9th Class Chemistry Notes Khyber Pakhtunkhwa (KPK) Full Book Notes

Complete chapter-wise notes with original practice questions

Based on the Chemistry book for Class 9 prescribed by the
Khyber Pakhtunkhwa Textbook Board, Peshawar
For all KPK Boards: Peshawar, Mardan, Swat, Abbottabad,
Bannu, Kohat, Malakand and D.I. Khan

8 chapters • Key concepts, definitions, equations and worked examples
80 original MCQs • 40 short questions • 16 long questions
with an MCQ answer key at the end

Zahid Notes — www.zahidenotes.com

About These Notes

These notes follow the eight chapters of the Chemistry book for Class 9 prescribed by the Khyber Pakhtunkhwa Textbook Board, Peshawar. The same book and syllabus are used by all eight Boards of Intermediate and Secondary Education in Khyber Pakhtunkhwa — Peshawar, Mardan, Swat, Abbottabad, Bannu, Kohat, Malakand and D.I. Khan — so one set of notes serves every KPK board.

Each chapter contains concise study notes with key concepts, definitions, chemical equations and worked examples, followed by original practice questions: 10 multiple-choice questions, 5 short questions and 2 long questions. The questions are written fresh for practice; they are not copied from the textbook or from any past paper. An MCQ answer key is given at the end of the book.

How to Use This Book

- Read the study notes of one chapter first, then try its questions without looking back.
- Learn every definition in your own words and practise writing the chemical equations.
- Work the numerical examples with a calculator, step by step, until you can do them alone.
- Revise the MCQs of all chapters together in the last week before the examination.

Contents

- Chapter 1 — Fundamentals of Chemistry
- Chapter 2 — Structure of Atoms
- Chapter 3 — Periodic Table and Periodicity of Properties
- Chapter 4 — Structure of Molecules
- Chapter 5 — Physical States of Matter
- Chapter 6 — Solutions
- Chapter 7 — Electrochemistry
- Chapter 8 — Chemical Reactivity
- MCQ Answer Key • Final Revision Checklist

> **A note on the chapter names:** the book's contents list the fifth chapter as "Physical States of Matter". A few notes websites call it "Physical Properties of Matter". The contents list is followed here. The other seven chapter names are the same in every source checked.

Paper Pattern — KPK Boards, Class 9 Chemistry

The written paper and the practical examination together carry 75 marks. The figures below are taken from official BISE documents: the theory split (Sections A, B and C) is printed on the official model papers of BISE Bannu and on the official Chemistry paper of BISE Malakand, and the practical marks are printed on the official practical date sheet of BISE Peshawar.

Part of the examination	Marks	What it contains
Section A — Objective (MCQs)	12	12 MCQs, 1 mark each; 15 minutes allowed
Section B — Short questions	32	Attempt any 8 questions out of 11; 4 marks each
Section C — Long questions	21	Attempt any 3 questions; each has two parts, 7 marks in total
Theory paper total	**65**	**Time allowed: 3 hours in total (15 minutes for Section A, then 2 hours 45 minutes for Sections B and C)**
Practical examination	10	Laboratory work, practical notebook and viva voce
Subject total	**75**	**Theory 65 + Practical 10**

Passing marks: 33% of the total is the usual pass standard for a subject; confirm the current rule on your own roll-number slip or board notification.

How to Prepare for the Practical (10 Marks)

- Keep the practical notebook complete, neat and signed; it is part of the practical marks.
- Practise the experiments listed in the book's practical section, especially tests for gases, preparation of solutions of known concentration, and simple electrolysis.
- Learn the observations and the result of every experiment by heart, and be ready to explain the apparatus and the precautions in the viva voce.

> **Pattern warning:** a graphic shared on social media in early 2026 claimed a new 30-mark practical for Physics and Chemistry. No BISE document found for this book gives 30 practical marks — the official BISE Peshawar practical date sheet for the SSC Annual-II Examination 2026 prints Chemistry practical as 10 marks, and the same 10-mark figure appears for the other KPK boards. This book therefore uses 10 practical marks and 75 total marks. If your board issues a new notification, follow the notification.

Sources checked for the chapter list and pattern: the Khyber Pakhtunkhwa Textbook Board book contents as listed by eilm.com.pk (KPK Class 9 books page) and tehkals.com (Chemistry Notes 9th Class, KPK textbook); the BISE Bannu official model papers for Chemistry 9th (objective and subjective); the BISE Malakand official Chemistry 9th paper; and the BISE Peshawar official SSC Annual-II 2026 date sheet. Pattern figures are reproduced from those documents; they are not the publisher's own figures.

Chapter 1 — Fundamentals of Chemistry

1.1 What is Chemistry?

Chemistry is the branch of science that studies matter: its composition, structure, properties and the changes it undergoes. Matter is anything that has mass and occupies space. Chemistry helps us understand the air we breathe, the water we drink, the food we eat, medicines, fertilisers, fuels and the materials used in industry.

1.2 Branches of Chemistry

- **Organic chemistry** — the study of compounds of carbon, especially hydrocarbons and their derivatives. Examples: methane, ethanol, plastics, medicines.
- **Inorganic chemistry** — the study of all elements and their compounds other than organic compounds. Examples: salts, metals, acids such as sulphuric acid.
- **Physical chemistry** — the study of the physical laws and principles behind chemical changes: energy, rates of reaction, equilibrium and the states of matter.
- **Analytical chemistry** — the qualitative and quantitative analysis of matter: finding what is present and how much of it is present.
- **Biochemistry** — the chemistry of living things: proteins, carbohydrates, fats, enzymes and vitamins.
- **Industrial chemistry** — the large-scale manufacture of chemicals such as fertilisers, cement, soap and steel.
- **Nuclear chemistry** — the study of the nucleus, radioactivity and nuclear reactions.
- **Environmental chemistry** — the study of chemical substances in air, water and soil and their effect on the environment.

1.3 Muslim Scientists and Their Contributions

- **Jabir ibn Hayyan (721–803 AD)**, called the father of chemistry, prepared many acids, described distillation and crystallisation, and wrote about the preparation of chemicals.
- **Al-Razi (864–930 AD)** classified substances into vegetable, animal and mineral, described the preparation of alcohol and sulphuric acid, and used chemical methods in medicine.
- **Al-Biruni** measured the densities of metals and minerals and worked on the properties of matter.
- **Ibn al-Haytham** worked on optics and the scientific method of observation and experiment.

1.4 Matter: Elements, Compounds and Mixtures

- **Element:** a pure substance made of only one kind of atom, which cannot be broken down into simpler substances by chemical means. Examples: hydrogen, oxygen, iron, gold. There are 118 known elements.
- **Compound:** a pure substance formed when two or more elements combine chemically in a fixed ratio. It has properties different from its elements. Examples: water (H_2O), carbon dioxide (CO_2), common salt (NaCl).
- **Mixture:** a combination of two or more substances mixed together without a chemical reaction, in any ratio, keeping their own properties. Examples: air, sea water, brass. Mixtures can be separated by physical methods.

