9th Class Chemistry Notes Sindh Board (STBB) Full Book Notes

Complete chapter-wise notes with original practice questions

Based on the Sindh Textbook Board (STBB), Jamshoro, Chemistry book for Class 9

For all Sindh boards: Karachi, Hyderabad, Sukkur, Larkana and Mirpurkhas

8 chapters • Key concepts, definitions, equations and key terms

80 original MCQs • 40 short questions • 16 long questions

with an MCQ answer key at the end

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About These Notes

These notes follow the chapters of the Sindh Textbook Board (STBB), Jamshoro, Chemistry book prescribed for Class 9 in Sindh. They are written for students of all five Sindh boards — Karachi, Hyderabad, Sukkur, Larkana and Mirpurkhas. Each chapter has concise study notes followed by original practice questions: 10 multiple-choice questions, 5 short questions with answers and 2 long questions with answers. All questions are original study questions written for these notes; they are not copied from the textbook or from any past paper.

The book is published in English medium and in Urdu medium; these notes are written in **English**, matching the English-medium book.

Contents

- **Chapter 1:** Fundamentals of Chemistry
- **Chapter 2:** Structure of Atoms
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Board Examination Pattern (Sindh)

The pattern below is reproduced from third-party copies of the Board of Secondary Education Karachi (BSEK) Class 9 Chemistry (Theory) model paper and third-party marks tables. It has **not** been read directly off an official BSEK or STBB document — confirm it from your board's latest model paper.

Paper	Marks	Time
Theory paper — Section A: 12 MCQs from the whole syllabus (1 mark each)	12	3 hours
Section B: Short-answer questions — answer any 8 out of 12 (3 marks each)	24	
Section C: Descriptive questions — answer any 4 out of 6 (6 marks each)	24	
Theory total	**60**	
Practical examination (third-party figure, not officially confirmed)	15	
Chemistry total (theory + practical)	**75**	

Note: older 10-chapter lists and a new test edition with a different order are different editions; this book follows the 8-chapter list several independent sources agree on. See the note at the end.

How to Use These Notes

- Read one chapter's notes first, then answer its questions without looking at the answers.
- Learn the definitions and equations in the shaded boxes — they carry the most marks.
- Use the MCQ answer key at the end to check your work, and re-read any chapter where you score below 7 out of 10.

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.

1.2 Branches of Chemistry

- **Physical chemistry:** the study of the physical properties of matter and the laws that govern chemical change, such as energy changes and rates of reaction.
- **Organic chemistry:** the study of compounds of carbon, especially those containing carbon and hydrogen.
- **Inorganic chemistry:** the study of all elements and their compounds other than organic compounds.
- **Biochemistry:** the study of chemical processes inside living organisms.
- **Analytical chemistry:** the study of methods used to identify substances and measure how much of each is present.
- **Industrial chemistry:** the study of chemical processes used in factories to make useful products on a large scale.

1.3 Matter, Elements, Compounds and Mixtures

An **element** is a pure substance made of only one kind of atom; it cannot be broken down into simpler substances by chemical means. Examples are hydrogen, oxygen, iron and gold. A **compound** is a pure substance formed when two or more elements combine chemically in a fixed ratio, for example water (H_2O) and common salt (NaCl). A **mixture** contains two or more substances mixed together without a chemical reaction, in any ratio; its parts keep their own properties and can usually be separated by physical methods.

1.4 Symbols, Formulae and Valency

Each element is represented by a **symbol** — one or two letters, the first always capital. A **chemical formula** shows the elements in a compound and the number of atoms of each, for example CO_2 and H_2SO_4. **Valency** is the combining power of an element, the number of hydrogen atoms (or their equivalent) with which one atom of the element combines. The formula of a compound is written by crossing over the valencies of its parts, for example Ca^{2+} and Cl^{-} give $CaCl_2$.

1.5 Atomic Number, Mass Number and the Mole

The **atomic number (Z)** is the number of protons in the nucleus of an atom. The **mass number (A)** is the total number of protons and neutrons in the nucleus, so the number of neutrons is A − Z. The **relative atomic mass** compares the mass of an atom with one-twelfth of the mass of a carbon-12 atom. One **mole** of any substance contains 6.02×10^{23} particles (Avogadro's number) and has a mass in grams equal to its relative atomic or molecular mass.

> **Remember:** Element — one kind of atom. Compound — elements chemically joined in a fixed ratio. Mixture — substances physically mixed, separable by physical means. Mole — 6.02×10^{23} particles.

Chapter 1: Fundamentals of Chemistry — Questions

Multiple-Choice Questions

Q1. Chemistry is the study of

(a) matter and its changes (b) living things only (c) the stars and planets (d) forces and motion

Q2. Which branch of chemistry deals with carbon compounds?

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

Q3. Which of the following is an element?

(a) Water (b) Common salt (c) Iron (d) Air

Q4. Water is an example of a

(a) element (b) mixture (c) compound (d) metal

Q5. Air is an example of a

(a) compound (b) mixture (c) element (d) molecule

Q6. The symbol of sodium is

(a) S (b) So (c) Na (d) N

Q7. The atomic number of an element is the number of

(a) neutrons (b) protons (c) electrons and neutrons (d) protons and neutrons

Q8. An atom has atomic number 11 and mass number 23. Its neutrons are

(a) 11 (b) 12 (c) 23 (d) 34

Q9. One mole of any substance contains

(a) 6.02 x 10^23 particles (b) 6.02 x 10^22 particles (c) 12 particles (d) 100 particles

Q10. The valency of oxygen in water (H2O) is

(a) 1 (b) 2 (c) 3 (d) 4

Chapter 1: Fundamentals of Chemistry — Questions (continued)

Short Questions

Q1. Define chemistry.

Answer: Chemistry is the branch of science that studies matter — its composition, structure, properties and the changes it undergoes.

Q2. What is an element? Give two examples.

Answer: An element is a pure substance made of only one kind of atom and cannot be broken down into simpler substances by chemical means. Examples: iron and oxygen.

Q3. Distinguish between a compound and a mixture.

Answer: A compound is formed by chemical combination of elements in a fixed ratio, with new properties; a mixture is a physical mixing in any ratio, its parts keep their properties and can be separated physically.

Q4. What is meant by the mole?

Answer: One mole of a substance is the amount containing 6.02 x 10^23 particles (Avogadro's number), with a mass in grams equal to its relative atomic or molecular mass.

