THE RADIANT EDUCATION POINT
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METALS AND NON-METALS

Elements are generally classified into metals and non-metals based on their properties.

Physical properties of metals and non-metals:

Malleability: Metals can be converted into thin sheets. Metals are malleable while non-metals are not.

Ductility: Metals can be converted into long wires. Metals are ductile while non-metals are not.

Conduction of heat: Metals are good conductors of heat. ***Silver and copper*** are best conductors while ***lead and mercury*** are the poor conductors of heat. Non-metals are bad conductors of heat.

Conduction of electricity: Metals are good conductors of electricity while non- metals are bad conductors of electricity with an ***(exception of graphite which is a non-metal but is a good conductor of electricity).***

Appearance of surface: Metals, in their pure state, have a shining surface also called metallic lustre. Non-metals are generally non-lustrous with an ***(exception of iodine which is a non-metal but is lustrous)***. Metals on reacting with gases in atmosphere lose its shiny appearance when kept in air for a long time.

Hardness: Metals are generally hard with the ***(exceptions of sodium and potassium which are soft and can be cut by knife).***

Density: Metals have high density (except sodium and potassium which have low density.

Melting and boiling points: Generally, metals have high melting and boiling points ***(except for sodium, potassium, mercury, caesium, and gallium which are metals but have low melting and boiling points)***

State at room temperature: Metals are generally solid at room temperature ***(exception of mercury which is liquid).*** Non-metals are present in solids and gaseous state at room temperature ***(exception of bromine which is liquid).***

Sonority: Metals are sonorous while non-metals are not.

Comparison of physical and chemical properties of metals and non – metals:-

S.No	Property	Metals	Non-Metals
1	Physical State	Metals are solid at room temperature. Except mercury and gallium.	Non-metals generally exist as solids and gases, except Bromine.
2	Melting and boiling points	Metals generally have high m.pt and b.pt except gallium and cesium.	Non-metals have low m.pt and b.pt except diamond and graphite.
3	Density	Generally high.	Generally low.
4	Malleability and Ductility	Malleable and ductile.	Neither malleable nor ductile.
5	Electrical and thermal conductivity	Good conductors of heat and electricity.	Generally poor conductors of heat and electricity except graphite.
6	Luster	Poses shining luster.	Do not have luster except iodine.
7	Sonorous sound	Give sonorous sound when struck.	Does not give sonorous sound.
8	Hardness	Generally hard except Na, K	Solid non-metals are generally soft except diamond.

Exceptions in Physical Properties

1. **Graphite** a *non-metal* but is a good conductor of electricity.
2. **Iodine** is a *non-metal* but is a lustrous.
3. **Diamond,** an allotrope of carbon, which is a *non-metal* is the hardest substance while **sodium and potassium *are* *metals*** but soft enough to be cut by knife.
4. **Mercury**, which *is a metal* but is a liquid at room temperature.
5. **Sodium, potassium, mercury, caesium and gallium** *are metals* but they have low melting and low boiling points.
6. **Diamond** is *the non-metal* with high melting and boiling point.
7. **Sodium, potassium and lithium** are metals with low density.

Chemical properties of Metals

Reaction with oxygen:

→ Metals combine with oxygen to form metal oxide.

Metal + O_2 → Metal oxide

Examples:

(i) $2Cu + O_2 \rightarrow 2CuO$

Copper oxide (black)

(ii) $4Al + 3O_2 \rightarrow 2Al_2O_3$
Aluminium oxide
(iii) $2Mg + O_2 \rightarrow 2MgO$
Magnesium oxide

The reactivity of different metals are different with O_2.
→ ***sodium (Na) and potassium (K)*** react with oxygen so vigorously that they catch fire if kept in open so they are kept immersed in kerosene.
→ ***Surfaces of magnesium (Mg), aluminium (Al), zinc (Zn), lead (Pb)*** are covered with a thin layer of oxide which prevent them from further oxidation.
→ ***iron (Fe)*** does not burn on heating but iron fillings burn vigorously.
→ ***Copper (Cu)*** does not burn on heating but is coated with black copper oxide.
→ ***gold (Au) and silver (Ag)*** does not react with oxygen.

