

Chemical Reactions and Equations

COMPREHENSIVE STUDY NOTES - CLASS 10 RBSE

Chapter 1: Chemical Reactions and Equations

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1. INTRODUCTION TO CHEMICAL REACTIONS

What is a Chemical Reaction?

A chemical reaction is a process in which substances (reactants) undergo chemical change to form new substances (products) with different properties.

Observable Signs of Chemical Reactions

Chemical reactions can be identified by observing the following changes:

Observable Change	Example
Change in State	Solid to liquid, liquid to gas, etc.
Change in Colour	Magnesium ribbon burning produces white powder
Evolution of Gas	Zinc + dilute acid produces hydrogen gas
Change in Temperature	Calcium oxide + water releases heat

Key Activities:

- Activity 1.1: Magnesium ribbon burns with dazzling white flame in oxygen
 - Activity 1.2: Lead nitrate + potassium iodide produces yellow precipitate
 - Activity 1.3: Zinc + dilute sulphuric acid produces hydrogen gas with temperature increase
-

2. CHEMICAL EQUATIONS AND BALANCING

2.1 Representation of Chemical Reactions

Word Equation

A word equation represents a chemical reaction using the names of substances:

Format: Reactants → Products

Example: Magnesium + Oxygen → Magnesium oxide

Components:

- **Reactants (LHS):** Substances that undergo chemical change
- **Products (RHS):** New substances formed
- **Arrow (→):** Shows direction of reaction

Chemical Equation

Uses chemical formulae instead of words for more concise representation:

Example: $\text{Mg} + \text{O}_2 \rightarrow \text{MgO}$

2.2 Skeletal Chemical Equation

An unbalanced chemical equation where atom numbers may not be equal on both sides.

Example: $\text{Mg} + \text{O}_2 \rightarrow \text{MgO}$ (Unbalanced)

2.3 Balanced Chemical Equation

An equation where the number of atoms of each element is the same on both reactant side (LHS) and product side (RHS).

Law of Conservation of Mass: Mass cannot be created or destroyed in a chemical reaction. The total mass of elements in reactants = total mass of elements in products.

Example: $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$ (Balanced)

2.4 Steps to Balance a Chemical Equation (Hit-and-Trial Method)

Example: $\text{Fe} + \text{H}_2\text{O} \rightarrow \text{Fe}_3\text{O}_4 + \text{H}_2$

Step I: Draw Boxes Around Formulae

Helps organize the equation without changing formula contents.

Step II: List Number of Atoms

Create a table counting atoms of each element on LHS and RHS:

Element	Reactants (LHS)	Products (RHS)
Fe	1	3
H	2	2
O	1	4

Step III: Balance the Most Complex Compound

Start with the compound having maximum atoms. Balance the element with maximum atoms in that compound.

- Select Fe₃O₄ and oxygen
- Need 4 oxygen atoms on LHS: Use 4H₂O
- Equation becomes: Fe + 4H₂O → Fe₃O₄ + H₂

Step IV: Balance Other Elements

Next balance hydrogen atoms:

- 4H₂O gives 8 hydrogen atoms
- Need 8 H atoms on RHS: Use 4H₂
- Equation becomes: Fe + 4H₂O → Fe₃O₄ + 4H₂

Step V: Balance Remaining Elements

Balance iron atoms:

- 3 Fe atoms needed on LHS
- Final equation: 3Fe + 4H₂O → Fe₃O₄ + 4H₂

Step VI: Verify Balanced Equation

Count atoms on both sides:



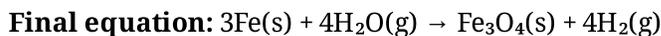
Element	LHS	RHS
Fe	3	3 ✓
H	8	8 ✓
O	4	4 ✓

Step VII: Add Physical State Symbols

Physical states are represented as:

- **(s)** = Solid
- **(l)** = Liquid
- **(g)** = Gas

- **(aq)** = Aqueous solution (dissolved in water)