Q5. Define valency and give the valency of calcium.

Answer: Valency is the combining power of an element. Calcium has valency 2, as in CaCl2, where one calcium atom combines with two chlorine atoms.

Long Questions

Q1. Explain the main branches of chemistry with one example of what each studies.

Answer: Physical chemistry studies energy and rates of change; organic chemistry studies carbon compounds such as methane; inorganic chemistry studies elements and their compounds such as salts and metals; biochemistry studies chemical processes in living things, such as digestion; analytical chemistry identifies substances and measures their amounts; industrial chemistry covers large-scale manufacture, such as fertiliser production; environmental chemistry studies pollutants in air and water. Together they cover every aspect of matter and its uses.

Q2. What is the difference between an element, a compound and a mixture? Explain with examples and describe how a compound is formed.

Answer: An element contains only one kind of atom, for example gold. A compound is formed when elements combine chemically in a fixed ratio, for example hydrogen and oxygen form water, H2O, with properties different from its elements. A mixture, such as air, is a physical combination in any ratio whose parts keep their properties. A compound is formed by a chemical reaction in which atoms rearrange and bonds form; its composition is fixed and it can be separated only by chemical methods, whereas a mixture can be separated by physical methods such as filtration or evaporation.

Chapter 2: Structure of Atoms

2.1 The Atom and its Particles

An atom is the smallest particle of an element that takes part in a chemical reaction. It is made of three subatomic particles: the **proton** (positive charge, mass 1, in the nucleus), the **neutron** (no charge, mass 1, in the nucleus) and the **electron** (negative charge, almost no mass, moving around the nucleus). An atom is electrically neutral because it has equal numbers of protons and electrons.

2.2 Dalton's Atomic Theory

Dalton 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 chemical reactions are rearrangements of atoms. Later discoveries showed that atoms are divisible into subatomic particles and that atoms of the same element can differ (isotopes), so the theory was modified, but its main ideas still guide chemistry.

2.3 Rutherford's Model and Bohr's Model

Rutherford fired alpha particles at a thin gold foil and found that most passed straight through, a few were deflected, and very few bounced back. He concluded that the atom is mostly empty space, with a small, dense, positively charged nucleus at the centre and electrons around it. Bohr improved this model by proposing that electrons revolve around the nucleus in fixed energy levels or shells (K, L, M, N), and that an electron does not radiate energy while it stays in its shell; it absorbs energy to jump to a higher shell and gives out energy when it falls back.

2.4 Electronic Configuration

Electrons fill shells in order of increasing energy. The maximum number of electrons in a shell is $2n^2$, where n is the shell number: the K shell (n = 1) holds 2, L holds 8, M holds 18. The electrons in the outermost shell are called **valence electrons**; they decide how an atom behaves in chemical reactions. For example, sodium (atomic number 11) has the configuration 2, 8, 1, so it has one valence electron and readily loses it.

2.5 Isotopes

Isotopes are atoms of the same element with the same atomic number but different mass numbers, because they have different numbers of neutrons. Hydrogen has three isotopes: protium (^1H), deuterium (^2H) and tritium (^3H). Carbon-12 and carbon-14 are isotopes of carbon; carbon-14 is radioactive and is used in dating ancient organic material. Isotopes of an element have the same chemical properties because they have the same number of electrons, but different physical properties.

> **Remember:** Protons and neutrons sit in the nucleus; electrons occupy shells. Atomic number = protons. Mass number = protons + neutrons. Shell capacity = $2n^2$.

Chapter 2: Structure of Atoms — Questions

Multiple-Choice Questions

Q1. The smallest particle of an element that takes part in a chemical reaction is the

(a) molecule (b) atom (c) ion (d) electron

Q2. The proton is found in the

(a) electron shell (b) nucleus (c) orbit (d) valence shell

Q3. The charge on a neutron is

(a) positive (b) negative (c) zero (d) double positive

Q4. An atom is electrically neutral because the number of protons equals the number of

(a) neutrons (b) electrons (c) ions (d) shells

Q5. The maximum number of electrons in the L shell is

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

Q6. The electronic configuration of sodium (atomic number 11) is

(a) 2, 8, 2 (b) 2, 8, 1 (c) 2, 9 (d) 2, 8, 8, 1

Q7. Isotopes of an element have the same

(a) mass number (b) number of neutrons (c) atomic number (d) number of nucleons

Q8. Carbon-14 is an isotope of carbon-12 with

(a) 6 protons, 8 neutrons (b) 8 protons, 6 neutrons (c) 6 protons, 6 neutrons (d) 14 protons

Q9. Who proposed that electrons revolve in fixed energy levels?

(a) Dalton (b) Rutherford (c) Bohr (d) Thomson

Q10. The valence electrons of chlorine (atomic number 17) are

(a) 1 (b) 2 (c) 7 (d) 8

Chapter 2: Structure of Atoms — Questions (continued)

Short Questions

Q1. Name the three subatomic particles and state their charges.
Answer: Proton: positive charge, in the nucleus. Neutron: no charge, in the nucleus. Electron: negative charge, moving around the nucleus.

Q2. What did Rutherford's gold-foil experiment show?
Answer: Most alpha particles passed through, so the atom is mostly empty space; a few were deflected and very few bounced back, showing a small dense positively charged nucleus at the centre.

Q3. State Bohr's main idea about electrons.
Answer: Electrons revolve around the nucleus in fixed energy levels (shells) and do not radiate energy while in a shell; they absorb energy to move up and emit energy when they fall to a lower shell.

Q4. What are valence electrons?
Answer: The electrons in the outermost shell of an atom. They take part in chemical bonding and decide the chemical behaviour of the element.

Q5. Define isotopes and give one example.
Answer: Isotopes are atoms of the same element with the same atomic number but different mass numbers, e.g. carbon-12 and carbon-14.

Long Questions

Q1. Describe the structure of an atom. Explain the position, charge and mass of each subatomic particle.
Answer: An atom has a small dense nucleus at the centre containing protons and neutrons, with electrons moving in shells around it. The proton carries a positive charge and has a relative mass of 1. The neutron has no charge and a relative mass of 1. The electron carries a negative charge and has almost no mass (about 1/1840 of a proton). The number of protons equals the number of electrons, so the atom is neutral. The nucleus contains almost all the mass of the atom, while most of the atom's volume is empty space occupied by electrons.

