Paper 1:

Topic 1 – Atomic Structure and the Periodic Table Retrieval Starters

4.1.1.1 Atoms, Elements and Compounds

1. What is an atom?
2. What is an element?
3. What is a compound?
4. What is the difference between an element and a compound?
5. How can compounds be separated into their elements?
6. Explain why a chemical reaction is needed to separate a compound but not a mixture.

4.1.1.2 Mixtures

1. What is a mixture?
2. Do substances in a mixture keep their own chemical properties?
3. Name four methods used to separate mixtures.
4. Which separation technique is used to separate dyes in ink?
5. Which separation technique is used to separate a soluble solid from a solution?
6. Explain why separating a mixture does not involve a chemical reaction.

4.1.1.3 The Development of the Model of the Atom

1. What model of the atom was accepted before the electron was discovered?
2. What did the plum pudding model suggest about the atom?
3. Which experiment led to the nuclear model of the atom?
4. What did Niels Bohr suggest about electrons?
5. Which scientist discovered the neutron?
6. Explain why Rutherford's alpha scattering experiment caused the atomic model to change.

4.1.1.4 Relative Electrical Charges of Subatomic Particles

1. What is the relative charge of a proton?
2. What is the relative charge of a neutron?

3. What is the relative charge of an electron?
4. What is the atomic number of an element?
5. Why does an atom have no overall charge?
6. Explain why atoms of different elements have different chemical properties.

4.1.1.5 Size and Mass of Atoms

1. What is the approximate radius of an atom?
2. Where is almost all of an atom's mass found?
3. What is the mass number of an atom?
4. What is an isotope?
5. Calculate the number of neutrons in an atom with mass number 35 and atomic number 17.
6. Explain why isotopes of the same element have similar chemical properties.

4.1.1.6 Relative Atomic Mass

1. What is relative atomic mass?
2. Why is relative atomic mass often not a whole number?
3. What does relative atomic mass take into account?
4. Which particles differ in isotopes of the same element?
5. How do you calculate the relative atomic mass from isotope abundance data?
6. Explain why chlorine has a relative atomic mass of 35.5 rather than exactly 35 or 37.

4.1.1.7 Electronic Structure

1. Where are electrons found in an atom?
2. Which shell fills first?
3. State the electronic structure of sodium.
4. State the electronic structure of oxygen.
5. Write the electronic structure of calcium.
6. Explain how electronic structure determines an element's position in the periodic table.

4.1.2.1 The Periodic Table

1. In what order are elements arranged in the periodic table?
2. What is a group?
3. What is a period?
4. Why do elements in the same group have similar chemical properties?
5. How can you predict the reactivity of an element using the periodic table?

6. Explain how an element's position is related to its electron arrangement.

4.1.2.2 Development of the Periodic Table

1. Who developed the first successful periodic table?
2. How did Mendeleev arrange the elements?
3. Why did Mendeleev leave gaps in his table?
4. What happened to the gaps in Mendeleev's table?
5. Why was arranging elements by atomic weight sometimes incorrect?
6. Explain why Mendeleev's periodic table was eventually accepted.

4.1.2.3 Metals and Non-metals

1. Where are metals found in the periodic table?
2. Where are non-metals found in the periodic table?
3. What type of ions do metals form?
4. What type of ions do non-metals form?
5. State two physical properties of metals.
6. Explain why metals and non-metals have different chemical properties.

4.1.2.4 Group 0

1. What are the Group 0 elements called?
2. Why are noble gases unreactive?
3. Which Group 0 element has only two electrons in its outer shell?
4. How does boiling point change down Group 0?
5. Which noble gas has the lowest boiling point?
6. Explain why the boiling points increase down Group 0.

4.1.2.5 Group 1

1. What are the Group 1 elements called?
2. How many electrons do Group 1 elements have in their outer shell?
3. How does reactivity change down Group 1?
4. What gas is produced when Group 1 metals react with water?
5. Name the products formed when sodium reacts with chlorine.
6. Explain why Group 1 metals become more reactive down the group.

4.1.2.6 Group 7

1. What are the Group 7 elements called?
2. How many electrons do Group 7 elements have in their outer shell?
3. How does reactivity change down Group 7?
4. What is a displacement reaction?
5. Which halogen can displace iodine from potassium iodide solution?

6. Explain why chlorine is more reactive than bromine.

4.1.3.1 Comparison with Group 1 Elements (Triple Only)

1. What are the transition elements?
2. Are transition metals more or less reactive than Group 1 metals?
3. Name one physical property in which transition metals differ from Group 1 metals.
4. Name one transition metal used in everyday life.
5. State two differences between transition metals and Group 1 metals.
6. Explain why transition metals are suitable for making structures and tools.

4.1.3.2 Typical Properties (Triple Only)

1. What is a catalyst?
2. Name one typical property of transition metals.
3. What colour are many transition metal compounds?
4. Can transition metals form ions with different charges?
5. Give one example of a transition metal used as a catalyst.
6. Explain why transition metals are widely used in industry.

Topic 1 Review

1. What is the difference between an element and a compound?
2. What particle determines the identity of an element?
3. State the relative charges of the three subatomic particles.
4. Explain why isotopes have different mass numbers.
5. Compare the properties of Group 1 and Group 7 elements.
6. Explain how the arrangement of electrons determines an element's position and reactivity in the periodic table.

Topic 2 – Bonding, Structure and the Properties of Matter Retrieval Starters

4.2.1.1 Chemical Bonds

1. What are the three types of strong chemical bond?
2. Which type of bonding occurs between a metal and a non-metal?
3. Which type of bonding occurs between non-metals?
4. What type of bonding is found in metals?
5. What holds atoms or ions together in chemical bonds?
6. Explain the difference between ionic, covalent and metallic bonding.

4.2.1.2 Ionic Bonding

1. What happens to electrons during ionic bonding?
2. What type of ion is formed when a metal loses electrons?
3. What type of ion is formed when a non-metal gains electrons?
4. Show aluminium oxide bonding using a dot and cross diagram
5. Why do atoms form ions during ionic bonding?
6. Explain how sodium and chlorine form sodium chloride.

4.2.1.3 Ionic Compounds

1. What structure do ionic compounds have?
2. What force holds ions together in an ionic lattice?
3. Why do ionic compounds have high melting points?
4. What is an empirical formula?
5. Why can molten ionic compounds conduct electricity?
6. Explain why solid ionic compounds do not conduct electricity.

4.2.1.4 Covalent Bonding

1. What happens to electrons in covalent bonding?
2. Between which types of elements does covalent bonding occur?
3. Show the bonding in carbon dioxide using a dot and cross diagram.
4. Name one giant covalent substance.
5. Name one substance made of small covalent molecules.
6. Explain why covalent bonds are described as strong bonds.

4.2.1.5 Metallic Bonding

1. What are metals made of?
2. What are delocalised electrons?
3. Show metallic bonding in calcium metal.
4. Why can electrons move through a metal?
5. What is meant by metallic bonding?
6. Explain why metals are good electrical conductors.

