9th Class Chemistry Notes Federal Board (FBISE) Full Book Notes

Complete chapter-wise notes with original practice questions

Based on the National Book Foundation (NBF) Chemistry book for Class 9
Federal Board of Intermediate and Secondary Education, Islamabad

19 chapters • Key concepts, definitions, equations and key terms
190 original MCQs • 95 short questions • 38 long questions
with an MCQ answer key at the end

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About These Notes

These notes follow the chapters of the National Book Foundation (NBF) Chemistry book prescribed for Class 9 by the Federal Board of Intermediate and Secondary Education (FBISE), Islamabad. Each chapter contains concise study notes with key concepts, definitions, equations and key terms, followed by original practice questions: 10 multiple-choice questions, 5 short questions and 2 long questions. The answers to all MCQs are given in the answer key at the end of the book.

> **About the paper:** The FBISE Class 9 Chemistry paper is SLO-based (Student Learning Outcomes). The paper pattern, marks distribution and timing can change, so students should always check the current model paper on the official FBISE website, **fbise.edu.pk**, before the examination. No marks figures or paper timings are given in these notes.

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How to use these notes: read the chapter notes first, learn the key terms, then attempt the practice questions without looking at the notes. Check your MCQ answers with the key at the end and revise any chapter where you score below 7 out of 10.

Chapter 1: Nature of Science in Chemistry

Key Concepts

- Science is organised knowledge gained through observation, experimentation and logical reasoning. Chemistry is the branch of science that studies the composition, structure, properties and changes of matter.
- Main branches of chemistry: organic chemistry (carbon compounds), inorganic chemistry (all other elements and their compounds), physical chemistry (energy and laws), analytical chemistry (identification and measurement), biochemistry (chemical processes in living things), environmental chemistry (chemicals in air, water and soil) and industrial chemistry (large-scale manufacture).
- Chemistry is closely linked with physics (energy and forces), biology (life processes), medicine (drugs and diagnosis), agriculture (fertilisers and pesticides) and industry (materials, fuels, plastics).
- The scientific method: observe a phenomenon, ask a question, form a hypothesis, design and perform experiments, record and analyse data, draw a conclusion, and report results so others can test them.
- A hypothesis is a testable explanation; a theory is a well-tested explanation supported by much evidence; a law is a statement that describes a pattern found in nature, such as the law of conservation of mass.
- Qualitative data describe qualities (colour, smell, state) while quantitative data are measured in numbers with units. Accurate measurements need correct instruments and standard units (SI units).
- Safety in the laboratory: wear goggles and gloves, never taste chemicals, read labels before use, handle acids with care, keep flammables away from flames, and know the first-aid and fire-safety steps.

Key Terms

- **Science**
- **Chemistry**
- **Hypothesis**
- **Theory**
- **Law**
- **Scientific method**
- **SI unit**

Chapter 1: Nature of Science in Chemistry — Practice Questions

Multiple Choice Questions

Q1. Which of the following is the best definition of chemistry?

(A) The study of living things (B) The study of matter, its composition, properties and changes (C) The study of forces and motion (D) The study of the Earth and rocks

Q2. The branch of chemistry that deals with carbon compounds is called

(A) inorganic chemistry (B) physical chemistry (C) organic chemistry (D) analytical chemistry

Q3. A testable explanation for an observation is called a

(A) law (B) theory (C) hypothesis (D) fact

Q4. Which of these is a qualitative observation?

(A) The solution is blue (B) The solution weighs 25 g (C) The temperature is 40 °C (D) The volume is 10 cm^3

Q5. The SI unit of mass is the

(A) gram (B) pound (C) kilogram (D) tonne

Q6. Which branch of chemistry studies the amount and identity of substances in a sample?

(A) Biochemistry (B) Analytical chemistry (C) Organic chemistry (D) Industrial chemistry

Q7. In the scientific method, the step that comes immediately after forming a hypothesis is

(A) drawing a conclusion (B) reporting results (C) designing and performing experiments (D) making an observation

Q8. A scientific law is best described as a statement that

(A) explains why something happens (B) describes a pattern found in nature (C) is a guess that has not been tested (D) changes with every new experiment

Q9. Which safety rule is correct in a chemistry laboratory?

(A) Taste a chemical to identify it (B) Wear safety goggles when handling chemicals (C) Pour water into concentrated acid quickly (D) Heat flammable liquids over an open flame

Q10. Biochemistry is mainly the study of

(A) chemicals in the atmosphere (B) chemical processes in living organisms (C) large-scale industrial manufacture (D) rocks and minerals

Short Questions

S1. Define chemistry and name its two most basic divisions.

S2. What is the difference between a hypothesis and a theory?

S3. List any four branches of chemistry with one example of what each studies.

S4. Why should a scientist repeat an experiment before accepting its result?

S5. Give two examples of qualitative and two of quantitative observations.

Long Questions

L1. Explain the steps of the scientific method in order, with a simple example from chemistry.

L2. Describe how chemistry is related to medicine and agriculture, giving two examples for each.

Chapter 2: Matter

Key Concepts

- Matter is anything that has mass and occupies space. All matter is made of tiny particles (atoms and molecules) that are always moving, with spaces between them.
- Physical properties (colour, state, density, melting and boiling points) can be observed without changing the substance. Chemical properties describe how a substance reacts to form new substances.
- States of matter: solids have a fixed shape and volume with particles in fixed positions; liquids have a fixed volume but take the shape of their container; gases have neither fixed shape nor fixed volume and spread to fill the container.
- Changes of state: melting, freezing, evaporation, condensation, sublimation and deposition. Evaporation happens at the surface at any temperature; boiling happens throughout the liquid at its boiling point.
- Elements are pure substances made of only one kind of atom (for example, iron, oxygen). Compounds are pure substances formed when two or more elements combine chemically in a fixed ratio (for example, water, H_2O). Mixtures contain two or more substances mixed physically and can be separated by physical methods.
- Homogeneous mixtures have a uniform composition throughout (salt solution, air). Heterogeneous mixtures have a non-uniform composition (sand and water, oil and water).
- Physical changes do not produce new substances (melting ice). Chemical changes produce new substances with new properties (burning wood, rusting of iron).

Key Terms

- **Matter**
- **Element**
- **Compound**
- **Mixture**
- **Homogeneous mixture**
- **Heterogeneous mixture**
- **Sublimation**

Chapter 2: Matter — Practice Questions

Multiple Choice Questions

Q1. Which of the following is matter?

(A) Light (B) Heat (C) Air (D) Sound

Q2. The particles of a solid

(A) move freely in all directions (B) vibrate about fixed positions (C) slide past one another easily (D) are far apart with large spaces

Q3. Which change is a chemical change?

(A) Melting of ice (B) Boiling of water (C) Rusting of iron (D) Dissolving salt in water

Q4. A compound differs from a mixture because a compound

(A) can be separated by filtration (B) has its components in a fixed ratio (C) has a variable composition (D) shows the properties of its parts

Q5. Which of the following is a heterogeneous mixture?

(A) Salt solution (B) Air (C) Oil and water (D) Sugar solution

Q6. Sublimation is the change of a solid directly into a

(A) liquid (B) gas (C) plasma (D) solution

Q7. The chemical symbol for water is

(A) CO_2 (B) O_2 (C) H_2O (D) NaCl

Q8. Which property is a chemical property?

(A) Density (B) Melting point (C) Flammability (D) Colour

Q9. Evaporation differs from boiling because evaporation

(A) occurs only at the surface and at any temperature (B) occurs throughout the liquid at a fixed temperature (C) needs a very high temperature (D) produces bubbles in the liquid

Q10. Which of the following is an element?

(A) Brass (B) Steel (C) Copper (D) Air

Short Questions

S1. Define matter and give three examples.

S2. How does a compound differ from a mixture? Give one example of each.

S3. Why do gases have neither a fixed shape nor a fixed volume?

S4. What is the difference between evaporation and boiling?

S5. Classify the following as element, compound or mixture: oxygen, carbon dioxide, sea water.

Long Questions

L1. Explain the three states of matter on the basis of particle arrangement and movement.

L2. Describe, with examples, how physical changes differ from chemical changes.

