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Metals and Non-metals Class 10 Notes | CBSE Chapter 3 Science

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Metals and Non-metals Class 10 Notes | CBSE

Chapter 3 Science

Metals and Non-metals is Chapter 3 of CBSE Class 10 Science. This chapter explores the physical and chemical properties of metals and non-metals, how metals react with various substances, and how metals are extracted from their ores through metallurgy. It also covers corrosion and its prevention.

This is an important chapter for board exams — expect 5–8 marks. Reactivity series, extraction of metals, and ionic bond formation are the most frequently tested topics.

Key Concepts

1. Physical Properties of Metals and Non-metals

Property	Metals	Non-metals
State at room temperature	Solid (except mercury — liquid)	Solid, liquid (bromine), or gas (oxygen, nitrogen)
Lustre (shine)	Shiny (lustrous)	Dull (non-lustrous), except iodine and diamond
Malleability (can be beaten into sheets)	Yes — gold and silver are most malleable	No — they are brittle
Ductility (can be drawn into wires)	Yes — gold is most ductile	No
Conductivity (heat and electricity)	Good conductors — silver is the best	Poor conductors (insulators), except graphite
Sonorous (produce sound when struck)	Yes — used in bells	No
Melting and boiling point	Generally high (except gallium, caesium)	Generally low
Density	Generally high	Generally low
Hardness	Generally hard (except sodium, potassium — cut with knife)	Generally soft (except diamond — hardest natural substance)

Exceptions to remember:

- Mercury is a metal but is liquid at room temperature
- Sodium and potassium are soft metals — can be cut with a knife
- Diamond (a non-metal, form of carbon) is the hardest natural substance
- Graphite (non-metal) is a good conductor of electricity
- Iodine (non-metal) has a lustrous (shiny) appearance

2. Chemical Properties of Metals

a) Reaction with Oxygen (Burning in Air)

Metals react with oxygen to form **metal oxides**, which are generally **basic** in nature.

Metal + Oxygen → Metal Oxide

- $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$ (magnesium burns with a bright white flame)
- $4\text{Al} + 3\text{O}_2 \rightarrow 2\text{Al}_2\text{O}_3$
- $2\text{Cu} + \text{O}_2 \rightarrow 2\text{CuO}$ (copper forms black copper oxide on heating)

Amphoteric oxides: Some metal oxides react with both acids and bases — e.g., Al_2O_3 and ZnO .

- $\text{Al}_2\text{O}_3 + 6\text{HCl} \rightarrow 2\text{AlCl}_3 + 3\text{H}_2\text{O}$ (acts as base)
- $\text{Al}_2\text{O}_3 + 2\text{NaOH} \rightarrow 2\text{NaAlO}_2 + \text{H}_2\text{O}$ (acts as acid)

b) Reaction with Water

Metal	Reaction with water	Product
Sodium (Na), Potassium (K)	React vigorously with cold water — catch fire	Metal hydroxide + H_2
Calcium (Ca)	Reacts with cold water — less vigorous	Ca(OH)_2 + H_2 (floats)
Magnesium (Mg)	Reacts slowly with cold water; fast with hot water/steam	Mg(OH)_2 or MgO + H_2
Aluminium (Al), Iron (Fe), Zinc (Zn)	React only with steam (not cold/hot water)	Metal oxide + H_2
Copper (Cu), Silver (Ag), Gold (Au)	Do NOT react with water at all	—

Na and K are stored in kerosene oil to prevent reaction with moisture in air.

c) Reaction with Dilute Acids

Metal + Dilute acid → Salt + Hydrogen gas

- $\text{Zn} + \text{H}_2\text{SO}_4 \rightarrow \text{ZnSO}_4 + \text{H}_2\uparrow$
- $\text{Fe} + 2\text{HCl} \rightarrow \text{FeCl}_2 + \text{H}_2\uparrow$
- $2\text{Na} + 2\text{HCl} \rightarrow 2\text{NaCl} + \text{H}_2\uparrow$

Important: Copper (Cu), Silver (Ag), and Gold (Au) do NOT react with dilute acids — they are below hydrogen in the reactivity series.

Note: Hydrogen gas is tested by bringing a burning splint near it — it burns with a “pop” sound.

d) Reaction with Salt Solutions (Displacement Reactions)

A more reactive metal displaces a less reactive metal from its salt solution.

