Chapter 3
Metals and Non-Metals

Elements are classified as metals and non-metals based on different properties.

Physical Properties of Metals

- Metals are lustrous, that is, they have a property to shine.
- All metals exist as solids except mercury which is liquid at room temperature.
- They can be drawn into wires, this is known as **Ductility**.
- Metals can be converted into sheets, this is known as **Malleability**, except mercury
- They are good conductors of electricity and heat. Except for Lead and mercury.
- They have high density and high melting point. Exception-sodium and potassium have low melting points.

Physical Properties of Non-Metals

- They are not lustrous, that is, they do not have a shining surface except, graphite and iodine.
- They are generally soft, except for diamonds.
- They are non-ductile.
- They are non-malleable
- They are poor conductors of electricity and heat. Exception-graphite is a good conductor of electricity
- They have low density compared to metals and low melting point except for Diamond which has a high melting point.

Chemical Properties of Metals

Metals reaction with air/oxygen:

- Metals react with oxygen to form metal oxide.

Metal + O_2 → Metal oxide

For Example, Copper reacts with oxygen to form copper oxide.

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

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

- **Amphoteric Oxides**: Oxides of metals can react with both acids and bases to produce salt and water. Such oxides are known as amphoteric oxides.

$$Al_2O_3 + 6HCl \rightarrow 2AlCl_3 + 3H_2O$$

$$Al_2O_3 + 2NaOH \rightarrow \underset{\text{(Sodium aluminate)}}{2NaAlO_2} + H_2O$$

- Some of these metal oxides dissolve in water to form alkalis. For Example

$$Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$$

$$K_2O(s) + H_2O(l) \rightarrow 2KOH(aq)$$

- **Potassium** and **Sodium**: Highly reactive metals, catch fire in the open. Stored in kerosene to prevent fires.
- **Oxide Layers on Some Metals**: Metals like magnesium, aluminium, zinc, and lead develop a protective oxide layer at normal temperatures. This layer prevents further oxidation of the metal.
- **Iron's Reaction to Heating**: Iron itself does not burn when heated but iron filings burn vigorously in a flame.
- **Copper on Heating**: Copper doesn't burn but forms a black copper(II) oxide layer when hot.
- **Silver and Gold**: Silver and gold don't react with oxygen even at high temperatures.

Metals reaction with water:

- Some metals react with **cold water** to produce metal hydroxides and hydrogen gas. For example, **sodium** and **potassium** react violently and catches **fire** in water.

$$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g) + \text{heat energy}$$
$$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g) + \text{heat energy}$$

- **Calcium** reacts with water to form calcium hydroxide and hydrogen gas starts floating because the **bubbles of hydrogen** gas formed stick to its surface.

$$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$$

- **Magnesium** reacts with **hot water** to form magnesium hydroxide and hydrogen. It also starts floating due to the bubbles of hydrogen gas sticking to its surface.

- Metals like **aluminium, iron** and **zinc** do not react with **steam** to form the metal oxide and hydrogen.

$$2Al(s) + 3H_2O(g) \rightarrow Al_2O_3(s) + 3H_2(g)$$
$$3Fe(s) + 4H_2O(g) \rightarrow Fe_3O_4(s) + 4H_2(g)$$

- Metals such as **lead, copper, silver** and **gold** DO NOT REACT with water at all.

Metals reaction with acids:

- Metals also react with dilute **acids** to form **salt** and **hydrogen** gas. For example, magnesium reacts with dilute hydrochloric acid to form magnesium chloride and hydrogen.

$$\text{Metal} + \text{Dilute acid} \rightarrow \text{Salt} + \text{Hydrogen}$$

$$\mathbf{Mg + 2HCl \rightarrow MgCl_2 + H_2}$$

- Hydrogen gas is not evolved when a metal reacts with **nitric acid** because HNO_3 is a strong oxidising agent.
- HNO_3 oxidizes H_2 produced to water and gets reduced to nitrogen oxides (N_2O, NO, NO_2). *However, magnesium (Mg) and manganese (Mn) react with very dilute HNO_3 to evolve H_2 gas.*
- Copper does not react with dilute hydrochloric acid (HCl) because it is a less reactive metal.