→ **Amphoteric Oxides:** Metal oxides which react with both acids as well as bases to produce salts and water are called amphoteric oxides.
Examples:
(i) $Al_2O_3 + 6HCl \rightarrow 2AlCl_3 + H_2O$
(acid)
(ii) $Al_2O_3 + 2NaOH \rightarrow 2NaAlO_2 + H_2O$
(base) Sodium Aluminate

Anodising: Anodising is a process of forming a thick oxide layer of aluminium. Aluminium develops a thin oxide layer when exposed to air. This aluminium oxide coat makes it resistant to further corrosion. The resistance can be improved further by making the oxide layer thicker. This oxide layer can be dyed easily to give aluminium articles an attractive finish.

Reaction with water:

Metals react with water to produce metal oxide and hydrogen gas. Metal oxide that are soluble in water dissolve in it to further form metal hydroxide. But all metals do not react with water.
Metal + water ⇨ Metal oxide + hydrogen
Metal + Water ⇨ Metal hydroxide + Hydrogen

Most of the metals do not react with water. However, alkali metals react vigorously with water.
Metals like potassium and sodium react with cold water. The reaction is an exothermic reaction due to which the evolved hydrogen catches fire.
Reaction of sodium metal with water: Sodium metal forms sodium hydroxide and liberates hydrogen gas along with lot of heat when reacts with water.
$Na + H_2O \Rightarrow NaOH + H_2$ + heat
Reaction of potassium metal with water: Potassium metal forms potassium hydroxide and liberates hydrogen gas along with lot of heat when reacts with water.
$K + H_2O \Rightarrow KOH + H_2$ + heat
Reaction of calcium metal with water: Calcium forms calcium hydroxide along with hydrogen gas and when reacts with water.
$Ca + 2H_2O \Rightarrow Ca(OH)_2 + H_2$
Calcium starts floating because the bubbles of hydrogen gas stick to the surface of calcium.

Magnesium do not react with cold water. It reacts with hot water.
Reaction of magnesium metal with water: Magnesium metal reacts with water slowly and forms magnesium hydroxide and hydrogen gas.

$Mg + 2H_2O \Rightarrow Mg(OH)_2 + H_2$

Metals like aluminum, iron and zinc do not react either with cold or hot water. But they react with the steam.

$2Al + 3H_2O \Rightarrow Al_2O_3 + 2H_2$

Reaction of zinc metal with water: Zinc metal produces zinc oxide and hydrogen gas when steam is passed over it. Zinc does not react with cold water.

$Zn + H_2O \Rightarrow ZnO + H_2$

Reaction of Iron with water: Reaction of iron with cold water is very slow and come into notice after a long time. Iron forms rust (iron oxide) when reacts with moisture present in atmosphere.

Iron oxide and hydrogen gas are formed by passing of steam over iron metal.

$3Fe + 4H_2O \Rightarrow Fe_3O_4 + 4H_2$

Other metals usually do not react with water or react very slowly.

Reaction with Dilute Acid:

• Most metals react with acids to give a salt and hydrogen gas.

Metal + Dilute acid $\longrightarrow$ Salt + Hydrogen

Examples:

(i) $Fe + 2HCl \rightarrow FeCl_2 + H_2$

(ii) $Mg + 2HCl \rightarrow MgCl_2 + H_2$

(iii) $Zn + 2HCl \rightarrow ZnCl_2 + H_2$

(iv) $2Al + 6HCl \rightarrow 2AlCl_3 + 3H_2$

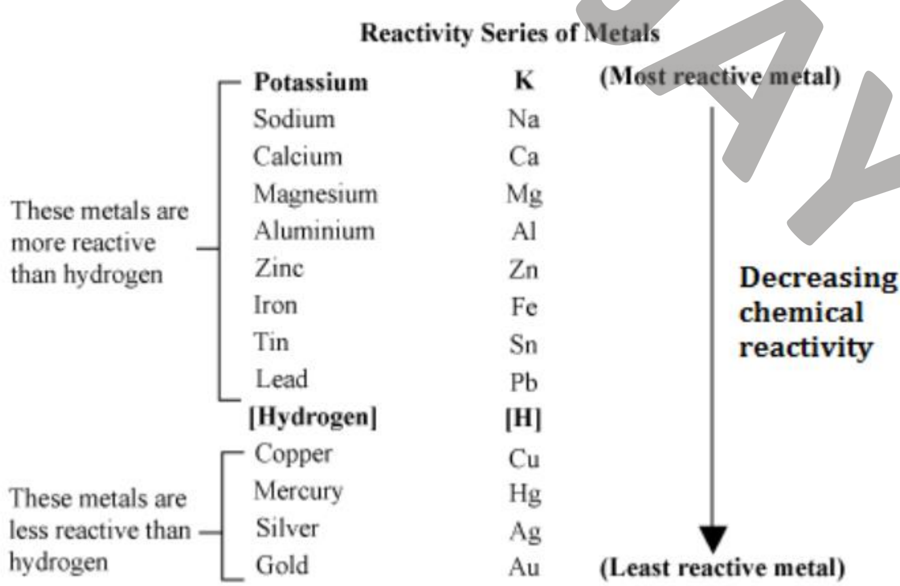
• Hydrogen gas is not evolved when a metal reacts with nitric acid. It's a strong oxidizing agent and oxidizes hydrogen produced to water and itself gets reduced to any of the nitrogen oxides (N_2O, NO, NO_2). But magnesium (Mg) and manganese (Mn) react with very dilute HNO_3 to evolve H_2 gas.