2.5 Reaction Conditions

Conditions like temperature, pressure, or catalyst are written above/below the arrow:

Example:

- $340 \text{ atm CO(g)} + 2\text{H}_2\text{(g)} \xrightarrow{\text{Pressure}} \text{CH}_3\text{OH(l)}$
- $6\text{CO}_2\text{(aq)} + 12\text{H}_2\text{O(l)} \xrightarrow{\text{Sunlight, Chlorophyll}} \text{C}_6\text{H}_{12}\text{O}_6\text{(aq)} + 6\text{O}_2\text{(aq)} + 6\text{H}_2\text{O(l)}$

3. TYPES OF CHEMICAL REACTIONS

3.1 Combination Reaction

Definition: A reaction in which two or more reactants combine to form a single product.

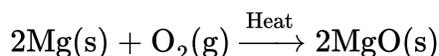
General Form: $\text{A} + \text{B} \rightarrow \text{AB}$

Characteristics:

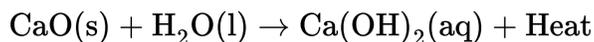
- Heat is usually released (exothermic)
- Number of products = 1
- Number of reactants ≥ 2

Examples:

1. Burning of Magnesium:

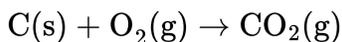


2. Reaction of Calcium Oxide with Water:

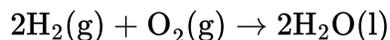


(Forms slaked lime, used for whitewashing)

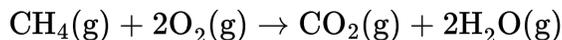
3. Burning of Coal:



4. Formation of Water:



5. Burning of Natural Gas:



3.2 Exothermic Reactions

Definition: Reactions that release heat energy along with product formation.

Heat is released because:

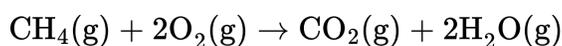
- Breaking old bonds requires energy (endothermic)
- Forming new bonds releases more energy (exothermic)
- Net result: Heat is released

Common Examples:

1. Respiration (Glucose oxidation in cells)



2. Combustion of Natural Gas



3. Decomposition of Vegetable Matter to Compost

- Slow oxidation process releasing heat

3.3 Decomposition Reaction

Definition: A reaction in which a single reactant breaks down into two or more simpler products.

General Form: $\text{AB} \rightarrow \text{A} + \text{B}$

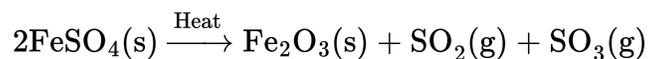
Characteristics:

- Requires energy input (heat, light, or electricity)
- Number of reactants = 1
- Number of products ≥ 2
- Opposite of combination reactions

Types of Decomposition:

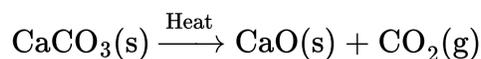
A. Thermal Decomposition (Heat Energy)

1. Decomposition of Ferrous Sulphate:



- Green crystals \rightarrow Reddish-brown powder + gases
- Characteristic burning sulphur smell

2. Decomposition of Calcium Carbonate:



- Used in cement manufacturing
- Calcium oxide called "quick lime" or "lime"

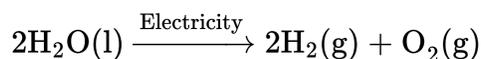
3. Decomposition of Lead Nitrate:



- Brown fumes of nitrogen dioxide observed
- Activity 1.6 demonstration

B. Electrolytic Decomposition (Electricity)

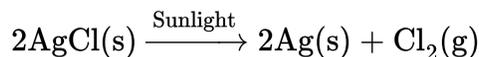
Electrolysis of Water:



- Activity 1.7 demonstration
- Hydrogen collects at cathode (negative electrode)
- Oxygen collects at anode (positive electrode)
- Volume of hydrogen gas = 2 × volume of oxygen gas
- Ratio 2:1

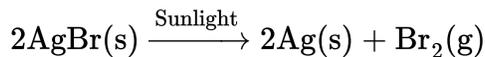
C. Photochemical Decomposition (Light Energy)

1. Decomposition of Silver Chloride:



- White AgCl → Gray powder (silver metal)
- Used in black and white photography

2. Decomposition of Silver Bromide:



3.4 Endothermic Reactions

Definition: Reactions that absorb energy from surroundings.

