Chapter 3: Metals And Non-Metals

Introduction

Metals and non-metals are very important in our daily life. They are used in making utensils, machines, wires, buildings, medicines and many other things.

Different elements show different properties, and based on these properties they are classified as **metals and non-metals**. Metals generally show properties like **lustre, malleability, ductility and good conductivity**, while non-metals usually do not show these properties.

Metals

Metals are elements that generally **lose electrons to form positive ions (cations)** and show properties like **lustre, malleability, ductility, and good conductivity of heat and electricity**.

Examples : Iron, Copper, Aluminium, Gold, Silver, Zinc

Non- Metals

Non-metals are elements that generally **gain electrons to form negative ions (anions)** and usually **do not show properties like malleability, ductility, or high electrical conductivity**.

Examples : Oxygen, Nitrogen, Carbon, Sulphur, Hydrogen

Physical Properties of Metals

Metals show some characteristic physical properties.

1. Lustre

Metals have a **shiny surface**. This property is called **metallic lustre**.

Example:
Gold, silver and aluminium are shiny metals.

Because of their lustre, metals are used in **jewellery and decorative items**.

2. Malleability

The property of metals by which they can be **beaten into thin sheets** is called **malleability**.

Example:
Gold and silver are highly malleable metals.

Gold is the **most malleable metal**.

Uses:
Aluminium sheets are used for **wrapping food**.

3. Ductility

The property of metals by which they can be **drawn into thin wires** is called **ductility**.

Example:
Copper and aluminium wires.

Gold is the **most ductile metal**.
A wire of about **2 km length can be drawn from 1 gram of gold**.

4. Hardness

Most metals are **hard**.

Example:
Iron is a hard metal.

However, hardness varies from metal to metal.

5. Conductivity

Metals are **good conductors of heat and electricity**.

Examples: Copper and aluminium are used in **electrical wires**.

Silver and copper are the **best conductors of heat**.

6. High Melting and Boiling Points

Most metals have **high melting and boiling points**.

Example:
Iron and copper have high melting points.

7. Sonorous Nature

Metals produce a **ringing sound when struck**.

This property is called **sonority**.

Example:
School bells are made of metals.

Exceptions in Metals

Some metals do not follow typical properties.

METAL	EXCEPTION
MERCURY	Liquid at room temperature
SODIUM AND POTASSIUM	Very soft metals

GALLIUM AND CAESIUM	Very low melting points

Physical Properties of Non-Metals

Non-metals generally show properties opposite to metals.

General Properties

1. Non-lustrous (dull appearance)
2. Poor conductors of heat and electricity
3. Brittle (break easily)
4. Not malleable
5. Not ductile
6. Low melting and boiling points

Examples:
Carbon, sulphur, oxygen, nitrogen.

Exception

Graphite (a form of carbon) is a **good conductor of electricity** even though it is a non-metal.

Diamond (another form of carbon) is the **hardest natural substance**.

Chemical Properties of Metals

Metals react with different substances such as **oxygen, water, acids and salt solutions**.

1. Reaction of Metals with Oxygen

Metals react with oxygen to form **metal oxides**.

Metal + Oxygen → Metal Oxide

Example: $2Mg + O_2 \rightarrow 2MgO$

Magnesium burns with a **bright white flame** to form magnesium oxide.

Metal oxides are generally **basic in nature**.

Amphoteric Oxides

Some metal oxides show both **acidic and basic nature**.
These are called **amphoteric oxides**.

Examples:
Aluminium oxide (Al_2O_3)
Zinc oxide (ZnO)

Reaction with acids:

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

Reaction with bases:

$Al_2O_3 + 2NaOH \rightarrow 2NaAlO_2 + H_2O$

2. Reaction of Metals with Water

Metals react with water to form **metal hydroxide and hydrogen gas**.

Metal + Water → Metal hydroxide + Hydrogen

Example:

$2Na + 2H_2O \rightarrow 2NaOH + H_2$

Some metals react violently with water.

Reactivity of Metals with Water

METAL	REACTION
POTASSIUM	Very violent reaction
SODIUM	Very vigorous reaction
CALCIUM	Reacts with cold water
MAGNESIUM	Reacts with hot water
ALUMINIUM, ZINC, IRON	React with steam
COPPER, SILVER, GOLD	Do not react

3. Reaction of Metals with Acids

Metals react with dilute acids to form **salt and hydrogen gas**.

Metal + Dilute Acid → Salt + Hydrogen

Example:

$Zn + 2HCl \rightarrow ZnCl_2 + H_2$

However, **nitric acid (HNO_3)** usually does not produce hydrogen gas because it is a **strong oxidising agent**.

4. Reaction with Metal Salt Solutions

A **more reactive metal displaces a less reactive metal** from its compound.

Example:

$Fe + CuSO_4 \rightarrow FeSO_4 + Cu$

Iron displaces copper from copper sulphate solution.

This is called a **displacement reaction**.

Reactivity Series

The **reactivity series** is a list of metals arranged in the order of **decreasing reactivity**.

