Metals & Non-metals

Chapter Outline

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- Non-metals
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- Occurrence of Metals
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Metal Machinery

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Metals & Non-metals

INTRODUCTION

Most of the elements are widely distributed in nature either in free state or in the combined state. The earth crust and the sea water are the treasure houses of elements. These elements are extracted from their different sources by different extraction methods.

The first classification of elements, categorized elements into metals and non-metals on the basis of physical and chemical properties.

ELECTRONIC CONFIGURATON OF METALS & NON-METALS

Metals :

Metals are the electropositive elements which possess tendency of losing one or more of their valence electrons attaining octet and thereby forming cations.

The metals usually have 1, 2 or 3 electrons (some have 4 electrons) in their valence shell, lose their valence electrons and form cations (positive ions) on gaining energy from an external source. e.g.,

$$\underset{(2,8,1)}{Na} \rightarrow \underset{(2,8)}{Na^+} + e^- \qquad (\text{Valency} = 1)$$

$$\underset{(2,8,2)}{Mg} \rightarrow \underset{(2,8)}{Mg^{2+}} + 2e^- \qquad (\text{Valency} = 2)$$

$$\underset{(2,8,3)}{Al} \rightarrow \underset{(2,8)}{Al^{3+}} + 3e^- \qquad (\text{Valency} = 3)$$

The number of electrons loss by an atom of metal to form the positive ion is called the valency of that metal. The ion of the metal carries the same units of positive charge. For example, sodium magnesium and aluminium atoms lose 1, 2 and 3 electrons respectively to form Na^+, Mg^{2+} and Al^{3+} and show valency of 1, 2 and 3 respectively.

Non-Metals :

Non-metals are the electronegative elements which possess tendency to form anion (negative ions) by gaining one or more electrons.

The atoms of non-metals usually have 4, 5, 6 or 7 electrons in their valence shell. In order to gain stable electronic configuration or to attain octet, they possess tendency to gain electrons and thereby form anions (negative ions). e.g.,

$$\underset{(2,5)}{N} + 3e^- \rightarrow \underset{(2,8)}{N^{3-}} \qquad (\text{Valency} = 3)$$

$$\underset{(2,6)}{O} + 2e^- \rightarrow \underset{(2,8)}{O^{2-}} \qquad (\text{Valency} = 2)$$

$$\underset{(2,8,7)}{Cl} + e^- \rightarrow \underset{(2,8,8)}{Cl^-} \qquad (\text{Valency} = 1)$$

The number of electrons gained by an atom of a non-metal to form the negative ion is called the valency of that non-metal. The ion thus formed carries the same units of negative charge.

For example, oxygen atom gain 2 electrons to form O^{2-} ions. Therefore, the valency of oxygen is 2 and O^{2-} ion carries two unit negative charge.

POSITION OF METALS AND NON-METALS IN THE PERIODIC TABLE

In periodic table, metals are placed on the left hand side, in the middle and at the bottom of the periodic table, whereas non-metals are placed on the extreme right hand side of the periodic table.

There are some elements which exhibit property of both metals and non-metals. Such elements are called metalloids. e.g., B, Si, Ge, As, Sb, Te, Po and At.

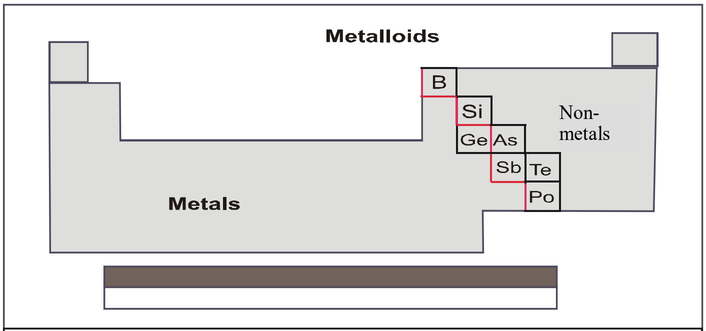

Fig. Position of metals, non-metals and metalloids in the periodic table.

Boost your knowledge

- Hydrogen, although a non-metals but resembles also to alkali metals is an exception which is placed on the extreme left of the periodic table.

PHYSICAL PROPERTIES OF METALS

(i) State : Metals are usually solids at ordinary temperature. For example, iron, copper, aluminium, silver and gold.

Except mercury which is the only liquid metal at room temperature. However, some metals like caesium and gallium have very low melting point and occur as liquid generally.

(ii) Metallic Lustre : They are lustrous in appearance and take a high polish. For example, gold, silver and copper are shiny metals and they can be polished. The shiny appearance of metals makes them useful in making jewellery and decoration. When a metal has been kept exposed to air for a long time, then it gets a dull appearance. The metals lose their shine or brightness due to the formation of a thin layer of oxide, carbonate or sulphide on the surface and thus, the metal surface gets corroded. But if we rub the dull surface of metal object with sand paper, then the outer corroded layer is removed and the metal becomes shiny and bright once again. For example, Cu, Ag, Au, Mg, Al, etc., all are lustrous in nature.

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☛ When light falls on the surface of a metal, the atoms absorb photons as energy. They get excited and start vibrating. These vibrating electrons release energy as light. Therefore, metals surface shine.

(iii) Hardness : Metals are generally hard e.g., copper, silver, gold, etc.

Exceptions : Sodium and potassium are soft metals which can be easily cut with a knife. The extent of hardness which a metal possess depends upon the strength of metallic bond. Stronger the metallic bond, harder is the metal.

(iv) Malleability : Metals are malleable i.e., they can be beaten into thin sheets. This property is called malleability. Gold and silver are the most malleable metals. Aluminium and copper metals are also highly malleable metals. All these metals can be beaten into a very thin sheets using a hammer. For example, silver can be hammered into thin silver foils which are used for decorating sweets. Aluminium is also highly malleable and can be converted into thin aluminium foils which are used for packing food items, medicines, cigarettes, etc. Aluminium sheets are also used for making cooking utensils. Copper is also highly malleable and so copper sheets are used to make utensils. Iron can also the hammered into sheets, which are used for making boxes, drums and water tanks, etc.

(v) Ductility : Metals are ductile i.e., metals can be drawn into wires. The ability of metals to be drawn into thin wires is called ductility. Gold is the most ductile metals. For examples, 1 gram of gold can be drawn into a thin wire of about 2 km long. Silver is also among the best ductile metals. Copper and aluminium metals are very ductile and can be drawn into wires which are used as electric wires. Iron, magnesium and tungsten can be drawn into thin wires. Iron wires are used for making wire gauges, magnesium wires are used in science experiments and tungsten wires are used for making the filament of electric bulbs.

(vi) Tensile strength : Metals have high tensile strength (ability to withstand strain i.e., capacity to bear considerable load without getting broken).

Iron has high tensile strength hence, it is used in the construction of bridges, buildings, railway lines, girders, machines, vehicles, chains, etc.

Exceptions : Zinc, mercury, gallium, sodium, potassium and calcium which are not tenacious.

(vii) Melting points and Boiling points : Metals have usually high melting points and boiling points.

For example:- Iron has high melting point of 1535ºC and copper has a melting point of 1083ºC. This is due to the reason that the constituent atoms are very closely packed and the attractive forces between them are very strong.

Exceptions : Sodium and potassium metals have low melting points of 98ºC and 64ºC respectively. Gallium and caesium metals also have very low melting point of 30ºC and 28ºC respectively. These two metals will melt if we keep them on our palm.

(viii) Density : Metals possess high densities i.e., metals are heavy substances. The density of iron is 7.8 g/cm^3 which is quite high.

Exception : Sodium and potassium have low densities (of 0.97 g/cm^3 and 0.86 g/cm^3) respectively, they are very light metals.

(ix) Brittleness : Metals are hard but not brittle i.e., they cannot break into pieces when hammered or stretched.

Exception : Zinc

(x) Thermal Conductivity : The conductivity of heat is called thermal conductivity. Metals are generally good conductor of heat. Copper ranks second after silver in thermal conductivity and is used in making cooking utensils and water boilers because it is a good conductor of heat.

Exception : The poorest conductor of heat among metals is lead.

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- The thermal conductivity of the metals is due to the kernels and the valence electrons in the electron sea. When we heat one end of the metal sheet, the kernels and valence electrons present in that portion move towards the colder region and transfer heat to that region by conduction.

(xi) **Electrical Conductivity :** Metals are good conductors of electricity i.e., they allow electricity to pass through them easily. Metals offer very little resistance to the flow of electric current and hence, show high electrical conductivity. Silver is the best conductor of electricity. Copper is next to silver in electrical conductivity followed by gold, aluminium and tungsten. The electric wires are made of copper and aluminium metals because they are good conductors of electricity. Metals are good conductors of electricity because they contain free electrons. These free electrons can move easily through the metal and conduct electric current.

Exceptions : Mercury, lead and bismuth are poor conductors of electricity.

The electric wires that carry current in our homes have a covering of plastic such as polyvinyl chloride (PVC). Polyvinyl chloride is an insulator and does not allow electric current to pass through it and protect us from electric shock if touched accidentally when an electric current is flowing through the wire.

(xii) **Sonority :** Metals are sonorous i.e., metals make a sound when hit with an object. The property of metals of being sonorous is called sonority. It is due to this property that metals are used for making bells and string of musical instruments like Sitar and Violin.

(xiii) **Colour :** Metals usually have a silver or grey colour except gold which has a yellow colour and copper which has a reddish brown colour.

PHYSICAL PROPERTIES OF NON-METALS

(i) **State :** Non-metals are liquids, gases or brittle solids. For example, carbon, sulphur and phosphorus are solid non-metals, bromine is a liquid non-metal whereas hydrogen, oxygen, nitrogen and chlorine are gaseous non-metals.

