Paper 1:

Topic 1 – Atomic Structure and the Periodic Table

4.1.1.1 Atoms, Elements and Compounds

1. An atom is the smallest part of an element that can exist and still have the properties of that element.
2. An element is a substance made of only one type of atom.
3. A compound is a substance made from two or more different elements chemically bonded together.
4. An element contains only one type of atom, whereas a compound contains different elements chemically bonded together.
5. Compounds can be separated into their elements by chemical reactions, such as electrolysis.
6. A chemical reaction is needed to separate a compound because chemical bonds must be broken, whereas mixtures are only physically combined and can be separated by physical methods.

4.1.1.2 Mixtures

1. A mixture is two or more substances mixed together without chemical bonding.
2. Yes, substances in a mixture keep their own chemical properties.
3. Filtration, crystallisation, distillation and chromatography.
4. Chromatography is used to separate dyes in ink.
5. Crystallisation is used to separate a soluble solid from a solution.
6. Separating a mixture does not involve a chemical reaction because no chemical bonds are broken or formed.

4.1.1.3 The Development of the Model of the Atom

1. The accepted model before the electron was discovered was the solid sphere model (Dalton's model).
2. The plum pudding model suggested that atoms were a ball of positive charge with negatively charged electrons embedded in it.
3. Rutherford's alpha scattering experiment led to the nuclear model of the atom.
4. Niels Bohr suggested that electrons exist in fixed energy levels or shells around the nucleus.
5. James Chadwick discovered the neutron.

6. Rutherford's experiment showed that most of the atom was empty space, with a small dense positive nucleus, causing the atomic model to change.

4.1.1.4 Relative Electrical Charges of Subatomic Particles

1. A proton has a relative charge of +1.
2. A neutron has a relative charge of 0.
3. An electron has a relative charge of -1.
4. The atomic number is the number of protons in an atom.
5. An atom has no overall charge because it has equal numbers of protons and electrons.
6. Atoms of different elements have different chemical properties because they have different numbers of protons and different electron arrangements.

4.1.1.5 Size and Mass of Atoms

1. The approximate radius of an atom is about 0.1 nanometres (1×10^{-10} m).
2. Almost all of an atom's mass is found in the nucleus.
3. The mass number is the total number of protons and neutrons in an atom.
4. An isotope is an atom of the same element with the same number of protons but a different number of neutrons.
5. Number of neutrons = mass number – atomic number = 35 – 17 = 18 neutrons.
6. Isotopes have similar chemical properties because they have the same number of electrons and the same electron arrangement.

4.1.1.6 Relative Atomic Mass

1. Relative atomic mass is the weighted average mass of atoms of an element compared with one-twelfth of the mass of carbon-12.
2. Relative atomic mass is often not a whole number because it is an average of different isotopes.
3. Relative atomic mass takes into account the masses and abundances of isotopes.
4. The particles that differ in isotopes are neutrons.
5. Relative atomic mass is calculated by multiplying each isotope's mass by its abundance, adding the results, then dividing by the total abundance.
6. Chlorine has a relative atomic mass of 35.5 because it is a mixture of chlorine-35 and chlorine-37 isotopes in different abundances.

4.1.1.7 Electronic Structure

1. Electrons are found in shells around the nucleus.
2. The first shell fills first.
3. Sodium has the electronic structure 2,8,1.
4. Oxygen has the electronic structure 2,6.
5. Calcium has the electronic structure 2,8,8,2.
6. The number of occupied electron shells determines the period, and the number of electrons in the outer shell determines the group.

4.1.2.1 The Periodic Table

1. Elements are arranged in order of increasing atomic number.
2. A group is a vertical column of elements in the periodic table.
3. A period is a horizontal row of elements in the periodic table.
4. Elements in the same group have similar chemical properties because they have the same number of electrons in their outer shell.
5. Reactivity can be predicted by observing trends down a group.
6. An element's position is related to its electron arrangement because the number of shells determines the period and outer electrons determine the group.

4.1.2.2 Development of the Periodic Table

1. Dmitri Mendeleev developed the first successful periodic table.
2. Mendeleev arranged elements by increasing atomic weight and grouped elements with similar properties.
3. Mendeleev left gaps because he predicted that undiscovered elements would fit there.
4. The gaps were later filled when new elements were discovered with properties close to Mendeleev's predictions.
5. Arranging elements by atomic weight was sometimes incorrect because isotopes have different masses.
6. Mendeleev's table was accepted because his predictions for undiscovered elements were accurate.

4.1.2.3 Metals and Non-metals

1. Metals are found on the left and centre of the periodic table.
2. Non-metals are found on the right side of the periodic table.
3. Metals form positive ions.
4. Non-metals form negative ions.
5. Metals are good conductors of electricity and are malleable.
6. Metals and non-metals have different chemical properties because they have different electron arrangements and form different ions.

4.1.2.4 Group 0

1. Group 0 elements are called noble gases.
2. Noble gases are unreactive because they have a full outer electron shell.
3. Helium has only two electrons in its outer shell.
4. Boiling points increase down Group 0.
5. Helium has the lowest boiling point.
6. Boiling points increase down Group 0 because atoms become larger and have stronger intermolecular forces.