Reactivity Series:

K Na Ca Mg Al	Electrolysis
Zn Fe Pb Cu	Reduction using carbon
Ag Au	Found in native state

Metals at the **top are most reactive** and metals at the **bottom are least reactive**.

Formation of Ionic Compounds

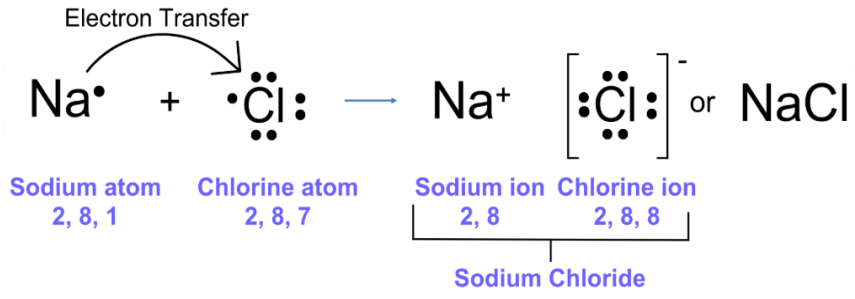
Metals lose electrons and form **positive ions (cations)**.

Non-metals gain electrons and form **negative ions (anions)**.

Example: Formation of Sodium Chloride

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

$Na^+ + Cl^- \rightarrow NaCl$



Compounds formed by transfer of electrons are called **ionic compounds or electrovalent compounds**.

Example:
NaCl, $MgCl_2$, CaO.

Properties of Ionic Compounds

1. Physical Nature

Ionic compounds are **hard and brittle solids**.

2. High Melting and Boiling Points

They have **high melting and boiling points** due to strong electrostatic forces.

Example:
NaCl, $CaCl_2$.

3. Solubility

Ionic compounds are generally **soluble in water** and **insoluble in organic solvents** like petrol and kerosene.

4. Conduction of Electricity

Ionic compounds conduct electricity:

- in **aqueous solution**
- in **molten state**

They do **not conduct electricity in solid state** because ions cannot move freely.

Occurrence of Metals

Metals occur in nature in the form of **minerals**.

Mineral

Naturally occurring substances in the earth's crust containing metals.

Ore

A mineral from which metal can be **extracted economically**.

Example:

ORE	METAL
BAUXITE	Aluminium
HAEMATITE	Iron
CINNABAR	Mercury

Extraction of Metals (Metallurgy)

Metallurgy:
The process of **extracting metals from their ores and refining them to obtain pure metal** is called metallurgy.

Main Steps:

1. Concentration of ore
2. Conversion of ore into oxide
3. Reduction of oxide to metal
4. Refining of metal

1. Concentration of Ore

In this step, **impurities (gangue)** such as sand, clay and stones are **removed from the ore**.

This increases the **percentage of metal in the ore**.

Common methods

- Hydraulic washing
- Froth flotation
- Magnetic separation

Example: Removing sand from iron ore by hydraulic washing.

2. Conversion of Ore into Oxide

In this step, the **concentrated ore is converted into metal oxide** because **metal oxides are easier to reduce**.

(a) Roasting

Heating of sulphide ores in the presence of oxygen.

Example:

$$2ZnS + 3O_2 \rightarrow 2ZnO + 2SO_2$$

(Zinc sulphide → Zinc oxide)

(b) Calcination

Heating of carbonate ores in limited or no oxygen.

Example:

$$ZnCO_3 \rightarrow ZnO + CO_2$$

(Zinc carbonate → Zinc oxide)

3. Reduction of Oxide

In this step, metal oxide is reduced to metal by removing oxygen.

Using carbon (for moderately reactive metals)

Example:

$$ZnO + C \rightarrow Zn + CO$$

Here carbon removes oxygen from zinc oxide to produce zinc metal.

4. Refining of Metal

The metal obtained after reduction is **not completely pure**.
It contains some impurities, so it is **purified by refining**.

Common method: Electrolytic Refining

Example: **Copper refining**

- **Anode:** Impure copper
- **Cathode:** Pure copper
- **Electrolyte:** Copper sulphate solution

Reaction:

At Anode:

$$Cu \rightarrow Cu^{2+} + 2e^{-}$$

At Cathode:

$$Cu^{2+} + 2e^{-} \rightarrow Cu$$

Pure copper gets deposited on the **cathode**.

Corrosion

Corrosion is the **slow destruction of metals due to reaction with air, water or chemicals**.

Example: Rusting of iron.

Rust is **hydrated iron(III) oxide**.

4Fe + 3O□ + xH□O → 2Fe□O□·xH□O

Examples of Corrosion
Silver → black coating (silver sulphide)

Copper → green coating (basic copper carbonate)

Iron → rust

Prevention of Corrosion

Corrosion can be prevented by:

1. Painting
2. Oiling and greasing
3. Galvanization (coating iron with zinc)
4. Electroplating
5. Alloying

Alloys

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

Examples:

ALLOY	COMPOSITION	USE
BRASS	Copper + Zinc	Utensils
BRONZE	Copper + Tin	Statues
STEEL	Iron + Carbon	Construction
SOLDER	Lead + Tin	Joining wires

Alloys improve properties such as **strength, hardness and corrosion resistance**.