7. METAL

OCCURRENCE OF METALS

Metals occur in nature, in the native (in free state) as well as in the combined state. Those occurring in the free state are very few e.g. gold, silver and platinum.

Whether a given metal occurs in nature in a native or combined state, it is always contaminated with impurities–mostly clays and siliceous matter. Naturally occurring materials containing metals are called minerals. A mineral from which a given metal is obtained economically is called on ore. The natural occurrence of some selected metals is given in Table .

Natural Occurrence of Selected Metals

Element	Nature of Ore	Mineral	Composition
	Sulphides		
Iron (Fe)		Iron pyrites	FeS_2
Copper (Cu)		Copper pyrites	$CuFeS_2$
Silver (Ag)		Argentite	Ag_2S
Zinc (Zn)		Zinc blende	ZnS
Mercury (Hg)		Cinnabar	Hgs
Lead (Pb)		Galena	PbS
	Oxides		
Iron (Fe)		Haematite	Fe_2O_3
		Magmetite	Fe_2O_4
Aluminium (Al)		Bauxite	$Al_2O_3.2H_2O$
Titanium (Ti)		Rutile	TiO_2
Manganese (Mn)		Pyrolusite	MnO_2
Copper (Cu)		Cuprite	Cu_2O
Tin (Sn)		Cassiterite	SnO_2
	Carbonates		
Magnesium (Mg)		Dolomite	$CaCO_3.MgCo_3$
		Magnesite	$MgCO_3$
Calcium (Ca)		Lime stone	$CaCO_3$
Zinc (Zn)		Calamine	$ZnCO_3$
	Halides		
Sodium (Na)		Rock salt	NaCl
Magnesium (Mg)		Camallite	$KCl.MgCl_2.6H_2O$
	Sulphates		
Calcium (Ca)		Gypsum	$CaSO_4.2H_2O$
Magnesium (Mg)		Epsom salt	$MgSO_4.7H_2O$

A variety of minerals are found in India, some in huge quantities and some in small.

Properties of Metals

Physical Properties Mercury is the only liquid metal, all other are solids. They have high melting and boiling points. They are malleable and ductile. They are quite lustrous and are good conductors of heat and electricity.

Chemical Properties Metals are electropositive and form positive ions by losing electrons. They combine with oxygen to form basic oxides. They react with dilute mineral acids to liberate hydrogen gas. With halogens they form halides. Their reactivity varies. The first group of elements, that is alkali metals are most reactive.

Mineral Wealth of India

State	Mineral
Andhra Pradesh	Diamonds, Limestone, asbestos, copper, mica, gold
Bihar	Asbestos, bauxite, chromite, coal, graphite, haematite, limestone, manganese ore, mica
Gujarat	Gypsum, manganese ore, fluorspar
Karnataka	Asbestos, gold, chromite, iron ore, magnesite, corundum
Kashmir and Himachal Pradesh	Gypsum
Kerala	Manazite, illmenite, rutile, garnet
Madhya Pradesh	Iron and manganess ores, bauxite, limestone, coal, dolomite
Maharashtra	Chromite and manganese ores
Orissa	Dolomite, limestone, graphite, haematite, chromite
Rajasthan	Copper, lead, manganese and zine ores, rock phosphate, barytes, gypsum, mica
Tamil Nadu	Iron and manganese ores, magnesite, mica, limestone, lignite
West Bengal	Coal, dolomite tungsten
Uttar Pradesh	Gypsum, magnesite, dolomite, rock phosphate

Extraction of Metals

Extraction of metals comprises the processing of natural raw materials to obtain elemental metals. The initial stages of this process involve three steps:

1. Concentration of the ore (or beneficiation)
2. Chemical or electrochemical reduction to the metal
3. Removal of impurities from the metal (i.e. purification or refining).

In certain cases these steps may overlap, for example, removal of impurities may be a part of the ore concentration process such that subsequent refining is not necessary.

Concentration of the Ore The ore usually contains rocky and siliceous matter called gangue. If the gangue is objectionable at a later stage, the first step in ore concentration is to remove it. To do so the ore is usually crushed and ground until the particles of the mineral are broken apart from the gangue. If possible, these particles are separated by physical means such as washing, magnetic attraction or froth floatation. These methods are based on differences between the physical properties of the mineral and the gangue.

If the ore cannot be sufficiently concentrated by these physical methods, chemical processes are used. Some examples are given below:

(i) In many cases the ore is roasted (heating in air) to drive off volatile impurities, burn off organic matter, and to from compounds that are more easily smelted. Roasting in air usually converts sulphides to oxides.

(ii) Concentration of bauxite ($A1_2 O_3. 2H_2O$) by hot aqueous sodium hydroxide. This method is based on the amphoteric nature of aluminium hydroxide. The crushed ore is treated with a hot aqueous solution of NaOH to dissolved alumina, $A1_2O_3$ or bauxite and impurities such as iron oxide and silicates remain unaffected. The solution is filtered, cooled and CO_2 gas is blown when aluminium hydroxide gets precipitated. On heating aluminium hydroxide, pure alumina is obtained.