aquaregia:

• Freshly prepared mixture of concentrated hydrochloric acid and concentrated nitric acid in the ratio of 3:1.

• Is a highly corrosive, fuming liquid and one of the few reagents able to dissolve gold and platinum.

Reactivity series



How do metals and non-metals react: (metal and non-metal react to form ionic compound)

• The compounds formed by the transfer of electrons from a metal to a non-metal are known as ionic or electrovalent compounds.

• **Properties of ionic compound:**

a. **Physical Nature**: Solid and hard due to strong inter-ionic force of attraction; generally brittle.

b. **Melting and boiling points**: High melting and boiling points since a considerable force is required to break the strong inter-ionic attraction.

c. **Solubility:** Generally soluble in water but insoluble in solvents such as kerosene, petrol, etc.

d. **Conduction of electricity:**

• Conducts electricity through solution due to involvement of charged particles (ions).

• As movement of ions is not possible in solid state, due to rigid structure, so ionic compounds do not conduct electricity in solid state.

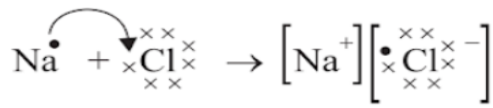
• In molten state ionic compounds conducts electricity.

E.g.: Formation of NaCl

$Na \rightarrow Na^+ + e^-$ **$Cl + e^- \rightarrow Cl^-$**

2, 8, 1   2, 8   **2, 8, 7   2, 8, 8**

Sodium cation   **Chloride anion**



Occurrence of Metals:

• **Mineral**: The elements or compounds, which occur naturally in the earth's crust.

• **Ore:** Mineral that contains high percentage of metal that can be extracted profitably from it.

• Every ore is a mineral but every mineral is not an ore.

Obtaining metal from ore:

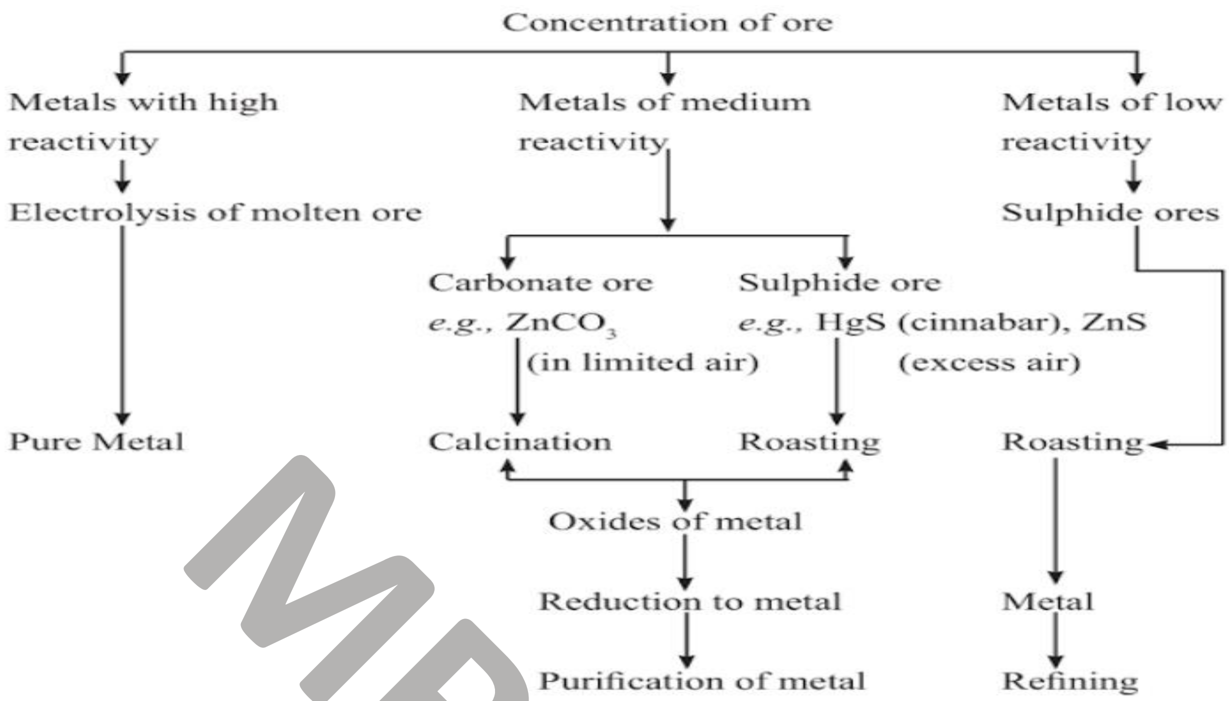
Different techniques are to be used for obtaining the metals from their ore on the basis of their reactivity.

