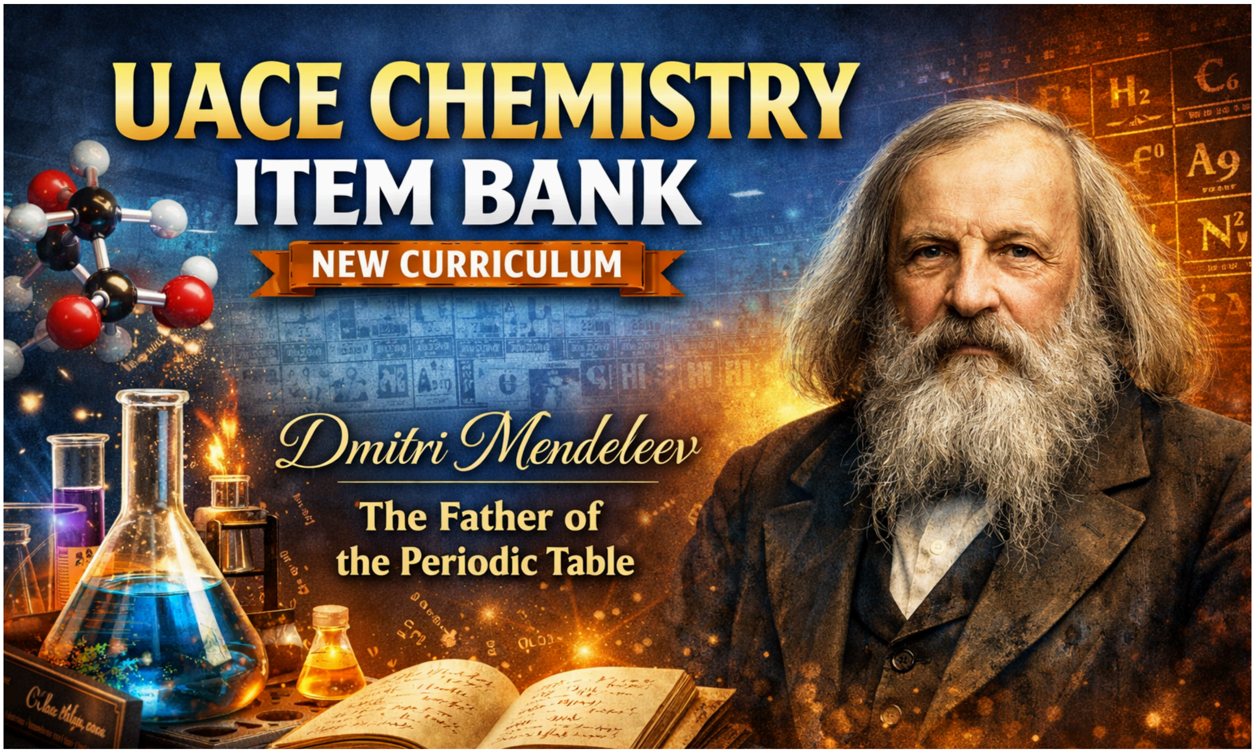

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For permissions, feedback, or inquiries, you can contact: [joelamanyire3@gmail.com or Whatsapp number +256788477510]

TERM 1

S5 CHEMISTRY – MOLES AND EQUATIONS

Item 1
A technician prepares oxygen gas in the laboratory by heating potassium chlorate(V) in the presence of manganese(IV) oxide as a catalyst.

a) Write a balanced chemical equation for the reaction.
b) Calculate the number of moles of potassium chlorate(V) that decompose when 4.90 g of the solid is heated.
c) Determine the volume of oxygen gas produced at room temperature and pressure.
d) State the role of manganese(IV) oxide in this reaction.

Item 2
A sample of calcium carbonate reacts completely with excess dilute hydrochloric acid to produce carbon dioxide gas.

a) Write a balanced chemical equation for the reaction.
b) If 10.0 g of calcium carbonate is used, calculate the mass of carbon dioxide formed.
c) Determine the volume of the carbon dioxide at r.t.p.
d) Explain why excess acid is used in this experiment.

Item 3
Nitrogen reacts with hydrogen in the Haber process to form ammonia.

a) Write a balanced equation for the reaction.
b) Calculate the number of moles of ammonia formed when 14.0 g of nitrogen reacts completely.
c) Determine the volume of ammonia produced at r.t.p.
d) State two conditions used in the Haber process and explain their importance.

Item 4
A student dissolves 5.30 g of sodium carbonate in water and makes the solution up to 250 cm^3.

a) Calculate the number of moles of sodium carbonate used.
b) Determine the concentration of the solution in mol dm^{-3}.
c) Calculate the mass of sodium ions present in the solution.
d) Explain why sodium carbonate is considered a basic salt.

Item 5
Excess zinc granules react with 100 cm^3 of 1.0 mol dm^{-3} hydrochloric acid.

a) Write a balanced chemical equation for the reaction.

b) Calculate the number of moles of hydrochloric acid used.
c) Determine the mass of zinc that reacts completely.
d) Calculate the volume of hydrogen gas produced at r.t.p.

Item 6
Sulfur dioxide is produced when copper reacts with hot concentrated sulfuric acid.

a) Write a balanced chemical equation for the reaction.
b) Calculate the mass of copper required to produce 0.50 mol of sulfur dioxide.
c) Determine the volume of sulfur dioxide produced at r.t.p.
d) State one environmental problem caused by sulfur dioxide.

Item 7
A mixture contains sodium chloride and sodium carbonate only. The total mass of the mixture is 10.6 g. When the mixture is treated with excess dilute hydrochloric acid, 2.24 dm^3 of carbon dioxide is produced at r.t.p.

a) Write the equation for the reaction producing carbon dioxide.
b) Calculate the number of moles of carbon dioxide formed.
c) Determine the mass of sodium carbonate in the mixture.
d) Calculate the mass of sodium chloride present.

Item 8
Iron reacts with steam at high temperature to form an oxide and hydrogen gas.

a) Write a balanced chemical equation for the reaction.
b) Calculate the number of moles of iron that react when 11.2 dm^3 of hydrogen is produced at r.t.p.
c) Determine the mass of iron used.
d) State one industrial importance of this reaction.

Item 9
A sample of hydrated copper(II) sulfate has the formula $CuSO_4 \cdot xH_2O$. When 5.00 g of the hydrated salt is heated strongly, 3.20 g of anhydrous copper(II) sulfate remains.

a) Calculate the mass of water lost on heating.
b) Determine the number of moles of water lost.
c) Calculate the value of x in the formula.
d) State one use of hydrated copper(II) sulfate in the laboratory.

Item 10
Magnesium ribbon burns in air to form magnesium oxide.

a) Write a balanced chemical equation for the reaction.

b) Calculate the mass of magnesium oxide formed when 2.40 g of magnesium burns completely.
c) Determine the number of moles of oxygen used.
d) Explain why magnesium burns with a bright white flame.

ANSWERS – MOLES AND EQUATIONS

Item 1 – Answers
a) $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$
b) Moles of $KClO_3$ = 4.90 ÷ 122.5 = 0.040 mol
c) Moles of O_2 = (3/2) × 0.040 = 0.060 mol
Volume of O_2 at r.t.p = 0.060 × 22.4 = 1.34 dm^3
d) MnO_2 acts as a catalyst to increase the rate of decomposition.

Item 2 – Answers
a) $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$
b) Moles of $CaCO_3$ = 10.0 ÷ 100 = 0.10 mol
Mass of CO_2 = 0.10 × 44 = 4.4 g
c) Volume of CO_2 at r.t.p = 0.10 × 22.4 = 2.24 dm^3
d) Excess acid ensures complete reaction of calcium carbonate.

Item 3 – Answers
a) $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
b) Moles of N_2 = 14.0 ÷ 28 = 0.50 mol
Moles of NH_3 = 2 × 0.50 = 1.0 mol
c) Volume of NH_3 at r.t.p = 1.0 × 22.4 = 22.4 dm^3
d) Conditions: high pressure (increases yield), iron catalyst (increases rate).

Item 4 – Answers
a) Moles of Na_2CO_3 = 5.30 ÷ 106 = 0.050 mol
b) Concentration = 0.050 ÷ 0.250 = 0.20 mol dm^{-3}
c) Mass of Na^+ = 0.10 mol × 23 = 2.3 g
d) Sodium carbonate forms OH^- ions in water.

Item 5 – Answers
a) $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
b) Moles of HCl = 1.0 × 0.100 = 0.10 mol
c) Moles of Zn = 0.10 ÷ 2 = 0.050 mol
Mass of Zn = 0.050 × 65 = 3.25 g
d) Volume of H_2 = 0.050 × 22.4 = 1.12 dm^3

Item 6 – Answers
a) $Cu(s) + 2H_2SO_4(conc) \rightarrow CuSO_4(aq) + SO_2(g) + 2H_2O(l)$
b) Moles of Cu = 0.50 mol

Mass of Cu = 0.50 × 63.5 = 31.8 g
c) Volume of SO_2 = 0.50 × 22.4 = 11.2 dm^3
d) Sulfur dioxide causes acid rain.

Item 7 – Answers
a) $Na_2CO_3(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$
b) Moles of CO_2 = 2.24 ÷ 22.4 = 0.10 mol
c) Mass of Na_2CO_3 = 0.10 × 106 = 10.6 g
d) Mass of NaCl = 10.6 − 10.6 = 0.0 g

Item 8 – Answers
a) $3Fe(s) + 4H_2O(g) \rightarrow Fe_3O_4(s) + 4H_2(g)$
b) Moles of H_2 = 11.2 ÷ 22.4 = 0.50 mol
Moles of Fe = (3/4) × 0.50 = 0.375 mol
c) Mass of Fe = 0.375 × 56 = 21.0 g
d) Used in hydrogen production.

Item 9 – Answers
a) Mass of water lost = 5.00 − 3.20 = 1.80 g
b) Moles of H_2O = 1.80 ÷ 18 = 0.10 mol
c) Moles of $CuSO_4$ = 3.20 ÷ 160 = 0.020 mol
x = 0.10 ÷ 0.020 = 5
d) Used to test for water.

Item 10 – Answers
a) $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$
b) Moles of Mg = 2.40 ÷ 24 = 0.10 mol
Mass of MgO = 0.10 × 40 = 4.0 g
c) Moles of O_2 = 0.10 ÷ 2 = 0.050 mol
d) Magnesium releases a large amount of energy on oxidation.

📘 Author: joelPCM

S5 CHEMISTRY – ATOMIC STRUCTURE

Item 1
At a national research laboratory, scientists were investigating an element extracted from a mineral found in western Uganda. The element was found to be highly reactive and quickly formed ions when dissolved in water. During analysis, scientists discovered that each atom of the element contained 11 protons and several energy levels occupied by electrons. When the element reacted, it easily lost one electron from its outermost shell, forming a stable ion. The researchers also noticed that the element's behavior was closely related to its electronic arrangement and position in the periodic table.

a) Determine the electronic configuration of the atom.
b) State the group and period of the element.
c) Explain why the element easily forms a positive ion.
d) Write the electronic configuration of the ion formed.

Item 2
A radioactive substance was stored in a hospital laboratory for use in cancer treatment. The technicians observed that after every 8 days, the amount of the radioactive material reduced to half its previous quantity. Over time, the radiation emitted decreased steadily, making the substance safer to handle. Doctors emphasized the importance of understanding radioactive decay so that the substance could be used effectively without exposing patients and workers to unnecessary danger.

a) Define the term half-life.
b) Calculate the fraction of the substance remaining after 24 days.
c) State one medical use of radioactive isotopes.
d) Explain why radioactive materials must be carefully shielded.

Item 3
Two students were given samples of the same element collected from different locations. Laboratory analysis showed that both samples had atoms with the same number of protons but different numbers of neutrons. Although the atoms behaved identically in chemical reactions, their masses were slightly different. This observation helped the students understand why atomic masses listed in the periodic table are not whole numbers.

a) Define the term isotope.
b) Explain why the two samples had identical chemical properties.
c) State one reason why isotopes have different mass numbers.
d) Explain why relative atomic mass is usually a decimal.

Item 4
During a practical lesson, learners were shown a mass spectrum produced from a gaseous element. The spectrum displayed two distinct peaks of different heights. The teacher explained that the peaks represented atoms of the same element with different masses. The relative heights of the peaks were used to calculate the average atomic mass of the element.

a) Identify the particles detected in a mass spectrometer.
b) Explain why two peaks appeared in the spectrum.
c) State what determines the height of each peak.
d) Explain how relative atomic mass is calculated from a mass spectrum.

Item 5

In a physics–chemistry collaboration, students passed an electric discharge through hydrogen gas at low pressure. When the light produced was viewed through a spectroscope, several distinct colored lines were observed instead of a continuous spectrum. The instructor explained that this experiment provided strong evidence about how electrons are arranged in atoms and how energy changes occur.

a) Define emission spectrum.
b) Explain why hydrogen produces a line spectrum.
c) State what happens to electrons when light is emitted.
d) Explain what the lines reveal about atomic energy levels.

Item 6
A senior chemist explained the Bohr model using diagrams drawn on the board. He described electrons moving in fixed circular orbits around the nucleus and stated that electrons could only exist in certain allowed energy levels. When an electron moved from a higher level to a lower one, energy was released as light. However, he also mentioned that the model could not explain the spectra of larger atoms.

a) State two postulates of the Bohr atomic model.
b) Explain how the model accounts for atomic spectra.
c) State one limitation of the Bohr model.
d) Suggest why the model works best for hydrogen.

Item 7
A student examined the electronic configuration of an atom and noticed that its outermost shell contained seven electrons. The atom readily gained one electron during reactions to achieve a stable noble gas configuration. This behavior explained why the element was highly reactive and commonly found as a negative ion in compounds.

a) Write the electronic configuration of the atom.
b) Explain why the atom forms a negative ion.
c) State the group to which the element belongs.
d) Write the electronic configuration of the ion formed.

Item 8
During nuclear studies, learners were introduced to alpha radiation. They were told that alpha particles consist of heavy, positively charged particles emitted from unstable nuclei. Although alpha radiation was found to be highly ionizing, it could be stopped easily by a sheet of paper.

a) Describe the nature of alpha particles.
b) Explain why alpha radiation has low penetrating power.
c) State one effect of alpha radiation on living tissue.
d) Give one practical use of alpha radiation.

Item 9

In another experiment, beta radiation was demonstrated using a radioactive source. Students learned that beta particles were emitted when a neutron inside the nucleus changed into another particle. The atomic number of the atom increased, yet the mass number remained unchanged.

a) Describe what happens during beta decay.
b) State the change in atomic number during beta decay.
c) Explain why the mass number remains constant.
d) Give one use of beta radiation.

Item 10

A chemistry teacher described how electrons occupy regions called orbitals rather than fixed paths. She emphasized that orbitals represent regions of high probability of finding electrons. The concept helped learners understand atomic structure more accurately than earlier models.

a) Define the term atomic orbital.
b) Distinguish between an orbital and an energy level.
c) State the maximum number of electrons in the third energy level.
d) Explain why electrons occupy lower energy levels first.

Item 11

Two ions were compared during a lesson: one had lost two electrons while the other had gained one electron. Although both ions had noble-gas configurations, their sizes were different. The teacher explained that nuclear charge and number of shells played a major role.

a) Define the term ionic radius.
b) Explain why cations are smaller than their parent atoms.
c) Explain why anions are larger than their parent atoms.
d) State one factor affecting ionic size.

Item 12

A radioactive element was observed to emit radiation continuously without any external influence. The decay rate was unaffected by temperature, pressure, or chemical reactions. This made radioactive decay unique compared to ordinary chemical reactions.

a) Define radioactive decay.
b) Explain why decay rate is unaffected by external conditions.
c) State one danger of radioactive materials.
d) State one safety precaution when handling radioactive substances.

Item 13

A student calculated the relative atomic mass of an element using data on isotopic abundances. The calculation involved multiplying each isotope's mass by its percentage abundance and dividing by 100. The result closely matched the value in the periodic table.

a) Define relative atomic mass.
b) Explain why isotopic abundance is considered in calculations.
c) State one reason why relative atomic mass is important.
d) State one limitation of relative atomic mass.

Item 14
During ionization experiments, energy was supplied to gaseous atoms to remove electrons. The teacher explained that the energy required depended on nuclear charge, atomic radius, and shielding effect.

a) Define ionization energy.
b) Explain why ionization energy increases across a period.
c) Explain why ionization energy decreases down a group.
d) State one use of ionization energy data.

Item 15
In a discussion on nuclear stability, students learned that heavy nuclei often undergo radioactive decay because the balance between nuclear forces becomes unstable. This instability leads to the emission of radiation as the nucleus attempts to become more stable.

a) Define nuclear stability.
b) Explain why heavy nuclei are often radioactive.
c) State two factors affecting nuclear stability.
d) Explain how radioactive decay leads to more stable nuclei.

📘 Author: joelPCM

ANSWERS – ATOMIC STRUCTURE

Item 1 – Answers
a) Electronic configuration: $1s^2\ 2s^2\ 2p^6\ 3s^1$
b) Group 1, Period 3
c) It loses one outer electron to attain a stable noble gas configuration
d) Ion configuration: $1s^2\ 2s^2\ 2p^6$

Item 2 – Answers
a) Half-life is the time taken for half the nuclei in a radioactive sample to decay
b) After 24 days = 3 half-lives → fraction remaining = $(1/2)^3 = 1/8$
c) Used in cancer treatment (radiotherapy)

d) To protect people from harmful ionizing radiation

Item 3 – Answers

a) Isotopes are atoms of the same element with the same atomic number but different mass numbers

b) They have the same electronic configuration

c) They contain different numbers of neutrons

d) Because it is a weighted average of isotopic masses

Item 4 – Answers

a) Positive ions (cations)

b) Due to presence of isotopes of different masses

c) The relative abundance of each isotope

d) By taking the weighted average of isotopic masses

Item 5 – Answers

a) Emission spectrum is a set of discrete wavelengths emitted by excited atoms

b) Because electrons exist in fixed energy levels

c) Electrons fall from higher to lower energy levels

d) Energy levels in atoms are quantized

Item 6 – Answers

Electrons move in fixed energy levels

Energy is emitted or absorbed when electrons change levels

b) Energy released corresponds to specific wavelengths

c) It cannot explain spectra of multi-electron atoms

d) Hydrogen has only one electron

Item 7 – Answers

a) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

b) It gains one electron to complete the outer shell

c) Group 17 (halogens)

d) Ion configuration: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Item 8 – Answers

a) Alpha particles are helium nuclei (2 protons, 2 neutrons)

b) They are heavy and lose energy quickly

c) They cause severe ionization of tissues

d) Used in smoke detectors

Item 9 – Answers

a) A neutron changes into a proton and an electron

b) Atomic number increases by 1

c) Total number of nucleons remains the same
d) Used in thickness control of materials

Item 10 – Answers
a) An atomic orbital is a region of space with high probability of finding an electron
b) Orbitals are regions; energy levels are shells
c) Maximum electrons = 18
d) Lower levels are more stable and require less energy

Item 11 – Answers
a) Ionic radius is the size of an ion
b) Loss of electrons reduces electron–electron repulsion
c) Gain of electrons increases repulsion
d) Nuclear charge

Item 12 – Answers
a) Radioactive decay is the spontaneous breakdown of unstable nuclei
b) It depends on nuclear forces, not external conditions
c) Can damage living tissues
d) Use lead shielding

Item 13 – Answers
a) Relative atomic mass is the weighted average mass of isotopes compared to 1/12 of carbon-12
b) Because isotopes occur in different proportions
c) Used in chemical calculations
d) It does not represent actual mass of a single atom

Item 14 – Answers
a) Ionization energy is the energy required to remove one mole of electrons from gaseous atoms
b) Nuclear charge increases across a period
c) Atomic radius and shielding increase down a group
d) Used to predict reactivity

Item 15 – Answers
a) Nuclear stability is the tendency of a nucleus to remain unchanged
b) Repulsion between protons becomes large
c) Neutron–proton ratio, nuclear size
d) Emission of radiation lowers energy and increases stability

📘 *Author: joelPCM*

S5 CHEMISTRY – BONDING AND STRUCTURE

Item 1

In a chemical manufacturing plant, engineers were comparing two solids used for different purposes. One solid was very hard, had a high melting point, and did not conduct electricity, while the other was soft, slippery, and could conduct electricity along certain directions. Both substances were made entirely of carbon atoms. The engineers explained that the difference in properties was due to the way atoms were bonded and arranged in space rather than the type of atoms present.

a) Name the two forms of carbon described.
b) Describe the type of bonding present in each substance.
c) Explain why one substance conducts electricity while the other does not.
d) Relate the structure of each substance to its physical properties.

Item 2

A student heated sodium chloride strongly in a crucible and observed that it did not melt easily. However, when the solid was dissolved in water or melted, it conducted electricity. The teacher used this observation to introduce the concept of ionic bonding and lattice structures.

a) Explain how ionic bonds are formed in sodium chloride.
b) Describe the structure of solid sodium chloride.
c) Explain why solid sodium chloride does not conduct electricity.
d) Explain why molten or aqueous sodium chloride conducts electricity.

Item 3

During a laboratory discussion, learners compared hydrogen chloride gas and sodium chloride solid. Both compounds contained chlorine, yet one existed as a gas at room temperature while the other was a solid with a very high melting point. The teacher emphasized that the difference lay in the type of bonding and forces between particles.

a) State the type of bonding in hydrogen chloride.
b) State the type of bonding in sodium chloride.
c) Explain why hydrogen chloride has a low boiling point.
d) Compare the intermolecular forces present in the two substances.