Mixture type	Meaning	Example
Homogeneous	Uniform composition throughout; one phase	Salt solution, air, brass
Heterogeneous	Non-uniform composition; two or more phases visible	Sand in water, oil in water

1.5 Atom, Molecule, Ion and Formula

- **Atom:** the smallest particle of an element that takes part in a chemical reaction and keeps the properties of the element.
- **Molecule:** two or more atoms joined by a chemical bond. A molecule of an element has atoms of the same kind (H_2, O_2); a molecule of a compound has atoms of different kinds (H_2O, CO_2, NH_3).
- **Ion:** an atom or group of atoms with an electric charge, formed by losing or gaining electrons. A cation is positively charged (Na^+, Ca^{2+}); an anion is negatively charged (Cl^-, SO_4^{2-}).
- **Atomic number (Z):** the number of protons in the nucleus of an atom. It decides the identity of the element.
- **Mass number (A):** the total number of protons and neutrons in the nucleus. A = Z + number of neutrons.
- **Isotopes:** atoms of the same element with the same atomic number but different mass numbers, because they have different numbers of neutrons. Examples: hydrogen has ^1H, ^2H and ^3H; carbon has ^{12}C and ^{14}C; chlorine has ^{35}Cl and ^{37}Cl.

1.6 Chemical Symbols, Formulae and Valency

A **chemical symbol** is a short way of writing the name of an element, usually one or two letters: H for hydrogen, O for oxygen, Na for sodium (from Latin *natrium*), Fe for iron (from *ferrum*). A **chemical formula** shows the number and kind of atoms in a molecule or compound, for example H_2SO_4 for sulphuric acid.

Valency is the combining capacity of an element — the number of electrons it loses, gains or shares to complete its outer shell. Sodium has valency 1, calcium 2, aluminium 3, chlorine 1, oxygen 2. To write a formula, cross the valencies: calcium (2) and chlorine (1) give $CaCl_2$; aluminium (3) and oxygen (2) give Al_2O_3.

Compound	Formula	How it is formed
Sodium chloride	NaCl	$Na^+ + Cl^-$
Calcium carbonate	$CaCO_3$	$Ca^{2+} + CO_3^{2-}$
Ammonium sulphate	$(NH_4)_2SO_4$	$2 \times NH_4^+ + SO_4^{2-}$
Aluminium oxide	Al_2O_3	$2 \times Al^{3+} + 3 \times O^{2-}$

1.7 Atomic Mass, Molecular Mass and the Mole

- **Relative atomic mass:** the average mass of an atom of an element compared with one-twelfth of the mass of a carbon-12 atom. Example: hydrogen 1.008, oxygen 16.0, carbon 12.0.
- **Molecular mass:** the sum of the atomic masses of all the atoms in a molecule. For water: 2 × 1 + 16 = 18. For carbon dioxide: 12 + 2 × 16 = 44.
- **Formula mass:** the same idea for ionic compounds, which have no separate molecules. For NaCl: 23 + 35.5 = 58.5.
- **Mole:** the amount of a substance that contains 6.02×10^{23} particles (Avogadro's number). One mole of any substance has a mass in grams equal to its atomic or molecular mass.
- **Avogadro's number (N_A):** 6.02×10^{23} particles per mole.

Worked example 1: How many moles are there in 88 g of carbon dioxide? Molecular mass of CO_2 = 12 + 32 = 44. Moles = mass ÷ molar mass = 88 ÷ 44 = 2 moles.

Worked example 2: How many molecules are there in 2 moles of hydrogen gas? Molecules = $2 \times 6.02 \times 10^{23} = 1.204 \times 10^{24}$ molecules.

1.8 Empirical and Molecular Formulae

The **empirical formula** is the simplest whole-number ratio of the atoms in a compound. The **molecular formula** shows the actual number of atoms in one molecule. For glucose, the empirical formula is CH_2O and the molecular formula is $C_6H_{12}O_6$. Molecular formula = (empirical formula) × n, where n = molecular mass ÷ empirical formula mass.

Percentage composition: the mass of each element in 100 g of the compound. For water: %H = (2 ÷ 18) × 100 = 11.1%; %O = (16 ÷ 18) × 100 = 88.9%.

> **Key terms to memorise:** matter, element, compound, mixture, atom, molecule, ion, atomic number, mass number, isotope, valency, mole, Avogadro's number, empirical formula, molecular formula, percentage composition.

Chapter 1 — Practice Questions

Multiple Choice Questions — choose one correct answer.

1. Which branch of chemistry deals with the compounds of carbon?

(a) Inorganic chemistry (b) Organic chemistry (c) Physical chemistry (d) Analytical chemistry

2. Which of the following is a compound?

(a) Air (b) Brass (c) Water (d) Gold

3. The atomic number of an element is the number of:

(a) Neutrons (b) Protons (c) Protons and neutrons (d) Electrons and neutrons

4. Isotopes of the same element differ in the number of:

(a) Protons (b) Electrons (c) Neutrons (d) Shells

5. The molecular mass of sulphuric acid (H_2SO_4) is:

(a) 49 (b) 98 (c) 96 (d) 82

6. One mole of oxygen gas (O_2) contains:

(a) 6.02×10^{23} atoms (b) 6.02×10^{23} molecules (c) 32 atoms (d) 16 molecules

7. The valency of aluminium in Al_2O_3 is:

(a) 1 (b) 2 (c) 3 (d) 6

8. Which of the following is a homogeneous mixture?

(a) Sand in water (b) Oil in water (c) Salt solution (d) Iron filings in sand

9. The empirical formula of hydrogen peroxide (H_2O_2) is:

(a) H_2O (b) HO (c) H_2O_2 (d) H_4O_2

10. The number of moles in 20 g of calcium (atomic mass 40) is:

(a) 0.5 (b) 2 (c) 20 (d) 40

Chapter 1 — Practice Questions (continued)

Short Questions

1. Define chemistry and name any four of its branches with one example of what each studies.

2. What is the difference between an element and a compound? Give two examples of each.

3. Define atomic number and mass number. An atom has 11 protons and 12 neutrons. What are its atomic number and mass number?

4. What are isotopes? Give the isotopes of hydrogen and chlorine.

5. Calculate the molecular mass of (a) ammonia (NH_3) and (b) calcium carbonate ($CaCO_3$). (Atomic masses: H = 1, N = 14, Ca = 40, C = 12, O = 16.)

Long Questions

1. (a) Explain the difference between homogeneous and heterogeneous mixtures with two examples of each and one method of separating each type. (b) How many molecules are present in 4.4 g of carbon dioxide? (Molecular mass of CO_2 = 44.)

2. (a) Describe the contributions of Jabir ibn Hayyan and Al-Razi to chemistry. (b) A compound contains 40% carbon, 6.7% hydrogen and 53.3% oxygen by mass, and its molecular mass is 180. Find its empirical and molecular formulae.

Chapter 2 — Structure of Atoms

2.1 Dalton's Atomic Theory

John Dalton (1808) proposed that all matter is made of tiny indivisible particles called atoms; atoms of the same element are alike in mass and size, while atoms of different elements differ; atoms combine in simple whole-number ratios to form compounds; and in a chemical reaction atoms are rearranged but never created or destroyed. Later work showed that atoms are divisible into smaller particles and that atoms of the same element can differ in mass (isotopes), so the theory was modified, but its main ideas still stand.

2.2 The Electron, Proton and Neutron

Particle	Symbol	Charge	Relative mass	Discovered by
Electron	e^-	−1	1/1836 (almost zero)	J. J. Thomson (1897)
Proton	p^+	+1	1	E. Goldstein / Rutherford
Neutron	n	0	1	James Chadwick (1932)

Thomson discovered the electron with cathode-ray (discharge) tubes and suggested the "plum pudding" model: the atom is a positively charged ball with electrons embedded in it, like plums in a pudding.

