CHEMICAL REACTIONS

Learning outcomes

By the end of this theme, students will be able to:

- describe the different types of chemical reactions with examples
- identify the type of chemical reaction
- explain the reactivity series of metals
- predict reactions based on the reactivity series of metals
- differentiate between exothermic and endothermic reactions
- describe neutralisation reactions
- describe how oxides are formed and classify them as acidic, basic, amphoteric and neutral
- explain the effect of heat on some oxides

Warm-up

Encourage students to complete the *Get going* section given at the beginning of the chapter.

GUIDELINES TO TEACH

Introduction

- Help students recall various examples of physical and chemical changes.

Chemical reactions and chemical equations

- **Recap:** Explain that a chemical equation represents a chemical reaction using symbols and formulae.
- Explain about all the factors that a chemical equation should inform us.

Ask students:

What is a symbol?

What is a formula?

What are reactants?

What are products?

What information can you get from a chemical reaction? What information can you not get?

Changes that accompany a chemical reaction

- Discuss chemical changes that happen around us and within us. Elicit examples from students.
- Highlight the fact that a chemical change involves the breaking of old bonds and the making of new ones.

- List the indicators of a chemical change (chemical reaction) other than the formation of new substances, as given in the coursebook.

Change in colour

- Carry out the experiment, given in the coursebook, to demonstrate a change in colour during a chemical reaction.

Ask student:

How does copper get deposited on the iron nail? (The iron displaces the copper in the copper sulphate solution.)

Change of state

- Give the example of water. Hydrogen and oxygen are both gases. But when two hydrogen atoms combine with an oxygen atom, water is formed which is a liquid with totally different properties from that of oxygen and hydrogen.

Evolution of a gas

- Define effervescence.
- Carry out the experiment, given in the coursebook, to demonstrate the evolution of hydrogen gas when sodium reacts with dil. hydrochloric acid.
- Discuss the information given in 'Spotlight'.

Formation of a precipitate

- Define precipitate.
- Carry out the experiment, given in the coursebook, to demonstrate the formation of a precipitate.

Emission or absorption of energy (exothermic or endothermic reactions)

- Recap what an exothermic reaction is with the help of the equation, given in the coursebook, and the role that fuels play in combustion reactions.
- With the help of the example, given in the coursebook, explain endothermic reactions.
- Discuss the information given in 'Know more'.

Production or absorption of heat during a chemical reaction

- Explain the breaking and formation of chemical bonds, using the simple explanatory equation, as given in the coursebook.

Ask students:

What are the key indicators of a chemical change?

Conditions for a chemical reaction to occur

Ask students:

Can you think of the conditions that are required for chemical reactions to take place?

(Give hints—the conditions for rusting, photosynthesis and combustion, presence of catalysts, changes in temperature and pressure and so on.)

- Discuss the conditions, as given in the coursebook, for a chemical reaction to happen.
- Using the information from the 'Eco corner', discuss the formation and effects of photochemical smog.
- Use '**Stop and check**' for a quick recap.

Types of chemical reactions

- Carry out experiments to demonstrate combination, decomposition, displacement and double displacement reactions in the laboratory. Provide more examples of such reactions for better understanding.
- Explain the importance of the metal reactivity series in predicting how a metal will react with the salt solution or oxide of another metal.
- Use the activity series of metals to make students understand that metals are arranged in order of decreasing reactivity (from the most reactive metal to the least reactive metal).
- Let students predict the outcome of a few reactions involving metals and salt solutions and give their equations.

Ask students:

Name the metals which do not react with acids to liberate hydrogen.

Will there be any reaction if copper (Cu) is added to iron sulphate solution?
Explain (refer the reactivity series).

- Share the 'SciTech' information on the uses of potassium chlorate and lead oxides.
- Discuss the information given in 'Know more'.
- Use '**Stop and check'** for a quick recap.
- With the help of the equations, given in the coursebook, explain double displacement reactions.
- Discuss the information given in 'Spotlight'.
- Demonstrate precipitation and neutralisation reactions in the laboratory.
- Discuss how neutralisation reactions play an important role in our life.

Ask students:

In what kind of reaction is only a single product formed?

In what kind of reaction is only one reactant involved?

Are the products of decomposition reactions elements or compounds?

What are displacement reactions also known as?

How do neutralisation reactions help in the treatment of acidity?

What happens when solutions of potassium iodide and lead(II) nitrate are mixed?

Why can't copper replace zinc in any reaction?

What happens when calcium carbonate is heated? How can the presence of carbon dioxide be tested?

