Metals and Non-metals

Comprehensive Study Notes for RBSE Class 10 Science Chapter 3

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1. CLASSIFICATION OF ELEMENTS

Definition

Elements can be classified based on their properties into two main categories: **Metals** and **Non-metals**.

Metals vs Non-metals: Overview

Metals	**Non-metals**
Malleable (can be hammered into thin sheets)	Brittle (break into pieces when hammered)
Ductile (can be drawn into thin wires)	Not ductile
Lustrous (shiny appearance)	Dull appearance (except bromine)
Good conductors of heat and electricity	Poor conductors (except graphite for electricity)
Sonorous (produce sound when struck)	Not sonorous
Generally solid at room temperature	Solid, liquid, or gas at room temperature

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2. PHYSICAL PROPERTIES OF METALS

2.1 Malleability

Definition

Ability of a metal to be beaten or hammered into thin sheets without breaking.

Examples

- **Gold:** Most malleable metal
- **Silver, Copper, Aluminum:** Highly malleable
- **Tungsten, Molybdenum:** Less malleable

Application

- Gold is used for making jewelry and decorative items
- Aluminum foil for wrapping and covering

2.2 Ductility

Definition

Ability of a metal to be drawn into thin wires without breaking.

Examples

- **Gold:** Most ductile metal
- **Silver, Copper, Aluminum:** Highly ductile
- **1 gram of gold can be drawn into wire of length:** 2 km

Application

- Copper wires used in electrical installations
- Aluminum wires for various purposes

2.3 Lustre (Lusciousness)

Definition

Natural shine or brightness of a metal surface.

Characteristics

- Metals have a characteristic shine/glow
- **Polished metals:** Reflect light better
- **Oxidized metals:** Lose shine due to oxide layer

Examples

- Silver, gold, copper, aluminum show high lustre
- Iron oxidizes to form rust, losing shine

2.4 Conductivity of Heat and Electricity

Heat Conductivity

- Metals are excellent conductors of heat
- **Best conductor:** Silver
- **Second best:** Copper
- **Application:** Used in cooking vessels, radiators, heat exchangers

Electrical Conductivity

- Metals are excellent conductors of electricity
- **Best conductor:** Silver
- **Second best:** Copper
- **Application:** Wires for electrical transmission, circuits

2.5 Sonority

Definition

Ability to produce sound when struck or vibrated.

Characteristics

- Metals produce a ringing/bell-like sound
- **More sonorous metals:** Iron, Steel, Aluminum
- **Applications:** Bells, gongs, cymbals

2.6 State at Room Temperature

Observation

- **Almost all metals are solids** at room temperature (25°C)
- **Exception:** Mercury is liquid at room temperature

Other Metals with Low Melting Points

- **Gallium:** Melts at 30°C (can melt in hot water)
- **Cesium:** Melts at 28°C

2.7 Density and Melting/Boiling Points

Density

- Metals generally have high density
- **Densest metal:** Osmium
- **Lightest metal:** Lithium
- **Common metals:** Iron, copper, aluminum have moderate density

Melting and Boiling Points

- **Metals generally have high melting and boiling points**
- **Exceptions:** Gallium, Cesium (very low melting points)
- **Examples:**
    - Copper: Melting point = 1084°C
    - Iron: Melting point = 1538°C
    - Aluminum: Melting point = 660°C

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3. PHYSICAL PROPERTIES OF NON-METALS

3.1 Brittleness

Definition

Tendency to break into pieces when struck or stressed.

Examples

- Carbon (diamond, coal): Breaks when hammered
- Sulfur: Shatters into pieces
- Phosphorus: Very brittle

Difference from Malleability

- **Metals:** Can be hammered into sheets (malleable)
- **Non-metals:** Break or shatter (brittle)

3.2 Non-ductility

Definition

Non-metals cannot be drawn into thin wires.

Reason

- Lack flexible structure
- Break easily under stretching stress

Exception

- **Carbon fibers:** Can be made into very fine threads due to special structure

3.3 Appearance

Lustre

- **Non-metals are generally dull** (no shine)
- **Exception:** Diamond (non-metal) is very lustrous due to special crystal structure

Bromine

- Reddish-brown liquid at room temperature
- Has some shine due to liquid nature
- Exception to dull appearance rule

3.4 Poor Conductivity

Heat Conductivity

- Non-metals are **poor conductors of heat**
- **Exception:** Diamond conducts heat very well
- **Application:** Used as insulators (wood, plastic, rubber)

Electrical Conductivity

- Non-metals are **poor conductors of electricity**
- **Exception:** Graphite (carbon allotrope) conducts electricity
- **Application:** Graphite used for:
  - Electrodes in cells
  - Dry cells
  - Brushes in electric motors

3.5 Non-sonorous Nature

Definition

Non-metals do not produce ringing sound when struck.

Observation

- Produce dull thud sound
- No metallic ringing

3.6 States at Room Temperature

Variety of States

- **Solid non-metals:** Carbon, sulfur, phosphorus, iodine
- **Gaseous non-metals:** Oxygen, nitrogen, hydrogen, chlorine, fluorine
- **Liquid non-metal:** Bromine

Melting and Boiling Points

- **Generally low** compared to metals
- **Exceptions:**
  - Diamond: Very high melting point (3823°C)
  - Silicon: High melting point (1414°C)

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4. CHEMICAL PROPERTIES OF METALS

4.1 Metals are Electropositive

Definition

Metals have tendency to lose electrons from their valence shell to form positive ions (cations).