2.3 Rutherford's Model and Its Defects

Rutherford bombarded a thin gold foil with alpha particles. Most particles passed straight through, a few were deflected by small angles, and about one in 8,000 bounced straight back. He concluded that the atom is mostly empty space; its positive charge and almost all its mass are concentrated in a tiny central nucleus; and the electrons move around the nucleus. The model could not explain why the circling electrons do not lose energy and fall into the nucleus, and it could not explain the line spectrum of hydrogen.

2.4 Bohr's Atomic Model

Neils Bohr (1913) proposed that electrons revolve around the nucleus in fixed circular orbits or shells of definite energy, called energy levels. While an electron stays in one orbit it neither loses nor gains energy. Energy is absorbed when an electron jumps to a higher shell and emitted as light when it falls to a lower shell. The shells are named K, L, M, N from the nucleus outwards, and the maximum number of electrons in a shell is $2n^2$, where n is the shell number: K holds 2, L holds 8, M holds 18 and N holds 32.

2.5 Electronic Configuration

The **electronic configuration** is the arrangement of electrons in the shells of an atom. Electrons fill the lowest energy shell first. Examples: hydrogen (1) = 1; carbon (6) = 2, 4; oxygen (8) = 2, 6; sodium (11) = 2, 8, 1; magnesium (12) = 2, 8, 2; chlorine (17) = 2, 8, 7; argon (18) = 2, 8, 8; calcium (20) = 2, 8, 8, 2. The electrons in the outermost shell are the **valence electrons**; they decide how the atom reacts. Elements with 8 electrons in the outer shell (or 2 in the first shell, as in helium) are stable and do not react easily.

2.6 Shells, Sub-shells and Isotopes

- The atomic number Z is the number of protons; in a neutral atom it also equals the number of electrons.
- The mass number A = protons + neutrons, so neutrons = A − Z. For example, ^{23}Na has 11 protons, 11 electrons and 12 neutrons.
- Isotopes have the same chemical properties because they have the same number of electrons, but different physical properties because their masses differ. Carbon-14 is used for dating old organic material; iodine-131 is used in treating thyroid disease; cobalt-60 is used in cancer treatment.
- Atoms become ions by losing or gaining electrons: metals usually lose electrons to form cations ($Na \rightarrow Na^+ + e^-$), and non-metals usually gain electrons to form anions ($Cl + e^- \rightarrow Cl^-$).

> **Remember:** the nucleus contains protons and neutrons; almost all the mass of the atom is in the nucleus; the atom is mostly empty space; electrons decide chemical behaviour.

Chapter 2 — Practice Questions

Multiple Choice Questions — choose one correct answer.

1. Who discovered the electron?

(a) Rutherford (b) Bohr (c) J. J. Thomson (d) Chadwick

2. The neutron was discovered by:

(a) Thomson (b) Chadwick (c) Dalton (d) Goldstein

3. The maximum number of electrons in the M shell is:

(a) 2 (b) 8 (c) 18 (d) 32

4. The electronic configuration of chlorine (atomic number 17) is:

(a) 2, 8, 8 (b) 2, 8, 7 (c) 2, 7, 8 (d) 2, 8, 6

5. An atom with 12 protons and 12 neutrons has mass number:

(a) 12 (b) 24 (c) 6 (d) 36

6. In Rutherford's experiment, most alpha particles:

(a) Bounced back (b) Were deflected at large angles (c) Passed straight through the foil (d) Were absorbed

7. Which of the following is an isotope of carbon?

(a) ^{14}N (b) ^{14}C (c) ^{16}O (d) ^{23}Na

8. The number of valence electrons in magnesium (atomic number 12) is:

(a) 1 (b) 2 (c) 8 (d) 12

9. According to Bohr, an electron emits energy when it:

(a) Jumps to a higher shell (b) Falls to a lower shell (c) Stays in the same shell (d) Leaves the atom

10. Which particle has no electric charge?

(a) Electron (b) Proton (c) Neutron (d) Alpha particle

Chapter 2 — Practice Questions (continued)

Short Questions

1. State the main points of Dalton's atomic theory and name two ways in which it was later modified.

2. What did Rutherford conclude from the alpha-particle scattering experiment? Give three conclusions.

3. Write the electronic configurations of (a) oxygen (8), (b) sodium (11), (c) argon (18) and (d) calcium (20).

4. An atom of sulphur has atomic number 16 and mass number 32. How many protons, neutrons and electrons does it have?

5. Why do isotopes of the same element have the same chemical properties but different physical properties? Give two uses of isotopes.

Long Questions

1. (a) Describe Rutherford's alpha-particle scattering experiment, its observations and the conclusions drawn from it. (b) State two defects of Rutherford's model and explain how Bohr's model removed them.

2. (a) Explain Bohr's atomic model and the rule $2n^2$ for the maximum electrons in a shell. (b) Write the electronic configurations of carbon, nitrogen, neon and potassium (atomic numbers 6, 7, 10 and 19) and state the number of valence electrons in each.

Chapter 3 — Periodic Table and Periodicity of Properties

3.1 The Modern Periodic Table

The modern periodic law states that the properties of elements are a periodic function of their atomic numbers. Elements are arranged in order of increasing atomic number in the periodic table. A horizontal row is a **period** (there are 7 periods); a vertical column is a **group** (there are 18 groups). Elements in the same group have the same number of valence electrons and similar chemical properties.

Block	Groups	Outer-shell electrons	Examples
s-block	Groups 1 and 2	1–2	Na, K, Ca, Mg
p-block	Groups 13–18	3–8	C, N, O, Cl, Ar
d-block	Groups 3–12	Transition metals	Fe, Cu, Zn
f-block	Lanthanides, actinides	Inner shells filling	Ce, U

3.2 Groups and Their Names

- **Group 1 — Alkali metals:** Li, Na, K, Rb, Cs. Very reactive metals; they have one valence electron and form +1 ions. They react violently with water.
- **Group 2 — Alkaline earth metals:** Be, Mg, Ca, Sr, Ba. They have two valence electrons and form +2 ions.
- **Groups 3–12 — Transition metals:** Fe, Cu, Zn, Ni. Hard, strong metals with high melting points; used in construction, coins and machinery.
- **Group 17 — Halogens:** F, Cl, Br, I. Very reactive non-metals with seven valence electrons; they form −1 ions.
- **Group 18 — Noble gases:** He, Ne, Ar, Kr, Xe. Their outer shells are complete, so they are almost unreactive and exist as single atoms.

3.3 Periodicity of Properties

Atomic radius: the distance from the nucleus to the outermost electron shell. It **decreases** across a period (more protons pull the electrons closer) and **increases** down a group (a new shell is added each time).

Ionisation energy: the energy needed to remove the outermost electron from a gaseous atom. It **increases** across a period (electrons are held more tightly) and **decreases** down a group (the outer electron is farther from the nucleus and shielded by inner shells). Sodium has low ionisation energy; neon has high ionisation energy.

Electron affinity: the energy released when a gaseous atom gains an electron. Halogens have high electron affinity because they need only one electron to complete their shell.

Electronegativity: the power of an atom in a molecule to attract the shared pair of electrons towards itself. It **increases** across a period and **decreases** down a group. Fluorine is the most electronegative element; caesium is among the least.

Shielding effect: the inner-shell electrons reduce the pull of the nucleus on the outer-shell electrons. Shielding **increases** down a group as more shells are added, which is why atomic size grows and ionisation energy falls down a group.

