Paper 1:

Topic 1 – Atomic Structure and the Periodic Table

5.1.1.1 Atoms, Elements and Compounds

1. The smallest particle of an element that can exist and retain the properties of that element.
2. A substance made of only one type of atom.
3. By their chemical symbols.
4. Approximately 100 different elements.
5. A substance made from two or more different elements chemically bonded together.
6. Compounds contain elements that are chemically bonded, so chemical reactions are needed to break the bonds and separate them.

5.1.1.2 Mixtures

1. Two or more substances mixed together that are not chemically bonded.
2. The properties of substances in a mixture are not changed because no chemical bonds are formed.
3. A mixture contains substances that are not chemically bonded and can be separated by physical methods; a compound contains elements chemically bonded and requires a chemical reaction to separate.
4. Filtration.
5. Fractional distillation.
6. Physical separation methods do not produce new substances because no chemical bonds are broken or formed.

5.1.1.3 The Development of the Model of the Atom

1. Atoms were originally thought to be tiny solid spheres that could not be divided.
2. The discovery of electrons.
3. A positive sphere containing negatively charged electrons.

4. The atom was mostly empty space with a small, dense, positively charged nucleus.
5. The nuclear model was developed because the plum pudding model could not explain the results of the alpha particle scattering experiment.
6. Scientific evidence can show that existing models are incorrect, leading to new models being developed.

5.1.1.4 Relative Electrical Charges of Subatomic Particles

1. +1
2. 0
3. -1
4. Atoms have equal numbers of protons and electrons, so the positive and negative charges cancel out.
5. The number of protons in an atom.
6. All atoms of the same element have the same number of protons because the number of protons defines the element.

5.1.1.5 Size and Mass of Atoms

1. Approximately 0.1 nm.
2. The nucleus.
3. Protons and neutrons.
4. The total number of protons and neutrons in an atom.
5. Isotopes have the same number of protons but different numbers of neutrons.
6. Protons = 12, neutrons = 12, electrons = 12.

5.1.1.6 Relative Atomic Mass

1. The weighted mean mass of atoms of an element compared with one twelfth of the mass of carbon-12.
2. Because it is an average value that includes isotopes with different masses.
3. The mass and abundance of each isotope.
4. Because elements contain different isotopes.
5. (20 × 75/100) + (22 × 25/100) = 20.5.
6. Chlorine has different isotopes, mainly chlorine-35 and chlorine-37, giving an average relative atomic mass of approximately 35.5.

5.1.1.7 Electronic Structure

1. Energy levels or electron shells.
2. The lowest energy level first.
3. Two electrons.
4. Sodium: 2,8,1.
5. Oxygen: 2,6.
6. The outer shell determines the chemical properties and how an atom reacts.

5.1.2.1 The Periodic Table

1. Elements are arranged in order of increasing atomic number.
2. The number of protons in an atom.
3. Elements in the same group have the same number of electrons in their outer shell.
4. The group number tells you the number of electrons in the outer shell.
5. The period shows the number of occupied shells and the group shows the number of outer electrons.
6. Reactivity can be predicted from the element's position and trends within its group.

5.1.2.2 Development of the Periodic Table

1. Early scientists arranged elements by atomic weight and properties.
2. Some elements were placed incorrectly because atomic weight did not always match their properties.
3. Dmitri Mendeleev.
4. He left gaps for undiscovered elements and predicted their properties.
5. New elements were discovered that matched Mendeleev's predictions.
6. Isotopes showed that elements with different masses can have the same chemical properties, so atomic weight was not always suitable.

5.1.2.3 Metals and Non-metals

1. Metals are generally found on the left and centre of the periodic table.
2. Non-metals are generally found on the right side of the periodic table.
3. Positive ions.
4. Negative ions.
5. Metals are good conductors of electricity.
6. Metals lose electrons to form positive ions, while non-metals gain electrons to form negative ions.

5.1.2.4 Group 0

1. Noble gases.
2. They have a full outer electron shell.
3. Eight electrons, except helium which has two.
4. Helium has one shell that is full with two electrons.
5. Boiling points increase down Group 0.
6. Their full outer shells make them stable and unreactive.

5.1.2.5 Group 1

1. Alkali metals.
2. One electron.
3. Reactivity increases down the group.
4. Hydrogen.
5. Positive ions.
6. The outer electron is further from the nucleus and easier to lose.

5.1.2.6 Group 7

1. Halogens.
2. Seven electrons.
3. Non-metals.
4. Reactivity decreases down the group.
5. A more reactive halogen displaces a less reactive halogen from its salt.
6. Iodine has stronger intermolecular forces because it has larger molecules and more electrons.

Topic 1 Review

1. The atomic model changed as new evidence was discovered, including the discovery of electrons and the nucleus.
2. An atom contains protons and neutrons in the nucleus with electrons arranged in shells around it.
3. Electronic structure determines position because the number of shells gives the period and the number of outer electrons gives the group.
4. Group 1 metals become more reactive down the group, while Group 7 halogens become less reactive down the group.
5. The periodic table allows scientists to predict properties because elements in the same group have similar electron arrangements.
6. Atoms are single particles, elements contain one type of atom, compounds contain chemically bonded elements, and mixtures contain substances that are not chemically bonded.