Q2. What are isotopes? Explain with examples and describe one use of a radioactive isotope.
Answer: Isotopes are atoms of the same element with the same atomic number but different mass numbers because they contain different numbers of neutrons. Hydrogen has protium, deuterium and tritium; carbon has carbon-12 and carbon-14; chlorine has chlorine-35 and chlorine-37. Isotopes have the same chemical properties since they have the same electronic arrangement, but different physical properties such as mass and density. Radioactive isotopes are useful: carbon-14 is used to find the age of ancient wood, bone and other organic remains, because it decays at a known rate after the organism dies.

Chapter 3: Periodic Table and Periodicity of Properties

3.1 The Periodic Table

The periodic table arranges all known elements in order of increasing atomic number. It has horizontal rows called **periods** (7 in total) and vertical columns called **groups** (18 in total). Elements in the same group have the same number of valence electrons and similar chemical properties. The table is divided into metals (on the left and centre), non-metals (on the right) and metalloids (along the dividing line, such as silicon and boron).

3.2 Development of the Periodic Table

Early tables were arranged by atomic mass. Mendeleev arranged elements by atomic mass and left gaps for elements that had not yet been discovered, predicting their properties correctly. The modern periodic law states that the properties of elements are a periodic function of their atomic numbers. Arranging by atomic number removed the problems of Mendeleev's table, such as the positions of argon and potassium.

3.3 Periodicity of Properties

Properties change gradually across a period and repeat in the next period. **Atomic radius** decreases across a period (greater nuclear pull on the same shell) and increases down a group (new shells are added). **Ionisation energy** — the energy needed to remove an electron from a gaseous atom — generally increases across a period and decreases down a group. **Electron affinity** and **electronegativity** (the power of an atom to attract shared electrons) also generally increase across a period and decrease down a group.

3.4 Groups of the Periodic Table

Group 1, the **alkali metals** (lithium, sodium, potassium), are soft, reactive metals with one valence electron. Group 2, the **alkaline earth metals** (magnesium, calcium), have two valence electrons. Group 17, the **halogens** (fluorine, chlorine, bromine), are reactive non-metals with seven valence electrons. Group 18, the **noble gases** (helium, neon, argon), have full outer shells and are almost unreactive. The **transition metals** sit in the middle of the table and include iron, copper and zinc.

> **Remember:** Same group = same valence electrons = similar chemistry. Across a period: radius falls, ionisation energy rises. Down a group: radius rises, ionisation energy falls.

Chapter 3: Periodic Table and Periodicity of Properties — Questions

Multiple-Choice Questions

Q1. The modern periodic table arranges elements in order of increasing

(a) atomic mass (b) atomic number (c) density (d) melting point

Q2. The horizontal rows of the periodic table are called

(a) groups (b) periods (c) blocks (d) series

Q3. Elements in the same group have the same number of

(a) protons (b) neutrons (c) valence electrons (d) shells

Q4. Which of the following is a noble gas?

(a) Chlorine (b) Sodium (c) Argon (d) Oxygen

Q5. Atomic radius across a period generally

(a) increases (b) decreases (c) stays the same (d) first increases then decreases

Q6. Atomic radius down a group generally

(a) decreases (b) increases (c) stays the same (d) becomes zero

Q7. Ionisation energy down a group generally

(a) increases (b) decreases (c) stays the same (d) doubles

Q8. Which element is an alkali metal?

(a) Calcium (b) Sodium (c) Chlorine (d) Iron

Q9. The number of groups in the modern periodic table is

(a) 7 (b) 8 (c) 18 (d) 32

Q10. Electronegativity generally increases across a period because

(a) the atom gets bigger (b) the nuclear pull on bonding electrons increases (c) the atom loses shells (d) the number of neutrons falls

Chapter 3: Periodic Table and Periodicity of Properties — Questions (continued)

Short Questions

Q1. State the modern periodic law.

Answer: The properties of elements are a periodic function of their atomic numbers.

Q2. What is a period and what is a group?

Answer: A period is a horizontal row of the periodic table; a group is a vertical column. There are 7 periods and 18 groups.

Q3. Why do elements in the same group show similar chemical properties?

Answer: They have the same number of valence electrons, so they behave similarly in chemical reactions.

Q4. Define ionisation energy.

Answer: The energy needed to remove the outermost electron from a gaseous atom to form a positive ion.

Q5. Why does atomic radius decrease across a period?

Answer: The nuclear charge increases while electrons are added to the same shell, so the nucleus pulls the electrons more strongly inward.

Long Questions

Q1. Describe the structure of the modern periodic table and explain how elements are classified into metals, non-metals and metalloids.

Answer: The modern periodic table has 7 periods and 18 groups, arranged by increasing atomic number. Metals lie on the left and in the centre; they are shiny, good conductors of heat and electricity, malleable and ductile, and they lose electrons to form positive ions. Non-metals lie on the right; they are poor conductors, dull, brittle when solid, and they gain or share electrons. Metalloids such as silicon, boron and germanium lie along the border between metals and non-metals and show properties of both. Groups 1 and 2 are the alkali and alkaline earth metals, group 17 the halogens and group 18 the noble gases.

Q2. Explain the trends in atomic radius, ionisation energy and electronegativity across a period and down a group, giving reasons.

Answer: Across a period, atomic radius decreases because the increasing nuclear charge pulls the electrons of the same shell inward more strongly. Ionisation energy increases across a period because the smaller, more strongly held electrons are harder to remove. Electronegativity increases for the same reason. Down a group, atomic radius increases because a new electron shell is added in each period, and the outer electrons are shielded from the nucleus by inner shells. Ionisation energy and electronegativity therefore decrease down a group, because the outer electron is further away and more shielded, so it is easier to remove and less able to attract bonding electrons.

Chapter 4: Structure of Molecules (Chemical Bonding)

4.1 Why Atoms Form Bonds

Atoms form chemical bonds to achieve a stable electronic arrangement, usually a full outer shell of eight electrons — the **octet rule** (hydrogen and helium follow the duplet rule, two electrons). A **chemical bond** is the force that holds atoms together in a molecule or crystal. Atoms may transfer electrons, share electrons, or share their electrons as a sea of free electrons, giving ionic, covalent and metallic bonds.