3.4 Metals, Non-metals and Metalloids

- **Metals** (left side of the table) are shiny, good conductors of heat and electricity, malleable and ductile, and they lose electrons to form positive ions.
- **Non-metals** (right side) are dull, poor conductors (except graphite), brittle, and they gain or share electrons.
- **Metalloids** (boron, silicon, germanium, arsenic) lie along the zig-zag line between metals and non-metals and show properties of both. Silicon is used in computer chips.
- **Trends:** metallic character increases down a group and decreases across a period; non-metallic character does the opposite.

> **How to find a position:** the period number equals the number of shells in the atom; the group number (for groups 1, 2 and 13–18) equals the number of valence electrons. Sodium (2, 8, 1) is in period 3, group 1. Chlorine (2, 8, 7) is in period 3, group 17.

Chapter 3 — Practice Questions

Multiple Choice Questions — choose one correct answer.

1. The modern periodic law is based on the atomic:
(a) Mass (b) Number (c) Radius (d) Volume

2. How many groups are there in the modern periodic table?
(a) 7 (b) 8 (c) 18 (d) 32

3. Which group contains the noble gases?
(a) Group 1 (b) Group 2 (c) Group 17 (d) Group 18

4. Atomic radius increases down a group because:
(a) Nuclear charge increases (b) A new shell is added (c) Electrons are lost (d) Protons decrease

5. Which element has the highest electronegativity?
(a) Oxygen (b) Chlorine (c) Fluorine (d) Sodium

6. Ionisation energy generally decreases down a group due to:
(a) More protons (b) More neutrons (c) Increased shielding and larger size (d) Fewer shells

7. Which of the following is a metalloid?
(a) Iron (b) Silicon (c) Oxygen (d) Sodium

8. The element with electronic configuration 2, 8, 1 belongs to:
(a) Period 2, group 1 (b) Period 3, group 1 (c) Period 3, group 2 (d) Period 1, group 1

9. Halogens have seven electrons in their outermost shell, so they usually:
(a) Lose one electron (b) Gain one electron (c) Lose seven electrons (d) Share seven electrons

10. Which pair has similar chemical properties?
(a) Na and Mg (b) Li and Na (c) C and O (d) He and H

Chapter 3 — Practice Questions (continued)

Short Questions

1. State the modern periodic law and explain the difference between a period and a group.

2. Why does atomic radius decrease across a period but increase down a group?

3. Define ionisation energy and electronegativity. How does each change across a period?

4. What is the shielding effect? How does it explain the fall in ionisation energy down a group?

5. Magnesium has atomic number 12. Write its electronic configuration and find its period and group.

Long Questions

1. (a) Describe the main features of groups 1, 17 and 18, giving two elements and one property of each group. (b) Explain why fluorine is more reactive than iodine although both are halogens.

2. (a) Explain the trends in atomic radius, ionisation energy and electronegativity across a period and down a group, with reasons. (b) Distinguish between metals, non-metals and metalloids with two properties and one example of each.

Chapter 4 — Structure of Molecules

4.1 Why Do Atoms Combine?

Atoms combine to become stable. Most atoms have incomplete outer shells, so they lose, gain or share electrons to get the stable electronic configuration of a noble gas (usually 8 electrons in the outer shell). This is the **octet rule**. Hydrogen and helium are stable with 2 electrons (the **duplet rule**). The force that holds atoms together in a molecule or compound is a **chemical bond**.

4.2 Ionic (Electrovalent) Bond

An ionic bond is formed by the **complete transfer** of one or more electrons from a metal atom to a non-metal atom, producing oppositely charged ions that attract each other. Example: sodium (2, 8, 1) gives one electron to chlorine (2, 8, 7), forming Na^+ and Cl^-; the two ions attract and form sodium chloride. Magnesium gives two electrons to two chlorine atoms to form $MgCl_2$.

Properties of ionic compounds: high melting and boiling points; hard and brittle crystals; they conduct electricity when molten or dissolved in water, but not as solids; they dissolve in water but not in organic solvents.

4.3 Covalent Bond

A covalent bond is formed by the **mutual sharing** of one or more pairs of electrons between two non-metal atoms. A single covalent bond shares one pair (H—H, Cl—Cl); a double bond shares two pairs (O=O); a triple bond shares three pairs (N≡N). In methane, carbon shares four pairs of electrons with four hydrogen atoms; in water, oxygen shares one pair with each hydrogen atom.

Properties of covalent compounds: low melting and boiling points; they are usually gases, liquids or soft solids; they do not conduct electricity; they dissolve in organic solvents but usually not in water.

Property	Ionic compounds	Covalent compounds
Melting point	High	Low
Electrical conduction	Molten or in solution	Usually none
Solubility in water	Usually soluble	Usually insoluble
Example	NaCl, CaO	H_2O, CO_2, CH_4

4.4 Polar and Non-polar Covalent Bonds

When two identical atoms share electrons (H_2, O_2, N_2), the shared pair is exactly in the middle and the bond is **non-polar**. When two different atoms share electrons (HCl, H_2O), the more electronegative atom pulls the shared pair towards itself, so one end becomes slightly negative and the other slightly positive: the bond is **polar**. Water is a polar molecule, which is why it dissolves ionic compounds so well.

4.5 Metallic Bond and Intermolecular Forces

In a metal, the atoms lose their valence electrons, which move freely through the whole piece of metal as a "sea of electrons". The attraction between the positive metal ions and this sea of electrons is the **metallic bond**. It explains why metals conduct electricity and heat, and why they are malleable and ductile.

Intermolecular forces are the weak attractions between separate molecules. **Hydrogen bonding** is a strong type of intermolecular force found in water, where a hydrogen atom bonded to oxygen, nitrogen or fluorine is attracted to a lone pair on another molecule. It explains why water has an unusually high boiling point. **Dipole-dipole forces** act between polar molecules, and **London (dispersion) forces** act between all molecules.

4.6 Shapes of Molecules

- **Water (H_2O):** bent (angular) shape because oxygen has two lone pairs that push the two O—H bonds together.
- **Ammonia (NH_3):** pyramidal shape because nitrogen has one lone pair.
- **Methane (CH_4):** tetrahedral shape, with four C—H bonds pointing to the corners of a tetrahedron.
- **Carbon dioxide (CO_2):** linear shape, with the two C=O bonds in a straight line.
- **BCl_3:** triangular planar shape.

Chapter 4 — Practice Questions

Multiple Choice Questions — choose one correct answer.