(ii) **Lustre :** Non-metals do not have luster but have a dull appearance. **For example,** sulphur and phosphorus. **Exceptions :** Graphite and iodine are non-metals having a lustrous appearance.

(iii) **Hardness :** Non-metals are quite soft. For example, sulphur and phosphorus.

Exception : Diamond (which is an allotropic form of carbon) is the hardest substance known.

(iv) **Malleability:** Since, non-metals are not malleable, they cannot be beaten with a hammer to form their sheets. **For example**, sulphur and phosphorus are non-malleable.

Exception : Carbon fibre.

(v) **Ductility :** Since, non-metals are not ductile, they cannot be stretched to form their wires. For example, sulphur and phosphorus are non-ductile.

Exception : Carbon fibre.

(vi) **Tensile strength :** Non-metals have low tensile strength. For example, graphite a non-metal is not strong but breaks into pieces when a weight is put on a graphite sheet.

Exception : Carbon fibre.

(vii) **Melting points and Boiling points :** Non-metals usually have low melting points and boiling points.

Exception : Diamond, silicon, boron and graphite are non-metals but have high melting points.

(viii) **Density :** Non-metals have low densities i.e., non-metals are light. Densities of sulphur and phosphorous are quite low (about 1.8g/cm^3 and 2.0 gm/cm^3 respectively).

Can you think why ?

- Non-metals have low density.

(ix) **Brittleness :** Non-metals are brittle. Non-metals on being beaten with hammer break into pieces. The property of being brittle (breaking easily) is called brittleness.

(x) **Thermal conductivity :** Non-metals are poor conductors of heat.

Exception : Diamond and Graphite.

(xi) **Electrical conductivity :** Non-metals are non-conductors of electricity. The gases nitrogen, oxygen, carbon-dioxide, etc., which constitute the air are all poor conductors of electricity.

Exceptions : Graphite and gas carbon. Since, graphite is a good conductor of electricity, it is used for making electrodes.

(xii) **Sonority :** Non-metals are non-sonorous. They do not produce sound when hit with an object.

(xiii) **Colour :** Non-metals have different colours. **For example**, sulphur is yellow, phosphorus is white or red, graphite is black, chlorine is greenish-yellow etc.

Table : Comparison of Physical properties of Metals and Non-Metals

Property		Metals	Non-metals
1.	Physical state	All metals exist as solid except mercury (liquid).	Non-metals exist in all the three states of matter.
2.	Lustre	Metals have a characteristic metallic lustre. They can be polished to give a shining surface.	Except iodine and graphite, non-metals are non-lustrous and cannot be polished.
3.	Density	High density except lithium, sodium and potassium etc.	Low density except diamond (carbon)
4.	Melting and boiling points	High melting and boiling points, except sodium, potassium, mercury and gallium	Low melting and boiling points except carbon, boron and silicon.
5.	Malleability	Usually malleable.	Non-malleable.
6.	Ductility	Usually ductile and can be drawn into wires.	Not ductile. However, carbon fibres have been obtained.
7.	Brittleness	Usually not brittle except zinc.	Generally brittle.
8.	Hardness	Usually hard except alkali metals and lead.	Usually soft except diamond (carbon) which is extremely hard.
9.	Tensile strength	High tensile strength.	Low tensile strength except carbon fibers.
10.	Electrical conductivity and Thermal conductivity	Good conductors of electricity and heat.	Non-conductors of electricity and heat. Except graphite.
11.	Alloy formation	Metals form alloys. Mercury form amalgams with metals.	Only carbon forms alloy with iron.
12.	Sonorous nature	Sonorous nature.	Not sonorous.

CHEMICAL PROPERTIES OF METALS

(A) Electropositive Nature :

The atoms of metals possess four or less than four electrons in their valence shell. e.g.,

$_{11}$Na : 2, 8, 1
$_{20}$Ca : 2, 8, 8, 2
$_{13}$Al : 2, 8, 3

Metals lose their valence electrons and form stable positively charged ion or cation during the course of reaction e.g.,

$$\underset{(2,8,8,2)}{Ca} \rightarrow \underset{(2,8,8)}{Ca^{2+}} + 2e^-$$

The number of electrons lost by a metal atom to form a cation is called the valency of that metal.

Due to the formation of cations, metals are called electropositive in nature. Therefore metals show oxidation and metal ions show reduction.

$M \rightarrow M^{+n} + ne^-$ (oxidation)
$M^{+n} + ne^- \rightarrow M$ (Reduction)

(B) Reaction of Metals with oxygen :

On burning a metal in air, it combines with the oxygen of air to form a metal oxide and heat.

Metal + Oxygen → Metal oxide + heat

Almost all metals combine with oxygen to form metal oxides. However, the intensity of the chemical reactivity of the metal differ from metal to metal. Some examples are given below.

(i) **Lithium, sodium, potassium** etc. being highly reactive, react with oxygen even at room temperature to form their respective oxides.

$$\underset{\text{Lithium}}{4Li(s)} + O_2(g) \rightarrow \underset{\text{Lithium oxide}}{2Li_2O(s)}$$

$$2Na(s) + O_2(g) \rightarrow \underset{\text{sodium oxide}}{Na_2O(s)}$$

$$K(s) + O_2\,(g) \rightarrow \underset{\text{Potassium oxide}}{K_2O(s)}$$

Therefore these metals are always stored under kerosene to avoid their contact with air.

(ii) **Magnesium** burns in oxygen on heating to form magnesium oxide, giving a dazzling white flame.

$$\underset{\text{Magnesium}}{2Mg(s)} + O_2(g) \rightarrow \underset{\text{Magnesium oxide}}{2MgO(s)}$$

(iii) **Aluminium** burns on heating air to form aluminium oxide and liberates excessive heat.

$$\underset{\text{Aluminium}}{4Al(s)} + 3O_2\,(g) \rightarrow \underset{\text{Aluminium oxide}}{2Al_2O_3(s)}$$

(iv) **Zinc** burns on strong heating in air to form zinc oxide.

$$\underset{\text{Zinc}}{2Zn(s)} + O_2(g) \rightarrow \underset{\text{Zinc oxide}}{2ZnO(s)}$$

(v) **Copper** and **iron** on continuously heating in oxygen for a long time combines with oxygen to form black copper (II) oxide and brown iron oxide respectively.

$$\underset{\text{Copper}}{2Cu(s)} + O_2\,(g) \rightarrow \underset{\text{Copper (II) oxide}}{2CuO(s)}$$

$$2Fe\,(s) + O_2(g) \rightarrow 2\,FeO(s)$$

- **Solubility of oxides :** The oxides of some alkali metals such as sodium oxide, potassium oxide are soluble in water. However most of the metal oxides e.g. copper oxide, aluminium oxide, lead oxide, mercury oxide etc. are water insoluble or partially soluble.

$$\underset{\text{Sodium oxide}}{Na_2O(s)} + H_2O(\ell) \rightarrow \underset{\text{Sodium hydroxide}}{2NaOH(aq.)}$$

$$\underset{\text{Sodium oxide}}{K_2O(s)} + H_2O(\ell) \rightarrow \underset{\text{Potassium hydroxide}}{2KOH(aq)}$$

Magnesium oxide and calcium oxide are partially soluble in water and form a suspension when shaken vigorously with water. The clear solution can be obtained after filtration.

$$\underset{\substack{\text{Magnesium}\\\text{Oxide}}}{MgO(s)} + H_2O(\ell) \rightarrow \underset{\substack{\text{Magnesium hydroxide}\\\text{(partially soluble)}}}{Mg(OH)_2(aq.)}$$

$$\underset{\text{Calcium oxide}}{CaO(s)} + H_2O\,(\ell) \rightarrow \underset{\substack{\text{Calcium hydroxide}\\\text{(partially soluble)}}}{Ca(OH)_2(aq.)}$$

The surfaces of metals such as aluminium, magnesium, zinc, lead etc. are covered with a thin layer of their respective oxide at room temperature. This oxide layer protects the metal from further aerial oxidation.

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☛ The thick protective oxide layer on metal can also be made by anodizing. In anodizing a clean oxygen gas liberated at the anode reacts with metal to form a thicker oxide layer which protects the metal to a greater extent from further oxidation. This is done specially for metals like aluminium.

Nature of metal oxides :

(i) **Basic Nature :** Lithium oxide, sodium oxide, potassium oxide, magnesium oxide, calcium oxide, barium oxide, copper oxide etc. are all basic oxides and react with acids to form salts.

$$\underset{\substack{\text{Calcium oxide}\\\text{(a basic oxide)}}}{CaO(s)} + \underset{\substack{\text{Hydrochloric acid}\\\text{(an acid)}}}{2HCl(aq.)} \rightarrow \underset{\substack{\text{Calcium chloride}\\\text{(salt)}}}{CaCl_2(aq.)} + H_2O(\ell)$$

These also react with acidic non-metal oxides.

e.g. $$\underset{\text{basic}}{CaO} + \underset{\text{acidic}}{CO_2} \rightarrow \underset{\text{Calcium carbonate}}{CaCO_3}$$

The basic nature of Na_2O can be noticed by adding little amount of water in it. Now add litmus solution, blue colour of litmus in aqueous NaOH confirms the basic nature of Na_2O.

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- **Amphoteric Nature :** Aluminium oxide, zinc oxide, Beryullium oxide etc. are amphoteric and show both acidic as well as basic behaviour. These react with both acids and bases to form salts. For example.

