

Element are classified into two basic categories, metals and non-metals.



Before giving the definition of metal and non-metal, we must know some important terms:

- (i) **Malleability:** Ability of an element due to which it can be beaten with hammer into thin sheets is called malleability.
- (ii) **Ductility:** Ability of an element due to which it can be drawn into wires is called ductility.
- (iii) **Sonorous:** Property of an element to produce sound when it is struck with a hard substance.
- (iv) Brittleness: Property of an element to break easily into pieces on hammering or stretching.

1. METALS AND NON-METALS

Keeping these terms in mind, now we can define metal and non-metals.

Metals are defined as those elements which possess lusture, are malleable and ductile and good conductor of heat and electricity.

For e.g. sodium, magnesium, potassium, aluminium, copper, silver, gold etc.

Non-metals are defined as those elements which do not possess lusture and are neither good conductors of heat and electricity nor malleable. They are not ductile but are brittle. For e.g. carbon, hydrogen, oxygen, nitrogen, sulphur, bromine etc.

1.1 PHYSICAL PROPERTIES OF METALS

- (i) Metals are malleable i.e. they can be beaten into thin sheets with a hammer e.g. gold, silver, aluminium and copper.
- (ii) Metals are ductile i.e. they can be drawn into thin wires when hammered. Gold and silver are among the best ductile metals.
- (iii) Metals are good conductors of heat i.e. metals allow heat to pass through them easily. When a metal is heated, its atoms gain energy and vibrate more vigorously. They transfer energy to other electrons and atoms of the metal and heat is conducted. Silver is the best conductor of heat. Copper and aluminium are also good conductors of heat. Lead and mercury are poor conductors of heat.
- (iv) Metals are good conductors of electricity. This is because they contain free electrons. These free electrons can move easily through the metal and conduct electric current. Silver metal is the best conductor of electricity. Copper, gold and aluminium are also good conductors.
- (v) Metals are shiny. They can be polished. On exposure to air, metals lose their shine due to formation of oxides, sulphides, carbonates etc by reaction with gases present in air.
- (vi) Metals are generally hard. However, sodium and potassium are soft metals which can easily be cut with a knife.





Sodium can be easily cut with a knife

- (vii) Metals are strong and have high tensile strength i.e. heavy weight can be suspended from metallic wires without breaking the wire.
- (viii) Metals have high densities except sodium and potassium.
- (ix) Metals are sonorous i.e. they produce sound when hit with some solid object.
- (x) Metals are solid at room temperature except mercury which is liquid. Gallium and caesium have very low melting point. These two melt when held in hand.



Liquid Metal : Mercury

(xi) Metals have high melting and boiling points except sodium and potassium.

1.2 PHYSICAL PROPERTIES OF NON-METALS

- (i) Non-metals are non-malleable and non-ductile. Non-metals are brittle. The property of being brittle (breaking easily) when hammered is called brittleness.
- (ii) Non-metals are bad conductors of heat and electricity. Carbon (graphite) is an exception. It is conductor of electricity and is used in making electrodes.
- (iii) Non-metals are non-lustrous and dull except graphite and iodine.
- (iv) Non-metals are generally soft except diamond, allotrope of carbon, which is the hardest natural substance.





Diamond is lustrous and hard non-metal

- (v) Non-metals have low tensile strength and they are not strong.
- (vi) Non-metals possess low densities. They are light as compared to metals.
- (vii) Non-metals are non-sonorous.

(viii) Non-metals may exist in solid, liquid or gaseous state at room temperature. For example sulphur and carbon are solid at room temperature, bromine is liquid and nitrogen, oxygen are gaseous non-metals.



Liquid non-metal : Bromine

(ix) All non-metals possess low melting and boiling points except graphite.

1.3 CHEMICAL PROPERTIES OF METALS

Metals are highly electropositive elements, i.e. they form positive ions by losing electrons.

M –→ Mⁿ⁺ + ne⁻

For example:

Na \rightarrow Na⁺ + e⁻

Mg \rightarrow Mg²⁺ + 2e⁻

The characteristic chemical properties of metal are due to their electropositive character.

1.3.1 Reaction with oxygen (of air)

Metals combine with oxygen to form their respective oxides. The metal oxides are ionic compounds. Metals lose their valence shell electrons and form positive ions while oxygen



atoms receive electrons to form the negative oxide ion. The two oppositely charged ions combine to produce ionic metal oxide.

Metal + Oxygen \rightarrow Metal oxide

Metal oxides are basic in nature.

The reactivity of metals towards oxygen are different e.g.

(i) Sodium and potassium react with oxygen at room temperature to form basic oxides.

4Na	+	$O_2 \longrightarrow$	2Na ₂ O
Sodium		Oxygen	Sodium oxide

 $\begin{array}{cccc} 4K & + & O_2 & \longrightarrow & 2K_2O \\ \text{Potassium} & & \text{Oxygen} & \text{Potassium oxide} \end{array}$

Most of the metal oxides are insoluble in water. But some of the metal oxides dissolve in water to form alkalis.

Sodium and potassium oxide are the two metal oxides which are soluble in water and form alkalis.

Na ₂ O +	$H_2O \rightarrow 2NaOH$
Sodium oxide	Sodium hydroxide
(Basic oxide)	(An alkali)
K ₂ O +	$H_2O \longrightarrow 2KOH$
Potassium oxide	Potassium hydroxide
(Basic oxide)	(An alkali)

(ii) Magnesium does not react with oxygen at room temperature. It combines with oxygen only on heating. On heating, magnesium burns in air producing intense heat and light.

 $\begin{array}{cccc} 2Mg & + & O_2 & - \rightarrow & 2MgO \\ \hline Magnesium & & Oxygen & Magnesium oxide \end{array}$

Since reaction of magnesium with oxygen takes place at higher temperature than the reaction of sodium with oxygen, it means magnesium is less reactive than sodium.

(iii) Zinc metal burns in air only on strong heating to form zinc oxide.

 $2Zn + O_2 \rightarrow 2ZnO$

Oxygen Zinc oxide

(iv) Iron does not burn in air even on strong heating. The reaction of iron with oxygen takes place less readily than that of zinc.

 $\begin{array}{rcl} 2Fe & + & O_2 & \longrightarrow & 2Fe_3O_4\\ & & & \\ Iron & & & \\ Oxygen & & & \\ Iron oxide \end{array}$ Fe₃O₄ is a mixture of ferrous oxide (FeO) and ferric oxide (Fe₂O₃) and is known as iron (II, III) oxide.

(v) Copper is very less reactive towards oxygen. Copper burns in air on prolonged strong heating.

 $\begin{array}{cccc} 2Cu & + & O_2 & - \rightarrow & 2CuO\\ Copper & Oxygen & Copper(II) oxide\\ the order of reactivity of these metals with oxy$

So, the order of reactivity of these metals with oxygen is:

Na > Mg > Zn > Fe > Cu

Zinc



Silver and gold do not react with oxygen even at high temperature.

1.3.2 Amphoteric oxides

Though most of the metal oxides are basic in nature but some of them show basic as well as acidic nature. Those metal oxides which show basic as well as acidic behaviour are known as amphoteric oxides. Aluminium metal and zinc metal forms amphoteric oxides. Amphoteric oxides react with both, acids as well as bases to form salt and water.

(i) Aluminium oxide as basic oxide: Aluminium oxide reacts with hydrochloric acid to form aluminium chloride and water.

Al ₂ O ₃ +	- 6HCI –→	2AICI ₃ +	H_2O
Aluminium oxide	Hydrochloric acid	Aluminium chloride	Water
	(Acid)	(Salt)	

(ii) Aluminium oxide as acidic oxide: Aluminium oxide reacts with sodium hydroxide to form sodium aluminate and water.

Al ₂ O ₃ +	2NaOH –→	2NaAlO ₂ +	H_2O
Aluminium oxide	Sodium Hydroxide	Sodium aluminate	Water
	(Base)	(Salt)	

Zinc oxide is an amphoteric oxide which reacts with acids as well as with bases to form salt and water.

(iii) **Zinc oxide as basic oxide:** Zinc oxide reacts with hydrochloric acid to form zinc chloride (salt) and water.

ZnO	+	2HCI	\rightarrow	ZnCl ₂	+	H_2O
Zinc oxide		Hydrochloric a	icid	Zinc chloride		Water
		(Acid)		(Salt)		

(iv) Zinc oxide as acidic oxide: It reacts with sodium hydroxide to form sodium zincate and water.

ZnO	+ 2NaOH \rightarrow	Na ₂ ZnO ₂ +	H ₂ O
Zinc oxide	Sodium hydroxide	Sodium zincate	Water
	(Base)	(Salt)	

1.3.3 Reaction of metals with water

Metals react with water and produce metal oxide or metal hydroxide and hydrogen gas. However, the reactivities of metal with water are different.

Metal + Water \rightarrow Metal oxide or metal hydroxide + Hydrogen (i) Metals such as potassium, sodium and calcium react with cold water.

2K	+	2H₂O	\rightarrow	2KOH	+	H_2
Potassium		Water		Potassium hydrox	ide	
2Na	+	$2H_2O$	\rightarrow	2NaOH	+	H_2
Sodium		Water		Sodium hydroxid	le	
Ca	+	2H ₂ O	\rightarrow	Ca(OH) ₂	+	H_2
Calcium		Water		Calcium hydroxi	ide	

Potassium decomposes water more vigorously than sodium. Hence, potassium is more reactive than sodium. Sodium decomposes water more vigorously than calcium. Hence, sodium is more reactive than calcium.



 H_2

(ii) Magnesium does not react with cold water. It reacts with hot water to form magnesium hydroxide and hydrogen.

