Dev Academy
Yamunanagar (Haryana)

Chapter - 3

Metals and Non-metals

Class 10

Revision Notes

SCIENCE

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Part – 1

Physical Properties of Metals and Non-metals

1.1 Physical Properties of Metals

- Metals are good conductor of heat and electricity because of free electrons in their metallic lattice.
  - Gold (Au), Silver (Ag) and Copper (Cu) are the best conductor of heat and electricity.
  - Lead (Pb), Mercury (Hg) and Stainless steel (Alloy) are the poor conductor of heat.
  - Tungsten (W) is the poor conductor of electricity.
- They are good conductor of sound. Hence, they are **sonorous**.
- In pure state, metals have a shining surface and can be polished. This property is called **metallic lustre**.
- All metals are generally hard (Except sodium (Na) and potassium (K), which are soft and can be cut with knife). However, hardness varies from metal to metal.
- Metals can be beaten into thin sheets. This property is called **malleability**.
- Metals can be drawn into thin wires. This property is called **ductility**.
  - Gold (Au) is the most ductile metal. A wire of about 2 km length can be drawn from one gram of gold.
- Metals are solid at room temperature (except mercury (Hg), which is liquid at room temperature)
- Metals have high melting and boiling points.
  - Tungsten (W) has the highest melting and boiling points whereas the mercury (Hg) has the lowest.
  - Gallium (Ga) and Cesium (Cs) also have low melting and boiling points. These two metals will melt on human body temperature.
- Metals have high density.
  - Iridium (Ir : 22.56 g/cc) and Osmium (Os : 22.57 g/cc) have the highest densities.
  - Lithium (Li : 0.54 g/cc) has the lowest density.
  - Iron (Fe) has density = 7.9 g/cc
- Metal oxides and metal hydroxides are basic in nature. (i.e., CaO and $NaOH$)
- Due to large atomic size, metals have low ionization enthalpies. Hence, they are **electropositive** in nature. They exist as cation in their compounds.
- Examples of metals: Gold (Au), Silver (Ag), Copper (Cu), Zinc (Zn), Mercury (Hg), Lead (Pb), Tungsten (W), Sodium (Na), Potassium (K), Iron (Fe) etc.

1.2 Physical Properties of Non-metals

- Non-metals are insulator of heat and electricity (except graphite).
- They are **dull**, they do not have metallic lustre (except Iodine).
- They are **brittle** in nature (therefore, they lack malleability and ductility).
- They usually have low density.
- Due to small atomic size, they have high ionization enthalpies. Hence, they are **electronegative**. This means that they have a strong tendency to hold electrons. They exist as anion in their compounds.
- Most of non-metals are gases at room temperature (i.e., Hydrogen (H_2), Oxygen (O_2), Nitrogen (N_2), Fluorine (F_2), Chlorine (Cl_2), and noble gases),
    - Few non-metals are solid at room temperature (i.e., Carbon (C), Sulphur (S), Phosphorus (P), and Iodine (I_2)).
    - Bromine (Br_2) is the only non-metal which is liquid at room temperature.
- They are not sonorous and do not produce a deep ringing sound when hit by another material.
- Most of non-metals produce acidic oxides (i.e., SO_2) when dissolved in water.
- Some non-metals show **allotropy** or polymorphism. Allotropy is the phenomenon in which a substance adopts different structural arrangements under different environmental conditions. All the allotropes of the substance are identical in their chemical properties, but they have significant differences in their physical properties.
    - Carbon has two crystalline allotropic form diamond and graphite.
    - Phosphorus has three allotropic forms; white phosphorus, red phosphorus, and Black phosphorus.
- Examples of non-metals: Fluorine (F), Chlorine (Cl), Bromine (Br), Iodine (I), Hydrogen (H), Oxygen (O), Nitrogen (N), and Noble gases.

NCERT Questions

Q1.: Give an example of a metal which;

(i) is a liquid at room temperature.
(ii) can be easily cut with a knife.
(iii) is the best conductor of heat.
(iv) is a poor conductor of heat.

Q2.: Explain the meanings of malleable and ductile.

Part – 2

Chemical Properties of Metals

2.1 Burning of Metals in Air

- Almost all metals combine with oxygen to form metal oxides.

$$\text{Metal} + \text{Oxygen} \longrightarrow \text{Metal oxide}$$

- All metals do not react with oxygen at the same rate. Reactivity varies metal to metal.
  - Sodium (Na) and potassium (K) react so rapidly with oxygen and catch fire if kept in the open. To prevent accidental fires, they are kept immersed in kerosene oil.
  - At ordinary temperature, the surface of some metals (i.e., Mg, Zn, Pb, Al) are covered with a thin layer of oxide. This oxide layer prevents the metal from further oxidation.
  - Iron (Fe) does not burn but iron filings can burn vigorously.
  - Copper does not burn, but the hot copper metal is coated with black coloured layer of copper (II) oxide (CuO).
- Example:
  - Magnesium (Mg) burns in air with dazzling white flame and produce white colour ash of magnesium oxide.

    $$Mg + O_2 \longrightarrow MgO$$

    [NOTE: Before burning, Magnesium ribbon is rubbed with sand paper in order to remove pre-existing protective layer of magnesium oxide from the surface of magnesium ribbon.]
  - When Copper (Cu) is heated in air, it combines with oxygen to form black layer of Copper (II) oxide.

    $$2Cu + O_2 \longrightarrow 2CuO$$

  - When Aluminium is heated in air, it combines with oxygen to form Aluminium oxide.

    $$4Al + 3O_2 \longrightarrow 2Al_2O_3$$

  - When Zinc is heated in air, it combines with oxygen to form Zinc Oxide.