Basic Nature $Al_2O_3(s) + 6HCl(aq.) \rightarrow \underset{\substack{\text{Aluminium chloride}\\\text{(salt)}}}{2AlCl_3(aq)} + 3H_2O(\ell)$

Acidic Nature $Al_2O_3(s) + 2NaOH(aq.) \rightarrow \underset{\text{Sod. aluminium (salt)}}{2NaAlO_2(aq)} + H_2O(\ell)$

Basic Nature $\underset{\substack{\text{Zinc oxide}\\\text{(acting as basic oxide)}}}{ZnO(s)} + \underset{\substack{\text{Hydrochloric acid}\\\text{(an acid)}}}{2HCl(aq.)} \rightarrow \underset{\substack{\text{Znic chloride}\\\text{(salt)}}}{ZnCl_2(aq.)} + H_2O(\ell)$

Acidic Nature $\underset{\substack{\text{Zinc oxide}\\\text{(acting as an acidic oxide)}}}{ZnO(s)} + \underset{\substack{\text{Sod. hydroxide}\\\text{(base)}}}{2NaOH(aq.)} \rightarrow \underset{\substack{\text{Sodium zincate}\\\text{(salt)}}}{Na_2ZnO_2(aq.)} + H_2O(\ell)$

- **Acidic Nature :** Some metal oxides (transition metal oxides) in their higher oxidation state e.g. Mn_2O_7, CrO_3 and WO_3 etc) are acidic in nature.

(C) Reaction of Metals with Water

Metals react with water to form metal oxide or metal hydroxide and give hydrogen. The reactivity of metals towards water depends upon the nature of the metals. Some metals react even with cold water, some react with water only on heating while some metals do not react even with steam. For example.

Metal + Water → Metal oxide + Hydrogen

or

Metal + Water → Metal hydroxide + Hydrogen

The intensity of the reaction of a metal with water depends upon the chemical reactivity of the metal.

(a) Sodium and potassium metals react vigorously with cold water to form sodium hydroxide liberating hydrogen gas.

$$\underset{\text{Sodium}}{2Na(s)} + 2H_2O(\ell) \rightarrow \underset{\text{Sodium hydroxide}}{2NaOH(aq.)} + \underset{\text{Hyhdrogen}}{H_2(g)}\uparrow$$

$$\underset{\text{Potassium}}{2K(s)} + 2H_2O(\ell) \rightarrow \underset{\text{Potassium hydroxide}}{2KOH(aq.)} + \underset{\text{Hydrogen}}{H_2(g)}\uparrow$$

(b) Calcium reacts with cold water to form calcium hydroxide and hydrogen gas. The reaction is less violent. The heat produced in reaction is not sufficient for H_2 to catch fire.

$$\underset{\text{Calcium}}{Ca(s)} + \underset{\text{Cold water}}{2H_2O(s)} \rightarrow \underset{\text{Calcium hydroxide}}{Ca(OH)_2(aq.)} + H_2(g)$$

(c) Magnesium reacts very slowly with cold water but reacts rapidly with hot boiling water forming magnesium hydroxide and hydrogen.

$$\underset{\text{Magnesium}}{Mg(s)} + \underset{\text{Boiling water}}{2H_2O(\ell)} \rightarrow \underset{\text{Magnesium oxide}}{Mg(OH)_2(aq.)} + H_2(g)$$

(d) Metals like zinc and aluminium react only with steam to form their corresponding oxides and hydrogen.

$$\underset{\text{Zinc}}{Zn(s)} + \underset{\text{Steam}}{H_2O(g)} \rightarrow \underset{\text{Zinc oxide}}{ZnO(aq.)} + H_2(g)$$

$$\underset{\text{Aluminium}}{2Al(s)} + \underset{\text{Steam}}{3H_2O(g)} \rightarrow \underset{\text{Aluminium oxide}}{Al_2O_3(aq.)} + 3H_2(g)$$

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

$$\underset{\text{Iron (Red hot)}}{3Fe(s)} + \underset{\text{Steam}}{4H_2O} \rightarrow \underset{\text{Iron (II, III) oxide}}{Fe_3O_4(aq.)} + \underset{\text{Hydrogen}}{4H_2(g)}$$

The order of reactivities of different metals with water decreases in the order :

Na > Mg > Zn > Fe > Cu

(D) Reaction of Metals with Acids :

Metals react with acids to form a salt liberating hydrogen gas.

Metal + Dilute acid → Salt + Hydrogen

(a) Reaction of Metals with dilute hydrochloric acid and dilute sulphuric acid :

(i) Potassium and sodium react violently with dilute acids to form respective salts with the liberation of hydrogen gas.

$$\underset{\text{Potassium}}{2K(s)} + \underset{\substack{\text{Hydrochloric acid}\\ \text{(dilute)}}}{2HCl(aq)} \rightarrow \underset{\text{Potassium chloride}}{2KCl(aq.)} + \underset{\text{Hydrogen}}{H_2(g)\uparrow}$$

$$\underset{\text{Sodium}}{2Na(s)} + \underset{\substack{\text{Hydrochloric acid}\\ \text{(dilute)}}}{2HCl(aq)} \rightarrow \underset{\text{Sodium chloride}}{2NaCl(aq.)} + \underset{\text{Hydrogen}}{H_2(g)}$$

$$\underset{\text{Sodium}}{2Na(s)} + \underset{\text{Dil. sulphuric acid}}{H_2SO_4(aq.)} \rightarrow \underset{\text{Sodium sulphate}}{Na_2SO_4(aq.)} + \underset{\text{Hydrogen}}{H_2(g)\uparrow}$$

(ii) Magnesium reacts with dilute acids less violently and forms magnesium salts and hydrogen.

$$\underset{\text{Magnesium}}{Mg(s)} + \underset{\substack{\text{Hydrochloric acid}\\ \text{(dilute)}}}{2HCl(aq)} \rightarrow \underset{\text{Magnesium chloride}}{MgCl_2(aq.)} + \underset{\text{Hydrogen}}{H_2(g)\uparrow}$$

$$\underset{\text{Magnesium}}{Mg(s)} + \underset{\text{dil. sulphuric acid}}{H_2SO_4(aq.)} \rightarrow \underset{\text{Magnesium sulphate}}{MgSO_4(aq.)} + \underset{\text{Hydrogen}}{H_2(g)\uparrow}$$

(iii) Aluminium reacts relatively less rapidly than magnesium with dilute acids to form salts with the evolution of hydrogen gas.

$$\underset{\text{Aluminium}}{2Al(s)} + \underset{\substack{\text{Hydrochloric acid}\\ \text{(dilute)}}}{6HCl(aq)} \rightarrow \underset{\text{Aluminium chloride}}{2AlCl_3(aq.)} + \underset{\text{Hydrogen}}{3H_2(g)}$$

$$\underset{\text{Aluminium}}{2Al(s)} + \underset{\text{dil. sulphuric acid}}{3H_2SO_4(aq.)} \rightarrow \underset{\text{Aluminium sulphate}}{Al_2(SO_4)_3} + \underset{\text{Hydrogen}}{3H_2(g)}$$

The aluminium metal possesses a thin protective layer of aluminium oxide on its surface.

Therefore, aluminium first reacts slowly with dilute hydrochloric acid. After sometime, when the oxide layer get dissolved, the reaction becomes fast.

(iv) Zinc reacts less rapidly with dilute acid in comparison to that with aluminium.

$$\underset{\text{Zinc}}{Zn(s)} + \underset{\substack{\text{Hydrochloric acid}\\ \text{(dilute)}}}{2HCl(aq)} \rightarrow \underset{\text{Zinc chloride}}{ZnCl_2(aq.)} + \underset{\text{Hydrogen}}{H_2(g)\uparrow}$$

$$\underset{\text{Zinc}}{Zn(s)} + \underset{\text{dil. sulphuric acid}}{H_2SO_4(aq.)} \rightarrow \underset{\text{Zinc sulphate}}{ZnSO_4(aq.)} + \underset{\text{Hydrogen}}{H_2(g)\uparrow}$$

(v) Iron reacts slowly with cold and dilute acids to from iron salts and hydrogen gas.

$$\underset{\text{Iron}}{Fe(s)} + \underset{\substack{\text{Hydrochloric acid}\\ \text{(dilute)}}}{2HCl(aq)} \rightarrow \underset{\text{Iron (II) chloride}}{FeCl_2(aq.)} + \underset{\text{Hydrogen}}{H_2(g)\uparrow}$$

$$\underset{\text{Iron}}{Fe(s)} + \underset{\text{dil. sulphuric acid}}{H_2SO_4(aq.)} \rightarrow \underset{\text{Iron (II) sulphate}}{FeSO_4(aq.)} + \underset{\text{Hydrogen}}{H_2(g)\uparrow}$$

The reaction of metals with dilute acids gives an idea of the relative chemical reactivity of metals. On the basis of intensity of a metal with dilute acid, the metals under consideration have decreasing reactivity towards water in the order : K > Na > Mg > Al > Zn > Fe > Cu.

(b) Reaction of Metals with dilute Nitric acid :

Metals on reacting with dilute nitric acid do not give hydrogen gas due to strong oxidizing nature of nitric acid which oxidizes the hydrogen (H_2) produced in the reaction to water (H_2O). Nitric acid itself gets reduced to an oxide of nitrogen such as nitrous oxide (N_2O), nitric oxide (NO) or nitrogen dioxide (NO_2).

$$\underset{\text{Zinc}}{4Zn(s)} + \underset{\substack{\text{Nitric acid}\\ \text{(dilute, 20\%)}}}{10HNO_3(aq.)} \rightarrow \underset{\text{Zinc nitrate}}{4Zn(NO_3)_2(aq.)} + \underset{\text{Nitrous oxide}}{N_2O(g)} + 5H_2O(\ell)$$

$$\underset{\text{Copper}}{3Cu(s)} + \underset{\substack{\text{Nitric acid}\\ \text{(cold and dilute)}}}{8HNO_3(aq.)} \rightarrow \underset{\text{Copper (II) nitrate}}{3Cu(NO_3)_2(aq.)} + \underset{\text{Nitric oxide}}{2NO(g)} + 4H_2O(\ell)$$

(E) Reaction of Metals with solutions of other metal salts :

Some metals are highly reactive, some possess moderate reactivity and some have very poor chemical reactivity. The relative chemical reactivity of metal is decided by its property of displacing other metals from their salt solutions.