Mg	+	$2H_2O$	\rightarrow	Mg(OH) ₂ +
Magnesium		Water		Magnesium hydroxide
		(hot)		

(iii) Metals like aluminium, iron and zinc do not react either with cold or hot water. They react with steam to form metal oxide and hydrogen.

2AI	+	2H ₂ O	\rightarrow	AI_2O_3	+	$3H_2$
Aluminium		Steam		Aluminium oxid	de	
Zn	+	H ₂ O	\rightarrow	ZnO	+	H_2
Zinc		Steam		Zinc oxide		
3Fe	+	4H ₂ O	\rightarrow	Fe ₃ O ₄	+	$4H_2$
Iron		Steam		lron(II, III) ox	ide	
(Red hot)						

(Red hot)

The metals such as lead, copper, silver and gold do not react with water at all.

Metals evolve hydrogen from water

Water is a very weak electrolyte. It ionizes to give hydrogen ions (H^+) and hydroxide ions (OH^-).

 $H_2O \longrightarrow H^+ + OH^-$

When a metal reacts with water, it loses electrons to reduce H⁺ ions to H atoms.

 $2H^+$ + $2e^- \rightarrow H_2$

The H atom then unites to give H_2 gas.

1.3.4 Reaction of metals with Acids

Metals usually displace hydrogen from dilute acids.

Metal + Dilute acid \rightarrow Salt + Hydrogen

All metals, however do not react with dilute acids.

(a) Reaction of metals with hydrochloric acid

2HCI

(i) Sodium metal reacts with dilute hydrochloric acid vigorously to form sodium chloride and hydrogen gas.

 H_2

2Na +

2NaCl Sodium chloride

This shows that sodium is a very reactive metal.

(ii) Magnesium reacts rapidly with dilute hydrochloric acid to form magnesium chloride and hydrogen gas.

 $\begin{array}{cccc} Mg & \textbf{+} & 2HCI & - \rightarrow & MgCl_2 & \textbf{+} & H_2 \\ & & & & \\ & & & \\ & & & & \\ & & & & \\ & & & & \\ & & & & \\ & & & & \\ & & & & \\ & & & & \\ & & & & \\ & & & & \\ & & & & \\ & & & & \\ & & & & \\ & & & & \\ & & & & \\ & & & & \\ & & & & \\ & & & & \\ & & & \\ & & & \\ & & & & & \\ & & & & \\ & & & & \\ & & & & \\ & & & & \\ & & & & \\ & & & & \\ & &$

(iii) Aluminium metal reacts rapidly with dilute hydrochloric acid to form aluminium chloride and hydrogen gas.



Fe

Cu

Metals and Non-Metals (Notes)

 H_2

At first, it reacts slowly due to the presence of tough protective layer of aluminium oxide on its surface. But when the outer oxide layer gets dissolved in acid, then fresh aluminium metal is exposed which reacts rapidly.

(iv) Zinc reacts with dilute hydrochloric acid to produce zinc chloride and hydrogen but less rapidly than that of aluminium.

Zn + 2HCl \rightarrow $ZnCl_2$ + H_2 Zinc chloride

This reaction shows that zinc is less reactive than aluminium.

(v) Iron reacts slowly with dilute hydrochloric acid to form ferrous chloride

+ 2HCl \rightarrow FeCl₂ + Ferrous chloride

(vi) Copper does not react with dilute hydrochloric acid.

+ HCI \rightarrow No reaction

Gold and silver metals also do not react with dilute acids.

Metals displace hydrogen from dilute acids

Those metals which are more reactive than hydrogen, i.e., those metals which lose electrons more easily than hydrogen, displace hydrogen from dilute acids to produce hydrogen gas. This is due to the fact that the more reactive metals give electrons easily and those electrons reduce hydrogen ions of acids to hydrogen gas.

The metals like copper and silver which are less reactive than hydrogen, do not displace hydrogen from dilute acids. This is because they do not give out electrons required for the reduction of hydrogen ion present in acids.

(b) Reaction of a metal with dilute sulphuric acid

Metals displace hydrogen from dilute sulphuric acid. For example :

2Na (s) Sodium		+ H_2SO_4 Sulphuric acid	(aq)	$-\rightarrow$ Na ₂ S Sodium sulph	O₄ (aq) ate	+ Hydrog	H ₂ (g) en
Mg (s) Magnesium	+	H ₂ SO ₄ (aq) Sulphuric acid	\rightarrow	MgSO ₄ (aq) Magnesium sul	+ phate	H ₂ (g) Hydroge	n
2AI (s) Aluminium	+	3H ₂ SO ₄ (aq) Sulphuric acid	\rightarrow	Al ₂ (SO ₄) ₃ (ac Aluminium sulp	l) + hate	3H ₂ (g) Hydroge	n
Zn (s) ^{Zinc}	+	H ₂ SO ₄ (aq) Sulphuric acid	\rightarrow	ZnSO ₄ (aq) Zinc sulphate	+	H ₂ (g) Hydroge	n
Cu (s) Copper	+	H ₂ SO ₄ (aq) Sulphuric acid	\rightarrow	No reaction			

(c) Reaction of a metal with dilute nitric acid

When a metal reacts with dilute nitric acid, then hydrogen gas is not evolved. Nitric acid is a strong oxidizing agent. So, **as soon as hydrogen gas is formed in the reaction between a metal and dilute nitric acid, the nitric acid oxidizes this hydrogen to water.** So, in the reactions of metals with dilute nitric acid, no hydrogen



gas is evolved. Now, when nitric acid oxidizes hydrogen to water, then nitric acid itself is reduced to any of the nitrogen oxides (such as dinitrogen monoxide, N_2O ; nitrogen monoxide, NO; or nitrogen dioxide, NO_2). The type of oxide formed depends on the nature of metal, the temperature of reaction and concentration of nitric acid.

Very dilute nitric acid, however, reacts with magnesium and manganese metals to evolve hydrogen gas. This is because the very dilute nitric acid is a weak oxidizing agent which is not able to oxidize hydrogen to water.

(i) **Magnesium** reacts with very dilute nitric acid to form magnesium nitrate and hydrogen gas.

Mg(s) +	$2HNO_3$ (aq) \rightarrow	$Mg(NO_3)_2$ (aq) +	H ₂ (g)
Magnesium	Nitric acid	Magnesium nitrate	Hydrogen
	(very dilute)		

(ii) Manganese reacts with very dilute nitric acid to form manganese nitrate and hydrogen gas.

Mn (s) +	2HNO₃ (aq)	\rightarrow	Mn(NO ₃) ₂ (aq)	+	$H_2(g)$
Manganese	Nitric acid		Manganese nitrate		Hydrogen
	(very dilute)				

1.3.5 Reaction of metal with solutions of other metal salts

(a) A more reactive metal displaces a less reactive metal from its salt solution. Such reactions are called displacement reactions.

Metal A + Salt solution of B \rightarrow Salt solution of metal A + Metal B For example :

 $\begin{array}{c} {\sf Zn}(s)\\ {\sf Zinc} \end{array} + \begin{array}{c} {\sf CuSO}_4(aq)\\ {\sf Copper \ Sulphate}\\ ({\sf Blue}) \end{array} \to \begin{array}{c} {\sf ZnSO}_4(aq)\\ {\sf Zinc \ Sulphate}\\ ({\sf Colourless}) \end{array} + \begin{array}{c} {\sf Cu}(s)\\ {\sf Copper \ ({\sf Reddish \ brown})} \end{array}$

(i)

Thus, when zinc pieces are added to copper sulphate solution, then, zinc being more reactive metal than copper, displaces copper from its solution ($CuSO_4$) so that copper is set free. The blue colour of $CuSO_4$ solution fades due to formation of $ZnSO_4$ (colourless). A reddish brown deposit of copper metal is formed on the surface of zinc.

 $\begin{array}{ccc} Fe(s) + & CuSO_4(aq) \rightarrow & FeSO_4(aq) + & Cu(S) \\ & \text{iron sulphate} \\ & (Blue \ solution) & (Brownish \ deposit \ on \ iron) \end{array}$

(ii)

(b) When iron nail is added to copper sulphate solution, then, iron being more reactive metal than copper, displaces copper from its solution (CuSO₄) so that Cu is set free. The blue colour of CuSO₄ fadeds and it changes into light green due to formation of iron sulphate (FeSO₄).

Relative activities or reactivities of metals

Metals have been arranged in decreasing order of their activities (or reactivities) in the activity series. After performing displacement experiments, the following series known as the reactivity or activity series has been developed as follows :



It is clear from the series that the metals lying above the hydrogen are more reactive than the metals lying below the hydrogen. Thus any metal can displace the metals lying below it from its solution.

(c) Significance of activity series

- (i) The metals above hydrogen in the activity series have greater tendency than hydrogen to give up electrons in their solutions. Such metals are called electropositive metals. The electropositive character of metals becomes less pronounced as we go down the series. For example, potassium (K), the first metal in the series is the most electropositive, while platinum (Pt), the last metal of the series is the least electropositive.
- (ii) The metals above hydrogen in the series can liberate hydrogen when treated with an acid solution. Thus, magnesium and zinc react with dilute solutions of sulphuric acid to produce hydrogen gas.

In these reactions, electrons released by metals are accepted by H^+ or H_3O^+ ions present in the acid solution.

2e-

Mg \longrightarrow Mg²⁺ +



Thus, H^{+} (or $H_{3}O^{+}$) ions act as oxidizing agents while the metal acts as reducing agent.

(iii) A more electropositive metal can replace a less electropositive metal from the solution of a salt of the less electropositive metal. For example, when an iron rod is dipped into a solution of copper sulphate, reddish coloured copper is deposited on the iron rod. Fe + $CuSO_4 \longrightarrow FeSO_4 + Cu$

This is because iron is more electropositive than copper.

