + H_2O$
22. $CaSO_4 \cdot 2H_2O \xrightarrow{373K} CaSO_4 \cdot \frac{1}{2}H_2O + \frac{3}{2}H_2O$
23. $S + \underset{\text{(conc.)}}{2H_2SO_4} \longrightarrow 3SO_2 + 2H_2O$
24. $H_2 + S \longrightarrow H_2S$
25. $2NaOH + SI + H_2O \longrightarrow Na_2SiO_3 + 2H_2$
26. $C + H_2O\ (\text{Steam}) \longrightarrow Na_2SiO_3 + 2H_2$
27. $P_4 + 3NaOH + 3H_2O$

    $\longrightarrow \underset{\text{Phosphine}}{PH_3} + \underset{\text{Sodium hypophosphite}}{3NaH_2PO_2}$
28. $\underset{\text{Cinnabar}}{2HgS} + 3O_2 \longrightarrow 2HgO + 2SO_2$
29. $P_4 + 6Cl_2 \longrightarrow 4PCl_3$

30. $P_4 + 10Cl_2 \longrightarrow 4PCl_5$

31. $2Mg + O_2 \longrightarrow 2MgO$

32. $Na_2O + H_2O \longrightarrow 2NaOH$

33. $K_2O + H_2O \longrightarrow 2KOH$

34. $CaO + H_2O \longrightarrow Ca(OH)_2$

35. $2Na + 2H_2O \longrightarrow 2NaOH + H_2$

36. $2K + 2H_2O \longrightarrow 2KOH + H_2$

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

38. $Mg + H_2O \xrightarrow{\text{boil}} MgO + H_2$

39. $Zn + H_2O \xrightarrow{\text{boil}} ZnO + H_2$

40. $3Fe + 4H_2O\ (\text{Steam}) \longrightarrow Fe_3O_4 + 4H_2$

41. $Zn + H_2SO_4\ (\text{dil}) \longrightarrow ZnSO_4 + H_2$

42. $2Na + 2HCI \longrightarrow 2NaCl + H_2$

43. $Cu + \underset{\text{Conc.}}{2H_2SO_4} \xrightarrow{\text{Heat}} CuSO_4 + 2H_2O + SO_2$