A more reactive metal displaces the lesser reactive metal from its solution.

If metal A is more reactive than metal B, A will displace B from its solution.

e.g., on addition of iron pieces in (aqucous) $CuSO_4$, the blue colour of copper sulphate solution fades away showing the following reaction.

$$\underset{\substack{\text{Iron}\\\text{(more reactive)}}}{Fe(s)} + \underset{\text{Copper sulphate}}{CuSO_4(aq.)} \rightarrow \underset{\text{Iron (II) sulphate}}{FeSO_4(aq.)} + \underset{\substack{\text{Copper}\\\text{(less reactive)}}}{Cu(s)}$$

On the other hand, copper turnings on addition to aqueous ferrous sulphate solution show no change in green colour of $FeSO_{4(aq)}$ i.e. copper fails to displace iron from iron sulphate solution. Therefore iron is more reactive than copper.

Such reactions are called displacement reactions. Some other displacement reactions are :

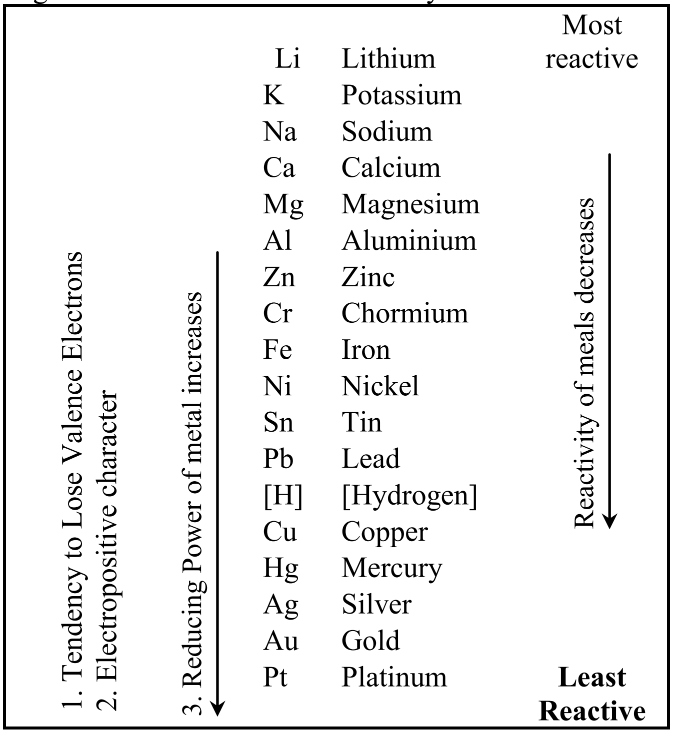
$$\underset{\text{Zinc}}{Zn(s)} + \underset{\text{Copper sulphate}}{CuSO_4(aq.)} \rightarrow \underset{\text{Zinc sulphate}}{ZnSO_4(aq.)} + \underset{\text{Copper}}{Cu(s)}$$

$$\underset{\text{Zinc}}{Zn(s)} + \underset{\text{Iron (II) sulphate}}{FeSO_4(aq.)} \rightarrow \underset{\text{Zinc sulphate}}{ZnSO_4(aq.)} + \underset{\text{Iron}}{Fe(s)}$$

$$\underset{\text{Copper}}{Cu(s)} + \underset{\text{Silver nitrate}}{2AgNO_3(aq.)} \rightarrow \underset{\text{Copper (II) nitrate}}{Cu(NO_3)_2(aq.)} + \underset{\text{Silver}}{2Ag(s)}$$

REACTIVITY OR ACTIVITY SERIES OF METALS

It is a series in which metals are arranged in the order of their decreasing chemical reactivities. On the basis of the displacement of metal from its salt solution by some other metal, all the metals are arranged in a series known as the reactivity series or the activity series. In the activity series, metals are placed in the decreasing order of their chemical reactivity.



Li	Lithium
K	Potassium
Na	Sodium
Ca	Calcium
Mg	Magnesium
Al	Aluminium
Zn	Zinc
Cr	Chormium
Fe	Iron
Ni	Nickel
Sn	Tin
Pb	Lead
[H]	[Hydrogen]
Cu	Copper
Hg	Mercury
Ag	Silver
Au	Gold
Pt	Platinum

Following information can be derived from activity series :

(i) The electropositive character of metal decreases on moving down the series.

(ii) The reactivity of metal decreases on moving down the series.

(iii) The tendency of metal to get oxidized, i.e., to lose valence electrons decreases on moving down the series. Thus, Lithium (placed at the top of series) gets most readily oxidized.

(iv) The reducing power of metal decrease on moving down the series. Thus, Lithium is the strongest reducing agent, whereas gold the weakest.

(v) A metal can displace any other metal placed below it in the series from its salt solution.

(vi) Metals placed above hydrogen in the series can displace hydrogen from dilute acids, while those placed below hydrogen cannot displace hydrogen from dilute acids.

(vii) The thermal stability of the oxide of metals decreases on moving down the series. The oxides of potassium, sodium, calcium, magnesium are very stable and do not decompose on heating. On the other hand, the oxides of metals like Hg, Ag, Au etc. are not much stable and decompose on heating.

CHEMICAL PROPERTIES OF NON-METALS

(A) Electronegative Nature :

Atoms of non-metals usually possess 4, 5, 6, 7 or 8 electrons in their valence shells. For example.

$_{6}C : 2, 4$
$_{7}N : 2, 5$
$_{8}O : 2, 6$
$_{9}F : 2, 7$
$_{10}Ne : 2, 8$
$_{17}Cl : 2, 8, 7$

Non-metals (except noble gases) usually gains electron in order to complete their octet and form negative ions or anions e.g.,

$$\underset{\substack{\text{Nitrogen}\\(2,5)}}{N} + 3e^- \rightarrow \underset{\substack{\text{Nitride ion}\\(2,8)}}{N^{3-}}$$

$$\underset{\substack{\text{Oxygen}\\(2,6)}}{O} + 2e^- \rightarrow \underset{\substack{\text{Oxide ion}\\(2,8)}}{O^{2-}}$$

$$\underset{\substack{\text{Fluorine}\\(2,7)}}{F} + e^- \rightarrow \underset{\substack{\text{Fluoride ion}\\(2,8)}}{F^-}$$

The number of electrons gained by non-metal to form stable negatively charged ion or anion is called the valency of that non-metal. Therefore, non-metals are electronegative in nature as they possess more affinity for electrons. Also during electrolysis anions are migrated towards anode where they are discharged and oxidized e.g.,

Anode $2Cl^- \rightarrow Cl_2 + 2e^-$

(B) Reaction of Non-metals with Oxygen :

Non-metals burn in oxygen (on air) to give their respective oxides.

Boost your knowledge

- The oxides of non-metals are either acidic or neutral in nature. These oxides dissolve in water to form respective oxyacids and therefore oxide of non-metals are known as anhydrides of that acid.

Some examples are given below.

(i) Carbon burns in air to give carbon dioxide (an acidic oxide) which on dissolution in water gives an acidic solution.

$$\underset{\substack{\text{Carbon}\\(\text{a non}-\text{metal})}}{C(s)} + O_2\,(g) \rightarrow \underset{\substack{\text{Carbon dioxide}\\(\text{an acidic oxide})}}{CO_2(g)}$$

$$SO_2(g) + H_2O\,(\ell) \rightarrow H_2SO_3(aq.)$$

SO_2 is called acid anhydride of sulphurous acid.

(ii) Carbon also forms a neutral oxide CO in limited supply of oxygen.

$$\underset{\text{Carbon}}{C(s)} + \frac{1}{2}O_2(g) \rightarrow \underset{\substack{\text{Carbon monoxide}\\(\text{a neutral oxide})}}{2CO(g)}$$

Boost your knowledge

- Although CO is neutral oxide but it behaves as acid when passed over heated NaOH(s) to give HCOONa.

$$\underset{(\text{Base})}{NaOH} + \underset{\text{acid}}{CO} \xrightarrow[\text{high T}]{\text{high P}} \underset{\text{salt}}{HCOONa}$$

Nitric oxide (NO) and nitrous oxide (N_2O) are also neutral in nature. N_2O_3, N_2O_4 and N_2O_5 are acidic oxides of nitrogen. N_2O_3 is acid anhydride of nitrous acid (HNO_2) whereas N_2O_5 is acid anhydride of nitric acid (HNO_3), N_2O_4 is mixed acid anhydride of two acids (HNO_2 and HNO_3).

(C) Reaction of Non-Metals with Water :

Non-metals do not react with water or steam.

(D) Reaction of Non-metals with dilute Acids :

Non-metals do not react with dilute acids, e.g., dilute HCl, dilute H_2SO_4 etc.

(E) Reaction of Non-Metals with other Non-Metal Salt Solutions :

Reaction of non-metals with salt solutions are significant only in case of halogens (F_2, Cl_2, Br_2 and I_2). A more reactive halogen is able to displace a less reactive halogen from its salt solution.

$$\underset{\text{Chlorine}}{Cl_2(g)} + \underset{\text{Sodium bromide}}{2NaBr(aq.)} \rightarrow \underset{\text{Sodium chloride}}{2NaCl(aq.)} + \underset{\text{Bromine}}{Br_2(\ell)}$$

$$\underset{\text{Bromine}}{Br_2(\ell)} + \underset{\text{Sodium iodide}}{2NaI(aq.)} \rightarrow \underset{\text{Sodium bromide}}{2NaBr(aq.)} + \underset{\text{Iodine}}{I_2(s)}$$

The reactivity order of halogens is :

$F_2 > Cl_2 > Br_2 > I_2$.