1. An ionic bond is formed by the:

(a) Sharing of electrons (b) Transfer of electrons (c) Sharing of protons (d) Transfer of neutrons

2. Which compound has a covalent bond?

(a) NaCl (b) MgO (c) HCl (d) CaF_2

3. The shape of a methane molecule is:

(a) Linear (b) Bent (c) Tetrahedral (d) Pyramidal

4. Which of the following conducts electricity in the molten state?

(a) Sugar (b) Sodium chloride (c) Kerosene oil (d) Diamond

5. A triple covalent bond is present in:

(a) O_2 (b) N_2 (c) Cl_2 (d) H_2O

6. Water is a polar molecule because:

(a) It has ionic bonds (b) Oxygen attracts the shared electrons more strongly (c) It has no lone pairs (d) It is linear

7. The shape of an ammonia molecule is:

(a) Tetrahedral (b) Linear (c) Pyramidal (d) Triangular planar

8. Hydrogen bonding is strongest in:

(a) CH_4 (b) H_2O (c) H_2 (d) CO_2

9. Which property belongs to ionic compounds?

(a) Low melting point (b) No conduction when molten (c) High melting point (d) Insoluble in water

10. The metallic bond is the attraction between metal ions and:

(a) Neutrons (b) A sea of free electrons (c) Protons (d) Other metal ions only

Chapter 4 — Practice Questions (continued)

Short Questions

1. What is the octet rule? Why do atoms form chemical bonds?

2. Explain the formation of sodium chloride by electron transfer, with dot-and-cross diagrams described in words.

3. Give three differences between ionic and covalent compounds.

4. What is a polar covalent bond? Explain why HCl is polar but Cl_2 is not.

5. What is hydrogen bonding? Explain how it affects the boiling point of water.

Long Questions

1. (a) Describe the formation of ionic bonds in NaCl and $MgCl_2$. (b) Explain four properties of ionic compounds with reasons.

2. (a) Describe covalent bonding in H_2, O_2 and N_2 and state how many pairs of electrons each shares. (b) Explain the shapes of water, ammonia and methane and the role of lone pairs in each.

Chapter 5 — Physical States of Matter

5.1 The Three States of Matter

Property	Solid	Liquid	Gas
Shape	Fixed	Takes the shape of the container	Takes the shape and volume of the container
Volume	Fixed	Fixed	Changes with the container
Particles	Closely packed, vibrate in place	Close but can slide past each other	Far apart, move freely and fast
Compressibility	Almost none	Almost none	High
Density	High	High	Low

Matter changes state when it gains or loses heat: melting (solid → liquid), freezing (liquid → solid), boiling or evaporation (liquid → gas), condensation (gas → liquid) and sublimation (solid → gas, for example dry ice and camphor). A fourth state, **plasma**, is a hot ionised gas found in the sun, stars and lightning.

5.2 Kinetic Molecular Theory

- All matter is made of tiny particles (atoms or molecules) that are always moving.
- The particles in a gas move in straight lines in all directions and collide with each other and with the walls of the container; these collisions cause gas pressure.
- The average kinetic energy of the particles increases with temperature, so gases expand when heated.
- There is almost no attraction between gas particles, while the attraction is strong in solids and moderate in liquids.

5.3 Gas Laws

Boyle's law: at constant temperature, the volume of a fixed mass of gas is inversely proportional to its pressure. $P_1V_1 = P_2V_2$. If the pressure is doubled, the volume becomes half.

Charles's law: at constant pressure, the volume of a fixed mass of gas is directly proportional to its absolute temperature (in kelvin). $V_1/T_1 = V_2/T_2$. Temperature in kelvin = temperature in °C + 273.

Avogadro's law: at the same temperature and pressure, equal volumes of all gases contain the same number of molecules. One mole of any gas occupies 22.4 dm^3 (22,400 cm^3) at standard temperature and pressure (STP: 0 °C and 1 atmosphere).

General gas equation: $P_1V_1/T_1 = P_2V_2/T_2$. For an ideal gas, $PV = nRT$, where n is the number of moles and R is the gas constant.

Worked example: A gas occupies 200 cm^3 at 300 K. What is its volume at 450 K at constant pressure? $V_2 = V_1 \times T_2 \div T_1 = 200 \times 450 \div 300 = 300\ cm^3$.

5.4 Liquids: Evaporation, Vapour Pressure and Boiling

- **Evaporation** takes place at the surface of a liquid at any temperature. It cools the liquid, increases with temperature, wind and surface area, and decreases with humidity.
- **Vapour pressure** is the pressure exerted by the vapour in equilibrium with the liquid. It increases with temperature.
- **Boiling point** is the temperature at which the vapour pressure of the liquid equals the outside atmospheric pressure. Water boils at 100 °C at sea level, but at a lower temperature on a mountain because the pressure is lower.
- **Melting point** is the temperature at which a solid changes into a liquid. Ice melts at 0 °C.

5.5 Solids: Crystalline and Amorphous

- **Crystalline solids** have a regular, repeating arrangement of particles and a sharp melting point. Examples: common salt, diamond, ice, quartz.
- **Amorphous solids** have no regular arrangement and soften over a range of temperature. Examples: glass, rubber, plastic.
- **Allotropy:** the same element existing in different forms with different properties. Carbon exists as diamond (hard, used for cutting), graphite (soft, conducts electricity, used in pencils) and fullerene (C_{60}); sulphur and phosphorus also show allotropy.

Chapter 5 — Practice Questions

Multiple Choice Questions — choose one correct answer.

1. Which state of matter has a fixed volume but no fixed shape?

(a) Solid (b) Liquid (c) Gas (d) Plasma

2. According to Boyle's law, if the pressure of a gas is doubled at constant temperature, its volume becomes:

(a) Double (b) Half (c) Four times (d) Unchanged

3. The volume of one mole of a gas at STP is:

(a) 2.24 dm^3 (b) 22.4 dm^3 (c) 224 dm^3 (d) 11.2 dm^3

4. Which of the following is an amorphous solid?

(a) Diamond (b) Common salt (c) Glass (d) Ice

5. Evaporation takes place:

(a) Only at the boiling point (b) At the surface at any temperature (c) Only in a closed container (d) Only from solids

6. A gas occupies 100 cm^3 at 273 K. Its volume at 546 K at constant pressure is:

(a) 50 cm^3 (b) 100 cm^3 (c) 200 cm^3 (d) 273 cm^3

7. The boiling point of water decreases on a mountain because:

(a) Temperature is higher (b) Atmospheric pressure is lower (c) Water is impure (d) Air is dry

8. Which allotrope of carbon conducts electricity?

(a) Diamond (b) Graphite (c) Fullerene (d) Coal

9. Gas pressure is caused by:

(a) The weight of the gas (b) Collisions of gas particles with the walls (c) Attraction between particles (d) The temperature of the container

10. The change of a solid directly into a gas is called:

(a) Melting (b) Condensation (c) Sublimation (d) Freezing

Chapter 5 — Practice Questions (continued)

Short Questions

1. Give three differences between solids, liquids and gases in terms of particle arrangement and movement.

2. State Boyle's law and Charles's law, and write the formula for each.

3. A gas occupies 300 cm^3 at 27 °C. What volume will it occupy at 127 °C at constant pressure?

4. What is vapour pressure? Why does water boil at a lower temperature on a mountain?

5. Distinguish between crystalline and amorphous solids with two examples of each.

Long Questions

1. (a) State the kinetic molecular theory of gases and explain gas pressure with it. (b) A gas at 2 atm pressure occupies 400 cm^3. What will its volume be at 4 atm at constant temperature? What law did you use?

2. (a) Explain evaporation, boiling point and melting point, and state four factors that affect the rate of evaporation. (b) Describe the allotropes of carbon (diamond, graphite and fullerene) with their properties and uses.

Chapter 6 — Solutions

6.1 What is a Solution?

A **solution** is a homogeneous mixture of two or more substances. The substance that dissolves is the **solute**; the substance in which it dissolves is the **solvent**. Water is the most common solvent and is called the universal solvent. A solution has only one phase, its particles are too small to be seen, and the solute cannot be separated by filtration.