1.3.6 Reaction of metals with chlorine

Most of the metals react with chlorine to form chlorides. These chlorides are ionic in character.

Metal + Chlorine \rightarrow Metal chloride

During this reaction, metal loses electrons and become positively charges whereas chlorine atoms accept electrons and become negatively charged ions. Metals form ionic chlorides because they give electrons to chlorine atoms to form ions.

Metal chlorides are usually solid and conduct electricity in solution or molten state. Thus metal chlorides are electrolytes.

(i) Sodium readily reacts with chlorine to form an electrovalent chloride called sodium chloride

2NaCl

- 2Na + $Cl_2 \rightarrow$ Sodium Chlorine
 - Chlorine Sodium chloride
- (ii) Calcium reacts vigorously with chlorine to form an electrovalent chloride called calcium chloride.

Са + CaCl₂ Chlorine Calcium chloride Calcium

(iii) Magnesium on heating forms magnesium chloride, which is an electrovalent chloride.

Mg + $Cl_2 \longrightarrow MgCl_2$

Magnesium Chloride Magnesium chloride

- (iv) Zinc combines directly with chlorine to form zinc chloride.

(v) Iron combines with chlorine, when heated, to form iron (III) chloride or ferric chloride

```
2 Fe + 3Cl_2 \xrightarrow{heat} 2 FeCl_3
```

Iron Chlorine Ferric Chloride

(vi) On heating, copper reacts with chlorine to form copper (II) chloride or cupric chloride $Cu + Cl_2 \longrightarrow CuCl_2$ Copper Chlorine Cupric chloride

1.3.7 Reaction of metals with hydrogen

Most of the metals do not react with hydrogen. A few reactive metals such as sodium, potassium and calcium react with hydrogen to form ionic hydride.



For example:

Sodium

2K

(i) When hydrogen gas is passed over heated sodium, then sodium hydride is formed. 2Na + H₂ \rightarrow 2NaH

Sodium hydride

(ii) When hydrogen gas is over heated potassium, then potassium hydride is formed.

+ $H_2 \longrightarrow 2KH$

Potassium Potassium Hydride

(iii) When hydrogen gas is passed over heated calcium, then calcium hydride is formed.

 $Ca \qquad + \qquad H_2 \quad - \rightarrow \quad CaH_2$

Calcium Calcium Hydride

The comparatively less reactive metals like zinc, copper and iron do not react with hydrogen.

1.3.8 Reaction of metal with metal oxides

A more reactive metal can displace a less reactive metal from its oxide. The reactive metals are good reducing agents because they can easily give the electrons needed for reduction. For example:

(i) Aluminium can displace iron from ferric oxide.

2AI	+	Fe ₂ O ₃ ––	→ 2Fe	+ Al ₂ O ₃
Aluminium		Ferric oxide	Iron	Aluminium oxide

(ii) Magnesium can displace copper from copper oxide.

CuO +	Mg	_→	MgO	+	Cu
Copper (II) Oxide	Magnesium		Magnesium o	xide	Copper

1.4 CHEMICAL PROPERTIES OF NON-METALS

The non-metals are electronegative elements, i.e. they form negative ions by gaining electrons. In other words, non-metals have tendency to gain electrons and hence undergo reduction.

1.4.1 Reaction of non-metals with Oxygen

Non-metals react with oxygen to form oxides. These oxides are covalent in nature and are generally gaseous. Most of these oxides when dissolved in water give acidic solution. These solutions turn blue litmus red. For example, sulphur when burnt in the presence of oxygen forms sulphur dioxide.

(a) S + $O_2 \rightarrow SO_2$

Sulphur dioxide

Sulphur dioxide when dissolved in water forms sulphurous acid.

SO ₂	+	H_2O	\rightarrow	H_2SO_3
Sulphur d	ioxide			Sulphurous acid

(b) Similarly, carbon forms carbon dioxide and phosphorus forms phosphorus pentoxide. These oxides also give acidic solutions when dissolved in water.





On the other hand, certain oxides of non-metals are neutral. Some examples are nitrous oxide (N_2O), carbon monoxide (CO), etc. These oxides do not have any effect on litmus paper.

1.4.2 Reaction of non-metals with water

Non-metals do not react with water or steam to evolve hydrogen gas because they do not lose electrons to reduce hydrogen ions (H^+) into hydrogen gas.

1.4.3 Reaction of non-metals with acids

Non-metals do not displace hydrogen from acids. For the liberation of hydrogen the non-metals should be able to reduce H^+ ions to H_2 gas by supplying electrons. However, non-metals are electronegative elements and hence have more tendency to accept electrons rather than donating. Hence, non-metals do not produce hydrogen gas on reaction with acids.

1.4.4 Reaction of non-metals with salts solution

A more reactive non-metal displaces a less reactive non-metal from its salt solution. For example :

When chlorine is passed through a solution of sodium bromide, then sodium chloride and bomine are formed :

 Cl_2

1.4.5 Reaction of non-metals with chlorine

+

Non-metals react with chlorine to form chlorides. These chlorides are covalent in nature and are generally gaseous or liquids.

(a) Hydrogen reacts with chlorine to form covalent chloride called hydrogen chloride.

 H_2

2HCI
Hydrogen chloride

 Br_2

Bromine

(b) Phosphorus reacts with chlorine to form covalent chloride called phosphorus trichloride and phosphorus pentachloride.

P ₄	+	$6Cl_2 \longrightarrow$	4PCl ₃
			Phosphorus trichloride
P ₄	+	$10Cl_2 \rightarrow$	
			Phosphorus pentachloride

1.4.6 Reaction of non-metals with hydrogen

Non-metals on reaction with hydrogen form hydrides which are covalent in nature. Most of these hydrides are gases or liquids. Some examples are :

(a) 2H ₂	+	O ₂	\rightarrow	$2H_2O$
				Water



Science (Class X)

Metals and Non-Metals (Notes)

(b) N ₂	+	$3H_2$	\rightarrow	$2NH_3$
				Ammonia
(c) H ₂	+	S	\rightarrow	H_2S
				Hydrogen sulphide

1.4.7 Oxidising Nature

Since non-metals have great tendency to accept electrons, therefore, many non-metals act as good oxidizing agents. For example, fluorine is the strongest oxidizing agent among all the non-metals.

(a) Zn	+	S	\rightarrow	Zn ²⁺ S ^{2–}
				Zinc sulphide
(b) 2Na	+	Cl_2	\rightarrow	2Na⁺ Cl⁻
				Sodium chloride

1.5 DIFFERENCE BETWEEN METALS AND NON-METALS

Properties	Metals	Non-metals		
Physical Properties				
1. State	Metals are solids at ordinary	Non-metals exist in all the three states,		
	temperature. (except mercury,	that is, solid, liquid and gas.		
	which is a liquid.)			
2. Lustre	They possess lustre or shine.	They possess no lustre.		
		(except lodine and graphite.)		
3. Malleability and	Metals are generally malleable	Non-metals are neither malleable nor		
Ductility	and ductile.	ductile.		
4. Hardness	Metals are generally hard.	Non-metals possess varying hardness.		
	Alkali metals are exception.	Diamond is an exception. It is the		
		hardest substance known to occur in		
		nature.		
5. Density	They have high densities.	They generally possess low densities.		
6. Conductivity	Metals are good conductors of	Non-metals are poor conductors of		
	heat and electricity.	heat are electricity. The only exception		
		is graphite which is a good conductor		
		of electricity.		
7. Melting and boiling	They usually have high melting	Their melting and boiling point are		
point	and boiling point.	usually low. The exceptions are boron,		
		carbon and silicon.		
Chemical Properties				
1. Action with mineral	Metals generally react with dilute	Non-metals do not displace hydrogen		
acids	mineral acids to liberate	on reaction with dilute minerals acids.		
	hydrogen gas.			
2. Nature of oxides	They form basic oxides. For	Non-metals form acidic or neutral		
	example, Na ₂ O, MgO, etc. These	oxides. For examples, SO ₂ , CO ₂ , P ₂ O ₅ ,		
	oxides are ionic in nature.	etc. are acidic whereas CO, N_2O , etc.		
		are neutral. These oxides are covalent		
		in nature.		



Science (Class X)

Metals and Non-Metals (Notes)

<u> </u>		Matela and the last the state of the second se	
3.	Complination with	Metals generally do not combine	Non-metals combine with hydrogen to
	hydrogen	with hydrogen. However Li, Na,	form stable hydrides. For example,
		Ca, etc. form unstable hydrides.	HCI, H_2S , CH_4 , NH_3 , PH_3 , etc. These
		For examples, LiH, NaH, CaH ₂	hydrides are covalent.
		etc. These hydrides are ionic in	
		character.	
4.	Combination with	They combine with halogens to	Non-metals form halides which are
	halogens	form well defined and stable	unstable and undergo hydrolysis
		crystalline solids. For example,	readily. For example, PCl ₅ , PCl ₃ , etc.
		NaCl, KBr, etc.	
5.	Electrochemical	Metals are electropositive in	Non-metals are electronegative in
	behaviour	character. They form cations in	character. They form anions in solution
		solution and are deposited on	and are liberated at the anode when
		the cathode when electricity is	their salt solutions are subjected to
		passed through their solution.	electrolysis. Hydrogen in an exception.
			It usually forms positive ions and is
			liberated at cathode.
6.	Oxidising or	Metals behave as reducing	Non-metals generally behave as
	reducing behaviour	agents. This is because of their	oxidising agents since they have the
		tendency to lose electrons.	tendency to gain electrons.
		Na –→ Na⁺ + e⁻	1
			$\frac{1}{2}$ Cl ₂ + e ⁻ \rightarrow Cl ⁻
			2

1.6 USES OF METALS

Metals are used for a large number of purposes. Some of the uses of metals are given below:

- (i) Copper and aluminium metals are used to make wires to carry electric current. This is because copper and aluminium have very low electrical resistance and hence very good conductors of electricity.
- (ii) Iron, copper and aluminium metals are used to make house-hold utensils and factory equipments.
- (iii) Iron is used as a catalyst in the preparation of ammonia gas by Haber's process.
- (iv) Zinc is used for galvanizing iron to protect it from rusting.
- (v) Chromium and nickel metals are used for electroplating and in the manufacture of stainless steel.
- (vi) The aluminium foils are used in packaging of medicines, cigarettes and food materials.