Item 4

A group of students examined the structures of methane, ammonia, and water using molecular models. They noticed that although all three molecules had similar numbers of electrons around the central atom, their shapes were different. The teacher explained that lone pairs played a significant role in determining molecular shape.

a) State the shapes of methane, ammonia, and water.
b) Explain why ammonia is not tetrahedral.

c) Explain why water has a bent shape.
d) State the effect of lone pairs on bond angles.

Item 5
In a materials science lesson, students learned about metals used in construction. They were told that metals could be bent, stretched, and shaped easily without breaking. This behavior was explained using the model of metallic bonding.

a) Describe metallic bonding.
b) Explain why metals are good conductors of electricity.
c) Explain why metals are malleable and ductile.
d) State one limitation of the metallic bonding model.

Item 6
A chemist compared diamond and silicon dioxide (sand) and noted that both substances were extremely hard and had very high melting points. Despite being made of different elements, they shared similar bonding characteristics.

a) Describe the type of structure present in diamond.
b) Describe the type of structure present in silicon dioxide.
c) Explain why both substances have high melting points.
d) State one difference between the two structures.

Item 7
During a lesson on covalent bonding, learners studied chlorine gas. They were told that each chlorine atom contributed one electron to form a shared pair. This sharing allowed each atom to achieve a stable electronic configuration.

a) Define a covalent bond.
b) Explain why chlorine forms a diatomic molecule.
c) Draw a dot-and-cross diagram for chlorine gas.
d) State one property of simple covalent substances.

Item 8
In an experiment, solid iodine was gently heated and observed to change directly into a purple vapor without forming a liquid. The teacher used this example to explain intermolecular forces.

a) Name the process observed when iodine changes directly into vapor.
b) Describe the forces between iodine molecules.
c) Explain why iodine sublimes easily.
d) Compare these forces with those in ionic compounds.

Item 9

A student dissolved ethanol in water and observed that the two liquids mixed completely. However, when oil was added to water, the liquids formed separate layers. The teacher explained this using intermolecular forces and polarity.

a) State the type of bonding present in ethanol.
b) Explain why ethanol is soluble in water.
c) Explain why oil is insoluble in water.
d) Define hydrogen bonding.

Item 10
In a chemistry lecture, the teacher explained that some molecules have uneven charge distribution even though they are electrically neutral overall. Water was used as an example to demonstrate polarity.

a) Define a polar molecule.
b) Explain why water is a polar molecule.
c) State one consequence of polarity in water.
d) Give one example of a non-polar molecule.

Item 11
A factory used calcium oxide to remove acidic impurities from waste gases. Students were asked to explain why calcium oxide readily reacted with acids and had a high melting point.

a) State the type of bonding in calcium oxide.
b) Describe the structure of calcium oxide.
c) Explain why calcium oxide has a high melting point.
d) State one use of calcium oxide based on its properties.

Item 12
During revision, learners compared giant covalent structures with simple molecular substances. They observed that giant structures were solids with very high melting points, while simple molecules were often gases or liquids.

a) Define a giant covalent structure.
b) Give one example of a giant covalent substance.
c) Explain why giant covalent structures have high melting points.
d) Contrast this with simple molecular substances.

Item 13
A student noticed that graphite could be used as a lubricant and also as an electrode. The teacher explained that this was due to its unique layered structure.

a) Describe the bonding within layers of graphite.
b) Describe the forces between layers.

c) Explain why graphite is slippery.
d) Explain why graphite conducts electricity.

Item 14
In a bonding experiment, magnesium reacted with oxygen to form magnesium oxide. The product was a white solid with a very high melting point.

a) Describe how magnesium and oxygen form ions.
b) Explain the electrostatic attraction in magnesium oxide.
c) Describe the lattice structure formed.
d) State one property resulting from this structure.

Item 15
A teacher summarized bonding by explaining that the type of bond formed between atoms depends on their electronic configuration and electronegativity difference. This ultimately determines the physical properties of substances.

a) State two factors that influence the type of bonding.
b) Explain how electronegativity difference affects bond type.
c) Distinguish between ionic and covalent bonding.
d) Explain why bonding determines physical properties.

ANSWERS – BONDING AND STRUCTURE

Item 1 – Answers
a) Diamond and graphite
b) Both have giant covalent bonding
c) Graphite conducts electricity due to delocalized electrons; diamond does not because all electrons are used in bonding
d) Diamond has a rigid 3-D structure making it hard; graphite has layers that slide over each other making it soft

Item 2 – Answers
a) Sodium loses one electron; chlorine gains one electron forming Na^+ and Cl^- ions
b) A giant ionic lattice of oppositely charged ions
c) Ions are fixed in position and cannot move
d) Ions are free to move in molten or aqueous state

Item 3 – Answers
a) Covalent bonding
b) Ionic bonding

c) Weak intermolecular forces between HCl molecules
d) HCl has van der Waals forces; NaCl has strong electrostatic forces

Item 4 – Answers
a) Methane – tetrahedral; Ammonia – trigonal pyramidal; Water – bent
b) Presence of one lone pair distorts shape
c) Two lone pairs cause greater repulsion
d) Lone pairs reduce bond angles more than bonding pairs

Item 5 – Answers
a) Attraction between positive metal ions and delocalized electrons
b) Free electrons move and carry charge
c) Layers of ions slide while bonding remains intact
d) Does not explain differences in metal hardness

Item 6 – Answers
a) Giant covalent structure
b) Giant covalent network
c) Many strong covalent bonds must be broken
d) Diamond is pure carbon; SiO_2 contains silicon and oxygen

Item 7 – Answers
a) A bond formed by sharing electrons
b) Each chlorine needs one electron to complete its octet
c) Diagram shows shared pair between two Cl atoms
d) Low melting and boiling points

Item 8 – Answers
a) Sublimation
b) Weak van der Waals forces
c) Weak intermolecular forces require little energy to overcome
d) Ionic compounds have strong electrostatic forces

Item 9 – Answers
a) Covalent bonding
b) Hydrogen bonding forms between ethanol and water
c) Oil is non-polar and cannot form hydrogen bonds
d) Attraction between hydrogen and highly electronegative atoms

Item 10 – Answers
a) A molecule with uneven charge distribution
b) Oxygen is more electronegative than hydrogen and molecule is bent
c) Strong hydrogen bonding and high boiling point
d) Oxygen (O_2) or methane (CH_4)

Item 11 – Answers
a) Ionic bonding
b) Giant ionic lattice of Ca^{2+} and O^{2-} ions
c) Strong electrostatic attractions require much energy
d) Neutralizing acidic gases

Item 12 – Answers
a) A structure with many atoms joined by covalent bonds
b) Diamond or silicon dioxide
c) Strong covalent bonds throughout structure
d) Simple molecules have weak intermolecular forces

Item 13 – Answers
a) Strong covalent bonds within layers
b) Weak van der Waals forces between layers
c) Layers slide easily over each other
d) Delocalized electrons move within layers

Item 14 – Answers
a) Magnesium loses two electrons; oxygen gains two
b) Strong attraction between Mg^{2+} and O^{2-} ions
c) Giant ionic lattice
d) High melting point and hardness

Item 15 – Answers
a) Electronic configuration and electronegativity difference
b) Large difference favors ionic bonding
c) Ionic involves electron transfer; covalent involves sharing
d) Bond strength and structure affect melting point, conductivity, and solubility

📘 Author: joelPCM

TERM 2
S5 CHEMISTRY – PERIODICITY

Item 1
During a visit to an industrial chemical store, students observed several elements arranged according to increasing atomic number. They noticed that elements with similar chemical properties appeared at regular intervals. The instructor explained that this repeating pattern was not accidental but was linked to electronic configuration and atomic structure. As they moved across one row of the table, the elements became less metallic and more reactive as non-metals.

a) Define the term periodicity.
b) Explain why elements show periodic repetition of properties.
c) Describe how metallic character changes across a period.
d) Explain the electronic reason for this trend.

Item 2
A laboratory technician compared sodium and chlorine, two elements found in the same period but at opposite ends of the periodic table. Sodium reacted violently with water, while chlorine was a toxic green gas. The technician emphasized that despite being in the same period, their properties were extremely different due to differences in their outer-shell electrons.

a) State the electronic configuration of sodium and chlorine.
b) Explain why their chemical properties differ greatly.
c) Describe how atomic radius changes across the period.
d) Explain how nuclear charge influences this change.

Item 3
In a revision lesson, students compared lithium, sodium, and potassium. When placed in water, lithium reacted slowly, sodium reacted vigorously, and potassium reacted explosively. The teacher linked this observation to atomic size and ease of electron loss.

a) State the group to which these elements belong.
b) Describe the trend in reactivity down the group.
c) Explain this trend in terms of atomic radius.
d) Explain the role of shielding effect.

Item 4
A student noticed that the first ionization energy of magnesium was higher than that of sodium. However, aluminium showed a lower ionization energy than magnesium, which seemed unexpected. The teacher explained that electronic configuration and subshells must be considered.

a) Define first ionization energy.
b) Explain why ionization energy increases from sodium to magnesium.
c) Explain why aluminium has a lower ionization energy than magnesium.
d) State one general trend of ionization energy across a period.

Item 5
During a practical session, students observed that chlorine gained electrons easily during reactions, while sodium readily lost electrons. The teacher explained that electronegativity played a key role in this behavior.

a) Define electronegativity.

b) Describe how electronegativity changes across a period.
c) Explain why chlorine has a high electronegativity.
d) State one consequence of high electronegativity in bonding.

Item 6
A chemist studying halogens compared fluorine, chlorine, bromine, and iodine. She observed that fluorine was the most reactive, while iodine reacted least readily. The explanation involved atomic size and attraction for electrons.

a) State the group to which halogens belong.
b) Describe the trend in reactivity down the group.
c) Explain why reactivity decreases down the group.
d) Explain the role of atomic radius in this trend.

Item 7
When examining oxides of elements across Period 3, students noticed that sodium oxide was strongly basic, magnesium oxide was weakly basic, aluminium oxide was amphoteric, and sulfur dioxide was acidic. The teacher used this to demonstrate periodic trends in oxide behavior.

a) Define the term amphoteric.
b) Describe the trend in nature of oxides across Period 3.
c) Explain why aluminium oxide is amphoteric.
d) Relate this trend to bonding type.

Item 8
A science club investigated atomic radii using data from the periodic table. They observed that atomic radius decreased steadily across a period but increased down a group. This pattern was explained using nuclear charge and number of shells.

a) Define atomic radius.
b) Describe the trend of atomic radius across a period.
c) Explain why atomic radius decreases across a period.
d) Explain why atomic radius increases down a group.

Item 9
During a lesson on electron affinity, students learned why some atoms readily gain electrons while others do not. The teacher explained that energy changes occur when electrons are added to atoms.

a) Define electron affinity.
b) Describe the trend of electron affinity across a period.
c) Explain why halogens have high electron affinity.
d) State one limitation of electron affinity data.

Item 10
A teacher explained why noble gases are unreactive and exist as single atoms. Their electronic structures were highlighted as the main reason for this unusual stability.

a) State the electronic configuration of a noble gas.
b) Explain why noble gases are chemically inert.
c) Explain why noble gases have very high ionization energies.
d) State one practical use of noble gases.

Item 11
In comparing metallic elements across Period 3, students observed that electrical conductivity decreased from sodium to aluminium, and then dropped sharply as non-metals were reached.

a) Define metallic character.
b) Describe the trend in metallic character across a period.
c) Explain this trend using electron loss.
d) Explain why metals conduct electricity.

Item 12
A student noticed that melting points of elements varied across Period 3. Sodium had a low melting point, silicon had a very high melting point, while argon had a very low melting point.

a) Describe the bonding in silicon.
b) Explain why silicon has a high melting point.
c) Explain why argon has a very low melting point.
d) Relate melting point trends to structure and bonding.

Item 13
When studying chlorides of Period 3 elements, learners observed that sodium chloride was ionic, magnesium chloride had some covalent character, and aluminium chloride was largely covalent.

a) Explain why bonding changes across the period.
b) Describe the trend in polarization power across the period.
c) Explain why aluminium chloride is covalent.
d) State one property resulting from covalent character.

Item 14
A chemistry teacher emphasized that periodic trends are interrelated. Changes in atomic radius, ionization energy, and electronegativity all influence reactivity.

a) Explain the relationship between atomic radius and ionization energy.
b) Explain how electronegativity affects bond type.
c) Explain how ionization energy affects metallic character.

d) Give one example showing interrelation of periodic trends.

Item 15
At the end of the topic, students were asked to justify why the periodic table is one of the most important tools in chemistry. The teacher explained that understanding periodicity allows chemists to predict properties of unknown elements.

a) State the modern periodic law.
b) Explain how electronic configuration determines position in the periodic table.
c) Explain how periodic trends help predict chemical behavior.
d) State one limitation of the periodic table.

📘 Author: joelPCM

ANSWERS – PERIODICITY

Item 1 – Answers
a) Periodicity is the regular repetition of physical and chemical properties of elements when arranged in order of increasing atomic number.
b) Because elements with similar outer-shell electronic configurations occur at regular intervals.
c) Metallic character decreases across a period.
d) Nuclear charge increases across the period, making electron loss more difficult.

Item 2 – Answers
a) Sodium: $1s^2\ 2s^2\ 2p^6\ 3s^1$
Chlorine: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$
b) They have different numbers of valence electrons.
c) Atomic radius decreases across the period.
d) Increasing nuclear charge pulls electrons closer to the nucleus.

Item 3 – Answers
a) Group 1 (alkali metals).
b) Reactivity increases down the group.
c) Atomic radius increases down the group, making electron loss easier.
d) Increased shielding reduces nuclear attraction on the outer electron.

Item 4 – Answers
a) First ionization energy is the energy required to remove one mole of electrons from one mole of gaseous atoms.
b) Increased nuclear charge in magnesium holds electrons more strongly.
c) Aluminium loses an electron from a higher-energy 3p subshell.
d) Ionization energy generally increases across a period.

Item 5 – Answers

a) Electronegativity is the ability of an atom to attract electrons in a covalent bond.

b) It increases across a period.

c) Chlorine has high nuclear charge and small atomic radius.

d) Leads to formation of polar covalent or ionic bonds.

Item 6 – Answers

a) Group 17 (halogens).

b) Reactivity decreases down the group.

c) Increased atomic radius reduces attraction for an incoming electron.

d) Larger atoms shield the nucleus more effectively.

Item 7 – Answers

a) Amphoteric substances react with both acids and bases.

b) Oxides change from basic → amphoteric → acidic across the period.

c) Aluminium oxide has both ionic and covalent character.

d) Bonding changes from ionic to covalent across the period.

Item 8 – Answers

a) Atomic radius is half the distance between nuclei of two bonded atoms.

b) Atomic radius decreases across a period.

c) Increased nuclear charge pulls electrons closer.

d) Addition of new energy levels increases size down a group.

Item 9 – Answers

a) Electron affinity is the energy change when an atom gains an electron.

b) Electron affinity becomes more negative across a period.

c) Halogens readily gain electrons to complete their octet.

d) Noble gases do not form stable negative ions.

Item 10 – Answers

a) A full outer shell ($ns^2\ np^6$).

b) They have stable electronic configurations.

c) Strong nuclear attraction makes electron removal difficult.

d) Argon is used in welding; neon in advertising signs.

Item 11 – Answers

a) Metallic character is the tendency of an atom to lose electrons.

b) Metallic character decreases across a period.

c) Increasing ionization energy makes electron loss difficult.

d) Metals have delocalized electrons that move freely.

Item 12 – Answers

a) Giant covalent bonding.
b) Strong covalent bonds throughout the lattice.
c) Argon has weak van der Waals forces only.
d) Stronger bonding leads to higher melting points.

Item 13 – Answers
a) Increasing nuclear charge increases polarization.
b) Polarizing power increases across the period.
c) Al^{3+} strongly polarizes the chloride ion.
d) Low melting point and volatility.

Item 14 – Answers
a) Smaller atomic radius increases ionization energy.
b) High electronegativity favors covalent bonding.
c) Low ionization energy increases metallic character.
d) Alkali metals: large radius, low IE, high reactivity.

Item 15 – Answers
a) Properties of elements are periodic functions of their atomic numbers.
b) Electronic configuration determines group and period.
c) Trends allow prediction of reactivity and bonding.
d) Does not explain all transition-metal behavior.

📘 Author: joelPCM

S5 CHEMISTRY – THERMOCHEMISTRY

Item 1
At a rural secondary school, students carried out an experiment where magnesium ribbon was burned in excess oxygen inside a crucible. The laboratory temperature rose sharply, and the students observed a bright white flame. After the experiment, the teacher explained that heat changes accompanying chemical reactions are studied under thermochemistry and are vital in understanding energy flow in reactions.

a) Define thermochemistry.
b) State the type of reaction occurring when magnesium burns.
c) Explain why the temperature of the surroundings increased.
d) Sketch and label an energy profile diagram for the reaction.

Item 2
During an industrial visit, students learned how cement factories rely heavily on energy changes during chemical reactions. Limestone was heated strongly, and carbon dioxide gas was released. The process required continuous heating to proceed.

a) Write a balanced chemical equation for the reaction described.
b) State whether the reaction is exothermic or endothermic.
c) Explain your answer using bond energy concepts.
d) State one industrial importance of this reaction.

Item 3
In a calorimetry experiment, a student dissolved ammonium nitrate in water and noticed that the temperature of the solution dropped significantly. The beaker felt cold to touch, surprising many learners.

a) State the type of reaction involved.
b) Explain why the temperature decreased.
c) Define enthalpy change of solution.
d) Explain one practical application of this process.

Item 4
A chemistry teacher demonstrated Hess's Law using several reactions involving carbon, carbon monoxide, and carbon dioxide. The students were told that direct measurement of some enthalpy changes is difficult, hence indirect methods are used.

a) State Hess's Law.
b) Explain why Hess's Law is valid.
c) Using suitable equations, explain how enthalpy change of formation of CO can be determined.
d) State one limitation of calorimetric measurements.

Item 5
Students compared the burning of methane and ethanol in air. They noticed differences in flame color, heat released, and soot formation. The teacher related this to enthalpy changes of combustion.

a) Define enthalpy change of combustion.
b) Explain why complete combustion releases more energy than incomplete combustion.
c) State two conditions required for complete combustion.
d) Explain why ethanol produces less soot than methane.

Item 6
In a revision class, learners were told that breaking chemical bonds requires energy, while forming bonds releases energy. This explanation helped them understand why some reactions absorb energy while others release it.

a) Define bond energy.
b) Explain how bond energies can be used to calculate enthalpy change.

c) State the formula used in such calculations.
d) Explain why calculated enthalpy values are approximate.

Item 7
A student attempted to determine the enthalpy change of neutralization between hydrochloric acid and sodium hydroxide using a simple calorimeter made from a polystyrene cup. Heat losses were expected.

a) Define enthalpy change of neutralization.
b) State the expected sign of the enthalpy change.
c) Explain why polystyrene is used as a calorimeter.
d) Suggest one source of experimental error.

Item 8
At a sugar factory, heat exchangers are used to control temperature during processing. The engineers explained that some reactions must be cooled to prevent damage to equipment.

a) Explain why temperature control is important in chemical reactions.
b) State the relationship between temperature and kinetic energy.
c) Explain how temperature affects reaction enthalpy measurements.
d) State one industrial consequence of uncontrolled exothermic reactions.

Item 9
Students were introduced to standard conditions when studying enthalpy changes. They learned that comparisons are only meaningful when measurements are made under the same conditions.

a) Define standard enthalpy change.
b) State the standard conditions used.
c) Explain why standard conditions are necessary.
d) State one example of a standard enthalpy change.

Item 10
During an experiment, zinc reacted with dilute hydrochloric acid in an open beaker. Heat was released, but not all of it was captured by the measuring device.

a) Write the equation for the reaction.
b) Explain why not all heat released was measured.
c) State one method of reducing heat loss.
d) Explain how heat loss affects calculated enthalpy values.

Item 11
A teacher explained that fuels are chosen based on the amount of energy they release per mole or per gram when burned. Students compared coal, petrol, and ethanol.

a) Define calorific value.
b) Explain why fuels with high calorific values are preferred.
c) State one disadvantage of using fossil fuels.
d) Explain why ethanol is considered a cleaner fuel.

Item 12
Learners observed that dissolving sodium hydroxide pellets in water caused the solution temperature to rise rapidly. The container became warm.

a) State the type of enthalpy change involved.
b) Explain why heat is released during dissolution.
c) Define hydration energy.
d) Explain the role of hydration energy in this process.

Item 13
In a classroom discussion, students compared enthalpy change of formation and enthalpy change of combustion. The teacher emphasized that definitions must be precise.

a) Define enthalpy change of formation.
b) State the standard conditions for this enthalpy change.
c) Explain why elements must be in their standard states.
d) Give one example of an enthalpy change of formation.

Item 14
A chemistry club investigated why some reactions start slowly despite being exothermic. The teacher introduced the concept of activation energy using energy profile diagrams.

a) Define activation energy.
b) Explain why activation energy is required.
c) Explain how catalysts affect activation energy.
d) State one industrial advantage of using catalysts.

Item 15
At the end of the topic, students were asked to explain the importance of thermochemistry in real life, especially in industries and environmental studies.

a) State one importance of thermochemistry in industry.
b) Explain how thermochemistry helps in fuel selection.
c) Explain the role of thermochemistry in environmental protection.
d) State one limitation of thermochemical data.

📘 Author: joelPCM

ANSWERS – THERMOCHEMISTRY

Item 1 – Answers

a) Thermochemistry is the study of heat changes that accompany chemical reactions.
b) An exothermic reaction.
c) Heat energy is released to the surroundings when new bonds form.
d) Energy profile shows products at lower energy than reactants with ΔH negative.