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 positive and negative ions that attract each other. Sodium loses one electron to form Na^+ and chlorine gains it to form Cl^-; the two ions attract to form sodium chloride. Ionic compounds have high melting and boiling points, are usually soluble in water, conduct electricity when molten or in solution (but not as solids), and are hard but brittle.

4.3 Covalent Bond

A **covalent bond** is formed when two atoms share one or more pairs of electrons. A single bond shares one pair (H_2, HCl), a double bond shares two pairs (O_2) and a triple bond shares three pairs (N_2). A **coordinate (dative) covalent bond** is a covalent bond in which both shared electrons come from one atom, as in the ammonium ion (NH_4^+). Covalent compounds usually have low melting and boiling points, do not conduct electricity, and may be gases, liquids or soft solids.

4.4 Metallic Bond and Intermolecular Forces

In a **metallic bond**, metal atoms share a sea of free (delocalised) electrons. This explains why metals conduct electricity and heat, are malleable and ductile, and have a shine. Between molecules there are weaker forces called **intermolecular forces**, including dipole-dipole forces and **hydrogen bonding** — the strong attraction between a hydrogen atom bonded to fluorine, oxygen or nitrogen and a lone pair on a nearby atom of one of those elements. Hydrogen bonding explains why water has a high boiling point and why ice floats.

4.5 Polar and Non-Polar Covalent Bonds

When two atoms share electrons equally, the bond is **non-polar**, as in H_2 and Cl_2. When the atoms differ in electronegativity, the shared pair is pulled towards the more electronegative atom, giving a **polar** covalent bond with a slightly positive end and a slightly negative end, as in HCl and H_2O. The greater the difference in electronegativity, the more ionic the character of the bond.

> **Remember:** Ionic = transfer of electrons (metal + non-metal). Covalent = sharing of electrons (non-metal + non-metal). Metallic = sea of free electrons. Hydrogen bonding needs H bonded to F, O or N.

Chapter 4: Structure of Molecules (Chemical Bonding) — Questions

Multiple-Choice Questions

Q1. Atoms form chemical bonds to achieve

(a) a full outer shell (b) more protons (c) a bigger nucleus (d) more neutrons

Q2. The octet rule states that atoms tend to have

(a) two electrons in the outer shell (b) eight electrons in the outer shell (c) eight shells (d) two shells

Q3. An ionic bond is formed by

(a) sharing of electrons (b) complete transfer of electrons (c) sea of electrons (d) hydrogen bonding

Q4. Sodium chloride is an example of

(a) a covalent compound (b) an ionic compound (c) a metallic compound (d) a mixture

Q5. A double covalent bond is present in

(a) H_2 (b) O_2 (c) HCl (d) NaCl

Q6. In a coordinate covalent bond

(a) electrons are transferred (b) both shared electrons come from one atom (c) no electrons are shared (d) electrons are free to move

Q7. Which of the following is a property of ionic compounds?

(a) Low melting point (b) Conduct electricity when molten (c) Do not dissolve in water (d) Are soft solids

Q8. Metals conduct electricity because of

(a) covalent bonds (b) the sea of free electrons (c) hydrogen bonds (d) ionic bonds

Q9. Hydrogen bonding occurs when hydrogen is bonded to

(a) carbon (b) fluorine, oxygen or nitrogen (c) sodium (d) chlorine only

Q10. HCl has a polar covalent bond because

(a) chlorine is more electronegative than hydrogen (b) hydrogen is a metal (c) both atoms are identical (d) no electrons are shared

Chapter 4: Structure of Molecules (Chemical Bonding) — Questions (continued)

Short Questions

Q1. Define a chemical bond.

Answer: A chemical bond is the force of attraction that holds atoms together in a molecule or a crystal.

Q2. What is the octet rule?

Answer: Atoms tend to gain, lose or share electrons so that their outermost shell contains eight electrons, like a noble gas.

Q3. How is an ionic bond formed? Give one example.

Answer: It is formed by the complete transfer of electrons from a metal to a non-metal, producing oppositely charged ions that attract each other, e.g. NaCl.

Q4. What is a coordinate covalent bond?

Answer: A covalent bond in which the shared pair of electrons is supplied by only one of the two bonded atoms, e.g. the bond formed when NH3 accepts H+ to give NH4+.

Q5. Why do metals conduct electricity?

Answer: Because of the sea of free (delocalised) electrons in the metallic bond, which can move through the metal and carry charge.

Long Questions

Q1. Explain ionic and covalent bonding with examples, and compare the properties of ionic and covalent compounds.

Answer: An ionic bond forms by complete transfer of electrons from a metal to a non-metal, as in NaCl, where Na+ and Cl- attract strongly. A covalent bond forms by sharing electron pairs between non-metals, as in H2, O2 and HCl. Ionic compounds have high melting and boiling points, are usually soluble in water and conduct electricity when molten or dissolved, but not as solids. Covalent compounds usually have low melting and boiling points, are often insoluble in water, and do not conduct electricity because they contain no free ions or electrons. Ionic compounds are hard and brittle; covalent compounds are often gases, liquids or soft solids.

Q2. What is hydrogen bonding? Explain its importance with reference to water.

Answer: Hydrogen bonding is the strong attraction between a hydrogen atom covalently bonded to a highly electronegative atom (fluorine, oxygen or nitrogen) and a lone pair of electrons on a nearby atom of fluorine, oxygen or nitrogen. In water, each molecule forms hydrogen bonds with its neighbours. This makes water have a much higher boiling point than expected for such a small molecule, gives it high surface tension, and makes ice less dense than liquid water because the hydrogen bonds hold the molecules in an open structure — so ice floats, protecting aquatic life in winter.

Chapter 5: Physical States of Matter

5.1 The Three States of Matter

Matter exists in three common states: **solid**, **liquid** and **gas**. In a solid, particles are closely packed in a fixed pattern and vibrate about fixed positions, so a solid has a fixed shape and volume. In a liquid, particles are close together but can move past each other, so a liquid has a fixed volume but takes the shape of its container. In a gas, particles are far apart and move freely, so a gas fills its container and can be compressed easily. A fourth state, **plasma**, is ionised gas found in the Sun and in lightning.

5.2 Kinetic Molecular Theory

The kinetic molecular theory explains the behaviour of matter in terms of moving particles. All particles are in constant motion; the motion increases with temperature. In gases, particles move rapidly in straight lines, collide with each other and with the walls of the container, and the pressure of a gas is caused by these collisions. There are large spaces between gas particles, which is why gases are easily compressed and have low density.