4.2.2.1 The Three States of Matter

1. What are the three states of matter?
2. What happens during melting?
3. What happens during condensation?
4. What happens to particles when a substance is heated?
5. Why do different substances have different melting points?

6. Explain why more energy is needed to melt some substances than others.

4.2.2.2 State Symbols

1. What state symbol represents a solid?
2. What state symbol represents a liquid?
3. What state symbol represents a gas?
4. What does (aq) mean?
5. Why are state symbols included in chemical equations?
6. Write the equation of hydrochloric acid reacting with solid magnesium including correct state symbols for both reactants and products.

4.2.2.3 Properties of Ionic Compounds

1. Why do ionic compounds have high melting points?
2. Why do ionic compounds have high boiling points?
3. When do ionic compounds conduct electricity?
4. Why do molten ionic compounds conduct electricity?
5. Why do solid ionic compounds not conduct electricity?
6. Explain how the structure of ionic compounds affects their properties.

4.2.2.4 Properties of Small Molecules

1. What type of forces exist between small molecules?
2. Are intermolecular forces stronger or weaker than covalent bonds?
3. Why do small molecular substances have low melting points?
4. Do small molecular substances conduct electricity?
5. How does molecule size affect boiling point?
6. Explain why covalent bonds are not broken when a molecular substance melts.

4.2.2.5 Polymers

1. What is a polymer?
2. What type of bonding joins atoms within a polymer?
3. Why are polymers solids at room temperature?
4. What type of molecules do polymers contain?
5. Give one example of a polymer.
6. Explain why polymers generally have higher melting points than small molecular substances.

4.2.2.6 Giant Covalent Structures

1. What is a giant covalent structure?
2. Name three giant covalent substances.
3. Why do giant covalent substances have very high melting points?
4. What type of bond must be broken to melt a giant covalent structure?
5. Which giant covalent substance conducts electricity?
6. Explain why diamond and graphite have different properties.

4.2.2.7 Properties of Metals and Alloys

1. Why do metals have high melting points?
2. Why can pure metals be bent easily?
3. What is an alloy?
4. Why are alloys harder than pure metals?
5. Give one example of an alloy.
6. Explain why alloys are often used instead of pure metals.

4.2.2.8 Metals as Conductors

1. Why do metals conduct electricity?
2. What particles carry electrical charge through a metal?
3. Why do metals conduct thermal energy well?
4. What are delocalised electrons free to do?
5. Name one property that makes metals useful for electrical wiring.
6. Explain how metallic bonding allows metals to conduct heat and electricity.

4.2.3.1 Diamond

1. How many covalent bonds does each carbon atom form in diamond?
2. What type of structure does diamond have?
3. Why is diamond very hard?
4. Why does diamond have a very high melting point?
5. Does diamond conduct electricity?
6. Explain why diamond does not conduct electricity.

4.2.3.2 Graphite

1. How many covalent bonds does each carbon atom form in graphite?
2. What shape are the layers in graphite?
3. Why can graphite conduct electricity?
4. Why can the layers in graphite slide over each other?
5. State one use of graphite.
6. Explain why graphite is soft but diamond is hard.

4.2.3.3 Graphene and Fullerenes

1. What is graphene?
2. What is a fullerene?
3. What shape is Buckminsterfullerene (C_{60})?
4. What are carbon nanotubes?
5. Give one use of graphene or carbon nanotubes.
6. Explain why graphene is useful in electronics.

4.2.4.1 Sizes of Particles and Their Properties (Triple Only)

1. What size range do nanoparticles have?
2. What happens to surface area to volume ratio as particles get smaller?
3. Why do nanoparticles have different properties from bulk materials?
4. What does the prefix nano- mean?
5. Why are smaller quantities of nanoparticles often needed?
6. Explain why nanoparticles are more reactive than larger particles.

4.2.4.2 Uses of Nanoparticles (Triple Only)

1. Give one use of nanoparticles in medicine.
2. Give one use of nanoparticles in cosmetics.
3. Give one industrial use of nanoparticles.
4. State one advantage of using nanoparticles.
5. State one possible risk of using nanoparticles.
6. Explain why scientists continue to research nanoparticle applications.

Topic 2 Review

1. Name the three types of strong chemical bonding.
2. What is the difference between ionic and covalent bonding?
3. Explain why ionic compounds only conduct electricity when molten or dissolved.
4. Compare the structures of diamond and graphite.
5. Explain why alloys are harder than pure metals.
6. Describe how bonding and structure determine the physical properties of substances.

Topic 3 – Quantitative Chemistry Retrieval Starters

4.3.1.1 Conservation of Mass and Balanced Chemical Equations

1. What does the law of conservation of mass state?
2. Why must chemical equations be balanced?
3. What does a coefficient in a chemical equation show?
4. What is the difference between a coefficient and a subscript?
5. Balance the equation: $H_2 + O_2 \rightarrow H_2O$.
6. Explain why mass stays the same during a chemical reaction.

4.3.1.2 Relative Formula Mass

1. What does relative formula mass (Mr) mean?
2. How is relative formula mass calculated?
3. What information is needed to calculate Mr?
4. Calculate the Mr of H_2O.
5. Calculate the percentage by mass of oxygen in water.
6. Explain why relative formula mass is useful in chemistry.

4.3.1.3 Mass Changes When a Reactant or Product is a Gas

1. Why can the mass increase when magnesium burns?
2. Why can the mass decrease when a metal carbonate is heated?
3. What gas is produced during the thermal decomposition of a metal carbonate?
4. What happens to the mass in a closed system?
5. Why do some reactions appear to lose mass?
6. Explain why mass is conserved even when gases are produced.

4.3.1.4 Chemical Measurements

1. What is meant by uncertainty in a measurement?
2. What is meant by the mean of a set of results?
3. What is the range of a set of results?
4. Why are repeat measurements carried out?
5. How can the uncertainty of repeated measurements be estimated?
6. Explain why measurements in science always have some uncertainty.

4.3.2.1 Moles (Higher Tier)

1. What is the unit used to measure the amount of substance?

2. What is one mole?
3. What is the Avogadro constant?
4. Write the equation linking moles, mass and relative formula mass.
5. Calculate the number of moles in 36 g of water.
6. Calculate the mass of 0.25 moles of sodium chloride ($NaCl$).

4.3.2.2 Amounts of Substances in Equations (Higher Tier)

1. What does a balanced equation show about moles?
2. Why must equations be balanced before calculations are carried out?
3. How many moles of hydrogen are produced when one mole of magnesium reacts with hydrochloric acid?
4. What is the first step in a mole calculation?
5. Calculate the mass of magnesium oxide produced from 24 g of magnesium.
6. Calculate the mass of aluminium oxide produced when 135g of aluminium is burned in air.