Aluminium Foil

- (vii)Silver and gold metals are used to make jewellery. The thin foils made of silver and gold are used to decorate sweets.
- (viii) The liquid metal 'mercury' is used in thermometers.
- (ix) Sodium, titanium and zirconium metals are used in atomic energy (nuclear energy) and space science projects.
- (x) Zirconium metal is used in making bullet-proof alloy steels.

1.7 USES OF NON-METALS

The important uses of non-metals are as follows:

- (i) Hydrogen is used in the hydrogenation of vegetable oils to make vegetable *ghee* (or vanaspati ghee).
- (ii) Hydrogen is used in the manufacture of ammonia (whose compounds are used as fertilisers).
- (iii) Liquid hydrogen is used as a rocket fuel.
- (iv) Carbon (in the form of graphite) is used for making the electrodes of electrolytic cells and dry cells.
- (v) Nitrogen is used in the manufacture of ammonia, nitric acid and fertilisers.
- (vi) Due to its inertness, nitrogen is used to preserve food materials.
- (vii)Compounds of nitrogen like Tri Nitro Toluene (TNT) and nitroglycerine are used as explosives.
- (viii) Sulphur is used for manufacturing sulphuric acid.
- (ix) Sulphur is used as a fungicide and in making gun powder.
- (x) Sulphur is used to the vulcanisation of rubber.

2. WHY DO METALS AND NON-METALS REACT?

When metals react with non-metals, they form ionic compounds (which contain ionic bonds). On the other hand, when non-metals react with other non-metals, they form covalent compounds (which contain covalent bonds).

The force which holds a number of atoms together within the compound or the molecule is called a chemical bond.



To understand the formation of chemical bond 'between the atoms of metals and non-metals' or 'between the atoms of two non-metals', it is necessary to know the reason for the unreactive nature (or inertness) of noble gases.

There are some elements in group 18 of the periodic table which do not combine with other elements. These elements are : Helium, Neon, Agron, Krypton, Xenon and Radon. They are known as noble gases or inert gases because they do not react with other elements to form compounds. In other words, inert gases do not form chemical bonds.

That is why they are called noble gases or inert gases and exist as monoatomic gases, i.e., as He, Ne, Ar, Kr, Xe and Rn.

S.No.	Element	Symbol	Atomic No.	Electronic configuration					
				К	L	М	Ν	0	Р
1.	Helium	Не	2	2					
2.	Neon	Ne	10	2,	8				
3.	Argon	Ar	18	2,	8,	8			
4.	Krypton	Kr	36	2,	8,	18,	8		
5.	Xenon	Хе	54	2,	8,	18,	18,	8	
6.	Radon	Rn	86	2,	8,	18,	32,	18,	8

Electronic Configurations of noble gases

When we look at their electronic configuration, as given in the above table, we observe that except in case of helium, where there are two electrons present in the outermost shell, all other noble gases have eight electrons present in the outermost shell. The presence of two electrons in the outermost shell is called a duplet whereas the presence of eight electrons in the outermost shell is called an octet. Thus, the cause of stability of noble gases is due to presence of duplet of electrons in case of helium and presence of octet of electrons in case of other noble gases.

Name of the	Symbol	Atomic No.	Electronic c	onfiguratio	on	
metal atom			K	L	М	Ν
Lithium	Li	3	2	1		
Sodium	Na	11	2	8	1	
Magnesium	Mg	12	2	8	2	
Aluminium	Al	13	2	8	3	
Potassium	К	19	2	8	8	1
Calcium	Са	20	2	8	8	2
Name of the	Symbol	Atomic No.	Electronic c	onfigurati	on	
non-metal atom			17			NI
			К	L	IVI	IN
Nitrogen	N	7	K 2	5	M	N
Nitrogen Oxygen	N O	7 8	K 2 2	5 6	M	N
Nitrogen Oxygen Fluorine	N O F	7 8 9	K 2 2 2	5 6 7	M	N
Nitrogen Oxygen Fluorine Phosphorus	N O F P	7 8 9 15	R 2 2 2 2 2	5 6 7 8	5	N
Nitrogen Oxygen Fluorine Phosphorus Sulphur	N O F P S	7 8 9 15 16	K 2 2 2 2 2 2	5 6 7 8 8	5 6	N

Electronic Configurations of some common metals and non-metals



When we look at the electronic configuration of few metals and non-metals, as given in the above table, we observe that their outermost shell has less than eight electrons. So, all atoms tend to complete their octets (i.e., outermost shell with eight electrons) or duplet (i.e., outermost shell with two electrons) if K-shell is the outermost shell to acquire stability. Thus, the cause of chemical combination is the tendency of the atoms to complete their octet (i.e., outermost shell with eight electrons) or duplet (i.e., outermost shell with eight electrons) or duplet their octet (i.e., outermost shell with eight electrons) or duplet their octet (i.e., outermost shell with eight electrons) or duplet their octet (i.e., outermost shell with eight electrons) or duplet (i.e., outermost shell with two electrons in case of elements having only K-shell) so that they acquire the stable nearest noble gas configuration.

Modes of Chemical combination

As explained above, when atoms combined, they complete their octets (or duplets) by any one of the following methods :

- (i) by losing one or more electrons (to another atom).
- (ii) by gaining one or more electrons (from another atom).
- (iii) by sharing one or more electrons (with another atom).





2.1 IONS

An ion is an electrically charged atom (or group of atoms). Examples of the ions are : sodium ion, Na⁺, magnesium ion, Mg²⁺, chloride ion, Cl⁻, and oxide ion, O²⁻. An ion is formed by the loss or gain of electrons by an atom.



2.1.1 A positively charged ion is known as cation

Sodium ion, Na^+ , and magnesium ion, Mg^{2+} , are cations because they are positively charged ions. A cation is formed by the loss of one or more electrons by an atom. For example, sodium atom loses one electron to form a sodium ion, Na^+ , which is a cation.

	Na	_	e-	\rightarrow	Na⁺
S	odium atom		Electron		Sodium ion
Electronic	KLM				K L
Configuration:	2, 8, 1				2, 8
	(unstable electron			(5	Stable, neon gas electron
	arrangement)				arrangement)

2.1.2 A negatively charged ion is known as anion

Chloride ion, CI^- , and oxide ion, O^{2-} , are anion because they are negatively charged ion. An anions is formed by the gain of one or more electrons by an atom. For example, a chlorine atom gains (accepts) one electron to form a chlorine ion, CI^- , which is an anion

	CI	+	e–	\rightarrow	Cl⁻	
	Chlorine atom		Electron		Chloride ion	
Electronic	KLM				KLM	
Configuration	n: 2, 8, 7				2, 8, 8	
	(unstable ele	ctron		((Stable, neon gas elec	tron
	arrangemer	nt)			arrangement))

2.2 ELECTRONIC DOT REPRESENTATION (LEWIS SYMBOLS)

In the formation of a chemical bond between two atoms, only the electrons of the outermost shell are involved (as the inner shell electrons are well protected). These electrons present in the outermost shell are called valence electrons.

G.N. Lewis introduced a simple method of representing the valence electrons by dots or small crosses around the symbol of the atom. These symbols are known as **electron dot symbols** or **Lewis symbols**. These symbols ignore the inner shell electrons.

Element	Symbol	Atomic	Electronic configuration			ration	Valence	Lewis symbol
		No.					electrons	
			К	L	М	N		
Hydrogen	Н	1	1				1	H • or H [×]
Helium	Не	2	2				2	He : or H [×]

Electron dot (Lewis symbols) of some common elements

Educor	Science (Class X) Metals and Non-Metals (Notes)							
Lithium	Li	3	2	1			1	Li • or Li [×]
Beryllium	Ве	4	2	1			1	Be : or Be [×]
Boron	В	5	2	3			3	•B• or ×B×
Caron	с	6	2	4			4	•C• or ×C×
Nitrogen	N	7	2	5			5	°××××××××××××××××××××××××××××××××××××
Oxygen	0	8	2	6			6	O: or ×O×
Fluorine	F	9	2	7			7	•F• or $\stackrel{\times}{_{\!$
Neon	Ne	10	2	8			8	Net or XNex
Sodium	Na	11	2	8	1		1	Na• or Na [×]
Magnesium	Mg	12	2	8	2		2	Mg: or Mg $_{\times}^{\times}$
Aluminium	AI	13	2	8	3		3	•Àl• or ×Ål×
Chlorine	CI	17	2	8	7		7	:Cit or $\overset{\times \times \times}{\underset{\times \times}{\overset{\times \times}{\overset{\times}{$
Potassium	к	19	2	8	8	1	1	K• or K [×]
Calcium	Са	20	2	8	8	2	2	Ca: or Ca [×]

2.3 TYPES OF CHEMICAL BONDS

There are two types of chemical bonds :

- (i) Ionic bond, and
- (ii) Covalent bond.

lonic bonds are formed by the transfer of electrons from one atom to another whereas covalent bonds are formed by the sharing of electrons between two atoms. Ionic bond is also called electrovalent bond.