5.3 Diffusion, Evaporation and Boiling

Diffusion is the movement of particles from a region of higher concentration to a region of lower concentration, until they are evenly spread; it is fastest in gases and slowest in solids, and it speeds up with temperature. **Evaporation** is the change of a liquid into vapour at its surface at any temperature, and it causes cooling because the fastest molecules escape and take energy with them. **Boiling** is the rapid change of a liquid into vapour throughout the liquid at a fixed temperature, the boiling point, when the vapour pressure equals the atmospheric pressure.

5.4 Melting, Freezing and Sublimation

Melting is the change of a solid into a liquid at its melting point; **freezing** is the reverse change. During a change of state the temperature stays constant because the heat supplied is used to overcome the forces between particles — this is called **latent heat**. **Sublimation** is the direct change of a solid into a gas without passing through the liquid state, shown by camphor, naphthalene and solid carbon dioxide (dry ice). The reverse process, gas to solid, is called deposition.

5.5 Gas Laws (Qualitative)

Boyle's law states that, at constant temperature, the volume of a fixed mass of gas decreases as the pressure increases. Charles's law states that, at constant pressure, the volume of a fixed mass of gas increases as the temperature increases. Together these laws describe how the pressure, volume and temperature of a gas are related, and they follow from the kinetic molecular theory.

> **Remember:** Solid — fixed shape and volume. Liquid — fixed volume, shape of container. Gas — fills container, compressible. Diffusion: gas > liquid > solid. Boiling happens at one fixed temperature; evaporation happens at all temperatures.

Chapter 5: Physical States of Matter — Questions

Multiple-Choice Questions

Q1. Which state of matter has a fixed shape and a fixed volume?

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

Q2. In a liquid, the particles

(a) are fixed in position (b) can move past each other (c) are far apart (d) do not move

Q3. The pressure of a gas is caused by

(a) the weight of the gas (b) particles colliding with the walls (c) the temperature alone (d) the colour of the gas

Q4. Diffusion is fastest in

(a) solids (b) liquids (c) gases (d) plasma

Q5. Evaporation takes place

(a) only at the boiling point (b) at the surface at any temperature (c) only in solids (d) only at 0 C

Q6. Boiling differs from evaporation because boiling occurs

(a) only at the surface (b) throughout the liquid at a fixed temperature (c) only at night (d) without heat

Q7. The direct change of a solid into a gas is called

(a) melting (b) condensation (c) sublimation (d) freezing

Q8. During melting, the temperature of the solid

(a) rises steadily (b) falls (c) stays constant (d) doubles

Q9. Which of the following sublimes?

(a) Iron (b) Camphor (c) Water (d) Copper

Q10. According to Boyle's law, at constant temperature, increasing the pressure on a gas

(a) increases its volume (b) decreases its volume (c) does not change its volume (d) liquefies it at once

Chapter 5: Physical States of Matter — Questions (continued)

Short Questions

Q1. Name the three common states of matter and state one property of each.

Answer: Solid: fixed shape and volume. Liquid: fixed volume but takes the shape of its container. Gas: no fixed shape or volume; it fills its container.

Q2. What is diffusion? Give one example.

Answer: Diffusion is the movement of particles from a region of higher concentration to a region of lower concentration, e.g. the smell of perfume spreading across a room.

Q3. Define evaporation and explain why it causes cooling.

Answer: Evaporation is the change of a liquid into vapour at its surface at any temperature. The fastest molecules escape, taking energy with them, so the remaining liquid cools.

Q4. What is sublimation? Give one example.

Answer: Sublimation is the direct change of a solid into a gas without becoming a liquid first, e.g. camphor or naphthalene.

Q5. State Boyle's law.

Answer: At constant temperature, the volume of a fixed mass of gas is inversely proportional to its pressure — as pressure rises, volume falls.

Long Questions

Q1. Describe the three states of matter using the kinetic molecular theory, and explain how they change into each other.

Answer: In a solid, particles are closely packed in a regular pattern and vibrate about fixed positions, held by strong forces, so the solid keeps its shape and volume. In a liquid, particles are close but can slide past each other, so the liquid flows and takes the shape of its container while keeping its volume. In a gas, particles are far apart, move fast in all directions and collide with the walls, so the gas fills its container and is compressible. Heating a solid melts it to a liquid; further heating boils it to a gas. Cooling reverses the changes: gas condenses to liquid, liquid freezes to solid. During a change of state the temperature stays constant because heat is used as latent heat to overcome the forces between particles.

Q2. Explain diffusion, evaporation and boiling, and state the factors that affect the rate of evaporation.

Answer: Diffusion is the spreading of particles from high to low concentration until evenly mixed; it is fastest in gases because their particles move fastest, and it speeds up with temperature. Evaporation is the escape of molecules from the surface of a liquid at any temperature; boiling is the rapid formation of vapour throughout the liquid at a fixed temperature when vapour pressure equals atmospheric pressure. The rate of evaporation increases with temperature, with a larger surface area, with wind removing vapour, and with lower humidity, because each of these helps more molecules escape from the surface.

Chapter 6: Solutions

6.1 Solutions, Solute and Solvent

A **solution** is a homogeneous mixture of two or more substances, uniform throughout. The substance that dissolves is the **solute** and the substance that does the dissolving is the **solvent**. In a salt solution, salt is the solute and water is the solvent. Water is called the universal solvent because it dissolves many substances. Solutions may be solid, liquid or gaseous: brass is a solid solution of metals, seawater is a liquid solution, and air is a gaseous solution.

6.2 Types of Solutions and Solubility

Solutions are classified by the amount of solute they contain. An **unsaturated solution** can dissolve more solute at a given temperature. A **saturated solution** contains the maximum amount of solute that can dissolve at that temperature. A **supersaturated solution** contains more solute than a saturated solution and is unstable. **Solubility** is the maximum mass of solute that dissolves in 100 g of solvent at a given temperature. Solubility of most solids in water increases with temperature, while the solubility of gases decreases as temperature rises.

6.3 Concentration of Solutions

The **concentration** of a solution tells how much solute is present in a given amount of solution or solvent. Common ways to express it are mass percentage (mass of solute per 100 g of solution), volume percentage, and **molarity** — the number of moles of solute in one litre (1 dm^3) of solution. For example, dissolving 40 g of sodium hydroxide (1 mole) in water and making the volume up to 1 litre gives a 1 molar (1 M) solution.