    $$2Zn + O_2 \longrightarrow 2ZnO$$

2.1.1 Basic Nature of Metal Oxides

- As we know that metal oxides are basic in nature, they easily react with hydrochloric acid and sulphuric acid.

$$CaO + 2HCl \longrightarrow CaCl_2 + H_2O$$
$$CaO + H_2SO_4 \longrightarrow CaSO_4 + H_2O$$

  - Similarly:

$$Na_2O + 2\,HCl \longrightarrow 2NaCl + H_2O$$
$$Na_2O + H_2SO_4 \longrightarrow Na_2SO_4 + H_2O$$

- But some metal oxides (i.e., ZnO and Al_2O_3) are **amphoteric** in nature, it means they show both **acidic** and **basic** behavior, hence, they can react with both acid and base.

$$ZnO + 2HCl \longrightarrow ZnCl_2 + H_2O$$
$$ZnO + 2NaOH \longrightarrow Na_2ZnO_2 + H_2O$$

  - Similarly:

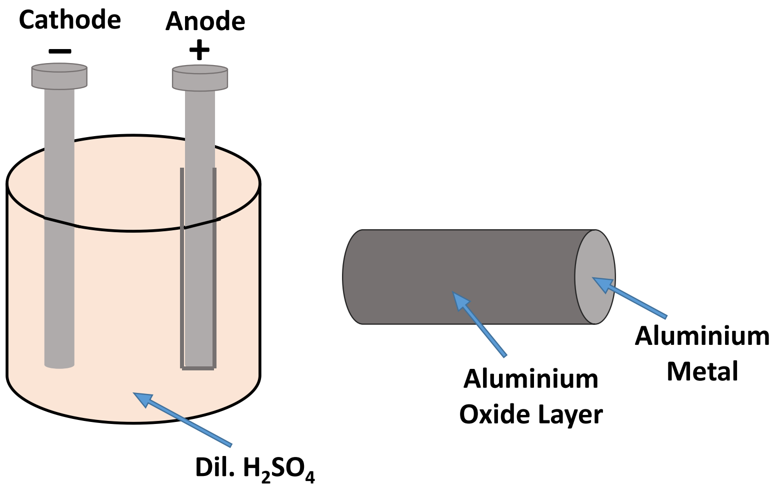
$$Al_2O_3 + 6HCl \longrightarrow 2AlCl_3 + H_2O$$
$$Al_2O_3 + 2NaOH \longrightarrow 2NaAlO_2 + H_2O$$

- Most of metal oxides (i.e., Fe_2O_3, Al_2O_3, CuO, ZnO etc.) are **insoluble** in water but some of metal oxides dissolves in water to form **base**. Sodium oxide, potassium oxide, and calcium oxide dissolve in water to produce sodium hydroxide, potassium hydroxide and calcium hydroxide respectively.

$$Na_2O + H_2O \longrightarrow 2NaOH$$
$$K_2O + H_2O \longrightarrow 2KOH$$
$$CaO + H_2O \longrightarrow Ca(OH)_2$$

2.1.2 Anodising

- ⇨ Anodising is a process of forming a thick layer of aluminium oxide over the surface of aluminium.
- ⇨ This aluminium oxide coat makes it resistant to further corrosion.
- ⇨ During anodising, a clean aluminium article is made the anode and is electrolysed with dilute sulphuric acid. The oxygen gas evolved at the anode reacts with aluminium to make a thicker protective oxide layer.



2.2 Reaction of Metals with Water

- ⇨ Metals react with water and produce a metal oxide and hydrogen gas.
- ⇨ Metal oxides that are soluble in water dissolve in it to further form metal hydroxide. But all metals & metal oxides do not react with water.
- ⇨ Example;

Metal + Water ⟶ Metal oxide + Hydrogen

Metal oxide + Water ⟶ Metal hydroxide

$$2Na + H_2O \longrightarrow Na_2O + H_2$$

$$Na_2O + H_2O \longrightarrow 2NaOH$$

$$2Na + 2H_2O \longrightarrow 2NaOH + H_2$$ Overall reaction

- ⇨ Potassium and Sodium react with cold water violently. The reaction is highly exothermic as a result evolved hydrogen gas immediately catches fire.

$$2K + 2H_2O \longrightarrow 2KOH + H_2 + \text{Heat}$$

$$2Na + 2\,H_2O \longrightarrow 2NaOH + H_2 + \text{Heat}$$

  - Due to low densities of Na and K, they float on water surface.

    Na = 0.97 gm/cc , K = 0.86 gm/cc

- Calcium also reacts with cold water but the reaction is less violent. Heat evolved is not sufficient for the hydrogen to catch fire.

$$Ca + 2H_2O \longrightarrow Ca(OH)_2 + H_2$$

  - The density of calcium (1.55 gm/cc) is higher than water, even that calcium start floating on water surface because the bubbles of hydrogen gas formed stick to the surface of metal, hence, start floating.

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

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

  - Density of magnesium (1.74 gm/cc) is higher than water, even that magnesium also starts floating due to bubbles of hydrogen gas sticking to its surface.

- Metals like Aluminium, Zinc and Iron do not react either with cold or hot water.
  - But they react with steam to form the metal oxide and hydrogen. They do not form hydroxides because their oxides are not soluble in H_2O.

$$2Al + 3H_2O_{(g)} \longrightarrow Al_2O_3 + 3H_2$$

$$3Fe + 4H_2O_{(g)} \longrightarrow FeO.Fe_2O_3 \text{ (or } Fe_3O_4) + 4H_2$$

$$Zn + H_2O_{(g)} \longrightarrow ZnO + H_2$$

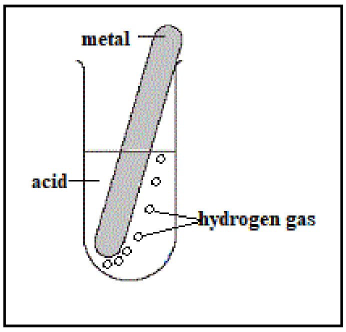
- Metals such as Lead, Copper, Silver, and Gold do not react with water at all.