4.3.2.3 Using Moles to Balance Equations (Higher Tier)

1. Why are moles used to balance chemical equations?
2. What must masses be converted into before balancing by calculation?
3. What type of ratio is used to balance equations?
4. Balance the chemical equation for photosynthesis using the following masses: 264 g carbon dioxide reacts with 108 g water to produce 180 g glucose and 192 g oxygen.
5. Balance the chemical equation for the complete combustion of methane using the following masses: 16 g methane reacts with 64 g oxygen to produce 44 g carbon dioxide and 36 g water.
6. Balance a chemical equation between magnesium and oxygen using the following masses: 24 g magnesium reacts with 16 g oxygen to produce 40 g magnesium oxide.

4.3.2.4 Limiting Reactants (Higher Tier)

1. What is a limiting reactant?
2. Why is one reactant often used in excess?
3. What happens when the limiting reactant is completely used up?
4. Which reactant determines the amount of product formed?
5. Identify the limiting reactant from given amounts.
6. Explain how the limiting reactant affects the yield of a reaction.

4.3.2.5 Concentration of Solutions

1. What is concentration?
2. What units are used for concentration in g/dm^3?

3. Write the equation linking concentration, mass and volume.
4. Calculate the concentration of a solution containing 20 g of solute in 2 dm^3.
5. Calculate the mass of solute in 0.5 dm^3 of a 10 g/dm^3 solution.
6. Explain how increasing the amount of solute affects concentration.

4.3.3.1 Percentage Yield (Triple Only)

1. What is meant by percentage yield?
2. Write the equation for percentage yield.
3. Give one reason why the percentage yield may be less than 100%.
4. What is the theoretical yield?
5. Calculate the percentage yield if the actual yield is 8 g and the theoretical yield is 10 g.
6. Explain why reactions rarely produce a 100% yield.

4.3.3.2 Atom Economy (Triple Only)

1. What is atom economy?
2. Why is a high atom economy desirable?
3. Write the equation for atom economy.
4. Calculate the atom economy for the production of iron from iron(III) oxide using the balanced equation: $2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$
5. Calculate the atom economy for the production of magnesium oxide using the balanced equation: $2Mg + O_2 \rightarrow 2MgO$
6. Explain why reactions with high atom economy are more sustainable.

4.3.4 Using Concentrations of Solutions in mol/dm^3 (Triple Higher Tier)

1. What units are used for concentration in moles?
2. Write the equation linking concentration, moles and volume.
3. What unit must volume be in when using the concentration equation?
4. Calculate the number of moles in 0.5 dm^3 of a 2 mol/dm^3 solution.
5. Calculate the concentration of a solution containing 1 mole in 0.25 dm^3.
6. Explain how titration can be used to determine concentration.

4.3.5 Use of Amount of Substance in Relation to Volumes of Gases (Triple Higher Tier)

1. What volume does one mole of gas occupy at room temperature and pressure?
2. What conditions are assumed when using 24 dm^3 per mole?
3. Write the equation linking gas volume and moles.
4. Calculate the volume occupied by 2 moles of gas at room temperature.

5. Calculate the number of moles in 48 dm^3 of gas.
6. Explain why equal numbers of moles of different gases occupy equal volumes under the same conditions.

Topic 3 Review

1. State the law of conservation of mass.
2. Write the equation used to calculate moles.
3. Explain why balanced equations are needed in quantitative chemistry.
4. State two reasons why percentage yield may be less than 100%.
5. Explain the difference between percentage yield and atom economy.
6. Describe how moles are used to calculate reacting masses and gas volumes.

Topic 4 – Chemical Changes Retrieval Starters

4.4.1.1 Metal Oxides

1. What is produced when a metal reacts with oxygen?
2. What is oxidation in terms of oxygen?
3. What is reduction in terms of oxygen?
4. Why are reactions between metals and oxygen called oxidation reactions?
5. Which substance is reduced when copper oxide is heated with carbon?
6. Explain why oxidation and reduction occur together in reactions involving metal oxides.

4.4.1.2 The Reactivity Series

1. What is the reactivity series?
2. Which metal is more reactive: magnesium or copper?
3. What happens when a more reactive metal is added to a compound containing a less reactive metal?
4. Which non-metals are included in the reactivity series?
5. Put the following metals in order of reactivity from most to least reactive: iron, potassium, zinc, copper.
6. Explain why potassium reacts more vigorously with water than magnesium.

4.4.1.3 Extraction of Metals and Reduction

1. How are unreactive metals such as gold found in the Earth's crust?
2. Which metals can be extracted from their oxides using carbon?

3. What happens to oxygen during the reduction of a metal oxide?
4. Why is carbon used to extract some metals from their oxides?
5. Why cannot aluminium be extracted from aluminium oxide using carbon?
6. Explain why the method used to extract a metal depends on its position in the reactivity series.

4.4.1.4 Oxidation and Reduction in Terms of Electrons (Higher Tier)

1. What is oxidation in terms of electrons?
2. What is reduction in terms of electrons?
3. Which species is oxidised in the reaction: $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$?
4. Which species is reduced in the reaction: $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$?
5. Write the ionic equation for the reaction between magnesium and copper(II) sulfate solution.
6. Explain why displacement reactions are examples of redox reactions.

4.4.2.1 Reactions of Acids with Metals

1. What gas is produced when an acid reacts with a metal?
2. What type of salt is produced when magnesium reacts with hydrochloric acid?
3. What is the general word equation of a metal reacting with acid?
4. Write the word equation for the reaction between zinc and sulfuric acid.
5. Which metal reacts more vigorously with dilute hydrochloric acid: magnesium or iron?
6. Explain why reactions between acids and metals are redox reactions.

4.4.2.2 Neutralisation of Acids and Salt Production

1. What is produced when an acid reacts with an alkali?
2. What is produced when an acid reacts with a metal carbonate?
3. Which acid is used to produce sulfate salts?
4. Which acid is used to produce nitrate salts?
5. What salt is produced when nitric acid reacts with potassium hydroxide?
6. Explain how the acid used determines the name of the salt produced.

4.4.2.3 Soluble Salts (Required Practical 1)

1. Why is excess insoluble solid added when making a soluble salt?
2. Why is the excess solid filtered from the solution when making a soluble salt?
3. Why is the salt solution heated after filtration when making a soluble salt?
4. When making a soluble salt, why should the salt solution not be heated to complete dryness?

5. Put the following steps in order: filtration, crystallisation, add excess solid, warm the acid.
6. Explain the required practical of producing a pure, dry sample of a soluble salt.

4.4.2.4 The pH Scale and Neutralisation

1. What is the pH of a neutral solution?
2. Which ions make a solution acidic?
3. Which ions make a solution alkaline?
4. What is produced when hydrogen ions react with hydroxide ions?
5. Which piece of equipment can be used to measure the approximate pH of a solution?
6. Explain why hydrochloric acid and sodium hydroxide produce a neutral solution when they react in the correct proportions.