2.3.1 Ionic bond

The chemical bond formed by the transfer of electrons from one atom to another is known as an ionic bond. The transfer of electrons takes place in such a way that the ions formed have the stable electron arrangement of an inert gas. The ionic bond is called so because it is a chemical bond between oppositely charged ions.

When a metal reacts with a non-metal, transfer of electrons takes place from metal atoms to the non-metal atoms, and an ionic bond is formed.

The compounds containing ionic bonds are called ionic compounds. Ionic compounds are made up of ions. Lets understand with the help of examples.



e-

(i) Formation of Sodium Chloride.

Atomic number of sodium (Na) = 11

 \therefore Its electronic configuration is 2, 8, 1.

It has only one electron in the valence shell. It loses this electron to acquire the stable electronic configuration 2, 8 (similar to that of neon) and form sodium ion (Na^{+}) .

Na [×]	\rightarrow	Na⁺	+
Sodium atom		Sodium ion	
(2, 8, 1)		(2, 8)	

Atomic number of chlorine (Cl) = 17

: Its electronic configuration is 2, 8, 7.

It has seven electrons in the valence shell. It gains one electron to acquire the stable electronic configuration 2, 8, 8 (similar to that of argon) and form chloride ion (CI^-)

Chloride ion (2, 8)

:[C|:]

Thus, when a sodium atom and a chlorine atom approach each other, an electron is transferred from sodium atom to chlorine atom. In other words, sodium loses one electron to form Na⁺ ion and chlorine gains that electron to form Cl⁻ ion. As a result, both acquire the stable nearest noble gas configuration. These oppositely charged ions are then held together by electrostatic forces of attraction forming the compound Na⁺Cl⁻ or simply written as NaCl. The transfer of electron may be represented in one step as follows:

→ Na⁺ [×Ci:]⁻ or NaCl

Sodium chloride

Sodium atom

Na

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

*CI:

Chlorine atom

(ii) Formation of Magnesium chloride.

Atomic number of magnesium (Mg) = 12

 \therefore Its electronic configuration = 2, 8, 2

It loses two electrons from the valence shell to acquire the nearest noble gas configuration of neon (2, 8) and form Mg²⁺ ion.

Atomic number of chlorine (CI) = 17

It electronic configuration = 2, 8, 7

It needs to gain only one electron in the valence shell to acquire the nearest noble gas configuration of argon (2, 8, 8) and form chloride ion (Cl⁻). Now, as magnesium atom has to lose two electrons and a chorine atom can gain only one electron, therefore, two chlorine atoms will be required to accept the two electrons, one by each chlorine atom. Thus, the transference of two electrons from one Mg atom to two Cl atoms may be represented as follows:



Thus, here one magnesium atom combines with two chlorine atom. Therefore, the formula of magnesium chloride is $Mg^{2+}Cl_2^-$ or $MgCl_2$. Here, again the oppositely charged ions are held together by electrostatic forces of attraction.

(iii) Formation of Magnesium Oxide.

Atomic number of magnesium (Mg) = 12

 \therefore Its electronic configuration = 2, 8, 2

It loses two electrons from the valence shell to acquire the stable configuration of neon (2, 8) and form Mg²⁺ ion.

Atomic number of oxygen (O) = 8

.. Its electronic configuration = 2, 6

It gains two electrons in the valence shell to acquire the stable configuration of neon again (2, 8) and form oxide ion (O^{2-}) .

Thus, in the formation of magnesium oxide, two electrons are transferred from magnesium atom to oxygen atom as represented below :



2.3.2 Covalent bond

The chemical bond formed by the sharing of electrons between two atoms is known as a covalent bond. The sharing of electrons takes place in such a way that each atom in the resulting molecule gets the stable electron arrangement of an inert gas.

Whenever a non-metal combines with another non-metal, sharing of electrons takes place between their atoms and a covalent bond is formed. A covalent bond can also be formed between two atoms of the same non-metal.

Covalent bonds are of three types :

- (i) Single covalent bond
- (ii) Double covalent bond
- (iii) Triple covalent bond

(i) Single Bond



Single bond is formed by the sharing of one pair of electrons between two atoms. A single covalent bond is formed by the sharing of two electrons between the atoms, each atom contributing one electron for sharing.

For example, a hydrogen molecule H_2 , contains a single covalent bond and it is written as H : H, the two dots drawn between the hydrogen atoms represent a pair of shared electrons which constitutes the single bond. A single covalent bond is denoted by written as H–H.

(a) Formation of a Hydrogen Molecule, H₂: A hydrogen atom is very reactive and cannot exist free because it does not have the stable, inert gas electron arrangement. So, hydrogen gas does not consist of single atoms, it consists of more stable H₂ molecules. The atomic number of hydrogen is 1, so its electronic configuration is K 1. Hydrogen atom has only one electron in the outermost shell (which is K shell), and this is not a stable arrangement of electrons. A stable arrangement is to have two electrons in the K shell because then the helium gas electron structure will be achieved. Thus, a hydrogen atom needs one more electron to become stable. It gets this electron by sharing with another hydrogen atom. So, two hydrogen atoms share one electron each to form a hydrogen molecule.



The formation of a hydrogen molecule from two hydrogen atoms can be shown by the following diagram:



(b) Formation of a Water Molecule, H₂O : Water is a covalent compound consisting of hydrogen and oxygen. It contains single covalent bonds. The formation of a water molecule from hydrogen and oxygen can be explained as follows :

The hydrogen atom has only one electron in its outermost K shell, so it needs one more electron to achieve the stable, two electron arrangement of the inert gas helium. The oxygen atom has six electrons in its outermost shell, and it needs two more electrons to complete the stable, eight electron arrangement of inert gas neon. So, one atom of oxygen shares its two electrons with two hydrogen atoms to form a water molecule.





(ii) Double Bond

Double bond is formed by the sharing of two pairs of electrons between two atoms. A double covalent bond is formed by the sharing of four electrons between two atoms, each atom contributing two electrons for sharing.

It is represented by putting two short line (=) between the two atoms. For example, oxygen molecule, O_2 , contains a double bond between two atoms and it can be written as O=O.

(a) Formation of Oxygen Molecule, O_2 : Oxygen atom is very reactive and cannot exist free because it does not have the stable, inert gas electron arrangement in its valence shell. The atomic number of oxygen is 8, so its electronic configuration is 2, 6. Thus, an oxygen atom has six electrons in its outermost shell. It requires two more electrons to achieve the stable, eight electron inert gas configuration. The oxygen atom gets these electrons by sharing its two electrons with the two electrons of another oxygen atom. So, two oxygen atoms share two electrons each and form a stable oxygen molecule.



Thus, in the oxygen molecule, the two oxygen atoms are held together by a double bond.

(b) Formation of Carbon Dioxide Molecule, CO₂ : Carbon dioxide is a covalent compound made up of carbon and oxygen elements and its contains covalent bonds in it.

Carbon atom has four valence electrons, so it needs two more electrons to achieve the eight-electron inert gas configuration and become stable. Oxygen atom has 6 valence electrons and it needs two more electrons to achieve the eight-electron configuration and become stable. So, one carbon atoms shares its four electrons with two oxygen atoms and forms a carbon dioxide molecule:

:C: + 2:
$$\ddot{O}$$
: \rightarrow : \ddot{O} :: \ddot{C} :: \ddot{O} : Or O=C=O
One carbon Two oxygen atom atoms molecule, CO₂

There are two double bonds in a carbon dioxide molecule.

(iii) Triple Bond

A triple bond is formed by the sharing of three pairs of electrons between two atoms. A triple bond is formed by the sharing of six electrons between two atoms, each atom contributing three electrons for sharing. A triple bond is actually a combination of three single bonds, so it is represented by putting three short line (\equiv) between the two atoms. Nitrogen molecule, N₂, contains a triple bond, so it can be written as N \equiv N.



(a) Formation of a Nitrogen Molecule, N_2 : A nitrogen atom is very reactive and cannot exist free because it does not have the stable electron arrangement of an inert gas. The atomic number of nitrogen is 7, so its electronic configuration is 2, 5. This means that a nitrogen atom has five electrons in its outermost shell. Since a nitrogen atom has five electrons in its outermost shell, it needs three more electrons to achieve the eight electron structure of an inert gas and become stable. So, two nitrogen atoms combine together by sharing three electrons each to form a molecule of nitrogen gas.



2.3.3 Properties of ionic compounds

(i) **Physical state** : Most of the ionic compounds are crystalline solids. They are relatively hard because of strong electrostatic forces of attraction between the oppositely charged ions. They are brittle and break into pieces on applying force.

Laboratory Test : Collect samples of sodium chloride, potassium chloride, barium chloride or any other salt from the laboratory. We see that they are all solids and brittle, i.e., they break into pieces if some force (e.g., hammer) is applied on them.

- (ii) Solubility : They are soluble in water (which is a polar solvent) but insoluble in organic solvents like benzene, alcohol, ether, chloroform etc.
 Laboratory Test : Shake the samples of the salts taken above one by one with water, alcohol, petrol and kerosene oil taken in four test tubes separately. We find that the salts are soluble in water but insoluble in other solvents.
- (iii) Melting points and boiling points : They have high melting and boiling points. This is because in the ionic compounds, the oppositely charged ions are held together by strong electrostatic forces of attraction. Hence, they require a lot of heat to cut off these forces of attraction and break them into ions.

Laboratory Test : Take any of the above salts in a dry pyrex or corning glass test tube (which can withstand high temperature). Hold in a test tube holder and heat it in the Bunsen burner. We will observe that it is difficult to melt it.

(iv) Colour in the flame : Most of the salts when brought into the flame, impart characteristic colours to the flame.

Laboratory Test : Take solid sodium chloride on a metal spatula and bring it into the blue flame of the Bunsen burner. We see that it imparts golden yellow colour to the flame. Similarly, potassium salts impart violet colour and barium salts impart green colour to the flame.