Type of solution	Solute	Solvent	Example
Gas in gas	Gas	Gas	Air (oxygen in nitrogen)
Gas in liquid	Gas	Liquid	Carbon dioxide in water (soda water)
Liquid in liquid	Liquid	Liquid	Alcohol in water
Solid in liquid	Solid	Liquid	Salt in water
Solid in solid	Solid	Solid	Brass (zinc in copper)

6.2 Types of Solutions by Concentration

- **Unsaturated solution:** it can still dissolve more solute at that temperature.
- **Saturated solution:** it contains the maximum amount of solute that can dissolve at that temperature; any extra solute remains undissolved.
- **Supersaturated solution:** it contains more solute than a saturated solution at that temperature, made by cooling a hot saturated solution carefully; it is unstable and the extra solute crystallises out easily.
- **Dilute solution:** a small amount of solute in a large amount of solvent. **Concentrated solution:** a large amount of solute in a small amount of solvent.

6.3 Solubility and Its Factors

Solubility is the maximum mass of solute, in grams, that dissolves in 100 g of solvent at a given temperature to form a saturated solution. It depends on the nature of the solute and solvent, and on temperature.

- **Temperature:** the solubility of most solids in water increases with temperature (for example, sugar and potassium nitrate). The solubility of gases in water decreases with temperature.
- **Pressure:** pressure has almost no effect on the solubility of solids and liquids, but the solubility of a gas increases with pressure (Henry's law) — this is why soda bottles are sealed under pressure.
- **"Like dissolves like":** polar solutes dissolve in polar solvents (salt in water) and non-polar solutes dissolve in non-polar solvents (grease in petrol). Benzene does not dissolve in water but dissolves in alcohol.

6.4 Concentration of a Solution

Percentage by mass (% m/m): mass of solute ÷ mass of solution × 100. **Percentage by volume (% v/v):** volume of solute ÷ volume of solution × 100.

Molarity (M): the number of moles of solute dissolved in one cubic decimetre (1 dm^3 = 1000 cm^3) of solution. Molarity = moles of solute ÷ volume of solution in dm^3.

Worked example 1: 10 g of salt is dissolved in 90 g of water. % m/m = 10 ÷ (10 + 90) × 100 = 10%.

Worked example 2: What is the molarity of a solution made by dissolving 4 g of sodium hydroxide (molar mass 40) in 500 cm^3 of solution? Moles = 4 ÷ 40 = 0.1. Volume = 0.5 dm^3. Molarity = 0.1 ÷ 0.5 = 0.2 mol/dm^3.

6.5 Colloids, Suspensions and the Tyndall Effect

- **True solution:** particles smaller than 1 nm; clear; solute does not settle; light passes through without scattering.
- **Colloid:** particles between 1 nm and 1000 nm; the mixture looks homogeneous but scatters a beam of light (the **Tyndall effect**). Examples: milk, fog, smoke, jelly, blood, mist. Colloids include aerosols (fog, smoke), emulsions (milk, mayonnaise), sols (paint, ink) and gels (jelly).
- **Suspension:** particles larger than 1000 nm; heterogeneous; the solid settles on standing and can be separated by filtration. Examples: muddy water, chalk in water.

Chapter 6 — Practice Questions

Multiple Choice Questions — choose one correct answer.

1. In a salt solution, the solvent is:

(a) Salt (b) Water (c) Both (d) Neither

2. A saturated solution is one that:

(a) Can dissolve more solute (b) Contains the maximum solute at that temperature (c) Has no solute (d) Is always hot

3. The solubility of a gas in water decreases when temperature:

(a) Decreases (b) Increases (c) Stays the same (d) Reaches 0 °C

4. Which of the following is a colloid?

(a) Salt solution (b) Milk (c) Muddy water (d) Sugar solution

5. The Tyndall effect is shown by:

(a) True solutions (b) Colloids (c) Pure liquids (d) Gases only

6. Molarity is the number of moles of solute in:

(a) 100 g of solvent (b) 1 dm^3 of solution (c) 1 kg of solution (d) 100 cm^3 of solvent

7. 20 g of sugar in 80 g of water gives a solution of:

(a) 20% (b) 25% (c) 80% (d) 10%

8. Which rule explains why benzene does not dissolve in water?

(a) Boyle's law (b) Like dissolves like (c) Avogadro's law (d) Henry's law

9. Brass is an example of a:

(a) Gas in gas solution (b) Solid in solid solution (c) Liquid in liquid solution (d) Suspension

10. In a suspension, the solid particles:

(a) Never settle (b) Settle on standing (c) Dissolve completely (d) Are smaller than 1 nm

Chapter 6 — Practice Questions (continued)

Short Questions

1. Define solution, solute and solvent, and give one example of a gas-in-liquid solution.

2. What is the difference between a saturated, an unsaturated and a supersaturated solution?

3. How does temperature affect the solubility of solids and of gases in water?

4. Calculate the percentage by mass of a solution made by dissolving 15 g of salt in 135 g of water.

5. Distinguish between a true solution, a colloid and a suspension, giving one example of each.

Long Questions

1. (a) Explain the rule "like dissolves like" with two examples. (b) A solution is made by dissolving 8 g of NaOH (molar mass 40) in water to make 400 cm^3 of solution. Calculate its molarity.

2. (a) Define solubility and explain the factors that affect it. (b) Describe colloids and suspensions, the Tyndall effect, and the four types of colloids with examples.

Chapter 7 — Electrochemistry

7.1 Oxidation and Reduction

- **Oxidation** is the loss of electrons, the gain of oxygen, or the loss of hydrogen. Example: $Na \rightarrow Na^{+} + e^{-}$; $2Mg + O_2 \rightarrow 2MgO$.
- **Reduction** is the gain of electrons, the loss of oxygen, or the gain of hydrogen. Example: $Cl_2 + 2e^{-} \rightarrow 2Cl^{-}$; $CuO + H_2 \rightarrow Cu + H_2O$.
- **Redox reaction:** oxidation and reduction always take place together. In $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$, zinc is oxidised (it loses electrons) and copper(II) ions are reduced (they gain electrons).
- **Oxidising agent:** the substance that causes oxidation and is itself reduced (it gains electrons). **Reducing agent:** the substance that causes reduction and is itself oxidised (it loses electrons). In the reaction above, $CuSO_4$ is the oxidising agent and zinc is the reducing agent.

7.2 Oxidation Number

The **oxidation number** is the charge an atom appears to have when the electrons in its bonds are counted as belonging to the more electronegative atom. Rules to remember:

- The oxidation number of an element in its free state is zero (Na, O_2, Fe = 0).
- The oxidation number of hydrogen is +1 (except in metal hydrides, where it is −1); of oxygen is −2 (except in peroxides, where it is −1).
- The oxidation number of a simple ion equals its charge ($Na^{+} = +1$, $Cl^{-} = -1$).
- The sum of the oxidation numbers in a neutral compound is zero; in a polyatomic ion it equals the charge of the ion.

Worked examples: In H_2SO_4, sulphur = 0 − (2 × +1) − (4 × −2) = +6. In $KMnO_4$, manganese = 0 − (+1) − (4 × −2) = +7. In CO_2, carbon = 0 − (2 × −2) = +4.

7.3 Electrochemical Cells

An **electrochemical cell** converts chemical energy into electrical energy. In the **Daniel cell**, a zinc rod stands in zinc sulphate solution and a copper rod stands in copper sulphate solution, and the two solutions are joined by a salt bridge.