4.1.2.5 Group 1

1. Group 1 elements are called alkali metals.
2. Group 1 elements have one electron in their outer shell.
3. Reactivity increases down Group 1.
4. Hydrogen gas is produced when Group 1 metals react with water.
5. Sodium reacts with chlorine to form sodium chloride (NaCl).
6. Group 1 metals become more reactive down the group because the outer electron is further from the nucleus and easier to lose.

4.1.2.6 Group 7

1. Group 7 elements are called halogens.
2. Group 7 elements have seven electrons in their outer shell.
3. Reactivity decreases down Group 7.
4. A displacement reaction occurs when a more reactive element replaces a less reactive element in a compound.
5. Chlorine can displace iodine from potassium iodide solution.
6. Chlorine is more reactive than bromine because chlorine atoms gain an electron more easily due to their smaller size.

4.1.3.1 Comparison with Group 1 Elements (Triple Only)

1. Transition elements are metals found in the central block of the periodic table.
2. Transition metals are less reactive than Group 1 metals.
3. Transition metals have higher melting points than Group 1 metals.
4. Iron is a transition metal used in everyday life.
5. Transition metals are less reactive and harder than Group 1 metals.
6. Transition metals are suitable for structures and tools because they are strong, hard and have high melting points.

4.1.3.2 Typical Properties (Triple Only)

1. A catalyst is a substance that increases the rate of reaction without being used up.
2. A typical property of transition metals is that they form coloured compounds.
3. Many transition metal compounds are coloured.
4. Yes, transition metals can form ions with different charges.
5. Iron is used as a catalyst in the Haber process.
6. Transition metals are widely used in industry because they are strong, have useful properties and can act as catalysts.

Topic 1 Review

1. An element contains one type of atom, while a compound contains two or more elements chemically bonded together.
2. The particle that determines the identity of an element is the proton.
3. Proton = +1, neutron = 0, electron = -1.
4. Isotopes have different mass numbers because they contain different numbers of neutrons.
5. Group 1 elements are reactive metals that become more reactive down the group; Group 7 elements are reactive non-metals that become less reactive down the group.
6. Electron arrangement determines an element's position because outer electrons determine its group and number of shells determine its period. It also determines reactivity because reactions involve gaining, losing or sharing electrons.

Topic 2 – Bonding, Structure and the Properties of Matter

4.2.1.1 Chemical Bonds

1. The three types of strong chemical bonding are ionic bonding, covalent bonding and metallic bonding.
2. Ionic bonding occurs between a metal and a non-metal.
3. Covalent bonding occurs between non-metals.
4. Metallic bonding is found in metals.
5. Chemical bonds are held together by strong electrostatic forces of attraction between particles.

6. Ionic bonding involves attraction between oppositely charged ions, covalent bonding involves sharing electrons, and metallic bonding involves attraction between positive metal ions and delocalised electrons.

4.2.1.2 Ionic Bonding

1. Electrons are transferred from one atom to another during ionic bonding.
2. A metal loses electrons to form a positive ion.
3. A non-metal gains electrons to form a negative ion.
4. Aluminium oxide forms when two aluminium atoms lose six total electrons (three each) and three oxygen atoms gain these electrons (2 each), forming Al^{3+} and O^{2-} ions in a giant ionic lattice.
5. Atoms form ions during ionic bonding to achieve a full outer electron shell and become more stable.
6. Sodium loses one electron to form Na^{+} and chlorine gains one electron to form Cl^{-}. The oppositely charged ions attract to form sodium chloride.

4.2.1.3 Ionic Compounds

1. Ionic compounds have a giant ionic lattice structure.
2. Ions are held together by strong electrostatic forces of attraction.
3. Ionic compounds have high melting points because strong forces between ions require lots of energy to overcome.
4. An empirical formula shows the simplest whole-number ratio of atoms or ions in a compound.
5. Molten ionic compounds conduct electricity because ions are free to move and carry charge.
6. Solid ionic compounds do not conduct electricity because the ions are fixed in place and cannot move.

4.2.1.4 Covalent Bonding

1. Electrons are shared between atoms during covalent bonding.
2. Covalent bonding occurs between non-metal elements.
3. Carbon dioxide has two carbon-oxygen double covalent bonds, with carbon sharing two pairs of electrons with each oxygen atom.
4. Diamond is an example of a giant covalent substance.
5. Water is an example of a substance made of small covalent molecules.
6. Covalent bonds are strong because they involve strong electrostatic attractions between the shared electrons and nuclei.

4.2.1.5 Metallic Bonding

1. Metals are made of positive ions arranged in a regular lattice surrounded by delocalised electrons.

2. Delocalised electrons are electrons that are free to move through the metal.
3. Calcium metal consists of Ca^{2+} ions surrounded by delocalised electrons.
4. Electrons can move through a metal because they are delocalised.
5. Metallic bonding is the strong attraction between positive metal ions and delocalised electrons.
6. Metals conduct electricity because delocalised electrons carry electrical charge through the structure.

4.2.2.1 The Three States of Matter

1. The three states of matter are solid, liquid and gas.
2. Melting occurs when a solid changes into a liquid because particles gain enough energy to move past each other.
3. Condensation occurs when a gas changes into a liquid because particles lose energy.
4. When a substance is heated, particles gain kinetic energy and move faster.
5. Different substances have different melting points because the forces between particles are different strengths.
6. More energy is needed to melt some substances because stronger forces between particles require more energy to overcome.