Topic 2 – Bonding, Structure and the Properties of Matter

5.2.1.1 Chemical Bonds

1. Ionic, covalent and metallic bonds.
2. Positive and negative ions.
3. Share electrons.
4. Delocalised electrons.
5. Ionic bonding.
6. Chemical bonds form when atoms transfer or share electrons to achieve stable electron arrangements.

5.2.1.2 Ionic Bonding

1. Electrons are transferred from the metal atom to the non-metal atom.
2. It loses electrons and forms a positive ion.
3. It gains electrons and forms a negative ion.
4. They gain or lose electrons to achieve a full outer shell.
5. Magnesium loses two electrons and each chlorine atom gains one electron, forming Mg^{2+} and Cl^- ions with a giant ionic lattice structure.
6. The group number shows the number of electrons lost or gained to form a stable ion.

5.2.1.3 Ionic Compounds

1. A giant ionic lattice structure.
2. Strong electrostatic forces of attraction between oppositely charged ions.
3. A lot of energy is needed to overcome the strong forces between ions.
4. The ions are free to move and carry charge.
5. The ions are free to move through the solution and carry charge.
6. Use the charges of the ions to find the simplest ratio of ions.

5.2.1.4 Covalent Bonding

1. They share electrons.
2. Strong electrostatic forces of attraction exist between the nuclei and shared electrons.
3. A small molecule usually contains only a few atoms.

4. Polymers are very large molecules made from repeating units; small molecules contain only a few atoms.
5. Diamond, graphite and graphene.
6. Carbon dioxide has a central carbon atom double-bonded to two oxygen atoms. Carbon shares two pairs of electrons with each oxygen, creating two double bonds to achieve a full outer shell

5.2.1.5 Metallic Bonding

1. A giant metallic structure.
2. Outer electrons become delocalised.
3. Electrons that are free to move through the structure.
4. Strong attraction between positive metal ions and delocalised electrons.
5. Positive calcium ions arranged in a lattice surrounded by delocalised electrons.
6. They are not attached to one atom and can move through the metal.

5.2.2.1 The Three States of Matter

1. Solid, liquid and gas.
2. Particles gain energy, move faster and become less ordered.
3. Particles gain enough energy to overcome forces and spread apart.
4. More energy is needed to overcome stronger forces between particles.
5. Changes of state occur when particles gain or lose energy, changing their movement and arrangement.
6. It assumes particles are solid spheres and does not explain particle interactions or forces accurately.

5.2.2.2 State Symbols

1. (s)
2. (l)
3. (g)
4. (aq)
5. To show the physical state of substances in a reaction.
6. (aq)

5.2.2.3 Properties of Ionic Compounds

1. They contain many oppositely charged ions arranged in a regular lattice.
2. Strong electrostatic forces between ions.
3. Because the ions have strong attractions between them.

4. Ions are free to move when molten.
5. Ions are free to move when dissolved in water.
6. The ions cannot move because they are fixed in position.

5.2.2.4 Properties of Small Molecules

1. Weak intermolecular forces require little energy to overcome.
2. Weak intermolecular forces.
3. Covalent bonds involve strong forces within molecules, while intermolecular forces are weaker forces between molecules.
4. Larger molecules have stronger intermolecular forces.
5. They do not contain charged particles that can move.
6. Intermolecular forces.

5.2.2.5 Polymers

1. Small molecules called monomers.
2. They contain thousands of atoms joined together.
3. Covalent bonds.
4. They contain long chains of atoms with strong covalent bonds.
5. A polymer has repeating units in its structure.
6. Polymers are very large molecules, while small molecules contain fewer atoms.

5.2.2.6 Giant Covalent Structures

1. A structure containing many atoms joined by covalent bonds in a giant network.
2. Many strong covalent bonds require large amounts of energy to break.
3. Covalent bonds extend throughout the whole structure.
4. Diamond, graphite and graphene.
5. It contains many strong covalent bonds between carbon atoms.
6. A giant covalent structure has a large network of atoms joined by covalent bonds.

5.2.2.7 Properties of Metals and Alloys

1. Strong metallic bonds require a lot of energy to break.
2. Regular layers of metal atoms.
3. Layers of atoms can slide over each other.
4. Different sized atoms disrupt the layers, making them harder to move.
5. They make the layers less able to slide over each other.

6. Alloys are stronger and harder than pure metals.

5.2.2.8 Metals as Conductors

1. Metals contain delocalised electrons that can move.
2. Delocalised electrons.
3. Delocalised electrons transfer energy through the metal.
4. They carry electrical charge and transfer thermal energy.
5. They are good electrical conductors because electrons can move freely.
6. Metallic bonding allows electrons to move through the structure.

5.2.3.1 Diamond

1. Four covalent bonds.
2. A giant covalent structure.
3. Each carbon atom is strongly bonded to four others.
4. Many strong covalent bonds require lots of energy to break.
5. It has no free electrons to carry charge.
6. Its giant covalent structure explains its hardness and high melting point.

5.2.3.2 Graphite

1. Three covalent bonds.
2. Layers of carbon atoms arranged in hexagonal structures.
3. It has delocalised electrons that can move.
4. Layers are held together by weak forces.
5. Weak forces between layers allow them to slide.
6. Its layered structure explains why it is soft and conducts electricity.

5.2.3.3 Graphene and Fullerenes

1. A single layer of graphite made of carbon atoms.
2. It is one layer of graphite.
3. It conducts electricity and is very strong.
4. Molecules made from carbon atoms arranged in hollow shapes.
5. A spherical molecule containing 60 carbon atoms.
6. Used in nanotechnology and strengthening materials.