Metals reaction with salt solutions of other metals:

Reactive metals can displace less reactive metals from their compounds in solution or molten form.

$$\text{Metal A} + \text{Salt solution of B} \rightarrow \text{Salt solution of A} + \text{Metal B}$$

$$\underset{\text{(Copper sulphate)}}{Fe(s) + CuSO_4(aq)} \rightarrow \underset{\text{(Iron sulphate)}}{FeSO_4(aq)} + Cu(s)$$

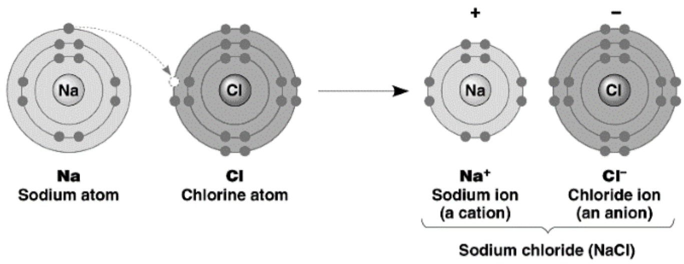
Reactivity Series

The series in which metals are arranged in the decreasing order of reactivity is known as the **reactivity series**.

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

Compounds of metals and non-metals: Ionic Compounds

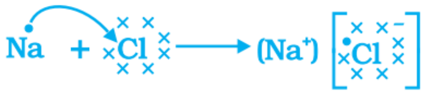
Compounds formed due to the transfer of electrons from a metal to a non-metal are known as **Ionic Compounds**.



Electron-dot structures showing the formation of NaCl:

$$\underset{2,8,1}{Na} \rightarrow \underset{\substack{2,8 \\ \text{(Sodium cation)}}}{Na^+} + e^-$$

$$\underset{2,8,7}{Cl} + e^- \rightarrow \underset{\substack{2,8,8 \\ \text{(Chloride anion)}}}{Cl^-}$$

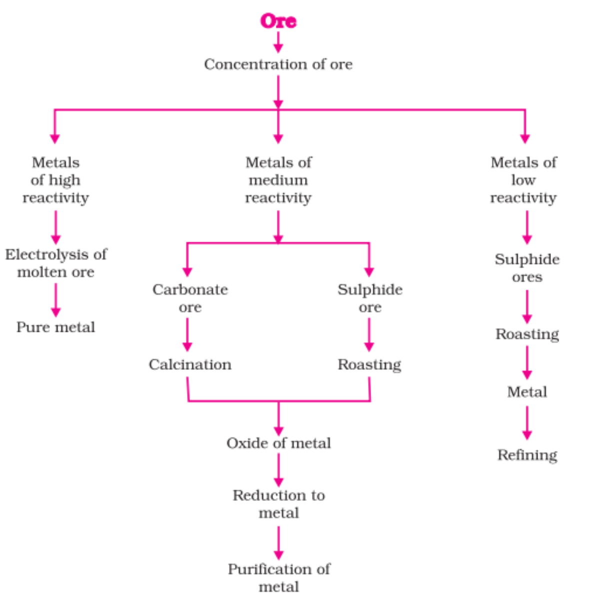


Properties of Ionic Compounds

- They are generally hard and solid.
- They have a high melting and boiling point.
- They are soluble in water but insoluble in inorganic solvents such as kerosene, petrol, etc.
- They are conductors of electricity in molten and solution states.

Occurrence of Metals

- Elements or compounds which occur naturally in earth crust are known as **minerals**. Minerals from which pure metals can be extracted are known as **ores**.



Steps for extraction of pure metals from its ores:

1. **Enrichment of the ore**: The first step is to increase the concentration of the metal in the ore. We want more of the valuable metal and less of the other impurities.
2. **Extraction of metals**: Different metals need different methods for extraction. For example, we might use heat to melt the metal out, or we might use chemical reactions to dissolve the metal and then separate it.
3. **Refining of metal:** Even after extraction, the metal may not be completely pure. It might still have some impurities. The third step involves refining the metal to make it as pure as possible. This can include processes like electrolysis or chemical reactions to remove any remaining impurities.