2.3 Reaction of Metals with Acids

- Metals react with acids to give a salt and liberate hydrogen gas.

$$\text{Metal} + \text{dil. Acid} \longrightarrow \text{Salt} + \text{Hydrogen}$$

- Examples;



$$Mg + 2HCl \longrightarrow MgCl_2 + H_2$$
$$Mg + H_2SO_4 \longrightarrow MgSO_4 + H_2$$

$$2Al + 6HCl \longrightarrow 2AlCl_3 + 3H_2$$
$$2Al + 3H_2SO_4 \longrightarrow Al_2(SO_4)_3 + 3H_2$$

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$$
$$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2$$

$$2\,Fe + 6HCl \longrightarrow 2\,FeCl_3 + 3H_2$$
$$2\,Fe + 3H_2SO_4 \longrightarrow Fe_2(SO_4)_3 + 3H_2$$

$$Cu + HCl \text{ or } H_2SO_4 \longrightarrow \text{No reaction}$$
$$Ag + HCl \text{ or } H_2SO_4 \longrightarrow \text{No reaction}$$

- If you carefully observe the rate of formation of bubbles, it was the fastest in case of magnesium and slowest in case of iron.
- The reactivity of metals with acid decreased in the order:

$$Mg > Al > Zn > Fe$$

- In case of copper, silver, and gold no bubbles were seen and temperature also remained unchanged. This shows that these metals do not react with acids.

- Hydrogen gas is not evolved when metals react with nitric acid (HNO_3). It is because HNO_3 is strong oxidising agent. It oxidises the hydrogen gas produced in the reaction to water and itself gets reduce to any of the nitrogen oxides (NO, NO_2, N_2O, N_2O_5 etc.).

$$Zn + 4HNO_3 \longrightarrow Zn(NO_3)_2 + 2NO_2 + 2H_2O$$
$$Al + 4HNO_3 \longrightarrow Al(NO_3)_3 + NO + 2H_2O$$

- However, Magnesium reacts with very dilute HNO_3 to liberate Hydrogen gas.

$$Mg + 2HNO_3 \longrightarrow Mg(NO_3)_2 + H_2$$

- Aqua regia (Royal water), is a freshly prepared mixture of concentrated hydrochloric acid and concentrated nitric acid in the ratio of 3 : 1. It is one of the few reagents that can dissolve noble metals like gold and platinum, even though neither of these acids can do so alone.

2.4 Reaction of Metals with solution of other metal salts

- ⇨ More reactive metals can displace less reactive metals from their compounds in solution or molten form. This reaction is called single displacement reaction.
- ⇨ General Reaction;

$$A_{(s)} + BZ_{(aq.)} \longrightarrow AZ_{(aq.)} + B_{(s)}$$

Where, A = More reactive metal and, B = Less reactive metal

- ⇨ Examples;

	Reactivity
$Zn_{(s)} + CuSO_{4\,(aq.)} \longrightarrow ZnSO_{4\,(aq.)} + Cu_{(s)}$	Zn > Cu
$Cu_{(s)} + 2AgNO_{3\,(aq.)} \longrightarrow Cu(NO_3)_{2\,(aq.)} + 2Ag_{(s)}$	Cu > Ag
$Fe_{(s)} + CuSO_{4\,(aq.)} \longrightarrow FeSO_{4\,(aq.)} + Cu_{(s)}$	Fe > Cu

- ⇨ By studying the displacement reaction of various metals, give better evidence about the reactivity of metals. It is simple and easy to study, if metal 'A' displaces metal 'B' from its solution, it means metal 'A' is more reactive than metal 'B.'

- ⇨ **The reactivity Series**:
  It is a list of metals arranged in the order of their decreasing activities.
  Here is the list of reactivity of common metals studied through displacement reaction:

K > Na > Ca > Mg > Al > Zn > Fe > Pb > H > Cu > Ag > Au

More reactive Metals   Reactivity decreases $\longrightarrow$ Less reactive metals

NCERT Questions

Q1.: Why is sodium kept immersed in kerosene oil?

Q2.: Write equations for the reactions of

(i) iron with steam

(ii) calcium and potassium with water

Q3.: Samples of four metals A, B, C and D were taken and added to the following solution one by one. The results obtained have been tabulated as follows.

Metal	Iron(II) sulphate	Copper(II) sulphate	Zinc sulphate	Silver nitrate
A	No reaction	Displacement		
B	Displacement		No reaction	
C	No reaction	No reaction	No reaction	Displacement
D	No reaction	No reaction	No reaction	No reaction

Use the Table above to answer the following questions about metals A, B, C, and D.

(i) Which is the most reactive metal?

(ii) What would you observe if B is added to a solution of Copper (II) sulphate?

(iii) Arrange the metals A, B, C and D in the order of decreasing reactivity.

Q4.: Which gas is produced when dilute hydrochloric acid is added to a reactive metal? Write the chemical reaction when iron reacts with dilute H_2SO_4.

Q5.: What would you observe when zinc is added to a solution of iron(II) sulphate? Write the chemical reaction that takes place.

Part – 3

3.1 How do Metals and Non-metals react?

⇨ From the study of electronic configuration of the noble gases, it is clear that, they have 8 electrons in their valence shell except in case of 'He' which has 2. Hence, it was suggested that they possess stable electronic configuration.