Topic 2 Review

1. Bonding and structure determine properties such as melting point, conductivity and hardness.
2. Ionic bonding involves electron transfer, covalent bonding involves electron sharing, and metallic bonding involves positive ions and delocalised electrons.
3. Solid ionic compounds cannot conduct because ions cannot move; molten ionic compounds conduct because ions are free to move.
4. Diamond has a giant structure with four bonds per carbon atom, while graphite has layers and delocalised electrons.
5. Metals conduct because they contain mobile delocalised electrons.
6. Scientists use bonding and structure knowledge to design materials with specific properties.

Topic 3 – Quantitative Chemistry

5.3.1.1 Conservation of Mass and Balanced Chemical Equations

1. Mass is not created or destroyed during a chemical reaction.
2. Atoms are rearranged during reactions, but no atoms are lost or gained.
3. Chemical equations must be balanced to show the same number of atoms of each element on both sides.
4. The numbers in front of chemical formulas show the ratio of reacting substances and products.
5. A multiplier is the number in front of a formula showing the number of molecules or moles; a subscript shows the number of atoms of an element in a formula.
6. $H_2 + O_2 \rightarrow H_2O$ becomes $2H_2 + O_2 \rightarrow 2H_2O$.

5.3.1.2 Relative Formula Mass

1. The sum of the relative atomic masses of all atoms in a compound.
2. Add together the relative atomic masses of all atoms in the chemical formula.
3. Relative atomic masses from the periodic table.
4. H_2O: Mr = (2 × 1) + 16 = 18.

5. The total mass of reactants equals the total mass of products because mass is conserved.
6. Percentage by mass = (mass of element in compound ÷ relative formula mass) × 100.

5.3.1.3 Mass Changes When a Reactant or Product is a Gas

1. A reaction can appear to change mass because gases can enter or leave the system.
2. The metal reacts with oxygen from the air, adding oxygen atoms to the metal.
3. Carbon dioxide gas escapes into the surroundings.
4. Gas entering or leaving an open system changes the measured mass.
5. The metal oxide contains the original metal plus oxygen atoms, increasing the mass.
6. Gas particles can move into or out of the system, causing changes in measured mass.

5.3.1.4 Chemical Measurements

1. Uncertainty is the amount by which a measurement may differ from the true value.
2. Every measurement has uncertainty because equipment and human measurements have limitations.
3. Uncertainty can be estimated by: (highest value – lowest value) ÷ 2.
4. The spread of results shows how close together the measurements are.
5. Repeated measurements improve confidence by allowing a mean to be calculated.
6. A smaller range suggests measurements are more precise and reliable.

5.3.2.1 Moles (HT Only)

1. The mole (mol).
2. The mass of one mole is equal to the relative formula mass in grams.
3. 6.02×10^{23} particles per mole.
4. 6.02×10^{23} particles.
5. Moles = mass ÷ Mr. Water: 18 ÷ 18 = 1 mol.
6. Mr of Cl_2 = 35.5 × 2 = 71. Mass = 0.5 × 71 = 35.5 g.

5.3.2.2 Amounts of Substances in Equations (HT Only)

1. A balanced equation shows the ratio of moles of reactants and products.
2. The coefficients in a balanced equation represent the mole ratio.
3. $Mg + 2HCl \rightarrow MgCl_2 + H_2$. One mole of magnesium reacts with 2 moles of hydrochloric acid.
4. Calculate moles of the known substance, use the mole ratio, then convert the product moles into mass.
5. Mole ratios show the exact amounts of substances reacting.
6. $4Al + 3O_2 \rightarrow 2Al_2O_3$. Moles of Al = 135 ÷ 27 = 5 mol. Ratio $Al:Al_2O_3$ = 4:2, so moles of Al_2O_3 = 5 ÷ 2 = 2.5 mol. Mr Al_2O_3 = (2 × 27) + (3 × 16) = 102. Mass = 2.5 × 102 = 255 g.

5.3.2.3 Using Moles to Balance Equations (HT Only)

1. Moles are used because equations must show the correct ratios of reacting particles.
2. Masses must be converted into moles.
3. Mole ratios.
4. CO_2: 264 ÷ 44 = 6 mol. H_2O: 108 ÷ 18 = 6 mol. Glucose: 180 ÷ 180 = 1 mol. O_2: 192 ÷ 32 = 6 mol. Equation: $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$.
5. CH_4: 16 ÷ 16 = 1 mol. O_2: 64 ÷ 32 = 2 mol. CO_2: 44 ÷ 44 = 1 mol. H_2O: 36 ÷ 18 = 2 mol. Equation: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$.
6. Mg: 24 ÷ 24 = 1 mol. O_2: 16 ÷ 32 = 0.5 mol. MgO: 40 ÷ 40 = 1 mol. Ratio = 2:1:2. Equation: $2Mg + O_2 \rightarrow 2MgO$.

5.3.2.4 Limiting Reactants (HT Only)

1. The reactant that is completely used up first.
2. It ensures the reaction uses the available reactants efficiently and determines the maximum product formed.
3. It limits the amount of product that can be produced.
4. The excess reactant remains after the reaction has finished.
5. Compare the number of moles of reactants with the balanced equation ratio.
6. Increasing the amount of the limiting reactant allows more product to form.