Chapter 3: Atomic Structure

Key Concepts

- All matter is made of atoms. An atom has a tiny central nucleus containing protons (positive charge) and neutrons (no charge), surrounded by electrons (negative charge) moving in shells.
- Atomic number (Z) = number of protons in an atom. Mass number (A) = number of protons + number of neutrons. Number of neutrons = A − Z.
- John Dalton proposed that matter is made of indivisible atoms and that atoms of the same element are identical. J. J. Thomson discovered the electron. Ernest Rutherford's gold-foil experiment showed that the atom has a small, dense, positively charged nucleus.
- Niels Bohr proposed that electrons move in fixed energy levels or shells (K, L, M, N) around the nucleus and do not radiate energy while in these shells.
- The maximum number of electrons in a shell is $2n^2$, where n is the shell number: K = 2, L = 8, M = 18. Electrons fill shells from the lowest energy level upwards.
- Isotopes are atoms of the same element with the same number of protons but different numbers of neutrons (for example, ^{1}H, ^{2}H and ^{3}H; carbon-12 and carbon-14). Isotopes have the same chemical properties but different masses.
- Electronic configuration is written by distributing electrons shell by shell, for example, sodium (11): 2, 8, 1; chlorine (17): 2, 8, 7. The outermost shell is the valence shell and its electrons take part in chemical bonding.

Key Terms

- **Atom**
- **Proton**
- **Neutron**
- **Electron**
- **Atomic number (Z)**
- **Mass number (A)**
- **Isotope**
- **Electronic configuration**

Chapter 3: Atomic Structure — Practice Questions

Multiple Choice Questions

Q1. The atomic number of an element is the number of

(A) neutrons (B) protons (C) protons + neutrons (D) electrons + neutrons

Q2. Which particle has a negative charge?

(A) Proton (B) Neutron (C) Electron (D) Nucleus

Q3. The mass number of an atom is equal to the number of

(A) protons only (B) electrons only (C) protons + neutrons (D) protons + electrons

Q4. An atom of carbon-14 has 6 protons. How many neutrons does it have?

(A) 6 (B) 8 (C) 12 (D) 14

Q5. Who discovered the electron?

(A) Dalton (B) Rutherford (C) J. J. Thomson (D) Bohr

Q6. The maximum number of electrons in the L shell is

(A) 2 (B) 8 (C) 18 (D) 32

Q7. Isotopes of the same element have the same number of

(A) neutrons (B) protons (C) nucleons (D) none of these

Q8. The electronic configuration of sodium (atomic number 11) is

(A) 2, 8, 2 (B) 2, 8, 1 (C) 2, 7, 2 (D) 2, 9

Q9. Rutherford's gold-foil experiment proved that the atom has a

(A) uniform positive charge (B) small dense nucleus (C) cloud of electrons only (D) solid structure

Q10. Which of the following is the valence shell of magnesium (2, 8, 2)?

(A) K shell (B) L shell (C) M shell (D) N shell

Short Questions

S1. Define atomic number and mass number. How are they related to neutrons?

S2. What are isotopes? Give one example.

S3. State the maximum number of electrons in the K, L and M shells.

S4. Write the electronic configuration of oxygen (8) and calcium (20).

S5. What did Rutherford conclude from his gold-foil experiment?

Long Questions

L1. Describe the structure of an atom, naming each subatomic particle, its charge and its location.

L2. Explain Bohr's model of the atom and how electrons are arranged in shells, with examples.

Chapter 4: Periodic Table and Periodicity of Properties

Key Concepts

- The periodic table arranges elements in order of increasing atomic number. Horizontal rows are periods (7) and vertical columns are groups (18). Elements in the same group have similar chemical properties because they have the same number of valence electrons.
- Dmitri Mendeleev arranged elements by atomic mass and predicted the properties of undiscovered elements. The modern periodic law states that the properties of elements are a periodic function of their atomic numbers.
- Metals are on the left and in the centre of the table; they are shiny, good conductors of heat and electricity, malleable and ductile. Non-metals are on the right; they are dull, poor conductors and brittle as solids. Metalloids (such as silicon and boron) lie along the staircase line and show properties of both.
- Group 1 elements are the alkali metals (Li, Na, K) — soft, reactive metals with one valence electron. Group 2 are the alkaline earth metals. Group 17 are the halogens (F, Cl, Br, I) — reactive non-metals with seven valence electrons. Group 18 are the noble gases (He, Ne, Ar) — very unreactive with full outer shells.
- Atomic size decreases across a period (more protons pull the same shell inwards) and increases down a group (new shells are added).
- Ionisation energy is the energy needed to remove an electron from an atom. It generally increases across a period and decreases down a group.
- Electronegativity is the ability of an atom to attract a shared pair of electrons in a bond. It increases across a period and decreases down a group. Electron affinity is the energy released when an atom gains an electron.

Key Terms

- **Period**
- **Group**
- **Periodic law**
- **Metal**
- **Non-metal**
- **Metalloid**
- **Ionisation energy**
- **Electronegativity**

Chapter 4: Periodic Table and Periodicity of Properties — Practice Questions

Multiple Choice Questions

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

(A) atomic mass (B) atomic number (C) density (D) number of neutrons

Q2. How many groups are there in the modern periodic table?

(A) 7 (B) 8 (C) 16 (D) 18

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) Oxygen (C) Argon (D) Sodium

Q5. Atomic size generally increases when moving

(A) across a period from left to right (B) down a group (C) across a period from right to left (D) nowhere

Q6. Which group contains the halogens?

(A) Group 1 (B) Group 2 (C) Group 17 (D) Group 18

Q7. Which element is a metalloid?

(A) Iron (B) Silicon (C) Sulphur (D) Sodium

Q8. Ionisation energy generally decreases down a group because

(A) the atom gets smaller (B) new shells are added and the outer electron is further from the nucleus (C) the nuclear charge decreases (D) electrons are removed more easily from small atoms

Q9. Who first arranged the elements in a table by atomic mass and predicted missing elements?

(A) Bohr (B) Mendeleev (C) Rutherford (D) Thomson

Q10. The most electronegative element in the periodic table is

(A) oxygen (B) chlorine (C) fluorine (D) nitrogen

Short Questions

S1. State the modern periodic law.

S2. What are periods and groups? How many of each are in the modern periodic table?

S3. How do metals and non-metals differ in their physical properties?

S4. Why are noble gases very unreactive?

S5. Describe the trend of atomic size across a period and down a group.

Long Questions

L1. Explain the trends in ionisation energy and electronegativity across a period and down a group, with reasons.

L2. Describe the main features of Groups 1, 17 and 18 of the periodic table.

Chapter 5: Chemical Bonding

Key Concepts

- Atoms bond to achieve a stable electronic configuration, usually an octet (eight electrons) in the outer shell, like the noble gases. Atoms do this by losing, gaining or sharing electrons.
- An ionic bond forms when a metal atom transfers one or more electrons to a non-metal atom. The metal becomes a positive ion (cation) and the non-metal becomes a negative ion (anion); the opposite charges attract. Example: Na^+Cl^- in sodium chloride.
- A covalent bond forms when two non-metal atoms share one or more pairs of electrons. A single bond shares one pair, a double bond two pairs and a triple bond three pairs. Examples: H_2 (single), O_2 (double), N_2 (triple).
- Coordinate (dative) covalent bond: one atom provides both electrons of the shared pair, as in the ammonium ion NH_4^+.
- A metallic bond is the attraction between metal cations and the sea of free (delocalised) electrons around them. It explains why metals conduct electricity and are malleable and ductile.
- Polar covalent bonds form when two different non-metals share electrons unequally because of different electronegativity; water (H_2O) is a polar molecule. Non-polar covalent bonds share electrons equally, as in H_2 or Cl_2.
- Ionic compounds have high melting and boiling points, are hard and brittle, and conduct electricity when molten or dissolved in water, but not as solids. Covalent compounds usually have low melting and boiling points and do not conduct electricity.
- Lewis dot structures show valence electrons as dots around the symbol of the atom; they help to draw how atoms share or transfer electrons.

Key Terms

- **Ionic bond**
- **Covalent bond**
- **Metallic bond**
- **Cation**
- **Anion**
- **Octet rule**
- **Polar bond**
- **Lewis structure**

Chapter 5: Chemical Bonding — Practice Questions

Multiple Choice Questions

Q1. An ionic bond is formed by the

(A) sharing of electrons (B) transfer of electrons (C) sharing of protons (D) transfer of protons

Q2. Which of the following is an ionic compound?

(A) H_2O (B) CO_2 (C) NaCl (D) CH_4

Q3. A covalent bond is formed between two atoms of

(A) a metal and a non-metal (B) two metals (C) two non-metals (D) a metal and a metal

Q4. The number of electron pairs shared in a triple bond is

(A) 1 (B) 2 (C) 3 (D) 4

Q5. Which compound has a double covalent bond?

(A) H_2 (B) O_2 (C) N_2 (D) HCl

Q6. A cation is formed when an atom

(A) gains electrons (B) loses electrons (C) gains protons (D) loses neutrons

Q7. Metals conduct electricity because of

(A) free ions (B) free (delocalised) electrons (C) covalent bonds (D) ionic bonds

Q8. Which of these is a polar molecule?