- **Anode (negative, zinc):** oxidation takes place. $Zn \rightarrow Zn^{2+} + 2e^{-}$.
- **Cathode (positive, copper):** reduction takes place. $Cu^{2+} + 2e^{-} \rightarrow Cu$.
- Electrons flow through the outside wire from zinc to copper; the salt bridge keeps the solutions electrically neutral by letting ions move.
- Overall reaction: $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$. A Daniel cell produces about 1.1 volts.

7.4 Electrolysis

Electrolysis is the decomposition of a compound by passing an electric current through its molten state or its solution. The electrodes are usually graphite or platinum.

- **Electrolysis of water** (with a little acid added): hydrogen gas is collected at the cathode ($2H^{+} + 2e^{-} \rightarrow H_2$) and oxygen at the anode ($4OH^{-} \rightarrow 2H_2O + O_2 + 4e^{-}$). Twice as much hydrogen as oxygen is produced.
- **Electrolysis of molten sodium chloride:** sodium metal forms at the cathode ($Na^{+} + e^{-} \rightarrow Na$) and chlorine gas at the anode ($2Cl^{-} \rightarrow Cl_2 + 2e^{-}$).
- **Electrolysis of concentrated brine** gives hydrogen at the cathode, chlorine at the anode, and sodium hydroxide left in the solution.
- **Electroplating:** coating one metal with a thin layer of another metal by electrolysis. The object to be coated is the cathode and the coating metal is the anode, in a solution of a salt of the coating metal. Example: chromium plating of bicycle parts, silver plating of spoons, tin plating of iron cans.

7.5 Corrosion and Its Prevention

Corrosion is the slow eating away of a metal by air, water or chemicals. **Rusting** is the corrosion of iron: iron reacts with oxygen and water to form hydrated iron(III) oxide, $Fe_2O_3.xH_2O$. Rust is soft and crumbly, so it exposes fresh iron underneath and the damage continues.

- **Painting and oiling:** keep air and water away from the surface.
- **Galvanising:** coating iron with a layer of zinc. Zinc is more reactive than iron, so even if the coating is scratched, zinc corrodes first and protects the iron (sacrificial protection).
- **Electroplating with tin or chromium:** a protective metal layer.
- **Alloying:** mixing iron with chromium and nickel makes stainless steel, which does not rust.
- **Sacrificial anodes:** blocks of zinc or magnesium attached to ships and underground pipes corrode instead of the steel.

Chapter 7 — Practice Questions

Multiple Choice Questions — choose one correct answer.

1. Oxidation is the:

(a) Gain of electrons (b) Loss of electrons (c) Loss of oxygen (d) Gain of hydrogen

2. In the reaction $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$, the reducing agent is:

(a) $CuSO_4$ (b) Cu (c) Zn (d) $ZnSO_4$

3. The oxidation number of sulphur in H_2SO_4 is:

(a) +4 (b) +6 (c) −2 (d) +2

4. In a Daniel cell, oxidation takes place at the:

(a) Copper cathode (b) Zinc anode (c) Salt bridge (d) Copper anode

5. In the electrolysis of water, hydrogen gas is collected at the:

(a) Anode (b) Cathode (c) Salt bridge (d) Battery

6. The oxidation number of oxygen in hydrogen peroxide (H_2O_2) is:

(a) −2 (b) −1 (c) +1 (d) 0

7. Galvanising is the coating of iron with:

(a) Tin (b) Chromium (c) Zinc (d) Copper

8. The chemical formula of rust is:

(a) FeO (b) $Fe_2O_3.xH_2O$ (c) $FeSO_4$ (d) $FeCl_3$

9. In electroplating, the object to be coated is made the:

(a) Anode (b) Cathode (c) Salt bridge (d) Electrolyte

10. Which of the following is an oxidising agent?

(a) Zinc metal (b) Hydrogen gas (c) Chlorine gas (d) Sodium metal

Chapter 7 — Practice Questions (continued)

Short Questions

1. Define oxidation and reduction in terms of electrons, and give one example of each.

2. What is an oxidising agent and a reducing agent? Identify both in the reaction $CuO + H_2 \rightarrow Cu + H_2O$.

3. Calculate the oxidation number of (a) carbon in CO_2, (b) nitrogen in HNO_3 and (c) manganese in $KMnO_4$.

4. Describe the Daniel cell: label the anode, the cathode, and the reactions at each electrode.

5. What is galvanising? Explain why zinc protects iron even when the coating is scratched.

Long Questions

1. (a) Explain electrolysis with the example of molten sodium chloride, giving the reactions at both electrodes. (b) Describe the electrolysis of water and state the volumes of the gases produced.

2. (a) What is corrosion? Explain the rusting of iron and the conditions needed for it. (b) Describe four methods of preventing corrosion, explaining how each one works.

Chapter 8 — Chemical Reactivity

8.1 Metals and Non-metals

Metals are elements that lose electrons easily and form positive ions. They are usually solids (except mercury), shiny, malleable, ductile and good conductors. Non-metals gain or share electrons; they are usually gases or brittle solids, dull, and poor conductors (graphite is the exception). The reactivity of a metal is its tendency to lose electrons and form positive ions.

8.2 The Reactivity Series

Metals can be arranged in order of decreasing reactivity. The series used in this book is:

Potassium > Sodium > Calcium > Magnesium > Aluminium > Zinc > Iron > Lead > Hydrogen > Copper > Silver > Gold > Platinum

A metal higher in the series is more reactive: it displaces a metal below it from its salt solution, while a metal below cannot displace one above it. For example, zinc displaces copper from copper sulphate ($Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$), but copper does not displace zinc from zinc sulphate. Gold and platinum are at the bottom: they are noble metals and hardly react at all.

8.3 Reactions of Metals

- **With oxygen:** metals form metal oxides. $2Mg + O_2 \rightarrow 2MgO$ (magnesium burns with a bright white flame). $4Fe + 3O_2 \rightarrow 2Fe_2O_3$. Sodium and potassium react so violently with air that they are stored under kerosene oil.
- **With water:** potassium and sodium react violently with cold water, giving hydrogen and the metal hydroxide: $2Na + 2H_2O \rightarrow 2NaOH + H_2$. Calcium reacts slowly with cold water; magnesium reacts with steam; iron reacts with steam to give iron(II,III) oxide and hydrogen: $3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2$. Copper, silver and gold do not react with water.
- **With dilute acids:** metals above hydrogen in the reactivity series react with dilute hydrochloric or sulphuric acid to give a salt and hydrogen gas: $Zn + 2HCl \rightarrow ZnCl_2 + H_2$. Metals below hydrogen (copper, silver, gold) do not react with dilute acids.
- **With chlorine:** metals form chlorides: $2Na + Cl_2 \rightarrow 2NaCl$.

8.4 Reactions of Non-metals and Halogens

- **Halogens** (group 17) are the most reactive non-metals. Fluorine is the most reactive of all; reactivity decreases down the group: F > Cl > Br > I.
- A more reactive halogen displaces a less reactive one from its salt: $Cl_2 + 2NaBr \rightarrow 2NaCl + Br_2$. Bromine cannot displace chlorine from sodium chloride.
- **Oxygen** reacts with non-metals to form acidic oxides: $C + O_2 \rightarrow CO_2$; $S + O_2 \rightarrow SO_2$. Carbon dioxide and sulphur dioxide dissolve in water to form acids.
- **Noble gases** (group 18) have complete outer shells and are almost unreactive; they are used in light bulbs (argon) and in balloons (helium).