(F) Reaction of Non Metal with other Non-Metal :

Some of the reactions of this type are given below.

(i) Hydrogen reacts with chlorine to give hydrogen chloride

$$H_2(g) + Cl_2(g) \xrightarrow{\text{light}} 2HCl(g)$$

(ii) Hydrogen reacts with nitrogen to give ammonia

$$3H_2(g) + N_2(g) \xrightleftharpoons[\text{200 atm, 500 ºC}]{\text{Fe}} 2NH_3(g)$$

Comparison of Chemical properties of Metals and Non-metals :

Property	Metals	Non-Metals
1. Nature of ions	Metals are electropositive elements and hence, lose one or more electrons to form positive ions.	Non-metals are electronegative elements and hence, gain one or more electrons to form negative ions.
2. Nature of oxides	Metals form basic oxides	Non-metals form either acidic or neutral oxides.
3. Reaction with water	Most of the metals displace hydrogen from water or steam.	Non-metals do not react with water.
4. Reaction with dilute acids	Metals which lie above hydrogen in the activity series displace hydrogen from dilute acids.	Non-metals do not react with dilute acids and hence, do not displace hydrogen from dilute acids.
5. Nature of hydrides	Highly electropositive elements (i.e., K, Na, Ca, etc.,) react with hydrogen to form ionic hydrides which are generally unstable.	Non-metals form covalent hydrides which are quite stable.
6. Nature of chlorides	Metals generally combine with chlorine to form solid ionic chlorides which conduct electricity in the aqueous solution or in the molten state.	Non-metals combine with chlorine to form covalent hydrides. These are either gases or liquids. Some non-metal chlorides do not contain ions, therefore, they do not conduct electricity.
7. Oxidizing and reducing behaviour	Metals have a strong tendency to lose electrons and hence they behave as reducing agents.	Non-metals have a strong tendency to accept electrons and hence they behave as oxidizing agents.

ELEMENTARY IDEA OF BONDING

Chemical Bond : The force which links the atoms (ions) in a molecule is called a chemical bond.

Inertness of Noble Gases : Noble gases have stable electronic configuration, i.e., their outermost (valence) shell is complete. They all have 8 electrons in the outermost shell with the exception of helium (which has 2 electrons in its outermost shell).

They do not lose, gain or share electrons and are inert or unreactive. Since noble gases are chemically unreactive, it is clear that the electronic arrangement in their atoms are very stable and do not allow the outermost electrons to take part in chemical reactions.

Electronic Configuration of Noble Gases (or Inert Gases)

Noble gas (Inert gas)	Symbol	Atomic Number	Electronic Configuration K L M N O P	Number of Electrons In Outermost Shell (Valence Shell)
1. Helium	He	2	2	2
2. Neon	Ne	10	2, 8	8
3. Argon	Ar	18	2, 8, 8	8
4. Krypton	Kr	36	2, 8, 18, 8	8
5. Xenon	Xe	54	2, 8, 18, 18, 8	8
6. Radon	Rn	86	2, 8, 18, 32, 18, 8	8

Since noble gases having completely filled outermost shells or valence shells are chemically unreactive, we can explain the reactivity of elements as a tendency of their atoms to achieve a completely filled outermost shell or valence shell (just like those of noble gases) and become stable.

Causes of Chemical bonding : Atoms of elements other than noble gases have unstable electronic configuration i.e., their outermost shell is incomplete. They can lose, gain or share electrons and are chemically reactive. The driving force for atoms to combine is related to the tendency of each atom to attain stable electronic configuration of the nearest noble gas. For an atom to achieve stable electronic configuration it must have :

(i) two electrons in the outermost shell (nearest noble gas He) – Duplet rule.
(ii) eight electrons in the outermost shell (all noble gases other than He) – Octet Rule.

TYPES OF BONDS

Depending upon the mode of attainment of stable noble gas configuration, following types of chemical bonds are noticed.

1. **Electrovalent or Ionic bond :** Bonds formed by the transfer of one or more electrons from one atom to another are called ionic bonds. It has the strongest bonding among all bonds and occurs between electropositive element and electronegative element e.g., KCl.
2. **Covalent bond :** Bonds formed by the mutual sharing of one, two, or three pairs of electrons between two electronegative elements e.g., NH_3.
3. **Coordinate or Dative bond :** Bonds formed by the one sided sharing of electron pair i.e., the shared electron pair is contributed by only one of the combining atoms e.g., NH_4^+.

ELECTROVALENT OR IONIC BOND

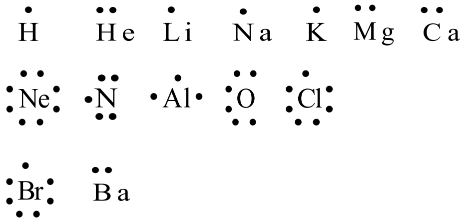
Definition and Formation of Electrovalent Bond :

1. The union of two or more atoms involving complete transfer of one or more electrons from the valence shell of an atom into the valence shell of other atoms so that each atom involved in bonding acquires nearest noble gas configuration to gain stability.
2. Ionic bonding usually involves a metal and a non-metal. The metal atom (the more electropositive atom) loses one or more electron to acquire octet and forms cation. The non-metal atom (the more electronegative atom) gains one or more electrons to acquire octet and forms anion.
3. The two ions in solid state are held together by strong attractive forces called coulombic forces or electrostatic forces.
4. This electrostatic force of attraction gives rise to the formation of the electrovalent bond or ionic bond.
5. The compounds containing electrovalent bonds are called electrovalent compounds or ionic compounds.

   The number of electrons donated or accepted by the valence shell of an atom of an element to achieve stable electronic configuration is called electrovalency of that element.

Boost your knowledge

- G.N. Lewis, an American chemist, introduced a simple notation to denote the valence electrons in an atom. These notations are called Lewis symbols or electron dot symbols. In these symbols, only the valence electrons are shown as dots surrounding the symbol of the given element. For example, Lewis symbols :

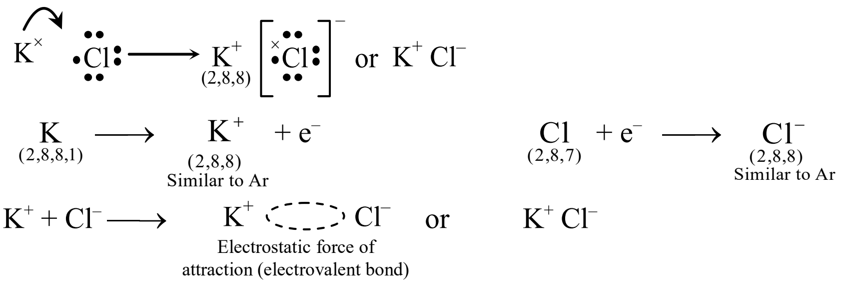


(i) Formation of potassium chloride (KCl) : Potassium is an electropositive element, and chlorine is an electronegative element as shown by their electronic configuration.

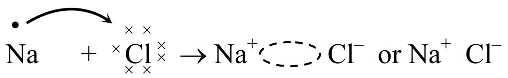
$_{19}K : 2, 8, 8, 1$

$_{17}Cl : 2, 8, 7$

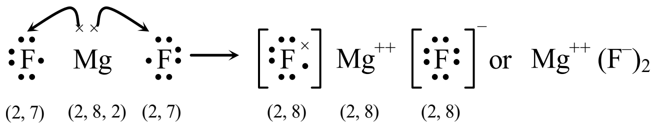
Potassium acquires the stable configuration of nearest noble gas argon (Ar : 2, 8, 8) by losing its valence electron. Chlorine acquires the stable electronic configuration of its nearest noble gas argon (Ar : 2, 8, 8) by gaining one electron. Therefore the potassium atom transfers its valence electrons to the chlorine atom during the combination of potassium and chlorine atoms. Potassium atom forming potassium ion (K^+) while the chlorine atom forming chloride ion (Cl^-) which are hold together by strong electrostatic force of attraction. The formation of ionic bond between K and Cl can be shown as follows :



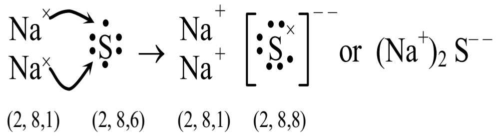
(ii) Formation of NaCl :



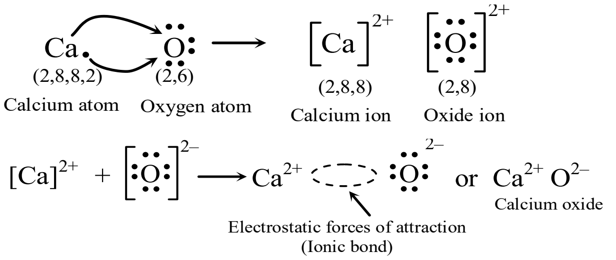
(iii) Formation of MgF_2 :



(iv) Formation of Na_2S :



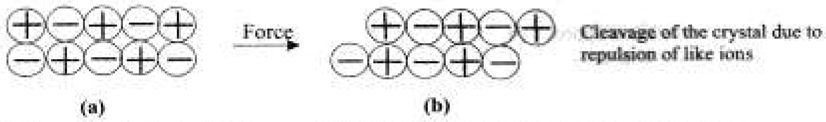
(v) Formation of CaO :



The arrows indicate the transfer of an electron from more electropositive atom to more electronegative atom. This type of representation of an ionic compound is called Lewis electron dot structure.