44. $Mg + 2HCl \longrightarrow MgCl_2 + H_2$

45. $2Al + 6HCl \longrightarrow 2AlCl_3 + 3H_2$

46. $Zn + CuSO_4 \longrightarrow ZnSO_4 + Cu$

47. $Fe + CuSO_4 \longrightarrow FeSO_4 + Cu$

48. $Cu + 2AgNO_3 \longrightarrow Cu(NO_3)_2 + 2Ag$

49 $Cu(s) + 2Ag^+ \longrightarrow Cu_2^+(aq) + 2Ag(s)$

50. $CaCO_3 + SiO_2 \longrightarrow CaSiO_3 + CO_2$

51 $CuCO_3 \xrightarrow{\text{Heat}} Cu + CO_2$

52. $ZnCO_3 \xrightarrow{\text{Heat}} ZnO + CO_2$

53. $2Na + H_2 \xrightarrow{\text{Heat}} 2NaH$

54. $NaH + H_2 \xrightarrow{\text{Heat}} NaOH + H_2$

55 $4Al + 3O_2 \longrightarrow 2AlP_3$

56. $Ca + Cl_2 \longrightarrow CaCl_2$

57. $Fe + S \xrightarrow{\text{Heat}} FeS$

58. $2Al + 3Cl_2 \longrightarrow 2AlCl_3$

59. $2Fe + 3Cl_2 \longrightarrow 2FeCl_3$

60. $Fe + H_2SO_4\ (\text{dil.}) \longrightarrow FeSO_4 + H_2$

61. $Fe + 2HCl\ (\text{dil.}) \longrightarrow FeCl_2 + H_2$

62. $Mg + H_2SO_4\ (\text{dil.}) \longrightarrow MgSO_4 + H_2$

63. $2Al + 2NaOH + 2H_2O \longrightarrow 2NaAlO_2 + 3H_2$

64. $Zn + S \xrightarrow{\text{heat}} ZnS$

65. $FeS + H_2SO_4\ (\text{dil.}) \longrightarrow FeSO_4 + H_2S$

66. $Cu + Cl_2 \longrightarrow CuCl_2$

67. $2Cu + O_z \longrightarrow 2CuO$

68. $CuO + H_2O \longrightarrow Cu(OH)_2$

69. $Cu(OH)_2 + CO_2 \longrightarrow CuCO_3 + H_2O$

70. $Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 + H_2O$

71. $2NaHCO_3 \xrightarrow{\Delta} Na_2CO_3 + H_2O + CO_2$

72. $2ZnS + 3O_2 \longrightarrow 2ZnO + 2SO_2$

73. $SnO_2 + C \longrightarrow Sn + CO_2$

74. $FeCO_3 \xrightarrow{\Delta} FeO + CO_2$

75. $Na_2CO_3 \cdot 10H_2O \xrightarrow{\Delta} Na_2CO_3 + 10H_2O$

76. $NH_3 + H_2O + CO_2 \longrightarrow (NH_4)HCO_3$

77. $(NH_4)HCO_3 + NaCl \longrightarrow NaHCO_3 + NH_4Cl$

78. $CO_2 + H_2O \longrightarrow H_2CO_3$

79. $SO_2 + H_2O \longrightarrow H_2SO_3$

80. $SO_3 + H_2O \longrightarrow H_2SO_4$

81. $S + 6HNO_3\ (\text{Conc.}) \longrightarrow H_2SO_4 + 6NO_2 + 2H_2O$

82. $P_2O_5 + 3H_2O \longrightarrow 2H_3PO_4$

83. $N_2O_5 + H_2O \longrightarrow 2HNO_3$

84. $Si + 2Cl_2 \xrightarrow{\text{heat}} SiCl_4$

85. $Si + 4HCl \longrightarrow SiCl_4 + 2H_2$

86. $P + 5HNO_3\ (\text{Cone.}) \longrightarrow H_3PO_4 + 5NO_2 + H_2O$

87. $C + 4HNO_3\ (\text{Cone.}) \longrightarrow CO_2 + 4NO_2 + 2H2O$

88. $Ca_3(PO_4)_2 + 3SiO_2 \longrightarrow 3CaSiO_3 + P_2O_5$

89. $2P_2O_5 + 10C \longrightarrow P_4 + 10\ CO$

90. $P_4 + 3O_2 \longrightarrow 2P_2O_3$

91. $P_4 + 5O_2 \longrightarrow 2P_2O_5$

92. $C + 2S \xrightarrow{\text{heat}} CS_2$

93. $Cu + \underset{\text{(Conc.)}}{4HNO_3} \longrightarrow Cu(NO_3)_2 + 2NO_2 + 2H_2O$

94. $3Zn + 8HNO_3(dil) \longrightarrow 3Zn(NO_3)_2 + 2NO + 4H_2O$

95. $Mg + 2HNO_3$ (5%) $\longrightarrow Mg(NO_3)_2 + H_2$

96. $Mn + 2HNO_3$ (5%) $\longrightarrow Mn(NO_3)_2 + H_2$

97. $4Zn + \underset{\text{(V.dil.)}}{10HNO_3} \longrightarrow 4Zn(NO_3)_2 + NH_4NO_3 + 3H_2O$

98. $Mg + ZnSO_4 \longrightarrow MgSO_4 + Zn$

99. $Mg + CuSO_4 \longrightarrow MgSO_4 + Cu$

100. $Mg + FeSO_4 \longrightarrow MgSO_4 + Fe$

➢ POINT TO REMEMBER

- Metals are lustrous, malleable, ductile and have high density.
- Metals are generally good conductors of heat and electricity.
- Metals form positive ions by losing electrons.
- All metals combine with oxygen to form metal oxides.
- Metal oxides are basic in nature.
- Metals react with chlorine to form electrovalent chlorides of metals.
- Metals like Li, Na and Ca form hydrides with hydrogen.
- There are metal-bearing substances below the earth's surface which are called minerals.
- The minerals from which metals can be obtained profitably are called ores.
- A flux is a substance which is added to the furnace charge to remove nonfusible impurities present in the ore.
- Flux combines with the nonfusible impurity to convert it into a fusible substance known as slag.
- Nonmetals are generally bad conductors of heat and electricity.
- All nonmetals are electronegative.
- The function of limestone in the extraction of iron is to provide calcium oxide (CaO) for the formation of the slag $CaSiO_3$.
- Alloy steels are prepared by adding a small quantity of nickel, cobalt, chromium, tungsten, molybdenum, manganese or silicon to steel.
- The heating of steel to redness and then cooling it slowly is called tempering of steel.
- Rusting of iron is an oxidation reaction which occurs in the presence of air and water.
- Aluminium is the most abundant metal in the earth's crust.
- The process of extracting metals from their ores and refining them for use is called metallurgy.
- The unwanted materials present in an ore are called gangue.
- Froth floatation is a method used for the concentration of some ores.
- Calcination is the process of heating an ore strongly so that volatile impurities are removed.
- Roasting is the process of heating an ore in a controlled supply of air at a controlled temperature.
- Smelting is the process of obtaining metals from their compounds.
- Bauxite is an ore from which aluminium metal is commercially obtained.
- The most important are of aluminium is bauxite ($Al_2O_3 \cdot 2H_2O$).
- Slow destruction of metals by the action of air, carbon dioxide, moisture, etc., is known as corrosion of metals.
- An alloy consists of two or more metals, or a metal and a nonmetal. Brass is an alloy of copper and zinc.
- The property of an element to exist in two or more ,... different forms is known as allotropy.

- The most important ore of copper is copper pyrite ($CuFeS_2$).
- Copper when heated in air at 300°C forms cupric oxide (CuO), while at 1000°C forms cuprous oxide (Cu_2O).
- Silver does not react with water and hydrochloric acid. It, however, reacts with nitric acid to produce NO_2 gas.
- Gold dissolves in aqua regia. Aqua regia is a mixture of conc. hydrochloric acid and conc. nitric acid in the volume ratio 3 : 1.
- Lead on rubbing leaves a mark on paper. Lead is the poorest conductor of heat.
- Zinc when heated with a concentrated solution of sodium hydroxide gives off hydrogen gas, while sodium zincate is left in solution.
- Sulphur is used in vulcanization of rubber.