(iii) Leaching of gold and silver ores with aqueous sodium cyanide in the presence of air to yield the metal cyano complexes.

(iv) Concentration of magnesium ions from sea water by calcium hydroxide. $Mg\,(OH)_2$ is converted into anhydrous $MgCl_2$ which is used for the electrolytic reduction for obtaining magnesium.

Calcination and Roasting When the ore has been sufficiently concentrated for the isolation of a metal, it is subjected either to calcination or roasting process depending on the nature of the ore. In these operations, volatile impurities are also removed.

Calcination is the heating of the ore in the absence of air. This method is employed for obtaining the metal oxides from carbonates and hydroxides.

Roasting is the heating of the ore in the presence of air. On roasting, part of the ore is oxidized to form an oxide. This oxide is then reduced to the metal.

Conversion of the ore into the oxide form is an important step in the final reduction to the metal. The reducing agent used should be easily available and cheap, the obvious choice is carbon.

Smelting The industrial reduction process for obtaining metal from the treated ore is called smelting. There are several chemical and electrochemical methods available which can be used to produce metals.

Ores generally contain considerable gangue even after concentration. To remove the last amounts of the gangue, a flux is added during the smelting process. Flux is a substance that combines with gangue and forms a molten material called slag as the mixture is heated in a furnace. At high temperatures, the slag is a liquid that is insoluble in the molten metal and it forms a separate layer. If the gangue is an acidic oxide, such as silica (SiO_2), a cheap basic oxide like quicklime (CaO) may be used as flux. Impurities and flux react in a furnace to form slag. If on the other hand, the gangue is basic, the flux to be added would be acidic.

Refining of Metals The purification or adjustment of the composition of impurities in crude metals is called refining. Metals with low boiling points, such as mercury, zinc and magnesium can be separated from most impurities by simple distillation. Like salts, metals can also be refined by fractional crystallization. Probably the most widely used refining method is the electrolytic method. Among modern purification methods, Zone refining, chromatographic, ion-exchanges and solvent extraction methods are also used for specific purpose.

COPPER

Symbol Cu **Atomic No.29**

Electronic *configuration* of copper

The atomic number of copper is 29. Thus, an atom of copper contains 29 electrons. The electronic configuration of copper is :

	K	L	M	N
Cu (29)	2	8	18	1

Copper atom has only one electron in its outermost shell.

Position in the periodic table.

Copper is a transition metal. It is placed in group IB along with silver and gold.

VIII			IB	IIB
Fe	Co	Ni	Cu	Zn
(26)	(27)	(28)	(29)	(30)
Ru	Rh	Pd	Ag	Cd
(44)	(45)	(46)	(47)	(48)
Os	Ir	Pt	Au	Hg
(76)	(77)	(78)	(79)	(80)

Being a member of the same subgroup **(IB),** copper shows some marked similarities in properties with those of silver and gold. Further, the properties of subgroup **(IB)** metals are entirely different from those of subgroup **(IA)** metals. Hence, the subdivision into **(IA)** and **(IB)** is justified.

Occurrence of copper

Copper is an unreactive metal. Hence, it occurs in\ nature in native or free state. It is also found in combined state in the form of sulphide, oxide, carbonate, etc. The important ores of copper are:

(i) Sulphide ores

Copper glance (chalcocite), Cu_2S Copper pyrites (chalcopyrite), $CuFeS_2$

(ii) Oxide ores Cuprite, Cu_2O

(iii) Carbonate ores Malachite, $CuCO_3.Cu(OH)_2$

Azurite, $2CuCO_3.Cu(OH)_2$

The most important ore of copper is copper pyrites ($CuFeS_2$).

Copper in India

In India, copper is found in Sikkim, Singhbhum in Bihar, Nellore in Andhra Pradesh and Khetri in Rajasthan. Copper is nowadays extracted through a modern process by The Indian Copper Corporation Ltd., Musabani, Ghatsila.

Extraction of copper from copper pyrites

Copper is obtained chiefly from the *pyrite ore.* This ore contains different materials. Hence, it needs to be concentrated or dressed.

1. **Dressing of the ore** The dressing of the ore is done by froth floatation process.

2. **Roasting** The dressed or the concentrated ore is roasted on the bed of a reverbveratory furnace in a current of air. As a result of this operation, a portion of sulphur burns to form SO_2 and the metals are partially oxidized. The residue *contains oxides and sulphides of copper and iron.*

$$2CuFeS_2 + O_2 \rightarrow Cu_2S + 2FeS + SO_2$$

3. **Smelting** The roasted ore is mixed with a little coke and sand. The mixture is heated strongly in a blast furnace. Most of the ferrous sulphide is converted into ferrous oxide. The ferrous oxide forms a slag with the sand and is removed. The residue is almost pure copper sulphide and is called the *white metal.*

$$2FeS + 3O_2 \rightarrow 2FeO + 2SO_2$$

$FeO + SiO_2$ ® $FeSiO_3$ (ferrous silicate)

slag

4. **Preparation of blister copper** The white metal is carefully heated in a reverberatory furnace where a portion of Cu_2S is oxidized to Cu_2O. As the temperature rises, Cu_2O reacts with the remaining portion of Cu_2S to form copper and SO_2.