4.2.2.2 State Symbols

1. A solid is represented by (s).
2. A liquid is represented by (l).
3. A gas is represented by (g).
4. (aq) means dissolved in water (aqueous solution).
5. State symbols show the physical states of reactants and products in chemical equations.
6. $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

4.2.2.3 Properties of Ionic Compounds

1. Ionic compounds have high melting points because strong ionic bonds require lots of energy to break.
2. Ionic compounds have high boiling points because strong forces between ions require large amounts of energy to overcome.
3. Ionic compounds conduct electricity when molten or dissolved in water.
4. Molten ionic compounds conduct electricity because ions are free to move.
5. Solid ionic compounds do not conduct electricity because ions cannot move.
6. The giant ionic lattice structure gives ionic compounds high melting points and allows conduction only when ions are mobile.

4.2.2.4 Properties of Small Molecules

1. Small molecules have weak intermolecular forces between molecules.
2. Intermolecular forces are weaker than covalent bonds.
3. Small molecular substances have low melting points because little energy is needed to overcome intermolecular forces.
4. Small molecular substances do not conduct electricity because they have no charged particles that can move.
5. Larger molecules have stronger intermolecular forces and higher boiling points.
6. Covalent bonds are not broken when a molecular substance melts; only intermolecular forces are overcome.

4.2.2.5 Polymers

1. A polymer is a long chain molecule made from many repeating units.
2. Atoms within a polymer are joined by strong covalent bonds.
3. Polymers are solids at room temperature because their long chains have strong forces between them.
4. Polymers contain very large molecules made from repeating units.
5. Poly(ethene) is an example of a polymer.
6. Polymers generally have higher melting points than small molecular substances because they have stronger forces between their long chains.

4.2.2.6 Giant Covalent Structures

1. A giant covalent structure is a structure where atoms are joined by many strong covalent bonds in a large network.
2. Diamond, graphite and graphene are giant covalent substances.
3. Giant covalent substances have very high melting points because many strong covalent bonds must be broken.
4. Covalent bonds must be broken to melt a giant covalent structure.
5. Graphite conducts electricity.
6. Diamond and graphite have different properties because their atoms are arranged differently.

4.2.2.7 Properties of Metals and Alloys

1. Metals have high melting points because strong metallic bonds require lots of energy to break.
2. Pure metals can be bent easily because layers of atoms can slide over each other.
3. An alloy is a mixture of a metal with one or more other elements.
4. Alloys are harder than pure metals because different-sized atoms disrupt the regular layers and prevent them sliding.

5. Brass is an example of an alloy.
6. Alloys are often used instead of pure metals because they are stronger and more useful.

4.2.2.8 Metals as Conductors

1. Metals conduct electricity because they contain delocalised electrons.
2. Delocalised electrons carry electrical charge through a metal.
3. Metals conduct thermal energy well because delocalised electrons transfer energy quickly.
4. Delocalised electrons are free to move through the metal structure.
5. Metals are useful for electrical wiring because they are good electrical conductors and are ductile.
6. Metallic bonding allows conduction because electrons are free to move and transfer charge and energy.

4.2.3.1 Diamond

1. Each carbon atom forms four covalent bonds in diamond.
2. Diamond has a giant covalent structure.
3. Diamond is very hard because each carbon atom is strongly bonded to four others.
4. Diamond has a very high melting point because many strong covalent bonds must be broken.
5. Diamond does not conduct electricity.
6. Diamond does not conduct electricity because all electrons are held in covalent bonds and none are delocalised.

4.2.3.2 Graphite

1. Each carbon atom forms three covalent bonds in graphite.
2. Graphite consists of layers of carbon atoms arranged in hexagonal shapes.
3. Graphite conducts electricity because it has delocalised electrons.
4. Layers in graphite slide over each other because weak forces exist between the layers.
5. Graphite is used as a lubricant.
6. Graphite is soft because layers slide easily, whereas diamond is hard because of its strong three-dimensional bonding.

4.2.3.3 Graphene and Fullerenes

1. Graphene is a single layer of carbon atoms arranged in a hexagonal structure.

2. A fullerene is a molecule made from carbon atoms arranged in a hollow shape.
3. Buckminsterfullerene (C_{60}) has a spherical shape.
4. Carbon nanotubes are cylindrical structures made from carbon atoms.
5. Graphene can be used in electronics or sensors.
6. Graphene is useful in electronics because it conducts electricity and is very thin and strong.

4.2.4.1 Sizes of Particles and Their Properties (Triple Only)

1. Nanoparticles have sizes between 1 and 100 nanometres.
2. Surface area to volume ratio increases as particles get smaller.
3. Nanoparticles have different properties because they have a much larger surface area compared with their volume.
4. Nano- means one billionth (10^{-9}).
5. Smaller quantities of nanoparticles are often needed because they have a large surface area to volume ratio.
6. Nanoparticles are more reactive because more particles are exposed at the surface.

4.2.4.2 Uses of Nanoparticles (Triple Only)

1. Nanoparticles can be used in medicine for targeted drug delivery.
2. Nanoparticles can be used in cosmetics such as sunscreen.
3. Nanoparticles can be used industrially as catalysts.
4. An advantage is that small amounts can be effective because of their large surface area.
5. A possible risk is that nanoparticles may enter cells and cause unknown health effects.
6. Scientists continue researching nanoparticles because they have many potential applications but their risks must be understood.