Examples:

- All decomposition reactions
- Barium hydroxide + ammonium chloride reaction (cooling effect felt)
- Photosynthesis

3.5 Displacement Reaction

Definition: A reaction in which a more reactive element displaces a less reactive element from its compound.

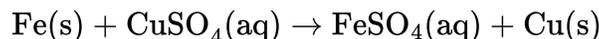
General Form: $\text{A} + \text{BC} \rightarrow \text{AC} + \text{B}$

Characteristics:

- One element and one compound as reactants
- A more reactive element displaces less reactive element
- Based on reactivity series

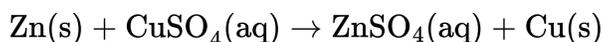
Examples:

1. Iron Displacing Copper:

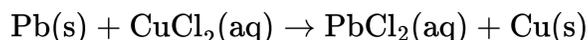


- Activity 1.9 demonstration
- Blue solution fades (copper sulphate reduced)
- Iron nail becomes brownish (copper deposited)

2. Zinc Displacing Copper:



3. Lead Displacing Copper:



Reactivity Series (Partial):

More Reactive → Less Reactive

K > Na > Ca > Mg > Al > Zn > Fe > Cu > Hg > Ag > Au

3.6 Double Displacement Reaction

Definition: A reaction in which two compounds exchange their ions to form two new compounds.

General Form: AB + CD → AD + CB

Characteristics:

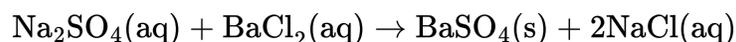
- Two compounds as reactants
- Exchange of ions/groups occurs
- Often forms a precipitate (insoluble solid)

Precipitation Reaction:

When an insoluble solid (precipitate) forms as a product.

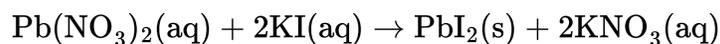
Examples:

1. Sodium Sulphate + Barium Chloride:



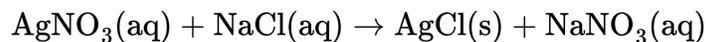
- Activity 1.10 demonstration
- White precipitate of barium sulphate forms
- Sodium and chloride remain in solution

2. Lead Nitrate + Potassium Iodide:



- Activity 1.2 demonstration
- Yellow precipitate of lead iodide forms

3. Silver Nitrate + Sodium Chloride:



- White precipitate of silver chloride
-

4. OXIDATION AND REDUCTION

4.1 Definitions

Oxidation:

- **Original Definition:** Gain of oxygen or loss of hydrogen
- **Broader Definition:** Loss of electrons

Reduction:

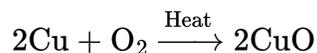
- **Original Definition:** Loss of oxygen or gain of hydrogen
- **Broader Definition:** Gain of electrons

Redox Reaction (Oxidation-Reduction Reaction):

A reaction in which both oxidation and reduction occur simultaneously.