$$2Cu_2O + Cu_2S \rightarrow 6Cu + SO_2$$

Due to the evolution of $S0_2$, copper obtained has a blister-like appearance. This copper is called *blister copper.*

5. **Refining of blister copper** Blister copper contains iron, sulphur, arsenic, lead, etc., as impurities. It is melted in a reverberatory furnace in a current of air. The impurities present are oxidized away as gases or they form **a** slag with **silica of** the furnace. The slag is removed. Anthracite coal is sprinkled on the surface. The molten mass is stirred with poles of green wood. Pure copper is obtained.

Very pure copper is, however, obtained by electrolysis. Sheets of impure copper are made the anode while thin plates of pure copper are made the cathode. A solution of copper sulphate is used as the electrolytic solution. On passing electric current, sheets of impure copper gradually dissolve. Pure copper is deposited on the cathode. The impurities settle down at the bottom or remain in the solution.

$$Cu^{2+} + 2e \rightarrow Cu \text{ (at cathode)}$$

$$Cu \rightarrow Cu^{2+} + 2e \text{ (at anode)}$$

At the cathode, Cu^{2+} ions are taken from the copper sulphate solution while Cu^{2+} ions go into the solution at the anode. In this way, pure copper is transferred from the anode to the cathode.

Properties of Copper

Physical properties

Copper is a reddish, shining metal. It melts at 1083°C. Its density is 8.95 g/cm^3. It is ductile and malleable. It is a good conductor of heat and electricity.

Chemical properties

1. **Valency** Copper exhibits variable valencies. The valencies of copper are 1 and 2. Thus, copper forms *monovalent ions* (Cu^{2+}) *and divalent ions (Cu) . The mono- valent ion (Cu) is* Known as cuprous ion and the divalent ion (Cu^{2+}) is known as *cupric* ion. Compounds corresponding to cuprous and cupric ions are called cuprous compounds and cupric compounds respectively .The cupric ion (Cu^{2+}) is more stable than cuprous (Cu^+).

An atom of copper does not easily lose its electrons to form ions. Thus, the reactivity of copper is very low and it is treated as an uncreative metal. Copper lies very low in the activity series of metals.

2. Action of air Dry air has no action on copper. When the metal is exposed to moist air, its surface gradually gets covered with a green layer of basic copper carbonate $CuCO_3\ Cu(OH)_2$. When copper is heated in air at about **300°C it** combines with oxygen to form black cupric oxide.

300°C

$$2Cu + 0_2 \rightarrow 2CuO$$

At a temperature of about 1000°C, cuprous oxide is formed.

1000°C

$$4Cu + O^2 \rightarrow 2Cu_2O$$

3. Action of water Copper does not react with water or steam.

4. Action of acids Copper lies below hydrogen in the electrochemical series. Hence, it cannot displace hydrogen from acids.

Copper, however reacts with boiling concentrated HCl in presence of air or some oxidizing agent, with the evolution of hydrogen.

$$2Cu + 2HCl \rightarrow 2CuCl + H_2$$

Copper does not react with dilute hydrochloric dilute sulphuric acid.

$$CuCl + Cl^- \rightarrow [CuCl_2]^-$$

This is the only reaction in which copper displaces H_2 from HCl.

Copper reacts with dilute H_2SO_4 only in the presence of air:

$$2Cu + 2H_2SO_4 + O_2 \rightarrow 2CuSO_4 + 2H_2O$$

When copper is heated with concentrated sulphruic acid, copper sulphate, sulphur dioxide and water are formed.

$$Cu + 2H_2SO_4 \rightarrow CuSO_4 + SO_2 + 2H_2O$$

With dilute nitric acid copper reacts to form nitrate and nitric oxide gas.

$$3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$$

5. Action of chlorine When copper is heated in a current of chlorine, cupric chloride is produced.

$$Cu + Cl_2 \rightarrow CuCl_2$$

6. **Action of sulphur** When copper is heated with sulphur, cupric sulphide is formed.

$$Cu + S \rightarrow CuS$$

7. Displacement of metals Copper displaces some such as silver, gold, platinum, etc., from the solutions of their salts.

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

It should be noted that only those metals are displaced from their salt solutions which lie below copper in the elecrochemical series.

8. Action of aqueous ammonia Copper dissolves in aqueous ammonia in the presence of air, forming complex compound, tetra ammine cooper (II) hydroxide.

$$Cu + H_2O + 1/2O_2 \rightarrow Cu^{2+} + 2OH^-$$

$$Cu^{2+} + 4NH^3 \rightarrow [Cu(NH_3)_4]^{2+}$$

$$[Cu(NH_3)_4]^{2+}{}_{+} 2OH^- \rightarrow [Cu(NH_3)_4]\text{-}(OH)_2$$

Test for copper (II) salts or cupric ions (Cu^{2+})

When a copper (II) salt solution or a solution containing Cu^{2+}ions is treated with ammonium hydroxide, a pale blue precipitate of copper (II) hydroxide is obtained , On adding excess of ammonium hydroxide, the precipitate dissolves to produce a deep blue solution.