What happens when lead (IV) oxide is heated to around 600°C?

- Use '**Stop and check**' for a quick recap.

Oxides

- Explain what an oxide is.

Metal oxides

Ask students:

In your notebook, write down as many metal oxides as you can think of along with their formulae.

- Tell students to check their list against the table '*Some metallic oxides and their formulae*', given in the coursebook.

The preparation of metal oxides

- Discuss how metal oxides are formed by the reaction of some metals with oxygen and thermal decomposition of metal carbonates, nitrates and hydroxides.
- Demonstrate the activity *'To observe the thermal decomposition of limestone ($CaCO_3$)'*, given in the coursebook.

Non-metal oxides

Ask students:

In your notebook, write down as many non-metal oxides as you can think of along with their formulae.

- Tell students to check their list against the table '*Some non-metallic oxides and their formulae*', given in the coursebook.

The preparation of non-metal oxides

- Help students to write down some simple equations to form non-metal oxides.

The nature of oxides

- Discuss basic, acidic, amphoteric and neutral oxides, as given in the coursebook.
- Share the information given in 'Career watch'.

CHEMICAL REACTIONS

Stop and check

1. False   2. True   3. False   4. False   5. True   6. False

Stop and check

1. Combination reaction: $A + B \rightarrow AB$
   Decomposition reaction: $AB \rightarrow A + B$
2. When electricity is passed through acidified water, water breaks up into hydrogen and oxygen.

$$2H_2O\ (l) \xrightarrow{\text{electricity}} O_2\ (g) + 2H_2\ (g)$$

3. Equation (ii) will not occur. Copper cannot displace zinc as copper is below zinc in the reactivity series.

Stop and check

1. Anions are negatively charged particles. Cations are positively charged particles.
2. The general representation of a double displacement reaction is:

$$AB + CD \rightarrow AD + CB$$

3. An ant bite can be treated at home by applying some baking soda, a base, on the bite. This base neutralises the formic acid that is left on the skin as a result of the ant bite.
4. Farmers neutralise acidic soil by adding lime, and alkaline soil by adding compost that is acidic.

Checkpoint

A. 1. b, decomposition reaction   2. d, double displacement reaction   3. c, $AB + C \rightarrow AC + B$
4. b, sodium   5. c, $2KI + Pb\ (NO_3)_2 \rightarrow PbI_2 + 2KNO_3$

B. 1. chemical equation   2. symbols   3. exothermic   4. precipitate
5. magnesium oxide   6. acids

C. 1. $Fe + CuSO_4 \rightarrow FeSO_4 + Cu$
2. $NH_3 + HCl \rightarrow NH_4Cl$
3. $2Na + 2HCl \rightarrow 2NaCl + H_2$
4. $BaCl_2 + Na_2SO_4 \rightarrow 2NaCl + BaSO_4$
5. $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

D. 1.

Decomposition reaction	Displacement reaction

In a decomposition reaction, a single reactant breaks down into two or more products.	In a displacement reaction, an element replaces another element. This type of reaction is also called a substitution reaction or a single replacement reaction.
The general representation of this type of reaction is: $AB \rightarrow A + B$	The general representation of this type of reaction is: $AB + C \rightarrow AC + B$

2.

Amphoteric oxides	Neutral oxides
These oxides react with both acids and bases to form a salt and water. The solutions of these oxides are neutral and do not change the colour of blue or red litmus paper.	These oxides do not react with acids or bases.
Examples include zinc oxide (ZnO) and aluminium oxide (Al_2O_3).	Examples include water (H_2O) and carbon monoxide (CO).

E. 1. Iron in the presence of oxygen and moisture forms hydrated iron (III) oxide (*rust*), which is reddish-brown in colour.

$4Fe\ (s) + 3O_2\ (g) \rightarrow 2Fe_2O_3\ (s)$

(iron) oxygen iron(III) oxide
(red deposit)

2. Copper reacts with concentrated nitric acid to form copper nitrate, nitrogen dioxide and water. Nitrogen dioxide is a pungent reddish-brown poisonous gas.