Item 2 – Answers

a) $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
b) Endothermic reaction.
c) More energy is required to break bonds than is released during bond formation.
d) Manufacture of quicklime for cement production.

Item 3 – Answers

a) Endothermic reaction.
b) Energy is absorbed from the surroundings to break ionic bonds.
c) Enthalpy change when one mole of solute dissolves in excess solvent.
d) Used in instant cold packs.

Item 4 – Answers

a) The total enthalpy change of a reaction is independent of the route taken.
b) Energy is conserved in all chemical reactions.
c) Using enthalpy changes of formation of CO_2 and C, enthalpy of CO is calculated indirectly.
d) Heat loss to surroundings.

Item 5 – Answers

a) Enthalpy change when one mole of a substance burns completely in oxygen.
b) Complete combustion forms stronger bonds and releases more energy.
c) Sufficient oxygen and high temperature.
d) Ethanol contains oxygen which aids complete combustion.

Item 6 – Answers

a) Bond energy is the energy required to break a bond in gaseous molecules.
b) By comparing energy absorbed and energy released.
c) $\Delta H = \Sigma(\text{bonds broken}) - \Sigma(\text{bonds formed})$
d) Average bond energies are used.

Item 7 – Answers

a) Enthalpy change when one mole of water is formed from an acid and base.
b) Negative (exothermic).
c) It minimizes heat loss due to insulation.

d) Heat loss to surroundings.

Item 8 – Answers

a) To prevent excessive heat that may damage equipment.
b) Temperature is directly proportional to kinetic energy.
c) Heat loss or gain affects accuracy of measurements.
d) Explosion or equipment failure.

Item 9 – Answers

a) Enthalpy change measured under standard conditions.
b) 298 K, 1 atm pressure, 1 mol dm^{-3} solutions.
c) To allow fair comparison of enthalpy values.
d) Standard enthalpy of combustion.

Item 10 – Answers

a) $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
b) Heat escapes to the surroundings.
c) Use an insulated calorimeter with a lid.
d) Enthalpy change is underestimated.

Item 11 – Answers

a) Energy released per unit mass or mole of fuel.
b) They produce more energy for the same amount of fuel.
c) Environmental pollution.
d) Produces less carbon monoxide and soot.

Item 12 – Answers

a) Enthalpy change of solution.
b) Hydration energy exceeds lattice energy.
c) Energy released when ions are hydrated.
d) It releases heat during dissolution.

Item 13 – Answers

a) Enthalpy change when one mole of compound forms from its elements.
b) 298 K and 1 atm.
c) To maintain consistency and accuracy.
d) Formation of water from hydrogen and oxygen.

Item 14 – Answers

a) Minimum energy required to start a reaction.
b) To break initial bonds.
c) Catalysts lower activation energy.
d) Increases reaction rate and saves energy.

Item 15 – Answers
a) Helps design efficient industrial processes.
b) Determines energy output of fuels.
c) Helps reduce pollution by choosing cleaner reactions.
d) Experimental errors affect accuracy.

📘 Author: joelPCM

S5 CHEMISTRY – ORGANIC CHEMISTRY I

Item 1
At a fuel storage depot near Kampala, students observed different petroleum fractions being stored in separate tanks. The engineer explained that these fractions are mixtures of organic compounds mainly composed of carbon and hydrogen. He further explained that organic chemistry is a branch of chemistry that focuses on such compounds and their reactions. During the discussion, the engineer emphasized that carbon forms a very large number of compounds compared to other elements, a fact that puzzled the learners.

a) Define organic chemistry.
b) Explain why carbon forms a large number of compounds.
c) State two properties that distinguish organic compounds from inorganic compounds.
d) Explain the concept of catenation.

Item 2
During a lesson on hydrocarbons, a teacher described a family of organic compounds obtained from crude oil that are saturated and relatively unreactive. The teacher explained that these compounds are used as fuels and lubricants because of their stability. A sample compound was said to contain 6 carbon atoms.

a) Name the homologous series described.
b) State the general formula of this homologous series.
c) Write the molecular formula of the compound described.
d) Calculate the relative molecular mass of the compound.

Item 3
A student visiting a gas station noticed that LPG burns with a clean blue flame while kerosene produces a yellow smoky flame. The station manager explained that the difference is due to molecular structure and completeness of combustion.

a) State the main type of bonding present in hydrocarbons.
b) Explain why LPG burns with a blue flame.
c) Explain why kerosene may produce a smoky flame.

d) State one environmental effect of incomplete combustion.

Item 4

In a chemistry practical, learners heated an organic liquid with bromine water in the absence of light. The brown color of bromine disappeared instantly. The teacher warned that the reaction must be done carefully and away from sunlight to avoid different reactions occurring.

a) State the class of organic compound present.
b) Name the type of reaction that occurred.
c) Explain why the reaction occurred rapidly.
d) State why sunlight must be avoided.

Item 5

A petrochemical engineer explained that alkenes are more reactive than alkanes due to the presence of a particular type of bond. He demonstrated this by reacting ethene with hydrogen in the presence of a nickel catalyst at high temperature.

a) State the general formula of alkenes.
b) Name the type of bond responsible for alkene reactivity.
c) Name the reaction taking place.
d) Write a balanced equation for the reaction.

Item 6 (CALCULATION)

A student burned 0.84 g of an alkane completely in excess oxygen and obtained 2.64 g of carbon dioxide and 1.08 g of water. The student was told that the compound belongs to the alkane homologous series.

a) Calculate the empirical formula of the alkane.
b) Determine the molecular formula if its molar mass is 72 g mol^{-1}.
c) Name the alkane.
d) State one use of this alkane.

Item 7

In a factory producing plastics, large quantities of ethene gas are converted into a solid polymer under controlled conditions. The engineer explained that the process involves opening of multiple bonds and formation of long chains.

a) Name the process described.
b) State the type of polymer formed.
c) Explain the role of the double bond in the reaction.
d) State two uses of the polymer formed.

Item 8

During a revision lesson, students compared alkanes, alkenes, and alkynes. The teacher stressed that despite being hydrocarbons, their chemical behavior differs significantly due to differences in bonding.

a) State the general formula of alkynes.
b) State the type of bond present in alkynes.
c) Explain why alkynes are more reactive than alkanes.
d) State one test that distinguishes alkenes from alkanes.

Item 9
At a roadside mechanical workshop, fuel leakage caused a fire that spread rapidly. Firefighters explained that fuels with shorter carbon chains ignite more easily than those with longer chains.

a) Explain how chain length affects volatility.
b) State the trend in boiling point with increasing chain length.
c) Explain this trend in terms of intermolecular forces.
d) State one safety precaution when handling volatile hydrocarbons.

Item 10 (CALCULATION)
A hydrocarbon contains 85.7% carbon and 14.3% hydrogen by mass. The compound is gaseous at room temperature and belongs to the alkene series.

a) Calculate the empirical formula of the compound.
b) Determine the molecular formula if the molar mass is 28 g mol^{-1}.
c) Name the compound.
d) State one industrial use of the compound.

Item 11
A chemistry teacher demonstrated homologous series using molecular models. Students observed that members of a series differed by a constant unit and showed gradual changes in physical properties.

a) Define a homologous series.
b) State the common difference between successive members.
c) Explain why boiling points increase down a homologous series.
d) State two characteristics of a homologous series.

Item 12
In a school laboratory, ethanol was dehydrated using concentrated sulfuric acid to produce a gas that turned bromine water colorless. The reaction required careful temperature control.

a) Name the gas produced.
b) State the type of reaction involved.
c) Write a balanced equation for the reaction.

d) State one condition necessary for the reaction.

Item 13
During a field trip, students learned that unsaturated hydrocarbons can be converted into saturated ones to improve fuel quality. The process involved heat, pressure, and a metal catalyst.

a) Name the process described.
b) State the catalyst used.
c) Explain why the process improves fuel quality.
d) State one disadvantage of the process.

Item 14
A teacher explained that organic compounds can undergo substitution or addition reactions depending on their structure. Examples were given using methane and ethene.

a) State the type of reaction methane undergoes with chlorine.
b) State the conditions required for this reaction.
c) State the type of reaction ethene undergoes with bromine.
d) Explain the difference between the two reaction types.

Item 15 (CALCULATION)
A hydrocarbon has a vapour density of 21 and belongs to the alkane homologous series. The compound is widely used as a fuel.

a) Calculate the molar mass of the hydrocarbon.
b) Determine the molecular formula of the compound.
c) Name the hydrocarbon.
d) State two uses of the compound.

📘 Author: joelPCM

ANSWERS – ORGANIC CHEMISTRY I

Item 1 – Answers
a) Organic chemistry is the branch of chemistry that deals with carbon compounds, excluding carbonates, oxides of carbon, and carbides.
b) Carbon is tetravalent and can form four covalent bonds, allowing formation of long chains and rings.
c) Organic compounds have low melting points and are usually insoluble in water.
d) Catenation is the ability of carbon atoms to bond to one another forming long chains.

Item 2 – Answers

a) Alkanes.
b) $C_{\square}H_{2\square+2}$
c) C_6H_{14}
d) Relative molecular mass = (6 × 12) + (14 × 1) = 72 + 14 = 86

Item 3 – Answers
a) Covalent bonding.
b) LPG undergoes complete combustion producing CO_2 and H_2O only.
c) Kerosene undergoes incomplete combustion producing carbon particles.
d) Incomplete combustion causes air pollution and carbon monoxide poisoning.

Item 4 – Answers
a) Alkene.
b) Addition reaction.
c) Presence of a carbon–carbon double bond which is reactive.
d) Sunlight causes substitution reactions with bromine.

Item 5 – Answers
a) $C_{\square}H_{2\square}$
b) Carbon–carbon double bond (C=C).
c) Hydrogenation.
d) $C_2H_4 + H_2 \rightarrow C_2H_6$

Item 6 – Answers (CALCULATION)
a)
CO_2 mass = 2.64 g → moles = 2.64 ÷ 44 = 0.06
Carbon moles = 0.06

H_2O mass = 1.08 g → moles = 1.08 ÷ 18 = 0.06
Hydrogen moles = 0.06 × 2 = 0.12

Ratio C:H = 0.06 : 0.12 = 1 : 2
Empirical formula = CH_2

b) Empirical formula mass = 14
Molecular mass = 72
n = 72 ÷ 14 ≈ 5
Molecular formula = C_5H_{10}

c) Pentene
d) Used in manufacture of plastics.

Item 7 – Answers

a) Addition polymerization.
b) Polyethene.
c) The double bond opens to form long chains.
d) Making plastic bags and bottles.

Item 8 – Answers
a) $C_{\square}H_{2\square-2}$
b) Carbon–carbon triple bond.
c) Presence of multiple bonds increases reactivity.
d) Bromine water test.

Item 9 – Answers
a) Shorter chains have weaker intermolecular forces.
b) Boiling point increases with chain length.
c) Stronger van der Waals forces require more energy to overcome.
d) Store away from open flames.

Item 10 – Answers (CALCULATION)

a)
C: 85.7 ÷ 12 = 7.14
H: 14.3 ÷ 1 = 14.3

Ratio ≈ 1 : 2
Empirical formula = CH_2

b) Empirical mass = 14
28 ÷ 14 = 2
Molecular formula = C_2H_4

c) Ethene
d) Used in manufacture of plastics.

Item 11 – Answers
a) A homologous series is a family of organic compounds with similar chemical properties.
b) $–CH_2–$
c) Increasing intermolecular forces.
d) Same functional group and same general formula.

Item 12 – Answers
a) Ethene.
b) Elimination (dehydration).
c) $C_2H_5OH \rightarrow C_2H_4 + H_2O$

d) Concentrated H_2SO_4 at 170°C.

Item 13 – Answers
a) Hydrogenation.
b) Nickel catalyst.
c) Converts unsaturated fuels into more stable saturated fuels.
d) Requires high temperature and pressure.

Item 14 – Answers
a) Substitution reaction.
b) Ultraviolet light.
c) Addition reaction.
d) Substitution replaces atoms; addition adds atoms across a double bond.

Item 15 – Answers (CALCULATION)
a) Vapour density = 21
Molar mass = 21 × 2 = 42 g mol^{-1}

b) Alkane formula: C_nH_{2n+2}
(12n + 2n + 2) = 42
14n = 40 → n ≈ 3
Molecular formula = C_3H_8

c) Propane
d) Cooking gas and fuel.

📘 Author: joelPCM
📘 UACE CHEMISTRY ITEM V

S5 CHEMISTRY – EQUILIBRIA I

Item 1
At a fertilizer manufacturing plant near Jinja, students on an industrial visit observed a large reaction vessel operating continuously under controlled temperature and pressure. Inside the reactor, nitrogen and hydrogen gases were constantly being introduced, while ammonia gas was continuously removed. The plant engineer explained that the reaction never goes to completion but instead reaches a state where the forward and backward reactions occur simultaneously. The students were surprised to learn that even when the amounts of reactants and products appear constant, reactions are still taking place at the molecular level. The engineer further explained that altering conditions such as pressure or temperature could shift the position of equilibrium to favor more ammonia production. The discussion later continued in class, where learners were asked to analyze the nature of this equilibrium system.

a) Define chemical equilibrium.
b) Write a balanced equation for the reaction described.
c) Explain what is meant by a dynamic equilibrium.
d) State two conditions necessary for equilibrium to be established.
e) Explain why ammonia is continuously removed in the industrial process.

Item 2
In a school laboratory, a sealed glass container held a mixture of nitrogen dioxide gas and dinitrogen tetroxide gas. Students observed that at low temperatures, the container appeared colorless, but when gently heated, it turned brown. On cooling again, the brown color faded. The teacher emphasized that no gases were escaping from the container and that the observed color changes were entirely due to changes occurring within the closed system. Learners were asked to relate the observations to equilibrium shifts and energy changes associated with the reaction.

a) Write the equilibrium equation for the system described.
b) State the color of nitrogen dioxide and dinitrogen tetroxide.
c) Explain the effect of increasing temperature on the equilibrium position.
d) State whether the forward reaction is endothermic or exothermic.
e) Explain your answer using Le Chatelier's principle.

Item 3
During revision, students studied the manufacture of sulfur trioxide in the Contact Process. They learned that sulfur dioxide reacts with oxygen in the presence of a catalyst inside a closed chamber. Despite using excess oxygen, the reaction does not proceed to completion. Engineers carefully adjust temperature, pressure, and catalyst conditions to obtain the best yield. Students were asked to relate these conditions to equilibrium concepts rather than reaction rates alone.

a) Write the balanced equation for the formation of sulfur trioxide.
b) State the catalyst used in the process.
c) Explain why high pressure favors the forward reaction.
d) Explain why a very high temperature is not used.
e) Distinguish between factors affecting equilibrium position and reaction rate.

Item 4
A chemistry teacher prepared an equilibrium mixture of iron(III) ions and thiocyanate ions, producing a deep red complex ion. Students noticed that adding more thiocyanate intensified the red color, while adding iron(III) chloride had a similar effect. However, when water was added, the red color faded significantly. The teacher challenged the students to explain these observations using equilibrium principles.

a) Write the equilibrium equation involved.
b) Explain the effect of adding thiocyanate ions.
c) Explain why dilution causes the red color to fade.

d) State the effect of adding iron(III) ions.
e) Name the principle used to explain these changes.

Item 5
In an advanced lesson, learners were introduced to the equilibrium constant, Kc. They were told that for a given reaction at constant temperature, the value of Kc remains unchanged even if concentrations are altered. A reaction mixture was allowed to reach equilibrium, and concentration data were provided for analysis.

a) Define the equilibrium constant, Kc.
b) Write a general expression for Kc for a reversible reaction.
c) State two conditions under which Kc remains constant.
d) Explain what a large value of Kc indicates about equilibrium position.
e) State one limitation of using Kc alone to predict reaction behavior.

Item 6 (CALCULATION)
In a closed container at constant temperature, the following equilibrium was established:

$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$

At equilibrium, the concentrations were:
$[H_2]$ = 0.20 mol dm^{-3},
$[I_2]$ = 0.20 mol dm^{-3},
$[HI]$ = 0.80 mol dm^{-3}.

Students were asked to analyze the system quantitatively.

a) Write the expression for Kc for the reaction.
b) Calculate the value of Kc.
c) State whether products or reactants are favored.
d) Explain how temperature affects the value of Kc.
e) State one assumption made in equilibrium calculations.

Item 7
A science club investigated the effect of pressure on gaseous equilibria. They used a syringe containing a reversible gaseous reaction and observed changes in volume. The club leader emphasized that pressure changes only affect systems involving gases and only when the number of gas molecules changes.

a) Explain how increasing pressure affects gaseous equilibrium.
b) State the condition under which pressure has no effect on equilibrium.
c) Explain the effect of decreasing volume on equilibrium position.
d) Relate pressure effects to molecular collisions.
e) State one industrial example where pressure control is important.

Item 8
In an agricultural chemistry lesson, students studied the solubility of sparingly soluble salts in water. They learned that although such salts appear insoluble, a small amount dissolves to establish an equilibrium between solid and ions in solution. The teacher emphasized that this equilibrium is also dynamic.

a) Write an equilibrium equation for a sparingly soluble salt.
b) Define solubility equilibrium.
c) Define solubility product, Ksp.
d) Explain the effect of adding a common ion.
e) State one application of solubility equilibria.

Item 9
A laboratory technician added dilute hydrochloric acid to a solution of calcium hydroxide until the pH changed very slowly. The teacher explained that the solution resisted sudden pH changes due to the presence of a buffer system. This sparked discussion about acid–base equilibria.

a) Define a buffer solution.
b) Explain how a buffer resists pH change.
c) State one example of an acidic buffer.
d) State one example of a basic buffer.
e) State one biological importance of buffers.

Item 10
Students compared strong and weak acids using conductivity experiments. Hydrochloric acid conducted electricity strongly, while ethanoic acid conducted weakly. The teacher explained that this difference was related to the extent of ionization and equilibrium position.

a) Explain what is meant by a weak acid.
b) Write an equilibrium equation for ethanoic acid in water.
c) Explain why weak acids establish equilibrium in solution.
d) State one factor affecting degree of ionization.
e) Explain why weak acids have lower conductivity.

Item 11
In a closed laboratory flask, ammonia gas was dissolved in water to form ammonium hydroxide. Students noticed that the solution did not fully ionize and that equilibrium was established between molecules and ions.

a) Write the equilibrium equation for ammonia in water.
b) Explain why ammonia is a weak base.
c) Define ionic equilibrium.
d) State one factor affecting base strength.

e) Explain the effect of dilution on the equilibrium.

Item 12
A chemistry class investigated the effect of catalysts on equilibrium systems. A catalyst was added to a reversible reaction, and students observed that equilibrium was reached faster, but the final amounts of reactants and products were unchanged.

a) Explain the effect of a catalyst on equilibrium position.
b) Explain why catalysts do not change Kc.
c) State how catalysts affect activation energy.
d) Distinguish between equilibrium position and equilibrium rate.
e) State one industrial advantage of catalysts.

Item 13
Learners were asked to compare reversible and irreversible reactions using real-life examples such as rusting and esterification. The teacher emphasized that equilibrium concepts apply only to certain systems.

a) Define a reversible reaction.
b) State one condition required for reversibility.
c) Give one example of a reversible reaction.
d) Explain why rusting is irreversible.
e) State one importance of reversible reactions in industry.

Item 14 (CALCULATION)
For the equilibrium:

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

At equilibrium, the concentrations were:
$[SO_2] = 0.40$ mol dm^{-3},
$[O_2] = 0.20$ mol dm^{-3},
$[SO_3] = 0.80$ mol dm^{-3}.

a) Write the Kc expression.
b) Calculate the value of Kc.
c) State the equilibrium position.
d) Explain how pressure affects this equilibrium.
e) State one industrial relevance of this reaction.

Item 15
At the end of the topic, students reflected on how equilibrium concepts apply in chemistry, biology, and industry. The teacher emphasized that equilibrium does not mean reactions stop but that balance is achieved at the molecular level.

a) Explain why equilibrium is described as dynamic.
b) State two characteristics of an equilibrium system.
c) Explain one real-life application of equilibrium.
d) State one limitation of equilibrium concepts.
e) Explain why closed systems are essential for equilibrium.

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ANSWERS – EQUILIBRIA I

Item 1 – Answers
a) Chemical equilibrium is the state in a reversible reaction where the rate of the forward reaction equals the rate of the backward reaction.
b) $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$
c) Dynamic equilibrium means reactions continue but concentrations remain constant.
d) Closed system and constant temperature.
e) Removing ammonia shifts equilibrium forward to produce more NH_3.

Item 2 – Answers
a) $2NO_2(g) \rightleftharpoons N_2O_4(g)$
b) NO_2 is brown; N_2O_4 is colorless.
c) Increasing temperature shifts equilibrium to the left.
d) Forward reaction is exothermic.
e) Equilibrium shifts to absorb added heat.

Item 3 – Answers
a) $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$
b) Vanadium(V) oxide, V_2O_5.
c) Fewer moles of gas on the product side.
d) High temperature favors reverse reaction.
e) Temperature affects equilibrium; catalyst affects rate only.

Item 4 – Answers
a) $Fe^{3+}(aq) + SCN^-(aq) \rightleftharpoons FeSCN^{2+}(aq)$
b) Equilibrium shifts to the right forming more complex.
c) Dilution favors side with more particles.
d) Increases red color intensity.
e) Le Chatelier's principle.