Noble Gas	Atomic Number	Electronic Configuration
He	2	2
Ne	10	2,**8**
Ar	18	2,8,**8**
Kr	36	2,8,18,**8**
Xe	54	2,8,18,18,**8**
Rn	86	2,8,18,32,18,**8**

⇨ G.N. Lewis independently developed a theory of atomic bonding, "Electronic Theory of Chemical Bonding." According to this theory;

- The atoms can combine either by transfer of their valence electrons from one atom to another (gaining or losing) or by sharing of valence electrons in order to complete their octets or duplet in case of H, Li and Be to attain stable nearest noble gas configuration. This is known as **octet rule**.

Type of Element	Element	Atomic Number	Electronic Configuration
Metals	Na	11	2,8,**1**
	Mg	12	2,8,**2**
	Al	13	2,8,**3**
	K	19	2,8,8,**1**
	Ca	20	2,8,8,**2**
Non-metals	N	7	2,**5**
	O	8	2,**6**
	F	9	2,**7**
	P	15	2,8,**5**
	S	16	2,8,**6**
	Cl	17	2,8,**7**

⇨ When metal and non-metals react with each other, they usually form ionic compound (Salts).

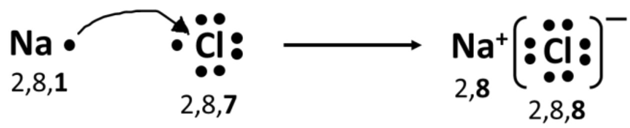
- When a bond is formed by complete transference of one or more electrons from one atom to another, hence, acquire the stable nearest noble gas configuration, the bond formed is called ionic bond or electrovalent bond.

- The metal atom, loses the electron, acquires a positive charge called cation. $\mathbf{A \longrightarrow A^+ + e^-}$
- The non-metal atoms, gain the electrons, acquire a negative charge called anion. $\mathbf{B + e^- \longrightarrow B^-}$
- The electrostatic attraction between the oppositely charge ions called ionic bond. $\mathbf{A^+ + B^- \longrightarrow AB}$

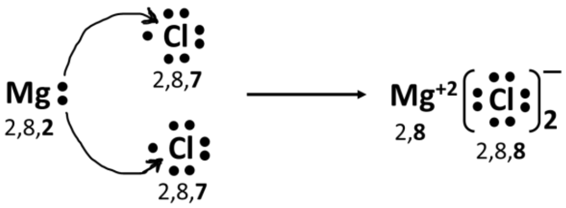
⇨ Example 1.: Formation of NaCl

$$\underset{2,8,1}{Na} \longrightarrow \underset{2,8}{Na^+} + e^-$$

$$\underset{2,8,7}{Cl} + e^- \longrightarrow \underset{2,8,8}{Cl^-}$$



⇨ Example 2.: Formation of $MgCl_2$



$$\underset{2,8,2}{Mg} \longrightarrow \underset{2,8}{Mg^{+2}} + 2e^-$$

$$\left.\underset{2,8,7}{Cl} + e^- \longrightarrow \underset{2,8,8}{Cl^-}\right] \times 2$$

3.2 Properties of Ionic Compounds

⇨ **Physical State:** Ionic compounds usually exist in solid state. They are generally hard and brittle in nature.

⇨ **Melting and Boiling Points:** Ionic compounds possess high melting and boiling points. This is because ions are tightly held together by strong electrostatic forces and hence huge amount of energy is required to break the crystal lattice.

⇨ **Solubility:** Ionic compounds are easily soluble in polar solvent like water because have high dielectric constant (about 80). Furthermore, the ions may interact with solvent molecules to liberate energy called hydration enthalpy, which is sufficient to break the attractive forces between ions.

   - However, ionic compounds are insoluble or slightly soluble in non-polar solvents like carbon tetrachloride, benzene, acetone, oils etc. because having low dielectric constants.

⇨ **Electrical conductivity:** Ionic compounds are good conductors of electricity in aqueous or molten state, because their ions free to move. But they are bad conductors of electricity in solid state, because ions are not free to move.
   - Note: They conduct electricity via movement of ions hence they are called electrolytic conductors.

NCERT Questions

Q1.: Give Answer;

(i) Write the electron-dot structures for sodium, oxygen, and magnesium.
(ii) Show the formation of Na_2O and MgO by the transfer of electrons.
(iii) What are the ions present in these compounds?

Q2.: Why do ionic compounds have high melting points?