$Cu(s) + 4HNO_3\ (conc) \rightarrow Cu(NO_3)_2\ (aq) + 2NO_2\ (g) + 2H_2O\ (l)$

3. When calcium oxide is dissolved in water, calcium hydroxide is formed.

$CaO(s) + H_2O\ (l) \rightarrow Ca(OH)_2\ (aq)$

4. When hydrogen peroxide is heated, it splits into water and oxygen.

$2H_2O_2\ (s) \rightarrow 2H_2O\ (l) + O_2\ (g)$

5. When copper is added to zinc sulphate, no reaction takes place as copper lies below zinc in the reactivity series.

6. When a base such as sodium hydroxide is added to hydrochloric acid, neutral sodium chloride and water are formed.

$HCl\ (aq) + NaOH\ (aq) \rightarrow NaCl\ (aq) + H_2O$

7. When carbon dioxide is bubbled through sulphuric acid, no reaction takes place.

F. 1. Reactions should be classified so that we can understand the chemical changes that occur in different reactions. We can also predict the products of other reactions.

2. The four main types of reactions are combination or synthesis reactions, decomposition reactions, displacement reactions and double displacement reactions.
3. A precipitation reaction is a double displacement reaction.
   When a solution of silver nitrate is added to a solution of sodium chloride, a white precipitate of silver chloride is formed.

   $NaCl\ (aq) + AgNO_3\ (aq) \rightarrow AgCl \downarrow + NaNO_3\ (aq)$

sodium chloride	silver nitrate	silver chloride	sodium nitrate

4. If the energy required to break chemical bonds in the reactants is more than that required to form chemical bonds in the products, the reaction is exothermic.
   If the energy required to form chemical bonds in the products is more than that required to break chemical bonds in the reactants, the reaction is endothermic.
5. When carbon dioxide is dissolved in water, carbonic acid is formed. When blue litmus paper is dipped in this solution, the paper turns red.

G. 1. The following changes are possible in a chemical reaction: change in colour, change in state, evolution of gas, formation of a precipitate and absorption or emission of energy.

   (i) *Change in colour:* We notice a change in colour when some chemical reactions occur. For example, in the presence of oxygen and moisture, iron forms iron (III) oxide or rust which is reddish-brown in colour.

   $4Fe\ (s) + 3O_2\ (g) \rightarrow 2Fe_2O_3\ (s)$

   (ii) *Change of state:* In some chemical reactions, the reactants may be in a different state from that of the product. For example, gaseous reactants (such as hydrogen and oxygen) give a liquid product (water).

   $2H_2\ (g) + O_2\ (g) \rightarrow 2H_2O\ (l)$

   (iii) *Evolution of a gas:* During some chemical reactions, a gas is one of the products.

   When a metal is added to a dilute acid, hydrogen gas is evolved and a salt is formed.

   $2Na + 2HCl\ (dil) \rightarrow 2NaCl + H_2\uparrow$

   (iv) *Formation of a precipitate:* One of the products formed during a chemical reaction may be an insoluble solid (precipitate) that settles at the bottom of the test tube.

   When a solution of silver nitrate is added to a solution of sodium chloride, a white precipitate of silver chloride is formed.

   $NaCl\ (aq) + AgNO_3\ (aq) \rightarrow AgCl\downarrow + NaNO_3\ (aq)$

   (v) *Emission or absorption of energy:* Besides forming products, some reactions release a large quantity of heat (exothermic), while some other reactions absorb heat from the surroundings (endothermic).
   For example, when methane (CH_4) gas burns in air, carbon dioxide and water vapour are formed, and a large quantity of heat is given out.

2. The following are some of the conditions that are required for a chemical reaction to occur.

   (i) The reactants should be in contact so that the chemical reaction can occur. Powdering the reactants and mixing the reactants are some ways of making sure that a chemical reaction occurs. For example, barium hydroxide and ammonium chloride react in the solid state. In many cases, the reactants are made into solutions and mixed.

(ii) Some reactions need heat to be applied to begin. This is the case with combustion reactions. For example, a lit match is required to light a candle.

(iii) Some reactions need a catalyst. For example, platinum acts as a catalyst in the reaction of carbon monoxide and oxygen to give carbon dioxide and increases the reaction rate.

(iv) Some reactions may need light to occur. They are called photochemical reactions. Photosynthesis and the formation of vitamin D in the skin are good examples. Photography also depends on photochemical reactions.

(v) Some reactions need an increase or a decrease in pressure to occur. This is because increasing the pressure brings molecules of reactants closer together.
This is especially so in the case of gases. For example, the reaction of nitrogen with hydrogen to produce ammonia yields larger amounts of ammonia when the reactants are under high pressure. (*Any three*.)