Item 5 – Answers
a) Kc is the ratio of product concentrations to reactant concentrations at equilibrium.

b) Kc = $[C]^c[D]^d / [A]^a[B]^b$
c) Constant temperature and same reaction.
d) Products are favored.
e) Does not indicate rate of reaction.

Item 6 – Answers (CALCULATION)
a) Kc = $[HI]^2 / ([H_2][I_2])$
b) Kc = $(0.80)^2 / (0.20 \times 0.20)$
$= 0.64 / 0.04$
$= 16$
c) Products are favored.
d) Temperature changes Kc.
e) System is at equilibrium.

Item 7 – Answers
a) Equilibrium shifts to side with fewer gas molecules.
b) When number of gas molecules is equal.
c) Same effect as increasing pressure.
d) More frequent collisions favor smaller volume.
e) Haber process.

Item 8 – Answers
a) $AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$
b) Equilibrium between solid and dissolved ions.
c) Product of ion concentrations at equilibrium.
d) Decreases solubility.
e) Water treatment.

Item 9 – Answers
a) A solution that resists pH change.
b) Neutralizes added acids or bases.
c) Ethanoic acid and sodium ethanoate.
d) Ammonia and ammonium chloride.
e) Maintains blood pH.

Item 10 – Answers
a) A weak acid partially ionizes in solution.
b) $CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$
c) Ionization is incomplete.
d) Concentration.
e) Fewer ions are present.

Item 11 – Answers
a) $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

b) It partially ionizes.
c) Equilibrium between ions and molecules.
d) Base strength.
e) Shifts equilibrium to the right.

Item 12 – Answers
a) No effect on equilibrium position.
b) Speeds both reactions equally.
c) Lowers activation energy.
d) Rate vs amount.
e) Saves energy and time.

Item 13 – Answers
a) Reaction that proceeds in both directions.
b) Closed system.
c) Esterification.
d) Rusting goes to completion.
e) Improves yield.

Item 14 – Answers (CALCULATION)

a) $K_c = [SO_3]^2 / ([SO_2]^2[O_2])$
b) $K_c = (0.80)^2 / (0.40^2 \times 0.20)$
$= 0.64 / (0.16 \times 0.20)$
$= 0.64 / 0.032$
$= 20$
c) Products favored.
d) High pressure favors forward reaction.
e) Contact Process.

Item 15 – Answers
a) Reactions continue at equal rates.
b) Constant concentrations and closed system.
c) Ammonia synthesis.
d) Applies only to reversible reactions.
e) Prevents loss of reactants/products.

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Senior Six – Equilibria II

Item 1:
During a dry season, a chemistry teacher in Soroti set up an experiment to demonstrate evaporation and condensation using a sealed glass container partially filled with pure ethanol. The container was placed on a laboratory bench and left undisturbed for several hours. At first, students observed that the level of liquid ethanol slowly decreased, while a faint mist appeared near the upper part of the container. As time went on, the mist disappeared, yet the liquid level no longer changed.

Later, the teacher gently warmed the container using a water bath and asked the students to carefully observe what happened inside the container. Some students noticed that droplets began forming on the cooler walls, while others argued that ethanol molecules were escaping permanently into the air space. When the container was cooled again, the liquid level increased slightly.

The teacher reminded the class that the container was sealed and that no ethanol could enter or leave the system. He then challenged the students to explain the observations using molecular ideas rather than simple definitions.

Task
a) Describe, in terms of molecular motion, what is meant by liquid–vapour equilibrium in this system.
b) Explain why the liquid level stopped changing even though evaporation was still taking place.
c) Using a step-by-step mechanism, describe what happens to ethanol molecules at the surface during evaporation and condensation.
d) Explain the effect of increasing temperature on the position of equilibrium.
e) State and explain two conditions necessary for physical equilibrium to be established.

Item 2:
In a school laboratory, excess potassium nitrate was added to a beaker of water at 60 °C and stirred until no more solid appeared to dissolve. The hot solution was then allowed to cool slowly to room temperature without disturbance. After some time, crystals of potassium nitrate were observed settling at the bottom of the beaker, while the solution above remained clear.

One student claimed that once crystals formed, all movement of ions stopped. Another insisted that even at room temperature, the system was still active at the microscopic level. To settle the disagreement, the teacher asked the class to analyze the situation using equilibrium concepts.

Task
a) Explain why crystals form when the solution cools.
b) Describe the equilibrium that exists between the solid potassium nitrate and its ions in solution.
c) Using an ionic mechanism, explain what happens simultaneously at the surface of the crystals.

d) Explain why stirring does not change the equilibrium position.
e) State how the equilibrium would be affected if more solid potassium nitrate were added.

Item 3:

Colligative Properties – Vapour Pressure Lowering

A soft-drink manufacturer in Kampala investigated why sugar solutions boil at higher temperatures than pure water. In one experiment, a known mass of sucrose was dissolved in water to prepare a concentrated syrup. The vapour pressure of the syrup was found to be significantly lower than that of pure water at the same temperature.

The quality-control officer explained to new interns that this effect depended on the number of particles present rather than their chemical nature. However, many interns struggled to visualize how dissolved particles could affect vapour pressure.

Task

a) Define vapour pressure in terms of molecular escape from a liquid.
b) Using a detailed molecular mechanism, explain why adding a non-volatile solute lowers vapour pressure.
c) Explain why sucrose does not appear in the vapour phase.
d) State two practical consequences of vapour pressure lowering in everyday life.
e) Explain why colligative properties depend on concentration and not chemical identity.

Item 4

Boiling Point Elevation – Stepwise Explanation

During a practical lesson, a student heated pure water and a concentrated sodium chloride solution under identical conditions. The student observed that the salt solution required more heating before it began to boil. This observation led to a heated debate about whether heat was being “absorbed by the salt” or whether a deeper explanation was required.

The teacher insisted that the explanation must be based on equilibrium between the liquid and vapour phases and not on vague ideas of heat absorption.

Task

a) Define boiling point in terms of vapour pressure and external pressure.
b) Explain, step by step, how the presence of dissolved sodium chloride raises the boiling point of water.
c) Using equilibrium ideas, explain why more energy is required to boil the solution.
d) State one assumption made when applying boiling point elevation theory.
e) Suggest one industrial application of boiling point elevation.

Item 5

Freezing Point Depression – Particle Mechanism

In a cold-storage facility, salt is spread on icy floors to prevent slipping. A chemistry student visiting the facility noticed that ice melted even though the surrounding temperature was below 0 °C. The facility manager explained that the salt interfered with the freezing process of water.

The student was asked to explain this phenomenon using chemical equilibrium concepts rather than simple memorization.

Task

a) Define freezing point in terms of phase equilibrium.
b) Explain, using a particle-level mechanism, how salt lowers the freezing point of water.
c) Describe what happens at the ice–solution interface when salt is added.
d) Explain why the effect increases with higher salt concentration.
e) State one limitation of using freezing point depression in real systems.

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Senior Six – Equilibria II (continued)

Item 6: **Osmotic Pressure in Living Cells**

During a biology–chemistry joint practical session, students placed strips of potato tissue into three different solutions: distilled water, dilute sodium chloride solution, and concentrated sodium chloride solution. After one hour, the strip placed in distilled water became firm, while the one in concentrated salt solution became limp. The strip in dilute salt solution showed little change.

The teacher explained that this was not a biological mystery but a chemical equilibrium involving solvent movement through a semi-permeable membrane. Some students struggled to connect the observation to molecular motion and equilibrium concepts.

Questions:

a) Define osmosis in terms of particle movement and equilibrium.
b) Using a step-by-step molecular mechanism, explain why the potato strip gained mass in distilled water.
c) Explain why the strip lost mass in concentrated salt solution.
d) Define osmotic pressure and relate it to this experiment.
e) Explain why osmotic pressure is considered a colligative property

Item 7: **Vapour Pressure and Raoult's Law**

A laboratory technician prepared a solution by mixing ethanol with water in a closed container fitted with a pressure sensor. As the mole fraction of ethanol increased, the total vapour pressure above the solution changed in a predictable manner. The technician noted that neither liquid behaved independently once mixed.

Students observing the experiment were asked to interpret the results using Raoult's law rather than intuition.

Questions:
a) State Raoult's law.
b) Explain, using molecular interactions, why vapour pressure changes with composition.
c) Describe how intermolecular forces affect deviations from Raoult's law.
d) Explain why ideal solutions obey Raoult's law closely.
e) State one limitation of Raoult's law.

Item 8

Dynamic Nature of Physical Equilibrium

A sealed flask containing water was placed inside a transparent box with a temperature sensor. Over time, the level of liquid water remained constant even though evaporation was expected to occur. A student claimed that evaporation had stopped, while another argued that evaporation and condensation were still occurring.

The teacher challenged the class to justify the correct view using kinetic theory.

Questions: a) Define dynamic equilibrium.
b) Explain why evaporation does not stop at equilibrium.
c) Using a molecular mechanism, explain condensation in this system.
d) Explain the effect of increasing surface area on equilibrium position.
e) State one industrial process where dynamic equilibrium is important.

Item 9:

Colligative Properties in Food Preservation

A food technologist explained why concentrated sugar solutions are used to preserve fruits in syrup. The syrup prevented microbial growth and altered the physical properties of water in the fruit tissues.

Students were asked to analyze the preservation method using equilibrium and colligative property concepts.

Questions:
a) Explain how sugar affects the chemical potential of water.
b) Describe the mechanism by which high solute concentration inhibits microbial growth.
c) Explain why sugar concentration matters more than sugar type.
d) Relate this process to osmotic pressure.

e) State one disadvantage of this preservation method.

Item 10:

Freezing Point Depression Calculations

In a winter experiment, a student dissolved 10.0 g of sodium chloride in 200 g of water to prepare an ice-melting mixture. The temperature at which the mixture froze was significantly lower than 0 °C.

The student was asked to quantify the effect using known constants and equilibrium concepts.

Questions:

a) Define freezing point depression.
b) State the formula used to calculate freezing point depression.
c) Explain why sodium chloride produces a larger effect than sugar.
d) State the assumptions made in this calculation.
e) Explain the physical meaning of the calculated temperature change.

Item 11

Boiling Point Elevation in Industrial Processes

In a sugar-processing factory, workers observed that concentrated sugar solutions required higher temperatures to boil compared to pure water. This affected energy costs and equipment design.

The factory engineer asked chemistry students to explain the phenomenon using equilibrium and thermodynamic principles.

Questions:

a) Explain boiling in terms of vapour pressure equilibrium.
b) Describe how solute particles affect vapour pressure.
c) Explain why more heat is required to boil concentrated solutions.
d) State how boiling point elevation affects industrial efficiency.
e) Give one limitation of boiling point elevation theory.

Item 12:

Solvent–Solute Interactions

A research student compared two solutions: one containing sodium chloride and another containing ethanol, both in water. The changes in vapour pressure and freezing point were different, even at similar concentrations.

The student was asked to explain the observations using particle behavior and intermolecular forces.

Questions:

a) Explain how electrolytes differ from non-electrolytes in solution.

b) Describe the role of ion–dipole interactions.
c) Explain why dissociation affects colligative properties.
d) State the van't Hoff factor and its significance.
e) Explain one reason for deviation from ideal behavior.

Item 13:
Physical Equilibrium in Closed Systems
A sealed container holding iodine crystals was gently heated until violet vapour filled the container. On cooling, some crystals reappeared on the walls of the container. The mass of iodine in the container remained constant throughout.

Students were asked to analyze the process using equilibrium concepts.

Questions:
a) Identify the type of equilibrium involved.
b) Explain sublimation using molecular motion.
c) Describe condensation of iodine vapour.
d) Explain the effect of temperature on the equilibrium.
e) State why the system must be closed.

Item 14
Colligative Properties and Molecular Mass
Two unknown solutes A and B were separately dissolved in water at the same mass concentration. Solution A produced a larger freezing point depression than solution B.

Students were asked to determine what could be inferred about the two solutes.

Questions:
a) Explain the relationship between freezing point depression and number of particles.
b) Deduce which solute has lower molar mass.
c) Explain the reasoning using equilibrium ideas.
d) State one assumption made in this comparison.
e) Suggest one experimental source of error.

Item 15:
Real-Life Application of Equilibria II
A municipal engineer explained how antifreeze solutions protect car engines in cold regions. The solution prevented freezing and controlled heat transfer during engine operation.

Students were asked to connect this explanation to Equilibria II concepts.

Questions:
a) Identify two colligative properties involved.
b) Explain how antifreeze affects freezing point and boiling point.

c) Describe the molecular mechanism involved.
d) Explain why concentration control is important.
e) State one environmental concern associated with antifreeze use.

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ANSWERS – EQUILIBRIA II

Item 1 – Answers
a) Liquid–vapour equilibrium is a dynamic state in which the rate of evaporation of ethanol molecules from the liquid surface equals the rate of condensation of ethanol molecules from the vapour back into the liquid.

b) The liquid level stopped changing because evaporation and condensation occurred at equal rates, so there was no net change in the amount of liquid.

c) During evaporation, high-energy ethanol molecules at the liquid surface overcome intermolecular forces and escape into the vapour phase. During condensation, vapour molecules collide with the liquid surface, lose kinetic energy, and are pulled back into the liquid by intermolecular forces.

d) Increasing temperature increases molecular kinetic energy, increasing evaporation and shifting equilibrium towards the vapour phase.

e) The system must be closed, and temperature must remain constant for equilibrium to be established.

Item 2 – Answers
a) Crystals form because solubility of potassium nitrate decreases as temperature decreases.

b) A dynamic equilibrium exists between solid KNO_3 and K^+ and NO_3^- ions in solution.

c) At the crystal surface, ions continuously leave the solid lattice and enter solution while other ions from solution reattach to the crystal at the same rate.

d) Stirring increases the rate at which equilibrium is reached but does not change the equilibrium position.

e) Adding more solid potassium nitrate does not affect the equilibrium concentration of ions.

Item 3 – Answers
a) Vapour pressure is the pressure exerted by vapour molecules when evaporation and condensation are at equilibrium.

b) Adding a non-volatile solute reduces the number of solvent molecules at the surface, decreasing the rate of evaporation and lowering vapour pressure.

c) Sucrose does not vaporize because it is non-volatile and has strong intermolecular forces.

d) Food preservation and boiling point elevation in industrial processes.

e) Colligative properties depend on the number of solute particles, not their identity.

Item 4 – Answers
a) Boiling point is the temperature at which vapour pressure equals external pressure.

b) Dissolved NaCl reduces vapour pressure, so higher temperature is needed to reach atmospheric pressure.

c) More energy is required to increase kinetic energy of solvent molecules enough to escape.

d) The solution behaves ideally at low concentration.

e) Used in antifreeze and industrial boiling systems.

Item 5 – Answers
a) Freezing point is the temperature at which solid and liquid phases are in equilibrium.

b) Salt ions disrupt the formation of the ice lattice, requiring lower temperature to freeze.

c) Salt dissolves in the thin liquid layer on ice, lowering freezing point and causing melting.

d) Higher concentration increases number of particles, increasing the effect.

e) At very high concentrations, deviations from ideal behavior occur.

Item 6 – Answers
a) Osmosis is the movement of solvent molecules through a semi-permeable membrane from higher to lower water potential until equilibrium is reached.

b) Water moves into the potato cells due to higher water potential outside.

c) Water leaves the potato cells into the concentrated salt solution.

d) Osmotic pressure is the pressure required to stop osmosis.

e) It depends on number of solute particles.

Item 7 – Answers
a) Raoult's law states that partial vapour pressure is proportional to mole fraction.

b) Increased ethanol mole fraction increases its contribution to vapour pressure.

c) Strong intermolecular forces cause deviations.

d) Ideal solutions have similar intermolecular forces.

e) Not valid for concentrated or reactive solutions.

Item 8 – Answers
a) Dynamic equilibrium is when forward and reverse processes occur at equal rates.

b) Evaporation continues but is balanced by condensation.

c) Vapour molecules lose energy on collision and return to liquid.

d) Increased surface area increases rates but not equilibrium position.

e) Distillation processes.

Item 9 – Answers
a) Sugar lowers the chemical potential of water.

b) Microbes lose water by osmosis.

c) Number of particles matters, not type.

d) High osmotic pressure dehydrates cells.

e) High sugar content is unhealthy.

Item 10 – Answers
a) Freezing point depression is the lowering of freezing temperature due to solute addition.

b) $\Delta Tf = iKf\ m$

c) NaCl dissociates into ions, increasing particles.

d) Solution behaves ideally.

e) Temperature change reflects reduced solid–liquid equilibrium.

Item 11 – Answers
a) Boiling occurs when vapour pressure equals external pressure.

b) Solute particles reduce surface solvent molecules.

c) More heat needed to raise vapour pressure.

d) Increases fuel consumption.

e) Limited to dilute solutions.

Item 12 – Answers
a) Electrolytes dissociate; non-electrolytes do not.

b) Ion–dipole attractions stabilize ions.

c) More particles increase colligative effects.

d) van't Hoff factor represents particle number.

e) Ion pairing causes deviation.

Item 13 – Answers
a) Solid–vapour equilibrium.

b) Sublimation involves molecules escaping directly to vapour.

c) Vapour loses energy and reforms solid.

d) Heating favors vapour formation.

e) Prevents loss of iodine.

Item 14 – Answers
a) Greater particle number causes larger depression.

b) Solute A has lower molar mass.

c) More moles per mass produce larger effect.

d) Solutes are non-electrolytes.

e) Impurities affect results.

Item 15 – Answers
a) Freezing point depression and boiling point elevation.

b) Antifreeze lowers freezing point and raises boiling point.

c) Solute particles disrupt solvent equilibrium.

d) Correct concentration ensures protection.

e) Environmental toxicity.

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📘 ***UACE CHEMISTRY ITEM***

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Senior Six – Electrochemistry

Item 1:
Oxidation and Reduction in a Village Workshop
In a small metal workshop in Mbarara, iron tools are often left exposed to moist air. Over several weeks, a reddish-brown coating forms on the surface of the tools. The workshop owner notices that tools kept near a leaking water pipe rust faster than those stored in a dry cupboard. During a school visit, a chemistry teacher uses this situation to introduce electrochemical ideas related to oxidation and reduction.

The teacher explains that rusting is not a simple reaction but a combination of several electrochemical processes occurring at different regions of the metal surface.

Questions: a) Define oxidation and reduction in terms of electron transfer.
b) Identify the anodic and cathodic regions during rusting of iron.
c) Using half-equations, explain the electrochemical mechanism of rust formation.
d) Explain the role of water and oxygen in the rusting process.
e) State two methods of preventing rusting and explain one using electrochemical principles.

Item 2:

Oxidation Numbers in Industrial Processes

In a fertilizer factory, ammonia is converted into nitric acid through a series of reactions. Students observing the process are told that oxidation numbers change at each stage of the reaction. Some students confuse oxidation number change with actual charge.

The factory chemist challenges them to analyze the reactions step by step.

Questions: a) Define oxidation number.
b) Determine the oxidation number of nitrogen in NH_3, NO, and NO_2.
c) Identify which stages involve oxidation and which involve reduction.
d) Explain why oxidation number is a bookkeeping tool and not a real charge.
e) State one importance of oxidation numbers in balancing redox equations.

Item 3:

Daniell Cell – Construction and Mechanism

A student constructs a Daniell cell using a zinc electrode in zinc sulfate solution and a copper electrode in copper(II) sulfate solution, connected by a salt bridge. When a voltmeter is connected, a steady potential difference is observed.

The student is curious about how electrons move through the circuit and how ions move inside the solutions to maintain electrical neutrality.

Questions: a) Draw a labeled diagram of the Daniell cell.
b) Identify the anode and cathode.
c) Write the half-equations occurring at each electrode.
d) Explain the role of the salt bridge using ion movement.
e) State why the cell eventually stops working.

Item 4:

Cell Potential and Feasibility

During a practical lesson, students are given standard electrode potential values and asked to predict whether certain redox reactions are feasible. Some reactions occur spontaneously, while others do not.

The teacher emphasizes that feasibility depends on the relative tendencies of substances to gain or lose electrons.

Questions: a) Define standard electrode potential.
b) Explain how a standard hydrogen electrode is used as a reference.
c) Calculate the standard cell potential for a given electrochemical cell.
d) State the condition for a reaction to be feasible.
e) Explain the relationship between cell potential and spontaneity.

Item 5:

Electrochemical Cells in Everyday Life

A mobile phone technician explains that batteries convert chemical energy into electrical energy. Different types of batteries are used depending on required voltage, capacity, and rechargeability.

Students are asked to relate these observations to electrochemical cell theory.

Questions: a) Distinguish between primary and secondary cells.
b) Explain the electrochemical reactions occurring in a dry cell.
c) State why rechargeable cells can be reused.
d) Explain one advantage and one disadvantage of secondary cells.
e) Give one environmental concern related to battery disposal.

Item 6:

Electrolysis of Molten Ionic Compounds

In an industrial plant, molten sodium chloride is electrolyzed to produce sodium metal and chlorine gas. The process requires high temperatures and a large supply of electrical energy.

A student is asked to explain why molten sodium chloride is used instead of an aqueous solution.

Questions: a) Define electrolysis.
b) Explain why molten NaCl conducts electricity.
c) Write half-equations for the reactions at the electrodes.
d) Explain why sodium cannot be produced from aqueous NaCl by electrolysis.
e) State one industrial use of sodium metal.

Item 7:

Electrolysis of Aqueous Solutions

During a laboratory experiment, dilute sulfuric acid is electrolyzed using inert electrodes. Bubbles of gas are observed at both electrodes.

Students are asked to identify the gases and explain the reactions involved.