(A) H_2 (B) Cl_2 (C) H_2O (D) N_2

Q9. Ionic compounds conduct electricity when

(A) solid only (B) molten or dissolved in water (C) in the gas state (D) never

Q10. In a Lewis structure, the dots around an atom represent its

(A) protons (B) neutrons (C) valence electrons (D) inner electrons

Short Questions

S1. Define an ionic bond and a covalent bond.

S2. Why do atoms form chemical bonds?

S3. What is a metallic bond? Name two properties it explains.

S4. Distinguish between polar and non-polar covalent bonds with an example of each.

S5. Why do ionic compounds not conduct electricity in the solid state?

Long Questions

L1. Explain the formation of sodium chloride and magnesium oxide by electron transfer, using diagrams described in words.

L2. Compare ionic and covalent compounds in terms of melting point, electrical conductivity and solubility.

Chapter 6: Stoichiometry

Key Concepts

- Stoichiometry is the study of the quantitative relationships between reactants and products in a chemical reaction, based on the balanced chemical equation.
- The mole is the SI unit for amount of substance. One mole contains Avogadro's number of particles, 6.022×10^{23}. The mass of one mole of a substance is its molar mass in grams (for example, H_2O = 18 g mol^{-1}).
- Number of moles = mass ÷ molar mass. Number of particles = number of moles × 6.022×10^{23}.
- A balanced chemical equation obeys the law of conservation of mass: the number of atoms of each element is the same on both sides.
- The empirical formula is the simplest whole-number ratio of atoms in a compound; the molecular formula is the actual number of atoms (molecular formula = empirical formula × n, where n is a whole number).
- The limiting reactant is the reactant that is used up first and limits the amount of product formed. The other reactants are in excess.
- Percentage yield = (actual yield ÷ theoretical yield) × 100. The theoretical yield is the maximum amount calculated from the equation; the actual yield is usually less because of incomplete reactions and losses.

Key Terms

- **Mole**
- **Avogadro's number**
- **Molar mass**
- **Empirical formula**
- **Molecular formula**
- **Limiting reactant**
- **Percentage yield**

Chapter 6: Stoichiometry — Practice Questions

Multiple Choice Questions

Q1. One mole of any substance contains

(A) 6.022×10^{22} particles (B) 6.022×10^{23} particles (C) 6.022×10^{24} particles (D) 6.022×10^{25} particles

Q2. The molar mass of water (H_2O) is

(A) 16 g mol^{-1} (B) 17 g mol^{-1} (C) 18 g mol^{-1} (D) 20 g mol^{-1}

Q3. How many moles are there in 8 g of oxygen gas O_2 (molar mass 32 g mol^{-1})?

(A) 0.25 mol (B) 0.5 mol (C) 2 mol (D) 4 mol

Q4. A balanced chemical equation must obey the law of conservation of

(A) energy only (B) mass (C) volume (D) charge only

Q5. The empirical formula of glucose $C_6H_{12}O_6$ is

(A) $C_6H_{12}O_6$ (B) CH_2O (C) $C_2H_4O_2$ (D) CHO

Q6. The reactant that is completely used up first in a reaction is called the

(A) excess reactant (B) limiting reactant (C) catalyst (D) product

Q7. Percentage yield is calculated as

(A) (theoretical yield ÷ actual yield) × 100 (B) (actual yield ÷ theoretical yield) × 100 (C) actual yield + theoretical yield (D) theoretical yield − actual yield

Q8. The molar mass of carbon dioxide CO_2 is

(A) 28 g mol^{-1} (B) 32 g mol^{-1} (C) 44 g mol^{-1} (D) 48 g mol^{-1}

Q9. Stoichiometry is mainly concerned with

(A) the speed of a reaction (B) the quantities of reactants and products (C) the colour changes in a reaction (D) the temperature of a reaction

Q10. If the theoretical yield is 50 g and the actual yield is 40 g, the percentage yield is

(A) 40% (B) 50% (C) 80% (D) 90%

Short Questions

S1. Define a mole. What is Avogadro's number?

S2. How do you calculate the number of moles from a given mass?

S3. What is the difference between empirical and molecular formula?

S4. Define limiting reactant and give an example.

S5. Why is the actual yield usually less than the theoretical yield?

Long Questions

L1. Calculate the number of moles and the number of molecules in 36 g of water (H = 1, O = 16).

L2. Explain, with a worked example, how a balanced chemical equation is used to find the mass of a product formed.

Chapter 7: Electrochemistry

Key Concepts

- Electrochemistry is the study of the interconversion of chemical energy and electrical energy. It includes electrochemical (galvanic) cells, electrolytic cells and electrolysis.
- Oxidation is the loss of electrons (increase in oxidation number); reduction is the gain of electrons (decrease in oxidation number). In a redox reaction, oxidation and reduction always occur together. Remember: OIL RIG — Oxidation Is Loss, Reduction Is Gain.
- An oxidising agent accepts electrons and is itself reduced; a reducing agent donates electrons and is itself oxidised.
- A galvanic (voltaic) cell converts chemical energy into electrical energy. It has two half-cells connected by a salt bridge. At the anode oxidation occurs; at the cathode reduction occurs. Electrons flow through the external wire from anode to cathode.
- A standard hydrogen electrode (SHE) is assigned an electrode potential of 0.00 V and is used as the reference for measuring other electrode potentials.
- Electrolysis is the use of electrical energy to bring about a chemical change. In the electrolysis of water, hydrogen is produced at the cathode and oxygen at the anode. In the electrolysis of molten sodium chloride, sodium forms at the cathode and chlorine gas at the anode.
- Corrosion is the slow eating away of a metal by air and moisture, for example rusting of iron: $4Fe + 3O_2 + 6H_2O \rightarrow 4Fe(OH)_3$. It is prevented by painting, oiling, galvanising (coating with zinc) and using sacrificial anodes.

Key Terms

- **Oxidation**
- **Reduction**
- **Redox reaction**
- **Galvanic cell**
- **Electrolytic cell**
- **Electrolysis**
- **Anode**
- **Cathode**
- **Corrosion**
- **Galvanising**

Chapter 7: Electrochemistry — Practice Questions

Multiple Choice Questions

Q1. Oxidation is defined as the

(A) gain of electrons (B) loss of electrons (C) gain of protons (D) loss of neutrons

Q2. In a redox reaction, the reducing agent is

(A) reduced (B) oxidised (C) unchanged (D) neutralised

Q3. In a galvanic cell, oxidation takes place at the

(A) cathode (B) anode (C) salt bridge (D) electrolyte

Q4. Electrons flow through the external circuit of a galvanic cell from

(A) cathode to anode (B) anode to cathode (C) salt bridge to anode (D) cathode to salt bridge

Q5. The standard hydrogen electrode is assigned a potential of

(A) 1.00 V (B) 0.50 V (C) 0.00 V (D) −1.00 V

Q6. In the electrolysis of water, hydrogen gas is produced at the

(A) anode (B) cathode (C) both electrodes (D) neither electrode

Q7. Which metal is used to galvanise iron to protect it from rusting?

(A) Copper (B) Aluminium (C) Zinc (D) Lead

Q8. Rusting of iron requires

(A) oxygen only (B) water only (C) oxygen and water (D) carbon dioxide only

Q9. An oxidising agent

(A) donates electrons (B) accepts electrons (C) donates protons (D) accepts protons

Q10. In the electrolysis of molten sodium chloride, chlorine gas is formed at the

(A) cathode (B) anode (C) salt bridge (D) battery

Short Questions

S1. Define oxidation and reduction in terms of electrons.

S2. What is a galvanic cell? Name its two electrodes and the reaction at each.

S3. What is the standard hydrogen electrode and why is it used?

S4. What is corrosion? Give the word equation for rusting of iron.

S5. Name three methods of preventing corrosion.

Long Questions

L1. Explain the working of a galvanic cell, including the roles of the salt bridge and the external circuit.

L2. Describe electrolysis, giving the products formed at each electrode in the electrolysis of water and of molten sodium chloride.

Chapter 8: Energetics

Key Concepts

- Energetics is the study of energy changes during chemical reactions. Energy is usually exchanged as heat.
- An exothermic reaction releases heat to the surroundings; the temperature of the surroundings rises and ΔH is negative. Examples: combustion of fuels, respiration, neutralisation.
- An endothermic reaction absorbs heat from the surroundings; the temperature of the surroundings falls and ΔH is positive. Examples: photosynthesis, melting of ice, thermal decomposition of calcium carbonate.
- Enthalpy change (ΔH) is the heat change of a reaction at constant pressure. The standard enthalpy change is measured at 298 K and 1 atm pressure.
- Hess's law: the total enthalpy change of a reaction is the same regardless of the route taken, provided the starting and ending conditions are the same. It allows enthalpy changes to be calculated from known values.
- Bond energy is the energy needed to break one mole of a particular bond in the gas phase. Energy is absorbed to break bonds and released when bonds form. ΔH = energy needed to break bonds − energy released when bonds form.
- Fuels release energy on combustion. Fossil fuels (coal, oil, natural gas) are non-renewable; solar, wind and hydroelectric energy are renewable. Calorimetry measures heat changes using a calorimeter.