Electron Loss

- Metals lose 1, 2, or 3 electrons (rarely more)
- Form stable electron configuration
- Become positively charged ions

Examples

Sodium loses 1 electron:

$$\mathrm{Na} \rightarrow \mathrm{Na}^{+} + e^{-}$$

Magnesium loses 2 electrons:

$$\mathrm{Mg} \rightarrow \mathrm{Mg}^{2+} + 2e^{-}$$

Aluminum loses 3 electrons:

$$\mathrm{Al} \rightarrow \mathrm{Al}^{3+} + 3e^{-}$$

4.2 Reaction of Metals with Oxygen

General Reaction

$$\text{Metal} + \mathrm{O_2} \rightarrow \text{Metal oxide}$$

Formation of Oxides

- **Metal oxides are basic in nature**
- Some metal oxides are **amphoteric** (act as both acid and base)

Characteristics

- Metals form **basic oxides** when burned in oxygen
- These oxides dissolve in water to form bases
- Some metal oxides are **amphoteric:** Al_2O_3, ZnO, PbO

Examples

Sodium burns with yellow flame:

$$2\mathrm{Na} + \mathrm{O_2} \rightarrow \mathrm{Na_2O_2}$$

(Sodium peroxide)

Magnesium burns with white flame:

$$2\mathrm{Mg} + \mathrm{O_2} \rightarrow 2\mathrm{MgO}$$

Iron burns to form iron oxide:

$$3\mathrm{Fe} + 2\mathrm{O_2} \rightarrow \mathrm{Fe_3O_4}$$

Copper turns black due to oxide:

$$2\mathrm{Cu} + \mathrm{O_2} \rightarrow 2\mathrm{CuO}$$

4.3 Reaction of Metals with Water

Metals React Differently with Water/Steam

Reaction with Cold Water

Very reactive metals (Group 1):

$$2\text{Na} + 2\text{H}_2\text{O} \rightarrow 2\text{NaOH} + \text{H}_2(g)$$

Characteristics:

- Sodium and potassium react vigorously with cold water
- Form metal hydroxide and hydrogen gas
- **Highly exothermic:** Hydrogen catches fire
- **Sodium:** Moves on water surface, melts to ball due to heat

Less reactive metals:

$$2\text{Ca} + 2\text{H}_2\text{O} \rightarrow 2\text{Ca(OH)}_2 + \text{H}_2(g)$$

Reaction with Hot Water

Magnesium reacts with hot water:

$$\text{Mg} + 2\text{H}_2\text{O(hot)} \rightarrow \text{Mg(OH)}_2 + \text{H}_2(g)$$

Reaction with Steam

Aluminum with steam:

$$2\text{Al} + 3\text{H}_2\text{O(steam)} \rightarrow \text{Al}_2\text{O}_3 + 3\text{H}_2(g)$$

Iron with steam:

$$3\text{Fe} + 4\text{H}_2\text{O(steam)} \rightarrow \text{Fe}_3\text{O}_4 + 4\text{H}_2(g)$$

Zinc with steam:

$$\text{Zn} + \text{H}_2\text{O(steam)} \rightarrow \text{ZnO} + \text{H}_2(g)$$

Metals that Do NOT React with Water

- **Lead, copper, silver, gold:** Do not react with water or steam
- **Reason:** Too unreactive/low in reactivity series

4.4 Reaction of Metals with Dilute Acids

General Reaction

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

Characteristics

- Only **more reactive metals** react with dilute acids
- **Less reactive metals** (copper, silver, gold) do NOT react
- Hydrogen gas evolved can be tested with **burning splint** (produces pop sound)

Examples

Zinc with dilute sulfuric acid:

$$\mathrm{Zn} + \mathrm{H_2SO_4} \rightarrow \mathrm{ZnSO_4} + \mathrm{H_2}(g)$$

Magnesium with dilute hydrochloric acid:

$$\mathrm{Mg} + 2\mathrm{HCl} \rightarrow \mathrm{MgCl_2} + \mathrm{H_2}(g)$$

Iron with dilute sulfuric acid:

$$\mathrm{Fe} + \mathrm{H_2SO_4} \rightarrow \mathrm{FeSO_4} + \mathrm{H_2}(g)$$

Test for Hydrogen Gas

- Bring a burning splint near the gas outlet
- Hydrogen burns with a **characteristic "pop" sound**
- This is the **definitive test for $\mathrm{H_2}$ gas**

4.5 Reaction of Metals with Salt Solutions

General Reaction (Displacement)

$$\mathrm{Metal_1} + \mathrm{Salt\ of\ Metal_2} \rightarrow \mathrm{Salt\ of\ Metal_1} + \mathrm{Metal_2}$$

Condition

- **$\mathrm{Metal_1}$ must be more reactive than $\mathrm{Metal_2}$** (higher in reactivity series)
- Only then can displacement occur

Examples

Copper displaces silver (Cu more reactive):

$$\mathrm{Cu} + 2\mathrm{AgNO_3} \rightarrow \mathrm{Cu(NO_3)_2} + 2\mathrm{Ag}$$