6.4 Suspensions and Colloids

A **suspension** is a heterogeneous mixture in which solid particles are spread through a liquid but do not dissolve; they settle on standing and can be separated by filtration, for example muddy water or chalk in water. A **colloid** is a mixture in which tiny particles are evenly spread and do not settle, for example milk, fog and jelly. Colloids scatter light (the Tyndall effect), while true solutions do not. Colloidal particles are too small to see but larger than the particles in a true solution.

6.5 Separation and Uses

Solutions are separated by methods such as evaporation, crystallisation and distillation. Pure water is obtained from seawater by distillation. Solutions are used everywhere: in medicines, in food and drinks, in cleaning, in agriculture as fertiliser solutions, and in industry as solvents and reactants.

> **Remember:** Solute + solvent = solution. Saturated = cannot dissolve more. Molarity = moles of solute per litre of solution. Suspension settles; colloid does not; true solution is uniform and clear.

Chapter 6: Solutions — Questions

Multiple-Choice Questions

Q1. A homogeneous mixture of two or more substances is called a

(a) suspension (b) solution (c) colloid (d) compound

Q2. In a salt solution, water is the

(a) solute (b) solvent (c) suspension (d) colloid

Q3. A solution that cannot dissolve any more solute at a given temperature is

(a) unsaturated (b) saturated (c) dilute (d) supersaturated

Q4. The solubility of most solid solutes in water

(a) decreases with temperature (b) increases with temperature (c) does not change (d) becomes zero on heating

Q5. Molarity is defined as moles of solute per

(a) 100 g of solvent (b) 1 litre of solution (c) 1 kg of solution (d) 100 cm3 of solute

Q6. Which of the following is a colloid?

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

Q7. Solid particles in a suspension

(a) dissolve completely (b) settle on standing (c) cannot be filtered out (d) are invisible

Q8. Brass is an example of a

(a) liquid solution (b) gaseous solution (c) solid solution (d) suspension

Q9. A solution of 1 mole of solute in 1 litre of solution is

(a) 0.1 M (b) 1 M (c) 10 M (d) 100 M

Q10. Which method gives pure water from seawater?

(a) Filtration (b) Distillation (c) Decantation (d) Sublimation

Chapter 6: Solutions — Questions (continued)

Short Questions

Q1. Define solution, solute and solvent.

Answer: A solution is a homogeneous mixture; the solute is the substance that dissolves and the solvent is the substance that does the dissolving, e.g. salt (solute) in water (solvent).

Q2. What is a saturated solution?

Answer: A solution that contains the maximum amount of solute that can dissolve in the solvent at a given temperature; it cannot dissolve any more solute.

Q3. Define solubility.

Answer: Solubility is the maximum mass of solute that dissolves in 100 g of solvent at a given temperature.

Q4. What is molarity?

Answer: Molarity is the number of moles of solute dissolved in one litre (1 dm3) of solution.

Q5. Distinguish between a suspension and a colloid.

Answer: In a suspension, solid particles do not dissolve and settle on standing, e.g. muddy water. In a colloid, tiny particles stay evenly spread and do not settle, e.g. milk.

Long Questions

Q1. Explain the terms saturated, unsaturated and supersaturated solutions, and describe how temperature affects the solubility of solids and gases.

Answer: An unsaturated solution can still dissolve more solute at a given temperature. A saturated solution holds the maximum amount of solute possible at that temperature, with any extra solute remaining undissolved. A supersaturated solution holds more solute than a saturated solution, usually made by cooling a hot saturated solution carefully; it is unstable and the extra solute crystallises out easily. The solubility of most solid solutes in water increases with temperature, because heating helps break the forces holding the solid together. The solubility of gases in liquids decreases with temperature, which is why warm water holds less dissolved oxygen than cold water.

Q2. What is meant by the concentration of a solution? Explain mass percentage and molarity, and work out the molarity of a solution made by dissolving 20 g of sodium hydroxide (molar mass 40 g/mol) in water to make 500 cm3 of solution.

Answer: Concentration is the amount of solute in a given amount of solution or solvent. Mass percentage is the mass of solute per 100 g of solution. Molarity is the number of moles of solute per litre of solution. For the example: moles of NaOH = 20 g / 40 g/mol = 0.5 mol. Volume = 500 cm3 = 0.5 litre. Molarity = 0.5 mol / 0.5 L = 1.0 mol/L, i.e. a 1 M solution.

Chapter 7: Electrochemistry

7.1 Oxidation and Reduction

Oxidation is the loss of electrons or the gain of oxygen; **reduction** is the gain of electrons or the loss of oxygen. Oxidation and reduction always happen together in a **redox reaction**. In the reaction $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$, zinc loses electrons and is oxidised, while copper(II) ions gain electrons and are reduced. The substance that causes oxidation (and is itself reduced) is the **oxidising agent**; the substance that causes reduction (and is itself oxidised) is the **reducing agent**.

7.2 Electrolysis

Electrolysis is the decomposition of a compound by passing an electric current through its molten state or its aqueous solution. The vessel is the electrolytic cell; the positive electrode is the **anode** and the negative electrode is the **cathode**. Positive ions (cations) move to the cathode and gain electrons (reduction), while negative ions (anions) move to the anode and lose electrons (oxidation). Electrolysis of molten sodium chloride gives sodium metal at the cathode and chlorine gas at the anode. Electrolysis of water (with a little acid added) gives hydrogen at the cathode and oxygen at the anode, in the volume ratio 2:1.

7.3 Electrolytes and Non-Electrolytes

An **electrolyte** is a substance that conducts electricity when molten or in aqueous solution because it contains free ions — for example sodium chloride, acids and alkalis. A **non-electrolyte**, such as sugar or pure water, does not conduct electricity because it has no free ions. Strong electrolytes are almost fully ionised in solution, while weak electrolytes are only slightly ionised.

7.4 Electrochemical Cells

An **electrochemical (galvanic) cell** converts chemical energy into electrical energy. In a simple cell, two different metals (for example zinc and copper) are dipped in an electrolyte and connected by a wire; zinc loses electrons (oxidation at the anode) and the electrons flow through the wire to the copper electrode, where ions gain them (reduction at the cathode), producing a current. Dry cells and car batteries are practical electrochemical cells. **Electroplating** uses electrolysis to coat one metal with a thin layer of another metal, for example chromium plating to prevent rusting and improve appearance.