Key Terms

- **Exothermic reaction**
- **Endothermic reaction**
- **Enthalpy change (ΔH)**
- **Hess's law**
- **Bond energy**
- **Calorimetry**
- **Fuel**

Chapter 8: Energetics — Practice Questions

Multiple Choice Questions

Q1. An exothermic reaction

(A) absorbs heat from the surroundings (B) releases heat to the surroundings (C) has ΔH positive (D) lowers the temperature of the surroundings

Q2. Which of the following is an endothermic process?

(A) Combustion of methane (B) Melting of ice (C) Respiration (D) Neutralisation of an acid

Q3. For an endothermic reaction, the enthalpy change ΔH is

(A) zero (B) negative (C) positive (D) always −100 kJ mol^{-1}

Q4. Hess's law states that the enthalpy change of a reaction is

(A) different for different routes (B) the same regardless of the route taken (C) always zero (D) dependent on the catalyst

Q5. Energy is absorbed when bonds are

(A) formed (B) broken (C) both formed and broken (D) neither formed nor broken

Q6. Which reaction releases the most familiar form of heat energy in daily life?

(A) Photosynthesis (B) Combustion of fuels (C) Dissolving ammonium nitrate (D) Melting wax

Q7. The standard conditions for enthalpy change are

(A) 273 K and 1 atm (B) 298 K and 1 atm (C) 300 K and 2 atm (D) 0 K and 1 atm

Q8. In the reaction $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$, heat is

(A) absorbed (B) released (C) neither absorbed nor released (D) absorbed then released equally

Q9. A calorimeter is used to measure

(A) mass (B) volume (C) heat changes (D) pressure

Q10. Which of the following is a renewable source of energy?

(A) Coal (B) Natural gas (C) Solar energy (D) Petrol

Short Questions

S1. Define exothermic and endothermic reactions, giving one example of each.

S2. What is enthalpy change? What are standard conditions for measuring it?

S3. State Hess's law in your own words.

S4. Why is energy released when new bonds form?

S5. Give two examples of fuels and state whether each is renewable or non-renewable.

Long Questions

L1. Explain how Hess's law is used to find an enthalpy change that cannot be measured directly, with an example.

L2. Describe bond energies and show how they are used to estimate the enthalpy change of a reaction.

Chapter 9: Chemical Equilibrium

Key Concepts

- Many chemical reactions are reversible: the products can react to re-form the reactants. At equilibrium, the forward and reverse reactions occur at the same rate, so the concentrations of reactants and products stay constant.
- Chemical equilibrium is dynamic — both reactions continue, but there is no net change. Equilibrium can only be reached in a closed system.
- The equilibrium constant Kc = [products] ÷ [reactants], with each concentration raised to the power of its coefficient in the balanced equation. A large Kc means the equilibrium lies to the right (more products); a small Kc means it lies to the left.
- Le Chatelier's principle: when a change is applied to a system at equilibrium, the system shifts to oppose the change.
- Effect of concentration: adding a reactant shifts the equilibrium to the right (more product); removing a product also shifts it to the right.
- Effect of temperature: increasing the temperature favours the endothermic direction; decreasing it favours the exothermic direction.
- Effect of pressure: increasing the pressure favours the side with fewer gas molecules. A catalyst speeds up both forward and reverse reactions equally, so it does not change the position of equilibrium; it only helps equilibrium to be reached faster.

Key Terms

- **Reversible reaction**
- **Dynamic equilibrium**
- **Equilibrium constant (Kc)**
- **Le Chatelier's principle**

Chapter 9: Chemical Equilibrium — Practice Questions

Multiple Choice Questions

Q1. At chemical equilibrium, the rates of the forward and reverse reactions are

(A) zero (B) equal (C) maximum (D) unequal

Q2. Chemical equilibrium can be reached only in a

(A) closed system (B) open system (C) any system (D) system with a catalyst

Q3. A large value of Kc means the equilibrium mixture contains

(A) more reactants than products (B) more products than reactants (C) equal amounts of both (D) no products at all

Q4. According to Le Chatelier's principle, adding more reactant to a system at equilibrium

(A) shifts it to the left (B) shifts it to the right (C) has no effect (D) stops the reaction

Q5. Increasing the temperature of an exothermic equilibrium reaction favours the

(A) forward reaction (B) reverse reaction (C) neither direction (D) formation of a catalyst

Q6. Increasing the pressure favours the side with

(A) more gas molecules (B) fewer gas molecules (C) more liquid molecules (D) more solid molecules

Q7. A catalyst in an equilibrium reaction

(A) changes the position of equilibrium (B) speeds up both directions equally without changing the equilibrium position (C) only speeds up the forward reaction (D) is used up in the reaction

Q8. For the reaction $N_2 + 3H_2 \rightleftharpoons 2NH_3$, increasing the pressure will

(A) decrease the yield of ammonia (B) increase the yield of ammonia (C) have no effect on yield (D) stop the reaction

Q9. Which statement about dynamic equilibrium is correct?

(A) Both reactions have stopped (B) Only the forward reaction continues (C) Both reactions continue at the same rate (D) Concentrations keep changing

Q10. Removing a product from an equilibrium mixture shifts the equilibrium

(A) to the left (B) to the right (C) nowhere (D) toward the catalyst

Short Questions

S1. What is a reversible reaction? Give one example.

S2. What is meant by dynamic equilibrium?

S3. Write the expression for Kc for the reaction $A + B \rightleftharpoons C + D$.

S4. State Le Chatelier's principle.

S5. How does a catalyst affect a system at equilibrium?

Long Questions

L1. Explain Le Chatelier's principle and how changes in concentration, temperature and pressure shift the position of equilibrium.

L2. Describe the industrial manufacture of ammonia (Haber process) and explain how temperature, pressure and catalyst are chosen, using Le Chatelier's principle.

Chapter 10: Acids, Bases and Salts

Key Concepts

- An acid is a substance that produces hydrogen ions (H^+) in water (Arrhenius definition) or donates a proton (Brønsted-Lowry definition). Common acids: hydrochloric acid (HCl), sulphuric acid (H_2SO_4), nitric acid (HNO_3), ethanoic acid (CH_3COOH).
- A base is a substance that produces hydroxide ions (OH^-) in water or accepts a proton. Alkalis are soluble bases, such as sodium hydroxide (NaOH) and potassium hydroxide (KOH).
- Acids taste sour, turn blue litmus red and react with metals to produce hydrogen gas and with carbonates to produce carbon dioxide. Bases feel slippery, turn red litmus blue and neutralise acids.
- Neutralisation: acid + base → salt + water, for example, HCl + NaOH → NaCl + H_2O.
- The pH scale runs from 0 to 14: pH < 7 is acidic, pH = 7 is neutral, pH > 7 is alkaline. pH = −log[H^+]. Strong acids and bases ionise completely; weak acids and bases ionise only partly.
- Salts are ionic compounds formed when the hydrogen of an acid is replaced by a metal or ammonium ion. Common salts: sodium chloride, calcium carbonate, ammonium sulphate, copper(II) sulphate.
- Uses: sulphuric acid in fertilisers and batteries; hydrochloric acid in cleaning metals; sodium hydroxide in soap; calcium carbonate in cement and as an antacid; ammonium salts as fertilisers.

Key Terms

- **Acid**
- **Base**
- **Alkali**
- **Salt**
- **Neutralisation**
- **pH scale**
- **Indicator**
- **Strong acid**
- **Weak acid**

Chapter 10: Acids, Bases and Salts — Practice Questions

Multiple Choice Questions

Q1. Which of the following is an acid?

(A) NaOH (B) HCl (C) NH_3 (D) CaO

Q2. An alkali is a base that is

(A) insoluble in water (B) soluble in water (C) a gas (D) a salt

Q3. Acids turn blue litmus paper

(A) blue (B) red (C) green (D) colourless

Q4. The reaction between an acid and a base produces

(A) salt and hydrogen (B) salt and water (C) salt and oxygen (D) water and hydrogen

Q5. A solution with pH 3 is

(A) strongly alkaline (B) neutral (C) acidic (D) strongly basic

Q6. Which ion do all acids produce in water?

(A) OH^- (B) H^+ (C) Na^+ (D) Cl^-

Q7. The pH of a neutral solution at room temperature is

(A) 0 (B) 7 (C) 14 (D) 1

Q8. Which of the following is a weak acid?

(A) Hydrochloric acid (B) Sulphuric acid (C) Ethanoic acid (D) Nitric acid

Q9. When an acid reacts with a metal carbonate, the gas produced is

(A) hydrogen (B) oxygen (C) carbon dioxide (D) nitrogen

Q10. Which salt is used as a fertiliser?