Uses of copper

1. Copper is largely used in electrical industry for making electric wires, electric motors, dynamos and several other electrical goods.
2. Copper is used in electroplating and electrotyping.
3. Copper is also used in making household utensils.
4. An important use of copper is in making alloys with metals. Some of the alloys and their compositions; given in Table.

Table Alloys of copper

	Alloy	Composition	Uses
(i)	Brass	Cu= 80%, Zn=20%	Utensils, cartridges, condenser tubes, etc
(ii)	Bronze	Cu= 80%, Zn=10%, Sn= 10%	Utensils, statues, coins, etc.
(iii)	German silver	Cu= 60%, Zn=20%, Ni= 20%	Utensils, resistance , coils, etc
(iv)	Bell metal	Cu= 80%, Sn=20%	Bells , etc.
(v)	Gun metal	Cu= 87%, Sn=10% , Zn= 3%	Bearings, machine parts Formerly use for cannons.

IRON

Symbol: Fe **Atomic no. 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:

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

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

Position of iron in the periodic table

The electronic configuration of iron shows that there are two electrons in its outermost shell or the valence shell. So, it can be reasonably expected that iron should have a place in group II of the periodic table. But it is not so. Since iron is a transition metal, its properties are different from the normal elements of group II. In fact, iron has been assigned a place in group VIII of the periodic table. It is a member of the first transition series in the fourth period known as a triad.

Group VIII

Fourth period	Fe	Co	Ni
Fifth period	Ru	Rh	Pd
Sixth period	Os	Ir	Pt

Occurrence of iron

Iron is the second most abundant metal in the earth's crust, the first one being aluminium. Free iron has been found in most meteorites. It is also believed that a core of iron is present in the centre of the earth.

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 on account of its high sulphur content.

Iron in India

In India, large deposits of iron ores are found in Bihar, Orissa, Madhya Pradesh, Tamilnadu and Sikkim. The important iron and steel plants are located at Bhillai, Bokaro, Jamshedpur, Rourkela, Durgapur, Asansole and Bhadravati.

Extraction of iron 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 the blast furnace. The following reactions occur in the blast furnace:

(i) As the charge comes down to the 600°C region, the iron oxide is reduced by the ascending carbon monoxide gas produced by the burning of coke.

$$2C + 0_2 \rightarrow 2CO$$

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

The iron thus obtained is called *spongy iron.*

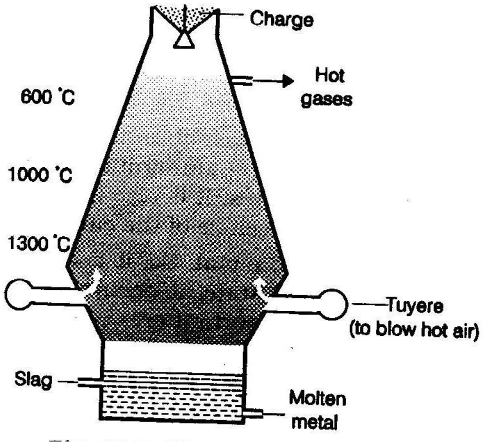
(ii) At the 1000°C region, the silica is converted to slag.

$$CaC0_3 \rightarrow CaO + CO_2$$

$$CaO + SiO_2 \;\; CaSiO_3$$

slag

(iii) At the 1300°C region, spongy 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.

$CaCO_3$ ® $CaO + CO_2$

quicklime

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

$CaO + SiO_2$ ® $CaSiO_3$

slag

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

2. **Wrought iron** It is almost pure form of iron. It contains only 0.*12* to 0.25% carbon. It melts at a higher temperature (1500°C) than that at which cast Iron melts. It is obtained by melting cast iron on a hearth lined with ferric oxide (Fe_2--O_3). The impurities, such as carbon, phosphorus, silicon and manganese, are oxidized by Fe_2O_3.

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

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, wires, anchors and cores of electromagnets.

3. **Steel** It is an alloy of iron and carbon. It contains about 0.1 to 0.15% of carbon. There are different types of steel.

(a) **Mild steel** It contains about 0.15% carbon. It is used for making sheets and wires.

(b) **Hard steel** It contains higher percentage (1.5%) 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 0.1% carbon. It is hard and is used in making rails, bridges, etc.

Table: Typical percentage compositions of the three types of iron

	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	0.12-0.25%	0.03-2.0%	0.03-2.0%	0.04-0.04%	0.1-4%
Mild steel	0.15%	0.03%	0.05%	0-0.05%	0-50%

Tempering

The hardness 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 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 *called* annealing.