8.5 Oxides: Acidic, Basic and Amphoteric

Type of oxide	Formed by	Reaction with	Examples
Basic	Metals	React with acids to give salt + water	Na_2O, CaO, MgO
Acidic	Non-metals	React with bases to give salt + water	CO_2, SO_2, NO_2
Amphoteric	Some metals	React with both acids and bases	Al_2O_3, ZnO
Neutral	Some non-metals	React with neither	CO, N_2O, H_2O

Examples: $CaO + 2HCl \rightarrow CaCl_2 + H_2O$ (basic oxide with acid); $CO_2 + Ca(OH)_2 \rightarrow CaCO_3 + H_2O$ (acidic oxide with base, the limewater test); $Al_2O_3 + 6HCl \rightarrow 2AlCl_3 + 3H_2O$ and $Al_2O_3 + 2NaOH \rightarrow 2NaAlO_2 + H_2O$ (amphoteric).

8.6 Uses of Common Metals and Non-metals

- **Iron:** construction, bridges, machinery, tools. **Aluminium:** aeroplanes, window frames, cooking utensils, electric cables (light and does not rust).
- **Copper:** electric wires, water pipes, coins (excellent conductor). **Zinc:** galvanising iron, dry cells. **Silver and gold:** jewellery, coins, electrical contacts.
- **Carbon (graphite):** pencils, electrodes, lubricant. **Sulphur:** sulphuric acid, fertilisers, medicines. **Chlorine:** disinfecting drinking water and swimming pools, making bleaching powder and PVC. **Iodine:** antiseptic, iodised salt.

Chapter 8 — Practice Questions

Multiple Choice Questions — choose one correct answer.

1. Which metal is the most reactive in the series?
(a) Zinc (b) Iron (c) Sodium (d) Potassium

2. Which metal will displace copper from copper sulphate solution?
(a) Silver (b) Gold (c) Zinc (d) Copper itself

3. Which metal reacts violently with cold water?
(a) Iron (b) Copper (c) Sodium (d) Magnesium

4. Which metal does NOT react with dilute hydrochloric acid?
(a) Zinc (b) Iron (c) Magnesium (d) Copper

5. Which oxide is amphoteric?
(a) Na_2O (b) CO_2 (c) Al_2O_3 (d) CaO

6. The order of halogen reactivity is:
(a) I > Br > Cl > F (b) F > Cl > Br > I (c) Cl > F > Br > I (d) Br > I > Cl > F

7. Carbon dioxide is an acidic oxide because it reacts with:
(a) Acids (b) Bases (c) Water only (d) Metals only

8. Which gas is released when zinc reacts with dilute sulphuric acid?
(a) Oxygen (b) Chlorine (c) Hydrogen (d) Carbon dioxide

9. Which metal is stored under kerosene oil?
(a) Iron (b) Sodium (c) Copper (d) Aluminium

10. The reaction $Cl_2 + 2NaBr \rightarrow 2NaCl + Br_2$ shows that chlorine is:
(a) Less reactive than bromine (b) More reactive than bromine (c) Equal in reactivity to bromine (d) A noble gas

Chapter 8 — Practice Questions (continued)

Short Questions

1. What is the reactivity series? Write the series from potassium to copper.

2. Why does zinc displace copper from copper sulphate solution, but copper does not displace zinc from zinc sulphate?

3. Write the word and chemical equations for the reaction of (a) sodium with water, (b) zinc with dilute hydrochloric acid.

4. What are amphoteric oxides? Give two examples and one reaction of each type (with an acid and with a base) for aluminium oxide.

5. Why are gold and platinum called noble metals? Give two uses of gold and two uses of copper.

Long Questions

1. (a) Describe the reactions of metals with oxygen, water and dilute acids, giving one example and one equation for each. (b) Explain why sodium is stored under kerosene oil while iron is not.

2. (a) Explain the reactivity of halogens and the displacement reactions between them, with equations. (b) Classify oxides as acidic, basic, amphoteric and neutral, giving two examples and one reaction for each of the first three types.

MCQ Answer Key

Check your answers below. If you got a question wrong, re-read the study notes for that chapter before moving on.

Chapter 1 — Fundamentals of Chemistry: 1. (b) 2. (c) 3. (b) 4. (c) 5. (b) 6. (b) 7. (c) 8. (c) 9. (b) 10. (a)

Chapter 2 — Structure of Atoms: 1. (c) 2. (b) 3. (c) 4. (b) 5. (b) 6. (c) 7. (b) 8. (b) 9. (b) 10. (c)

Chapter 3 — Periodic Table and Periodicity: 1. (b) 2. (c) 3. (d) 4. (b) 5. (c) 6. (c) 7. (b) 8. (b) 9. (b) 10. (b)

Chapter 4 — Structure of Molecules: 1. (b) 2. (c) 3. (c) 4. (b) 5. (b) 6. (b) 7. (c) 8. (b) 9. (c) 10. (b)

Chapter 5 — Physical States of Matter: 1. (b) 2. (b) 3. (b) 4. (c) 5. (b) 6. (c) 7. (b) 8. (b) 9. (b) 10. (c)

Chapter 6 — Solutions: 1. (b) 2. (b) 3. (b) 4. (b) 5. (b) 6. (b) 7. (a) 8. (b) 9. (b) 10. (b)

Chapter 7 — Electrochemistry: 1. (b) 2. (c) 3. (b) 4. (b) 5. (b) 6. (b) 7. (c) 8. (b) 9. (b) 10. (c)

Chapter 8 — Chemical Reactivity: 1. (d) 2. (c) 3. (c) 4. (d) 5. (c) 6. (b) 7. (b) 8. (c) 9. (b) 10. (b)

Final Revision Checklist

- Can you write the formulae of NaCl, $CaCO_3$, H_2SO_4, NH_3 and Al_2O_3 from memory?
- Can you calculate moles, molecular mass and percentage composition without help?
- Can you write the electronic configurations of the first 20 elements?
- Can you explain the trends in atomic radius, ionisation energy and electronegativity?
- Can you draw (in words or on paper) ionic bonding in NaCl and covalent bonding in H_2O, NH_3 and CH_4?
- Can you use Boyle's law, Charles's law and the general gas equation in numerical problems?
- Can you calculate percentage by mass and molarity of a solution?
- Can you find oxidation numbers and identify oxidising and reducing agents?
- Can you describe the Daniel cell, electrolysis of water, electroplating and galvanising?
- Can you write the reactivity series and predict displacement reactions?
- Is your practical notebook complete, with every experiment, observation and result written up?

Examination Tips

- In Section A, answer all 12 MCQs; there is no choice, so do not leave any blank.
- In Section B, read all 11 short questions first, then answer the 8 you know best. Write to the point: each question carries 4 marks.
- In Section C, attempt any 3 long questions. Give both parts of each question, with equations and a labelled diagram where the question asks for one.
- Always balance chemical equations, show the steps in numerical problems, and write the unit with every answer.

> **About these notes:** the study notes and questions in this book are original writing prepared for students of the Khyber Pakhtunkhwa Textbook Board Class 9 Chemistry course. They explain the same syllabus topics as the prescribed textbook in fresh words; they do not reproduce textbook text, and the practice questions are not past-paper questions. Paper-pattern figures are reproduced from official BISE documents (BISE Bannu model papers, BISE Malakand Chemistry 9th paper, BISE Peshawar SSC Annual-II 2026 date sheet) as explained on the pattern page.