(A) Sodium chloride (B) Ammonium sulphate (C) Calcium carbonate (D) Copper sulphate

Short Questions

S1. Define an acid and a base according to the Arrhenius theory.

S2. What is neutralisation? Write the equation for HCl + NaOH.

S3. What does the pH scale measure? What is the pH of a neutral solution?

S4. Give two properties of acids and two of bases.

S5. What is the difference between a strong acid and a weak acid?

Long Questions

L1. Explain the Brønsted-Lowry concept of acids and bases, giving examples of proton donation and acceptance.

L2. Describe the preparation and uses of three common salts.

Chapter 11: Environmental Chemistry-Air

Key Concepts

- The atmosphere is a mixture of gases: nitrogen about 78%, oxygen about 21%, argon about 0.9%, carbon dioxide about 0.04% and small amounts of other gases and water vapour.
- Air pollutants include carbon monoxide (CO), sulphur dioxide (SO_2), nitrogen oxides (NOx), particulate matter, ozone at ground level and volatile organic compounds. Main sources are vehicles, industry, burning of fossil fuels and crop residue.
- Carbon monoxide is a poisonous gas that binds with haemoglobin in blood and reduces oxygen transport. Sulphur dioxide and nitrogen oxides cause breathing problems and form acid rain.
- Acid rain forms when SO_2 and NOx dissolve in rainwater to make sulphuric and nitric acids. It damages buildings, harms crops and makes lakes acidic, killing aquatic life.
- The greenhouse effect is the warming of the Earth by gases (CO_2, methane, water vapour) that trap heat. Increased greenhouse gases cause global warming, rising sea levels and changing weather patterns.
- Ozone in the stratosphere (the ozone layer) absorbs harmful ultraviolet radiation. Chlorofluorocarbons (CFCs) damage the ozone layer; their use is now restricted.
- Smog is a mixture of smoke and fog, worsened by vehicle exhaust and sunlight. Air quality can be improved by using cleaner fuels, filters on chimneys, catalytic converters in vehicles, planting trees and reducing burning.

Key Terms

- **Atmosphere**
- **Air pollutant**
- **Acid rain**
- **Greenhouse effect**
- **Global warming**
- **Ozone layer**
- **Smog**
- **CFC**

Chapter 11: Environmental Chemistry-Air — Practice Questions

Multiple Choice Questions

Q1. The most abundant gas in air is

(A) oxygen (B) argon (C) nitrogen (D) carbon dioxide

Q2. Carbon monoxide is dangerous because it

(A) causes acid rain (B) binds with haemoglobin and reduces oxygen transport (C) damages the ozone layer (D) causes smog only

Q3. Acid rain is mainly caused by

(A) carbon dioxide only (B) sulphur dioxide and nitrogen oxides (C) oxygen and nitrogen (D) methane

Q4. Which gas is the main greenhouse gas produced by burning fossil fuels?

(A) Oxygen (B) Nitrogen (C) Carbon dioxide (D) Argon

Q5. The ozone layer protects life by absorbing

(A) infrared radiation (B) ultraviolet radiation (C) visible light (D) X-rays

Q6. Which substance damages the ozone layer?

(A) Carbon dioxide (B) Chlorofluorocarbons (CFCs) (C) Oxygen (D) Water vapour

Q7. Smog is formed by a mixture of

(A) smoke and fog (B) rain and snow (C) dust and oxygen (D) steam and carbon dioxide

Q8. The percentage of oxygen in clean dry air is about

(A) 78% (B) 21% (C) 0.9% (D) 0.04%

Q9. Global warming is mainly caused by an increase in

(A) oxygen (B) greenhouse gases (C) nitrogen (D) argon

Q10. Which action helps to reduce air pollution?

(A) Burning more coal (B) Using catalytic converters in vehicles (C) Cutting down trees (D) Burning crop residue

Short Questions

S1. Name the four main gases in air with their approximate percentages.

S2. What is acid rain? How is it formed?

S3. What is the greenhouse effect?

S4. Why is the ozone layer important?

S5. Name three air pollutants and one source of each.

Long Questions

L1. Explain how burning fossil fuels leads to acid rain and global warming, and describe their effects on the environment.

L2. Describe five practical measures that can improve air quality in a city.

Chapter 12: Environmental Chemistry-Water

Key Concepts

- Water is essential for life. Pure water is colourless, tasteless and odourless, boils at 100 °C and freezes at 0 °C. The water cycle (evaporation, condensation, precipitation, collection) keeps water moving between Earth and atmosphere.
- Water is a good solvent and is called the universal solvent because it dissolves many substances. Its polar molecules attract charged particles.
- Water pollution is caused by sewage, industrial waste, fertilisers and pesticides from farms, oil spills and plastic waste. Pollutants harm aquatic life and make water unsafe to drink.
- Biological oxygen demand (BOD) measures the oxygen used by micro-organisms to break down organic matter in water. High BOD means the water is heavily polluted with organic waste.
- Hard water contains dissolved calcium and magnesium ions. It forms scum with soap and scale in kettles and pipes. Temporary hardness (calcium hydrogencarbonate) is removed by boiling; permanent hardness is removed by ion-exchange or washing soda.
- Water purification steps: screening, sedimentation, filtration, chlorination (to kill germs) and sometimes fluoridation. Boiling also kills germs and is a simple household method.
- Saving water matters: fix leaks, use water wisely in homes and farms, recycle waste water where safe, and protect rivers and groundwater from pollution.

Key Terms

- **Water cycle**
- **Water pollution**
- **BOD**
- **Hard water**
- **Temporary hardness**
- **Permanent hardness**
- **Chlorination**
- **Filtration**

Chapter 12: Environmental Chemistry-Water — Practice Questions

Multiple Choice Questions

Q1. Pure water boils at

(A) 90 °C (B) 100 °C (C) 110 °C (D) 120 °C

Q2. Water is called the universal solvent because it

(A) dissolves all substances completely (B) dissolves many substances (C) is found everywhere (D) never freezes

Q3. A high BOD value in river water indicates

(A) clean water (B) heavy pollution by organic waste (C) hard water (D) salty water

Q4. Hard water is caused mainly by dissolved ions of

(A) sodium and potassium (B) calcium and magnesium (C) iron and zinc (D) chloride and nitrate

Q5. Temporary hardness of water can be removed by

(A) boiling (B) filtration only (C) adding salt (D) cooling

Q6. Chlorination of drinking water is done to

(A) remove hardness (B) kill germs (C) add taste (D) remove colour

Q7. Which of the following is a source of water pollution?

(A) Rainfall (B) Sewage and industrial waste (C) Evaporation (D) Condensation

Q8. Which step in water purification removes large suspended particles?

(A) Chlorination (B) Filtration (C) Fluoridation (D) Boiling

Q9. Hard water forms scum with soap because of

(A) dissolved oxygen (B) calcium and magnesium ions (C) dissolved carbon dioxide (D) chlorine

Q10. Which of these is part of the water cycle?

(A) Photosynthesis (B) Evaporation (C) Respiration (D) Combustion

Short Questions

S1. Why is water called the universal solvent?

S2. What is BOD and what does a high BOD tell us?

S3. What causes hard water? How is temporary hardness removed?

S4. List the main steps of purifying drinking water.

S5. Give two ways to save water at home.

Long Questions

L1. Describe the causes and effects of water pollution, and explain how polluted water can be made safe to drink.

L2. Explain the difference between temporary and permanent hardness of water and the methods used to remove each.

Chapter 13: Organic Chemistry

Key Concepts

- Organic chemistry is the study of carbon compounds, especially compounds of carbon with hydrogen (hydrocarbons) and their derivatives. Carbon forms the basis of all living things.
- Carbon can form four covalent bonds. Its atoms can join in chains, branches and rings, and can form single, double and triple bonds — this is why there are millions of organic compounds.
- A homologous series is a family of compounds with the same functional group, similar chemical properties, and each member differing from the next by a $–CH_2–$ group. Members show a gradual change in physical properties such as boiling point.
- Functional groups are atoms or groups of atoms that give a molecule its characteristic reactions, for example: –OH (alcohol), –COOH (carboxylic acid), –CHO (aldehyde), –CO– (ketone), $–NH_2$ (amine).
- Alkanes have the general formula C_nH_{2n+2} (single bonds only); alkenes C_nH_{2n} (one double bond); alkynes C_nH_{2n-2} (one triple bond).
- Structural formula shows how atoms are joined; molecular formula shows the number of each atom; condensed formula groups atoms together, for example, CH_3CH_2OH for ethanol.
- Organic compounds are used as fuels (natural gas, petrol), plastics, medicines, dyes, soaps, solvents and food ingredients.