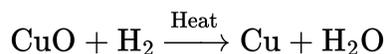
4.2 Examples of Oxidation-Reduction Reactions

1. Oxidation of Copper:



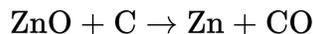
- Copper gains oxygen (oxidized)
- Oxygen is reduced

2. Reduction of Copper Oxide by Hydrogen:



- Copper oxide loses oxygen (reduced)
- Hydrogen gains oxygen (oxidized)

3. Reaction of Zinc with Carbon:



- Carbon oxidized to CO
- ZnO reduced to Zn

4. Reaction of MnO₂ with HCl:



- HCl oxidized to Cl₂
- MnO₂ reduced to MnCl₂

5. Reaction of Sodium with Oxygen:



- Sodium oxidized (loses electrons)
- Oxygen reduced (gains electrons)

4.3 How to Identify Oxidized and Reduced Substances

Process	Oxygen Perspective	Hydrogen Perspective
Oxidation	Gains oxygen	Loses hydrogen
Reduction	Loses oxygen	Gains hydrogen

Step-by-step Method:

1. Identify which element's oxidation state changed
2. If element gains oxygen OR loses hydrogen → **Oxidized**
3. If element loses oxygen OR gains hydrogen → **Reduced**

5. EFFECTS OF OXIDATION REACTIONS IN EVERYDAY LIFE

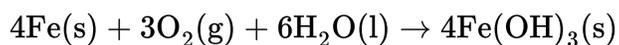
5.1 Corrosion

Definition: The process in which a metal is attacked by substances in the environment (moisture, oxygen, acids) leading to its degradation.

Rusting of Iron:

Iron articles appear shiny when new but develop a reddish-brown coating when exposed to air and moisture.

Chemical Reaction:



Other Examples of Corrosion:

Metal	Coating Color	Name
Iron	Reddish-brown	Rust
Silver	Black	Silver sulphide (Ag ₂ S)
Copper	Green	Copper carbonate/oxide

Problems Caused by Corrosion:

- Damage to car bodies
- Deterioration of bridges and iron railings
- Damage to ships and structures
- Enormous financial loss annually

Prevention Methods:

1. **Painting** - Prevents exposure to air and moisture
2. **Oiling/Greasing** - Creates protective layer
3. **Galvanization** - Coating with zinc
4. **Tinning** - Coating with tin
5. **Chromium Plating** - Protective shine
6. **Alloying** - Creating rust-resistant materials (stainless steel)

5.2 Rancidity

Definition: The process in which fats and oils undergo oxidation, causing change in their taste and smell.

Causes:

- Exposure to air (oxygen)
- Prolonged storage
- Exposure to light
- Bacterial action

Effects:

- **Taste** changes (becomes unpleasant)
- **Smell** changes (becomes unpleasant)
- **Nutritional value** decreases
- **Product spoilage**

Prevention Methods:

1. **Adding Antioxidants:**
 - Compounds that prevent oxidation
 - Common examples: BHA, BHT, vitamin E
 2. **Proper Storage:**
 - Airtight containers (prevents oxygen exposure)
 - Cool, dark places (prevents light exposure)
 - Refrigeration (slows oxidation)
 3. **Nitrogen Flushing:**
 - Chips/snacks packets filled with nitrogen gas
 - Displaces oxygen
 - Prevents oxidation of fats
 4. **Vacuum Packaging:**
 - Removes air/oxygen
 - Extends shelf life
-

6. EXAM-IMPORTANT QUESTIONS AND ANSWERS

SHORT ANSWER QUESTIONS (2-3 marks)

Q1. Why should a magnesium ribbon be cleaned before burning in air?

Answer: The surface of magnesium ribbon gets coated with magnesium oxide (MgO) and magnesium nitride (Mg₃N₂) which act as non-reactive layers. Cleaning with sand paper removes these coatings, allowing the pure magnesium to react directly with oxygen in air, producing a bright white flame.

Q2. Distinguish between exothermic and endothermic reactions with examples.

Answer:

Exothermic	Endothermic
Release heat to surroundings	Absorb heat from surroundings
Temperature of surroundings increases	Temperature of surroundings decreases
Example: Burning of coal	Example: Decomposition of lead nitrate
Example: Respiration	Example: Melting of ice

Q3. Define oxidation and reduction in terms of oxygen transfer.