5.3.2.5 Concentration of Solutions

1. The amount of solute dissolved in a given volume of solution.
2. g/dm³ or mol/dm³.
3. Concentration = mass ÷ volume.
4. Concentration = 20 ÷ 2 = 10 g/dm³.
5. Increasing the mass of solute increases concentration.
6. Increasing the volume of solution decreases concentration.

Topic 3 Review

1. The law of conservation of mass means atoms are rearranged but total mass remains constant.
2. Balanced equations show the quantities and ratios of reactants and products.
3. Relative formula mass is calculated by adding atomic masses and is used in chemical calculations.
4. Moles allow chemists to measure and compare amounts of substances.
5. The limiting reactant determines the maximum amount of product formed.
6. Concentration calculations show the amount of solute present in a solution.

Topic 4 – Chemical Changes

5.4.1.1 Metal Oxides

1. Metal oxides.
2. They involve the gain of oxygen by the metal.
3. Oxidation is the gain of oxygen.
4. Reduction is the loss of oxygen.
5. The metal atom loses electrons and forms a positive ion.
6. The metal gains oxygen, so it is oxidised.

5.4.1.2 The Reactivity Series

1. The ability of a metal to undergo reactions.
2. Metals lose electrons to form positive ions.
3. The order of metals from most reactive to least reactive.

4. The metal that displaces another metal from its compound is more reactive.
5. Reactions with water and acids show how easily metals react and allow comparison of reactivity.
6. A more reactive metal removes oxygen or other elements from a less reactive metal compound.

5.4.1.3 Extraction of Metals and Reduction

1. Some metals are found as compounds because they have reacted with other elements.
2. Metals below carbon in the reactivity series can be found naturally as the metal.
3. Carbon can remove oxygen from less reactive metal oxides.
4. Reduction is the loss of oxygen.
5. The metal oxide is reduced.
6. More reactive metals cannot be extracted using carbon because they form stronger bonds with oxygen.

5.4.1.4 Oxidation and Reduction in Terms of Electrons (HT Only)

1. Oxidation is the loss of electrons.
2. Reduction is the gain of electrons.
3. Electrons are lost during oxidation.
4. Electrons are gained during reduction.
5. Half equations show electrons being lost or gained.
6. The substance losing electrons is oxidised; the substance gaining electrons is reduced.

5.4.2.1 Reactions of Acids with Metals

1. A salt and hydrogen gas.
2. Measure the volume of gas produced over time or measure mass loss over time.
3. Hydrogen.
4. Electrons are transferred between substances.
5. The metal is oxidised.
6. Magnesium loses electrons to form Mg^{2+} ions, while hydrogen ions gain electrons to form hydrogen gas.

5.4.2.2 Neutralisation of Acids and Salt Production

1. A salt and water.
2. A salt, water and carbon dioxide.
3. Hydrochloric acid.
4. Nitric acid.
5. Sulfuric acid.
6. The acid and reactant determine the ions present in the salt produced.

5.4.2.3 Soluble Salts

1. React an acid with a suitable metal, metal oxide, metal hydroxide or metal carbonate.
2. To ensure all the acid reacts.
3. To remove the unreacted excess solid.
4. Evaporate some water and allow crystals to form.
5. Water is removed so the salt becomes more concentrated and crystals form.
6. Add excess insoluble solid to acid, filter to remove excess solid, evaporate water and dry the salt crystals.

5.4.2.4 The pH Scale and Neutralisation

1. Hydrogen ions (H^+).
2. Hydroxide ions (OH^-).
3. The concentration of hydrogen ions in a solution.
4. pH 7.
5. Acids have pH values below 7; alkalis have pH values above 7.
6. Hydrogen ions react with hydroxide ions to form water.

5.4.2.5 Strong and Weak Acids (HT Only)

1. An acid that completely ionises in water.
2. An acid that only partially ionises in water.
3. Strong acids release all their hydrogen ions.
4. Weak acids only release some hydrogen ions.
5. The hydrogen ion concentration increases by 1000 times.
6. Strength describes ionisation; concentration describes the amount of acid dissolved in a solution.

5.4.3.1 The Process of Electrolysis

1. The decomposition of an ionic compound using electricity.
2. A substance containing free-moving ions that conducts electricity.
3. Ions are free to move and carry charge.
4. The cathode.
5. The anode.
6. Electricity causes ions to gain or lose electrons, forming elements.

5.4.3.2 Electrolysis of Molten Ionic Compounds

1. The metal and non-metal elements.
2. Lead.
3. Bromine.
4. Ions must be free to move to carry charge.
5. They do not react with the products formed.
6. Positive ions form metals at the cathode; negative ions form non-metals at the anode.

5.4.3.3 Using Electrolysis to Extract Metals

1. For metals more reactive than carbon.
2. They cannot be reduced by carbon because they are too strongly bonded to oxygen.
3. Cryolite lowers the melting point of aluminium oxide and reduces energy costs.
4. Aluminium.
5. Large amounts of energy are needed to melt compounds and move ions.
6. The positive electrode reacts with oxygen to form carbon dioxide and wears away.

5.4.3.4 Electrolysis of Aqueous Solutions

1. The reactivity of the ions and their position in the reactivity series.
2. Hydrogen gas.
3. Oxygen gas.
4. Halogen gases are produced.
5. Water molecules can provide hydrogen and hydroxide ions.
6. Predict products by comparing the reactivity of ions present.