Questions: a) Identify the gas at the cathode and anode.
b) Write half-equations for both electrodes.
c) Explain why sulfuric acid is used instead of pure water.
d) State the overall reaction.
e) Explain one safety precaution during the experiment.

Item 8:

Faraday's First Law of Electrolysis

A technician passes a steady current through a copper(II) sulfate solution using copper electrodes. After a fixed time, the mass of copper deposited on the cathode is measured.

Students are required to relate mass deposited to charge passed.

Questions: a) State Faraday's first law of electrolysis.
b) Write the formula relating mass, charge, and time.
c) Explain the meaning of each symbol used.
d) State one assumption made in applying Faraday's law.
e) Explain one source of experimental error.

Item 9:
Faraday's Second Law

Two different electrolytes are electrolyzed using the same quantity of electricity. Different masses of substances are deposited at the electrodes.

Students are asked to interpret the results using Faraday's second law.

Questions: a) State Faraday's second law of electrolysis.
b) Explain the term electrochemical equivalent.
c) Relate deposited mass to molar mass and charge.
d) Explain why different substances produce different masses.
e) State one application of Faraday's laws.

Item 10:
Calculation Involving Electrolysis
A current of 2.0 A is passed through molten aluminium oxide for 30 minutes during aluminium extraction. Students are asked to calculate the mass of aluminium produced.

Questions: a) Write the cathode half-equation.
b) Calculate the total charge passed.
c) Determine the number of moles of electrons involved.
d) Calculate the mass of aluminium produced.
e) State one factor affecting energy consumption in the process.

Item 11:
Electroplating in Industry
A company electroplates iron objects with nickel to improve corrosion resistance and appearance. The iron object is connected to the negative terminal of a power supply and immersed in nickel sulfate solution.

Students are asked to explain the electrochemical principles involved.

Questions: a) Define electroplating.
b) Identify the cathode and anode.
c) Write the electrode reactions.
d) Explain why the plated object must be the cathode.
e) State one advantage of electroplating.

Item 12:

Choice of Electrodes

In electrolysis experiments, inert electrodes such as graphite or platinum are often used instead of reactive metals.

Students are asked to explain the importance of electrode choice.

Questions: a) Define inert electrodes.
b) Explain why graphite is commonly used.
c) State what could happen if reactive electrodes are used.
d) Explain the effect of electrode material on products formed.
e) Give one limitation of graphite electrodes.

Item 13:

Electrochemical Corrosion Protection

An underground pipeline is protected by attaching blocks of magnesium metal to it. Over time, the magnesium blocks wear away while the pipeline remains intact.

Students are required to explain the observation using electrochemical principles.

Questions: a) Define sacrificial protection.
b) Explain why magnesium is used.
c) Write the oxidation half-equation for magnesium.
d) Explain how electrons flow in this system.
e) State one advantage of sacrificial protection.

Item 14:

Fuel Cells

A fuel cell uses hydrogen and oxygen to produce electricity and water. Unlike conventional cells, the reactants are continuously supplied.

Students are asked to analyze the operation of the cell.

Questions: a) Define a fuel cell.
b) Write the half-equations for hydrogen–oxygen fuel cell.
c) Explain how electrical energy is produced.
d) State two advantages of fuel cells.
e) Give one limitation of fuel cells.

Item 15:
Electrochemistry and the Environment
A community is encouraged to recycle metals to reduce environmental damage caused by mining. Electrolysis plays a key role in metal extraction and purification.

Students are asked to link electrochemistry to environmental conservation.

Questions: a) Explain how electrolysis is used in metal purification.
b) State one environmental problem caused by mining.
c) Explain how recycling reduces energy consumption.
d) Relate electrochemistry to sustainable development.
e) State one challenge of large-scale recycling.

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ANSWERS – ELECTROCHEMISTRY

Item 1 – Answers (Rusting of Iron)
a) Oxidation is the loss of electrons, while reduction is the gain of electrons.

b) The anodic regions are areas where iron loses electrons, while cathodic regions are areas where oxygen gains electrons.

c) Anode:
$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

Cathode:
$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$

The Fe^{2+} ions react with OH^- ions to form iron(II) hydroxide, which is further oxidized to hydrated iron(III) oxide (rust).

d) Water acts as an electrolyte allowing ion movement, while oxygen acts as the oxidizing agent.

e) Painting, oiling, galvanizing, or sacrificial protection.
Galvanizing works because zinc oxidizes preferentially, protecting iron.

Item 2 – Answers (Oxidation Numbers)
a) Oxidation number is the apparent charge an atom would have if electrons were completely transferred.

b)
NH_3: N = −3
NO: N = +2
NO_2: N = +4

c) Nitrogen is oxidized from −3 to +2 and further to +4.

d) Oxidation number is imaginary and used only for tracking electron transfer.

e) Helps identify oxidizing and reducing agents and balance redox equations.

Item 3 – Answers (Daniell Cell)
a) Zinc electrode in $ZnSO_4$, copper electrode in $CuSO_4$, salt bridge connecting solutions, external circuit with voltmeter.

b) Anode: zinc
Cathode: copper

c)
Anode: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$
Cathode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

d) Salt bridge allows ion flow to maintain electrical neutrality.

e) The cell stops when reactants are used up or concentrations equalize.

Item 4 – Answers (Cell Potential)
a) Standard electrode potential is the e.m.f. of a half-cell measured under standard conditions relative to the hydrogen electrode.

b) The standard hydrogen electrode is assigned a potential of 0.00 V.

c)
$E°cell = E°cathode − E°anode$

d) A reaction is feasible if E°cell is positive.

e) Higher cell potential means greater spontaneity.

Item 5 – Answers (Batteries)
a) Primary cells are non-rechargeable; secondary cells are rechargeable.

b) Chemical reactions produce electrons that flow through an external circuit.

c) Reversible reactions allow recharging.

d) Advantage: reusable; disadvantage: expensive.

e) Heavy metal pollution.

Item 6 – Answers (Molten NaCl Electrolysis)
a) Electrolysis is the use of electricity to drive a non-spontaneous reaction.

b) Molten NaCl contains mobile Na^+ and Cl^- ions.

c)
Cathode: $Na^+ + e^- \rightarrow Na$
Anode: $2Cl^- \rightarrow Cl_2 + 2e^-$

d) In aqueous solution, water is reduced instead of Na^+.

e) Sodium is used in making organic compounds.

Item 7 – Answers (Aqueous Electrolysis)
a) Cathode: hydrogen gas
Anode: oxygen gas

b)
Cathode: $2H^+ + 2e^- \rightarrow H_2$
Anode: $4OH^- \rightarrow O_2 + 2H_2O + 4e^-$

c) Sulfuric acid increases conductivity.

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

e) Avoid sparks near hydrogen.

Item 8 – Answers (Faraday's First Law)
a) Mass deposited is proportional to quantity of electricity passed.

b) $m = ZIt$

c)
m = mass
Z = electrochemical equivalent
I = current
t = time

d) Current efficiency is assumed to be 100%.

e) Side reactions reduce mass deposited.

Item 9 – Answers (Faraday's Second Law)
a) Mass deposited is proportional to molar mass divided by charge.

b) Electrochemical equivalent is mass deposited per coulomb.

c) $m \propto M / z$

d) Different ions require different numbers of electrons.

e) Electroplating.

Item 10 – Answers (Calculation)
a) $Al^{3+} + 3e^- \rightarrow Al$

b) $Q = It = 2.0 \times 1800 = 3600$ C

c) Moles of electrons = 3600 / 96500 = 0.0373 mol

d) Moles of Al = 0.0373 / 3 = 0.0124
Mass = 0.0124 × 27 = 0.335 g

e) High temperature increases energy consumption.

Item 11 – Answers (Electroplating)
a) Electroplating is coating an object with metal using electricity.

b) Cathode: iron object
Anode: nickel

c)
Cathode: $Ni^{2+} + 2e^- \rightarrow Ni$
Anode: $Ni \rightarrow Ni^{2+} + 2e^-$

d) Reduction occurs at the cathode.

e) Prevents corrosion.

Item 12 – Answers (Electrodes)
a) Inert electrodes do not react.

b) Graphite is cheap and conducts electricity.

c) Reactive electrodes may dissolve.

d) Electrode material affects products formed.

e) Graphite may wear away.

Item 13 – Answers (Sacrificial Protection)
a) Sacrificial protection uses a more reactive metal to protect another.

b) Magnesium oxidizes easily.

c) $Mg \rightarrow Mg^{2+} + 2e^-$

d) Electrons flow from magnesium to iron.

e) Effective corrosion prevention.

Item 14 – Answers (Fuel Cells)
a) A fuel cell converts chemical energy directly into electrical energy.

b)
Anode: $2H_2 \rightarrow 4H^+ + 4e^-$
Cathode: $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$

c) Electron flow produces electricity.

d) Clean energy, high efficiency.

e) High cost.

Item 15 – Answers (Environment)
a) Electrolysis purifies metals by removing impurities.

b) Land degradation.

c) Recycling uses less electricity.

d) Supports sustainable resource use.

e) Collection and sorting difficulties.

UACE CHEMISTRY ITEM BANK

Senior Six – Periodicity II
(Group 14, Group 17, d-block elements)

Item 1:
Group 14 and Carbon Bonding in Real Life
In a materials science exhibition in Kampala, students observed different forms of carbon on display, including diamond jewellery, graphite rods, and carbon nanotube samples. A guide explained that although all these substances are made of carbon atoms, their physical properties differ greatly. Some students were surprised that one form is extremely hard while another is soft and slippery.

The guide emphasized that these differences arise from atomic structure, bonding, and periodic trends within Group 14.

Questions: a) State the position of carbon in the periodic table.
b) Explain why carbon forms giant covalent structures.
c) Describe the bonding in diamond and explain its hardness.
d) Explain why graphite conducts electricity but diamond does not.
e) Relate these properties to the electronic structure of carbon.

Item 2:
Silicon and the Semiconductor Industry
During a visit to a solar panel manufacturing plant, students learned that silicon is a key material used in photovoltaic cells. The engineer explained that silicon's usefulness depends on its position in Group 14 and its ability to conduct electricity under certain conditions.

Students were asked to link silicon's properties to periodic trends.

Questions: a) State two physical properties of silicon.
b) Explain why silicon is classified as a metalloid.
c) Describe how silicon's conductivity can be increased.
d) Compare the bonding in silicon with that in carbon.
e) Explain one use of silicon based on its electronic structure.

Item 3:
Trend in Group 14 Hydrides

A chemistry teacher prepared samples of methane, silane, germane, and stannane for discussion. Students noticed that the stability and physical properties of these compounds changed down the group.

The teacher asked students to analyze these changes using periodic trends.

Questions: a) Write the general formula of Group 14 hydrides.
b) Describe the trend in thermal stability down the group.
c) Explain the trend using bond energy considerations.
d) Describe the trend in boiling points.
e) Explain why methane is particularly stable.

Item 4:

Oxidation States in Group 14

In an industrial chemistry lecture, students learned that lead forms compounds in different oxidation states. Older paints containing lead compounds were banned due to toxicity.

Students were asked to relate oxidation states to periodic trends.

Questions: a) State the common oxidation states of Group 14 elements.
b) Explain the inert pair effect.
c) Describe how oxidation states change down the group.
d) Explain why Pb(II) compounds are more stable than Pb(IV).
e) State one industrial or environmental implication of this trend.

Item 5:

Group 17 – The Halogens in Daily Life

At a water treatment plant, chlorine gas is used to disinfect water. Nearby, iodine solutions are used in medical clinics, while fluorides are added to toothpaste.

A chemist explained that all these substances belong to the same group but behave differently.

Questions: a) Identify the group number of halogens.
b) Describe the trend in atomic radius down the group.
c) Explain the trend in electronegativity.
d) Relate reactivity trends to electronic structure.
e) Explain why fluorine is the most reactive halogen.

Item 6:

Physical Properties of Halogens

A sealed container displayed chlorine gas, bromine liquid, and iodine solid. Students observed the gradual change in physical state down the group.

The teacher emphasized intermolecular forces as the key explanation.

Questions: a) Describe the physical state of halogens at room temperature.
b) Explain the trend in melting and boiling points.
c) Describe the role of van der Waals forces.
d) Explain why iodine is solid at room temperature.
e) Predict the state of astatine.

Item 7:
Displacement Reactions of Halogens
In a laboratory experiment, chlorine gas was bubbled through potassium bromide solution, causing a color change. When bromine was added to potassium chloride, no reaction occurred.

Students were asked to explain these observations using periodic trends.

Questions: a) Define a displacement reaction.
b) Write the equation for chlorine reacting with potassium bromide.
c) Explain why bromine cannot displace chlorine.
d) Relate the reactions to oxidizing power.
e) State one practical application of halogen displacement.

Item 8:
Halogen Compounds and Bond Polarity
A student compared hydrogen fluoride, hydrogen chloride, hydrogen bromide, and hydrogen iodide. Despite similar formulas, their properties differed significantly.

The teacher emphasized bond polarity and bond strength.

Questions: a) Describe the trend in bond polarity down Group 17.
b) Explain the trend in bond strength.
c) Describe the trend in acidity of hydrogen halides.
d) Explain why HF is weak despite high polarity.
e) Relate these trends to atomic size.

Item 9:
d-Block Elements and Transition Metals
In a metal workshop, iron, copper, and chromium were observed to have very different appearances and uses. The instructor explained that these are transition metals with unique properties.

Students were asked to analyze these properties using periodic ideas.

Questions: a) Define a transition element.
b) State three general properties of transition metals.
c) Explain why transition metals form colored compounds.
d) Describe the role of partially filled d-orbitals.
e) Explain one industrial use of a transition metal.

Item 10:
Variable Oxidation States in Transition Metals

A chemistry student noticed that iron forms both Fe^{2+} and Fe^{3+} ions, each with different properties. Copper also forms Cu^{+} and Cu^{2+} ions.

Students were asked to explain why this occurs.

Questions: a) Explain variable oxidation states.
b) Relate this property to electronic configuration.
c) Explain why transition metals show this behavior.
d) Give one advantage of variable oxidation states in catalysis.
e) State one disadvantage of this property.

Item 11:
Catalytic Activity of Transition Metals

In a chemical industry, finely divided iron is used in the Haber process, while vanadium(V) oxide is used in the Contact process.

Students were asked to explain why transition metals make good catalysts.

Questions: a) Define catalysis.
b) Explain how transition metals act as catalysts.
c) Describe adsorption of reactants on catalyst surface.
d) Explain the role of variable oxidation states.
e) State one limitation of using catalysts.

Item 12:
Complex Ion Formation
A student added excess ammonia to a copper(II) sulfate solution and observed a deep blue solution forming.

The teacher explained that complex ions were responsible.

Questions: a) Define a complex ion.
b) Identify the ligand and central metal ion.
c) Explain how coordinate bonds are formed.

d) Explain why transition metals form complexes.
e) State one use of complex formation.

Item 13:
Magnetic Properties of Transition Metals
In a physics–chemistry demonstration, iron was strongly attracted to a magnet, while zinc was not.

Students were asked to explain the difference using electron arrangement.

Questions: a) Define paramagnetism.
b) Explain why iron is paramagnetic.
c) Describe the role of unpaired electrons.
d) Explain why zinc is diamagnetic.
e) Relate magnetism to d-orbital occupancy.

Item 14:
Trends Across the d-Block
A periodic table chart showed gradual changes in atomic radius across the first transition series. Students were surprised that the change was small compared to s-block elements.

The teacher explained this using shielding effects.

Questions: a) Describe the trend in atomic radius across the d-block.
b) Explain why the change is small.
c) Describe the role of d-electron shielding.
d) Explain the trend in ionization energy.
e) Compare this trend with s-block elements.

Item 15:
Environmental Impact of Periodicity II Elements
A community living near a mining area reported pollution caused by heavy metals such as lead and chromium. Environmental officers linked the issue to chemical properties of these elements.

Students were asked to relate environmental impact to periodic trends.

Questions: a) Explain why heavy metals are toxic.
b) Relate toxicity to electronic structure.
c) Explain why lead compounds persist in the environment.
d) State one method of reducing heavy metal pollution.
e) Explain the importance of understanding periodicity in environmental protection.

ANSWERS – PERIODICITY II

Item 1 – Answers (Carbon Allotropes)
a) Carbon is in Group 14 (IV) and Period 2 of the periodic table.

b) Carbon forms giant covalent structures because it has four valence electrons and can form four strong covalent bonds, leading to extensive networks.

c) In diamond, each carbon atom is covalently bonded to four other carbon atoms in a tetrahedral arrangement, forming a rigid three-dimensional lattice. This makes diamond extremely hard.

d) Graphite conducts electricity because each carbon atom forms three covalent bonds, leaving one delocalized electron that can move freely. Diamond has no free electrons.

e) Carbon's electronic configuration ($1s^2\ 2s^2\ 2p^2$) allows strong covalent bonding and catenation.

Item 2 – Answers (Silicon)
a) Silicon is a hard, grey, brittle solid with moderate electrical conductivity.

b) Silicon is a metalloid because it shows both metallic and non-metallic properties.

c) Conductivity is increased by doping with elements such as phosphorus or boron.

d) Silicon forms giant covalent structures similar to carbon but with weaker Si–Si bonds.

e) Used in solar cells because it can generate electric current when exposed to light.

Item 3 – Answers (Group 14 Hydrides)
a) General formula: EH_4

b) Thermal stability decreases down the group.

c) Bond energy decreases as atomic size increases, weakening E–H bonds.

d) Boiling points increase down the group due to stronger van der Waals forces.

e) Methane is stable due to strong C–H bonds and small atomic size.

Item 4 – Answers (Oxidation States)
a) Common oxidation states are +4 and +2.

b) Inert pair effect is the tendency of the s-electrons to remain unshared in bonding.

c) Stability of +2 state increases down the group.

d) Pb(II) is more stable because the $6s^2$ electrons are reluctant to participate in bonding.

e) Lead(II) compounds persist in environment and are toxic.

Item 5 – Answers (Halogens)
a) Halogens are in Group 17.

b) Atomic radius increases down the group.

c) Electronegativity decreases down the group.

d) Reactivity decreases down the group due to reduced attraction for electrons.

e) Fluorine is most reactive due to small size and high electronegativity.

Item 6 – Answers (Physical Properties of Halogens)
a) Fluorine and chlorine are gases, bromine is liquid, iodine is solid.

b) Melting and boiling points increase down the group.

c) van der Waals forces increase with molecular size.

d) Iodine has strong intermolecular attractions.

e) Astatine is predicted to be a solid.

Item 7 – Answers (Displacement Reactions)
a) A displacement reaction is where a more reactive element replaces a less reactive one.

b) $Cl_2 + 2KBr \rightarrow 2KCl + Br_2$

c) Bromine is less reactive than chlorine.

d) Stronger oxidizing agents displace weaker ones.

e) Used in extraction of bromine from seawater.

Item 8 – Answers (Hydrogen Halides)
a) Bond polarity decreases down the group.

b) Bond strength decreases down the group.

c) Acidity increases down the group.

d) HF is weak due to strong hydrogen bonding.

e) Larger atomic size weakens H–X bond.

Item 9 – Answers (Transition Metals)
a) A transition element forms at least one ion with partially filled d-orbitals.

b) They form colored compounds, act as catalysts, and have variable oxidation states.

c) Color arises from d–d electron transitions.

d) Partially filled d-orbitals allow multiple oxidation states.

e) Iron is used in construction.

Item 10 – Answers (Variable Oxidation States)
a) Ability to form ions with different charges.

b) Both 4s and 3d electrons are involved.

c) Small energy difference between orbitals allows electron loss.

d) Enables catalytic activity.

e) Can lead to corrosion.

Item 11 – Answers (Catalysis)
a) Catalysis is the increase in reaction rate without permanent change to catalyst.

b) Transition metals provide alternative reaction pathways.

c) Reactants adsorb onto catalyst surface.

d) Oxidation states change temporarily.

e) Catalysts can be poisoned.

Item 12 – Answers (Complex Ions)
a) A complex ion contains a central metal ion surrounded by ligands.

b) Ligand: NH_3, metal ion: Cu^{2+}

c) Coordinate bonds form when ligands donate lone pairs.

d) Transition metals have vacant orbitals.

e) Used in qualitative analysis.

Item 13 – Answers (Magnetism)
a) Paramagnetism is attraction to magnetic field due to unpaired electrons.

b) Iron has unpaired d-electrons.

c) Unpaired electrons create magnetic moments.

d) Zinc has fully paired electrons.

e) Depends on d-orbital electron arrangement.

Item 14 – Answers (d-Block Trends)
a) Atomic radius decreases slightly across the d-block.

b) Increased nuclear charge is balanced by d-electron shielding.

c) d-electrons shield poorly.

d) Ionization energy increases slightly.

e) s-block shows larger change.

Item 15 – Answers (Environmental Impact)
a) Heavy metals bind to enzymes and disrupt metabolism.

b) High charge density affects biological molecules.

c) Lead compounds are chemically stable.

d) Use chelation and controlled disposal.

e) Helps predict reactivity and toxicity.

📘 ***Author: joelPCM***
📘 ***UACE CHEMISTRY ITEM BANK***

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Senior Six – Organic Chemistry II

(Alcohols, Phenols, Carbonyl compounds, Carboxylic acids & derivatives)

Item 1:

Ethanol Production and Dehydration Mechanism

At a small ethanol-processing plant in eastern Uganda, sugarcane molasses is fermented and distilled to produce ethanol for industrial and laboratory use. During quality control, the chemist demonstrates that ethanol can be converted into ethene by heating it with concentrated sulfuric acid. As the reaction proceeds, a colorless gas is collected and tested using bromine water.