• Ores mined from earth contain large amount of impurities such as sand, soil. etc. called gangue.

• Prior to the extraction of metal, based on the differences between the physical or chemical properties of gangue and the ore, different processes are used to remove gangue.

Extraction of metals

- The metals at the bottom of the reactivity series are least reactive. They are often found in a free state. For ex- gold, silver, platinum and copper are found in the free state.
- The metals at the top of the reactivity series are so reactive that they are never found in the free state.
- The metals in the middle of reactivity series are moderately reactive so they are found in the form as oxides, sulphides and carbonates.



I. **Extraction of metals low in the Activity Series:**

Cinnabar:- (HgS) is an ore of mercury. Mercury found in the sulphide form is named as cinnabar. Cinnabar (HgS) is first heated in the presence of oxygen and converted into mercuric oxide (HgO). Mercuric oxide is then reduced to mercury on heating.

$$2HgS(s) + 3O_2(g) \xrightarrow{\text{Heat}} 2HgO(s) + 2SO_2(g)$$

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

Similarly, copper which is found as Cu_2S in nature can be obtained from its ore by heating.

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

$$2Cu_2O + Cu_2S \xrightarrow{\text{Heat}} 6Cu(s) + SO_2(g)$$

II. Extraction of Metals in the middle of the Activity Series:

The metals in the middle of the activity series such as iron, zinc, lead, copper, etc., are moderately reactive. These are usually present as sulphides or carbonates in nature. It is easier to obtain a metal from its oxide, as compared to its sulphides and carbonates.

Sulphide ores are converted into oxides by heating strongly in the presence of excess air. This process is known as roasting.

The carbonate ores are changed into oxides by heating strongly in limited air. This process is known as calcination.

The chemical reaction that takes place during roasting and calcination of zinc ores can be shown

• ***Roasting –***

• Roasting is a process of converting sulphide ores into oxides by heating strongly in the presence of excess air.

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

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Calcination

• Calcination is a process of converting carbonate ores into oxides by heating strongly in limited air.

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

The metal oxides are then reduced to the corresponding metals by using suitable reducing agents such as carbon. For example, when zinc oxide is heated with carbon, it is reduced to metallic zinc.

$ZnO(s) + C(s) \rightarrow Zn(s) + CO(g)$

Besides using carbon (coke) to reduce metal oxides to metals, sometimes displacement reactions can also be used. The highly reactive metals such as sodium, calcium, aluminium, etc., are used because they can displace metals of lower reactivity from their compounds. For example, when manganese dioxide is heated with aluminium powder, the following reaction takes place –

$3MnO_2(s) + 4Al(s) \rightarrow 3Mn(l) + 2Al_2O_3(s) + Heat$

These displacement reactions are highly exothermic. The amount of heat evolved is so large that the metals are produced in the molten state. In fact, the reaction of iron(III) oxide (Fe_2O_3) with aluminium is used to join railway tracks or cracked machine parts. This reaction is known as the thermit reaction.