5.4.3.5 Representation of Reactions at Electrodes as Half Equations (HT Only)

1. Positive ions gain electrons at the cathode.
2. They are reduction reactions because electrons are gained.
3. Negative ions lose electrons at the anode.
4. They are oxidation reactions because electrons are lost.
5. A half equation shows electrons transferred during a reaction.
6. Example: $Mg^{2+} + 2e^- \rightarrow Mg$.

Topic 4 Review

1. The reactivity series shows how easily metals react and allows predictions of displacement reactions.
2. Oxidation is loss of electrons or gain of oxygen; reduction is gain of electrons or loss of oxygen.
3. Metals can be extracted by reduction with carbon or by electrolysis depending on their reactivity.
4. Acids react with metals to produce salts and hydrogen, with bases to produce salts and water, and with carbonates to produce salts, water and carbon dioxide.
5. Electrolysis separates ionic compounds by moving ions to electrodes where they gain or lose electrons.
6. Carbon reduction is used for less reactive metals; electrolysis is needed for more reactive metals.

Topic 5 – Energy Changes

5.5.1.1 Energy Transfer During Exothermic and Endothermic Reactions

1. Energy cannot be created or destroyed; it is transferred between stores.
2. A reaction that transfers energy to the surroundings.
3. The temperature of the surroundings increases.
4. Combustion and neutralisation.
5. A reaction that takes in energy from the surroundings.
6. The temperature of the surroundings decreases.

5.5.1.2 Reaction Profiles

1. Particles must collide with enough energy and the correct orientation.
2. The minimum amount of energy needed for a reaction to occur.
3. The energy changes during a chemical reaction.
4. An exothermic reaction has products at a lower energy level than the reactants.
5. Endothermic profile: products are higher in energy than reactants; activation energy is shown as the energy needed to start the reaction; overall energy change is positive.
6. Exothermic profile: products are lower in energy than reactants; activation energy is shown as the energy needed to start the reaction; overall energy change is negative.

5.5.1.3 The Energy Change of Reactions (HT Only)

1. Energy is needed to break existing bonds and form new bonds.
2. Energy is absorbed when chemical bonds are broken.
3. Energy is released when new chemical bonds are formed.
4. More energy is released when bonds form than is absorbed when bonds break.
5. More energy is absorbed breaking bonds than is released forming bonds.
6. Overall energy change = energy needed to break bonds – energy released forming bonds.

Topic 5 Review

1. Exothermic reactions release energy to the surroundings; endothermic reactions absorb energy from the surroundings.
2. Energy is absorbed when bonds break and released when bonds form.
3. Reaction profiles show activation energy and whether energy is released or absorbed.
4. Activation energy is needed to start reactions by allowing particles to collide successfully.
5. Exothermic reactions have a negative energy change; endothermic reactions have a positive energy change.
6. Energy changes are used in applications such as fuels, hand warmers and cooling packs.

Paper 2:

Topic 6 – The Rate and Extent of Chemical Change

5.6.1.1 Calculating rates of reactions

1. Mean rate = quantity of reactant used ÷ time taken.
2. Mean rate = quantity of product formed ÷ time taken.
3. g/s or g/min.
4. cm^3/s or cm^3/min.
5. The gradient of the graph shows the rate; a steeper gradient means a faster reaction.
6. Draw a tangent at the point and calculate its gradient: change in amount ÷ change in time.

5.6.1.2 Factors which affect the rates of chemical reactions

1. Concentration, pressure, surface area, temperature and catalysts.
2. Increasing concentration increases the rate because there are more particles in the same volume.
3. Decreasing pressure decreases the rate because gas particles collide less often.
4. Increasing surface area increases the rate because more particles are exposed for collisions.
5. Increasing temperature increases the rate because particles move faster (due to more kinetic energy) and collide more successfully.
6. Catalysts increase the rate by providing an alternative pathway with lower activation energy.

5.6.1.3 Collision theory and activation energy

1. Particles must collide with enough energy and the correct orientation for a reaction to occur.
2. The minimum energy needed for particles to react.
3. More particles are present, causing more frequent successful collisions.

4. Particles move faster and have more energy, increasing successful collisions.
5. More surface particles are exposed, increasing collision frequency.
6. Higher pressure pushes gas particles closer together, causing more frequent collisions.

5.6.1.4 Catalysts

1. A substance that increases the rate of reaction without being used up.
2. Catalysts do not take part in the overall reaction and are not changed.
3. They provide an alternative reaction pathway with lower activation energy.
4. The activation energy decreases.
5. Different reactions require different catalysts because they involve different reactants and mechanisms.
6. Enzymes are biological catalysts that speed up reactions in living organisms.

5.6.2.1 Reversible reactions

1. A reaction where products can react to form the original reactants.
2. They are represented using a reversible arrow (⇋).
3. Products can react to form reactants again.
4. By changing the conditions such as temperature, pressure or concentration.
5. Changing temperature, pressure or concentration.
6. They can proceed in both directions, producing both reactants and products.

5.6.2.2 Energy changes and reversible reactions

1. It is endothermic in the opposite direction.
2. The energy change is equal but opposite.
3. Energy is transferred to the surroundings.
4. Energy is taken in from the surroundings.
5. The reverse reaction breaks and forms bonds in the opposite way.
6. Endothermic.