Zinc displaces copper (Zn more reactive):

$$\mathrm{Zn} + \mathrm{CuSO_4} \rightarrow \mathrm{ZnSO_4} + \mathrm{Cu}$$

Iron displaces copper (Fe more reactive):

$$\mathrm{Fe} + \mathrm{CuSO_4} \rightarrow \mathrm{FeSO_4} + \mathrm{Cu}$$

Aluminum displaces iron (Al more reactive):

$$2\mathrm{Al} + 3\mathrm{FeSO_4} \rightarrow \mathrm{Al_2(SO_4)_3} + 3\mathrm{Fe}$$

4.6 Metal + Chlorine Reaction

General Reaction

$$\text{Metal} + \text{Cl}_2 \rightarrow \text{Metal chloride (ionic)}$$

Products

- **Ionic chlorides** are formed
- These are salts with strong electrostatic bonds

Examples

Sodium with chlorine:

$$2\text{Na} + \text{Cl}_2 \rightarrow 2\text{NaCl}$$

Magnesium with chlorine:

$$\text{Mg} + \text{Cl}_2 \rightarrow \text{MgCl}_2$$

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5. CHEMICAL PROPERTIES OF NON-METALS

5.1 Non-metals are Electronegative

Definition

Non-metals have tendency to gain electrons in their valence shell to form negative ions (anions).

Electron Gain

- Non-metals gain 1, 2, or 3 electrons
- Form stable electron configuration
- Become negatively charged ions

Examples

Chlorine gains 1 electron:

$$\text{Cl} + e^- \rightarrow \text{Cl}^-$$

Oxygen gains 2 electrons:

$$\text{O} + 2e^- \rightarrow \text{O}^{2-}$$

Nitrogen gains 3 electrons:

$$\text{N} + 3e^- \rightarrow \text{N}^{3-}$$

5.2 Reaction of Non-metals with Oxygen

General Reaction

$$\text{Non-metal} + O_2 \rightarrow \text{Non-metal oxide}$$

Types of Oxides Formed

- **Acidic oxides:** Most common
- **Neutral oxides:** Some non-metals form neutral oxides

Characteristics

- Non-metal oxides are **acidic or neutral**
- Dissolve in water to form acids
- Turn litmus blue to red (acidic)

Examples

Carbon burns to form carbon dioxide:

$$C + O_2 \rightarrow CO_2$$

(Acidic oxide)

Sulfur burns to form sulfur dioxide:

$$S + O_2 \rightarrow SO_2$$

(Acidic oxide)

Nitrogen forms nitrogen dioxide:

$$N_2 + 2O_2 \rightarrow 2NO_2$$

(Acidic oxide)

Carbon can form carbon monoxide (neutral oxide):

$$2C + O_2 \rightarrow 2CO$$

(Neutral oxide)

5.3 Reaction of Non-metals with Water

General Statement

Non-metals do NOT react with water or steam to produce hydrogen gas.

Reason

- Non-metals are electronegative
- Cannot displace hydrogen from water
- Some non-metals dissolve in water without reacting chemically

Exception

- Some non-metals dissolve to form acids
- **Chlorine reacts with water:**

$$Cl_2 + H_2O \rightarrow HCl + HClO$$

(Forms hypochlorous acid and hydrochloric acid)

5.4 Reaction of Non-metals with Dilute Acids

General Statement

Non-metals do NOT react with dilute acids to produce hydrogen gas.

Reason

- Non-metals cannot displace hydrogen from acids
- Already more electronegative than hydrogen

Exception

- Some non-metals may react but do not produce H_2 gas
- **Example:** Carbon reacts with hot concentrated sulfuric acid (redox reaction, not typical acid-base)

5.5 Reaction of Non-metals with Salt Solutions

General Statement

Non-metals do NOT displace hydrogen or other metals from salt solutions.

Reason

- Non-metals gain electrons, not lose them
- Cannot replace positive ions from solutions

5.6 Non-metals with Chlorine

General Reaction

$$\text{Non-metal} + Cl_2 \rightarrow \text{Non-metal chloride (covalent)}$$

Products

- **Covalent chlorides** are formed
- Molecules with covalent bonds
- Generally volatile (can evaporate)

Examples

Hydrogen with chlorine:

$$H_2 + Cl_2 \rightarrow 2HCl$$

(Covalent, volatile)

Phosphorus with chlorine:

$$2P + 5Cl_2 \rightarrow 2PCl_5$$

(Covalent)

Sulfur with chlorine:

$$S + Cl_2 \rightarrow SCl_2$$

(Covalent)

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6. REACTIVITY SERIES OF METALS

6.1 Definition and Purpose

Definition

Reactivity series is a list of metals arranged in decreasing order of their reactivity.