Heating a salt sample on a spatula

(v) Electrical conductivity : When dissolved in water, ionic compounds dissociate to produce free ions in the solution. Similarly, on melting, the ionic compound splits to produce free ions in the melt. As ions can conduct electricity, therefore, ionic compounds conduct electricity in the aqueous solution or in the molten state.

Laboratory Test : Take water in a beaker and dissolve any one of the above salts into it. Set up the electrical circuit as shown in figure below :



Testing the conductivity of a salt solution

On switching on the current, the bulb is found to glow. This shows that the salt solution conducts electricity.

(vi) lonic reactions : As ionic compounds produce free ions in the solution, therefore, reaction between two ionic compounds is the reaction between their ions. As oppositely charged ions combine very quickly, therefore, these reactions are quite fast.

Laboratory Test : Add silver nitrate solution into sodium taken in a test tube. It is observed that a precipitate of silver chloride is formed immediately alongwith sodium nitrate in the solution as follows :

Ag⁺ NO ³ (aq) + Na⁺Cl- (aq) \rightarrow AgCl (s) + Na⁺NO ³

(aq)	
------	--

Silver nitrate Sodium chloride Silver chloride (ppt.) Sodium nitrate

2.3.4 PROPERTIES OF COVALENT COMPOUNDS

(i) Covalent compounds are usually liquids or gases. Only some of them are soids. For example, alcohol, ether, benzene, carbon disulphide, carbon tetrachloride and bromine are liquids; methane, ethane, ethene, ethyne, and chlorine are gases.



Glucose, cane sugar, urea, naphthalene and iodine are, however, solid covalent compounds. The covalent compounds are usually liquids or gases due to the weak force of attraction between their molecules.

- (ii) Covalent compounds have usually low melting points and low boiling points. For example, naphthalene has a low melting point of 80°C.
- (iii) Covalent compounds are usually insoluble in water but they are soluble in organic solvents. For example, naphthalene is insoluble in water but dissolves in organic solvents like ether.
- (iv) Covalent compounds do not conduct electricity. This means that covalent compounds are non-electrolytes. Covalent compounds do not conduct electricity because they do not contain ions.

2.3.5 DIFFERENCE BETWEEN IONIC AND COVALENT COMPOUNDS

Ionic compounds			valent compounds
1.	lonic compounds are usually crystalline	1.	Covalent compounds are usually liquids or
	solids.		gases. Only some of them are solids.
2.	Ionic compounds have high melting points	2.	Covalent compounds have usually low
	and boiling points. That is, ionic	-	melting points and boiling points. That is,
	compounds are non-volatile.		covalent compounds are usually volatile.
3.	lonic compounds conduct electricity when	3.	Covalent compounds do not conduct
	dissolved in water or melted.		electricity.
4.	lonic compounds are usually soluble in	4.	Covalent compounds are usually insoluble
	water.		in water (except, glucose, sugar, urea,
			etc.).
5.	lonic compounds are insoluble in organic	5.	Covalent compounds are soluble in
	solvents (like alcohol, ether, acetone, etc.).		organic solvents.

3. OCCURRENCE OF METALS

The earth's crust is the major source of metals. Sea water also contains some soluble salts of metals like sodium chloride, magnesium chloride etc.

Some metals are found in the earth's crust in free state where as some are found in the form of their compounds. The metals at the bottom of activity series are least reactive. Therefore, they are often found in free state. For e.g. gold, silver, platinum, copper.

The metals at the top of activity series (K, Na, Ca, Mg etc.) are so reactive, that they are never found in nature as free elements but are found in the form of their compounds. The metals in the middle of activity series (Al, Zn, Fe, Pb etc.) are moderately reactive. They are found in earth's crust mainly as oxides, sulphides or carbonates.

Thus, the metals occur in nature in two forms :





- (a) Native State : A metal is said to exist in native state if it exists in the elementary form in the crust of the earth.
- (b) Combined State : A metal is said to exist in combined state if it is found in the earth's crust in the form of its stable compounds.

3.1 MINERALS AND ORES

In the combined state, the metals are found in the crust of earth as oxides, carbonates, sulphides, silicates, phosphates etc.

The inorganic elements or compounds which occur naturally in the earth's crust are known as minerals. At some places, the minerals contain a very high percentage of a particular metal and the metal can be profitably extracted from it. Those minerals from which a metal can be profitably extracted are called ores.

When an ore is mined from earth, it is always found to be contaminated with sand and rocky materials. These impurities of sand and rocky materials present in the ore are known as gangue.

3.2 TYPES OF ORES

Most of the ores are either oxides or sulphides. Some of the ores are also carbonates and halides. The various types of ores and their examples are given below in the table :

	T di			
S.No.	Туре	Examples		
1.	Oxide ores	Haematite (Fe_2O_3)		
		Bauxite (AI_2O_3 . $2H_2O$)		
		Magnetite (Fe ₃ O ₄)		
2	Sulphide ores	Copper pyrites (CuFeS ₂)		
		Iron pyrites (FeS ₂)		
		Zinc blende (ZnS)		
		Cinnabar (HgS)	3	
3.	Carbonate ores	Limestone (CaCO ₃)		
		Siderite (FeCO ₃)		
		Calamine (ZnCO ₃)		
4.	Halide ores	Rock salt (NaCl)		
		Horn silver (AgCl)		
		Fluorspar (CaF ₂)		
		Fluorspar (CaF ₂)		

Various Types of Ores

Some common ores are given in table below :

Some Common Ores

S.No.	Elements	Ores
1.	Iron	Haematite (Fe ₂ O ₃)
		Magnetite (FeO . Fe ₂ O ₃)
		Iron pyrites (FeS ₂)
2	Aluminium	Bauxite (Al ₂ O ₃ . 2H ₂ O)
3.	Manganese	Pyrolusite (MnO ₂)
4.	Calcium	Limestone (CaCO ₃)
5.	Limestone	Dolomite (MgCO ₃ . CaCO ₃)



6.	Copper	Copper pyrites (CuFeS ₂)
7.	Mercury	Cinnabar (HgS)
8.	Zinc	Zinc blende (ZnS),
		Calamine (ZnCO ₃)
9.	Lead	Lead glance (PbS)
10.	Sodium	Rock salt (NaCl)
11.	Silver	Horn silver (AgCl)

4. METALLURGY

After the mining of the ore from the ground, it must be decomposed to give pure metal. This is known as extraction.

The extraction of metals from their ores and then refining them for use is known as metallurgy. The process of metallurgical operations consists of mainly three steps :

- 1. Concentration of ore or enrichment of ore
- 2. Reduction of concentrated ore
- 3. Refining

4.1 ENRICHMENT OF ORE

Ores mined from the earth usually contain large amounts of earthy impurities like soil, sand, etc., called gangue. The impurities must be removed from the ore prior to the extraction of the metal. The processes used for removing the gangue from the ore are based on the differences between the physical or chemical properties of the gangue and the ore. The various methods used for the enrichment of ores are explained below :

4.1.1 Hydraulic washing

It is used for the enrichment of oxide ores. It is based on the difference in the densities of the gangue and the ore particles. The gangue particles are generally lighter as compared to ore particles. In this process, the crushed and finely powdered ore is washed with a stream of water. The lighter gangue particles are washed away, leaving behind the heavier ore particles.





Hydraulic Washing

4.1.2 Froth Floatation Process

It is used to separate the gangue from the sulphide ores, especially those of copper, zinc and lead. In this process, the finely powdered ore is mixed with water in a large tank to form a slurry. Then some pine oil is added to it. The sulphide ores are preferentially wetted by the pine oil, whereas the gangue particles are wetted by the water. When air is blown through the mixture, the lighter oil froth carrying the metal sulphides rises to the top of the tank and floats as scum. It is then skimmed off and dried. The gangue particles being heavier, sink to the bottom of the tank.



4.1.3 Electromagnetic Separation

It is used to enrich the magnetic ores (iron ore). The separation is done by using magnetic separators. A magnetic separator consists of a leather belt that moves over two rollers, out of which one is an electromagnet. The finely powdered ore is dropped over the moving belt at one end. When the ore falls down from the moving belt at the other end, the magnetic portion in the ore is attracted by the magnet and forms a heap nearer to the roller. The non-magnetic impurities fall away from the magnetic ore and form a separate heap.





Magnetic separation : The finely powdered ore is dropped over a moving belt which passes over a magnetic roller. The particles attracted by the magnet form a separate pile.

4.1.4 Chemical Separation

The process of chemical separation makes use of differences between the chemical properties of the gangue and the ore.

For example, bauxite, $AI_2O_3 \cdot 2H_2O$, is an impure form of aluminium oxide. The impurities mainly present in it are iron (III) oxide (Fe₂O₃) and sand (SiO₂). The iron (III) oxide gives it a brown-red colour. Baeyer's method is used to obtain pure aluminium oxide from bauxite ore. In this method, the finely powdered ore is treated with hot sodium hydroxide solution. The aluminium oxide present in bauxite ore reacts with sodium hydroxide to form water soluble sodium aluminate.

 $\begin{array}{rrrr} AI_2O_3(s) & + & 2NaOH\,(aq) & \rightarrow & 2NaAIO_2\,(aq) & + & H_2O(I) \\ (From impure bauxite ore) & & & & & & \\ \end{array}$

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

SiO ₂	+	2NaOH	\rightarrow	$Na_2 SiO_3$	+	H ₂ O
				Sodium silicate		

The filtrate containing sodium aluminate and sodium silicate is then stirred with some aluminium hydroxide to induce the precipitation of aluminium hydroxide. The impurity, silica, remains dissolved as sodium silicate in the solution.