5.6.2.3 Equilibrium

1. A state where the forward and reverse reactions occur at the same rate.
2. A closed system and constant conditions.
3. Reactants and products must not escape or the amounts will change.

4. The forward and reverse reactions occur at the same rate.
5. It is dynamic because reactions continue happening.
6. The amounts of reactants and products remain constant.

5.6.2.4 The effect of changing conditions on equilibrium (HT only)

1. The equilibrium position shifts to oppose the change.
2. Le Chatelier's Principle.
3. If conditions are changed, the equilibrium moves to reduce the effect of that change.
4. Temperature, pressure and concentration changes shift equilibrium.
5. The system responds to restore equilibrium.
6. Increase the amount of desired product by changing conditions to favour the forward reaction.

5.6.2.5 The effect of changing concentration (HT only)

1. The equilibrium shifts to use up the increased reactant.
2. More reactant particles are available, increasing successful collisions and producing more products.
3. The equilibrium shifts to replace the removed product.
4. The system adjusts to restore the balance between reactants and products.
5. Increasing reactant concentration shifts equilibrium towards products; increasing product concentration shifts it towards reactants.
6. $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$: increasing hydrogen concentration shifts equilibrium to the right, producing more ammonia.

5.6.2.6 The effect of temperature changes on equilibrium (HT only)

1. More products are formed because the equilibrium shifts in the endothermic direction.
2. Less product is formed because the equilibrium shifts in the endothermic direction.
3. Less product is formed because the equilibrium shifts away from the endothermic direction.
4. More product is formed because the equilibrium shifts in the exothermic direction.

5. Temperature changes alter the balance between energy absorbed and released.
6. $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ is exothermic, so increasing temperature shifts equilibrium left, reducing ammonia yield.

5.6.2.7 The effect of pressure changes on equilibrium (HT only)

1. Increasing pressure shifts equilibrium towards the side with fewer gas molecules.
2. This reduces pressure because fewer gas particles occupy less space.
3. Decreasing pressure shifts equilibrium towards the side with more gas molecules.
4. More gas molecules increase pressure and oppose the change.
5. Count the number of gas molecules on each side of the equation.
6. $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$: increasing pressure shifts equilibrium right because there are fewer gas molecules on the product side.

Topic 6 Review

1. Increasing concentration, pressure, temperature and surface area increases reaction rate; catalysts also increase rate.
2. Higher temperature gives particles more energy, causing more successful collisions.
3. Catalysts lower activation energy and are not used up during reactions.
4. Dynamic equilibrium occurs in a closed system where forward and reverse reactions happen at equal rates.
5. Increasing pressure favours the side with fewer gas molecules; temperature changes favour the endothermic or exothermic direction; concentration changes oppose the change.
6. Reversible reactions can go both directions; equilibrium is when the forward and reverse rates are equal; irreversible reactions only go one direction.

Topic 7 – Organic Chemistry

5.7.1.1 Crude oil, hydrocarbons and alkanes

1. Crude oil is a mixture of hydrocarbons formed from the remains of ancient organisms.
2. Crude oil is finite because it takes millions of years to form and cannot be replaced quickly.
3. Hydrocarbons are compounds made only of hydrogen and carbon.
4. Carbon and hydrogen.
5. C_nH_{2n+2}.
6. C_4H_{10} follows the formula C_nH_{2n+2} because 2(4)+2 = 10, so it is an alkane.

5.7.1.2 Fractional distillation and petrochemicals

1. Fractional distillation.
2. Crude oil is heated and separated into fractions based on different boiling points.
3. Different hydrocarbons have different chain lengths and intermolecular forces, giving different boiling points.
4. Fuels, solvents and chemicals for making products.
5. Petrol and diesel.
6. Crude oil contains many useful hydrocarbons that can be used to make petrochemicals.

5.7.1.3 Properties of hydrocarbons

1. Boiling point increases as hydrocarbon molecules get larger.
2. Viscosity increases as hydrocarbon molecules get larger.
3. Flammability decreases as hydrocarbon molecules get larger.
4. Small hydrocarbons ignite easily and release useful energy.
5. Carbon dioxide and water.
6. $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$.

5.7.1.4 Cracking and alkenes

1. Cracking is breaking large hydrocarbon molecules into smaller, more useful molecules.
2. To produce more useful fuels and chemicals from larger hydrocarbons.
3. Alkanes and alkenes.
4. Bromine water.

5. Orange/brown bromine water turns colourless.
6. Cracking produces useful smaller hydrocarbons and alkenes needed for fuels and polymers.

Topic 7 Review

1. Crude oil is separated by fractional distillation using differences in boiling points.
2. Larger hydrocarbons have higher boiling points, higher viscosity and lower flammability.
3. Alkanes have only single carbon-carbon bonds; alkenes contain at least one carbon-carbon double bond.
4. Cracking breaks large hydrocarbons into smaller alkanes and alkenes.
5. Complete combustion produces carbon dioxide and water: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$.
6. Hydrocarbons are important because they are used as fuels and as raw materials for making many products.

Topic 8 – Chemical Analysis

5.8.1.1 Pure substances

1. A substance containing only one element or compound.
2. A pure substance contains only one substance; a mixture contains two or more substances not chemically bonded.
3. A pure substance has a fixed melting point; impurities cause melting over a range of temperatures.
4. A pure substance has a fixed boiling point; impurities change the boiling point.
5. Pure substances contain particles that are all the same, so they change state at a fixed temperature.
6. It suggests the substance is impure because impurities cause a range of melting temperatures.