Purpose

- Predicts which metal will displace another
- Helps in metal extraction from ores
- Understands corrosion and prevention
- Predicts displacement reactions

6.2 Activity Series (Reactivity Series)

Complete Series (from most to least reactive)

Most Reactive:

1. **K** (Potassium) - Extremely reactive
2. **Na** (Sodium) - Very reactive
3. **Ca** (Calcium) - Very reactive
4. **Mg** (Magnesium) - Quite reactive
5. **Al** (Aluminum) - Moderately reactive
6. **Zn** (Zinc) - Moderately reactive
7. **Fe** (Iron) - Moderately reactive
8. **Ni** (Nickel) - Less reactive
9. **Sn** (Tin) - Less reactive
10. **Pb** (Lead) - Less reactive
11. **H** (Hydrogen) - Reference point
12. **Cu** (Copper) - Low reactivity
13. **Hg** (Mercury) - Low reactivity
14. **Ag** (Silver) - Very low reactivity
15. **Au** (Gold) - Least reactive

Least Reactive:

6.3 Hydrogen Position in Series

Why H is Included?

- Reference point for metal reactivity
- Metals above H displace hydrogen from dilute acids
- Metals below H do NOT react with dilute acids

Displacement from Acids

- **More reactive than H:** React with dilute acids
  - $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2\uparrow$
  - $Mg + 2HCl \rightarrow MgCl_2 + H_2\uparrow$
- **Less reactive than H:** Do NOT react
  - Cu does not react with dilute acids
  - Au does not react with dilute acids

6.4 Application: Metal Displacement Reactions

Rule

A more reactive metal can displace a less reactive metal from its salt solution.

Example 1: Zn and $CuSO_4$

- Zn is more reactive than Cu

$$\mathrm{Zn} + \mathrm{CuSO_4} \rightarrow \mathrm{ZnSO_4} + \mathrm{Cu}$$

- **Occurs:** Yes (displacement happens)
- **Cu becomes:** Red solid

Example 2: Cu and $ZnSO_4$

- Cu is less reactive than Zn

$$\mathrm{Cu} + \mathrm{ZnSO_4} \rightarrow \mathrm{CuSO_4} + \mathrm{Zn}$$

- **Occurs:** No (no displacement)
- Reason: Cu cannot displace Zn (less reactive)

Example 3: Fe and $CuSO_4$

- Fe is more reactive than Cu

$$\mathrm{Fe} + \mathrm{CuSO_4} \rightarrow \mathrm{FeSO_4} + \mathrm{Cu}$$

- **Occurs:** Yes (displacement happens)

Example 4: Al and $FeSO_4$

- Al is more reactive than Fe

$$2\mathrm{Al} + 3\mathrm{FeSO_4} \rightarrow \mathrm{Al_2(SO_4)_3} + 3\mathrm{Fe}$$

- **Occurs:** Yes (displacement happens)

6.5 Corrosion and Reactivity

What is Corrosion?

Gradual wearing away or deterioration of metal due to chemical reaction with surrounding substances.

Metals that Corrode Easily

- **Highly reactive metals:** K, Na, Ca
- **Moderately reactive:** Mg, Al, Zn, Fe
- **Iron especially:** Rusts due to oxidation (forms $Fe_2O_3.xH_2O$)

Metals that Do NOT Corrode Easily

- **Unreactive metals:** Cu, Ag, Au
- **Platinum, Palladium:** Extremely resistant

Prevention of Corrosion

1. **Oil/grease coating:** Prevents air and water contact
2. **Paint coating:** Isolates metal from environment
3. **Galvanization:** Zinc coating on iron
4. **Alloying:** Making stainless steel (iron + chromium + nickel)

6.6 Metal Extraction and Reactivity Series

General Principle

- **Reactive metals:** Difficult to extract (require electrolysis)
- **Less reactive metals:** Easier to extract (can use reducing agents)

Extraction Methods

For Most Reactive Metals (K, Na, Ca):

- Electrolysis of molten salt compounds
- Example: Electrolysis of molten NaCl gives Na metal

For Moderately Reactive (Al, Zn, Fe):

- Electrolysis of ore compounds or
- Reduction with carbon/carbon monoxide
- Example: Fe_2O_3 reduced by CO or C at high temperature

For Less Reactive (Cu, Ag, Au):

- Reduction with carbon or carbon monoxide
- Sometimes found as free metals in nature
- Example: Cu_2O reduced by carbon

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7. METALS AND NON-METALS REACTIONS

7.1 Reaction Between Metals and Non-metals

General Reaction

$$\text{Metal} + \text{Non-metal} \rightarrow \text{Ionic compound}$$

Nature of Reaction

- Metals lose electrons (oxidation)
- Non-metals gain electrons (reduction)
- **Transfer of electrons** from metal to non-metal
- Forms **ionic compound** with strong electrostatic attraction

Examples

Sodium and chlorine (violent reaction):

$$2\text{Na} + \text{Cl}_2 \rightarrow 2\text{NaCl}$$

- Sodium loses electrons: Na → $Na^+ + e^-$
- Chlorine gains electrons: $Cl_2 + 2e^- \rightarrow 2Cl^-$
- Ionic compound formed: NaCl

Magnesium and oxygen (combustion):

$$2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$$

- Mg loses 2 electrons: $Mg \rightarrow Mg^{2+} + 2e^-$
- O gains 2 electrons: $O_2 + 4e^- \rightarrow 2O^{2-}$
- Ionic compound: MgO

Iron and sulfur (heated together):

$$\text{Fe} + \text{S} \rightarrow \text{FeS}$$

- Fe loses 2 electrons: $Fe \rightarrow Fe^{2+} + 2e^-$
- S gains 2 electrons: $S + 2e^- \rightarrow S^{2-}$
- Ionic compound: FeS (black)

7.2 Ionic Bonds and Ionic Compounds

Ionic Bond Definition

Electrostatic attraction between oppositely charged ions (cations and anions).