4.4.2.5 Titrations (Required Practical 2 - Chemistry Only)

1. What is the purpose of a titration?
2. Which piece of apparatus is used to deliver the acid accurately during a titration?
3. What is the end point of a titration?
4. Why is a suitable indicator added during a titration?
5. Calculate the mean titre for the following concordant results: 24.60 cm^3, 24.70 cm^3 and 24.65 cm^3.
6. Explain why concordant titres are used when calculating the mean titre.

4.4.2.6 Strong and Weak Acids (Higher Tier)

1. What is a strong acid?
2. What is a weak acid?
3. Which has the lower pH at the same concentration: a strong acid or a weak acid?
4. What is the difference between a concentrated acid and a dilute acid?
5. By what factor does the hydrogen ion concentration increase when pH decreases from 5 to 3?
6. Explain why a dilute hydrochloric acid can be a stronger acid than a concentrated ethanoic acid.

4.4.3.1 The Process of Electrolysis

1. What is an electrolyte?
2. Which electrode is the cathode?
3. Which ions move towards the cathode during electrolysis?
4. Which ions move towards the anode during electrolysis?

5. Why must an ionic compound be molten or dissolved before it can be electrolysed?
6. Explain why electrolysis causes elements to be produced at the electrodes.

4.4.3.2 Electrolysis of Molten Ionic Compounds

1. What is produced at the cathode during the electrolysis of molten lead bromide?
2. What is produced at the anode during the electrolysis of molten lead bromide?
3. Why are inert electrodes used during electrolysis?
4. Predict the products formed during the electrolysis of molten zinc chloride.
5. Write the formula of the substance produced at the anode during the electrolysis of molten sodium chloride.
6. Explain why molten ionic compounds conduct electricity.

4.4.3.3 Using Electrolysis to Extract Metals

1. Which metals are extracted using electrolysis?
2. Why is aluminium extracted by electrolysis instead of using carbon?
3. Why is cryolite mixed with aluminium oxide during electrolysis?
4. Why must the carbon anodes be replaced regularly during aluminium extraction?
5. Which electrode produces aluminium during electrolysis?
6. Explain why extracting aluminium by electrolysis is expensive.

4.4.3.4 Electrolysis of Aqueous Solutions (Required Practical 3)

1. What is produced at the cathode when aqueous copper(II) sulfate is electrolysed using inert electrodes?
2. What is produced at the anode when aqueous sodium chloride is electrolysed using inert electrodes?
3. Which gas is produced at the cathode if the metal is more reactive than hydrogen?
4. Which gas is usually produced at the anode if no halide ions are present?
5. Predict the products formed during the electrolysis of aqueous potassium chloride using inert electrodes.
6. Explain how the reactivity series helps predict the products of electrolysis in aqueous solutions.

4.4.3.5 Representation of Reactions at Electrodes as Half Equations (Higher Tier)

1. What happens to positive ions at the cathode during electrolysis?
2. What happens to negative ions at the anode during electrolysis?

3. Write the half equation for the formation of hydrogen from hydrogen ions.
4. Write the half equation for the formation of chlorine from chloride ions.
5. Write the half equation for the formation of copper from Cu^{2+} ions.
6. Explain why reduction occurs at the cathode and oxidation occurs at the anode.

Topic 4 Review

1. What is the difference between oxidation and reduction in terms of oxygen?
2. Explain why metals above carbon in the reactivity series cannot be extracted using carbon.
3. State the products formed when an acid reacts with a metal carbonate.
4. Describe how a pure, dry sample of a soluble salt is prepared from an insoluble base.
5. Explain why molten ionic compounds conduct electricity but solid ionic compounds do not.
6. Compare the products formed during the electrolysis of molten sodium chloride and aqueous sodium chloride.

Topic 5 – Energy Changes Retrieval Starters

4.5.1.1 Energy Transfer During Exothermic and Endothermic Reactions

1. What is an exothermic reaction?
2. What is an endothermic reaction?
3. Name one example of an exothermic reaction.
4. Name one example of an endothermic reaction.
5. What happens to the temperature of the surroundings during an endothermic reaction?
6. Explain why the products of an exothermic reaction have less energy than the reactants.

Required Practical 4: Investigating Temperature Changes

1. What is measured during the required practical investigating temperature changes in reacting solutions?

2. Why is a lid placed on the reaction cup during the required practical investigating temperature changes?
3. Why should the reactants be mixed quickly during the required practical investigating temperature changes?
4. Why should the same volumes of reactants be used when comparing temperature changes in different reactions?
5. Why are repeat measurements carried out during the required practical investigating temperature changes?
6. Explain why reducing heat loss improves the accuracy of the required practical investigating temperature changes.

4.5.1.2 Reaction Profiles

1. What is activation energy?
2. What does a reaction profile show?
3. Draw a labelled reaction profile for the endothermic reaction A + B → C + D, showing the reactants, products, activation energy and overall energy change.
4. Draw a labelled reaction profile for the exothermic reaction A + B → C + D, showing the reactants, products, activation energy and overall energy change.
5. What does the highest point on a reaction profile represent?
6. Explain why all chemical reactions require activation energy.

4.5.1.3 The Energy Change of Reactions (Higher Tier)

1. What happens to energy when chemical bonds are broken?
2. What happens to energy when new chemical bonds are formed?
3. In an exothermic reaction, is more energy released from bond formation or needed to break bonds?
4. In an endothermic reaction, is more energy needed to break bonds or released from bond formation?
5. Calculate the overall energy change for a reaction if 820 kJ is needed to break bonds and 950 kJ is released when new bonds are formed.
6. Explain why bond breaking and bond making occur together during a chemical reaction.

4.5.2.1 Cells and Batteries (Chemistry Only)

1. What produces electricity in a chemical cell?
2. How can a simple chemical cell be made?
3. What is the difference between a cell and a battery?
4. Why do non-rechargeable batteries eventually stop working?
5. Why can rechargeable batteries be used many times?
6. Explain why connecting cells in series increases the voltage of a battery.

4.5.2.2 Fuel Cells (Chemistry Only)

1. What fuel is used in a hydrogen fuel cell?
2. What is the overall product formed in a hydrogen fuel cell?
3. Where does the oxygen used in a hydrogen fuel cell come from?
4. How is hydrogen converted into electrical energy in a hydrogen fuel cell?
5. State one advantage of hydrogen fuel cells compared with rechargeable batteries.
6. Explain one disadvantage of using hydrogen fuel cells compared with rechargeable batteries.

Topic 5 Review

1. What is the difference between an exothermic reaction and an endothermic reaction?
2. Explain why combustion reactions are exothermic.
3. Describe how a reaction profile for an endothermic reaction differs from a reaction profile for an exothermic reaction.
4. Explain why energy is required to break chemical bonds.
5. Compare rechargeable batteries with hydrogen fuel cells.
6. Explain how the breaking and formation of chemical bonds determine whether a reaction is exothermic or endothermic.