- **Gangue** - Ores contain different impurities in it such as sand, soil etc. These impurities are known as **Gangue**.

Extracting Metals which are *low* in activity series

Metals which are low in the activity series are unreactive. The oxides of such metals can be reduced to metals by heating alone.

For Example,

Cinnabar (HgS), an ore of mercury.

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

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

Cu_2S, an ore of copper.

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

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

Extracting Metals in the *middle* of the activity Series

- These metals are moderately reactive. They exist as sulphides or carbonates in nature.
- Before **reduction**, metal sulphides and carbonates must be converted into metal oxides.
- Sulphide ores are converted into oxides by heating strongly in the presence of excess air, this is known as **Roasting**.
- Carbonate ores are converted into oxides by heating in limited air. This is known as **Calcination**.

  Roasting

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

  Calcination

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

- **Reduction**:

Metal oxides can be reduced to metals using a reducing agent such as **carbon**.

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

Sometimes displacement reactions can also be used to reduce metal oxides to metals. The highly reactive metals such as sodium, calcium, aluminium, etc., are used as reducing agents.

For example:

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

Thermit Reaction

These displacement reactions are highly *exothermic*. The amount of heat evolved is so large that the metals are produced in the *molten* state.

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*.

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

Extracting metals towards the *top* of the activity series

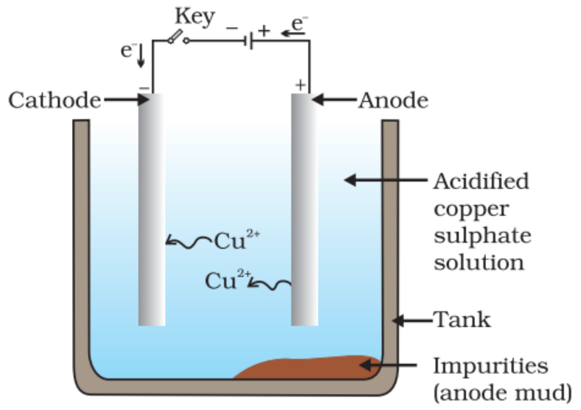
The metals are highly reactive. They cannot be obtained by heating. For Example, Sodium, magnesium and calcium are obtained by the electrolysis of their *molten chlorides*.

The metals are deposited at the cathode (the negatively charged electrode), whereas, chlorine is liberated at the anode (the positively charged electrode)

At cathode $Na^+ + e^- \rightarrow Na$
At anode $2Cl^- \rightarrow Cl_2 + 2e^-$

Refining of Metals

- Refining of impure metal is done using electrolytic refining. Impure copper is used as anode and a strip of pure copper is used as **cathode**.
- Acidified **copper sulphate** is used as an electrolyte. When an electric current is passed, the pure metal from the anode gets deposited in the electrolyte solution and same amount of pure metal from the electrolyte is deposited at 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.**



Corrosion

Corrosion is the gradual breakdown of metals caused by chemical reactions with their environment, leading to the weakening and decay of the metal's surface.

For example,

- **Silver** reacts with moist air and becomes **black** in colour due to *silver sulphide* coating.
- **Copper** reacts with moist carbon dioxide in the air and slowly loses its shiny brown surface and gains a **green** coat. This green substance is *basic copper carbonate*.
- **Iron** when exposed to moist air for a long time acquires a coating of a **brown** substance called *rust*.

Prevention of Corrosion

- Rusting of iron can be prevented by oiling, galvanising, painting, greasing etc.
- To protect steel and iron from rusting, a thin layer of zinc is coated on them, this is known as **galvanization**.

Alloy

Mixture of two or more metals or metal and non-metal is known as **alloy**. **For Example**,

- Brass is an alloy of copper and zinc.
- Bronze is an alloy of copper and tin.
- Solder is an alloy of lead and tin.