Formation Process

1. Metal atom loses electrons → becomes cation (positive ion)
2. Non-metal atom gains electrons → becomes anion (negative ion)
3. Cation and anion attract each other → **ionic bond**

Properties of Ionic Compounds

- **Soluble in water:** Most ionic compounds dissolve in water
- **Conduct electricity:** In solution or molten state (due to ions)
- **High melting points:** Strong ionic bonds
- **Hard and brittle:** Rigid crystal structure
- **Formed by metals and non-metals:** Typical ionic compounds

Ions and Noble Gas Configuration

- Ions form to achieve **nearest noble gas electron configuration**
- **Metals lose electrons** to achieve stable configuration
- **Non-metals gain electrons** to achieve stable configuration

Examples

Sodium (11 electrons):

- Loses 1 electron → Na^+ (10 electrons, like Ne)
- Configuration: 2,8 → 2,8 (like Neon)

Magnesium (12 electrons):

- Loses 2 electrons → Mg^{2+} (10 electrons, like Ne)
- Configuration: 2,8,2 → 2,8 (like Neon)

Chlorine (17 electrons):

- Gains 1 electron → Cl^- (18 electrons, like Ar)
- Configuration: 2,8,7 → 2,8,8 (like Argon)

Oxygen (8 electrons):

- Gains 2 electrons → O^{2-} (10 electrons, like Ne)
- Configuration: 2,6 → 2,8 (like Neon)

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8. ALLOYS

8.1 Definition

What is an Alloy?

Alloy is a homogeneous mixture of two or more metals or a metal and a non-metal.

Key Characteristics

- **Homogeneous mixture:** Uniform composition throughout
- **Not a compound:** No fixed proportions (can vary)
- **Mixed at atomic/molecular level:** Not just physical mixing

8.2 Preparation of Alloys

Method

1. **Melt the primary metal** (main component)
2. **Dissolve other metals/non-metals** in molten state
3. **Cool to room temperature** to solidify
4. **Alloy is formed** with mixed properties

Advantages

- Properties can be customized
- Better strength and hardness
- Better corrosion resistance
- Better heat/electrical properties

8.3 Common Alloys and Their Uses

Steel

Composition: Iron + Carbon (0.1% to 1.5% carbon)

- **Properties:** Very strong, hard, durable
- **Uses:**
  - Building construction
  - Bridges and railway tracks
  - Machinery and tools
  - Automobiles
- **Special types:**
  - **Stainless steel:** Iron + Chromium + Nickel (resists corrosion)
  - **High carbon steel:** Hard, brittle (for cutting tools)
  - **Low carbon steel:** Soft, ductile (for structures)

Brass

Composition: Copper + Zinc (15-45% zinc)

- **Properties:**
  - Yellow color
  - Good corrosion resistance
  - Good malleability
- **Uses:**
  - Decorative items and jewelry
  - Utensils and vessels
  - Electrical components
  - Plumbing fixtures

Bronze

Composition: Copper + Tin (typically 10-20% tin)

- **Properties:**
  - Reddish-brown color
  - Hard and strong
  - Good corrosion resistance

- **Uses:**
  - Statues and sculptures
  - Bells and gongs
  - Bearings in machinery
  - Historical artifacts

Pewter

Composition: Tin + Copper + Lead (and sometimes silver)

- **Properties:**
  - Low melting point
  - Easy to shape
  - Silver-gray color
- **Uses:**
  - Decorative items
  - Tableware
  - Ornaments

Duralumin

Composition: Aluminum + Copper + Magnesium + Manganese

- **Properties:**
  - Very light
  - Very strong
  - Good corrosion resistance
- **Uses:**
  - Aircraft bodies and structures
  - Aerospace components
  - High-strength parts

Solder

Composition: Lead + Tin (or lead-free variants)

- **Properties:**
  - Very low melting point
  - Easy to melt and bond
- **Uses:**
  - Joining electrical circuits
  - Electronics soldering
  - Metal joining

8.4 Why Alloys are Better Than Pure Metals

Property	Pure Metals	Alloys
Strength	Moderate	Much higher
Hardness	Lower	Much higher
Corrosion Resistance	Variable	Often better
Malleability	Often high (unwanted)	Controlled
Heat Resistance	Moderate	Can be improved
Cost	Varies	Can be economical
Customization	Fixed properties	Tailored properties

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9. AMPHOTERIC OXIDES AND HYDROXIDES

9.1 Definition

Amphoteric Nature

Substances that can act as both **acids and bases**.

Amphoteric Oxides

Metal oxides that react with both acids and bases.

Amphoteric Hydroxides

Metal hydroxides that react with both acids and bases.