Characteristic properties of electrovalent or Ionic compounds :

1. **Physical state and stability :** Ionic compounds are usually solid, brittle, hard and crystalline in nature at room temperature due to strong electrostatic forces existing between the oppositely charged ions holding them closer in the specified positions in the crystal lattice. The constituent units in a crystal of ionic solid are ions and not molecules. Crystal structure of ionic solid has a large number of cations and anions held together by strong electrostatic forces in definite geometry pattern. For example crystal of NaCl possesses simple cubic lattice in which each Na^+ ion is surrounded by six Cl^- ions and each Cl^- ion is surrounded by six Na^+ ions.
   There is nothing exist like NaCl in sodium chloride or in other words ionic compounds do not have any molecule and for the sake of representation we write NaCl. It may also be therefore suggested that sodium chloride has no molecular mass, however 58.5 is its formula mass.
2. **High melting and boiling points :** The oppositely charged ions are tightly held together in crystal lattice by strong electrostatic forces and require a very high temperature to overcome these forces. Only at high temperatures, ions are set free to move. Therefore, the melting points and boiling points of ionic compounds are generally very high.
3. **Hard and brittle nature :** The ionic solids are hard and brittle. An ionic solid consists of closely packed positive and negative ions in space lattice. On shear, the layers shift and similar charged ions come in front of each other as shown below. Thus, the two layers repel each other and the crystal gets cleaved.



(a) The layers in an ionic crystal (b) The slipping of one layer on applying force showing the cleavage of the crystal.

4. **Solubility in water :** Most of the ionic compounds are water soluble due to high dielectric constant of water which weakens the electrostatic force of attraction between the ions. This leads to separation of ions and pass them in solution state.

Boost your knowledge

- Many of the ionic solid e.g. AgCl, AgBr, PbI_2, $BaSO_4$, CaF_2 etc. are insoluble. In such cases hydration energy being lesser than ionization energy.

5. **Electrical conductivity :** Although ionic compounds have ions yet they are unable to conduct electricity in the solid state. This is because in the solid state, the ions are held together in specified positions and are unable to move. However, on fusion or on dissolution in water ions are set free to move and conduct current. Thus, ionic compounds are poor conductors of electricity in the solid state but are good conductors either in the fused state or in the dissolved state.
6. **Ionic reactions :** The reaction of ionic compound (ionic reaction) are very fast or instantaneous as they involve the mutual interaction of ions.

OCCURRENCE OF METALS

Elements occur in nature either in the free state or in the combined state.

1. **Free state or native state :** An element in uncombined state i.e., the elemental form in which it exist in nature is called free state or native state. The less reactive elements having less affinity for oxygen, moisture and other chemical reagents (e.g., metals like gold, silver, platinum etc; and non-metals like nitrogen, oxygen, carbon etc., ordinarily inert elements like helium, neon argon, etc.) occur in free state in nature.
2. **Combined state :** The elements having more affinity towards oxygen, moisture and other chemical reagents occur in the combined form. All the reactive elements occur in combined state (e.g., metals like sodium, potassium, calcium etc; non-metals like fluorine, chlorine, bromine etc.) always occur in the combined state in nature.

Boost your knowledge

- The elements having moderate or average reactivity occur both in the free state as well as in the combined state in nature e.g. copper, iron, sulphur, etc.

(a) Minerals and Ores :

Most of the metals found in combined state in the earth's crust are associated with sandy matter and other impurities and obtained by mining are called minerals. A mineral is either a single compound or a complex mixture of several compounds of the same or different metals with sandy matter and other impurities.

A mineral having very low percentage of the metal is not much profitable because of the high cost of the recovery of metal. The minerals from which a metal can be extracted profitably are called ores. Therefore all the ores are minerals but all minerals are not ores. e.g., both bauxite ($Al_2O_3.2H_2O$) and clay ($Al_2O_3.2SiO_2.2H_2O$) are minerals of aluminum. However, extraction of aluminium is usually made from bauxite ore and not from clay. Thus bauxite is an ore but clay is just a mineral of Al.

(b) Classification of Ores :

(i) Native ores : These contain less reactive metals e.g., Ag, Au, Pt etc., in their elemental form associated with impurities like clay, sand etc.

(ii) Sulphide Ores : Metals are present in the form of sulphides in these ores e.g., Copper pyrites ($CuFeS_2$), chalcocite (Cu_2S), Zinc blende (ZnS), Galena (PbS), Argentite (Ag_2S), Cinabar (HgS) etc.

(iii)Oxidised Ores : Metals are present in the form of oxides, carbonates, sulphates, silicates, phosphates or in hydrated forms of these compounds in these ores. e.g.

(a) Oxide Ores : Bauxite ($Al_2O_3 . 2H_2O$), Corundum (Al_2O_3), Zincite (ZnO), Pyrolusite (MnO_2), Haematite (Fe_2O_3), Chromite ($FeO . Cr_2O_3$) etc.

(b) Carbonate Ores : Magnesite ($MgCO_3$), Calamine ($ZnCO_3$), Malachite [$Cu(OH)_2.CuCO_3$], Dolomite ($CaCO_3 . MgCO_3$) etc.

(c) Sulphate Ores : Barytes ($BaSO_4$), Epsomite ($MgSO_4$), Gypsum ($CaSO_4 . 2H_2O$) etc.

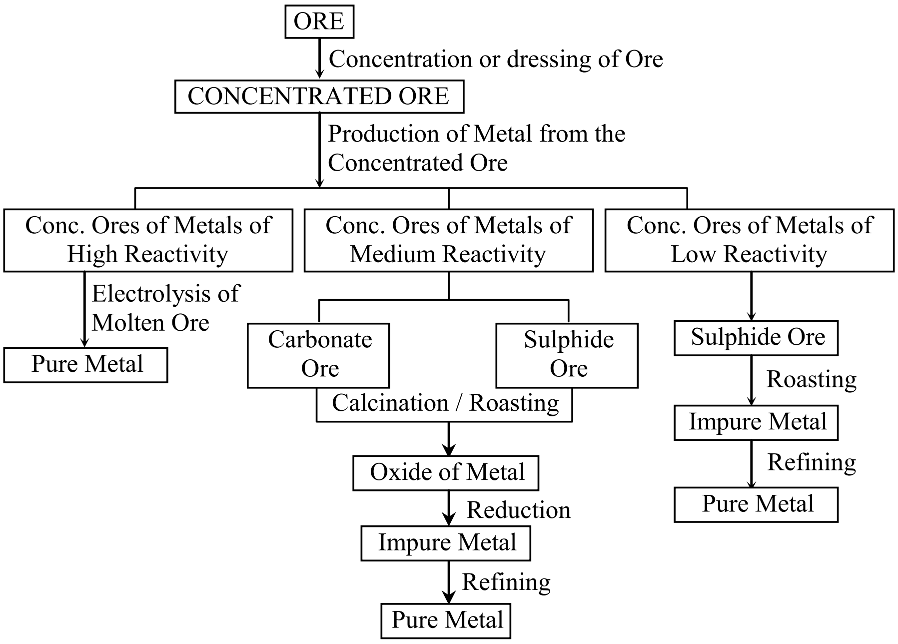
(iv) Halide Ores : Rock salt (NaCl), Carnalite ($KCl.MgCl_2.6H_2O$), Fluorspar (CaF_2), Cryolite (Na_3AlF_6), Horn silver (AgCl) etc.

Common Ores and Their Occurrence in India

Metal	Common Ores	Place of Occurrence in India
Iron	(i) Haematite, Fe_2O_3	Bihar, Karnataka
	(ii) Magnetite, $Fe_2O_3.FeO$	–
	(iii) Iron pyrites, FeS_2	–
Aluminium	(i) Bauxite, $Al_2O_3 . 2H_2O$	Madhya Pradesh, Maharaastra.
	(ii) Corundum, (Al_2O_3)	Orissa, Tamilnadu
Calcium	(i) Gypsum, $CuSO_4 . 2H_2O$	Rajasthan, Tamil Nadu, Jammu and Kashmir, Gujarat
	(ii) Limestone (chalk or marble), $CaCO_3$	Rajasthan, Madhya Pradesh
	(iii) Fluorspar, CaF_2	Madhya Pradesh
Zinc	Zinc Blende, Zns	Rajasthan
Copper	(i) Copper pyrites, $CuFeS_2$	Bihar, Orissa, Madhya Pradesh
	(ii) Malachite, $CuCO_3 . Cu(OH)_2$	Bihar, Madhya Pradesh
Lead	Galena, PbS	Rajasthan, Bihar (not much abundant in India)

METALLURGY

The process of obtaining a metal from its ores and then refining it for use is knows as metallurgy and involves several physical and chemical methods. A general flow chart for extraction of metal is given below.



(A) Crushing and Grinding of Ore : Larger pieces of ores obtained from earth crust are broken up into smaller pieces in the crushers. This process is known as crushing of ore. The crushed ore is further grinded to convert it into a fine powder.

(B) Enrichment or concentration of Ores : Presence of earth matter, rock matter, sand, limestone, mica, etc., in ores is known as gangue or matrix. Removal of these impurities from ores is called dressing or concentration.

(C) Conversion of the Concentrated Ore into its oxidised Form

The conversion of the concentrated ore into its oxidized form either removes the impurities as their volatile oxides or facilitates their removal in the later stages. Moreover, if the metal is present as carbonate or sulphide in the ore, its gets converted into the oxide. This helps in the production of metal in the later stages involving the following operations.