Key Terms

- **Organic chemistry**
- **Homologous series**
- **Functional group**
- **Hydrocarbon**
- **Structural formula**
- **Molecular formula**

Chapter 13: Organic Chemistry — Practice Questions

Multiple Choice Questions

Q1. Organic chemistry is mainly the study of compounds of

(A) silicon (B) carbon (C) nitrogen (D) sulphur

Q2. Carbon can form a maximum of how many covalent bonds?

(A) 2 (B) 3 (C) 4 (D) 6

Q3. Members of a homologous series differ from each other by a

(A) $-CH_3$ group (B) $-CH_2-$ group (C) –OH group (D) –COOH group

Q4. The functional group of alcohols is

(A) –COOH (B) –CHO (C) –OH (D) $-NH_2$

Q5. The general formula of alkanes is

(A) C_nH_{2n} (B) C_nH_{2n+2} (C) C_nH_{2n-2} (D) C_nH_{2n+1}

Q6. Which of the following is a hydrocarbon?

(A) Ethanol (B) Ethanoic acid (C) Methane (D) Glucose

Q7. The functional group –COOH belongs to

(A) alcohols (B) carboxylic acids (C) aldehydes (D) ketones

Q8. Which statement about a homologous series is true?

(A) Members have different functional groups (B) Members have similar chemical properties (C) Members differ by $-CH_3$ (D) Boiling points are all the same

Q9. Ethanol is an example of a compound containing the functional group

(A) –OH (B) –COOH (C) –CHO (D) $-NH_2$

Q10. A structural formula shows

(A) only the number of atoms (B) how atoms are joined together (C) only the mass of the molecule (D) the colour of the compound

Short Questions

S1. What is organic chemistry?

S2. Why can carbon form so many compounds?

S3. What is a homologous series? Give one example.

S4. Name four functional groups with their formulae.

S5. Write the general formulae of alkanes, alkenes and alkynes.

Long Questions

L1. Explain homologous series and functional groups, giving examples to show how they help classify organic compounds.

L2. Describe the bonding in carbon compounds and why carbon can form chains, branches and rings.

Chapter 14: Hydrocarbons

Key Concepts

- Hydrocarbons are organic compounds containing only carbon and hydrogen. They are classified as saturated (alkanes, single bonds only) and unsaturated (alkenes and alkynes, with double or triple bonds).
- Alkanes: general formula C_nH_{2n+2}. Methane (CH_4), ethane (C_2H_6), propane (C_3H_8). They are unreactive except for combustion and substitution reactions.
- Alkenes: general formula C_nH_{2n}. Ethene (C_2H_4) is the simplest. They undergo addition reactions because of the double bond. The test for unsaturation is decolourisation of bromine water.
- Alkynes: general formula C_nH_{2n-2}. Ethyne (acetylene, C_2H_2) is the simplest; it is used in welding because it burns with a very hot flame.
- Combustion: complete combustion of a hydrocarbon gives carbon dioxide and water; incomplete combustion gives carbon monoxide (poisonous) and soot.
- Petroleum (crude oil) is a mixture of hydrocarbons. Fractional distillation separates it into fractions: refinery gas, petrol, kerosene, diesel, lubricating oil and bitumen, in order of increasing boiling point.
- Uses: methane as fuel (natural gas); ethene to make polythene and to ripen fruit; petrol and diesel as fuels; bitumen for roads; kerosene for lamps and jet fuel.

Key Terms

- **Hydrocarbon**
- **Saturated hydrocarbon**
- **Unsaturated hydrocarbon**
- **Alkane**
- **Alkene**
- **Alkyne**
- **Fractional distillation**
- **Combustion**

Chapter 14: Hydrocarbons — Practice Questions

Multiple Choice Questions

Q1. Which of the following is a saturated hydrocarbon?

(A) Ethene (B) Ethyne (C) Methane (D) Benzene

Q2. The general formula of alkenes is

(A) C_nH_{2n+2} (B) C_nH_{2n} (C) C_nH_{2n-2} (D) C_nH_{2n+1}

Q3. Unsaturation in a hydrocarbon can be tested with

(A) limewater (B) bromine water (C) litmus paper (D) sodium hydroxide

Q4. Complete combustion of methane produces

(A) carbon monoxide and water (B) carbon dioxide and water (C) carbon and hydrogen (D) carbon dioxide only

Q5. Ethyne (acetylene) is mainly used in

(A) making soap (B) welding (C) making glass (D) food preservation

Q6. Which fraction of crude oil has the lowest boiling point?

(A) Diesel (B) Kerosene (C) Petrol (D) Bitumen

Q7. The molecular formula of propane is

(A) CH_4 (B) C_2H_6 (C) C_3H_8 (D) C_4H_{10}

Q8. Incomplete combustion of a hydrocarbon produces the poisonous gas

(A) carbon dioxide (B) carbon monoxide (C) nitrogen dioxide (D) sulphur dioxide

Q9. Crude oil is separated into fractions by

(A) filtration (B) fractional distillation (C) evaporation (D) chromatography

Q10. Which alkene is used to make polythene?

(A) Methane (B) Ethene (C) Propene (D) Butene

Short Questions

S1. What are hydrocarbons? How are they classified?

S2. Write the molecular formulae of the first three alkanes.

S3. How can you test whether a hydrocarbon is unsaturated?

S4. What are the products of complete and incomplete combustion?

S5. Name four fractions obtained from crude oil.

Long Questions

L1. Describe fractional distillation of crude oil and the main uses of each fraction.

L2. Compare alkanes, alkenes and alkynes in terms of general formula, bonding and typical reactions.

Chapter 15: Biochemistry

Key Concepts

- Biochemistry is the study of the chemical processes and substances in living organisms. The main biomolecules are carbohydrates, proteins, lipids (fats and oils), nucleic acids and vitamins.
- Carbohydrates are compounds of carbon, hydrogen and oxygen, with the general pattern $C_n(H_2O)_m$. They are the main source of energy. Examples: glucose, starch, cellulose. Glucose is the monomer; starch is a polymer of glucose.
- Proteins are polymers of amino acids joined by peptide bonds. They build and repair body tissues, make enzymes and hormones, and provide energy when needed. Meat, eggs, milk and pulses are rich in protein.
- Lipids include fats and oils. They store energy, insulate the body and form cell membranes. Saturated fats (animal fats) have single bonds; unsaturated fats (vegetable oils) have double bonds.
- Enzymes are biological catalysts — proteins that speed up reactions in the body without being used up. They work best at a specific temperature and pH. Examples: amylase (digests starch), pepsin (digests protein).
- Vitamins are organic compounds needed in small amounts for health. Vitamin C prevents scurvy; vitamin D is needed for strong bones; vitamin A for good eyesight; B vitamins help release energy from food.
- Nucleic acids (DNA and RNA) store and pass on genetic information. DNA is the molecule of heredity; RNA helps make proteins.

Key Terms

- **Biochemistry**
- **Carbohydrate**
- **Protein**
- **Lipid**
- **Enzyme**
- **Vitamin**
- **Nucleic acid**
- **Amino acid**

Chapter 15: Biochemistry — Practice Questions

Multiple Choice Questions

Q1. The main energy-giving biomolecule is

(A) protein (B) carbohydrate (C) vitamin (D) mineral

Q2. Proteins are polymers of

(A) glucose (B) fatty acids (C) amino acids (D) nucleotides

Q3. Which enzyme digests starch?

(A) Pepsin (B) Amylase (C) Lipase (D) Trypsin

Q4. Lack of vitamin C causes

(A) rickets (B) scurvy (C) night blindness (D) anaemia

Q5. DNA is an example of a

(A) carbohydrate (B) protein (C) nucleic acid (D) lipid

Q6. Which of the following is a rich source of protein?

(A) Rice (B) Eggs (C) Sugar (D) Butter

Q7. Enzymes are best described as

(A) biological catalysts (B) energy stores (C) structural fats (D) vitamins

Q8. Starch is a polymer made from units of

(A) amino acids (B) glucose (C) fatty acids (D) glycerol

Q9. Vitamin D is needed for

(A) good eyesight (B) strong bones and teeth (C) healthy blood (D) energy release

Q10. Fats and oils belong to the group called

(A) carbohydrates (B) proteins (C) lipids (D) nucleic acids

Short Questions

S1. Define biochemistry and name the four main groups of biomolecules.

S2. What are enzymes? Give two examples with their functions.

S3. What is the difference between saturated and unsaturated fats?

S4. Name three vitamins, their sources and the disease caused by their deficiency.

S5. What is the function of DNA in a cell?

Long Questions

L1. Describe carbohydrates, proteins and lipids: their building units, functions and food sources.

L2. Explain the role of enzymes in the body and the factors that affect their activity.