Paper 2:

Topic 6 – The Rate and Extent of Chemical Change Retrieval Starters

4.6.1.1 Calculating Rates of Reactions

1. What is meant by the rate of a chemical reaction?
2. How is the mean rate of reaction calculated using the quantity of reactant used?
3. How is the mean rate of reaction calculated using the quantity of product formed?
4. What units can be used for the rate of reaction when measuring mass change?
5. How can a graph showing product formed against time be used to determine the rate of reaction?

6. How can the gradient of a tangent to a reaction graph be used to calculate the rate of reaction at a specific time?

4.6.1.2 Factors Which Affect the Rates of Chemical Reactions

1. What five factors can affect the rate of a chemical reaction?
2. How does increasing the concentration of reactants in solution affect the rate of reaction?
3. How does decreasing the pressure of reacting gases affect the rate of reaction?
4. How does increasing the surface area of a solid reactant affect the rate of reaction?
5. How does increasing the temperature affect the rate of a chemical reaction?
6. How does the presence of a catalyst affect the rate of a chemical reaction?

Required Practical 5: Investigating How Changes in Concentration Affect the Rates of Reactions

1. What is measured when investigating how changes in concentration affect the rate of reaction using gas production?
2. Why can the volume of gas produced be used to measure the rate of reaction?
3. What variable is changed when investigating the effect of concentration on reaction rate?
4. Why should only one variable be changed when investigating the effect of concentration on reaction rate?
5. Why are repeat measurements carried out when investigating the effect of concentration on reaction rate?
6. Why should a hypothesis be made before carrying out an investigation into reaction rates?

4.6.1.3 Collision Theory and Activation Energy

1. What does collision theory state about chemical reactions?
2. Why must reacting particles collide for a chemical reaction to occur?
3. What is activation energy?
4. How does increasing the concentration of reactants increase the rate of reaction according to collision theory?
5. How does increasing temperature increase the rate of reaction according to collision theory?

6. Why does increasing the surface area of a solid reactant increase the rate of reaction?

4.6.1.4 Catalysts

1. What is a catalyst?
2. Why are catalysts not used up during a chemical reaction?
3. How do catalysts increase the rate of a chemical reaction?
4. What happens to the activation energy when a catalyst is used?
5. Why do different chemical reactions require different catalysts?
6. How can catalysts be identified from a chemical equation and reaction rate?

4.6.2.1 Reversible Reactions

1. What is a reversible reaction?
2. How are reversible reactions represented using a symbol equation?
3. What happens in a reversible reaction when products react to form reactants?
4. How can the direction of a reversible reaction be changed?
5. What is meant by the forward reaction in a reversible reaction?
6. What is meant by the reverse reaction in a reversible reaction?

4.6.2.2 Energy Changes and Reversible Reactions

1. What type of reaction is the reverse of an exothermic reaction?
2. What happens to energy when an exothermic reaction is reversed?
3. How much energy is transferred in each direction of a reversible reaction?
4. Why is the forward and reverse reaction in a reversible reaction linked by energy changes?
5. What type of energy change occurs when an endothermic reaction is reversed?
6. Explain why a reversible reaction can be both exothermic and endothermic.

4.6.2.3 Equilibrium

1. What is meant by equilibrium in a reversible reaction?
2. What conditions are needed for equilibrium to be reached?
3. Why must a reversible reaction be carried out in a closed system to reach equilibrium?

4. What happens to the forward reaction rate and reverse reaction rate at equilibrium?
5. Why do the concentrations of reactants and products remain constant at equilibrium?
6. Explain why equilibrium is described as a dynamic process.

4.6.2.4 The Effect of Changing Conditions on Equilibrium (HT Only)

1. What does Le Chatelier's Principle state?
2. What happens to an equilibrium system when a change is made to its conditions?
3. Why does a system at equilibrium respond when a condition is changed?
4. What factors can be changed to affect the position of equilibrium?
5. How can Le Chatelier's Principle be used to predict changes in equilibrium?
6. Why does an equilibrium system shift to counteract a change?

4.6.2.5 The Effect of Changing Concentration (HT Only)

1. What happens when the concentration of a reactant is increased in an equilibrium system?
2. Why does increasing the concentration of a reactant produce more products?
3. What happens when the concentration of a product is decreased in an equilibrium system?
4. Why do concentrations of all substances change after the concentration of one substance is altered?
5. How can changing concentration affect the position of equilibrium?
6. For the equilibrium reaction $N_2(g) + 3H_2(g) \leftrightharpoons 2NH_3(g)$, explain what happens to the position of equilibrium when the concentration of hydrogen is increased.

4.6.2.6 The Effect of Temperature Changes on Equilibrium (HT Only)

1. What happens to the amount of products at equilibrium when the temperature is increased for an endothermic reaction?
2. What happens to the amount of products at equilibrium when the temperature is increased for an exothermic reaction?

3. What happens to the amount of products at equilibrium when the temperature is decreased for an endothermic reaction?
4. What happens to the amount of products at equilibrium when the temperature is decreased for an exothermic reaction?
5. Why does increasing temperature affect the position of equilibrium?
6. For the (exothermic) equilibrium reaction $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$, predict how increasing the temperature affects the equilibrium position and explain why.

4.6.2.7 The Effect of Pressure Changes on Equilibrium (HT Only)

1. How does increasing pressure affect the equilibrium position of a gaseous reaction?
2. Why does increasing pressure shift equilibrium towards the side with fewer gas molecules?
3. How does decreasing pressure affect the equilibrium position of a gaseous reaction?
4. Why does decreasing pressure shift equilibrium towards the side with more gas molecules?
5. How can the symbol equation be used to predict the effect of pressure changes on equilibrium?
6. For the equilibrium reaction $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$, predict how increasing the pressure affects the equilibrium position and explain why.

Topic 6 Review

1. How is the mean rate of reaction calculated?
2. Explain how concentration, pressure, surface area and temperature affect the rate of a chemical reaction.
3. Explain how collision theory explains the effect of temperature on reaction rate.
4. Explain how catalysts increase the rate of reaction without being used up.
5. What is the difference between a reversible reaction and a reaction that goes to completion?
6. Explain how changing concentration, temperature and pressure affects the position of equilibrium.

Topic 7 – Organic Chemistry Retrieval Starters

4.7.1.1 Crude Oil, Hydrocarbons and Alkanes

1. What is crude oil?
2. Why is crude oil described as a finite resource?
3. What are hydrocarbons?
4. What is the general formula for alkanes?
5. What are the names of the first four alkanes?
6. A hydrocarbon has the formula C_5H_{12}. Explain why this compound is an alkane.

4.7.1.2 Fractional Distillation and Petrochemicals

1. What is fractional distillation used to separate?
2. How does fractional distillation separate the hydrocarbons in crude oil?
3. Why do different fractions condense at different temperatures during fractional distillation?
4. What are the main uses of fractions obtained from crude oil?
5. What is meant by a petrochemical?
6. Explain why crude oil is an important feedstock for the petrochemical industry.