Topic 2 Review

1. The three types of strong chemical bonding are ionic, covalent and metallic.
2. Ionic bonding involves electron transfer and attraction between ions; covalent bonding involves sharing electrons between atoms.
3. Ionic compounds only conduct electricity when molten or dissolved because ions must be free to move.
4. Diamond has a three-dimensional giant covalent structure and is hard; graphite has layers with delocalised electrons and is soft.
5. Alloys are harder than pure metals because different atoms disrupt the layers and stop them sliding.

6. Bonding and structure determine physical properties because the type and strength of forces between particles affect melting point, conductivity, hardness and flexibility.

Topic 3 – Quantitative Chemistry

4.3.1.1 Conservation of Mass and Balanced Chemical Equations

1. The law of conservation of mass states that mass is not created or destroyed during a chemical reaction.
2. Chemical equations must be balanced because the same number of each type of atom must be present on both sides of the equation.
3. A coefficient shows the number of molecules or moles of a substance.
4. A coefficient is a number placed in front of a formula and changes the amount of a substance; a subscript is part of the formula and shows the number of atoms in one molecule.
5. $2H_2 + O_2 \rightarrow 2H_2O$.
6. Mass stays the same because atoms are rearranged but not created or destroyed.

4.3.1.2 Relative Formula Mass

1. Relative formula mass (Mr) is the sum of the relative atomic masses of all atoms in a formula.
2. Relative formula mass is calculated by adding together the relative atomic masses of all atoms in the formula.
3. The relative atomic masses of all elements in the compound are needed.
4. Mr of $H_2O = (2 \times 1) + 16 = 18$.
5. Percentage by mass of oxygen in water = $(16 \div 18) \times 100 = 88.9\%$.
6. Relative formula mass is useful because it allows chemists to calculate reacting amounts and masses.

4.3.1.3 Mass Changes When a Reactant or Product is a Gas

1. The mass can increase when magnesium burns because magnesium reacts with oxygen from the air.
2. The mass can decrease when a metal carbonate is heated because carbon dioxide gas escapes.
3. Carbon dioxide is produced during the thermal decomposition of a metal carbonate.
4. In a closed system, the total mass remains constant.

5. Some reactions appear to lose mass because gases escape into the surroundings.
6. Mass is conserved because gases still contain the same atoms; they have just left the reaction container.

4.3.1.4 Chemical Measurements

1. Uncertainty is the amount by which a measurement may vary from the true value.
2. The mean is the average of a set of results.
3. The range is the difference between the highest and lowest results.
4. Repeat measurements are carried out to improve reliability and identify anomalies.
5. The uncertainty of repeated measurements can be estimated using half the range.
6. Measurements always have some uncertainty because instruments have limits of accuracy and human measurements vary.

4.3.2.1 Moles (Higher Tier)

1. The unit used to measure the amount of substance is the mole (mol).
2. One mole is the amount of substance containing 6.02×10^{23} particles.
3. The Avogadro constant is 6.02×10^{23} particles per mole.
4. Moles = mass ÷ relative formula mass.
5. Moles in 36 g of water = 36 ÷ 18 = 2 mol.
6. Mass of 0.25 mol NaCl = 0.25 × 58.5 = 14.6 g.

4.3.2.2 Amounts of Substances in Equations (Higher Tier)

1. A balanced equation shows the ratio of moles of reactants and products.
2. Equations must be balanced before calculations because mole ratios depend on the coefficients.
3. One mole of magnesium reacts with hydrochloric acid to produce one mole of hydrogen.
4. The first step is to calculate the number of moles of the known substance.
5. 24 g magnesium = 1 mol Mg, so it produces 1 mol MgO = 40 g magnesium oxide.
6. 135 g aluminium = 5 mol Al, producing 2.5 mol Al_2O_3 = 255 g aluminium oxide.

4.3.2.3 Using Moles to Balance Equations (Higher Tier)

1. Moles are used because equations show ratios of particles and reacting amounts.
2. Masses must be converted into moles before balancing by calculation.

3. A mole ratio is used to balance equations.
4. Photosynthesis:
   $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$.
5. Methane combustion:
   $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$.
6. Magnesium reaction:
   $2Mg + O_2 \rightarrow 2MgO$.

4.3.2.4 Limiting Reactants (Higher Tier)

1. A limiting reactant is the reactant that is completely used up first in a reaction.
2. One reactant is often used in excess to ensure another reactant reacts completely.
3. When the limiting reactant is used up, the reaction stops.
4. The limiting reactant determines the amount of product formed.
5. The limiting reactant is identified by comparing mole ratios of reactants.
6. The limiting reactant controls the yield because no more product can form once it is exhausted.

4.3.2.5 Concentration of Solutions

1. Concentration is the amount of solute dissolved in a given volume of solution.
2. Concentration in g/dm³ is measured in grams per cubic decimetre.
3. Concentration = mass ÷ volume.
4. Concentration = 20 ÷ 2 = 10 g/dm³.
5. Mass = concentration × volume = 10 × 0.5 = 5 g.
6. Increasing the amount of solute increases the concentration if the volume stays the same.

4.3.3.1 Percentage Yield (Triple Only)

1. Percentage yield compares the actual amount of product made with the maximum possible amount.
2. Percentage yield = (actual yield ÷ theoretical yield) × 100.
3. Reasons include incomplete reactions, product loss during separation or side reactions.
4. The theoretical yield is the maximum amount of product predicted from calculations.
5. Percentage yield = (8 ÷ 10) × 100 = 80%.
6. Reactions rarely produce 100% yield because products can be lost and reactions may not go to completion.