Quenching of steel

Hard steel is heated to a high temperature. It is then suddenly cooled by plunging it into oil or water. Steel becomes as hard and brittle as glass. This is known as *quenched steel,* and the process of obtaining such a steel is known as *quenching or hardening of steel*

Properties of Iron

Physical properties

Pure iron has a grey color. It is malleable and ductile, conductor of heat and electricity. It melts at 1525°C and has a density of 7.86 g/cm^3.

Chemical Properties

1. **Valency** Iron shows variable valency. It shows valancies of 2 and 3, i.e., it forms divalent ion (Fe^{2+}) as well trivalent ion (Fe^{3+}). The compounds in which iron shows are known as ferrous compounds, whereas the compounds in which iron shows trivalent 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. So, 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 ($2Fe_2O_3.3H_2O$).

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

$3Fe +$	$4H_2O \rightarrow$	Fe_3O_4	$+ 4H_2$
Iron	steam	ferrosoferric oxide	hydrogen

4. **Action of acids** Iron lies above hydrogen in the chemical series. So, it can displace hydrogen from the 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.

$$Fe + 2HC1 \rightarrow FeCl_2 + H_2$$

(b) Concentrated hydrochloric acid also evolves 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 oxide. There is no formation of hydrogen gas.

$$Fe + 2H_2SO_4 \rightarrow FeSO_4 + SO\text{-}_2 + 2H_2O$$

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

$$2FeSO_4 + 2H_2SO_4 \rightarrow Fe_2(SO_4)_3 + 2H_2O + SO_2$$

(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.

$$2Fe + 3C1_2 \rightarrow 2FeCl_3$$

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 in 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.

$$Fe + CuSO_4 \rightarrow Cu + FeS0_4$$

or

$$Fe + Cu^{2+} \rightarrow Cu + Fe^{2+}$$

Tests to distinguish between ferrous and ferric salts

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

$$FeSO_4 + 2NaOH \rightarrow Fe(OH)_2 + Na_2SO_4$$

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

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

2. 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 $Fe(OH)_3$, by water and atmospheric oxygen, is known as rusting.

Rusting of iron is an oxidation reaction which 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 due to 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 equipment.
2. Wrought Iron and cast iron are largely used in the manufacture of locomotives, railway lines, springs, tubes, etc.

1\. Iron finds wide application in house construction, e.g., in the reinforcement of roofs and other parts of buildings.

ALUMINIUM

Symbol Al **Atomic no. 13**

History Aluminium was first prepared in 1827. Impressed by its properties, scientists all over the world attempted to develop a process for the commercial production of the metal. 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. Since then, large-scale production of aluminium began in several parts of the world. It came into large-scale use towards the end of the nineteenth century.

Electronic configuration of aluminium

The atomic number of aluminium is 13. This means that an atom of aluminium contains 13 electrons in its shells. The electronic configuration is then,

	K	L	M
A1 (13)	2	8	3

Thus, the outermost shell or valence shell of an aluminium atom has 3 electrons.

Position in the periodic table

	IA	IIA	IIIA
First period	H (1)		
Second period	Li (3)	Be (4)	B(5)
Third period	Na (11)	Mg (12)	Al (13)
Fourth period	K(19)	Ca (20)	Ga (31)

Occurrence of aluminium

Aluminium comes third amongst the most abundant elements in nature. But aluminium is the most abundant metal in the earth's crust. It is a reactive metal, hence does not occur native or free in nature. In the state, it is widely distributed in nature in the form of compounds. Some of its important minerals are:

i. Cryolite, $Na3AlF\text{-}_6$

ii. Corundum, $A1_2O_3$

iii. Bauxite, $A1_2O_3.2H_2O$

iv. Feldspar or *Felspar,* $KAlSi_3O_8$

However, bauxite is the only mineral from which aluminium is obtained profitably. Hence, bauxite is an ore of aluminum.

In India, Hindustan Aluminium Corporation, Indian Aluminium Company and the Bharat Aluminium Corporation produce aluminium from bauxite.

Extraction of aluminium from bauxite

The extraction of aluminium from bauxite consists of the following steps:

1. **Preparation of alumina** Crude bauxite is finely powdered and treated with a concentrated caustic soda at 150°C. The bauxite goes into forming sodium aluminate while the behind undissolved.

$$Al_2O_3.2H_2O + 2NaOH \rightarrow 2NaAlO_2 + 3H_2O$$

bauxite **sodium** **sodium** **water**

hydroxide aluminate

The solution is filtered. The filtrate is diluted with water and a little freshly precipitated aluminium hydroxide is added to cause precipitation. Sodium aluminate gets hydrolyzed to produce a white precipitate of aluminium hydroxide.