Chapter 16: Empirical Data Collection and Analysis

Key Concepts

- Science depends on good data. Data are facts and measurements collected by observation or experiment. Qualitative data describe qualities; quantitative data are numbers with units.
- Before an experiment, identify the variables: the independent variable is changed by the experimenter, the dependent variable is measured, and controlled variables are kept constant. A fair test changes only one variable at a time.
- Measurements are made with instruments such as a balance (mass), measuring cylinder (volume), thermometer (temperature) and stopwatch (time). Every measurement has a unit and some uncertainty.
- Errors: systematic errors affect all readings in the same way (for example, a faulty balance); random errors vary from reading to reading. Repeating measurements and taking an average reduces random error.
- Results are organised in tables with headings and units. Data can be shown in graphs: a bar chart for categories, a line graph for a continuous variable, and a pie chart for parts of a whole.
- Mean (average) = sum of values ÷ number of values. Median is the middle value when data are arranged in order. Range = highest value − lowest value. These describe the centre and spread of data.
- Always draw conclusions from the data, not from opinions. A conclusion should answer the original question and state whether the hypothesis was supported.

Key Terms

- **Data**
- **Independent variable**
- **Dependent variable**
- **Controlled variable**
- **Systematic error**
- **Random error**
- **Mean**
- **Median**
- **Range**

Chapter 16: Empirical Data Collection and Analysis — Practice Questions

Multiple Choice Questions

Q1. The variable that the experimenter changes on purpose is the

(A) dependent variable (B) independent variable (C) controlled variable (D) constant

Q2. The variable that is measured in an experiment is the

(A) independent variable (B) dependent variable (C) controlled variable (D) error

Q3. A fair test changes

(A) all variables at once (B) only one variable at a time (C) no variables (D) only the controlled variables

Q4. Repeating a measurement and taking the average mainly reduces

(A) systematic error (B) random error (C) the number of readings (D) the size of the sample

Q5. The mean of 4, 6 and 8 is

(A) 4 (B) 6 (C) 8 (D) 18

Q6. The median of 3, 7, 9, 12, 15 is

(A) 7 (B) 9 (C) 12 (D) 3

Q7. The range of the data 5, 9, 14, 20 is

(A) 5 (B) 9 (C) 15 (D) 20

Q8. Which graph is best for showing how temperature changes with time?

(A) Pie chart (B) Bar chart (C) Line graph (D) Table only

Q9. A faulty balance that always reads 2 g too high causes a

(A) random error (B) systematic error (C) correct reading (D) zero error only

Q10. Which of these is quantitative data?

(A) The liquid is blue (B) The liquid smells sweet (C) The liquid has a volume of 25 cm^3 (D) The liquid is thick

Short Questions

S1. What is the difference between qualitative and quantitative data?

S2. Define independent, dependent and controlled variables with an example.

S3. What is the difference between a systematic error and a random error?

S4. How is the mean of a set of readings calculated?

S5. When should you use a line graph and when a bar chart?

Long Questions

L1. Design a simple experiment to find how temperature affects the rate at which salt dissolves, naming all the variables and explaining how you would make it a fair test.

L2. Explain how data are collected, organised, presented in graphs and analysed (mean, median, range) to reach a conclusion.

Chapter 17: Separation Techniques

Key Concepts

- Mixtures can be separated by physical methods because their components keep their own properties. The method chosen depends on the properties of the substances, such as particle size, solubility, boiling point or magnetism.
- Filtration separates an insoluble solid from a liquid using filter paper; the solid stays as the residue and the liquid passes through as the filtrate.
- Evaporation is used to recover a dissolved solid from a solution by heating the solvent away. Crystallisation is a slower, controlled method that gives pure crystals of the solid.
- Distillation separates liquids with different boiling points, or a solvent from a dissolved solid. Simple distillation is used for one liquid; fractional distillation for a mixture of liquids.
- Sublimation separates a substance that sublimes (such as ammonium chloride or iodine) from one that does not, by heating the mixture and collecting the vapour on a cold surface.
- Magnetic separation uses a magnet to remove iron or steel from a mixture. Decantation pours off a liquid from a settled solid. A separating funnel separates two immiscible liquids such as oil and water.
- Chromatography separates dissolved substances that move at different speeds through a medium; it is covered in detail in Chapter 19.

Key Terms

- **Filtration**
- **Residue**
- **Filtrate**
- **Evaporation**
- **Crystallisation**
- **Distillation**
- **Sublimation**
- **Decantation**

Chapter 17: Separation Techniques — Practice Questions

Multiple Choice Questions

Q1. Filtration is used to separate

(A) a dissolved solid from a liquid (B) an insoluble solid from a liquid (C) two miscible liquids (D) two gases

Q2. In filtration, the liquid that passes through the filter paper is called the

(A) residue (B) filtrate (C) solute (D) solvent

Q3. Which method is best for obtaining pure crystals of salt from sea water?

(A) Filtration (B) Crystallisation (C) Magnetic separation (D) Decantation

Q4. Simple distillation separates liquids on the basis of their different

(A) densities (B) boiling points (C) colours (D) masses

Q5. Sublimation can separate ammonium chloride from sand because ammonium chloride

(A) is magnetic (B) dissolves in water (C) changes directly from solid to gas on heating (D) melts easily

Q6. A separating funnel is used to separate

(A) two immiscible liquids (B) a solid and a liquid (C) two soluble solids (D) two gases

Q7. Which method would remove iron filings from a mixture of sand and iron filings?

(A) Filtration (B) Evaporation (C) Magnetic separation (D) Distillation

Q8. Evaporation is used to recover

(A) a liquid from a mixture of liquids (B) a dissolved solid from its solution (C) a gas from air (D) an insoluble solid

Q9. Decantation works best when the solid has

(A) dissolved in the liquid (B) settled at the bottom (C) floated on top (D) evaporated

Q10. The choice of separation method depends mainly on the

(A) colour of the mixture (B) physical properties of the components (C) amount of the mixture (D) smell of the components

Short Questions

S1. What is filtration? What are the residue and the filtrate?

S2. How is evaporation different from crystallisation?

S3. When is simple distillation used? Give an example.

S4. How can ammonium chloride be separated from sand?

S5. How would you separate oil from water?

Long Questions

L1. Describe, with a diagram in words, how you would obtain pure water from salty water by distillation.

L2. A mixture contains sand, salt and iron filings. Describe step by step how you would separate all three components.

Chapter 18: Qualitative Analysis

Key Concepts

- Qualitative analysis identifies which substances are present in a sample, without measuring how much of each there is. It uses observations such as colour changes, gas formation and precipitates.
- The flame test identifies some metal cations: lithium gives a crimson-red flame, sodium a golden-yellow flame, potassium a lilac flame, calcium a brick-red flame and copper a blue-green flame. A clean platinum wire dipped in the sample is held in a blue Bunsen flame.
- Ammonium ions (NH_4^+) are detected by warming the sample with sodium hydroxide: ammonia gas is released, which turns damp red litmus paper blue and has a sharp smell.
- Chloride ions (Cl^-) are detected by adding dilute nitric acid followed by silver nitrate solution: a white precipitate of silver chloride forms, which dissolves in dilute ammonia solution.
- Sulphate ions (SO_4^{2-}) are detected by adding dilute hydrochloric acid followed by barium chloride solution: a white precipitate of barium sulphate forms.
- Carbonate ions (CO_3^{2-}) are detected by adding dilute acid: carbon dioxide gas is released, which turns limewater milky.
- Hydrogen gas burns with a 'pop' sound. Oxygen relights a glowing splint. Carbon dioxide turns limewater milky. Ammonia turns damp red litmus blue.

Key Terms

- **Qualitative analysis**
- **Flame test**
- **Precipitate**
- **Cation**
- **Anion**
- **Limewater**

Chapter 18: Qualitative Analysis — Practice Questions

Multiple Choice Questions

Q1. Qualitative analysis is used to find out

(A) how much of a substance is present (B) which substances are present (C) the mass of a sample (D) the temperature of a reaction

Q2. A golden-yellow flame in the flame test shows the presence of

(A) potassium (B) sodium (C) calcium (D) lithium

Q3. Which ion gives a white precipitate with silver nitrate solution?

(A) Sulphate (B) Carbonate (C) Chloride (D) Nitrate

Q4. The white precipitate formed in the sulphate test is

(A) silver chloride (B) barium sulphate (C) calcium carbonate (D) lead chloride

Q5. Carbon dioxide gas is identified because it

(A) burns with a pop (B) relights a glowing splint (C) turns limewater milky (D) turns blue litmus red

Q6. Ammonia gas turns damp red litmus paper

(A) red (B) blue (C) white (D) green

Q7. Which flame colour is shown by potassium ions?