(a) Calcination : Calcination involves simple decomposition of ore on heating below its melting point usually in absence of air to produce new compounds having higher percentage of metal as well as removing the moisture, organic matter and volatile impurities, e.g., CO_2, etc. Calcination makes the ore porous and is made in reverberatory furnace.

Limestone : $CaCO_3 \xrightarrow{\Delta} CaO + CO_2$

Bauxite : $Al_2O_3.2H_2O \xrightarrow{\Delta} Al_2O_3 + 2H_2O$

Magnesite : $MgCO_3 \xrightarrow{\Delta} MgO + CO_2$

Haematite : $Fe_2O_3.3H_2O \xrightarrow{\Delta} Fe_2O_3 + 3H_2O$

(b) Roasting : Roasting involves action of heat in limited supply of air on ore below its melting point to produce other chemical changes along with decomposition.
Roasting is usually made in reverberatory furnace or blast furnace.

$2ZnS + 3O_2 \xrightarrow{\Delta} 2ZnO + 2SO_2$

Boost your knowledge

☛ In roasting definite chemical changes like oxidation, chlorination etc., take place while in calcinations, normally expulsion of organic matter and decomposition of ore takes place however both the process are fruitful in getting oxide form of metal.

In chloride roasting, the ore is changed into metal chloride by heating with common salts in presence of air.

$Ag_2S + 2NaCl \rightarrow 2AgCl + Na_2S$

(D) Reduction of Oxidised Ore into free metal

(a) Reduction of oxides of less reactive metals-air reduction or self reduction :

Metal oxides with less active metals Hg, Cu and Pb are easily reduced with air, The ores are roasted in air to separate metals. The process is named as auto reduction.

$$\begin{cases} \underset{\text{Cinnabar}}{2HgS} + 3O_2 \rightarrow 2HgO + 2SO_2; \\ 2HgO + HgS \rightarrow 3Hg + SO_2 \end{cases}$$

$$\begin{cases} \underset{\text{Galena}}{2PbS} + O_2 \rightarrow 2PbO + 2SO_2; \\ 2PbO + PbS \rightarrow 3Pb + SO_2 \end{cases}$$

$$\begin{cases} \underset{\text{Copper glance}}{2Cu_2S} + 3O_2 \rightarrow 2CuO + 2SO_2; \\ Cu_2S + 2Cu_2O \rightarrow 6Cu + SO_2 \end{cases}$$

(b) Reduction of oxides of moderately reactive metals-reduction with powerful reducing agents :

(i) Reduction by aluminium (Al) :

Extraction of less electropositive metals say Cr, Mn, Cu, Ni etc., can be made by heating their oxides with strong reducing agents, e.g., CO, CO + H_2, Na, Al, Mg. If Al is used as reducing agent, the process is known as Gold-Schmidt aluminothermic process and the mixture containing ore and aluminium (in the ratio 1 : 3) is known as thermite mixture.

$$\underbrace{Cr_2O_3 + 2Al}_{\text{Thermite mixture}} \rightarrow Al_2O_3 + 2Cr$$

The Gold-Schmidt aluminothermic process is commonly used for those metals which have very high melting point and are to be extracted from their oxides and their reduction with carbon is not satisfactory. Large amount of heat is released during reduction, which fuses both metal and alumina.

(ii) Carbon reduction process :

The oxides of less electropositive metals as well as the elements which do not form carbides (like Pb, Fe, Zn, Sb and Cu) are reduced by strongly heating with coal or coke.

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

$ZnO + C \rightarrow Zn + CO$

In addition to carbon, carbonmonoxide also acts as reducing agent

$6PbO + 3CO \rightarrow 6Pb + 3CO_2$

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

(c) Reduction of oxides of highly reactive metals -electrolytic reduction :

The highly electropositive metals such as alkali metals (Li, Na, K etc.), alkaline earth metals (Mg, Ca etc.) and aluminium are obtained by electrolytic reduction. During electrolytic reduction the oxides, hydroxides or chlorides of these metals are taken in the fused state and are subjected to electrolysis using inert electrodes. For example, extraction of aluminium is done by the electrolysis of fused alumina (Al_2O_3) mixed with cryolite (Na_3AlF_6). The addition of cryolite lowers the melting point of alumina.

$$2Al_2O_3 \xrightarrow[\text{Electric current}]{\text{(Cryolite)}} \underset{\text{(At cathode)}}{4Al} + \underset{\text{(At anode)}}{3O_2}$$

(E) Refining or purification of metal :

The metals extracted by either of the above operations is known as crude form.

To remove these impurities, the crude metal is subjected to the process of purification or refining.

Electrolytic refining :

The impure metal is made anode of an electrolytic cell having pure metal plate of cathode. The solution consists of same metal salt. On electrolysis pure metal from crude form dissolves at anode whereas at cathode pure metal is deposited. The soluble impurities pass into solution while the insoluble ones collect below the anode as anodic mud or anode sludge.

Highly electropositive elements are obtained by the electrolysis of their oxides, hydroxides or chlorides in fused state. Due to their highly reactive nature, neither aqueous solution can be electrolysed nor other reduction process can be applied, e.g., sodium is obtained by the electrolysis of fused sodium chloride. A small amount of some other salt say $MgCl_2$ is added to ore, in order to lower its fusion point and enhance conductance.

$$NaCl \overset{\text{Fused}}{\rightleftharpoons} Na^+ + Cl^-$$

Anode : $Cl^- \rightarrow Cl_2 + e$

Cathode : $Na^+ + e \rightarrow Na$

Similarly, aluminium is obtained by the electrolysis of alumina mixed with cryolite which acts as electrolyte and lowers the melting point of alumina to enhance its fusion and thereby electrolysis.

$$2Al_2O_3 \rightarrow \underset{\text{(At cathode)}}{4Al} + \underset{\text{(At anode)}}{3O_2}$$

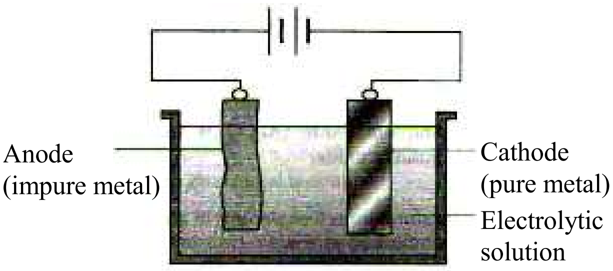
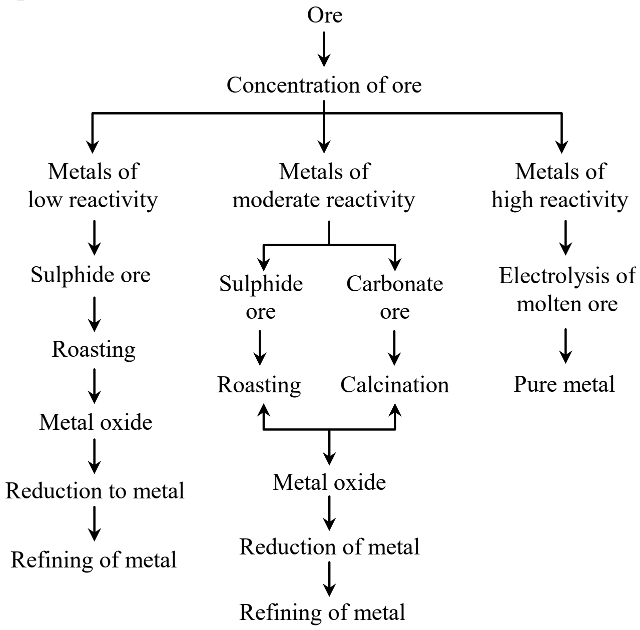


Fig. : Electrolytic refining

Flow chart indicating steps in metallurgy :

A flow chart indicating the steps involved in the extraction of metals from ores :



CORROSION

Most of the metals get degraded when kept exposed to moist air.

The process of slowly eating up of metals due to this conversion into oxides, carbonates, sulphide, etc., by the action of atmospheric gases and moisture is called corrosion.

For example,

(i) Corrosion of Iron (Called Rusting) :

Iron when exposed to moist air for a long time acquires a coating of a brown flaky substance called rust. Rust in mainly hydrated ferric oxide, $Fe_2O_3 . xH_2O$.

$$\underset{\text{Iron}}{2Fe} + \frac{3}{2}O_2 + xH_2O \rightarrow \underset{\substack{\text{Hydrated ferric oxide} \\ \text{(Rust)}}}{Fe_2O_3 . xH_2O}$$

The corrosion of iron is called rusting.

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- Gold metal is a yellow, shining metal and does not corrode when exposed to air because it is a highly unreactive metal which remains unaffected by air, moisture and other gases in the atmosphere. Since gold does not corrode and retains its luster for years. It is used to make jewellery, coins, etc.
  Platinum is also a metal which is highly resistant to corrosion. Platinum is a white metal with a silver shine. It is used to make jewellery because of its bright shiny surface and resistance to corrosion.
  Titanium is another metal resistant to corrosion because it forms a thin impermeable film of titanium oxide (TiO_2) on the surface which protects it from further damage.

Conditions Necessary for Rusting :

The two conditions which are necessary for rusting are :

1. Presence of air (or oxygen) and
2. Presence of water vapour (moisture).

Iron rusts when placed in moist air (damp air) or when placed in water. This is because moist air supplies both the things, air and water required for the rusting. Again ordinary water has some dissolved air in it. So it also supplies both the things, air and water needed for rusting. Let us perform an experiment to prove this.

Activity : To prove that air and moisture are essential for rusting :

- Take three test tubes as A, B and C and place clean iron nails in each of them.
- In test tube A, pour some tap water and cork it.
- In test tube B, pour boiled distilled water, add about 1 mL of oil and cork it. The oil will float on water and prevent the air from dissolving in the water.
- In test tube C, put some anhydrous calcium chloride and cork it. Anhydrous calcium chloride will absorb the moisture, if any from the air.
- Leave these test tubes for a few days and then observe.