5.8.1.2 Formulations

1. A mixture designed as a useful product with specific amounts of components.
2. To ensure the product has the correct properties and works effectively.
3. Paint and medicines.

4. Medicines contain carefully measured ingredients to provide the correct effect and dosage.
5. Formulations contain carefully selected amounts of several substances, while simple mixtures do not have controlled compositions.
6. Paint is a formulation because pigments, solvents and other chemicals are combined in specific quantities to produce the desired properties.

5.8.1.3 Chromatography

1. To separate and identify substances in a mixture.
2. Stationary phase: the paper. Mobile phase: the solvent.
3. Different substances travel at different speeds depending on their solubility and attraction to the paper.
4. The Rf value shows the distance travelled by a substance compared with the solvent front.
5. Rf = distance travelled by substance ÷ distance travelled by solvent
   Rf = 4 ÷ 8
   Rf = 0.5
6. It shows the compound contains only one substance.

5.8.2.1 Test for hydrogen

1. Place a burning splint at the gas sample.
2. The gas burns with a squeaky pop.
3. A squeaky pop confirms hydrogen.
4. To prevent the gas mixture escaping and to make the test safer.
5. Hydrogen.
6. The gas is hydrogen.

5.8.2.2 Test for oxygen

1. Place a glowing splint into the gas.
2. The glowing splint relights.
3. Relighting confirms oxygen.
4. Oxygen supports combustion.
5. Oxygen.
6. The gas is oxygen.

5.8.2.3 Test for carbon dioxide

1. Bubble the gas through limewater.
2. Limewater.

3. Limewater turns cloudy or milky.
4. A cloudy appearance confirms carbon dioxide.
5. Carbon dioxide.
6. The gas is carbon dioxide.

5.8.2.4 Test for chlorine

1. Place damp litmus paper into the gas.
2. Damp litmus paper is bleached and turns white.
3. Bleaching confirms chlorine.
4. Water allows chlorine to react and produce acidic substances that bleach the paper.
5. Chlorine.
6. The gas is chlorine.

Topic 8 Review

1. Melting point and boiling point data can identify pure substances because pure substances have fixed temperatures.
2. Chromatography separates substances because different substances move different distances through a stationary phase.
3. Rf value = distance travelled by substance ÷ distance travelled by solvent.
4. A formulation is a carefully designed mixture, while a pure substance contains only one substance.
5. Hydrogen gives a squeaky pop, oxygen relights a glowing splint, carbon dioxide turns limewater cloudy and chlorine bleaches damp litmus paper.
6. Chemical analysis helps identify substances in areas such as forensic science, medicine and industry.

Topic 9 – Chemistry of the Atmosphere

5.9.1.1 The proportions of different gases in the atmosphere

1. Approximately 78%.
2. Approximately 21%.
3. Argon and carbon dioxide.
4. Approximately 4:1 nitrogen to oxygen.

5. The proportions have changed due to processes such as photosynthesis, volcanic activity and human activity.
6. It is similar because Earth's atmosphere is mainly nitrogen and oxygen in these proportions.

5.9.1.2 The Earth's early atmosphere

1. Carbon dioxide.
2. Volcanoes released gases including carbon dioxide, water vapour and nitrogen.
3. It condensed to form oceans.
4. There is limited direct evidence because the early atmosphere existed billions of years ago.
5. Carbon dioxide dissolved in oceans and became locked into sedimentary rocks.
6. The early atmosphere had more carbon dioxide and little or no oxygen compared with today.

5.9.1.3 How oxygen increased

1. Plants and algae.
2. Photosynthesis.
3. Carbon dioxide + water → glucose + oxygen.
4. Photosynthesis increased the amount of oxygen in the atmosphere.
5. More oxygen allowed aerobic organisms to evolve.
6. Plants and algae remove carbon dioxide and release oxygen through photosynthesis.

5.9.1.4 How carbon dioxide decreased

1. Plants and algae removed carbon dioxide through photosynthesis.
2. Photosynthesis.
3. Carbon dioxide dissolved in oceans and formed carbonate compounds in sedimentary rocks.
4. Carbon dioxide was stored in fossil fuels formed from ancient organisms.
5. Limestone formed from carbon-containing sediments; coal, crude oil and natural gas formed from buried remains of organisms.
6. Carbon dioxide levels decreased as it was removed by photosynthesis and stored in rocks and fossil fuels.

5.9.2.1 Greenhouse gases

1. Gases that absorb infrared radiation and contribute to the greenhouse effect.
2. Carbon dioxide, methane and water vapour.
3. They absorb and re-radiate infrared radiation, keeping the Earth warm.
4. They absorb the energy and transfer it back towards the Earth.
5. It keeps the Earth warm enough for life.
6. Short wavelength radiation comes from the Sun; long wavelength radiation is emitted by the Earth.

5.9.2.2 Human activities which contribute to an increase in greenhouse gases

1. Burning fossil fuels and deforestation.
2. Farming and decomposition of waste.
3. Burning fossil fuels releases carbon that was stored underground as carbon dioxide.
4. Livestock and farming processes release methane.
5. Peer-reviewed evidence is checked by other scientists to improve reliability.
6. Predictions have uncertainty because future human activity and natural processes are difficult to predict.