7.5 Corrosion and its Prevention

Corrosion is the slow eating away of a metal by air, water and other substances. Rusting of iron needs both oxygen and water: $4Fe + 3O_2 + 6H_2O \rightarrow 4Fe(OH)_3$ (rust). Rusting is prevented by painting, oiling, galvanising (coating iron with zinc), tin-plating, and by sacrificial protection, in which a more reactive metal such as magnesium is attached to the iron and corrodes instead.

> **Remember:** OIL RIG — Oxidation Is Loss, Reduction Is Gain (of electrons). Cathode: reduction, cations gain electrons. Anode: oxidation, anions lose electrons. Rusting needs oxygen AND water.

Chapter 7: Electrochemistry — Questions

Multiple-Choice Questions

Q1. Oxidation is defined as

(a) gain of electrons (b) loss of electrons (c) gain of hydrogen (d) loss of oxygen

Q2. In a redox reaction, the oxidising agent is

(a) oxidised (b) reduced (c) unchanged (d) precipitated

Q3. Electrolysis is the decomposition of a compound by

(a) heat (b) light (c) electric current (d) a catalyst

Q4. In electrolysis, the positive electrode is called the

(a) cathode (b) anode (c) cell (d) ion

Q5. Cations move towards the

(a) anode (b) cathode (c) battery (d) solution

Q6. At the cathode, ions undergo

(a) oxidation (b) reduction (c) neutralisation (d) sublimation

Q7. Which of the following is an electrolyte?

(a) Sugar solution (b) Pure water (c) Sodium chloride solution (d) Kerosene

Q8. Electrolysis of water gives hydrogen and oxygen in the volume ratio

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

Q9. Galvanising protects iron by coating it with

(a) tin (b) zinc (c) copper (d) chromium

Q10. Rusting of iron needs

(a) oxygen only (b) water only (c) oxygen and water (d) carbon dioxide only

Chapter 7: Electrochemistry — Questions (continued)

Short Questions

Q1. Define oxidation and reduction in terms of electrons.

Answer: Oxidation is the loss of electrons; reduction is the gain of electrons. They always occur together in a redox reaction.

Q2. What is electrolysis?

Answer: Electrolysis is the decomposition of a compound by passing an electric current through its molten state or aqueous solution.

Q3. Name the electrodes in electrolysis and state what happens at each.

Answer: The anode is the positive electrode, where anions lose electrons (oxidation). The cathode is the negative electrode, where cations gain electrons (reduction).

Q4. What is an electrolyte? Give one example.

Answer: An electrolyte is a substance that conducts electricity when molten or in aqueous solution because it contains free ions, e.g. sodium chloride solution.

Q5. What is galvanising?

Answer: Galvanising is coating iron or steel with a layer of zinc to protect it from rusting, because zinc is more reactive and corrodes first.

Long Questions

Q1. Explain electrolysis with the example of molten sodium chloride, and describe what happens at each electrode.

Answer: Electrolysis is the breakdown of a compound by an electric current. In molten sodium chloride, Na+ and Cl- ions are free to move. When current passes, Na+ ions move to the negative cathode, gain electrons and form sodium metal: Na+ + e- -> Na (reduction). Cl- ions move to the positive anode, lose electrons and form chlorine gas: 2Cl- -> Cl2 + 2e- (oxidation). So sodium metal collects at the cathode and chlorine gas is released at the anode. The same principles apply to the electrolysis of aqueous solutions, except that water may also take part in the electrode reactions.

Q2. What is rusting? Explain the conditions needed for rusting and describe three methods of preventing it.

Answer: Rusting is the corrosion of iron — its slow reaction with oxygen and water to form hydrated iron(III) oxide, Fe(OH)3, a brown flaky solid. Both oxygen and water are needed; iron kept dry, or kept away from air, does not rust. Rusting is prevented by barrier methods such as painting, oiling and greasing, which keep air and water away from the surface; by coating with another metal, such as galvanising with zinc or tin-plating; and by sacrificial protection, in which a more reactive metal like magnesium is attached to the iron so that it corrodes instead of the iron. Rust is a problem because it is porous and flakes off, exposing fresh iron, unlike the protective oxide layer on aluminium.

Chapter 8: Chemical Reactivity

8.1 The Reactivity Series

Metals differ in how readily they react. Arranged in order of decreasing reactivity they form the **reactivity series**: potassium, sodium, calcium, magnesium, aluminium, zinc, iron, lead, hydrogen, copper, silver, gold. A more reactive metal displaces a less reactive metal from its salt solution; for example, zinc displaces copper from copper(II) sulfate solution, but copper cannot displace zinc from zinc sulfate. The most reactive metals react violently with cold water; less reactive metals react only with steam or acids, and the least reactive metals hardly react at all.

8.2 Reactions of Metals

Very reactive metals such as potassium and sodium react with cold water to give the metal hydroxide and hydrogen gas, with much heat. Calcium reacts slowly with cold water. Magnesium reacts with steam to give magnesium oxide and hydrogen. Metals above hydrogen in the series react with dilute acids to give a salt and hydrogen gas, for example $Zn + 2HCl \rightarrow ZnCl_2 + H_2$. Metals below hydrogen, such as copper, silver and gold, do not react with dilute acids. Metals react with oxygen to form metal oxides; sodium burns with a yellow flame, while iron rusts slowly.

8.3 Types of Chemical Reactions

In a **combination (synthesis) reaction**, two or more substances join to form one product, for example $2H_2 + O_2 \rightarrow 2H_2O$. In a **decomposition reaction**, one substance breaks into two or more products, for example $CaCO_3 \rightarrow CaO + CO_2$ on heating. In a **displacement reaction**, a more reactive element takes the place of a less reactive one, for example $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$. In a **double displacement reaction**, the ions of two compounds exchange partners, for example $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$. In a **neutralisation reaction**, an acid and a base react to form a salt and water.

8.4 Reactivity of Non-Metals

Non-metals show a reactivity trend too. The halogens become less reactive down the group: fluorine is the most reactive and iodine the least. A more reactive halogen displaces a less reactive halogen from its salt solution, for example chlorine displaces bromine from potassium bromide solution. Carbon is used as a reducing agent to extract metals from their oxides: for example, carbon reduces zinc oxide to zinc in a blast furnace, but it cannot reduce very reactive metals such as aluminium, which need electrolysis.