Students are told that sulfuric acid plays more than one role in the reaction and that the transformation does not occur in a single step but through a reaction mechanism.

Task

Identify the functional group present in ethanol.

b) Write the equation for the dehydration of ethanol.

c) Using a step-by-step mechanism, explain how ethanol is converted into ethene.

d) State the role of concentrated sulfuric acid in the reaction.

e) Explain how the product gas can be identified experimentally.

Item 2: Substitution Reactions of Alcohols

In a school laboratory, ethanol is reacted separately with concentrated hydrochloric acid in the presence of anhydrous zinc chloride. A sweet-smelling liquid is formed, which does not mix well with water. The teacher explains that the reaction involves replacement of one group by another through a specific mechanism.

Students are asked to analyze the reaction carefully rather than memorizing equations.

Task

a) Name the reagent mixture used in this reaction.

b) Write the equation for the reaction between ethanol and hydrogen chloride.

c) Explain the reaction using a nucleophilic substitution mechanism.

d) State the function of zinc chloride.

e) Explain why the product is less soluble in water than ethanol.

Item 3: Oxidation of Alcohols in Industry

At a chemical manufacturing company, ethanol is oxidized under controlled conditions to produce ethanal, which is further oxidized to ethanoic acid. The production manager explains that careful control of reagents determines how far oxidation proceeds.

Students are asked to explain the chemistry behind this control.

Questions: a) Name a suitable oxidizing agent for converting ethanol to ethanal.
b) Write equations for the oxidation of ethanol to ethanal and then to ethanoic acid.
c) Explain why further oxidation of ethanal occurs easily.
d) Describe how reaction conditions can be controlled to stop oxidation at ethanal.
e) State one test to distinguish between ethanal and ethanoic acid.

Item 4: Phenol and Its Unusual Reactivity

During a laboratory demonstration, phenol reacts rapidly with bromine water at room temperature, producing a white precipitate. Students observe that ethanol does not react in the same way under similar conditions.

The teacher explains that this difference is due to the structure of phenol and the interaction between the hydroxyl group and the benzene ring.

Questions: a) State the functional group present in phenol.
b) Write the equation for the reaction between phenol and bromine water.
c) Using an electrophilic substitution mechanism, explain why phenol reacts readily.
d) Explain the role of the –OH group in activating the benzene ring.
e) State one use of phenol based on its reactivity.

Item 5: Acidity of Phenol

A student compares the acidity of phenol and ethanol by adding each to aqueous sodium hydroxide. Phenol reacts, while ethanol does not. The teacher emphasizes that this difference can only be explained by considering reaction mechanisms and stability of ions formed.

Questions: a) Write equations showing the reaction of phenol and ethanol with NaOH.
b) Explain why phenol is acidic but ethanol is not.
c) Describe how resonance stabilizes the phenoxide ion.
d) Explain why ethanol does not form a stable ethoxide ion in water.
e) State one consequence of phenol's acidity in chemical reactions.

Item 6:
Formation of Aldehydes from Alkenes

In an organic synthesis experiment, an alkene is converted into an aldehyde using ozonolysis followed by reduction. The chemist explains that the reaction proceeds through an unstable intermediate.

Students are required to focus on the mechanism rather than memorizing reagents.

Task
a)Name the reaction used to cleave the alkene.
b) Write a general equation for ozonolysis of an alkene.
c) Describe the formation of the ozonide intermediate.
d) Explain how aldehydes are produced after reduction.
e) State one advantage of this method of preparing aldehydes.

Item 7: Nucleophilic Addition to Carbonyl Compounds

In a practical lesson, hydrogen cyanide adds to ethanal to form a hydroxynitrile. The teacher stresses that the reaction occurs because of the polarity of the carbonyl group.

Students are guided to explain the process step by step.

Task
a) Identify the functional group in ethanal.
b) Explain why the C=O bond is polar.
c) Describe the nucleophilic addition mechanism of HCN to ethanal.
d) Explain the role of cyanide ion in the reaction.
e) State one use of hydroxynitriles.

Item 8: Reduction of Carbonyl Compounds

In a pharmaceutical laboratory, sodium borohydride is used to reduce aldehydes and ketones to alcohols. The chemist explains that the reagent delivers hydride ions to the carbonyl carbon.

Students are asked to explain the mechanism involved.

Questions: a) Name a reducing agent used for aldehydes and ketones.
b) Write equations for reduction of an aldehyde and a ketone.
c) Describe the hydride transfer mechanism.
d) Explain why carboxylic acids are not reduced by sodium borohydride.
e) State one advantage of using $NaBH_4$.

Item 9: Carboxylic Acids in Food Chemistry

Ethanoic acid is widely used in food preservation. A chemist explains that its properties arise from the carboxyl functional group and hydrogen bonding.

Students are required to link structure to behavior.

Questions: a) Identify the functional group in ethanoic acid.
b) Explain why carboxylic acids have high boiling points.

c) Describe hydrogen bonding between acid molecules.
d) Write the reaction of ethanoic acid with sodium carbonate.
e) State one preservative role of ethanoic acid.

Item 10: Ester Formation Mechanism

In a fragrance factory, ethanoic acid reacts with ethanol in the presence of concentrated sulfuric acid to form ethyl ethanoate. The sweet-smelling ester is collected after heating under reflux.

Students are asked to explain how the ester is formed step by step.

Questions: a) Name the reaction between an alcohol and a carboxylic acid.
b) Write the equation for formation of ethyl ethanoate.
c) Describe the esterification mechanism step by step.
d) State the role of concentrated sulfuric acid.
e) Explain why the reaction is reversible.

Item 11: Hydrolysis of Esters

In soap and detergent chemistry, esters are broken down by heating with acids or alkalis. Students observe that alkaline hydrolysis goes to completion, while acid hydrolysis does not.

The teacher asks students to explain this difference using mechanisms.

Questions: a) Define ester hydrolysis.
b) Write equations for acidic and alkaline hydrolysis of an ester.
c) Explain the mechanism of alkaline hydrolysis.
d) Explain why alkaline hydrolysis is irreversible.
e) State one industrial application of ester hydrolysis.

Item 12: Acyl Chlorides and Their Reactivity

In an advanced practical, ethanoyl chloride reacts violently with water, producing fumes of hydrogen chloride. The teacher explains that acyl chlorides are highly reactive derivatives of carboxylic acids.

Students are asked to explain this behavior.

Questions: a) Identify the functional group in ethanoyl chloride.
b) Write the equation for its reaction with water.
c) Describe the nucleophilic substitution mechanism involved.
d) Explain why acyl chlorides are more reactive than esters.
e) State one use of acyl chlorides.

Item 13: Amide Formation

In pharmaceutical synthesis, carboxylic acids are converted into amides through reactions with ammonia or amines. The chemist emphasizes the importance of reaction conditions.

Students are required to explain how amides are formed.

Questions: a) Write the equation for formation of an amide from an acyl chloride and ammonia.
b) Describe the reaction mechanism involved.
c) Explain why direct reaction of acids with ammonia is difficult.
d) State one property of amides.
e) Give one use of amides.

Item 14: Comparison of Carboxylic Acid Derivatives

A student compares esters, acyl chlorides, and amides and notices differences in reactivity. The teacher explains that these differences arise from the nature of the substituent attached to the carbonyl carbon.

Questions: a) Arrange acyl chlorides, esters, and amides in order of reactivity.
b) Explain the trend using electron-withdrawing ability.
c) Describe how leaving group ability affects mechanism.
d) Explain why amides are least reactive.
e) State one consequence of this trend in synthesis.

Item 15: Organic Mechanisms in Everyday Life

A chemistry teacher explains that many everyday substances—soaps, perfumes, plastics, and medicines—are produced through organic reactions involving specific mechanisms.

Students are asked to reflect on the importance of understanding mechanisms.

Questions: a) Explain what is meant by a reaction mechanism.
b) State why mechanisms are important in organic chemistry.
c) Explain how mechanisms help predict products.
d) Relate organic mechanisms to industrial chemistry.
e) State one challenge students face when learning mechanisms.

📘 Author: joelPCM
📘 UACE CHEMISTRY ITEM BANK

ANSWERS – ORGANIC CHEMISTRY II

Item 1 – Ethanol Dehydration

a) Functional group: Hydroxyl (–OH)

b) Equation:
$C_2H_5OH \rightarrow C_2H_4 + H_2O$ (conc. H_2SO_4, heat)

c) Mechanism (Elimination):

The –OH group is protonated by sulfuric acid forming $-OH_2^+$

Water leaves, forming a carbocation

A proton is removed from the adjacent carbon

A C=C double bond forms producing ethene

d) Roles of H_2SO_4:

Protonates –OH

Acts as a dehydrating agent

e) Ethene decolourises bromine water

Item 2 – Substitution of Alcohols

a) Reagent: Lucas reagent (conc. HCl + $ZnCl_2$)

b) Equation:
$C_2H_5OH + HCl \rightarrow C_2H_5Cl + H_2O$

c) Mechanism (Nucleophilic substitution):

–OH group is protonated

Water leaves

Cl^- attacks the carbocation forming chloroethane

d) $ZnCl_2$ acts as a Lewis acid to polarise the C–O bond

e) Chloroethane is less soluble due to lack of hydrogen bonding

Item 3 – Oxidation of Alcohols

a) Oxidising agent: Acidified potassium dichromate(VI)

b) Equations:
Ethanol → Ethanal → Ethanoic acid

c) Ethanal oxidises easily because aldehydes have hydrogen attached to the carbonyl carbon

d) Distilling off ethanal prevents further oxidation

e) Ethanal gives silver mirror test, ethanoic acid does not

Item 4 – Phenol Reactivity

a) Functional group: Phenolic –OH

b) Equation:
Phenol + $3Br_2$ → 2,4,6-tribromophenol + 3HBr

c) Mechanism (Electrophilic substitution):

–OH donates electron density into ring

Ring attracts Br^+

Substitution occurs rapidly

d) –OH activates the benzene ring by resonance donation

e) Used as disinfectant and antiseptic

Item 5 – Acidity of Phenol

a) Equations:
Phenol + NaOH → Sodium phenoxide + H_2O
Ethanol + NaOH → No reaction

b) Phenol ion is resonance-stabilized

c) Phenoxide ion is stabilised by delocalisation of charge

d) Ethoxide ion is unstable in water

e) Phenol reacts with bases but ethanol does not

Item 6 – Ozonolysis

a) Reaction: Ozonolysis

b) General equation:
Alkene + O_3 → Ozonide → Aldehydes/Ketones

c) Ozone adds across C=C forming ozonide

d) Reduction splits ozonide into carbonyl compounds

e) Produces pure aldehydes cleanly

Item 7 – Nucleophilic Addition

a) Functional group: Aldehyde

b) C=O is polar due to electronegativity difference

c) Mechanism:

CN^- attacks carbonyl carbon

C=O bond breaks

Protonation occurs forming hydroxynitrile

d) CN^- is the nucleophile

e) Hydroxynitriles used in drug synthesis

Item 8 – Reduction of Carbonyls

a) Reducing agent: $NaBH_4$

b) Equations:

Aldehyde → Primary alcohol
Ketone → Secondary alcohol

c) Hydride (H^-) attacks carbonyl carbon

d) Acids are not reactive enough

e) $NaBH_4$ is mild and selective

Item 9 – Carboxylic Acids

a) Functional group: –COOH

b) High boiling point due to hydrogen bonding

c) Molecules form dimers

d) Equation:
$2CH_3COOH + Na_2CO_3 \rightarrow 2CH_3COONa + CO_2 + H_2O$

e) Prevents bacterial growth

Item 10 – Esterification

a) Reaction: Esterification

b) Equation:
$CH_3COOH + C_2H_5OH \rightleftharpoons CH_3COOC_2H_5 + H_2O$

c) Mechanism:

Protonation of C=O

Nucleophilic attack by alcohol

Elimination of water

Deprotonation

d) H_2SO_4 is catalyst and dehydrating agent

e) Reaction is reversible

Item 11 – Ester Hydrolysis

a) Ester hydrolysis is breaking ester into acid and alcohol

b) Acidic: reversible
Alkaline: irreversible

c) Mechanism:
OH^- attacks carbonyl carbon
Tetrahedral intermediate forms
Alcohol leaves

d) Salt formation prevents reverse reaction

e) Soap manufacture

Item 12 – Acyl Chlorides

a) Functional group: –COCl

b) Equation:
$CH_3COCl + H_2O \rightarrow CH_3COOH + HCl$

c) Mechanism:
Water attacks carbonyl carbon
Cl^- leaves

d) Cl^- is excellent leaving group

e) Used to prepare esters and amides

Item 13 – Amide Formation

a) Equation:
$CH_3COCl + 2NH_3 \rightarrow CH_3CONH_2 + NH_4Cl$

b) Nucleophilic substitution

c) Acids react poorly due to salt formation

d) Amides have high melting points

e) Used in drugs and polymers

Item 14 – Reactivity of Derivatives

a) Reactivity order:
Acyl chlorides > Esters > Amides

b) Electron withdrawal decreases down the series

c) Leaving group ability controls rate

d) –NH_2 is poor leaving group

e) Controls synthetic pathways

Item 15 – Importance of Mechanisms

a) A mechanism shows step-by-step bond changes

b) Helps predict products

c) Explains reaction conditions

d) Essential in industrial synthesis

e) Students struggle with electron movement

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Senior Six – Reaction Kinetics

Item 1
Rate of Reaction in a School Laboratory
During a practical chemistry lesson at a Senior Six laboratory, students investigate the reaction between magnesium ribbon and excess dilute hydrochloric acid. They notice that at the beginning, hydrogen gas is produced rapidly, but as time goes on, the reaction slows down and finally stops even though acid is still present. The teacher asks the students to think deeply about what controls the speed of a chemical reaction rather than focusing only on the final products.

Task

a) Define the rate of a chemical reaction.
b) Explain why the reaction is fastest at the beginning.
c) State two factors that affect the rate of this reaction.
d) Sketch a graph of volume of hydrogen gas against time.
e) Explain why the reaction eventually stops.

Item 2: Concentration and Collision Theory

In an industrial laboratory, a chemist increases the concentration of hydrochloric acid used to react with calcium carbonate during the production of carbon dioxide. It is observed that when a more concentrated acid is used, the reaction finishes much faster. The chemist explains that this effect can only be understood using collision theory.

Task
a) State the collision theory of reaction rates.
b) Explain how concentration affects the rate of reaction.
c) Describe what is meant by an effective collision.
d) Explain why increasing concentration does not change the amount of product formed.
e) State one industrial advantage of using high concentration.

Item 3: Temperature Effect on Reaction Rate

A student heats sodium thiosulfate solution before adding hydrochloric acid and notices that the time taken for a cross under the beaker to disappear becomes shorter as temperature increases. The teacher insists that the explanation must involve particle energy and not just “heat makes reactions faster”.

Questions: a) Describe how temperature affects reaction rate.
b) Explain the effect of temperature using collision theory.
c) Define activation energy.
d) Explain why only some collisions lead to reaction.
e) Sketch a Maxwell–Boltzmann distribution curve showing temperature effect.

Item 4: Surface Area in Industrial Processes

At a cement factory, limestone is crushed into fine powder before reacting with acid during quality testing. The factory manager explains that large limestone blocks react too slowly and are unsuitable for rapid analysis.

Questions: a) Define surface area in relation to reaction rate.
b) Explain why powdered limestone reacts faster than large lumps.
c) Describe the effect of surface area using collision theory.
d) State one industrial advantage of increasing surface area.
e) Give one situation where large surface area may be dangerous.

Item 5: Catalysts in Chemical Industry

In the Contact Process for the manufacture of sulfuric acid, vanadium(V) oxide is used as a catalyst. Engineers observe that the catalyst increases the speed of the reaction without being used up.

Questions: a) Define a catalyst.
b) Explain how a catalyst increases the rate of reaction.
c) Describe the effect of a catalyst on activation energy.
d) Sketch an energy profile diagram showing catalysed and uncatalysed reactions.
e) State two advantages of using catalysts in industry.

Item 6: Rate Equation and Order of Reaction

A reaction between substances A and B is studied, and it is found that doubling the concentration of A doubles the rate, while doubling the concentration of B has no effect on the rate. Students are told that the rate equation must be deduced experimentally.

Questions: a) Write the general form of a rate equation.
b) Determine the order of reaction with respect to A.
c) Determine the order of reaction with respect to B.
d) Write the overall order of the reaction.
e) Explain why the rate equation cannot be obtained from the balanced equation.

Item 7: Rate Calculations

In a reaction, the rate is given by:
Rate = $k[A]^2$

At a certain time, the concentration of A is 0.20 mol dm^{-3} and the rate is 4.0×10^{-3} mol dm^{-3} s^{-1}.

Questions: a) Calculate the value of the rate constant, k.
b) State the units of k.
c) Calculate the new rate when concentration of A is doubled.
d) Explain the effect of doubling concentration on rate.
e) State one limitation of rate calculations.

Item 8: Half-life of a Reaction

A radioactive substance used in medical treatment decomposes with first-order kinetics. The hospital chemist explains that the half-life of the substance is constant regardless of its initial concentration.

Questions: a) Define half-life.
b) State one characteristic of a first-order reaction.
c) Explain why half-life is constant for first-order reactions.
d) Sketch a graph of concentration against time.
e) State one application of half-life in science.

Item 9: Determining Rate Experimentally

Students are asked to determine the rate of reaction between sodium thiosulfate and hydrochloric acid using the disappearing cross method. The teacher warns them about experimental errors.

Questions: a) Describe how the disappearing cross method works.
b) State the measurement taken in the experiment.
c) Explain how rate is calculated from the measurement.
d) State two sources of experimental error.
e) Suggest one way to improve accuracy.

Item 10: Activation Energy and Energy Profiles

In a lecture, a teacher explains that reactions do not occur simply because particles collide, but because they must overcome an energy barrier.

Questions: a) Define activation energy.
b) Explain why activation energy is necessary.
c) Describe the shape of an energy profile diagram.
d) Explain the effect of catalysts on the diagram.
e) State one factor that does not affect activation energy.

Item 11: Enzymes as Biological Catalysts

In biology-related chemistry, enzymes are described as highly specific catalysts that work best at certain temperatures and pH values.

Questions: a) Define an enzyme.
b) Explain enzyme specificity.
c) Describe the effect of temperature on enzyme activity.
d) Explain enzyme denaturation.
e) State one difference between enzymes and inorganic catalysts.

Item 12: Reaction Mechanism and Rate-Determining Step

A complex reaction occurs in several steps, but only one step is slow. The chemist explains that this step controls the overall rate.

Questions: a) Define a reaction mechanism.
b) Explain what is meant by the rate-determining step.
c) Explain why the slowest step controls the rate.
d) Relate mechanism to rate equation.
e) State one importance of studying mechanisms.

Item 13: Effect of Pressure on Gaseous Reactions

In the Haber Process, nitrogen and hydrogen gases react under high pressure. Engineers observe that increasing pressure increases the reaction rate.

Questions: a) Explain how pressure affects gaseous reaction rates.
b) Relate pressure to concentration.
c) Explain using collision theory.
d) State one industrial disadvantage of high pressure.
e) Name one gaseous reaction where pressure is important.

Item 14: Comparing Fast and Slow Reactions

Students compare the explosion of hydrogen gas with the rusting of iron. Both are chemical reactions but occur at very different rates.

Questions: a) Define a fast reaction.
b) Define a slow reaction.
c) Explain reasons for difference in rates.
d) State two factors that speed up rusting.
e) Explain why some slow reactions are useful.

Item 15: Importance of Reaction Kinetics

A chemical engineer emphasizes that understanding reaction rates is essential in industry, medicine, and environmental chemistry.

Questions: a) State two industrial applications of kinetics.
b) Explain why reaction rate control is important.
c) Relate kinetics to safety.
d) Explain how kinetics saves cost in industry.
e) State one challenge in studying reaction rates.

ANSWERS – REACTION KINETICS

Item 1 – Rate of Reaction

a) The rate of a chemical reaction is the change in concentration of reactants or products per unit time.

b) The reaction is fastest at the beginning because the concentration of reactants is highest, leading to more frequent effective collisions.

c) Factors affecting rate:

Concentration of acid

Surface area of magnesium

Temperature

d) Graph: A curve that rises steeply at first and then gradually levels off.

e) The reaction stops because magnesium is completely used up, making it the limiting reagent.

Item 2 – Concentration Effect

a) Collision theory states that reactions occur when particles collide with sufficient energy and correct orientation.

b) Increasing concentration increases the number of particles per unit volume, leading to more frequent collisions.

c) An effective collision is one with energy equal to or greater than activation energy.

d) Concentration affects rate, not the final amount of product, since stoichiometry remains unchanged.

e) High concentration reduces reaction time in industry.

Item 3 – Temperature Effect

a) Increasing temperature increases reaction rate.

b) Higher temperature increases kinetic energy, causing more frequent and energetic collisions.

c) Activation energy is the minimum energy required for a reaction to occur.

d) Only collisions with energy ≥ activation energy result in reaction.

e) Curve shifts right and flattens at higher temperature.

Item 4 – Surface Area

a) Surface area refers to the exposed area of a solid reactant.

b) Powdered limestone has more exposed particles.

c) More surface area leads to more frequent collisions.

d) Faster reactions improve industrial efficiency.

e) High surface area increases explosion risk in powders.

Item 5 – Catalysts

a) A catalyst increases reaction rate without being used up.

b) It provides an alternative reaction pathway.

c) Catalyst lowers activation energy.

d) Catalysed pathway has lower energy peak.

e) Advantages:

Lower energy cost

Faster production

Item 6 – Rate Equation

a) Rate = k[A]^m[B]^n

b) Order with respect to A = 1

c) Order with respect to B = 0

d) Overall order = 1

e) Rate equations are determined experimentally, not from stoichiometry.