Answer:

- **Oxidation:** When a substance gains oxygen or loses hydrogen during a reaction.
 - **Reduction:** When a substance loses oxygen or gains hydrogen during a reaction.
-

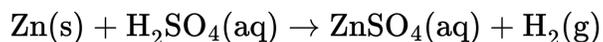
Q4. Why are decomposition reactions called opposite of combination reactions?

Answer:

- **Combination Reaction:** $A + B \rightarrow AB$ (2 or more substances combine to form 1)
 - **Decomposition:** $AB \rightarrow A + B$ (1 substance breaks into 2 or more substances)
 - They are reverse processes of each other
-

Q5. What happens when zinc granules are added to dilute sulphuric acid?

Answer:



Observations:

- Effervescence observed (bubbles of hydrogen gas)

- Temperature of test tube increases
- Zinc gradually dissolves
- This is a **displacement reaction** and **exothermic reaction**

Q6. What is the difference between displacement and double displacement reactions?

Answer:

Displacement	Double Displacement
Form: $A + BC \rightarrow AC + B$	Form: $AB + CD \rightarrow AD + CB$
Reactants: 1 element + 1 compound	Reactants: 2 compounds
Process: Element displaces another element	Process: Ions exchange between compounds
Example: $Fe + CuSO_4 \rightarrow FeSO_4 + Cu$	Example: $BaCl_2 + Na_2SO_4 \rightarrow BaSO_4 + NaCl$

LONG ANSWER QUESTIONS (5-6 marks)

Q7. Explain the steps of balancing a chemical equation with example.

Answer: [See Section 2.4 for detailed explanation with $Fe + H_2O \rightarrow Fe_3O_4 + H_2$ example]

Q8. What is corrosion? How can it be prevented?

Answer:

Definition: Corrosion is the process by which a metal is attacked by substances in the surrounding environment (oxygen, moisture, acids) and gradually deteriorates.

Example - Rusting of Iron:



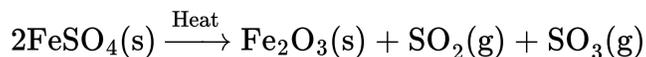
Prevention Methods:

1. **Painting** - Creates barrier against air and moisture
2. **Oiling/Greasing** - Protective liquid layer
3. **Galvanization** - Coating with zinc metal
4. **Chromium plating** - Protective metallic coating
5. **Alloying** - Making stainless steel (Fe + Cr + Ni)
6. **Proper storage** - Dry, cool conditions

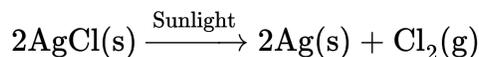
Q9. Explain with equations: (a) Thermal decomposition, (b) Photochemical decomposition, (c) Electrolytic decomposition

Answer:

(a) Thermal Decomposition (Heat):

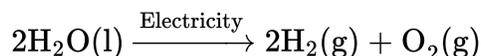


(b) Photochemical Decomposition (Light):



(Used in photography)

(c) Electrolytic Decomposition (Electricity):



Q10. Identify the oxidized and reduced substances in the reaction:



Answer:

- **Sodium (Na):** Oxidized (loses electrons, forms Na^+ ions)
- **Oxygen (O_2):** Reduced (gains electrons, forms O^{2-} ions)
- This is a **redox reaction** where Na loses electrons and O_2 gains them

7. QUICK REVISION POINTS

Key Formulas and Balanced Equations:

Combination Reactions:

- $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$
- $\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca}(\text{OH})_2$
- $\text{C} + \text{O}_2 \rightarrow \text{CO}_2$
- $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$

Decomposition Reactions:

- $2\text{FeSO}_4 \rightarrow \text{Fe}_2\text{O}_3 + \text{SO}_2 + \text{SO}_3$
- $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$
- $2\text{Pb}(\text{NO}_3)_2 \rightarrow 2\text{PbO} + 4\text{NO}_2 + \text{O}_2$
- $2\text{H}_2\text{O} \rightarrow 2\text{H}_2 + \text{O}_2$