$Fe_2O_3(s) + 2Al(s) \rightarrow 2Fe(l) + Al_2O_3(s) + Heat$

Calcination	Roasting
It is done for carbonate ores.	It is done for sulphide ores.
Heating of ores in absence of oxygen.	Heating of S. ore in presence of oxygen.
CO_2 gas is released and Metal oxide is obtained	SO_2 gas is released and Metal oxide is obtained.
$ZnCO_2(s) \xrightarrow{heat} ZnO(s) + CO_2(g)$	$2ZnS(s) + 3O_2(g) \xrightarrow{heat} 3ZnO(s) + SO_2(g)$

Thermit reaction: Reaction of iron oxide with aluminium used to join railway tracks or cracked machine parts.
$Fe_2O_3(s) + 2Al(s)$ ------------ $2Fe(l) + Al_2O_3(s) + Heat$

Extraction of metals high In the Activity Series:

• Since these are very reactive metals and thus cannot be obtained by displacement reactions. These metals are obtained by electrolytic refining.

• They are generally obtained by electrolysis of their molten chlorides. Metals are deposited at cathode (negatively charged), while chlorine is liberated at anode.

At cathode $Na^+ + e^-$ ------------ Na

At anode $2C1^-$ -----------------$Cl_2 + 2e^-$

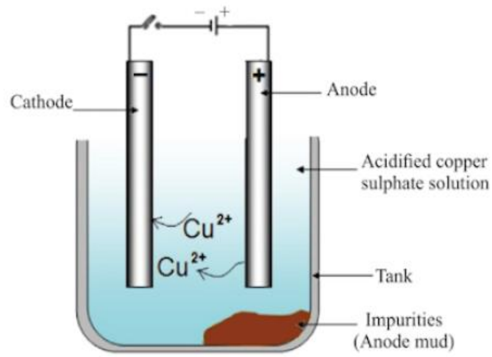
• Aluminium is obtained by electrolytic reduction of aluminium oxide.

Electrolytic refining:-

The most widely used method for refining impure metal is electrolytic refining.

Metals such as copper, zinc, tin, nickel, silver, gold etc are refined electrolytically. In this process impure metal is made anode and a thin strip of pure metal is made the cathode. A solution of metal salt is used as the electrolyte.

On passing the current, the pure metal from the anode dissolves into the electrolyte. An equivalent amount of pure metal from the electrolyte is deposited on the cathode. Insoluble impurities settle down at the bottom of the anode and are known are as anode mud



(i) Anode: Impure copper
(ii) Cathode: Strip of pure copper
(iii) Electrolyte: Solution of acidified copper sulphate
→ On passing the current through electrolyte, the impure metal from anode dissolves into the electrolyte.
→ An equivalent amount of pure metal from the electrolyte is deposited at the cathode.
→ The insoluble impurities settle down at the bottom of the anode and is called anode mud.

Corrosion

The surface of some metals get corroded when they are exposed to moist air for a long period of time. This is called corrosion.
Examples:
(i) Silver becomes black when exposed to air as it reacts with sulphur present in air to form a coating of silver sulphide.
(ii) Copper reacts with moist carbon dioxide in the air and gains a green coat of copper carbonate.
(iii) Iron when exposed to moist air acquires a coating of a brown flaky substance called rust.

• **Prevention of Corrosion**

→The rusting of iron can be prevented by painting, oiling, greasing, galvanizing, chrome plating, anodizing or making alloys.
→ **Galvanization**: It is a method of protecting steel and iron from rusting by coating them with a thin layer of zinc. The galvanised article is protected against rusting even if the zinc coating is broken. It is because zinc is more reactive then iron. So zinc reacts with the oxygen present in the atmosphere more easily and iron remain free of corrosion.
→ **Alloy**: An alloy is a homogeneous mixture of two or more metals, or a metal and a nonmetal. It is prepared by first melting the primary metal, and then, dissolving the other elements in it in definite proportions. It is then cooled to room temperature.
→ Examples of alloy:
(i) Iron : Mixed with small amount of carbon becomes hard and strong.
(ii) Steel : Iron + Nickel and chromium
(iii) Brass : Copper + Zinc
(iv) Bronze : Copper + Tin (Sn)
(v) Solder : Lead + tin
(vi) Amalgam : If one of the metal is mercury (Hg).

Note :-

1. Iron is the most widely used metal. But it is never used in its pure state. 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. Thus, if iron is mixed with some other substance, its properties change.
2. 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 parts of pure gold is alloyed with 2 parts of either copper or silver.

Properties of alloys:

- Alloy of a metal with mercury is known as **amalgam**.
- The electrical conductivity and melting point of an alloy is less than that of pure metals. For example, brass, an alloy of copper and zinc (Cu and Zn), and bronze, an alloy of copper and tin (Cu and Sn), are not good conductors of electricity whereas copper is used for making electrical circuits.
- Solder which is an alloy of lead and tin has a low melting point and is used for welding electrical wires.