Part – 4

Basic Metallurgical Processes

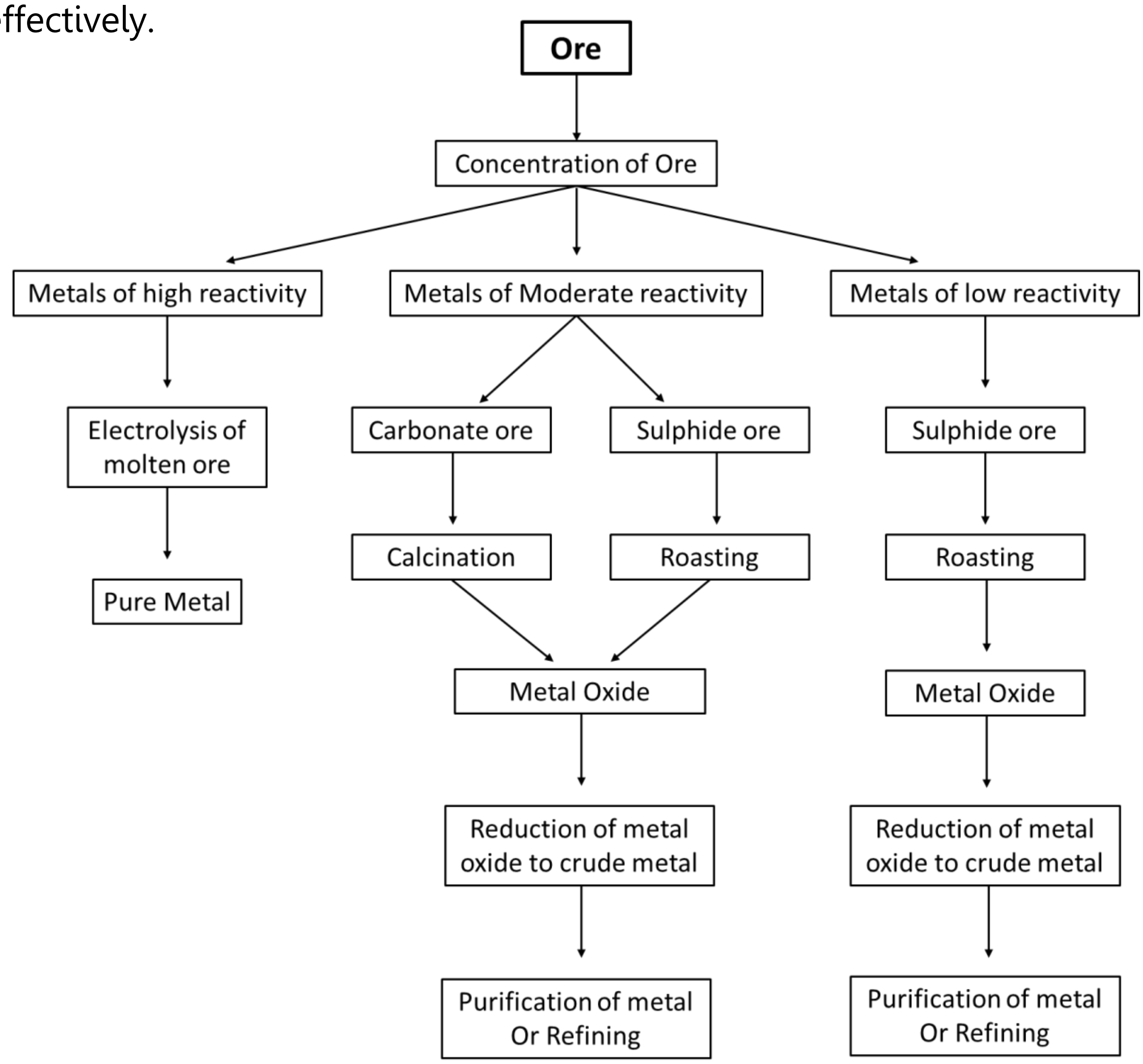
4.1 Occurrence of Metals

- The occurrence of metals is primarily associated with the presence of mineral deposits in the Earth's crust.
- The minerals from which metals can be extracted are called ore.
- Ore and minerals are related terms, but they have distinct meanings.
- Minerals are naturally occurring inorganic solid substances with a specific chemical composition. On the other hand, ore refers to a special type of mineral deposit that contains a high concentration of metals, which can be extracted through various metallurgical processes.
- Therefore, we can say that all ores are minerals, but not all minerals are ores.
- Here is a list of some common ores of aluminium, iron, zinc, and copper along with their chemical formulas:

Metals	Ore	Chemical Formula
Aluminium	Bauxite	$Al_2O_3{\cdot}2H_2O$
Iron	Hematite Magnetite Siderite	Fe_2O_3 Fe_3O_4 $FeCO_3$
Zinc	Sphalerite (Zinc blende) Smithsonite	ZnS ZnCO3
Copper	Chalcopyrite Chalcocite Malachite	$CuFeS_2$ Cu_2S $Cu_2CO_3(OH)_2$
Mercury	Cinnabar	HgS

4.2 Extraction of Metals

- Extraction of metals refers to the process of obtaining pure metals from their naturally occurring mineral sources (known as ores).
- Metals can be found in two forms in nature based on their reactivity:
  - **Free State:** Some metals, like gold, silver, copper, and platinum, are very less reactive or almost inert. So that they exist naturally in their pure form (free state) without combining with other elements.
  - **Compound State:** Other metals, like Iron, zinc, magnesium, aluminum, and calcium, are somewhat reactive and combine with elements like sulphur and oxygen to form compounds (called ores).
    - Note: Copper and silver are also found in the combined state as their sulphide or oxide ore.
- Based on their reactivity, metals can be classified into three categories: (i) Metals of low reactivity, (ii) Metals of medium reactivity, and (iii) Metals of high reactivity. Each category requires specific techniques for extracting the metals effectively.



⇨ **Concentration of ore / Enrichment of ore / Ore beneficiation:**

- Ores extracted from the earth often contain significant amounts of impurities like soil, sand, etc., collectively referred to as **gangue**.
- To extract the valuable metal from the ore, it is essential to eliminate these impurities. This removal of gangue from the ore is known as ore concentration or beneficiation.
- The process involves utilizing the differences in physical or chemical properties between the gangue and the ore.
- Following techniques are employed for ore concentration;
  - Crushing and Grinding
  - Gravity Separation
  - Froth Flotation
  - Magnetic Separation
  - Leaching

4.2.1 Extraction of Metals of low reactivity

⇨ Metals with low reactivity are usually found in their elemental or pure form in nature, but occasionally they can also be found combined with sulphur to form metal sulphides. To extract these metals from their compounds, we employ two main techniques.

⇨ In the first technique, certain metal sulphides like copper(I) sulphide and mercury(II) sulphide need to be converted into metal oxides before obtaining the metal. This conversion is achieved through a process known as **roasting**, where the metal sulphides are heated in the presence of oxygen. As a result, the sulphur combines with oxygen to form sulphur dioxide, leaving behind metal oxides. After roasting, we proceed with reducing the metal oxides to obtain the pure metal in its metallic form using a suitable method.