Na AlO₂ (aq) + 2H₂O (I)
$$\rightarrow$$
 Al(OH)₃(s) + NaOH(aq)
Aluminium hydroxide

Aluminium hydroxide is then filtered off, washed, dried and ignited to get pure aluminium oxide, which is called alumina.

2AI (OH)₃ (s)
$$\longrightarrow$$
 Al₂O₃ (s) + 3H₂O (g)

11--4

Metal

K Na Ca

Mg Al - Alumina

4.2 REDUCTION OR CONVERSION OF CONCENTRATED ORE INTO METAL

The method used for the extraction of the metal from the concentrated ore depends upon the nature of the metal. Based on their reactivity, the metals have been grouped into three categories :

- (a) Metals of low reactivity (b) Metals of medium reactivity
- (c) Metals of high reactivity

The activity series and related metallurgy is given below :

Method of extraction

Electrolysis of molten chloride or oxide



4.2.2 Extraction of Metals in the Middle of Activity Series

The metals in the middle of the reactivity series such as iron, zinc, lead, copper etc. are moderately reactive.

These metals are found in nature in the form of their oxide, sulphide or carbonate ores. Further, as it is easier to reduce oxides than sulphides or carbonates, therefore, the sulphide and carbonate ores are first converted into the corresponding metal oxides. Thus, the different steps involved for the extraction of the metal from the concentrated ore are as follows :

(a) Conversion of the carbonate or sulphide ore into metal oxide. This is done by either of the following two methods :



(i) Calcination (For carbonate ores). It is the process of heating the ore strongly in the absence of air. The metal carbonate decomposes to form metal oxide. For example,

 $ZnCO_{3}(s) \xrightarrow[(Absence of air)]{Heat} (Absence of air) \\ (Calamine-ore of Zn) ZnO(s) + CO_{2}(g)$

(ii) Roasting (For sulphide ores). It is the process of heating the ore strongly in the presence of excess of air. As a result, the sulphide ore is converted into metal oxide. For example,

2 ZnS(s)	+	$3O_{2}(g)$	Heat	2ZnO(s)	+	2SO ₂ (g)
Zinc sulphide		2.007	Presence of	. ,		2.007
(Calamine-ore of Zn)		From air	excess of air	Zinc oxide		Sulphur dioxide

- (b) Reduction of the metal oxide to metal. As these metals are moderately reactive, their oxides cannot be reduced by heating alone. Hence, their oxides are reduced to metals by using a suitable reducing agent such as carbon or some highly reactive metals like sodium, calcium, aluminium etc. A few examples are :
 - (i) Reduction by heating with carbon. Carbon in the form of coke or coal is the cheapest reducing agent. It combines with the oxygen of the metal oxide forming carbon monoxide. As a result, metal oxide is reduced to metal. For example, oxides of iron and zinc are reduced to their respective metals by heating with coke.

Carbon monoxide formed also acts as reducing agent and further reduces the metal oxide to metal.

 $\begin{array}{c|c} Fe_2O_3(s) + & 3CO(g) \\ \hline Ferric \ oxide \end{array} \xrightarrow[low dashed beta]{Heat}} & 2Fe(s) + & 3CO_2(g) \\ \hline & Iron \end{array}$

 $\begin{array}{ccc} ZnO(s) & + CO(g) & \xrightarrow{Heat} & Zn(s) & + & CO_2(g) \\ Zinc & zinc & & Zinc & Carbon \ dioxide \end{array}$

(ii) Reduction by heating with aluminium. Oxides of certain metals, e.g., manganese oxide (MnO₂), chromium oxide (Cr₂O₃), etc. cannot be satisfactorily reduced by heating with carbon. However, these oxides are easily reduced to their corresponding metals by hearing with aluminium powder (as aluminium is more reactive than manganese or chromium). As a result, aluminium is converted into aluminium oxide whereas the metal oxide is reduced to the metal. The reaction is highly exothermic. The heat evolved is so high that the metal is obtained in the molten state.



Heat 4 Al(s) $3 \text{MnO}_2(s)$ 3Mn(I) + $2 Al_2 O_3(s)$ Aluminium Manganese dioxide Aluminiúm Manganese (powder) (melt) oxide $Cr_2O_3(s)$ 2 Cr (I) 2 Al Aluminium $AI_2O_3(s)$ Aluminiùm Chromium oxide Chromium (melt) (powder) oxide

Similarly, when iron (III) oxide is heated with aluminium powder (the mixture is ignited by inserting a magnesium ribbon and then burning it), the heat evolved is so high that iron obtained melts.

 $\begin{array}{c|c} \mathsf{Fe}_2\mathsf{O}_3(s) &+& 2\,\mathsf{AI}(s) & \xrightarrow{ \mathsf{Ignited} } 2\,\mathsf{Fe}\,(\mathsf{I}) &+& \mathsf{AI}_2\mathsf{O}_3\,(s) \\ \mathsf{Iron}\,(\mathsf{III})\,\mathsf{oxide} & & \mathsf{Aluminium} & \xrightarrow{ \mathsf{Ignited} } 2\,\mathsf{Fe}\,(\mathsf{I}) &+& \mathsf{AI}_2\mathsf{O}_3\,(s) \\ \end{array}$

This reaction is, therefore, used for welding the broken parts of iron machinery, railway tracks etc. The reaction is known as **thermite reaction**.

The reduction of metal oxides to metal using aluminium as the reducing agent is called aluminothermy.



Thermit process for joining railway tracks

4.2.3 Extraction of Metals towards the Top of Activity Series

The metals high up in the reactivity series are very reactive. They cannot be obtained from their compounds by heating with carbon. For example, carbon cannot reduce the oxides of sodium, magnesium, calcium, aluminium, etc., to the respective metals. This is because these metals have more affinity for oxygen than carbon. These metals are obtained by electrolytic reduction. For example, sodium, magnesium and calcium are obtained by the electrolysis of their molten chlorides.

During electrolysis, the metal ions, being positive are liberated at the cathode (the negatively charged electrode), whereas the chlorine is liberated at the anode (the positively charged electrode). So, during the electrolysis of molten salts, the metals are always produced at the cathode (negative electrode).

For example :



(i) Electrolysis of molten sodium chloride. When sodium chloride melts, it splits into sodium ion (Na⁺) and chloride ions (Cl⁻).



When electricity is passed through the melt, Na⁺ ions go to the cathode whereas CI^- ions are liberated at the anode. Na⁺ ions gain electrons at the cathode and are thus reduced to sodium atoms. CI^- ions lose electrons at the anode and are thus oxidized to chlorine atoms. These chlorine atoms then combine with each other to form chlorine (Cl₂) gas. The reactions taking place are as follows :

Reduction e⁻ Na⁺(I) Na(s) Electron Sodium ions Sodium metal At Cathode : $\xrightarrow{\text{Oxidation}}$ $CI^{-}(I)$ CI(g) + e⁻ Chloride ion Chlorine atom At Anode : $\begin{array}{c} Cl\left(l\right) & + & Cl(g) & \longrightarrow & Cl_2\left(g\right) \\ Chloride atoms & & & Chlorine gas \end{array}$

Thus, sodium metal is obtained at the cathode whereas chlorine gas is liberated at the anode.

Similarly, calcium and magnesium are also obtained by the electrolysis of their fused chlorides.



(ii) Electrolysis of molten alumina. Alumina (Al₂O₃) is a stable compound. Hence, it cannot be reduced to aluminium by heating with carbon. It is obtained by electrolytic reduction of the molten alumina. Molten alumina (Al₂O₃) contains aluminium ions (Al³⁺) and oxide ions (O²⁻). On passing electricity, the reactions taking place at the electrodes are as follows :



Thus, aluminium is deposited at the cathode (due to reduction of aluminium ions) whereas oxygen gas is liberated at the anode.

4.3 REFINING OF METALS

The metals produced by various reduction processes described above are not very pure. They contain impurities. Therefore, they must be purified to obtain pure metals. *The process of purification of impure metals is known as refining of metals.* The method used for refining an impure metal depends upon (i) the nature of the element, and (ii) the nature of impurities present in it. Therefore, different methods are used for refining different metals. The impure metals are purified by any one of the following methods:

4.3.1 Liquation

This method is used for refining those metals which have low melting points such as tin, lead, etc. In this method, a sloping hearth is used. The hearth is maintained at a



temperature that is slightly above the melting point of the metal. When the impure metal is placed at the top of the hearth, it melts and flows down the hearth. The solid impurities, whose melting point is higher than the melting point of the metal, are left behind on the hearth. The pure metal is collected at the bottom of the hearth.

4.3.2 Distillation

The more volatile metals like mercury and zinc are purified by this method. The impure metal is heated in a retort. The pure metal distils over and is condensed in a receiver. The impurities are left behind in the retort.





4.3.3 Electrolytic Refining

This is the most widely used method for refining impure metals. Many metals, such as copper, zinc, tin, nickel, silver, gold, etc., are refined electrolytically. In this process, the impure metal is made as anode and a thin strip of pure metal is made as cathode. A solution of the metal salt is used as an electrolyte. On passing the current through the electrolyte, the pure metal from the anode dissolves into the electrolyte. An equivalent amount of pure metal from the electrolyte is deposited on the cathode. The soluble impurities go into the solution, whereas the insoluble impurities settle down at the bottom of the anode and are known as anode mud.



Electrolytic refining of copper. The electrolyte is a solution of acidified copper sulphate. The anode is impure copper, whereas, the cathode is a strip of pure copper. On passing electric current, pure copper is deposited on the cathode.