- $\text{Fe} + \text{CuSO}_4 \rightarrow \text{FeSO}_4 + \text{Cu}$ (iron displaces copper — blue solution turns green)
 - $\text{Zn} + \text{CuSO}_4 \rightarrow \text{ZnSO}_4 + \text{Cu}$
 - $\text{Cu} + 2\text{AgNO}_3 \rightarrow \text{Cu}(\text{NO}_3)_2 + 2\text{Ag}$ (copper displaces silver)
 - $\text{Cu} + \text{FeSO}_4 \rightarrow \text{No reaction}$ (copper is less reactive than iron)
-

3. The Reactivity Series

The **reactivity series** is an arrangement of metals in decreasing order of their reactivity:

K > Na > Ca > Mg > Al > Zn > Fe > Ni > Sn > Pb > (H) > Cu > Hg > Ag > Au > Pt

Memory trick: Kindly Note Calcium Magnesium Aluminium Zinc Iron Nickel Tin Lead (Hydrogen) Copper Mercury Silver Gold Platinum

Key points:

- Metals above hydrogen in the series displace hydrogen from dilute acids
 - Metals below hydrogen (Cu, Ag, Au, Pt) do NOT react with dilute acids
 - A metal can displace any metal below it in the series from its salt solution
 - More reactive metals are harder to extract from their ores
-

4. Chemical Properties of Non-metals

a) Reaction with Oxygen

Non-metals react with oxygen to form **non-metal oxides**, which are **acidic** in nature.

- $\text{C} + \text{O}_2 \rightarrow \text{CO}_2$ (acidic — turns lime water milky)
- $\text{S} + \text{O}_2 \rightarrow \text{SO}_2$ (acidic — causes acid rain)

b) Reaction with Water

Non-metals generally do NOT react with water.

c) Reaction with Dilute Acids

Non-metals do NOT react with dilute acids (they cannot displace hydrogen).

5. Ionic Bonds — How Metals and Non-metals React with Each Other

Metals react with non-metals by **transferring electrons** to form **ionic compounds** (also called electrovalent compounds).

- Metals **lose electrons** → become positively charged ions (cations)
- Non-metals **gain electrons** → become negatively charged ions (anions)
- The electrostatic attraction between cations and anions holds the compound together — this is an **ionic bond**

Example — Formation of NaCl (sodium chloride):

- Na (2,8,1) loses 1 electron → Na^+ (2,8) — achieves noble gas configuration
- Cl (2,8,7) gains 1 electron → Cl^- (2,8,8) — achieves noble gas configuration
- Na^+ and Cl^- are held together by strong electrostatic force

Example — Formation of MgCl_2 :

- Mg (2,8,2) loses 2 electrons → Mg^{2+} (2,8)
- Each Cl (2,8,7) gains 1 electron → Cl^- (2,8,8)
- One Mg^{2+} needs two Cl^- ions → MgCl_2

Properties of Ionic Compounds

Property	Explanation
High melting and boiling points	Strong electrostatic force between ions requires a lot of energy to break
Solid at room temperature	Ions are tightly packed in a regular crystal lattice
Conduct electricity when dissolved in water or molten	Ions become free to move and carry charge
Do NOT conduct electricity in solid state	Ions are fixed in position — cannot move
Soluble in water	Water molecules can separate the ions
Insoluble in organic solvents (petrol, kerosene)	These solvents cannot break ionic bonds

6. Occurrence and Extraction of Metals

How Metals Occur in Nature

- **Native/Free state:** Unreactive metals like gold, silver, and platinum are found as pure metals in nature
- **Combined state:** Most metals are found as compounds (ores) — usually as oxides, sulphides, or carbonates

Mineral vs Ore: A mineral is any naturally occurring compound of a metal. An ore is a mineral from which the metal can be profitably extracted.

Common Ores

Metal	Ore	Formula
Aluminium	Bauxite	$\text{Al}_2\text{O}_3 \cdot 2\text{H}_2\text{O}$
Iron	Haematite	Fe_2O_3
Zinc	Zinc blende (Sphalerite)	ZnS
Copper	Copper pyrites	CuFeS_2
Sodium	Rock salt	NaCl

Steps of Extraction

Step 1 — Enrichment of ore (Concentration): Remove impurities (gangue/matrix) from the ore.