4.3.3.2 Atom Economy (Triple Only)

1. Atom economy measures how much of the reactant atoms are converted into useful product.
2. A high atom economy reduces waste and makes reactions more sustainable.
3. Atom economy = (Mr of desired product ÷ total Mr of products) × 100.
4. Iron production:
   $2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$
   Atom economy = (4 × 56) ÷ [(4 × 56) + (3 × 44)] × 100 = 63.6%.
5. Magnesium oxide production:
   $2Mg + O_2 \rightarrow 2MgO$
   Atom economy = (2 × 40) ÷ (2 × 40) × 100 = 100%.
6. High atom economy reactions are more sustainable because they produce less waste.

4.3.4 Using Concentrations of Solutions in mol/dm³ (Triple Higher Tier)

1. Concentration in moles is measured in mol/dm³.
2. Concentration = moles ÷ volume.
3. Volume must be measured in dm³.
4. Moles = concentration × volume = 2 × 0.5 = 1 mol.
5. Concentration = 1 ÷ 0.25 = 4 mol/dm³.
6. Titration determines concentration by reacting a known volume of one solution with another of known concentration.

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

1. One mole of gas occupies 24 dm³ at room temperature and pressure.
2. The conditions assumed are room temperature and pressure.
3. Gas volume = moles × 24 dm³.
4. Volume of 2 moles = 2 × 24 = 48 dm³.
5. Moles in 48 dm³ = 48 ÷ 24 = 2 mol.
6. Equal numbers of moles of gases occupy equal volumes because they contain the same number of particles under the same conditions.

Topic 3 Review

1. The law of conservation of mass states that mass is conserved during chemical reactions.
2. Moles = mass ÷ relative formula mass.
3. Balanced equations are needed because they show the correct ratios of reacting substances.
4. Percentage yield may be less than 100% because of product loss and incomplete reactions.

5. Percentage yield measures the amount of product obtained compared with the maximum possible amount; atom economy measures how much of the reactant atoms become useful products.
6. Moles are used to calculate reacting masses and gas volumes by using mole ratios from balanced equations.

Topic 4 – Chemical Changes

4.4.1.1 Metal Oxides

1. A metal reacts with oxygen to produce a metal oxide.
2. Oxidation is the gain of oxygen.
3. Reduction is the loss of oxygen.
4. Reactions between metals and oxygen are called oxidation reactions because the metal gains oxygen.
5. Copper oxide is reduced when it is heated with carbon.
6. Oxidation and reduction occur together because when one substance gains oxygen, another substance must lose oxygen.

4.4.1.2 The Reactivity Series

1. The reactivity series is a list of metals arranged in order of their reactivity, from most reactive to least reactive.
2. Magnesium is more reactive than copper.
3. A more reactive metal displaces a less reactive metal from its compound.
4. The non-metals included in the reactivity series are carbon and hydrogen.
5. Potassium, iron, zinc, copper.
6. Potassium reacts more vigorously with water than magnesium because potassium is higher in the reactivity series and loses electrons more easily.

4.4.1.3 Extraction of Metals and Reduction

1. Unreactive metals such as gold are found as the metal itself in the Earth's crust.
2. Metals below carbon in the reactivity series can be extracted from their oxides using carbon.
3. Oxygen is removed from the metal oxide during reduction.
4. Carbon is used because it is more reactive than some metals and can remove oxygen from their oxides.
5. Aluminium cannot be extracted using carbon because it is more reactive than carbon.
6. The extraction method depends on the reactivity series because more reactive metals require electrolysis, while less reactive metals can be extracted by reduction with carbon.

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

1. Oxidation is the loss of electrons.
2. Reduction is the gain of electrons.
3. Zinc is oxidised in the reaction $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$.
4. Copper ions (Cu^{2+}) are reduced in the reaction $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$.
5. $Mg + Cu^{2+} \rightarrow Mg^{2+} + Cu$.
6. Displacement reactions are redox reactions because electrons are transferred between atoms and ions.

4.4.2.1 Reactions of Acids with Metals

1. Hydrogen gas is produced when an acid reacts with a metal.
2. Magnesium reacts with hydrochloric acid to produce magnesium chloride.
3. Metal + acid → salt + hydrogen.
4. Zinc + sulfuric acid → zinc sulfate + hydrogen.
5. Magnesium reacts more vigorously with dilute hydrochloric acid than iron.
6. Acid-metal reactions are redox reactions because the metal loses electrons and hydrogen ions gain electrons.

4.4.2.2 Neutralisation of Acids and Salt Production

1. An acid reacts with an alkali to produce salt and water.
2. An acid reacts with a metal carbonate to produce salt, water and carbon dioxide.
3. Sulfuric acid is used to produce sulfate salts.
4. Nitric acid is used to produce nitrate salts.
5. Nitric acid + potassium hydroxide → potassium nitrate + water.
6. The acid used determines the salt name because the acid's negative ion becomes part of the salt.

4.4.2.3 Soluble Salts (Required Practical 1)

1. Excess insoluble solid is added to ensure all the acid reacts.
2. Excess solid is filtered out to remove unreacted material.
3. The salt solution is heated to evaporate some water and concentrate the solution.
4. The solution should not be heated to complete dryness because crystals may decompose or become impure.
5. Warm the acid → add excess solid → filtration → crystallisation.
6. Add excess insoluble base to warm acid, filter to remove excess solid, evaporate some water, allow crystals to form, then filter and dry the crystals.