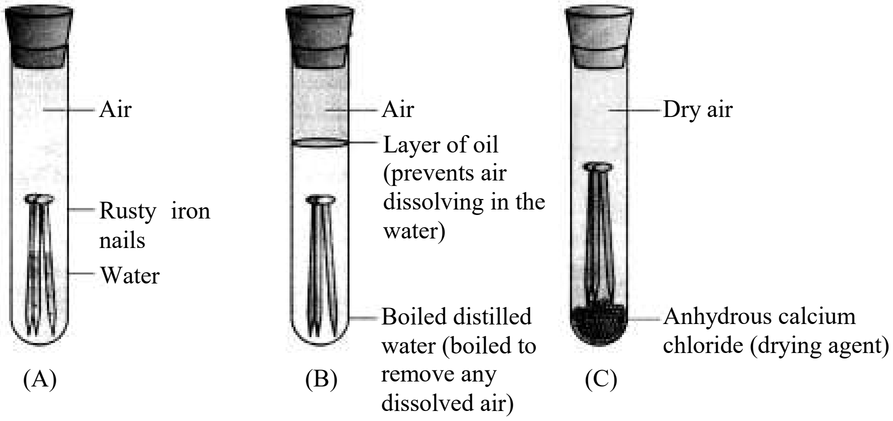

Fig. : Investigating the conditions under which Iron rusts. In tube A, both air and water are present. In tube B, there is no air dissolved in the water. In tube C, the air is dry.

Observation :

It is observed that iron nails rust in test tube A, but they do not rust in test tubes B and C. In test tube A, the nails are exposed to both air and water. Therefore, the nails becomes rusted. In test tube B, the nails are exposed to only water and the nails in the test tube C are exposed to dry air.

Conclusion :

This experiment shows that both air (oxygen) and moisture are essential for rusting to take place.

Prevention of Rusting :

We have seen that both air and water are needed for rusting. In order to check it, precautions must be taken so that they may not come in contact with the surface of the metal. The various method used for preventing the rusting are given below :

1. **Painting :** In materials like railings, window, grills, lamp posts, iron gates, iron bridge, railway coaches, bodies of cars, buses and trucks, etc., painting the metal surface does not allow them to come in contact with the moist air and thus, prevents rusting.
2. **Greasing and oiling :** When some grease or oil is applied on the surface of an iron object, then moisture and air cannot come in contact with it and hence, rusting is prevented. For example, the tools and machine parts made up of iron are covered with a film of grease or oil to prevent them from rusting.
3. **Galvanization :** Articles that are exposed to extreme moisture such as roof sheets and pipes are protected from rusting by galvanizing. Galvanization is a method of protecting steel and iron from rusting by coating them with a thin layer of zinc. The iron coated with zinc is called galvanized iron or simply GI.

   As zinc is more reactive than iron, it reacts with oxygen in the presence of moisture to form an invisible layer of basic zinc carbonate $ZnCO_3.Zn(OH)_2$ which protects it from further rusting. It is interesting to note that galvanometer iron articles remain protected from rusting even if zinc coating is broken. This is because zinc is more reactive than iron.
4. **Electroplating :** Electroplating is another technique used to prevent articles from rusting. It not only prevents from rusting but also beautifies the articles. Bathroom fittings and vehicle parts like bicycle handle bars, car bumpers, etc., are some of the articles that are chromium plated.

Fig. : Chromium plated bathroom fittings.

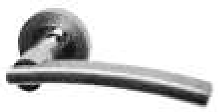

Fig. : Nickel plated door handles.

5. **Alloying :** Iron can be prevented from rusting by alloying. Also stainless steel (iron alloyed with chromium and nickel) does not rust and is used for making cutlery, cooking utensils, surgical instruments. etc.

(ii) Corrosion of Copper :

Copper reacts with moist carbon dioxide in the air and slowly loses its shiny brown surface and acquires a green coating due to the formation of basic copper carbonate, $CuCO_3.Cu(OH)_2$.

$$\underset{\text{Copper}}{2Cu} + \underbrace{CO_2(g) + O_2(g) + H_2O(\ell)}_{\text{From moist air}} \rightarrow \underset{\substack{\text{Basic copper carbonate}\\ \text{(Green)}}}{CuCO_3.Cu(OH)_2}$$

(iii) Corrosion of Silver :

Silver articles become black after sometime when exposed to air. This is due to formation of a coating of black silver sulphide (Ag_2S) on its surface by the action of H_2S gas present in the air.

$$\underset{\text{Silver}}{2Ag(s)} + \underset{\substack{\text{Hydrogen}\\ \text{sulphide}\\ \text{(From air)}}}{H_2S(g)} \rightarrow \underset{\substack{\text{Silver sulphide}\\ \text{(Black)}}}{Ag_2s_{(s)}} + H_2(g)$$

(iv) Corrosion of Aluminium :

Due to the formation of a dull layer of aluminium oxide when exposed to moist air, the aluminium metal loses its shine very soon after use. This aluminium oxide layer is very tough and prevents the metal underneath from further corrosion (because moist air is not able to pass through this aluminium oxide layer). This means sometimes corrosion is useful.

ALLOYS

Alloying is a very good method of improving the properties of a metal. We can get the desired properties by this method. For example, iron is the most widely used metal. But it is never used in the pure form. This is because pure iron is very soft and stretches easily when hot. But, if it is mixed with a small amount of carbon (about 0.05 %) it becomes hard and strong. When iron is mixed with nickel and chromium, we get stainless steel which is hard and does not rust. In fact, the properties of any metal can be changed if it is mixed with some other substance. The substance added may be a metal or a non-metal.

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

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☛ If one of the metals is mercury, then the alloy is known as an amalgam.

The properties of an alloy are different from the properties of the constituent metals from which it is made. The electrical conductivity and melting point of an alloy is less than that of pure metals. Alloys are stronger, harder and more resistant to corrosion.

Objectives of Alloy Making :

Alloys are prepared to develop certain specific properties not possessed by constituent elements. The main objectives of alloy making are :

1. **To increase hardness :** Alloys are generally more harder than pure metals. **For example,** stainless steel and chrome steel are harder than pure iron.
2. **To increase tensile strength :** Alloy steel have greater tensile strength as compared to iron. For example, chrome steel used in making axle, ball bearings and cutting tools such as files, etc, have greater tensile strength than iron.
3. **To increases resistance to corrosion :** Iron gets rusted easily but stainless steel resists corrosion.
4. **To lower the melting point :** Normally alloys have lower melting point than pure metals. **For example,** solder which is an alloy of lead and tin has lower melting point than both the metals.
5. **To modify chemical reactivity :** Alloys are formed to modify the chemical reactivity of metal. For example, sodium metal catches fire immediately when added to water. However, sodium amalgam which is an alloy of sodium with mercury does not catch fire.

Preparation of alloys : Alloys are prepared by melting the main constituent metal and then adding definite proportions of the other elements (metals or non-metals) also in the molten state. The molten mass is uniformly stirred and cooled to form a solid mass. This is known as alloy.

Type of alloys :

Alloys have been divided into the following three types :

1. **Ferrous alloys :** An alloy in which iron is present as one of the constituents is called a ferrous alloy. For example, manganese steel (Fe = 86 %, Mn = 13 % and C = 1 %) and nickel steel (Fe = 96-98 % and NI = 4-2 %) etc.
2. **Non-ferrous alloys :** An alloy which does not contain iron as one of the constituents is called non-ferrous alloy. For example, brass (Cu = 80 %, Zn = 20 %) and bronze (Cu = 90 %, Sn = 10 %).
3. **Amalgams :** An alloy containing mercury as one of the constituent metals is known as amalgam. For example, sodium amalgam (Na-Hg), zinc amalgam (Zn-Hg), etc.

Some Common Alloys :

Alloy	Composition	Properties and Uses
Brass	Cu (60–80%) Zn (20–40%)	Brass is used for decoration purposes, for making many scientific instruments, telescopes, microscopes, barometers etc.
Bronze	Cu (75-90%) Sn (10-25%)	For making statues, cooking utensils and coins.
German silver	Cu (30–60%) Zn (25–35%) Ni (15–35%)	It is silvery white as silver, malleable and ductile. It is used as imitation silver, in making ornaments and utensils and also for decoration.
Gun metal	Cu (88%) Sn (10%) Zn (2%)	It is used for making gears and bearings, and gun barrels.
Bell metal	Cu (80%) Sn (20%)	It is used for casting bells.
Stainless steel	Fe (74%) Cr (18%) Ni (8%)	Stainless steel is hard, tenacious and corrosion resistant. For making cutlery, utensils, ornamental pieces and other instrument and apparatus.
Nickel steel	Fe (96–98%) Ni (2–4%)	Nickel steel is hard, elastic and corrosion resistant. For making electric wire cables, automobile and aeroplane parts, watches, armour plates, propeller shafts, etc.
Alnico	Fe (60%) Al (12%) Ni (20-%) Co (8%)	Highly magnetic. For making permanent magnets.
Duralium or Duralumin	Al (95%) Cu (4%) Mg (0.5%) Mn (0.5%)	For making aeroplane, spacecrafts, ships and pressure cookers. In strength, it is as good as steel but it is very light. It is hard, corrosion-resistant and highly ductile.
Magnalium	Al (90–95%) Mg (5–10%)	For making light instruments and balance beams. It is hard and tough.
Solder	Pb (50%) Sn (50%)	It has a low-melting point. Used for soldering purposes.
Type metal	Pb (75–80%) Sb (15–20%) Sn (5%)	For making printing type.