9.2 Common Amphoteric Compounds

Aluminum Oxide (Al_2O_3)

Reaction with acid (as base):

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

Reaction with base (as acid):

$$Al_2O_3 + 2NaOH + 3H_2O \rightarrow 2NaAl(OH)_4$$

Zinc Oxide (ZnO)

Reaction with acid (as base):

$$ZnO + 2HCl \rightarrow ZnCl_2 + H_2O$$

Reaction with base (as acid):

$$ZnO + 2NaOH \rightarrow Na_2ZnO_2 + H_2O$$

Aluminum Hydroxide $Al(OH)_3$

Reaction with acid (as base):

$$Al(OH)_3 + 3HCl \rightarrow AlCl_3 + 3H_2O$$

Reaction with base (as acid):

$$Al(OH)_3 + NaOH \rightarrow NaAl(OH)_4$$

9.3 Significance

- Explains dual nature of some metal compounds
- Important in metal extraction and purification
- Used in water treatment ($Al(OH)_3$ as clarifying agent)
- Demonstrates concept of acids and bases beyond traditional definitions

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10. OXIDATION AND REDUCTION

10.1 Definitions

Oxidation

- **Original meaning:** Loss of oxygen or gain of hydrogen
- **Modern meaning:** Loss of electrons by a substance
- **Example:** Mg loses electrons when reacting with O_2

Reduction

- **Original meaning:** Gain of oxygen or loss of hydrogen
- **Modern meaning:** Gain of electrons by a substance
- **Example:** O_2 gains electrons when reacting with Mg

10.2 Redox Reactions

Definition

Reactions involving simultaneous oxidation and reduction.

Both Occur Together

- One substance loses electrons (oxidized)
- Another substance gains electrons (reduced)
- Always occur simultaneously

Example: Magnesium Burning

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

- **Magnesium (Mg):** Loses electrons (oxidized)
  - $Mg \rightarrow Mg^{2+} + 2e^-$
- **Oxygen (O_2):** Gains electrons (reduced)
  - $O_2 + 4e^- \rightarrow 2O^{2-}$

10.3 Oxidizing and Reducing Agents

Oxidizing Agent

- Substance that causes oxidation
- Accepts electrons
- Gets reduced itself
- Example: Oxygen (O_2) in combustion

Reducing Agent

- Substance that causes reduction
- Donates electrons
- Gets oxidized itself
- Example: Carbon in: $C + O_2 \rightarrow CO_2$

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11. USES OF METALS AND NON-METALS

11.1 Uses of Important Metals

Iron

- **Uses:** Building, bridges, machinery, vehicles, tools
- **Why:** Strong, abundant, affordable
- **Forms:** Steel (more useful than pure iron)

Copper

- **Uses:** Electrical wires, plumbing, alloys, cooking vessels
- **Why:** Good conductor, corrosion resistant, ductile
- **Properties:** Red color, good malleability

Aluminum

- **Uses:** Aircraft, utensils, foil, window frames
- **Why:** Light, corrosion resistant, good conductor
- **Advantage:** Develops protective oxide layer

Gold and Silver

- **Uses:** Jewelry, currency, electronics
- **Why:** Beautiful, corrosion resistant, rare
- **Value:** Precious and expensive

Calcium

- **Uses:** Building materials (concrete, cement)
- **Compounds:** Used in construction as $CaCO_3$, $Ca(OH)_2$

11.2 Uses of Important Non-metals

Oxygen

- **Uses:** Respiration, combustion, industrial processes
- **Important for:** All aerobic life, oxidation reactions
- **Percentage:** 21% in air

Nitrogen

- **Uses:** Fertilizers (ammonia), inert atmosphere
- **Percentage:** 78% in air
- **Compounds:** NH_3, NO, NO_2

Carbon

- **Forms:** Diamond, graphite, coal, coke
- **Uses:**
  - Diamond: Jewelry, cutting tools, industrial
  - Graphite: Electrodes, pencils, lubricant
  - Coal: Fuel, carbon source
  - Coke: Reduction in metal extraction

Sulfur

- **Uses:** Matches, gunpowder, sulfuric acid
- **Compounds:** Important in industrial chemistry
- **Chemical:** Sulfur dioxide (SO_2) for bleaching

Chlorine

- **Uses:** Water purification, disinfectant, bleach
- **Compounds:** Used to make many chemicals
- **Property:** Toxic but essential in industry

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12. IMPORTANT REACTIONS SUMMARY

Metal Reactions

Metal + Oxygen:

$$\text{Metal} + O_2 \rightarrow \text{Metal oxide (basic)}$$

Metal + Cold water:

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

Metal + Dilute acid:

$$\text{Metal} + \text{Dilute acid} \rightarrow \text{Salt} + H_2 \uparrow$$

Metal displacement (from salt solution):

$$\text{More reactive metal} + \text{Salt of less reactive metal} \rightarrow \text{Salt of more reactive metal} + \text{L}$$

Metal + Chlorine:

$$\text{Metal} + Cl_2 \rightarrow \text{Metal chloride (ionic)}$$

Non-metal Reactions

Non-metal + Oxygen:

$$\text{Non-metal} + O_2 \rightarrow \text{Non-metal oxide (acidic/neutral)}$$

Metal + Non-metal:

$$\text{Metal} + \text{Non-metal} \rightarrow \text{Ionic compound}$$

Non-metal + Chlorine:

$$\text{Non-metal} + Cl_2 \rightarrow \text{Non-metal chloride (covalent)}$$

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13. KEY CONCEPTS AND DEFINITIONS