5.9.2.3 Global climate change

1. Long-term changes in the average temperature and climate of Earth.
2. Increased greenhouse gases from human activities.
3. Rising sea levels, changing weather patterns, loss of habitats and melting ice.
4. Warmer temperatures cause ice to melt and oceans to expand, increasing sea levels.
5. Climate systems are complex and affected by many variables.
6. It affects ecosystems, agriculture, resources, health and human populations.

5.9.2.4 The carbon footprint and its reduction

1. The total amount of greenhouse gases released by a person, organisation or activity.
2. Carbon dioxide and other greenhouse gases such as methane.
3. Using renewable energy and reducing fossil fuel use.
4. Reducing livestock farming and improving waste management.

5. Many activities depend on fossil fuels and changing them can be difficult.
6. Renewable energy produces little or no carbon dioxide during generation, reducing emissions.

5.9.3.1 Atmospheric pollutants from fuels

1. Carbon monoxide, particulates, sulfur dioxide and nitrogen oxides.
2. It forms during incomplete combustion of fuels.
3. They form when fuels burn incompletely.
4. Sulfur dioxide.
5. Limited oxygen causes incomplete combustion.
6. Fuels contain impurities and combustion conditions vary, producing different pollutants.

5.9.3.2 Properties and effects of atmospheric pollutants

1. Carbon monoxide reduces the blood's ability to carry oxygen.
2. It is colourless and odourless.
3. Sulfur dioxide causes acid rain and respiratory problems.
4. Nitrogen oxides contribute to acid rain and breathing problems.
5. Particulates can damage the lungs and cause health problems.
6. Pollutants can cause environmental problems such as acid rain, global dimming and damage to ecosystems.

Topic 9 Review

1. Earth's atmosphere changed from being rich in carbon dioxide to mainly nitrogen and oxygen.
2. Photosynthesis removed carbon dioxide and released oxygen.
3. Greenhouse gases absorb infrared radiation and keep Earth warm.
4. Human activities such as burning fossil fuels and farming increase greenhouse gas levels.
5. Major pollutants include carbon monoxide, sulfur dioxide, nitrogen oxides and particulates.
6. Reducing emissions, using renewable energy and protecting natural carbon sinks can reduce atmospheric impacts.

Topic 10 – Using Resources

5.10.1.1 Using the Earth's resources and sustainable development

1. Natural resources are materials obtained from the Earth that humans use.
2. Metals, crude oil, water and rocks.
3. A finite resource is limited and will eventually run out; a renewable resource can be replaced naturally.
4. Wood can be replaced by synthetic materials.
5. Sustainable development means meeting the needs of the present without preventing future generations meeting their needs.
6. Reducing finite resource use helps conserve supplies and protects resources for future generations.

5.10.1.2 Potable water

1. Water that is safe to drink.
2. It contains dissolved substances and minerals, so it is not chemically pure.
3. It must contain no harmful microorganisms and have safe levels of dissolved substances.
4. Filter the water to remove solids and sterilise it to kill microorganisms.
5. Chlorine, ozone and ultraviolet light.
6. Desalination removes salt from seawater but requires large amounts of energy.

5.10.1.3 Waste water treatment

1. To remove harmful substances before releasing water back into the environment.
2. Organic matter, solids and microorganisms.
3. Large solids are removed by screening and grit is removed by settling.
4. Sewage sludge.
5. Anaerobic bacteria break down sludge and produce methane.
6. Aerobic biological treatment uses microorganisms to break down organic waste.

5.10.1.4 Alternative methods of extracting metals (HT Only)

1. Traditional extraction methods may use large amounts of energy and damage the environment.
2. Phytomining is using plants to absorb metal compounds from the ground.
3. Plants absorb metal ions and are then burned to produce ash containing metal compounds.
4. Bioleaching uses bacteria to produce solutions containing metal ions.
5. Copper can be extracted from the solution by displacement or electrolysis.
6. Advantage: uses less energy. Disadvantage: slower than traditional methods.

5.10.2.1 Life cycle assessment

1. A life cycle assessment evaluates the environmental impacts of a product throughout its life.
2. Raw material extraction, manufacture, transport, use and disposal.
3. Energy and resource use can be measured more easily than effects such as pollution.
4. They involve choices about which factors are included and how important they are.
5. Results could be selected to support a particular viewpoint.
6. An LCA compares the resources used, energy needed and environmental impacts of plastic and paper bags throughout their lifetimes.

5.10.2.2 Ways of reducing the use of resources

1. Reduce, reuse and recycle.
2. Recycling reduces the need for extracting new raw materials.
3. Glass bottles can be cleaned and used again.
4. Metals are collected, melted and reshaped into new products.
5. Recycling metals reduces the need to mine new ores.
6. Reducing, reusing and recycling conserve resources and reduce environmental damage.

Topic 10 Review

1. Sustainable development ensures resources are used without preventing future generations from accessing them.
2. Fresh water is treated by filtration and sterilisation to make it potable.

3. Waste water treatment removes solids, organic matter and harmful microorganisms.
4. Traditional extraction uses methods such as carbon reduction or electrolysis; phytomining and bioleaching use biological processes.
5. Life cycle assessments compare environmental impacts at different stages of a product's life.
6. Reducing, reusing and recycling resources helps conserve Earth's limited supplies and reduce pollution.