4.4.2.4 The pH Scale and Neutralisation

1. A neutral solution has a pH of 7.
2. Hydrogen ions (H^+) make a solution acidic.
3. Hydroxide ions (OH^-) make a solution alkaline.
4. Hydrogen ions react with hydroxide ions to produce water.
5. Universal indicator or a pH probe can be used to measure approximate pH.
6. Hydrochloric acid and sodium hydroxide produce a neutral solution because H^+ ions react with OH^- ions to form water.

4.4.2.5 Titrations (Required Practical 2 - Chemistry Only)

1. The purpose of a titration is to find the exact volume of one solution needed to react with another.
2. A burette is used to deliver acid accurately.
3. The end point is when the indicator changes colour.
4. An indicator is added to show when neutralisation is complete.
5. Mean titre = (24.60 + 24.70 + 24.65) ÷ 3 = 24.65 cm^3.
6. Concordant titres are used because they are close together and give a reliable mean result.

4.4.2.6 Strong and Weak Acids (Higher Tier)

1. A strong acid completely ionises in water.
2. A weak acid only partially ionises in water.
3. A strong acid has a lower pH at the same concentration.
4. A concentrated acid contains a large amount of acid per volume; a dilute acid contains less acid per volume.
5. Hydrogen ion concentration increases by 100 times when pH decreases from 5 to 3.
6. A dilute hydrochloric acid can be stronger because it completely ionises, while concentrated ethanoic acid only partially ionises.

4.4.3.1 The Process of Electrolysis

1. An electrolyte is a substance containing ions that can move and conduct electricity.
2. The cathode is the negative electrode.
3. Positive ions move towards the cathode.
4. Negative ions move towards the anode.
5. Ionic compounds must be molten or dissolved so ions are free to move.
6. Electrolysis produces elements because ions gain or lose electrons at the electrodes.

4.4.3.2 Electrolysis of Molten Ionic Compounds

1. Lead is produced at the cathode during electrolysis of molten lead bromide.
2. Bromine is produced at the anode during electrolysis of molten lead bromide.
3. Inert electrodes are used because they do not react with the products.
4. Molten zinc chloride produces zinc at the cathode and chlorine at the anode.
5. Chlorine (Cl_2) is produced at the anode during molten sodium chloride electrolysis.
6. Molten ionic compounds conduct electricity because ions are free to move and carry charge.

4.4.3.3 Using Electrolysis to Extract Metals

1. Metals above carbon in the reactivity series are extracted using electrolysis.
2. Aluminium is extracted by electrolysis because it is more reactive than carbon.
3. Cryolite lowers the melting point of aluminium oxide.
4. Carbon anodes are replaced because they react with oxygen to form carbon dioxide.
5. Aluminium is produced at the cathode.
6. Extracting aluminium is expensive because electrolysis requires large amounts of electricity.

4.4.3.4 Electrolysis of Aqueous Solutions (Required Practical 3)

1. Copper is produced at the cathode when aqueous copper(II) sulfate is electrolysed using inert electrodes.
2. Chlorine gas is produced at the anode when aqueous sodium chloride is electrolysed.
3. Hydrogen is produced at the cathode if the metal is more reactive than hydrogen.
4. Oxygen is usually produced at the anode if no halide ions are present.
5. Aqueous potassium chloride produces hydrogen at the cathode and chlorine at the anode.
6. The reactivity series helps predict whether a metal or hydrogen will be produced at the cathode.

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

1. Positive ions gain electrons at the cathode.
2. Negative ions lose electrons at the anode.
3. $2H^+ + 2e^- \rightarrow H_2$.

4. $2Cl^- \rightarrow Cl_2 + 2e^-$.
5. $Cu^{2+} + 2e^- \rightarrow Cu$.
6. Reduction occurs at the cathode because ions gain electrons, and oxidation occurs at the anode because ions lose electrons.

Topic 4 Review

1. Oxidation is gain of oxygen and reduction is loss of oxygen.
2. Metals above carbon cannot be extracted using carbon because they are more reactive and their oxides are too stable.
3. Acid + metal carbonate produces salt, water and carbon dioxide.
4. Add excess insoluble base to acid, filter, evaporate some water, crystallise and dry the soluble salt.
5. Molten ionic compounds conduct electricity because ions can move; solid ionic compounds do not because ions are fixed.
6. Molten sodium chloride produces sodium and chlorine, while aqueous sodium chloride produces hydrogen and chlorine because water affects which ions are discharged.

Topic 5 – Energy Changes

4.5.1.1 Energy Transfer During Exothermic and Endothermic Reactions

1. An exothermic reaction is a reaction that transfers energy to the surroundings, usually as heat.
2. An endothermic reaction is a reaction that takes in energy from the surroundings.
3. Combustion is an example of an exothermic reaction.
4. Thermal decomposition is an example of an endothermic reaction.
5. The temperature of the surroundings decreases during an endothermic reaction.
6. The products of an exothermic reaction have less energy than the reactants because more energy is released when bonds form than is needed to break bonds.

Required Practical 4: Investigating Temperature Changes

1. The temperature change of reacting solutions is measured.
2. A lid is placed on the reaction cup to reduce heat loss to the surroundings.
3. Reactants should be mixed quickly to ensure the reaction starts at the same time and the maximum temperature change is recorded.
4. The same volumes of reactants should be used to make a fair comparison.