- $2HgS_{(s)} + 3O_{2\,(g)} \xrightarrow{Heat} 2HgO_{(s)} + 2SO_{2\,(g)}$ (Roasting)

  $2HgO_{(s)} \xrightarrow{Heat} 2Hg_{(l)} + O_{2\,(g)}$ (Thermal decomposition)

  $2HgO + HgS \xrightarrow{Heat} 3Hg + SO_2$ (Self-Reduction during roasting)

- $2Cu_2S + O_2 \xrightarrow{Heat} 2Cu_2O + 2SO_2$ (Roasting)

  $2Cu_2O + Cu_2S \xrightarrow{Heat} 6Cu + SO_2$ (Self-Reduction during roasting)

- In the second technique, some metal sulphides, such as gold or silver sulphides, are directly converted into the metal itself. This process is called **leaching**. In leaching, the metal sulphides are dissolved in a suitable solvent, forming metal-containing solutions. From these solutions, we can extract the pure metal by using a displacement reaction with another electropositive metal, usually zinc.
    - $Ag_2S_{(s)} + 4NaCN_{(aq.)} \longrightarrow 2Na[Ag(CN)_2]_{(aq.)} + Na_2S_{(s)}$
      $2Na[Ag(CN)_2]_{(aq.)} + Zn_{(s)} \longrightarrow Na_2[Zn(CN)_4]_{(aq.)} + 2Ag_{(s)}$
    - $Au_2S_{(s)} + 4NaCN_{(aq.)} \longrightarrow 2Na[Au(CN)_2]_{(aq.)} + Na_2S_{(s)}$
      $2Na[Au(CN)_2]_{(aq.)} + Zn_{(s)} \longrightarrow Na_2[Zn(CN)_4]_{(aq.)} + 2Au_{(s)}$

4.2.2 Extraction of Metals of moderate or medium reactivity

- Metals with moderate reactivity, including iron, zinc, lead, copper, and aluminium, are commonly found in nature as sulphides, oxides, or carbonates.
- To extract these metals, it is more convenient to obtain them from their oxides rather than from their sulphides or carbonates. Thus, before the reduction process can take place, the metal sulphides and carbonates must be transformed into metal oxides.
- The sulphide ores are converted into oxides through a process called **roasting**. During roasting, the sulphide ores are strongly heated in the presence of excess air. This high-temperature treatment causes the sulphur in the sulphide ores to combine with oxygen from the air, forming sulphur dioxide gas. Consequently, metal sulphides are converted into metal oxides.
    - $2ZnS_{(s)} + 3O_{2\,(g)} \xrightarrow{Heat} 2ZnO_{(s)} + 2SO_{2\,(g)}$ (Roasting)
- Similarly, the carbonate ores are converted into oxides through a process known as **calcination**. In calcination, the carbonate ores are strongly heated, but this time in a limited supply of air. As a result, the metal carbonates lose carbon dioxide, leaving behind the metal oxides.
    - $ZnCO_{3\,(s)} \xrightarrow{Heat} ZnO_{(s)} + CO_{2\,(g)}$ (Calcination)

Calcination	Roasting
1. It is employed for carbonate ores 2. Carbonate ores are heated strongly in a limited supply of air or in the absence of air (in a closed furnace) 3. CO_2 gas is evolved during calcination. 4. Self-reduction is not taking place during calcination.	1. It is employed for sulphide ores 2. Sulphide ores are heated strongly in a full supply of air or in the presence of air (in a opened furnace) 3. SO_2 gas is evolved during roasting 4. Self-reduction is taking place during roasting process.

⇨ The metal oxides are then reduced to the corresponding metals by using suitable reducing agents such as carbon (coke).

- $ZnO_{(s)} + C_{(s)} \xrightarrow{Heat} Zn_{(s)} + CO_{(g)}$

  The Reduction of metal oxide using carbon is called **smelting**.

⇨ In addition to using carbon (coke) as a reducing agent to convert metal oxides into metals, another method involves the use of displacement reactions.

⇨ Highly reactive metals like sodium, calcium, and aluminium are utilized as reducing agents in these reactions because they have the ability to displace metals of lower reactivity from their compounds.

⇨ These displacement reactions are highly exothermic. The tremendous heat produced during these displacement reactions causes the metals to melt and turn into a molten state.

- **Example 1.:** Manganese can be extracted from manganese dioxide by heating it together with aluminium powder.

  $$3MnO_{2\,(s)} + 4Al_{(s)} \longrightarrow Mn_{(l)} + 2Al_2O_{3\,(s)} + Heat$$

- **Example 2.:** Iron can be extracted from iron(III) oxide by heating it along with aluminium powder.

  $$Fe_2O_{3\,(s)} + 2Al_{(s)} \longrightarrow 2Fe_{(l)} + Al_2O_{3\,(s)} + Heat$$

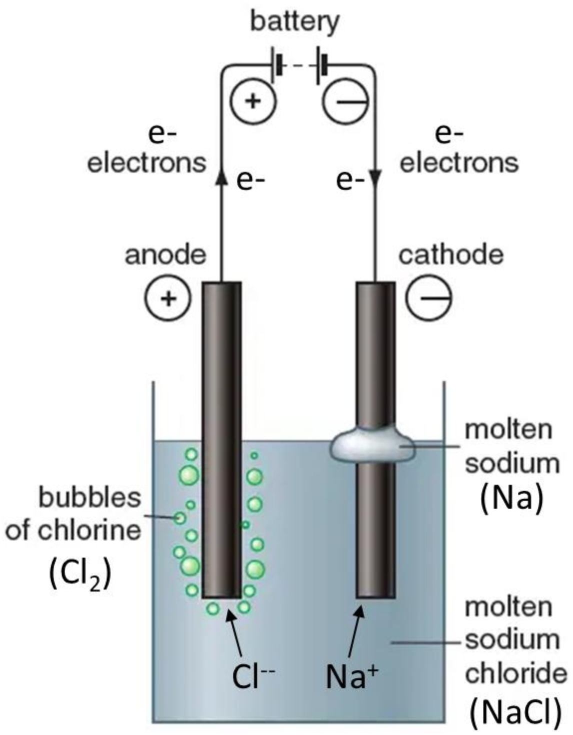
  As a result of this reaction, iron is produced in a molten state, giving it significant practical importance in various applications, such as joining railway tracks or repairing cracked machine parts. This reaction is commonly referred to as the **thermite reaction**.