Item 7 – Rate Calculations

Given:
Rate = $k[A]^2$
Rate = 4.0×10^{-3} mol dm^{-3} s^{-1}
$[A] = 0.20$ mol dm^{-3}

a)
$k = Rate / [A]^2$
$k = (4.0 \times 10^{-3}) / (0.20)^2$
$k = (4.0 \times 10^{-3}) / 0.04$
$k = 0.10$

b) Units of k: dm^3 mol^{-1} s^{-1}

c) New $[A] = 0.40$
New rate = $k(0.40)^2$
$= 0.10 \times 0.16$
$= 0.016$ mol dm^{-3} s^{-1}

d) Doubling concentration increases rate four times.

e) Assumes ideal conditions.

Item 8 – Half-life

a) Half-life is the time taken for concentration to reduce to half its initial value.

b) First-order reactions have constant half-life.

c) Rate depends only on concentration of one reactant.

d) Exponential decay curve.

e) Used in radioactive dating and medicine.

Item 9 – Experimental Rate

a) A cross disappears as sulfur precipitate forms.

b) Time taken for cross to disappear.

c) Rate $\propto$ 1/time

d) Errors:

Human reaction time

Uneven lighting

e) Repeat experiment and average results.

Item 10 – Activation Energy

a) Minimum energy needed to start reaction.

b) Needed to break bonds.

c) Curve rises to peak then falls.

d) Catalyst lowers peak height.

e) Concentration does not affect activation energy.

Item 11 – Enzymes

a) Enzymes are biological catalysts.

b) Each enzyme is specific to a substrate.

c) Activity increases then decreases with temperature.

d) Denaturation destroys active site.

e) Enzymes are affected by pH, inorganic catalysts are not.

Item 12 – Mechanism and RDS

a) A mechanism is the step-by-step pathway of a reaction.

b) Rate-determining step is the slowest step.

c) It limits overall speed.

d) Rate equation reflects slow step.

e) Helps control reaction conditions.

Item 13 – Pressure Effect

a) Increasing pressure increases reaction rate.

b) Higher pressure increases gas concentration.

c) More frequent collisions occur.

d) High pressure is costly and risky.

e) Haber Process.

Item 14 – Fast vs Slow Reactions

a) Fast reaction occurs almost instantly.

b) Slow reaction takes long time.

c) Differences due to activation energy and conditions.

d) Moisture and oxygen speed rusting.

e) Slow reactions allow control.

Item 15 – Importance of Kinetics

a) Applications:

Industrial synthesis

Drug metabolism

b) Prevents dangerous reactions.

c) Controls explosions.

d) Saves energy and time.

e) Measuring very fast reactions.

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Senior Six – Polymers and Amines

POLYMERS
Item 1:
Plastic Waste and Polymer Formation

In a town near Lake Victoria, a recycling company collects large amounts of plastic waste such as bottles, polythene bags, and food containers. During a school visit, the plant manager explains that most of these plastics are formed from small molecules that join together in long chains. He further explains that different plastics are made using different polymerisation mechanisms, which determine whether the plastic is flexible, rigid, or resistant to heat.

Students are asked to analyze the chemistry behind plastic formation.

Task
a) Define a polymer and a monomer.
b) Identify the type of polymerisation involved in making polythene.
c) Write the equation for the formation of polythene from ethene.
d) Describe the mechanism of addition polymerisation.
e) Explain why polythene is chemically unreactive.

Item 2: Nylon Production in Textile Industry

A textile factory manufactures nylon fibres for making clothes and fishing nets. The factory chemist explains that nylon is produced by reacting two different monomers, and that a small molecule is eliminated during the reaction. The strength of nylon is attributed to forces between polymer chains.

Students are expected to explain both the chemistry and properties involved.

Questions: a) Identify the type of polymer formed in nylon manufacture.
b) Name the two functional groups involved in nylon formation.
c) Write a general equation for nylon formation.

d) Describe the condensation polymerisation mechanism.
e) Explain why nylon fibres are strong.

Item 3: Comparing Addition and Condensation Polymers

During revision, students compare polythene, PVC, nylon, and terylene. The teacher insists that the comparison must be based on reaction mechanism, structure, and by-products formed.

Questions: a) Distinguish between addition and condensation polymerisation.
b) Give one example of each type.
c) Explain why condensation polymers can be hydrolysed.
d) Describe the role of functional groups in condensation polymerisation.
e) State one environmental concern related to each polymer type.

Item 4: Rubber and Elasticity

At a rubber-processing plant, natural rubber is vulcanised before being used to make tyres. The engineer explains that untreated rubber becomes soft in hot weather and brittle in cold weather, but vulcanisation improves its properties.

Students are required to explain these observations.

Questions: a) Name the monomer used to make natural rubber.
b) Describe the structure of natural rubber.
c) Explain the process of vulcanisation.
d) Explain how sulfur improves elasticity.
e) State two advantages of vulcanised rubber.

Item 5: Biodegradable Polymers

Environmental scientists promote biodegradable plastics to reduce pollution. They explain that some polymers can be broken down by hydrolysis or microorganisms because of their chemical structure.

Questions: a) Define a biodegradable polymer.
b) Explain why some polymers are biodegradable.
c) Name one biodegradable polymer.
d) Explain the role of ester or amide links in biodegradability.
e) State one limitation of biodegradable polymers.

AMINES

Item 6: Amines in Fish Processing

In a fish market, spoiled fish produces a strong unpleasant smell. A chemistry teacher explains that this smell is due to basic nitrogen-containing compounds formed during protein decomposition.

Students are asked to investigate the chemistry behind these compounds.

Questions: a) Define an amine.
b) Classify amines into primary, secondary, and tertiary.
c) Explain why amines have unpleasant smells.
d) Describe the shape around the nitrogen atom in amines.
e) Explain why amines are basic.

Item 7: Preparation of Amines

In a laboratory synthesis, a student prepares an amine from a haloalkane using excess ammonia in ethanol. The teacher warns that side reactions may occur if conditions are not controlled.

Students must explain the chemistry involved.

Questions: a) Write the equation for preparation of a primary amine from a haloalkane.
b) Describe the nucleophilic substitution mechanism involved.
c) Explain why excess ammonia is used.
d) State one side reaction that may occur.
e) Explain how the amine product can be separated.

Item 8: Basicity of Amines

A student compares the basicity of ammonia, methylamine, and phenylamine. The teacher explains that the trend cannot be understood without considering electron donation and structure.

Questions: a) Define basicity in terms of proton acceptance.
b) Arrange ammonia, methylamine, and phenylamine in order of increasing basicity.
c) Explain the effect of alkyl groups on basicity.
d) Explain why phenylamine is less basic than methylamine.
e) State one use of amines based on their basic nature.

Item 9: Reactions of Amines

In a qualitative analysis test, an amine reacts with nitrous acid to produce different products depending on the type of amine present. The teacher explains that this reaction is important for distinguishing amines.

Questions: a) Name the reagent used to generate nitrous acid in situ.
b) Describe the reaction of a primary aliphatic amine with nitrous acid.
c) State the observation made during the reaction.
d) Explain why nitrogen gas is evolved.
e) State one analytical use of this reaction.

Item 10: Amines in Medicine and Industry

A pharmaceutical chemist explains that many drugs contain amine groups, which influence solubility and biological activity. Students are encouraged to appreciate the importance of amines beyond the classroom.

Questions: a) Explain why amines increase solubility of drugs in water.
b) Describe salt formation between amines and acids.
c) Explain how amines act as intermediates in synthesis.
d) State one example of a drug containing an amine group.
e) Explain one safety concern when handling amines.

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ANSWERS – POLYMERS & AMINES

POLYMERS

Item 1 – Polymer Formation (Polythene)

a) A polymer is a large molecule formed by joining many small molecules called monomers.

b) Polythene is formed by addition polymerisation.

c) Equation:
$nCH_2{=}CH_2 \rightarrow (-CH_2-CH_2-)_n$

d) Mechanism (Addition polymerisation):

The C=C double bond in ethene opens

Free radicals initiate the reaction

Monomers add successively to form a long chain

No small molecule is eliminated

e) Polythene is unreactive because it contains strong C–C and C–H bonds and has no polar functional groups.

Item 2 – Nylon Production

a) Nylon is a condensation polymer.

b) Functional groups involved: –COOH (carboxylic acid) and $–NH_2$ (amine).

c) General equation:
Diamine + Dicarboxylic acid → Nylon + Water

d) Mechanism (Condensation polymerisation):

$–NH_2$ group attacks the carbonyl carbon

Amide linkage (–CONH–) forms

Water molecule is eliminated

Process repeats forming long chains

e) Nylon is strong due to hydrogen bonding between chains and close packing.

Item 3 – Addition vs Condensation Polymers

a) Addition polymerisation involves joining monomers without loss of small molecules, while condensation polymerisation involves elimination of small molecules like water.

b) Examples:
Addition – Polythene
Condensation – Nylon

c) Condensation polymers can be hydrolysed because they contain ester or amide links.

d) Functional groups allow bond formation between monomers.

e) Environmental concerns:
Addition – non-biodegradable waste
Condensation – slow hydrolysis in environment

Item 4 – Rubber and Vulcanisation

a) Monomer: Isoprene

b) Natural rubber has long coiled polymer chains with C=C bonds.

c) Vulcanisation involves heating rubber with sulfur.

d) Sulfur forms cross-links between chains, preventing slippage.

e) Advantages:

Increased strength

Improved elasticity

Resistant to temperature changes

Item 5 – Biodegradable Polymers

a) A biodegradable polymer can be broken down by microorganisms.

b) They contain hydrolysable links.

c) Example: Polylactic acid

d) Ester and amide links are easily hydrolysed.

e) Limitation: Often less durable.

AMINES

Item 6 – Amines in Fish

a) Amines are organic compounds containing –NH_2, –NHR, or –NR_2.

b) Primary (RNH_2), Secondary (R_2NH), Tertiary (R_3N).

c) They have strong fishy smells due to volatility.

d) Nitrogen has a trigonal pyramidal shape.

e) Amines are basic due to lone pair on nitrogen.

Item 7 – Preparation of Amines

a) Equation:
$R–X + 2NH_3 \rightarrow R–NH_2 + NH_4X$

b) Mechanism (Nucleophilic substitution):

NH_3 attacks carbon atom

Halide ion leaves

Proton transfer occurs

c) Excess ammonia prevents formation of secondary amines.

d) Side reaction: Formation of secondary or tertiary amines.

e) Product separated by acid–base extraction.

Item 8 – Basicity of Amines

a) Basicity is the ability to accept a proton.

b) Increasing basicity:
Phenylamine < Ammonia < Methylamine

c) Alkyl groups donate electrons, increasing basicity.

d) Phenylamine is less basic due to delocalisation of lone pair into benzene ring.

e) Used in manufacture of dyes and drugs.

Item 9 – Reaction with Nitrous Acid

a) Reagent: Sodium nitrite and hydrochloric acid.

b) Primary aliphatic amine forms alcohol.

c) Effervescence observed.

d) Nitrogen gas is released due to instability of diazonium intermediate.

e) Used to identify primary amines.

Item 10 – Amines in Medicine

a) Amines form soluble salts with acids.

b) They react with acids forming ammonium salts.

c) Amines act as intermediates in synthesis.

d) Example: Morphine, caffeine.

e) Amines can be toxic and irritating.

📘 ***Author: joelPCM***
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UACE CHEMISTRY – ADDITIONAL ITEMS

(Integrated / Combined Topics)

Item 1: Moles & Thermochemistry

A rural health centre uses calcium oxide to generate heat for emergency warming by adding water. During a science outreach program, students are asked to calculate how much heat is released when a known mass of calcium oxide reacts completely with excess water. The nurse also wants to know how much water is required to ensure complete reaction without wastage.

Questions:
a) Write the balanced equation for the reaction.
b) Calculate the number of moles of calcium oxide used.
c) Calculate the heat released using enthalpy data.
d) Explain why the reaction is exothermic.
e) State one safety precaution during the reaction.

Item 2: Atomic Structure & Periodicity

A chemistry teacher explains to students why sodium is highly reactive while magnesium is less reactive, even though they are in the same period. The discussion involves electron arrangement, ionization energy, and atomic size, with reference to real-life uses of these metals.

Questions:
a) Write the electronic configurations of sodium and magnesium.
b) Explain the trend in ionization energy across the period.
c) Explain the difference in reactivity.
d) Relate atomic size to reactivity.
e) State one use of each metal based on its properties.

Item 3: Bonding & Thermochemistry

During the manufacture of ammonia, engineers monitor energy changes and bond formation. Students are told that understanding bond energies helps explain why heat is released or absorbed during reactions.

Questions:
a) Write the equation for ammonia formation.
b) Identify the type of bonding in ammonia.
c) Use bond energies to explain the enthalpy change.
d) Explain why the reaction is reversible.
e) State one industrial condition that affects yield.

Item 4: Periodicity & Bonding

A student compares the bonding in sodium chloride and magnesium oxide while revising trends across Period 3. The teacher emphasizes the role of charge density and lattice energy.

Questions:
a) Identify the type of bonding in each compound.
b) Explain the difference in melting points.
c) Relate charge and ionic radius to lattice energy.
d) Explain why magnesium oxide is less soluble in water.
e) State one use of each compound.

Item 5: Equilibria I & Acids–Bases

A water treatment plant uses buffer solutions to maintain stable pH during purification. The chemist explains how weak acids and their salts resist pH changes.

Questions:
a) Define a buffer solution.
b) Identify the components of an acidic buffer.
c) Explain how equilibrium shifts when acid is added.
d) Write the relevant equilibrium equation.
e) State one importance of buffers in daily life.

Item 6: Reaction Kinetics & Equilibria

In the Haber Process, engineers carefully control temperature to balance rate and yield. Students are asked to analyze this compromise.

Questions:
a) Write the equation for the Haber Process.
b) Explain the effect of temperature on reaction rate.
c) Explain the effect of temperature on equilibrium position.
d) State why a compromise temperature is used.
e) Explain the role of the catalyst.

Item 7: Organic Chemistry I & Thermochemistry

An alkene undergoes combustion in a laboratory demonstration, producing a large amount of heat. Students compare this with the combustion of an alkane.

Questions:
a) Write the general formulae of alkenes and alkanes.
b) Write a balanced combustion equation for an alkene.
c) Explain why combustion is exothermic.
d) Compare the enthalpy of combustion of alkenes and alkanes.
e) State one environmental concern of combustion reactions.

Item 8: Organic Chemistry II & Reaction Mechanisms

Ethanol is converted into ethene and then polymerized to make polythene in an industrial setup. Students are guided to follow the chemistry step by step.

Questions:
a) Name the reaction converting ethanol to ethene.
b) Describe the reaction mechanism involved.
c) Write the equation for polymerization of ethene.
d) State the type of polymer formed.
e) Give one use of the polymer.

Item 9: Polymers & Environmental Chemistry

An environmental scientist explains why plastic waste persists in soil for many years, while some newer plastics degrade faster.

Questions:
a) Define a polymer.
b) Explain why addition polymers are non-biodegradable.

c) Identify functional groups that allow biodegradation.
d) Name one biodegradable polymer.
e) State one limitation of biodegradable plastics.

Item 10: Amines & Acid–Base Chemistry

In a pharmaceutical lab, an amine drug is converted into its salt form to improve solubility. Students are asked to explain the chemistry involved.

Questions:
a) Define an amine.
b) Explain why amines are basic.
c) Write an equation showing salt formation.
d) Explain why the salt is more soluble.
e) State one advantage of this process in medicine.

Item 11: Electrochemistry & Redox

A car battery operates through redox reactions. Students are asked to relate electron flow to oxidation and reduction processes.

Questions:
a) Define oxidation and reduction.
b) Identify the anode and cathode.
c) Explain electron flow in the circuit.
d) State one factor affecting cell voltage.
e) Give one use of electrochemical cells.

Item 12: Periodicity II & Transition Metals

A factory uses iron as a catalyst and copper for electrical wiring. Students analyze how periodic trends explain these uses.

Questions:
a) Define a transition metal.
b) Explain catalytic behavior of iron.
c) Explain high conductivity of copper.
d) Relate d-electrons to properties.
e) State one limitation of transition metals.

Item 13: Moles & Gas Laws

Oxygen gas is collected over water in a laboratory experiment. Students must calculate the volume of dry gas produced.

Questions:
a) State the ideal gas equation.
b) Calculate moles of gas produced.
c) Explain correction for water vapor.
d) State assumptions made.
e) Give one source of error.

Item 14: Kinetics & Catalysts

An enzyme speeds up digestion at body temperature but fails at high temperatures. Students compare this with industrial catalysts.

Questions:
a) Define a catalyst.
b) Explain enzyme specificity.
c) Explain denaturation.
d) Compare enzymes with inorganic catalysts.
e) State one industrial application of catalysis.

Item 15: Equilibria II & Colligative Properties

Salt is added to icy roads to melt ice. Students are asked to explain this using equilibrium concepts.

Questions:
a) Define freezing point depression.
b) Explain effect of salt on ice–water equilibrium.
c) Relate to number of particles.
d) State one limitation of this method.
e) Give one environmental concern.

Item 16: Organic Chemistry I & Kinetics

An alkene reacts faster than an alkane in an addition reaction. Students analyze why.

Questions:
a) Identify the functional group in alkenes.
b) Explain reaction mechanism difference.
c) Relate bond type to activation energy.
d) Explain effect of temperature.
e) State one industrial implication.

Item 17: Thermochemistry & Fuels

Different fuels release different amounts of heat when burned. Students compare ethanol and petrol.

Questions:
a) Define enthalpy of combustion.
b) Write equations for combustion.
c) Explain difference in heat output.
d) Relate structure to energy content.
e) State one advantage of biofuels.

Item 18: Polymers & Bonding

Kevlar is used in bulletproof vests due to its strength. Students analyze the bonding involved.

Questions:
a) Identify the type of polymer.
b) Explain intermolecular forces present.
c) Relate structure to strength.
d) Explain resistance to heat.
e) State one limitation of such polymers.

Item 19: Amines & Organic Reactions

A primary amine reacts differently from a tertiary amine during qualitative analysis.

Questions:
a) Classify amines.
b) Describe reaction with nitrous acid.
c) State observations.
d) Explain difference in reactivity.
e) State one analytical use.

Item 20: Integrated Chemistry in Industry

A chemical plant produces fertilizers, plastics, and fuels. Engineers rely on moles, energy changes, rates, and equilibria to run the plant efficiently.

Questions:
a) Explain the role of stoichiometry.
b) Explain importance of thermochemistry.
c) Explain why kinetics matters.
d) Explain role of equilibria.

e) State one consequence of poor control.

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ADDITIONAL ITEMS – ANSWERS

Item 1: Moles & Thermochemistry

a) $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s)$
b) Number of moles = mass ÷ molar mass of CaO
c) Heat released, $q = n \times \Delta H$ (negative sign indicates heat released)
d) The reaction forms stronger bonds in $Ca(OH)_2$ than those broken, releasing excess energy
e) Wear gloves and goggles because CaO reacts violently with water

Item 2: Atomic Structure & Periodicity

a) Na: $1s^2\ 2s^2\ 2p^6\ 3s^1$; Mg: $1s^2\ 2s^2\ 2p^6\ 3s^2$
b) Ionization energy increases across the period due to increasing nuclear charge
c) Sodium loses one electron easily, magnesium requires more energy
d) Atomic radius decreases across the period, increasing attraction to electrons
e) Sodium is used in sodium lamps; magnesium is used in flares

Item 3: Bonding & Thermochemistry

a) $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$
b) Covalent bonding
c) More energy is released forming N–H bonds than is absorbed breaking N≡N and H–H bonds
d) The reaction is reversible because ammonia decomposes at high temperature
e) High pressure increases yield

Item 4: Periodicity & Bonding

a) Both have ionic bonding
b) MgO has a higher melting point due to higher lattice energy
c) Mg^{2+} has higher charge and smaller radius than Na^+
d) MgO has stronger ionic bonds, making it less soluble
e) NaCl is used as table salt; MgO is used as a refractory material

Item 5: Equilibria I & Acids–Bases

a) A buffer resists changes in pH
b) Weak acid and its conjugate base
c) Added H^+ reacts with the conjugate base

d) $CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$
e) Maintaining blood pH

Item 6: Reaction Kinetics & Equilibria

a) $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$
b) Higher temperature increases collision frequency
c) Higher temperature shifts equilibrium to the left
d) A compromise temperature balances rate and yield
e) Catalyst lowers activation energy

Item 7: Organic I & Thermochemistry

a) Alkanes: C_nH_{2n+2} ; Alkenes: C_nH_{2n}
b) $C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$
c) Formation of strong C=O and O–H bonds releases energy
d) Alkenes generally have higher enthalpy of combustion
e) Air pollution

Item 8: Organic II & Mechanisms

a) Dehydration
b) Elimination mechanism involving protonation and loss of water
c) $nCH_2{=}CH_2 \rightarrow (-CH_2-CH_2-)_n$
d) Addition polymer
e) Making plastic bags

Item 9: Polymers & Environment

a) A large molecule made from repeating units
b) They lack functional groups attacked by microbes
c) Ester or amide groups
d) Polylactic acid
e) High cost

Item 10: Amines & Acid–Base

a) Organic compounds derived from ammonia
b) Lone pair on nitrogen accepts protons
c) $RNH_2 + HCl \rightarrow RNH_3^+Cl^-$
d) Ionic salts dissolve easily in water
e) Better drug absorption

Item 11: Electrochemistry & Redox

a) Oxidation is loss of electrons; reduction is gain
b) Anode: oxidation; Cathode: reduction
c) Electrons flow from anode to cathode
d) Concentration of electrolyte
e) Car batteries

Item 12: Periodicity II & Transition Metals

a) A metal with partially filled d-orbitals
b) Variable oxidation states allow surface reactions
c) Free electrons allow charge flow
d) d-electrons enable variable oxidation states
e) Corrosion

Item 13: Moles & Gas Laws

a) $PV = nRT$
b) n = mass ÷ molar mass
c) Subtract vapor pressure of water
d) Gas behaves ideally
e) Gas leakage

Item 14: Kinetics & Catalysts

a) A substance that increases rate without being used up
b) Enzymes are specific due to active sites
c) High temperature destroys protein structure
d) Enzymes work at mild conditions
e) Manufacture of sulfuric acid

Item 15: Equilibria II & Colligative Properties

a) Lowering of freezing point due to solute
b) Salt shifts ice–water equilibrium
c) Depends on number of particles
d) Causes corrosion
e) Soil pollution

Item 16: Organic I & Kinetics

a) C=C double bond
b) Alkenes undergo electrophilic addition

c) π-bond breaks easily
d) Higher temperature increases rate
e) Polymer manufacture

Item 17: Thermochemistry & Fuels

a) Heat released when one mole burns completely
b) Balanced combustion equations
c) Petrol has more C–H bonds
d) More bonds release more energy
e) Renewable source

Item 18: Polymers & Bonding

a) Condensation polymer
b) Hydrogen bonding
c) Strong intermolecular forces
d) High thermal stability
e) Expensive

Item 19: Amines & Organic Reactions

a) Primary, secondary, tertiary
b) Primary amines form alcohols and nitrogen gas
c) Effervescence observed
d) Steric hindrance in tertiary amines
e) Identification of amines

Item 20: Integrated Industrial Chemistry

a) Determines correct reactant amounts
b) Controls energy efficiency
c) Controls production speed
d) Maximizes yield
e) Economic loss
-
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UACE CHEMISTRY – APPLIED ORGANIC CHEMISTRY ITEMS

(Additional Items – Questions Only)

Item 1: Fuels, Alkanes & Cracking

At a petroleum depot in Kampala, engineers explain to visiting students how crude oil is separated into useful fractions before being converted into fuels suitable for vehicles. The students observe that heavy fractions are further processed to produce petrol, diesel, and LPG, which are in high demand due to increasing transport needs.