Step 2 — Conversion to metal oxide:

- Carbonate ores → **Calcination** (strong heating in absence of air): $\text{ZnCO}_3 \rightarrow \text{ZnO} + \text{CO}_2$
- Sulphide ores → **Roasting** (heating in excess air): $2\text{ZnS} + 3\text{O}_2 \rightarrow 2\text{ZnO} + 2\text{SO}_2$

Step 3 — Reduction to metal: Method depends on the metal's position in the reactivity series.

Position in Reactivity Series	Metals	Extraction Method
Top (highly reactive)	K, Na, Ca, Mg, Al	Electrolytic reduction (electrolysis of molten ore)
Middle (moderately reactive)	Zn, Fe, Ni, Sn, Pb	Reduction with carbon (coke): $\text{ZnO} + \text{C} \rightarrow \text{Zn} + \text{CO}$
Bottom (least reactive)	Cu, Hg, Ag, Au	Heating alone or self-reduction: $2\text{HgS} + 3\text{O}_2 \rightarrow 2\text{HgO} + 2\text{SO}_2$, then $2\text{HgO} \rightarrow 2\text{Hg} + \text{O}_2$

Thermite reaction: Highly reactive metals like aluminium can reduce less reactive metal oxides. This is used in welding railway tracks:



Step 4 — Refining (Purification)

The most common method is **electrolytic refining**:

- **Anode:** Impure metal
 - **Cathode:** Thin strip of pure metal
 - **Electrolyte:** Solution of metal salt
 - On passing current, pure metal deposits on the cathode; impurities settle as **anode mud**
-

7. Corrosion

Corrosion is the slow destruction of a metal by the action of air, moisture, and chemicals on its surface.

- **Rusting of iron:** Iron reacts with oxygen and moisture to form hydrated iron(III) oxide (rust): $4\text{Fe} + 3\text{O}_2 + x\text{H}_2\text{O} \rightarrow 2\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}$
- **Tarnishing of silver:** Silver reacts with H_2S in air to form black silver sulphide (Ag_2S)
- **Green coating on copper:** Copper reacts with CO_2 and moisture to form green copper carbonate [$\text{CuCO}_3 \cdot \text{Cu}(\text{OH})_2$]

Prevention of Corrosion

- **Painting / Oiling / Greasing:** Coating prevents air and moisture from reaching the metal
 - **Galvanisation:** Coating iron with a layer of zinc — the most common method for iron
 - **Electroplating:** Coating with another metal (e.g., tin, chromium) using electrolysis
 - **Alloying:** Mixing with other metals to make it corrosion-resistant (e.g., stainless steel = iron + chromium + nickel + carbon)
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8. Alloys

An **alloy** is a homogeneous mixture of a metal with one or more other metals or non-metals.

Alloy	Composition	Use
Brass	Copper + Zinc	Utensils, decorations
Bronze	Copper + Tin	Statues, medals, coins
Stainless steel	Iron + Chromium + Nickel + Carbon	Utensils, surgical instruments
Solder	Lead + Tin	Welding electrical wires
Amalgam	Mercury + other metal	Dental fillings

Why make alloys? Pure metals are often too soft. Alloying increases hardness, strength, and corrosion resistance. For example, pure gold (24 carat) is very soft — it is alloyed with silver or copper to make jewellery (22 carat = 22 parts gold, 2 parts other metal).

Important Definitions

Term	Definition
Malleability	Property of metals to be beaten into thin sheets
Ductility	Property of metals to be drawn into thin wires
Sonorous	Property of metals to produce a ringing sound when struck
Reactivity series	Arrangement of metals in decreasing order of their chemical reactivity
Ionic bond	Bond formed by transfer of electrons from a metal to a non-metal
Mineral	Naturally occurring compound of a metal found in the earth's crust
Ore	Mineral from which a metal can be extracted profitably
Gangue	Impurities (sand, clay, rock) present in an ore
Calcination	Heating an ore strongly in the absence of air (for carbonate ores)
Roasting	Heating an ore in excess of air (for sulphide ores)
Thermite reaction	Reaction between aluminium and iron oxide producing molten iron + intense heat
Corrosion	Slow destruction of a metal surface by air, moisture, and chemicals
Galvanisation	Coating iron with zinc to prevent rusting
Alloy	Homogeneous mixture of a metal with other metals or non-metals
Amalgam	Alloy of mercury with another metal

Solved Examples (NCERT-Based)

Example 1

Why is sodium stored in kerosene oil?

Answer: Sodium is a highly reactive metal — it reacts vigorously with oxygen in air and even with moisture (water vapour). When exposed to air, it catches fire. Kerosene oil does not react with sodium and prevents contact with air and moisture, keeping it safe.