$$NaAlO\text{-}_2 + 2H_2O \rightarrow Al(OH)_3 + NaOH$$

Sodium water aluminium sodium

Aluminate hydroxide hydroxide

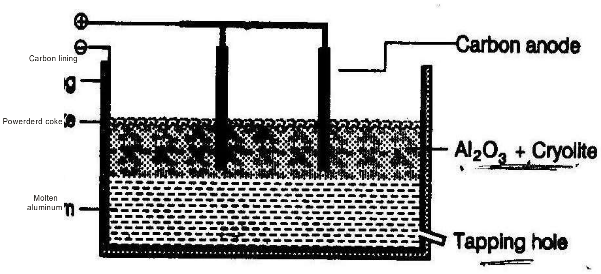
The precipitate of $Al(OH)_3$ is filtered off, washed, dried and ignited to get pure aluminium oxide.

$$2Al(OH)_3 \rightarrow Al_2O_3 + 3H_2O$$

Aluminium aluminium water

Hydroxide oxide

1. **Electrolytic reduction of alumina** Alumina is mixed with cryolite. The mixture is melted in an iron cell.



Alumina melts at 200°C. 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 900-950°C. 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, Al^{3+} ions are discharged at the cathode and the O^{2-} ions at the anode.

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

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

Some quantity of oxygen formed escapes and some reacts with the anode to form CO_2.

$c + o_2 \rightarrow co_2$

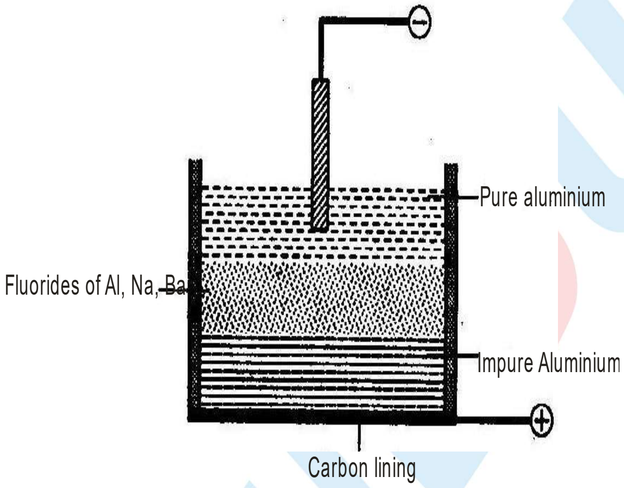
Hence, the anode has to be replaced from time to time. This increases the cost of production of aluminum. Molten aluminium sinks to the bottom, and is taken out from there.

3. **Electrolytic refining of aluminium** Aluminium thus obtained is 99% pure. It contains iron and silicon as impurities. It is purified by further electrolysis using *Hope's method.* The electrolysis is carried out in an iron box lined with carbon. Three layers, differing in densities, are placed in the box.

Bottom layer It consists of impure aluminium. It acts anode.

Middle layer It consists of a mixture of the fluorides of Al, Na and Ba. It serves as electrolyte.

Upper layer It consists of pure molten aluminium and serves as cathode.



On passing electric current, aluminium ions from the middle layer are discharged at the cathode. An equivalent amount of aluminium goes into the middle layer from the bottom layer. The impurities are left behind. Oxygen is liberated at the anode. Oxygen reacts with the carbon anode. Hence, the anode is gradually consumed.

Properties of Aluminium

Physical properties

Aluminium possesses a bluish white lustre. It is a light metal of density 2.7. It is malleable and ductile. It melts at 660°C. It is a good conductor of heat and electricity.

Chemical properties

1. **Valency** The outermost shell or the valence shell of aluminium contains 3 electrons. These three electrons can be given up to form aluminium ion, Al^{3+}. Thus, the valency of aluminium is 3, i.e., aluminium is a trivalent metal.

2. **Action of air** Aluminium is a very reactive metal. Yet it does not get corroded when exposed to air. Dry air has no action on aluminium. In moist air, a thin film of its oxide is formed over its surface. This prevents the metal from further attack by air or water or even by dilute acids. The film of the oxide can be removed by rubbing the surface of the metal with mercury. Then rapid corrosion by moist air occurs. The oxide film may also be removed by sodium chloride. That is why aluminium rusts in coastal areas.

   When heated to 800°C in air, aluminium burns with a brilliant light. In this reaction, aluminium oxide and a little aluminium nitride are formed.

$4Al + 3O_2 \rightarrow 2\,Al_2O_3 + Heat$

aluminium oxygen aluminium oxide

$2Al + N_2$ ® $2AlN$

Aluminium nitrogen aluminium nitride

3. **Action of water** Aluminium is a reactive metal. But it does not react with water. This is because the surface of aluminium remains covered with a thin film of unreactive oxide. The thin oxide film does not allow water to come in contact with aluminium.

   If the thin oxide film is removed by rubbing with a sand paper then aluminium reacts with boiling water to produce aluminium hydroxide and hydrogen.

$2A1 + 6H_2O \rightarrow 2Al(OH)_3 + 3H_2$

aluminium water aluminium hydroxide hydrogen

4. **Action of acids** Aluminium reacts with acids to produce aluminium salts and hydrogen.

(i) **Reaction with hydrochloric acid** Aluminium dissolves in hydrochloric acid to form aluminium chloride and hydrogen.