4.2.3 Extraction of Metals of high reactivity

- When we attempt to extract metals of high reactivity (i.e., Na, Mg, Ca, Al etc.) from their compounds using traditional methods like heating with carbon, it proves ineffective because these metals have a strong affinity for oxygen. As a result, carbon is unable to displace oxygen from their compounds.
- To overcome this challenge, we turn to a different and powerful technique called **electrolytic reduction**.
- In electrolytic reduction method, metals like sodium, magnesium, calcium, and others are obtained by subjecting their molten chlorides and oxides to electrolysis.
- Electrolysis involves passing an electric current (DC) through the molten compound, causing the metal ions to migrate towards the negatively charged electrode (cathode), where they get deposited as pure metals. Meanwhile, the negatively charged ions, like chloride ions, move towards the positively charged electrode (anode) and get discharged, liberating substances like chlorine gas.
  - Extraction of Sodium or Electrolysis of Molten NaCl

    Dissociation into ions; $2NaCl \rightleftharpoons 2Na^{+} + 2Cl^{-}$

    - At Cathode (-): $2Na^{+} + 2e^{-} \longrightarrow 2Na$
    - At Anode (+): $2Cl^{-} \longrightarrow Cl_2 + 2e^{-}$

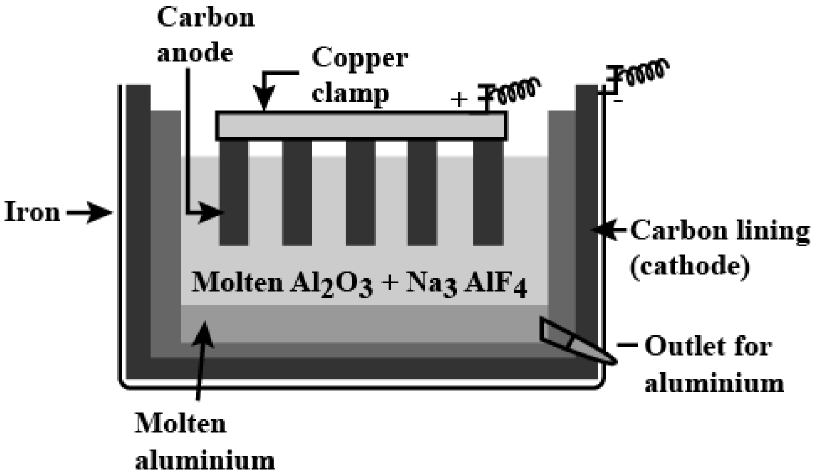


- Extraction of Aluminium or Electrolysis of Al_2O_3 or Hall Heroult process

  Dissociation into ions; $Al_2O_3 \rightleftharpoons 2Al^{+3} + 3O^{-2}$

  - At Cathode (-): $2Al^{+3} + 6e^- \longrightarrow 2Al$
  - At Anode (+): $3O^{2-} \longrightarrow 3O + 6e^-$

    $3C + 3O \longrightarrow 3CO$

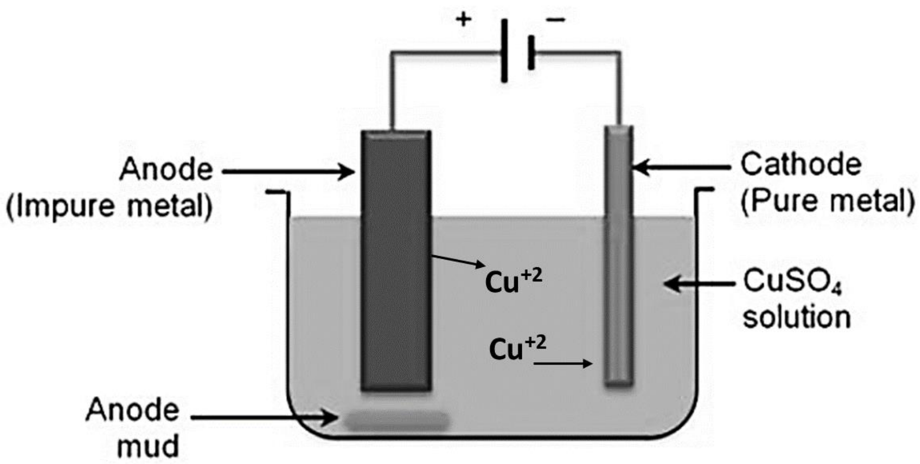


4.3 Refining of Metals

- Metals obtained through the reduction processes mentioned earlier are not entirely pure as they still contain impurities.
- To achieve high purity levels, these impurities must be eliminated from the metals. The process of purifying raw metal to high quality pure metal is called refining of metal.
- The refining process varies depending on the type of metal and the nature of the impurities. Common methods of refining include: (i) Electrolytic Refining (ii) Distillation (iii) Zone Refining (iv) Cupellation (v) Liquation

- The most commonly employed and widely accepted method for refining impure metals is known as **electrolytic refining**.
  - Electrolytic refining is commonly used for metals like copper, zinc, tin, aluminum, nickel etc.
  - In this process, the impure metal is made the anode, and a pure metal sheet acts as the cathode.

- These electrodes are dipped into an aqueous solution containing metal ions of the same metal that needs to be purified.
- When an electric current is passed through the electrolyte solution, the impure metal is oxidized at the anode and dissolves into the solution, whereas the soluble metal ions are reduced at the cathode and get deposited as pure metal on the cathode.
- The insoluble impurities settle down at the bottom of the anode and are known as anode mud. Whereas, the soluble impurities go into the solution.