Displacement Reactions:

- $\text{Fe} + \text{CuSO}_4 \rightarrow \text{FeSO}_4 + \text{Cu}$
- $\text{Zn} + \text{CuSO}_4 \rightarrow \text{ZnSO}_4 + \text{Cu}$
- $\text{Pb} + \text{CuCl}_2 \rightarrow \text{PbCl}_2 + \text{Cu}$

Double Displacement Reactions:

- $\text{Na}_2\text{SO}_4 + \text{BaCl}_2 \rightarrow \text{BaSO}_4\downarrow + 2\text{NaCl}$
- $\text{Pb}(\text{NO}_3)_2 + 2\text{KI} \rightarrow \text{PbI}_2\downarrow + 2\text{KNO}_3$
- $\text{AgNO}_3 + \text{NaCl} \rightarrow \text{AgCl}\downarrow + \text{NaNO}_3$

Redox Reactions:

- $2\text{Cu} + \text{O}_2 \rightarrow 2\text{CuO}$
- $\text{CuO} + \text{H}_2 \rightarrow \text{Cu} + \text{H}_2\text{O}$
- $\text{ZnO} + \text{C} \rightarrow \text{Zn} + \text{CO}$

Essential Definitions:

Term	Definition
Chemical Reaction	Process where reactants change to products
Balanced Equation	Atoms of each element same on both sides
Combination	2+ reactants \rightarrow 1 product
Decomposition	1 reactant \rightarrow 2+ products
Displacement	$\text{A} + \text{BC} \rightarrow \text{AC} + \text{B}$
Double Displacement	$\text{AB} + \text{CD} \rightarrow \text{AD} + \text{CB}$
Exothermic	Releases heat energy
Endothermic	Absorbs heat energy
Oxidation	Gain O_2 or loss H_2
Reduction	Loss O_2 or gain H_2
Corrosion	Metal attack by environment
Rancidity	Oxidation of fats/oils

Physical State Symbols:

- **(s)** = Solid
- **(l)** = Liquid
- **(g)** = Gas
- **(aq)** = Aqueous solution
- \downarrow = Precipitate (insoluble solid)

Reactivity Series (Decreasing):

K > Na > Ca > Mg > Al > Zn > Fe > Cu > Hg > Ag > Au

Important Activities and Demonstrations:

Activity	Key Observation	Equation
Activity 1.1	Mg burns with white flame	$2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$
Activity 1.3	H ₂ gas from Zn + acid	$\text{Zn} + \text{H}_2\text{SO}_4 \rightarrow \text{ZnSO}_4 + \text{H}_2$
Activity 1.4	CaO + H ₂ O (heat released)	$\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2$
Activity 1.5	FeSO ₄ color change when heated	$2\text{FeSO}_4 \rightarrow \text{Fe}_2\text{O}_3 + \text{SO}_2 + \text{SO}_3$
Activity 1.7	Water electrolysis (H ₂ :O ₂ = 2:1)	$2\text{H}_2\text{O} \rightarrow 2\text{H}_2 + \text{O}_2$
Activity 1.8	AgCl white to gray in sunlight	$2\text{AgCl} \rightarrow 2\text{Ag} + \text{Cl}_2$
Activity 1.9	Cu deposits on Fe nail	$\text{Fe} + \text{CuSO}_4 \rightarrow \text{FeSO}_4 + \text{Cu}$
Activity 1.10	White BaSO ₄ precipitate forms	$\text{Na}_2\text{SO}_4 + \text{BaCl}_2 \rightarrow \text{BaSO}_4\downarrow + \text{NaCl}$
Activity 1.11	CuO reduced by H ₂	$\text{CuO} + \text{H}_2 \rightarrow \text{Cu} + \text{H}_2\text{O}$