3. i. $CuSO_4 + H_2S \rightarrow CuS + H_2SO_4$ (double displacement reaction)

   ii. $SO_2 + 2KOH \rightarrow K_2SO_3 + H_2O$ (neutralisation reaction)

   iii. $Mg + FeSO_4 \rightarrow MgSO_4 + Fe$ (displacement reaction)

   iv. $CuSO_4 + 2NaOH \rightarrow Cu(OH)_2 + Na_2SO_4$ (double displacement reaction)

   v. $CuCO_3 \rightarrow CuO + CO_2$ (decomposition reaction)

   vi. $PCl_3 + Cl_2 \rightarrow PCl_5$ (combination reaction)

4. The reactivity series lists metals in the order of their reactivity. Metals at the top of the list are more reactive than those at the lower end. The reactivity of the metals decreases down the list. A more reactive metal will replace a less reactive metal from a solution of its salt or from its oxide.

   The reactivity series can be used to predict what will happen when a metal reacts with the salt solution of another metal.
   For example, since zinc is higher in the list than copper, zinc can replace copper in compounds of copper but copper cannot replace zinc in compounds of zinc.

   $Zn\ (s) + CuSO_4\ (aq) \rightarrow ZnSO_4\ (aq) + Cu\ (s)$

   Also, all the metals that are more reactive than hydrogen replace it in acids and liberate hydrogen gas.

   $Fe + H_2SO_4 \rightarrow FeSO_4 + H_2$

5. The following metals have been arranged in the order of their reactivity, that is, the most reactive (from the left) to the least reactive (to the right).

   sodium, calcium, magnesium, aluminium, zinc, tin, lead, copper, silver, gold
   $\longrightarrow$

6. A neutralisation reaction is a type of double displacement reaction. In this reaction, an acid and a base react to form a salt and water. These reactions are exothermic. It is called a neutralisation reaction because the solution after the reaction is neutral.
   The general representation of this type of reaction is:

   acid + base $\rightarrow$ salt + water

Examples:

(i) When a base such as sodium hydroxide is added to hydrochloric acid, neutral sodium chloride and water are formed.

$HCl\ (aq) + NaOH\ (aq) \rightarrow NaCl\ (aq) + H_2O$

(ii) When potassium hydroxide (a base) is added to sulphuric acid, potassium sulphate (neutral salt) and water are formed.

$2KOH\ (aq) + H_2SO_4\ (aq) \rightarrow K_2SO_4\ (aq) + 2H_2O$

7. Metal oxides can be prepared in the following ways: by reaction with oxygen and thermal decomposition.

Reaction with oxygen:

Some metals form an oxide when they react with oxygen with or without heating.

$$\underset{\text{sodium}}{4Na} + \underset{\text{oxygen}}{O_2} \rightarrow \underset{\text{sodium oxide}}{2Na_2O}$$

Copper reacts with oxygen on heating to form copper(II) oxide.

$$\underset{\text{copper}}{2Cu} + \underset{\text{oxygen}}{O_2} \xrightarrow{\text{heat}} \underset{\text{copper(II) oxide}}{2CuO}$$

Thermal decomposition:

When the carbonates, nitrates or hydroxides of certain metals are heated, the compounds decompose to form oxides.

Carbonates give the oxide and carbon dioxide.

$$\underset{\substack{\text{copper(II)}\\\text{carbonate}}}{CuCO_3} \xrightarrow{\text{heat}} \underset{\substack{\text{copper(II)}\\\text{oxide}}}{CuO} + \underset{\substack{\text{carbon}\\\text{dioxide}}}{CO_2}$$

Nitrates give oxides, nitrogen dioxide and oxygen.

$$\underset{\substack{\text{magnesium}\\\text{nitrate}}}{2Mg(NO_3)_2} \xrightarrow{\text{heat}} \underset{\substack{\text{magnesium}\\\text{oxide}}}{2MgO} + \underset{\substack{\text{nitrogen}\\\text{dioxide}}}{4NO_2} + \underset{\text{oxygen}}{O_2}$$

$$\underset{\substack{\text{zinc}\\\text{nitrate}}}{2Zn(NO_3)_2} \xrightarrow{\text{heat}} \underset{\substack{\text{zinc}\\\text{oxide}}}{2ZnO} + \underset{\substack{\text{nitrogen}\\\text{dioxide}}}{4NO_2} + \underset{\text{oxygen}}{O_2}$$

Hydroxides give the oxide and water.

$$\underset{\substack{\text{iron(III)}\\\text{hydroxide}}}{2Fe(OH)_3} \xrightarrow{\text{heat}} \underset{\substack{\text{iron(III)}\\\text{oxide}}}{Fe_2O_3} + \underset{\text{water}}{3H_2O}$$

Think and Answer

Calcium carbonate was heated in the test tube and calcium oxide was left behind in the test tube. When calcium oxide reacts with water it forms calcium hydroxide.