8.5 Chemical Reactivity in Daily Life

Reactivity decides how metals are extracted, stored and used. Reactive metals are stored under oil to keep them away from air and moisture. Gold and platinum are used for jewellery because they are unreactive and do not tarnish. Iron is painted or galvanised because it rusts. Fertilisers, medicines and fuels all depend on controlling chemical reactions so that they happen at the right rate and give the wanted products.

> **Remember:** Reactivity series: K, Na, Ca, Mg, Al, Zn, Fe, Pb, H, Cu, Ag, Au. A metal displaces only metals below it. Combination joins; decomposition splits; displacement swaps one; double displacement swaps two.

Chapter 8: Chemical Reactivity — Questions

Multiple-Choice Questions

Q1. The most reactive metal in the series is

(a) gold (b) copper (c) potassium (d) iron

Q2. Zinc displaces copper from copper sulfate solution because zinc is

(a) less reactive than copper (b) more reactive than copper (c) a non-metal (d) insoluble

Q3. Which metal reacts with cold water violently?

(a) Copper (b) Iron (c) Sodium (d) Silver

Q4. The reaction Zn + 2HCl -> ZnCl2 + H2 is an example of a metal reacting with

(a) water (b) an acid (c) oxygen (d) a base

Q5. Which metal does NOT react with dilute hydrochloric acid?

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

Q6. In a combination reaction

(a) one substance splits into two (b) two substances join to form one product (c) ions exchange partners (d) an acid and a base react

Q7. The decomposition of calcium carbonate on heating gives

(a) CaO and CO2 (b) Ca and CO (c) Ca(OH)2 (d) CaCl2

Q8. The reaction AgNO3 + NaCl -> AgCl + NaNO3 is a

(a) combination reaction (b) displacement reaction (c) double displacement reaction (d) decomposition reaction

Q9. Which halogen is the most reactive?

(a) Fluorine (b) Chlorine (c) Bromine (d) Iodine

Q10. Gold is used in jewellery mainly because it is

(a) cheap (b) very reactive (c) unreactive and does not tarnish (d) light

Chapter 8: Chemical Reactivity — Questions (continued)

Short Questions

Q1. What is the reactivity series?

Answer: It is the list of metals arranged in order of decreasing chemical reactivity, from potassium (most reactive) to gold (least reactive).

Q2. Why does zinc displace copper from copper sulfate solution?

Answer: Zinc is more reactive than copper, so it takes the sulfate from copper, giving zinc sulfate and copper metal.

Q3. Write the equation for zinc reacting with dilute hydrochloric acid.

Answer: Zn + 2HCl -> ZnCl2 + H2 (zinc chloride and hydrogen gas).

Q4. What is a displacement reaction? Give one example.

Answer: A reaction in which a more reactive element takes the place of a less reactive element in a compound, e.g. Zn + CuSO4 -> ZnSO4 + Cu.

Q5. What is neutralisation?

Answer: The reaction between an acid and a base to form a salt and water, e.g. HCl + NaOH -> NaCl + H2O.

Long Questions

Q1. Describe the reactivity series of metals and explain, with examples, how metals react with water, acids and oxygen.

Answer: The reactivity series runs from the most reactive metals (potassium, sodium, calcium) through magnesium, aluminium, zinc, iron and lead, past hydrogen, to the least reactive (copper, silver, gold). Potassium and sodium react violently with cold water, giving the hydroxide and hydrogen; calcium reacts slowly with cold water; magnesium reacts with steam; zinc and iron react with steam but not cold water; copper, silver and gold do not react with water. Metals above hydrogen react with dilute acids to give a salt and hydrogen, while metals below hydrogen do not. With oxygen, reactive metals burn vigorously to form oxides, while iron rusts slowly and gold does not react at all. A metal displaces any metal below it in the series from its salt solution, but never one above it.

Q2. Explain the four main types of chemical reactions — combination, decomposition, displacement and double displacement — with one balanced equation for each.

Answer: A combination reaction joins two or more substances into one product: 2H2 + O2 -> 2H2O. A decomposition reaction splits one substance into two or more products, often on heating: CaCO3 -> CaO + CO2. A displacement reaction replaces one element in a compound with a more reactive element: Zn + CuSO4 -> ZnSO4 + Cu. A double displacement reaction exchanges the ions of two compounds, often forming a precipitate: AgNO3 + NaCl -> AgCl + NaNO3. Neutralisation, an acid reacting with a base to give a salt and water, is a common special case of double displacement.

MCQ Answer Key

Answers to all 80 multiple-choice questions, chapter by chapter. The correct option letter is shown.

Chapter 1: 1a 2b 3c 4c 5b 6c 7b 8b 9a 10b

Chapter 2: 1b 2b 3c 4b 5b 6b 7c 8a 9c 10c

Chapter 3: 1b 2b 3c 4c 5b 6b 7b 8b 9c 10b

Chapter 4: 1a 2b 3b 4b 5b 6b 7b 8b 9b 10a

Chapter 5: 1c 2b 3b 4c 5b 6b 7c 8c 9b 10b

Chapter 6: 1b 2b 3b 4b 5b 6b 7b 8c 9b 10b

Chapter 7: 1b 2b 3c 4b 5b 6b 7c 8b 9b 10c

Chapter 8: 1c 2b 3c 4b 5c 6b 7a 8c 9a 10c

A Note on the Chapter List

Three different chapter lists for Sindh Class 9 Chemistry circulate online, and they caused a genuine conflict during research. The list used in this book — 8 chapters: Fundamentals of Chemistry; Structure of Atoms; Periodic Table and Periodicity of Properties; Structure of Molecules (Chemical Bonding); Physical States of Matter; Solutions; Electrochemistry; Chemical Reactivity — is the one given by several independent sources that describe the current STBB book (a Sindh notes guide describing the latest syllabus, the 2026 Sindh textbook chapter list, and a textbook listing updated for the 2025-26 syllabus). An older STBB list has 10 chapters (adding Chemical Combinations, Acids, Bases and Salts, and Chemical Energetics as separate chapters). A new test edition seen in late 2026 uses yet another order, starting with States of Matter and including Stoichiometry and Equilibria. If your school is using the new test edition, the topics in this book still cover the same chemistry — only the chapter order and grouping differ. Check with your teacher which edition your board examination follows.