(A) Golden-yellow (B) Crimson-red (C) Lilac (D) Brick-red

Q8. Hydrogen gas is tested with a

(A) glowing splint (B) lighted splint, giving a pop sound (C) limewater (D) litmus paper

Q9. Before the chloride test with silver nitrate, dilute nitric acid is added to

(A) dissolve the precipitate (B) remove carbonate ions that would interfere (C) make the solution alkaline (D) heat the sample

Q10. A precipitate is best described as

(A) a gas that escapes (B) an insoluble solid formed in a solution (C) a liquid that evaporates (D) a colour change only

Short Questions

S1. What is qualitative analysis?

S2. What flame colours are given by sodium and potassium ions?

S3. How would you test for chloride ions in a solution?

S4. How would you test for carbonate ions?

S5. Name the test and the observation for hydrogen gas.

Long Questions

L1. Describe the flame test and explain how it is used to identify metal cations, giving the colours for lithium, sodium, potassium and calcium.

L2. Describe the chemical tests for chloride, sulphate and ammonium ions, giving the reagent used and the observation in each case.

Chapter 19: Chromatography

Key Concepts

- Chromatography is a separation technique used to separate and identify the components of a mixture. It works because different substances move at different speeds through a medium.
- In paper chromatography, a spot of the mixture is placed on chromatography paper and the bottom edge is dipped in a solvent. The solvent rises up the paper by capillary action, carrying the dissolved substances with it. Substances that dissolve well and stick weakly to the paper travel furthest.
- The Rf value (retention factor) identifies a substance: Rf = distance travelled by the substance ÷ distance travelled by the solvent front. Rf is always less than 1 and has no unit.
- Two substances are the same if they have the same Rf value under the same conditions. Chromatography can show whether a sample is pure (one spot) or a mixture (several spots).
- Types of chromatography: paper chromatography, thin-layer chromatography (TLC), column chromatography and gas chromatography. Gas chromatography is used for volatile substances and in forensic and environmental testing.
- Uses: checking the purity of medicines and food colours, separating dyes in inks, identifying drugs, testing water and food samples, and forensic analysis of samples from crime scenes.

Key Terms

- **Chromatography**
- **Rf value**
- **Solvent front**
- **Mobile phase**
- **Stationary phase**
- **Chromatogram**

Chapter 19: Chromatography — Practice Questions

Multiple Choice Questions

Q1. Chromatography separates the components of a mixture because they

(A) have different masses only (B) move at different speeds through the medium (C) have different colours (D) react with each other

Q2. In paper chromatography, the solvent moves up the paper by

(A) evaporation (B) capillary action (C) filtration (D) condensation

Q3. The Rf value is calculated as

(A) distance of solvent front ÷ distance of substance (B) distance of substance ÷ distance of solvent front (C) distance of substance + distance of solvent front (D) distance of substance × distance of solvent front

Q4. An Rf value is always

(A) greater than 1 (B) less than 1 (C) equal to 1 (D) equal to 0

Q5. If a spot travels 4 cm and the solvent front travels 8 cm, the Rf value is

(A) 0.2 (B) 0.5 (C) 2.0 (D) 4.0

Q6. Two substances are likely to be the same if they have the same

(A) colour (B) Rf value under the same conditions (C) spot size (D) solvent

Q7. A pure substance in paper chromatography gives

(A) one spot (B) two spots (C) many spots (D) no spots

Q8. Which type of chromatography is used for volatile substances?

(A) Paper chromatography (B) Thin-layer chromatography (C) Gas chromatography (D) Column chromatography

Q9. Chromatography is used in the food industry to

(A) cook food (B) check the purity of food colours (C) add flavour (D) preserve food

Q10. In paper chromatography, the stationary phase is the

(A) solvent (B) chromatography paper (C) mixture (D) beaker

Short Questions

S1. What is chromatography?

S2. Define the Rf value and give its formula.

S3. How can chromatography show whether a sample is pure?

S4. Name three types of chromatography.

S5. Give three uses of chromatography.

Long Questions

L1. Describe, step by step, how paper chromatography is carried out, and explain how the Rf value is used to identify substances.

L2. A student spots black ink on chromatography paper and finds three coloured spots. Explain what this shows and how the student could identify the dyes.

MCQ Answer Key (Chapters 1–10)

Answers to the multiple-choice questions, chapters 1 to 10.

Chapter 1 (Nature of Science in Chemistry): Q1: **B** Q2: **C** Q3: **C** Q4: **A** Q5: **C** Q6: **B** Q7: **C** Q8: **B** Q9: **B** Q10: **B**

Chapter 2 (Matter): Q1: **C** Q2: **B** Q3: **C** Q4: **B** Q5: **C** Q6: **B** Q7: **C** Q8: **C** Q9: **A** Q10: **C**

Chapter 3 (Atomic Structure): Q1: **B** Q2: **C** Q3: **C** Q4: **B** Q5: **C** Q6: **B** Q7: **B** Q8: **B** Q9: **B** Q10: **C**

Chapter 4 (Periodic Table and Periodicity of Properties): Q1: **B** Q2: **D** Q3: **C** Q4: **C** Q5: **B** Q6: **C** Q7: **B** Q8: **B** Q9: **B** Q10: **C**

Chapter 5 (Chemical Bonding): Q1: **B** Q2: **C** Q3: **C** Q4: **C** Q5: **B** Q6: **B** Q7: **B** Q8: **C** Q9: **B** Q10: **C**

Chapter 6 (Stoichiometry): Q1: **B** Q2: **C** Q3: **A** Q4: **B** Q5: **B** Q6: **B** Q7: **B** Q8: **C** Q9: **B** Q10: **C**

Chapter 7 (Electrochemistry): Q1: **B** Q2: **B** Q3: **B** Q4: **B** Q5: **C** Q6: **B** Q7: **C** Q8: **C** Q9: **B** Q10: **B**

Chapter 8 (Energetics): Q1: **B** Q2: **B** Q3: **C** Q4: **B** Q5: **B** Q6: **B** Q7: **B** Q8: **B** Q9: **C** Q10: **C**

Chapter 9 (Chemical Equilibrium): Q1: **B** Q2: **A** Q3: **B** Q4: **B** Q5: **B** Q6: **B** Q7: **B** Q8: **B** Q9: **C** Q10: **B**

Chapter 10 (Acids, Bases and Salts): Q1: **B** Q2: **B** Q3: **B** Q4: **B** Q5: **C** Q6: **B** Q7: **B** Q8: **C** Q9: **C** Q10: **B**

MCQ Answer Key (Chapters 11–19)

Answers to the multiple-choice questions, chapters 11 to 19.

Chapter 11 (Environmental Chemistry-Air): Q1: **C** Q2: **B** Q3: **B** Q4: **C** Q5: **B** Q6: **B** Q7: **A** Q8: **B** Q9: **B** Q10: **B**

Chapter 12 (Environmental Chemistry-Water): Q1: **B** Q2: **B** Q3: **B** Q4: **B** Q5: **A** Q6: **B** Q7: **B** Q8: **B** Q9: **B** Q10: **B**

Chapter 13 (Organic Chemistry): Q1: **B** Q2: **C** Q3: **B** Q4: **C** Q5: **B** Q6: **C** Q7: **B** Q8: **B** Q9: **A** Q10: **B**

Chapter 14 (Hydrocarbons): Q1: **C** Q2: **B** Q3: **B** Q4: **B** Q5: **B** Q6: **C** Q7: **C** Q8: **B** Q9: **B** Q10: **B**

Chapter 15 (Biochemistry): Q1: **B** Q2: **C** Q3: **B** Q4: **B** Q5: **C** Q6: **B** Q7: **A** Q8: **B** Q9: **B** Q10: **C**

Chapter 16 (Empirical Data Collection and Analysis): Q1: **B** Q2: **B** Q3: **B** Q4: **B** Q5: **B** Q6: **B** Q7: **C** Q8: **C** Q9: **B** Q10: **C**

Chapter 17 (Separation Techniques): Q1: **B** Q2: **B** Q3: **B** Q4: **B** Q5: **C** Q6: **A** Q7: **C** Q8: **B** Q9: **B** Q10: **B**

Chapter 18 (Qualitative Analysis): Q1: **B** Q2: **B** Q3: **C** Q4: **B** Q5: **C** Q6: **B** Q7: **C** Q8: **B** Q9: **B** Q10: **B**

Chapter 19 (Chromatography): Q1: **B** Q2: **B** Q3: **B** Q4: **B** Q5: **B** Q6: **B** Q7: **A** Q8: **C** Q9: **B** Q10: **B**

> Scoring guide: 9–10 correct in a chapter means the chapter is well prepared. 7–8 correct means a quick revision of the notes is enough. Below 7, re-read the chapter notes and try the questions again.

These notes and questions are original study material prepared for students of FBISE Class 9 Chemistry, following the NBF book chapters. For the official syllabus and the current model paper, visit fbise.edu.pk.