5. Repeat measurements are carried out to improve reliability and calculate a mean result.
6. Reducing heat loss improves accuracy because more of the energy change is measured rather than being lost to the surroundings.

4.5.1.2 Reaction Profiles

1. Activation energy is the minimum energy needed for a reaction to occur.
2. A reaction profile shows the energy changes during a chemical reaction, including activation energy and overall energy change.
3. In an endothermic reaction profile, the products are at a higher energy level than the reactants, with a positive overall energy change.
4. In an exothermic reaction profile, the products are at a lower energy level than the reactants, with a negative overall energy change.
5. The highest point on a reaction profile represents the activated complex and the activation energy level.
6. All chemical reactions require activation energy because bonds must be broken before new bonds can form.

4.5.1.3 The Energy Change of Reactions (Higher Tier)

1. Energy is needed to break chemical bonds.
2. Energy is released when new chemical bonds are formed.
3. In an exothermic reaction, more energy is released during bond formation than is needed to break bonds.
4. In an endothermic reaction, more energy is needed to break bonds than is released during bond formation.
5. Energy change = energy needed to break bonds – energy released when bonds form = 820 – 950 = –130 kJ.
6. Bond breaking and bond making occur together because atoms must separate before rearranging into new substances.

4.5.2.1 Cells and Batteries (Chemistry Only)

1. Electricity is produced in a chemical cell by chemical reactions transferring electrons.
2. A simple chemical cell can be made using two different metals and an electrolyte.
3. A cell is a single unit that produces electricity; a battery is two or more cells connected together.
4. Non-rechargeable batteries stop working because the chemicals inside are used up.
5. Rechargeable batteries can be used many times because the chemical reactions can be reversed.

6. Connecting cells in series increases voltage because the voltages of the cells add together.

4.5.2.2 Fuel Cells (Chemistry Only)

1. Hydrogen is the fuel used in a hydrogen fuel cell.
2. The overall product formed is water.
3. Oxygen comes from the air.
4. Hydrogen reacts with oxygen in the fuel cell, transferring electrons through an external circuit to produce electricity.
5. An advantage is that hydrogen fuel cells only produce water, so they do not release carbon dioxide during operation.
6. A disadvantage is that hydrogen is difficult and expensive to store and transport.

Topic 5 Review

1. An exothermic reaction releases energy to the surroundings, while an endothermic reaction absorbs energy from the surroundings.
2. Combustion reactions are exothermic because forming products such as carbon dioxide and water releases more energy than is required to break the fuel bonds.
3. An endothermic reaction profile has products at a higher energy level than reactants, while an exothermic profile has products at a lower energy level.
4. Energy is required to break bonds because energy is needed to overcome the forces holding atoms together.
5. Rechargeable batteries can be reused because their reactions are reversible, while hydrogen fuel cells continuously produce electricity while supplied with hydrogen and oxygen.
6. A reaction is exothermic when more energy is released making bonds than is used breaking bonds, and endothermic when more energy is needed to break bonds than is released forming them.

Paper 2:

Topic 6 – The Rate and Extent of Chemical Change

4.6.1.1 Calculating Rates of Reactions

1. The rate of a chemical reaction is the amount of reactant used or product formed per unit of time.
2. Mean rate = quantity of reactant used ÷ time taken.
3. Mean rate = quantity of product formed ÷ time taken.

4. Units for rate when measuring mass change include g/s or g/min.
5. A graph of product formed against time can be used by calculating the gradient; a steeper gradient shows a faster reaction.
6. The gradient of a tangent gives the rate at a specific time by calculating change in amount ÷ change in time.

4.6.1.2 Factors Which Affect the Rates of Chemical Reactions

1. The five factors are temperature, concentration, pressure, surface area and catalysts.
2. Increasing concentration increases the rate because there are more reacting particles in the same volume, causing more frequent collisions.
3. Decreasing pressure of reacting gases decreases the rate because gas particles are further apart and collide less often.
4. Increasing surface area increases the rate because more particles are exposed and collisions happen more frequently.
5. Increasing temperature increases the rate because particles move faster and more collisions have enough energy to react.
6. A catalyst increases the rate by providing an alternative pathway with a lower activation energy.

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

1. The volume of gas produced is measured when investigating reaction rate using gas production.
2. Gas volume can be used to measure rate because faster reactions produce gas more quickly.
3. The concentration of a reactant is changed.
4. Only one variable should be changed to make the investigation a fair test.
5. Repeat measurements are carried out to improve reliability and calculate a mean.
6. A hypothesis is made before the investigation to predict the expected relationship between variables.

4.6.1.3 Collision Theory and Activation Energy

1. Collision theory states that particles must collide with enough energy and the correct orientation for a reaction to occur.
2. Reacting particles must collide to allow bonds to break and new bonds to form.
3. Activation energy is the minimum energy needed for successful collisions.

4. Increasing concentration increases reaction rate because there are more particles, causing more frequent successful collisions.
5. Increasing temperature increases reaction rate because particles move faster and more collisions have enough activation energy.
6. Increasing surface area increases reaction rate because more particles are exposed, causing more frequent collisions.