4.7.1.3 Properties of Hydrocarbons

1. How does boiling point change as the size of hydrocarbon molecules increases?
2. How does viscosity change as the size of hydrocarbon molecules increases?
3. How does flammability change as the size of hydrocarbon molecules increases?
4. Why are hydrocarbons used as fuels?
5. What are the products of the complete combustion of a hydrocarbon?
6. Write the balanced symbol equation for the complete combustion of methane (CH_4).

4.7.1.4 Cracking and Alkenes

1. What is cracking?
2. Why are hydrocarbons cracked?
3. What conditions are used for catalytic cracking?
4. What type of hydrocarbon is produced alongside alkanes during cracking?
5. What colour change occurs when bromine water reacts with an alkene?
6. Balance the equation for the cracking of $C_{10}H_{22}$ to produce C_8H_{18} and C_2H_4.

4.7.2.1 Structure and Formulae of Alkenes

1. What is the functional group of alkenes?
2. What is the general formula for alkenes?
3. Why are alkenes described as unsaturated hydrocarbons?
4. What are the names of the first four alkenes?
5. How can an alkene be identified from its formula?
6. A compound has the formula C_4H_8. Explain why this compound could be an alkene.

4.7.2.2 Reactions of Alkenes

1. Why are alkenes more reactive than alkanes?
2. What happens when an alkene reacts with hydrogen?
3. What conditions are needed for hydrogen to react with an alkene?
4. What happens when an alkene reacts with a halogen such as bromine?
5. What happens to the carbon-carbon double bond during addition reactions?
6. Complete the equation for the reaction between ethene and hydrogen: $C_2H_4 + H_2 \rightarrow ?$

4.7.2.3 Alcohols

1. What is the functional group of alcohols?
2. What are the names of the first four alcohols?
3. What happens when ethanol burns in oxygen?
4. What happens when alcohols react with sodium?
5. What conditions are needed for the fermentation of sugar to produce ethanol?
6. Write the balanced symbol equation for the complete combustion of ethanol (C_2H_5OH).

4.7.2.4 Carboxylic Acids

1. What is the functional group of carboxylic acids?
2. What are the names of the first four carboxylic acids?
3. What happens when carboxylic acids react with carbonates?
4. Why do carboxylic acids dissolve in water?
5. What type of reaction occurs when a carboxylic acid reacts with an alcohol?
6. Explain why carboxylic acids are weak acids in terms of ionisation. (HT only)

4.7.3.1 Addition Polymerisation

1. What is a polymer?
2. What is a monomer?
3. What type of monomers are used in addition polymerisation?
4. What happens to the carbon-carbon double bond during addition polymerisation?
5. Why does an addition polymer have the same atoms as its monomer?
6. Draw the repeating unit formed from the monomer ethene.

4.7.3.2 Condensation Polymerisation (HT Only)

1. What type of monomers are needed for condensation polymerisation?
2. Why are condensation polymerisation reactions different from addition polymerisation reactions?
3. What small molecule is usually produced during condensation polymerisation?
4. Draw the displayed formula equation showing the condensation polymerisation reaction between propanediol and butanedioic acid, including the monomers, the ester link formed, and the repeating unit of the polyester produced.
5. How are polyesters formed from two different monomers?
6. Explain how ethanediol and hexanedioic acid form a polyester by condensation polymerisation.

4.7.3.3 Amino Acids (HT Only)

1. What two functional groups are present in amino acids?
2. What type of polymer is formed when amino acids join together?
3. What type of reaction joins amino acids together?

4. What small molecule is produced when amino acids polymerise?
5. What is the formula of glycine?
6. Explain how different amino acids can combine to form different proteins.

4.7.3.4 DNA and Other Naturally Occurring Polymers

1. What is DNA?
2. What are the monomers that make up DNA?
3. What shape is the structure of most DNA molecules?
4. What information does DNA encode?
5. What are the monomers that make up proteins?
6. Name the monomers from which starch and cellulose are made.

Topic 7 Review

1. What is the difference between alkanes and alkenes?
2. Explain how fractional distillation separates crude oil into useful fractions.
3. Write the balanced symbol equation for the complete combustion of a hydrocarbon.
4. Explain why cracking is used to produce more useful hydrocarbons.
5. Compare addition polymerisation with condensation polymerisation.
6. Explain how carbon atoms can form a wide range of organic compounds.

Topic 8 – Chemical Analysis Retrieval Starters

4.8.1.1 Pure Substances

1. What is meant by a pure substance in chemistry?
2. What is the difference between a pure substance and a mixture?
3. Why do pure substances have specific melting points and boiling points?
4. How can melting point data be used to identify an impure substance?
5. How can boiling point data be used to distinguish pure substances from mixtures?
6. A substance melts over a range of temperatures rather than at one fixed temperature. Explain what this suggests about the substance.

4.8.1.2 Formulations

1. What is a formulation?
2. Why are formulations made by mixing components in carefully measured quantities?
3. Name three examples of products that are formulations.
4. Why does each chemical in a formulation have a specific purpose?
5. How can a formulation be identified from information about its components?
6. Explain why a paint is considered a formulation rather than a pure substance.

4.8.1.3 Chromatography

1. What is chromatography used to separate?
2. What are the stationary phase and mobile phase in paper chromatography?
3. How does paper chromatography separate substances in a mixture?
4. How is the Rf value of a substance calculated?
5. A spot moves 6 cm and the solvent moves 12 cm. Calculate the Rf value of the substance.
6. How can chromatography be used to distinguish between a pure substance and an impure substance?

Required Practical 6: Investigating Paper Chromatography

1. What is the aim of the required practical investigating paper chromatography?
2. Why must the solvent level be below the spots on the chromatography paper?
3. Why is pencil used to draw the baseline in paper chromatography?
4. Why are different solvents used in chromatography experiments?
5. How can Rf values be used to identify unknown substances?
6. Explain why a pure compound produces one spot on a chromatogram.

4.8.2.1 Test for Hydrogen

1. What test is used to identify hydrogen gas?
2. What happens when a burning splint is placed into hydrogen gas?
3. What sound is produced when hydrogen burns?
4. Where is the burning splint placed during the hydrogen test?
5. What observation confirms that a gas is hydrogen?
6. Explain why hydrogen produces a pop sound during the test.

4.8.2.2 Test for Oxygen

1. What test is used to identify oxygen gas?
2. What happens when a glowing splint is inserted into oxygen gas?
3. What observation confirms that a gas contains oxygen?
4. Why does oxygen relight a glowing splint?
5. Which type of splint is used to test for oxygen?
6. Explain why oxygen supports combustion.