- At Anode (+): $Cu_{(s)} \longrightarrow Cu^{+2} + 2e^-$
- At Cathode (-) $Cu^{+2} + 2e^- \longrightarrow Cu_{(s)}$

NCERT Questions

Q1.: Define the following terms.
(i) Mineral
(ii) Ore
(iii) Gangue

Q2.: Name two metals which are found in nature in the free state.

Q3.: What chemical process is used for obtaining a metal form its oxide?

Part – 5

Corrosion

5.1 Corrosion

- Corrosion is a natural process that affects various metals when they come in contact with air and moisture. In Chapter 1, you learned about some common examples of corrosion:
  - Silver articles turn black over time as they react with sulphur in the air, forming a layer of silver sulphide.
  - Copper, when exposed to moist carbon dioxide in the air, gradually loses its shiny brown surface and acquires a green coating of the basic copper carbonate.
  - Iron, on prolonged exposure to moist air, develops a brown flaky substance known as rust.

- To understand the conditions under which iron rusts, you can conduct a simple experiment using test tubes and iron nails:
  - Take three test tubes and place clean iron nails inside each of them.
  - Label the test tubes as A, B, and C.
  - Test tube A: Add some water and cork it.
  - Test tube B: Pour boiled distilled water and add a little oil. Cork it, and the oil will prevent air from dissolving in the water.
  - Test tube C: Place anhydrous calcium chloride inside and cork it. The calcium chloride will absorb any moisture from the air.
  - Leave the test tubes for a few days and then observe the results.
  - You will notice that the iron nails rust in test tube A, where they are exposed to both air and water.
  - However, the nails do not rust in test tubes B and C.
  - In test tube B, the nails are only exposed to water, and in test tube C, the nails are exposed to dry air.
  - This experiment tells us that both air and water are essential for the rusting of iron. When iron is exposed to both moisture and oxygen in the air, it undergoes a chemical reaction that leads to the formation of rust. However, if either the moisture or the oxygen is removed from the equation, the rusting process is hindered.

5.1.1 Prevention of Corrosion

- The process of rusting can be avoided through various techniques. Some effective methods to prevent rusting include painting, oiling, greasing, galvanizing, chrome plating, anodizing, or making alloys.

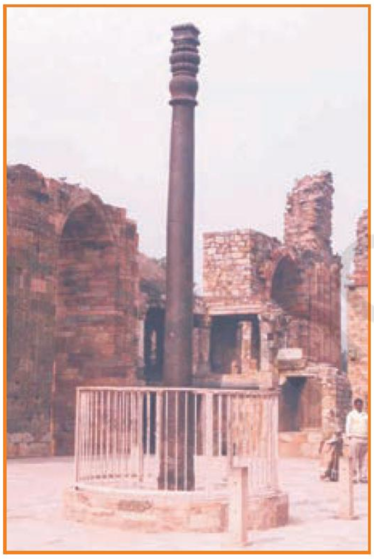
- **Galvanization** - Shielding against Rust:
  - Galvanization is an effective way to protect steel and iron from rusting. It involves coating the surface of iron or steel with a thin layer of zinc.
  - The galvanized coating acts as a sacrificial layer for the underlying metal.
  - If the galvanized metal gets scratched, the zinc coating corrodes first, protecting the metal underneath.
  - This is because zinc has a higher affinity for oxygen than iron or steel, so it readily reacts with oxygen and moisture to form zinc oxide, leaving the base metal protected.

- **Alloy Formation or Alloying** - Unleashing Metal Superpowers
  - An alloy is a homogeneous mixture of two or more metals (or a metal and a non-metal).
  - Alloying is an excellent method of improving the properties of a metal.
  - For instance, pure iron is quite soft and stretches easily when heated. However, when we mix a small amount of carbon (around 0.05%) with iron, it transforms into steel, which is strong and durable.
  - Furthermore, the addition of nickel and chromium to iron creates stainless steel, a powerful metal that never rusts.
  - Not even iron, any metal's properties can be altered by mixing it with another substance, which could be either a metal or a non-metal.

  - **Here are some other examples of alloys:**
    - Pure gold (known as 24 carat gold) is soft and not suitable for making jewellery. To make it hard and durable, pure gold is alloyed with either silver or copper. In India, 22 carat gold is commonly used for making ornaments. It means that 22 parts of pure gold is alloyed with 2 parts if either copper or silver.
    - Brass is an alloy of copper and zinc (Cu and Zn), while bronze is an alloy of copper and tin (Cu and Sn).
      - Compared to pure copper, brass and bronze are not as good conductors of electricity.

  - Despite this, brass is valued for its corrosion resistance and is commonly used in fittings, musical instruments, and decorative items.
  - Bronze is known for its strength, durability, and wear resistance, making it suitable for statues, sculptures, and bearings.
- Solder, which is an alloy of lead and tin (Pb and Sn), possesses a low melting point, making it ideal for joining electrical wires together through the process of soldering.
- Amalgam, an alloy of mercury, has a lower melting point and lower electrical conductance compared to pure mercury. Amalgam is useful in various applications, particularly in dentistry, where it is commonly used for dental fillings.

The wonder of ancient Indian metallurgy.



- The iron pillar near the Qutub Minar in Delhi is a wonder of ancient Indian metallurgy.
- It was built more than 1600 years ago by skilled Indian ironworkers.
- These ironworkers developed a process that prevented the iron pillar from rusting.
- Scientists from all over the world have examined it due to its impressive rust resistance.
- The iron pillar stands tall at 8 meters and weighs 6 tonnes (6000 kg), showcasing its remarkable craftsmanship and enduring quality.