Properties

Malleability: Ability to be beaten into thin sheets without breaking

Ductility: Ability to be drawn into thin wires without breaking

Lustre: Shine or brightness of material surface

Sonority: Ability to produce ringing sound when struck

Brittleness: Tendency to break into pieces when struck

Conductivity: Ability to allow flow of heat or electricity

Chemical Concepts

Electropositive: Tendency to lose electrons (property of metals)

Electronegative: Tendency to gain electrons (property of non-metals)

Cation: Positively charged ion (formed by metal losing electrons)

Anion: Negatively charged ion (formed by non-metal gaining electrons)

Ionic bond: Electrostatic attraction between cation and anion

Ionic compound: Compound formed by ionic bond between metal and non-metal

Alloy: Homogeneous mixture of two or more metals or metal and non-metal

Reactivity series: Arrangement of metals in decreasing order of reactivity

Corrosion: Gradual deterioration of metal due to chemical reaction

Amphoteric: Substance acting as both acid and base

Oxidation: Loss of electrons by a substance

Reduction: Gain of electrons by a substance

Redox reaction: Reaction involving simultaneous oxidation and reduction

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14. EXAM-FOCUSED PRACTICE QUESTIONS

Multiple Choice Questions

1. Which metal is liquid at room temperature?

- (a) Sodium
- (b) Mercury
- (c) Gallium
- **(d) b and c both**

2. Which non-metal is a good conductor of electricity?

- (a) Sulfur
- (b) Phosphorus
- **(c) Graphite**
- (d) Chlorine

3. In the reactivity series, which metal is most reactive?

- **(a) Potassium**
- (b) Sodium
- (c) Calcium
- (d) Aluminum

4. Which metal does NOT react with dilute acids?

- (a) Zinc
- (b) Iron
- **(c) Copper**
- (d) Magnesium

5. When a metal and non-metal react, which type of bond forms?

- (a) Covalent bond
- (b) Metallic bond
- **(c) Ionic bond**
- (d) Hydrogen bond

6. Which of the following is an amphoteric oxide?

- (a) CaO
- **(b) Al_2O_3**
- (c) Na_2O
- (d) MgO

7. What is an alloy?

- (a) Pure metal
- (b) Pure non-metal

- **(c) Homogeneous mixture of metals or metal and non-metal**
- (d) Compound of metals and non-metals

8. The reactivity series helps us to predict:

- (a) Density of metals
- (b) Melting points of metals
- **(c) Displacement reactions**
- (d) Color of metals

9. Which metal can displace copper from copper sulfate solution?

- (a) Silver
- (b) Gold
- **(c) Zinc**
- (d) Mercury

10. Non-metal oxides are generally:

- (a) Basic
- **(b) Acidic or neutral**
- (c) Amphoteric
- (d) Alkaline

Short Answer Questions

11. Differentiate between metals and non-metals on the basis of their appearance.

Answer:

- **Metals:** Generally lustrous (shiny), reflect light, have characteristic metallic shine
- **Non-metals:** Generally dull, do not reflect light, no metallic shine (exception: diamond is lustrous)

12. Why does magnesium react with steam but not with cold water?

Answer: Magnesium is moderately reactive. It requires higher temperature (steam) to react. At room temperature (cold water), the reaction is too slow or doesn't occur. Heat increases molecular kinetic energy, facilitating the reaction.

13. Give reason: Copper does not react with dilute HCl but zinc reacts.

Answer: Zinc is more reactive than copper and is positioned above hydrogen in the reactivity series. Only metals more reactive than hydrogen can displace it from dilute acids. Copper is less reactive than hydrogen, so it cannot displace hydrogen from acids.

14. Write the reaction of aluminum with hot concentrated sulfuric acid.

Answer:

$$2\text{Al} + 6\text{H}_2\text{SO}_4(\text{conc., hot}) \rightarrow \text{Al}_2(\text{SO}_4)_3 + 3\text{SO}_2 + 6\text{H}_2\text{O}$$

15. Define corrosion and give two ways to prevent it.

Answer:
Corrosion: Gradual deterioration of metal due to chemical reaction with surrounding environment.

Prevention methods:

1. **Galvanization:** Coating iron with zinc to prevent oxidation
2. **Oiling/Greasing:** Forms protective layer preventing air and water contact

16. Explain why alloys are preferred over pure metals.

Answer: Alloys have:

- Greater strength and hardness
- Better corrosion resistance
- Properties can be customized for specific uses
- Often cheaper than pure metals
- Better performance in demanding applications

Long Answer Questions

17. Describe the reactivity series of metals. How is it useful in predicting displacement reactions?

Answer:

Reactivity series: List of metals arranged in decreasing order of their reactivity (tendency to lose electrons).

Order (most to least reactive):
K > Na > Ca > Mg > Al > Zn > Fe > Ni > Sn > Pb > H > Cu > Hg > Ag > Au

Uses in displacement reactions:

A more reactive metal can displace a less reactive metal from its salt solution.