4.8.2.3 Test for Carbon Dioxide

1. What solution is used to test for carbon dioxide?
2. What happens when carbon dioxide is bubbled through limewater?
3. What observation confirms that a gas contains carbon dioxide?
4. What is the chemical name of limewater?
5. Why does limewater turn cloudy when carbon dioxide is present?
6. Write the word equation for the reaction between carbon dioxide and calcium hydroxide.

4.8.2.4 Test for Chlorine

1. What test is used to identify chlorine gas?
2. What happens when damp litmus paper is placed into chlorine gas?
3. What colour does bleached litmus paper become after reacting with chlorine?
4. Why must litmus paper be damp when testing for chlorine?
5. What observation confirms that a gas contains chlorine?
6. Explain why chlorine can be used as a bleaching agent.

4.8.3.1 Flame Tests

1. What are flame tests used to identify?
2. What flame colour is produced by lithium ions?

3. What flame colour is produced by sodium ions?
4. What flame colour is produced by potassium ions?
5. What flame colour is produced by copper ions?
6. A metal compound produces an orange-red flame during a flame test. Identify the metal ion present.

4.8.3.2 Metal Hydroxides

1. Why is sodium hydroxide solution added to solutions containing metal ions?
2. What colour precipitate is formed when copper(II) ions react with sodium hydroxide?
3. What colour precipitate is formed when iron(II) ions react with sodium hydroxide?
4. What colour precipitate is formed when iron(III) ions react with sodium hydroxide?
5. Which metal hydroxide precipitate dissolves in excess sodium hydroxide solution?
6. Write the balanced symbol equation for the formation of magnesium hydroxide from magnesium ions and hydroxide ions.

4.8.3.3 Carbonates

1. What gas is produced when carbonates react with dilute acids?
2. How can carbon dioxide produced from a carbonate be identified?
3. What happens when carbon dioxide is passed through limewater?
4. Which type of chemical test can identify carbonate ions?
5. Write the word equation for the reaction between a carbonate and a dilute acid.
6. Explain why carbonate ions can be identified by producing carbon dioxide gas.

4.8.3.4 Halides

1. What solution is used to test for halide ions?
2. Why is dilute nitric acid added before testing for halide ions?
3. What colour precipitate is formed when chloride ions react with silver nitrate?
4. What colour precipitate is formed when bromide ions react with silver nitrate?
5. What colour precipitate is formed when iodide ions react with silver nitrate?

6. A solution forms a cream precipitate when silver nitrate is added. Identify the halide ion present.

4.8.3.5 Sulfates

1. What solution is used to test for sulfate ions?
2. Why is dilute hydrochloric acid added when testing for sulfate ions?
3. What colour precipitate forms when sulfate ions react with barium chloride?
4. What observation confirms that sulfate ions are present?
5. Write the word equation for the reaction between sulfate ions and barium ions.
6. Explain why a white precipitate forms when sulfate ions are tested with barium chloride.

Required Practical 7: Identifying Ions in Unknown Ionic Compounds

1. What is the aim of the required practical identifying ions in unknown compounds?
2. Which tests can be used to identify metal ions?
3. Which tests can be used to identify negative ions such as halides and sulfates?
4. Why are several chemical tests needed to identify an unknown ionic compound?
5. How can observations from chemical tests be used to identify unknown ions?
6. Explain why using more than one test increases confidence when identifying an unknown ionic compound.

4.8.3.6 Instrumental Methods

1. What are instrumental methods used to identify?
2. Why are instrumental methods useful when only a small amount of chemical is available?
3. State three advantages of instrumental methods compared with chemical tests.
4. Why are instrumental methods used by forensic scientists?
5. How are instrumental methods different from simple chemical tests?

6. Explain why instrumental methods are considered more accurate than many chemical tests.

4.8.3.7 Flame Emission Spectroscopy

1. What is flame emission spectroscopy used to analyse?
2. How does flame emission spectroscopy identify metal ions?
3. What type of spectrum is produced in flame emission spectroscopy?
4. How can the concentration of metal ions be measured using flame emission spectroscopy?
5. Why must a reference set of spectra be used when interpreting flame emission spectroscopy results?
6. A sample produces the same line spectrum as a known lithium sample. Explain what this shows about the sample.

Topic 8 Review

1. Explain how melting point and boiling point data can be used to identify pure substances.
2. Explain how paper chromatography separates mixtures and identifies substances.
3. Describe the chemical tests used to identify hydrogen, oxygen, carbon dioxide and chlorine.
4. Explain how flame tests and precipitation reactions can be used to identify ions.
5. Compare instrumental methods with chemical tests for identifying substances.
6. Explain how different analytical tests provide evidence about the identity of unknown substances.

Topic 9 – Chemistry of the Atmosphere Retrieval Starters

4.9.1.1 The Proportions of Different Gases in the Atmosphere

1. What percentage of the modern atmosphere is nitrogen?
2. What percentage of the modern atmosphere is oxygen?
3. Name two gases found in small proportions in the atmosphere.

4. What are the approximate proportions of nitrogen and oxygen in the atmosphere?
5. Why is the composition of the atmosphere described as stable over the last 200 million years?
6. What is the most abundant gas in the Earth's atmosphere?

4.9.1.2 The Earth's Early Atmosphere

1. What gases are thought to have been present in the Earth's early atmosphere?
2. How did volcanic activity contribute to the formation of the early atmosphere?
3. Why was there little or no oxygen in the Earth's early atmosphere?
4. How did oceans form from water vapour in the early atmosphere?
5. How did the amount of carbon dioxide in the atmosphere decrease when oceans formed?
6. Explain why evidence about the Earth's early atmosphere is limited.

4.9.1.3 How Oxygen Increased

1. Which process produced the oxygen in the Earth's atmosphere?
2. Which organisms first produced oxygen by photosynthesis?
3. Approximately how long ago did algae first produce oxygen?
4. How did the increase in oxygen allow animals to evolve?
5. Write the word equation for photosynthesis.
6. Explain how plants caused the percentage of oxygen in the atmosphere to increase.

4.9.1.4 How Carbon Dioxide Decreased

1. How did plants and algae reduce the amount of carbon dioxide in the atmosphere?
2. How were sedimentary rocks formed from carbon dioxide?
3. How were fossil fuels formed from carbon-containing materials?
4. Why did the formation of limestone reduce carbon dioxide levels?
5. Name three carbon-containing deposits formed over millions of years.
6. Explain how photosynthesis and the formation of fossil fuels caused carbon dioxide levels to decrease.

4.9.2.1 Greenhouse Gases

1. What is a greenhouse gas?

2. Name three greenhouse gases.
3. Why are greenhouse gases important for life on Earth?
4. What happens when short wavelength radiation from the Sun reaches Earth?
5. What happens when long wavelength radiation is emitted from Earth?
6. Explain how greenhouse gases cause the greenhouse effect.