FREQUENTLY ASKED EXAM QUESTIONS

- 1. What are the signs of chemical reactions?**
 - Change in state, color, temperature, gas evolution
- 2. Difference between physical and chemical changes?**
 - Physical: Reversible, no new substance; Chemical: Irreversible, new substance formed
- 3. Why must chemical equations be balanced?**
 - Law of conservation of mass - atoms cannot be created or destroyed
- 4. Which type of reaction requires energy input?**
 - Decomposition reactions (thermal, photochemical, electrolytic)
- 5. Give an example of double displacement with precipitate formation:**
 - $\text{BaCl}_2 + \text{Na}_2\text{SO}_4 \rightarrow \text{BaSO}_4 + 2\text{NaCl}$

6. How is corrosion prevented?

- Painting, galvanization, chromium plating, alloying, proper storage

7. What prevents rancidity?

- Antioxidants, airtight containers, nitrogen flushing, cool storage

8. Name one thermal, photochemical, and electrolytic decomposition:

- Thermal: $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$
- Photochemical: $2\text{AgCl} \rightarrow 2\text{Ag} + \text{Cl}_2$
- Electrolytic: $2\text{H}_2\text{O} \rightarrow 2\text{H}_2 + \text{O}_2$

PRACTICE EXERCISE ANSWERS

Question Set 1: Multiple Choice

Q1. Statement (b) and (c) are incorrect. (Answer: ii)

Explanation: In reaction $2\text{PbO} + \text{C} \rightarrow 2\text{Pb} + \text{CO}_2$

- Lead is reduced (gains electrons) ✓
- Carbon is oxidized (loses electrons) ✓
- CO_2 is not oxidized further (it's the product)

Q2. Answer: (d) Displacement reaction

Q3. Answer: (a) Hydrogen gas and iron chloride are produced

Numerical and Balancing Problems:

Problem 1: $\text{H}_2 + \text{Cl}_2 \rightarrow \text{HCl}$

Balanced: $\text{H}_2 + \text{Cl}_2 \rightarrow 2\text{HCl}$

Problem 2: $\text{BaCl}_2 + \text{Al}_2(\text{SO}_4)_3 \rightarrow \text{AlCl}_3 + \text{BaSO}_4$

Balanced: $3\text{BaCl}_2 + \text{Al}_2(\text{SO}_4)_3 \rightarrow 2\text{AlCl}_3 + 3\text{BaSO}_4\downarrow$

Problem 3: $2\text{HNO}_3 + \text{Ca}(\text{OH})_2 \rightarrow \text{Ca}(\text{NO}_3)_2 + 2\text{H}_2\text{O}$

(Double displacement/Neutralization)

IMPORTANT NOTES FOR EXAM PREPARATION

- 1. Always balance equations** - Every numerical answer must have balanced equation
 - 2. Show all steps** - Write word equation → skeletal equation → balanced equation
 - 3. Include physical states** - Use (s), (l), (g), (aq) symbols
 - 4. Know activity series** - Important for displacement reactions
 - 5. Practice balancing** - Use hit-and-trial method systematically
 - 6. Identify reaction types** - Can be asked in MCQ or descriptive questions
 - 7. Real-world applications** - Corrosion and rancidity have practical importance
 - 8. Activities explanation** - Be ready to explain what happens in Activities 1.1-1.11
-

ADDITIONAL IMPORTANT CONCEPTS

Conservation of Mass:

- Total mass of reactants = Total mass of products
- Atoms are neither created nor destroyed
- This is why equations must be balanced

Energy in Chemical Reactions:

- **Exothermic:** Heat released (ΔH negative)
- **Endothermic:** Heat absorbed (ΔH positive)

Respiration (Important Exothermic Reaction):



This provides energy for all living activities.

Photography Application:

- Silver halides decompose in light
- This forms metallic silver on photographic film
- Creates photographic image

END OF COMPREHENSIVE NOTES

These notes cover all topics from Chapter 1 of Class 10 RBSE Science textbook in an exam-oriented format. Study these thoroughly and practice all balancing equations and reaction types for maximum marks.