In case of copper, the block of impure copper is made as anode. The cathode is made of a thin plate of pure copper. The electrolyte is a solution of copper sulphate, containing a small amount of dilute sulphuric acid. On passing electric current, copper dissolves from the anode into the electrolyte. An equivalent amount of copper from the electrolyte is then deposited on the cathode. Crude copper contains very small amounts of iron, silver and gold. The more reactive metals, such as iron, dissolve in the solution and remain there. The less reactive metals, such as silver and gold, fall below the anode in the electrolytic cell as anode mud. These metals are recovered in the native state. The following reactions take place :



A summary of the methods used and the steps applied for the extraction of different types of metals is given below :





5. CORROSION

One of the most destructive and annoying processes that occur in nature is the corrosion of metals. This process takes place on the surface of metal when they are exposed to air. Due to corrosion, small holes appear on the surface of the metal and the strength of the metal goes on decreasing Thus, corrosion brings about slow destruction of the metal . Corrosion is defined as :

The process of slow eating up of metals due to their conversion into oxides, carbonates, sulphides, sulphates etc by the action of atmospheric gases and moisture.

Generally, more reactive the metal, the more easily it corrodes. Metals which lie high up in the activity series (K, Mg, Al, Zn, Fe etc.) easily undergo corrosion while noble metals (such as gold and platinum) which lie at the bottom of the activity series do not corrode readily.

5.1 CORROSION OF IRON

The corrosion of iron is known as rusting. When ordinary iron is exposed to air in presence of moisture, it gets corroded and forms reddish brown scale which is known as rust. Rust is mainly hydrated ferric oxide ($Fe_2O_3 \cdot xH_2O$). Rust once formed, causes more and more rusting until the whole of the metal is eaten up.



$$4Fe + 3O_2 \xrightarrow{Moisture} 2Fe_2O_3.xH_2O \xrightarrow{(Rust)} Hydrated ferric oxide$$

The two conditions necessary for rusting of iron to occur are :

- 1. Presence of moisture, and
- 2. Presence of air (oxygen)

Experiment to show that Rusting of Iron requires both, Air and Water

Take three test tubes and add a few clean iron nails to each of them. To the tube A add distilled water and place a cork in its mouth. To the tube B add boiled distilled water. Then pour about 1 ml oil and place a cork in its mouth. The oil will float on the surface of water and prevent the air from entering. In the tube C put some anhydrous calcium chloride. Anhydrous calcium chloride will absorb the moisture from the air in the test tube.



Investigating the conditions under which iron rusts. In tube A, both air and water are present. In tube B, there is no air dissolved in the water. In tube C, the air is dry.

After a few days it is observed that the nails in the tube A rust but the nails in the tubes B and C do not rust. In tube A both air and water are present. In tube B, nails are exposed



only to water while in tube C, nails are exposed to dry air. Thus, presence of water and air both is essential for rusting.

The following factors further catalyse the process of rusting.

- 1. Presence of carbon dioxide,
- 2. Presence of acids,
- 3. Presence of impurities in the iron.

Due to rusting, iron object loses its strength.

5.2 PREVENTION OF RUSTING

The wasting of iron objects due to rusting causes a big loss to the country's economy, so it must be prevented. The various common methods of preventing the rusting of iron (or corrosion of iron) are given below :

- 1. Rusting of iron can be prevented by painting. The most common method of preventing the rusting of iron (or corrosion of iron) is to coat its surface with a paint. When a coat of paint is applied to the surface of an iron object, then air and moisture cannot come in contact with the iron object and hence no rusting takes place. The iron articles such as window grills, railings, steel furniture, iron bridges, railway coaches, ships, and bodies of cars, buses and trucks, etc., are all painted to protect them from rusting.
- 2. Rusting of iron can be prevented by applying grease or oil. When some grease or oil is applied to the surface of an iron object, then air and moisture cannot come in contact with it and hence rusting is prevented. For example, the tools and machine parts made of iron and steel are smeared with grease or oil to prevent their rusting.
- 3. Rusting of iron can be prevented by galvanization. The process of depositing a thin layer of zinc metal on iron objects is called galvanization. Galvanisation is done by dipping an iron object in molten zinc metal. A thin layer of zinc metal is then formed all over the iron object. This thin layer of zinc metal on the surface of iron objects protects them from rusting because zinc metal does not corrode on exposure to damp air. The iron sheets used for making buckets, drums, dust-bins and sheds (roofs) are galvanized to prevent their rusting. The iron pipes used for water supply are also galvanized to prevent rusting.
- 4. Rusting of iron can be prevented by tin-plating and chromium-plating. Tin and chromium metals are resistant to corrosion. So, when a thin layer of tin metal (or chromium metal) is deposited on iron and steel objects by electroplating, then the iron and steel objects are protected from rusting. For example, tiffin-boxes made of steel are nickel-plated from inside and outside to protect them from rusting. Tin is used for plating tiffin-boxes because it is non-poisonous and hence does not contaminate the food kept in them. Chromium-plating is done on bicycle handle bars and car bumpers made of iron and steel to protect them from rusting and give them a shiny appearance.
- **5.** Rusting of iron can be prevented by alloying it to make stainless steel. When iron is alloyed with chromium and nickel, then stainless steel is obtained. Stainless steel does not rust at all. Cooking utensils, knives, scissors and surgical instruments, etc.,



are made of stainless steel and do not rust at all. But stainless steel is too expensive to be used in large amounts.

6. ALLOYS

An alloy is a homogeneous mixture of two or more metals or metals and non-metals.

The alloys are prepared to improve the mechanical properties like hardness and tensile strength, to resist the atmospheric and chemical corrosion, to produce sound and workable casting, and to lower the melting point.

An alloy containing mercury as one of the metals is known as an amalgam.

Some examples of common alloys are :

- (a) Brass containing copper and zinc metals.
- (b) Bronze containing copper and tin metals.
- (c) Stainless steel containing iron, carbon and chromium.

6.1 PREPARATIONS OF ALLOYS

Alloys are generally prepared by any of the following methods :

- 1. By first melting the main metal and then dissolving the other elements in it in a definite proportion. It is then cooled to room temperature.
- 2. By melting the different components of the alloy together and then allowing it solidify.
- 3. By compressing together under high pressure.
- 4. By simultaneous electrodeposition of metals.

6.2 TYPES OF ALLOYS

There are two types of alloys :

- 1. **Ferrous alloys.** These are the alloys which contain iron as one of the constituents, such as nickel steel, stainless steel, etc.
- 2. **Non-ferrous alloys.** These are the alloys which do not contain iron as one of the constituents, such as bronze, brass, etc.

6.3 OBJECTIVES OF PREPARATION OF ALLOYS

The main objects of making alloys are given below :

- 1. *To increase the hardness.* For example, gold is alloyed with copper to increase the hardness. The ornaments made from pure gold are easily deformed. In order to avoid this, gold is alloyed with copper to make ornaments.
- 2. *To increase the tensile strength.* For example, magnalium which is an alloy of magnesium and aluminium has greater tensile strength than the constituent metals.
- 3. *To increase resistance to corrosion.* For example, stainless steel which is an alloy of iron, chromium, nickel and carbon does not undergo corrosion.
- 4. To modify chemical reactivity. For example, sodium amalgam is less reactive than sodium.
- 5. To lower the melting point. For example, soldering metal (an alloy of tin and lead).
- 6. To produce good casting. For example type metal.

6.4 ALLOYS OF GOLD



:

Pure, gold, known as 24 carat gold is very soft. It is, therefore, not suitable for making jewellery. It is alloyed with either silver or copper to make it hard. Generally, in India, 22 carat gold is used for making ornaments. It means that 22 carat pure gold is alloyed with 2 parts of either copper or silver.

6.5 COMPOSITION, PROPERTIES AND USES OF SOME COMMON ALLOYS

The composition, properties and uses of some common alloys are given in the table below

S.No.	Alloy	Percentage	Properties	Uses
		composition		
1	Steel	Iron-98 5 to 99 5%	Hard tough strong	Construction of
	otool	Carbon_0 1 to 1 5%		shins bridges
		Traces of other		vehicles cables
		impurities		machine narts
		Impunites		nachine parts,
				other alloy steels
2	Stainlass staal	Iron 74%	Hard Doos not rust	Surgical
Ζ.	Stamless steel	11011—74% Chromium 190/	naru, Does not rust	Surgical
		Chromium—18%	or corrode	Instruments. For
		NICKEI—8%		nousenoia
				utensiis, snaving
				blades, watch
				cases.
3. Alnico	Alnico	Aluminium—12%	Highly magnetic	For permanent
		Nickel—20%		magnets
		Cobalt—5%		
4. Magnalium		Magnesium—5%	Tough, very light	Machined articles,
		Aluminium—95%	does not stick to the	cheap balances
			tools	
5.	Duralumin or Duralium	Magnesium—2%	Light, tough, resists	For making air
		Aluminium—92%	corrosion	crafts, pressure
				cooker.
6. Brass		Copper—80%	Malleable, Resistant	Screws, nuts and
		Zinc—20%	to corrosion, good	bolts, Utensils and
			ductility, high	cartridge caps
			strength	
7. Bronze		Copper—90%	Resistant to	Statues, ships,
		Tin—10%	corrosion, hard,	medals, coins
			high tensile strength	
8. Solder		Lead—50%,	Low melting point	For welding
		Tin—50%		electric wires and
				by dentists for
				filling in teeth.

6.6 THE WONDER OF ANCIENT INDIAN METALLURGY



The iron pillar near the Qutub Minar in Delhi was made around 400 BC by the iron workers of India. They had developed a process which prevented wrought iron from rusting. This is likely because of formation of a thin film of magnetic oxide (Fe_3O_4) on the surface, as a result of finishing treatment given to the pillar, painting it with a mixture of different salts, then heating and quenching. The iron pillar is 8 m high and weighs 6 tonnes (6000 kg).