CHEMICAL REACTIONS

Answer the questions.

1. The reaction of barium hydroxide with ammonium chloride is endothermic. What does this mean? Write down the balanced chemical equation for this reaction.

   Ans: Some reactions absorb heat from the surroundings. The reaction vessel becomes cooler as the reaction progresses. Such reactions are known as endothermic reactions.

   The reaction of barium hydroxide with ammonium chloride is an example of such a reaction.

   The equation is:

   $$\underset{\text{barium hydroxide}}{Ba(OH)_2\,(s)} + \underset{\text{ammonium chloride}}{2NH_4Cl\,(s)} \rightarrow \underset{\text{barium chloride}}{BaCl_2(s)} + \underset{\text{ammonia}}{2NH_3\,(g)} + \underset{\text{water}}{2H_2O}$$

2. What are photochemical reactions? Give examples.

   Ans: Photochemical reactions are reactions that need light to occur. Examples include photosynthesis and the formation of Vitamin D in the skin. Photography also depends on photochemical reactions.

3. What is a catalyst? Give two examples.

   Ans: A catalyst is a substance that causes a change in the rate of a reaction without undergoing any change itself.

   Examples:

   (i) Platinum acts as a catalyst in the reaction of carbon monoxide and oxygen to give carbon dioxide and increases the reaction rate.

   $$\underset{\text{carbon monoxide}}{2CO} + \underset{\text{oxygen}}{O_2} \xrightarrow{\text{platinum}} \underset{\text{carbon dioxide}}{2CO_2}$$

   (ii) Enzymes are examples of biological catalysts. Enzymes in the digestive system speed up the digestive process.

4. What is a displacement reaction? Explain with an example. Write a balanced equation for your answer.

   Ans: A more reactive element displaces a less reactive element from its compound in a displacement reaction. For example, iron is more reactive than copper. When an iron nail is placed in a solution of copper sulphate, iron displaces the copper in the copper sulphate solution. We can see copper as a reddish-brown coat on the nail. Also, the blue copper sulphate solution turns into light green iron (II) sulphate solution.

   $Fe + CuSO_4 \rightarrow FeSO_4 + Cu$

5. How does an antacid tablet help people who have acidity?

   Ans: Antacid tablets contain a mild base like magnesium hydroxide. They provide relief from acidity by neutralising the hydrochloric acid in the stomach.

CHEMICAL REACTIONS

1. Write the chemical equations for the following reactions.

a. Potassium (K) and water (H_2O) react to form potassium hydroxide (KOH) and hydrogen gas (H_2).

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b. Hydrogen peroxide (H_2O_2), when exposed to light, splits into water (H_2O) and oxygen (O).

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c. Zinc metal reacts with silver nitrate ($AgNO_3$) to form zinc nitrate ($Zn(NO_3)_2$) and silver.

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2. Answer the following.

a. What is effervescence? Give an example.

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b. How does toothpaste prevent tooth decay?

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c. What is neutralisation reaction? Explain one way it is useful in daily life.

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3. Identify the type of reaction in:

a. $2Na + Br_2 \rightarrow 2NaBr$: ______________________

b. $2Al_2O_3 \rightarrow 4Al + 3O_2$: ______________________

c. $2NaBr + Cl_2 \rightarrow 2NaCl + Br_2$: ______________________

d. $Fe + S \rightarrow FeS$: ______________________

ANSWER KEY FOR THE WORKSHEET

CHEMICAL REACTIONS

1. a. $K + H_2O \rightarrow KOH + H_2$

$2K + 2H_2O \rightarrow 2KOH + H_2$

b. $2H_2O_2 (S) \rightarrow 2H_2O (l) + O_2 (g)$

hydrogen peroxide — water — oxgen

c. $Zn + 2AgNO_3 \rightarrow Zn(NO_3)_2 + 2Ag$

2. a. The evolution of gas in a chemical reaction is termed effervescence. For example, when copper reacts with concentrated nitric acid, copper nitrate, nitrogen dioxide and water are formed. Nitrogen dioxide is a pungent reddish-brown poisonous gas.

b. The bacteria in the mouth act on the sugars left behind and break them down to acids. Toothpaste contains bases to neutralise these acids.

c. The reaction between an acid and a base resulting in the formation of salt and water is called a neutralisation reaction. Farmers neutralise acidic soil by adding bases like quicklime to it.

3. a. combination b. decomposition c. displacement d. combination