$2Al + 6HCl$ - $2AlCl_3 + 3H_2$

aluminium hydrochloric aluminium chloride hydrogen

The reaction is slow with dilute hydrochloric acid but rapid with concentrated hydrochloric acid.

(ii) **Reaction with sulphuric acid** Aluminium dissolves in dilute sulphuric acid to form aluminium sulphate and hydrogen. The reaction is, however, slow.

$$2Al + 3H_2SO_4 \rightarrow Al_2(SO_4)_3 + 3H_2$$

Aluminium dissolves in hot concentrated sulphuric *acid to form aluminium sulphate with the evolution of* sulphur dioxide gas.

$2Al + 6H_2SO_4 \rightarrow Al_2(SO_4)_3 + 6H_2O + 3SO_2$

aluminium sulphuric acid aluminium sulphate water sulphur dioxide

(iii) **Reaction with nitric acid** In dilute or concentrated nitric acid, aluminium is rendered passive due to the formation of an insoluble layer of oxide. The oxide layer does not permit further reaction.

(iv) **Action of alkalis** Aluminium dissolves in caustic soda or caustic potash solution with the evolution of hydrogen, forming soluble sodium or potassium aluminate respectively.

$$2Al + 2NaOH + 2H_2O \rightarrow 2NaAlO_2 + 3H_2$$

$$2Al + 2KOH + 2H_2O \rightarrow 2KaAlO_2 + 3H_2$$

On account of these reactions, aluminum containers cannot be used to store alkali solutions.

(v). **Action of chlorine** When chlorine gas is passed over heated finely powdered aluminum, aluminium chloride is produced.

$$2Al + 3Cl_2 \rightarrow 2AlCl_3$$

(vi) **Reducing action** Aluminum is a powerful reducing agent. It reduces CO and CO_2 to Carbon .

$$3CO + 2Al \rightarrow Al_2 + 3C$$

$$3CO_2 + 4Al \rightarrow 2Al_2O_3 + 3C$$

It reduces many metal oxides to metals. For example, it reduces ferric oxide (Fe_2O_3) to iron metal.

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

A large amount of heat is liberated in this reaction. As a result, temperature rises to about 3500^0C.

Tests for aluminium salts

1. **Charcoal test** A solid aluminium salt is heated the cavity of a charcoal block. A white residue is obtained. When this residue is moistened with cobalt nitrate solution, a blue mass is obtained.
2. When the aqueous solution of an aluminium salt is treated with ammonium hydroxide, a white gelatinous precipitate of aluminium hydroxide is produced.

$$AlCl_3 + 3NH_4OH \rightarrow Al(OH)_3 + 3NH_4Cl$$

Aluminium chloride ammonium hydroxide aluminium hydroxide ammonium chloride

Uses of aluminium

1. Aluminium is a good conductor of heat and it does not get corroded. Hence, it is widely used to make utensils like pans, kettles, etc.
2. Aluminium is a good conductor of electricity; it is extensively used in making electrical wires.
3. Aluminium foils are used for wrapping foodstuffs.
4. Aluminium is a light metal. Hence, it is used in making bodies and parts of aircraft, buses furniture, etc. Since pure aluminium is not very strong, alloys of aluminium are used for these purposes.
5. Aluminium powder is used in anti-corrosion paints and in explosives.
6. Aluminium powder is also used in alumino-thermic process for welding two iron objects.

USES OF METALS

From time immemorial, metals have been in the service of man for various purposes. We know has been using metals for the manufacture of tools and weapons. With the passage of time, the use became widespread.

Metals are used in three different forms:

1. As pure metals
2. As alloys
3. As compounds of metals

Uses of pure metals

1. Metals are good conductors of electricity. copper is used to make good quality electric wires. Aluminium is also used to make electric wires which are cheaper than copper wires.
2. Iron, aluminium and copper are used for making cooking utensils and factory equipment.
3. Mercury is used in thermometers and in the action of some medicines.
4. Silver and gold are used in jewellery.
5. Thin foils of aluminium are used for packaging
6. Iron, in the form of steel, wrought iron and cast iron is used for the manufacture of locomotives, railway lines, implements of war, springs, tubes, cranes, etc.
7. Gold, silver, copper, nickel, etc., are being used for coinage all over the world.
8. Zinc and tin are used for coating iron objects to prevent rusting.
9. Cadmium, titanium and zirconium are used in atomic energy and space research projects.
10. Nowadays, titanium is used in aerospace and aircraft industries for making aircraft frames and engines where its high strength to weight ratio is useful, in the manufacture of military hardware, and in chemical reactors and industries. It finds application in making marine instruments. It is also used for hardening of steel.

The large-scale use of titanium is due to (i) its high strength to weight ratio, (ii) its resistance to corrosion and (iii) its high melting point. Due to its specific (especially military and aerospace) application called a strategic element.