Questions:
a) Name the process used to separate crude oil into fractions.
b) Explain why heavy fractions are not directly useful as fuels.
c) Describe the cracking process and state its purpose.
d) Write an equation for the cracking of a long-chain alkane producing an alkane and an alkene.
e) State one environmental concern associated with fuel combustion.

Item 2: Alcohols & Industrial Ethanol

A distillery produces ethanol for use as fuel, solvent, and antiseptic. The chemist explains that ethanol can be produced both biologically and industrially, depending on its intended use. Students are asked to compare the two methods.

Questions:
a) Name the biological method of producing ethanol.
b) Write the equation for the fermentation process.
c) State two conditions required for fermentation.
d) Explain one industrial use of ethanol based on its properties.
e) Explain why ethanol is blended with petrol in some countries.

Item 3: Alcohols, Oxidation & Breathalysers

Traffic police use breathalysers to test drivers suspected of drunk driving. The device relies on the chemical properties of ethanol when it reacts with oxidizing agents under controlled conditions.

Questions:
a) Identify the functional group present in ethanol.
b) Describe what happens when ethanol is oxidized.
c) Name a suitable oxidizing agent used in breathalysers.
d) Explain the chemical principle behind alcohol detection.
e) State one limitation of chemical breath testing.

Item 4: Esters, Perfumes & Food Flavourings

In a food-processing factory, esters are synthesized to produce artificial fruit flavours used in soft drinks and sweets. The quality control officer emphasizes the importance of reaction conditions and purity.

Questions:
a) Define an ester.
b) Name the type of reaction used to form esters.
c) Write a general equation for esterification.
d) State two conditions required for ester formation.
e) Give one reason why esters are suitable as food flavourings.

Item 5: Soaps, Detergents & Environmental Impact

A community near a lake complains that soap foam persists on the water surface for long periods. Environmental scientists investigate the chemistry behind soap and detergent use.

Questions:
a) Describe how soap is manufactured from fats or oils.
b) Explain why soap does not work well in hard water.
c) State one chemical difference between soaps and detergents.
d) Explain why detergents persist longer in the environment.
e) State one environmental problem caused by detergents.

Item 6: Polymers & Packaging Industry

A packaging company replaces glass containers with plastic bottles due to cost and durability. Engineers discuss the chemical structure of plastics and how this affects their properties.

Questions:
a) Define a polymer.
b) Distinguish between addition and condensation polymers.
c) Identify the type of polymer used in plastic bottles.
d) Explain why plastics are chemically inert.
e) State one disadvantage of plastic packaging.

Item 7: Polymers & Recycling

At a recycling plant, workers sort plastics using recycling codes. Students learn that not all plastics can be recycled easily due to differences in chemical structure.

Questions:
a) Explain why thermoplastics can be recycled.
b) Explain why thermosetting plastics cannot be recycled.

c) Relate polymer structure to melting behavior.
d) State one challenge in plastic recycling.
e) Suggest one solution to plastic waste management.

Item 8: Amines & Pharmaceutical Applications

A pharmaceutical company converts an amine drug into a salt before packaging it as tablets. The chemist explains that this improves drug performance in the human body.

Questions:
a) Define an amine.
b) Explain why amines are basic.
c) Describe how amine salts are formed.
d) Explain why salt formation improves drug solubility.
e) State one advantage of using amines in medicine.

Item 9: Amines & Dyes Industry

In a textile factory, brightly coloured dyes are produced from aromatic amines. The factory manager emphasizes strict safety procedures due to the reactive nature of these compounds.

Questions:
a) Distinguish between aliphatic and aromatic amines.
b) Explain why aromatic amines are useful in dye manufacture.
c) State one health hazard associated with aromatic amines.
d) Explain why protective equipment is necessary.
e) Give one economic importance of synthetic dyes.

Item 10: Organic Acids & Food Preservation

A juice-processing company uses organic acids to extend shelf life. The quality assurance officer explains how acidity prevents spoilage.

Questions:
a) Name one organic acid used in food preservation.
b) Identify the functional group present.
c) Explain how acids prevent microbial growth.
d) Explain why controlled concentration is important.
e) State one disadvantage of excessive acid use in food.

Item 11: Petrochemicals & Alkenes

An industrial plant converts alkenes into plastics, alcohols, and detergents. Students observe that alkenes are more reactive than alkanes.

Questions:
a) Identify the functional group in alkenes.
b) Explain why alkenes are more reactive than alkanes.
c) Name one industrial reaction involving alkenes.
d) Explain the importance of alkenes in petrochemicals.
e) State one safety concern in alkene handling.

Item 12: Organic Solvents & Safety

In a laboratory, organic solvents such as ethanol and propanone are used for cleaning and extraction. The lab technician stresses fire safety and ventilation.

Questions:
a) State one property that makes organic solvents effective cleaners.
b) Explain why many organic solvents are flammable.
c) Relate volatility to molecular structure.
d) State one safety precaution when using solvents.
e) Give one environmental concern related to solvent disposal.

Item 13: Polymers & Fibres

A textile manufacturer produces nylon fibres used in fishing nets and clothing. Engineers explain how molecular structure affects strength and elasticity.

Questions:
a) Name the type of polymer formed in nylon.
b) Identify the linkage present in nylon.
c) Explain why nylon fibres are strong.
d) State one use of nylon based on its properties.
e) State one disadvantage of synthetic fibres.

Item 14: Organic Chemistry & Environmental Pollution

An environmental officer investigates air pollution caused by incomplete combustion of fuels in urban areas.

Questions:
a) Define incomplete combustion.
b) Identify one organic pollutant formed.
c) Explain how the pollutant affects human health.
d) Suggest one chemical solution to reduce emissions.
e) State one role of chemistry in environmental protection.

Item 15: Integrated Applied Organic Chemistry

A chemical industry produces fuels, plastics, medicines, and detergents. Chemists must understand structure, reactions, and environmental impact to ensure sustainability.

Questions:
a) Explain the importance of functional groups in organic chemistry.
b) Explain how organic reactions support industrial production.
c) Explain the role of polymers in modern life.
d) Explain why applied organic chemistry must consider the environment.
e) State one career related to applied organic chemistry.

APPLIED ORGANIC CHEMISTRY – ANSWERS

Item 1: Fuels, Alkanes & Cracking

a) Fractional distillation
b) They have high boiling points and burn inefficiently
c) Cracking breaks long-chain alkanes into shorter alkanes and alkenes to increase fuel supply
d) Example:
$C_{10}H_{22} \rightarrow C_8H_{18} + C_2H_4$
e) Air pollution due to carbon dioxide and carbon monoxide

Item 2: Alcohols & Industrial Ethanol

a) Fermentation
b) $C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$
c) Yeast and absence of oxygen (temperature about 30–40 °C)
d) Ethanol dissolves many organic compounds, making it a good solvent
e) Ethanol burns cleanly and reduces reliance on fossil fuels

Item 3: Alcohols, Oxidation & Breathalysers

a) Hydroxyl (–OH) group
b) Ethanol is oxidized to ethanal or ethanoic acid
c) Acidified potassium dichromate(VI)
d) Colour change occurs when ethanol is oxidized
e) Other substances may give false results

Item 4: Esters, Perfumes & Food Flavourings

a) An ester is an organic compound formed from an acid and an alcohol
b) Esterification (condensation reaction)
c) $RCOOH + R'OH \rightleftharpoons RCOOR' + H_2O$
d) Concentrated sulfuric acid and heating
e) Esters have pleasant fruity smells

Item 5: Soaps, Detergents & Environmental Impact

a) By saponification of fats or oils with alkali
b) Soap reacts with Ca^{2+} or Mg^{2+} ions forming scum
c) Detergents contain sulfonate groups instead of carboxylate groups
d) Detergents are not easily biodegradable
e) Water pollution and eutrophication

Item 6: Polymers & Packaging Industry

a) A polymer is a large molecule made from repeating units
b) Addition polymers form from alkenes; condensation polymers eliminate small molecules
c) Addition polymer (e.g. poly(ethene))
d) Strong covalent bonds make plastics unreactive
e) Environmental pollution

Item 7: Polymers & Recycling

a) Thermoplastics soften on heating and can be reshaped
b) Thermosetting plastics have cross-linked structures
c) Cross-linking prevents melting
d) Sorting and contamination
e) Recycling programs and biodegradable plastics

Item 8: Amines & Pharmaceutical Applications

a) Organic derivatives of ammonia
b) Lone pair of electrons on nitrogen accepts protons
c) Reaction of amines with acids to form salts
d) Ionic salts dissolve better in water
e) Improved drug effectiveness

Item 9: Amines & Dyes Industry

a) Aliphatic amines have alkyl groups; aromatic amines have aryl groups
b) They form coloured azo compounds
c) Some are carcinogenic
d) To prevent inhalation and skin contact
e) Textile and clothing industry

Item 10: Organic Acids & Food Preservation

a) Ethanoic acid or citric acid
b) Carboxyl (–COOH) group
c) Low pH inhibits microbial growth
d) Excess acid affects taste and health
e) Tooth decay or irritation

Item 11: Petrochemicals & Alkenes

a) Carbon–carbon double bond (C=C)
b) Presence of π-bond makes them reactive
c) Polymerization
d) They are starting materials for many products
e) Fire hazard

Item 12: Organic Solvents & Safety

a) They dissolve grease and organic substances
b) They evaporate easily and ignite
c) Weak intermolecular forces increase volatility
d) Use in well-ventilated areas
e) Soil and water contamination

Item 13: Polymers & Fibres

a) Condensation polymer
b) Amide (–CONH–) linkage
c) Strong hydrogen bonding between chains
d) Making ropes and fabrics
e) Non-biodegradable

Item 14: Organic Chemistry & Environmental Pollution

a) Burning fuel in limited oxygen
b) Carbon monoxide
c) It reduces oxygen transport in blood
d) Use of catalytic converters

e) Pollution control

Item 15: Integrated Applied Organic Chemistry

a) They determine chemical properties and reactions
b) Enable synthesis of useful products
c) Provide materials for daily use
d) To reduce environmental damage
e) Chemical engineer or industrial chemist

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Item 1: Stoichiometry, Gases & Industrial Chemistry

A fertilizer factory in eastern Uganda produces ammonia using nitrogen from air and hydrogen obtained from natural gas. Engineers carefully calculate reactant quantities to avoid wastage. During one production run, nitrogen is supplied in slight excess while hydrogen is the limiting reactant. Temperature and pressure are continuously monitored to ensure efficiency.

Questions:
a) Write the balanced equation for ammonia formation.
b) Define the term limiting reagent.
c) Explain why hydrogen is made the limiting reagent.
d) Calculate the maximum amount of ammonia formed given suitable data.
e) State one economic reason for controlling reactant ratios.

Item 2: Atomic Structure & Spectroscopy

During a physics–chemistry joint lesson, students observe emission spectra of hydrogen using a discharge tube. The teacher explains that the coloured lines correspond to electron transitions between energy levels. The discussion links these observations to atomic stability.

Questions:
a) Explain why hydrogen produces a line spectrum.
b) Define an energy level.
c) Explain what happens when an electron falls to a lower energy level.
d) State one limitation of the Bohr model.
e) Explain one application of atomic spectra.

Item 3: Periodicity & Reactivity

A metals workshop uses sodium, magnesium, and aluminium for different purposes. Students are asked to explain why these metals show different chemical reactivities despite being in the same period of the periodic table.

Questions:
a) Write the electronic configuration of each metal.
b) Describe the trend in atomic radius across the period.
c) Explain the trend in metallic reactivity.
d) Relate ionization energy to reactivity.
e) State one use of aluminium based on its properties.

Item 4: Bonding, Structure & Properties

A civil engineer compares sodium chloride and silicon dioxide while selecting materials for road construction. One compound dissolves easily in water, while the other is extremely hard and insoluble.

Questions:
a) Describe the bonding in sodium chloride.
b) Describe the bonding in silicon dioxide.
c) Explain the difference in melting points.
d) Relate structure to solubility.
e) State one industrial use of silicon dioxide.

Item 5: Thermochemistry & Fuels

Different fuels are tested in a laboratory to compare the heat released during combustion. Students burn equal masses of ethanol and paraffin and observe different temperature rises in water.

Questions:
a) Define enthalpy of combustion.
b) Write balanced equations for the combustion of ethanol and paraffin.
c) Explain why the fuels release different amounts of heat.
d) State two sources of experimental error.
e) Explain one advantage of using ethanol as a fuel.

Item 6: Chemical Equilibria I

In a chemical plant, sulfur dioxide is converted to sulfur trioxide in a reversible reaction. Engineers must adjust conditions to maximize yield without slowing production.

Questions:
a) Write the equilibrium equation involved.

b) State Le Chatelier’s principle.
c) Explain the effect of increasing pressure.
d) Explain the effect of increasing temperature.
e) State the role of a catalyst in the process.

Item 7: Equilibria II & Colligative Properties

During cold seasons in mountainous regions, salt is spread on icy roads to prevent accidents. Students are asked to explain the chemistry behind this practice.

Questions:
a) Define freezing point depression.
b) Explain how adding salt affects ice–water equilibrium.
c) Explain why calcium chloride is more effective than sodium chloride.
d) State one disadvantage of using salts on roads.
e) Give one environmental concern.

Item 8: Acids, Bases & Buffers

Blood maintains a nearly constant pH despite metabolic reactions producing acidic substances. Medical students study how buffer systems operate.

Questions:
a) Define a buffer solution.
b) Identify the components of a blood buffer system.
c) Explain how the buffer resists pH change when acid is added.
d) Explain why pH control is vital in the body.
e) State one non-biological use of buffers.

Item 9: Electrochemistry & Redox Reactions

A car battery supplies electrical energy through chemical reactions. Over time, the battery weakens and requires recharging.

Questions:
a) Define oxidation and reduction in terms of electrons.
b) Identify the anode and cathode during discharge.
c) Explain how electrical energy is produced.
d) Explain what happens during recharging.
e) State one factor affecting battery efficiency.

Item 10: Electrolysis & Industrial Applications

Molten aluminium oxide is electrolysed to produce aluminium metal. The process requires high temperatures and large amounts of electricity.

Questions:
a) Explain why molten aluminium oxide is used instead of aqueous solution.
b) Write electrode reactions.
c) Explain the role of cryolite.
d) State two energy costs involved.
e) Give one use of aluminium based on its properties.

Item 11: Reaction Kinetics

In an experiment, magnesium reacts with dilute hydrochloric acid at different temperatures. Students observe changes in reaction rate.

Questions:
a) Define rate of reaction.
b) Explain the effect of temperature on rate.
c) Describe the collision theory.
d) Explain the effect of concentration.
e) State one industrial importance of reaction rates.

Item 12: Catalysis & Enzymes

Hydrogen peroxide decomposes slowly at room temperature but rapidly in the presence of manganese(IV) oxide or catalase enzyme.

Questions:
a) Define a catalyst.
b) Compare inorganic catalysts with enzymes.
c) Explain why enzymes are temperature-sensitive.
d) Explain how catalysts affect activation energy.
e) State one use of catalysis in industry.

Item 13: Alkanes & Fuels

Natural gas is increasingly used for cooking in urban areas. Its chemical properties make it a preferred domestic fuel.

Questions:
a) Define an alkane.
b) Write the general formula of alkanes.
c) Explain why alkanes are relatively unreactive.
d) Write an equation for complete combustion.

e) State one advantage and one disadvantage of natural gas.

Item 14: Alkenes & Reaction Mechanisms

Ethene is converted into many useful products through addition reactions. Students follow the mechanism of one such reaction.

Questions:
a) Identify the functional group in ethene.
b) Describe the mechanism of electrophilic addition.
c) Explain why alkenes are more reactive than alkanes.
d) State one industrial use of ethene.
e) Give one safety precaution when handling alkenes.

Item 15: Alcohols & Oxidation

Ethanol is widely used as a fuel, solvent, and antiseptic. Its chemical behavior depends on reaction conditions.

Questions:
a) Identify the functional group in ethanol.
b) Describe the oxidation of ethanol under mild conditions.
c) Describe oxidation under strong conditions.
d) State one test for ethanol.
e) Give one social impact of ethanol misuse.

Item 16: Carboxylic Acids & Esters

Ethanoic acid reacts with ethanol to form a sweet-smelling liquid used in flavourings.

Questions:
a) Name the reaction involved.
b) Write the chemical equation.
c) State conditions required.
d) Explain why the reaction is reversible.
e) State one use of the ester formed.

Item 17: Amines & Basicity

Amines are widely used in pharmaceuticals and dyes. Their basic nature influences their reactions.

Questions:

a) Define an amine.
b) Explain why amines are basic.
c) Compare basicity of ammonia and amines.
d) Describe salt formation with acids.
e) State one health concern related to aromatic amines.

Item 18: Polymers & Materials Science

Synthetic polymers have replaced traditional materials in many applications such as packaging and clothing.

Questions:
a) Define a polymer.
b) Distinguish between addition and condensation polymers.
c) Explain why polymers are lightweight.
d) State one environmental problem caused by plastics.
e) Suggest one solution to plastic pollution.

Item 19: Organic Chemistry & Environment

Incomplete combustion of fuels contributes to urban air pollution. Chemists study ways to reduce harmful emissions.

Questions:
a) Define incomplete combustion.
b) Identify two harmful products formed.
c) Explain their effects on health.
d) Describe how catalytic converters reduce pollution.
e) State one role of chemistry in environmental conservation.

Item 20: Integrated UACE Chemistry

A chemical industry combines stoichiometry, energy changes, reaction rates, equilibria, and organic chemistry to operate efficiently and sustainably.

Questions:
a) Explain the importance of mole calculations.
b) Explain the role of thermochemistry.
c) Explain why equilibrium control is essential.
d) Explain the importance of reaction kinetics.
e) State one ethical responsibility of chemists.

🎉 The End

Congratulations! You have reached the end of this item Bank — a resource created to support Uganda Advanced Certificate of Education students in mastering the new curriculum.

I sincerely hope this book has been helpful in your studies and exam preparation. Every effort has been made to ensure accuracy, clarity, and usefulness. However, should you notice any errors, or if you have suggestions for improvement, please feel free to contact me:

joelamanyire3@gmail.com or Whatsapp number 0788477510

Thank you for using this item bank. Keep studying, stay curious, and continue striving for excellence!

— JoelPCM

10 inspiring quotes specifically selected to motivate chemistry students and keep the fire of curiosity and perseverance burning:

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1. "Chemistry is not just a subject; it's the study of everything around you." – Unknown

2. "Science knows no country, because knowledge belongs to humanity." – Louis Pasteur

3. "In chemistry, as in life, small changes can lead to big reactions." – Unknown

4. "Research is creating new knowledge." – Neil Armstrong

5. "Chemistry is the melodies you can play on the periodic table." – Unknown

6. "Don’t be afraid to experiment; failure is just another reagent for success." – Unknown

7. "The important thing is not to stop questioning. Curiosity has its own reason for existing." – Albert Einstein

8. "Chemistry is the science of matter, but also the science of imagination." – Unknown

9. "Knowledge of chemistry is the key to understanding the world’s problems and solving them." – Unknown

10. "Stay curious. Stay determined. Every reaction brings you closer to discovery." – Unknown

Thank you for your support