Example 2

An iron nail is dipped in copper sulphate solution. What changes do you observe? Why?

Answer: The blue colour of CuSO_4 solution fades and turns green (FeSO_4 forms). A reddish-brown deposit of copper appears on the iron nail. This happens because iron is more reactive than copper (higher in the reactivity series), so it displaces copper: $\text{Fe} + \text{CuSO}_4 \rightarrow \text{FeSO}_4 + \text{Cu}$

Example 3

Why does aluminium not corrode even though it is a reactive metal?

Answer: Aluminium forms a thin, tough layer of aluminium oxide (Al_2O_3) on its surface when exposed to air. This oxide layer is very stable and prevents further reaction of the metal with air and moisture. So even though aluminium is reactive, its oxide layer protects it from corrosion.

Example 4

Explain why ionic compounds have high melting points but do not conduct electricity in the solid state.

Answer: Ionic compounds consist of positive and negative ions held together by very strong electrostatic forces of attraction. Breaking these bonds requires a large amount of energy, hence the high melting points. In the solid state, ions are fixed in a rigid crystal lattice and cannot move — so they cannot carry electric charge. When melted or dissolved in water, the ions become free to move and can conduct electricity.

Important Questions for Board Exams

1-Mark Questions

1. Name a metal that is liquid at room temperature.
2. What is the nature of metal oxides — acidic or basic?
3. Why is gold found in the free state in nature?
4. Name the ore of aluminium.
5. What is an alloy? Give one example.

2-Mark Questions

1. Write the electron dot structure for the formation of MgCl_2 .
2. What is the difference between calcination and roasting?
3. Why are food cans coated with tin and not zinc?
4. Name two metals that can displace hydrogen from dilute acids and two that cannot.
5. What is galvanisation? Why is it done?

3-Mark Questions

1. Explain the extraction of metals of medium reactivity from their sulphide ores with an example.

2. What are amphoteric oxides? Give two examples with reactions.
3. What are ionic compounds? List four properties of ionic compounds.
4. Describe the electrolytic refining of copper with a labelled diagram.
5. Arrange the following metals in order of their reactivity: Al, Cu, Fe, Mg, Zn. Justify with one reaction.

5-Mark Questions

1. Describe the steps involved in the extraction of metals from their ores. Explain how the method of extraction varies for metals of different reactivity.
 2. What is corrosion? Explain the rusting of iron. Describe four methods to prevent corrosion.
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Quick Revision Points

- Metals: lustrous, malleable, ductile, sonorous, good conductors. Exceptions: mercury (liquid), Na/K (soft)
 - Non-metals: dull, brittle, poor conductors. Exceptions: diamond (hard), graphite (conducts electricity), iodine (lustrous)
 - $\text{Metal} + \text{O}_2 \rightarrow \text{Metal oxide (basic)}$; $\text{Non-metal} + \text{O}_2 \rightarrow \text{Non-metal oxide (acidic)}$
 - Amphoteric oxides: Al_2O_3 and ZnO react with both acids and bases
 - Reactivity series: $\text{K} > \text{Na} > \text{Ca} > \text{Mg} > \text{Al} > \text{Zn} > \text{Fe} > (\text{H}) > \text{Cu} > \text{Ag} > \text{Au}$
 - More reactive metal displaces less reactive metal from salt solution
 - Ionic bond = transfer of electrons; ionic compounds have high MP, conduct when molten/dissolved
 - Ore enrichment $\rightarrow$ Convert to oxide (calcination/roasting) $\rightarrow$ Reduce to metal $\rightarrow$ Refine
 - Highly reactive metals: electrolysis | Medium: carbon reduction | Low: heating alone
 - Thermite reaction: $\text{Fe}_2\text{O}_3 + 2\text{Al} \rightarrow 2\text{Fe} + \text{Al}_2\text{O}_3$ (used in railway welding)
 - Corrosion prevention: painting, oiling, galvanisation (zinc coating), electroplating, alloying
 - Stainless steel = $\text{Fe} + \text{Cr} + \text{Ni} + \text{C}$; Brass = $\text{Cu} + \text{Zn}$; Bronze = $\text{Cu} + \text{Sn}$
 - Pure gold (24K) is too soft — alloyed for jewellery (22K)
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Previous Chapter: Chapter 2 — Acids, Bases and Salts

Next Chapter: Chapter 4 — Carbon and its Compounds



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