Uses of alloys

An alloy is a mixture of a metal with other metals or nonmetal. It has properties different from those of its constituents. The properties of a metal are much more improved and pronounced when the metal is transformed into an alloy. For example, aluminium is a light metal and it is not very strong. But an alloy of aluminium (duralumin) is light and very strong. Similarly, silver is soft, but a silver alloy is hard and more useful than pure silver.

In order to prepare an alloy we mix up various metals molten state. The molten mass is then cooled to room temperature. The solid substance formed is an alloy.

Some of the commonly used alloys, their composition various uses, are listed in Table.

	Alloy	**Composition**	**Uses**
1.	Brass	Copper = 80% zinc= 20%	Brass is harder than pure copper. It is used for making utensils, cartridges, etc.
2.	Bronze	Copper= 80%Zinc = 10%tin= 10%	It is used for making statues, medals, ships, coins, and machines,
3.	Solder	tin= 50%lead = 50%	It is used for joining metals, soldiering wires, electronic components. Etc.
4.	Duralumin	Aluminium= 94.4%silicon = 4%and Copper, Magnesium, manganese .	It is used for making bodies of aircraft and kitchenware
5.	German Silver	Copper = 60%Zinc=20%Nickel= 20%	It is used for making utensils, ornaments, etc.
6.	Monel metal	Nickel=70%Copper=30%Iron=2%	
7.	Nichrome	Nickel =67.5%Chromium=15%Iron =16% Manganese=1.5%	It is used for making coils of electric furnaces.
8.	Gun metal	CopperTinZinc	It is used for making bearing, valves, etc.
9.	Rose metal	SiliconLeadTin	It is used for making fuse wires.

21.5 SOME USEFUL COMPOUNDS OF METALS

Common Salt. Common salt is sodium chloride (NaCl). Common salt when kept open absorbs moisture from air. This is because of the presence of magnesium chloride in common salt. Magnesium chloride is deliquescent, that is, it absorbs water.

Sodium chloride is very important chemical. It is an essential ingredient in our daily meal. It can be converted into useful chemicals like caustic soda, washing soda and baking soda. It also yields hydrochloric acid, an important chemical used in chemical industry, and chlorine which is used as a bleaching agent.

Caustic Soda. Chemically caustic soda is sodium hydroxide (NaOH). It is obtained by electrolysis of brine (NaCI solution).

Sodium hydroxide is a deliquescent solid. On dissolving in water it gives a strong alkaline solution.

Washing Soda. Washing soda is sodium carbonate decahydrate Na2CO3 .10H_2O. Sodium carbonate (soda ash) is alkaline in nature and has detergent (cleaning) properties. In a crystalline state it possesses water of crystallization. On exposure to air, crystals lose water and turn into white opaque powder.

$$Na_2CO_3 . 10H_2O \rightarrow Na_2CO_3 . H_2O + 9H_2O$$

Such an action, that is the giving up of water of crystallization to the atmosphere is termed as efflorescence. Sodium carbonate is used as a cleansing agent for domestic purposes. It is also used for softening of hard water, in manufacturing of chemicals like borax, caustic soda, glass and water glass. It is a constituent of many dry soap powders.

Baking Soda. The chemical name of baking soda is sodium hydrogen carbonate (NaHCO3). Sodium bicarbonate is used in the manufacture of baking powders and health salts. In addition to sodium bicarbonate, baking powders contain tartaric acid (or a similar Compound.) When heated sodium bicarbonate reacts with acid producing carbon dioxide which causes cake to rise and become light.

Lime. Calcium oxide is commonly known as quicklime or just lime. It is prepared by heating lime stone ($CaCO_3$) in a kiln.

$$CaCO_3 \rightarrow CaO + CO_2$$

Calcium oxide reacts vigorously with water producing slaked lime [Ca(OH)2] which is used for making mortar and for white washing.

A solution of slaked lime is known as lime water which is used for testing carbon dioxide. Calcium oxide is used for making cement and glass.

Bleaching Powder. The chemical name of bleaching powder is calcium oxychloride ($CaOCl_2$). Bleaching powder is prepared by passing chlorine over solid slaked lime.

$$Ca(OH)_2 + Cl_2 \rightarrow \underset{\text{Bleaching powder}}{CaOCl_2} + H_2O$$

When exposed to air, bleaching powder deteriorates giving off chlorine.

Bleaching powder is commonly used for bleaching in laundries. It is used in paper and textile industries, for disinfecting water and for preparation of chloroform.

Plaster of Paris. It is a calcium sulphate hemihydrate $(CaSO\text{-}_4)_2\ H_2O$. It is prepared by heating gypsum ($CaSO_4$ $2H_2O$).

$$2(CaSO_4 .\ 2H_2O) \rightarrow (CaSO_4)_2\ H_2O + 3H_2O$$

Plaster of paris has a remarkable property of hardening when mixed with proper quantity of water. The hardening of plaster of paris is due to its hydration to form gypsum. It is used in the laboratories for sealing the gaps, where sunlight arrangement is required. It is also used for making casts for statues, and in surgery to maintain joints in a fixed position