Example 1:

$$\mathrm{Zn} + \mathrm{CuSO_4} \rightarrow \mathrm{ZnSO_4} + \mathrm{Cu}$$

(Occurs because Zn is more reactive than Cu)

Example 2:

$$\mathrm{Cu} + \mathrm{ZnSO_4} \rightarrow (\text{No reaction})$$

(Does not occur because Cu is less reactive than Zn)

With acids:

- Metals above H can displace H_2 from dilute acids
- Metals below H cannot displace H_2

17. Compare the reactions of metals with oxygen and non-metals with oxygen.

Answer:

Aspect	Metals with O_2	Non-metals with O_2
Type of oxide	Basic oxides	Acidic or neutral oxides
General reaction	Metal + O_2 → Metal oxide (basic)	Non-metal + O_2 → Non-metal oxide (acidic/neutral)
Example	$2Na + O_2 \rightarrow Na_2O_2$	$C + O_2 \rightarrow CO_2$
Property in water	Forms base/alkali	Forms acid
Reaction with indicators	Turns red litmus to blue	Turns blue litmus to red
Amphoteric examples	Al_2O_3, ZnO	None (purely acidic/neutral)

Examples:

Metal oxides (basic):

- $2Mg + O_2 \rightarrow 2MgO$ (basic oxide)
- $MgO + H_2O \rightarrow Mg(OH)_2$ (forms base)

Non-metal oxides (acidic):

- $S + O_2 \rightarrow SO_2$ (acidic oxide)
- $SO_2 + H_2O \rightarrow H_2SO_3$ (forms acid)

19. Explain ionic bond formation with an example. Show electron transfer in the formation of NaCl.

Answer:

Ionic bond: Electrostatic attraction between oppositely charged ions formed by electron transfer.

Electron transfer in NaCl formation:

Step 1: Sodium loses electron

$$\text{Na (11 electrons)} \rightarrow \text{Na}^+(10 electrons) + e^-$$

- Sodium has 1 valence electron
- Loses it to achieve stable configuration (like Ne)
- Becomes positively charged cation

Step 2: Chlorine gains electron

$$\text{Cl (17 electrons)} + e^- \rightarrow \text{Cl}^-(18 electrons)$$

- Chlorine has 7 valence electrons
- Gains 1 to achieve stable configuration (like Ar)
- Becomes negatively charged anion

Step 3: Ionic bond forms

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

- Electrostatic attraction between Na^+ and Cl^-
- Strong ionic bond holds them together

Properties of NaCl (ionic compound):

- Soluble in water (polar solvent)
- Conducts electricity when dissolved (ions free to move)
- High melting point (1414°C) due to strong ionic bonds
- Hard and brittle crystals

20. State the difference between oxidation and reduction. Give two examples of redox reactions.

Answer:

Oxidation:

- Loss of electrons by a substance
- Originally meant loss of oxygen (still true in some contexts)
- Substance getting oxidized loses electrons and becomes positively charged

Reduction:

- Gain of electrons by a substance
- Originally meant gain of oxygen (still true in some contexts)
- Substance getting reduced gains electrons and becomes negatively charged

Key Point: Both occur simultaneously in redox reactions

Example 1: Combustion of Magnesium

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

- **Magnesium:** Oxidized (loses 2 electrons: $Mg \rightarrow Mg^{2+} + 2e^-$)
- **Oxygen:** Reduced (gains 2 electrons: $O + 2e^- \rightarrow O^{2-}$)
- **Type:** Redox reaction

Example 2: Zinc Displacing Copper

$$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$$

- **Zinc:** Oxidized (loses 2 electrons: $Zn \rightarrow Zn^{2+} + 2e^-$)
- **Copper:** Reduced (gains 2 electrons: $Cu^{2+} + 2e^- \rightarrow Cu$)
- **Type:** Redox reaction (displacement)

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15. EXAM-FOCUSED KEY POINTS

MUST REMEMBER

✓ Metals are generally malleable, ductile, lustrous, and good conductors

✓ Non-metals are generally brittle, dull, and poor conductors (except graphite)

✓ Metals form basic oxides, non-metals form acidic/neutral oxides

✓ Reactivity series: K > Na > Ca > Mg > Al > Zn > Fe > H > Cu > Ag > Au

✓ Only metals above hydrogen in series can displace H_2 from dilute acids

✓ A more reactive metal can displace a less reactive metal from salt solution

✓ Metals lose electrons (form cations), non-metals gain electrons (form anions)

✓ Metal-non-metal reactions form ionic compounds with ionic bonds

✓ Oxidation = loss of electrons, Reduction = gain of electrons

✓ Both occur simultaneously in redox reactions

✓ Amphoteric oxides: Al_2O_3, ZnO (react with both acids and bases)

✓ Alloys are homogeneous mixtures with better properties than pure metals

✓ Steel is most important alloy (Fe + C)

✓ Corrosion is prevented by galvanization, painting, oiling

✓ Mercury is liquid metal at room temperature (exception)

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16. COMMON EXAM PATTERNS

- Name metals/non-metals from given properties
- Complete and balance chemical equations
- Predict products of displacement reactions using reactivity series
- Explain why some metals react and others don't (reactivity basis)
- Compare properties of metals vs non-metals
- Describe ionic bond formation with electron transfer diagrams
- Applications of alloys in real life
- Corrosion prevention methods
- Distinguish between oxidation and reduction
- Identify oxidizing and reducing agents
- Uses of important metals and non-metals

REFERENCES

- RBSE Class 10 Science Chapter 3: Metals and Non-metals
- NCERT Science Textbook Class 10
- Standard chemical equations and laboratory procedures
- Reactivity series of metals (IUPAC)