4.9.2.2 Human Activities Which Increase Greenhouse Gases

1. Name two human activities that increase carbon dioxide levels in the atmosphere.
2. Name two human activities that increase methane levels in the atmosphere.
3. How does burning fossil fuels increase carbon dioxide levels?
4. How does farming increase methane levels?
5. Why do scientists use peer review when studying climate change?
6. Explain why evidence about climate change can be difficult to interpret.

4.9.2.3 Global Climate Change

1. What is global climate change?
2. What is the main cause of the increase in average global temperature?
3. State four possible effects of global climate change.
4. How could rising global temperatures affect sea levels?
5. Why is climate change difficult to predict accurately?
6. Explain why scientists consider the risks and environmental impacts of global climate change.

4.9.2.4 The Carbon Footprint and Its Reduction

1. What is a carbon footprint?
2. What greenhouse gases are included when calculating a carbon footprint?
3. How can reducing fossil fuel use reduce a carbon footprint?
4. How can reducing methane emissions reduce a carbon footprint?
5. Give two actions that can reduce carbon dioxide emissions.
6. Explain why some methods of reducing carbon footprints are limited.

4.9.3.1 Atmospheric Pollutants from Fuels

1. Which process is a major source of atmospheric pollutants?

2. What gases can be produced when fuels containing carbon and hydrogen are burned?
3. How is carbon monoxide produced during combustion?
4. How are sulfur dioxide and oxides of nitrogen produced during combustion?
5. What are particulates?
6. Explain why incomplete combustion can produce carbon monoxide and soot.

4.9.3.2 Properties and Effects of Atmospheric Pollutants

1. Why is carbon monoxide dangerous to humans?
2. Why is carbon monoxide difficult to detect?
3. What problems are caused by sulfur dioxide?
4. What problems are caused by oxides of nitrogen?
5. What problems are caused by particulates in the atmosphere?
6. Explain how atmospheric pollutants can affect both human health and the environment.

Topic 9 Review

1. Describe how the composition of the Earth's atmosphere has changed over time.
2. Explain how photosynthesis increased oxygen levels and reduced carbon dioxide levels.
3. Explain how greenhouse gases affect the temperature of the Earth.
4. Describe how human activities contribute to climate change.
5. Explain how combustion of fuels produces atmospheric pollutants.
6. Evaluate methods used to reduce greenhouse gas emissions and atmospheric pollution.

Topic 10 – Using Resources Retrieval Starters

4.10.1.1 Using the Earth's Resources and Sustainable Development

1. What are natural resources used for by humans?

2. What is meant by a finite resource?
3. What is meant by a renewable resource?
4. Give one example of a natural product that can be replaced by an agricultural product.
5. Give one example of a natural product that can be replaced by a synthetic product.
6. Explain what is meant by sustainable development.

4.10.1.2 Potable Water

1. What is potable water?
2. What is the difference between potable water and pure water?
3. What makes water suitable for drinking as potable water?
4. What are the three main stages used to produce potable water from fresh water?
5. Name three substances or methods that can be used to sterilise water.
6. Explain why desalination of seawater requires large amounts of energy.

Required Practical 8: Analysis and Purification of Water Samples

1. What properties of water samples are investigated in the required practical analysing water?
2. How can the amount of dissolved solids in water be measured?
3. Why is distillation used when purifying water?
4. Why is pH measured when analysing water samples?
5. Explain why water from different sources may have different levels of dissolved substances.
6. Explain why distilled water is purer than potable water.

4.10.1.3 Waste Water Treatment

1. Why does waste water need to be treated before being released into the environment?
2. What is removed during the screening and grit removal stage of sewage treatment?
3. What is produced during the sedimentation stage of sewage treatment?
4. What happens during anaerobic digestion of sewage sludge?
5. What happens during aerobic biological treatment of effluent?
6. Explain why potable water is easier to obtain from fresh water than from waste water or seawater.

4.10.1.4 Alternative Methods of Extracting Metals (HT Only)

1. Why are alternative methods of metal extraction needed?
2. What is phytomining?
3. How does phytomining extract metals from the ground?
4. What is bioleaching?
5. How can copper be extracted from a solution of copper compounds?
6. Evaluate one advantage and one disadvantage of using biological methods of metal extraction.

4.10.2.1 Life Cycle Assessment

1. What is a life cycle assessment (LCA)?
2. What stages of a product's life are considered in a life cycle assessment?
3. Why are raw materials considered in a life cycle assessment?
4. Why are pollutant effects harder to include in a life cycle assessment than energy use?
5. Why are selective LCAs sometimes criticised?
6. Explain how a life cycle assessment can be used to compare plastic and paper bags.

4.10.2.2 Ways of Reducing the Use of Resources

1. What are the three main ways end users can reduce resource use?
2. Why does recycling reduce the use of limited resources?
3. Why does recycling metals reduce energy use?
4. How can glass bottles be reused or recycled?
5. How are metals recycled into new products?
6. Explain why recycling reduces environmental impacts caused by mining and quarrying.

4.10.3.1 Corrosion and Its Prevention

1. What is corrosion?
2. What is rusting?
3. What two substances are needed for iron to rust?
4. How can painting prevent iron from rusting?
5. What is sacrificial protection?
6. Explain why zinc protects iron from rusting when iron is galvanised.

4.10.3.2 Alloys as Useful Materials

1. What is an alloy?
2. What elements are present in bronze?
3. What elements are present in brass?
4. What is the percentage of gold in 18 carat gold?
5. Why are alloys often more useful than pure metals?
6. Explain why stainless steel is resistant to corrosion.

4.10.3.3 Ceramics, Polymers and Composites

1. What materials are used to make soda-lime glass?
2. Why does borosilicate glass have a higher melting point than soda-lime glass?
3. How are clay ceramics produced?
4. What is the difference between thermosoftening and thermosetting polymers?
5. What are the two components of a composite material?
6. Explain how the structure of a material affects its properties and uses.

4.10.4.1 The Haber Process

1. What is ammonia used to manufacture?
2. What are the raw materials needed for the Haber process?
3. What is the source of nitrogen used in the Haber process?
4. What is the source of hydrogen used in the Haber process?
5. What are the conditions needed for Haber process?
6. Explain why the Haber process uses high temperature and high pressure.

4.10.4.2 Production and Uses of NPK Fertilisers

1. What elements are present in NPK fertilisers?
2. Why are NPK fertilisers added to soil?
3. What are the names of the three elements represented by NPK?
4. Why can phosphate rock not be used directly as a fertiliser?
5. Which acids can be used to treat phosphate rock to make fertilisers?
6. Explain why NPK fertilisers are described as formulations.

Topic 10 Review

1. Explain how chemistry helps humans use resources sustainably.
2. Compare the production of potable water from fresh water and seawater.

3. Explain how waste water is treated before being released into the environment.
4. Describe methods used to reduce the use of limited resources.
5. Explain how the properties of alloys, ceramics, polymers and composites make them useful materials.
6. Explain how the Haber process produces ammonia and why the industrial conditions are a compromise.