4.6.1.4 Catalysts

1. A catalyst is a substance that increases the rate of a reaction without being used up.
2. Catalysts are not used up because they are unchanged at the end of the reaction.
3. Catalysts increase reaction rate by providing an alternative reaction pathway with lower activation energy.
4. The activation energy decreases when a catalyst is used.
5. Different reactions require different catalysts because catalysts work by specific interactions with reactants.
6. Catalysts can be identified because they appear unchanged in the chemical equation and increase reaction rate.

4.6.2.1 Reversible Reactions

1. A reversible reaction is a reaction where products can react to form the original reactants.
2. Reversible reactions are represented using the symbol ⇋.
3. Products react to form reactants in the reverse reaction.
4. The direction can be changed by altering conditions such as temperature, pressure or concentration.
5. The forward reaction is the reaction that forms products from reactants.
6. The reverse reaction is the reaction that forms reactants from products.

4.6.2.2 Energy Changes and Reversible Reactions

1. The reverse of an exothermic reaction is endothermic.
2. When an exothermic reaction is reversed, energy is absorbed.
3. The same amount of energy is transferred in each direction but in opposite ways.
4. Forward and reverse reactions are linked because reversing a reaction reverses its energy change.
5. The reverse of an endothermic reaction is exothermic.
6. A reversible reaction can be both exothermic and endothermic because the forward and reverse reactions have opposite energy changes.

4.6.2.3 Equilibrium

1. Equilibrium is reached when the forward and reverse reactions occur at the same rate.
2. Equilibrium requires a closed system and constant conditions.
3. A reversible reaction must be in a closed system so reactants and products cannot escape.
4. At equilibrium, the forward and reverse reaction rates are equal.
5. Concentrations remain constant because the amounts of reactants and products are no longer changing overall.
6. Equilibrium is described as dynamic because reactions continue happening even though concentrations stay constant.

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

1. Le Chatelier's Principle states that if conditions are changed, the equilibrium position shifts to oppose the change.
2. An equilibrium system responds by moving in the direction that reduces the effect of the change.
3. The system responds because it tries to restore equilibrium.
4. Temperature, pressure and concentration can affect equilibrium.
5. Le Chatelier's Principle predicts the direction equilibrium will shift after a change.
6. The system shifts to counteract a change by increasing the reaction that removes the effect of the change.

4.6.2.5 The Effect of Changing Concentration (HT Only)

1. Increasing reactant concentration shifts equilibrium towards the products.
2. More reactant particles cause more collisions, increasing the forward reaction rate.
3. Decreasing product concentration shifts equilibrium towards the products.
4. All concentrations change because the system adjusts until equilibrium is restored.
5. Changing concentration affects the position of equilibrium by favouring the reaction that removes the change.
6. For $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$, increasing hydrogen concentration shifts equilibrium to the right, producing more ammonia.

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

1. Increasing temperature in an endothermic reaction increases the amount of products.
2. Increasing temperature in an exothermic reaction decreases the amount of products.
3. Decreasing temperature in an endothermic reaction decreases the amount of products.

4. Decreasing temperature in an exothermic reaction increases the amount of products.
5. Increasing temperature affects equilibrium because the system shifts to absorb or release heat.
6. For the exothermic reaction $N_2 + 3H_2 \leftrightharpoons 2NH_3$, increasing temperature shifts equilibrium left because the reverse reaction absorbs heat.

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

1. Increasing pressure shifts equilibrium towards the side with fewer gas molecules.
2. The equilibrium shifts to fewer gas molecules because this reduces pressure.
3. Decreasing pressure shifts equilibrium towards the side with more gas molecules.
4. The equilibrium shifts to more gas molecules because this increases pressure.
5. The symbol equation shows the number of gas molecules on each side and predicts pressure effects.
6. For $N_2(g) + 3H_2(g) \leftrightharpoons 2NH_3(g)$, increasing pressure shifts equilibrium right because there are fewer gas molecules on the product side.

Topic 6 Review

1. Mean rate of reaction = quantity of reactant used or product formed ÷ time taken.
2. Increasing concentration, pressure, surface area and temperature increases reaction rate; catalysts also increase rate.
3. Collision theory explains temperature effects because higher temperature gives particles more energy and increases successful collisions.
4. Catalysts increase rate by lowering activation energy and are not used up.
5. A reversible reaction can go in both directions, while a reaction that goes to completion only forms products.
6. Changing concentration, temperature and pressure shifts equilibrium in the direction that opposes the change according to Le Chatelier's Principle.

Topic 7 – Organic Chemistry

4.7.1.1 Crude Oil, Hydrocarbons and Alkanes

1. Crude oil is a mixture of hydrocarbons found underground.

2. Crude oil is described as a finite resource because it is non-renewable and will eventually run out.
3. Hydrocarbons are compounds made only of hydrogen and carbon atoms.
4. The general formula for alkanes is C_nH_{2n+2}.
5. The first four alkanes are methane, ethane, propane and butane.
6. C_5H_{12} is an alkane because it follows the alkane general formula C_nH_{2n+2}.

4.7.1.2 Fractional Distillation and Petrochemicals

1. Fractional distillation is used to separate the hydrocarbons in crude oil.
2. It separates hydrocarbons because they have different boiling points.
3. Different fractions condense at different temperatures because they contain hydrocarbons with different chain lengths.
4. Fractions are used as fuels, lubricants and chemical feedstocks.
5. A petrochemical is a chemical product made from hydrocarbons obtained from crude oil.
6. Crude oil is an important feedstock because it provides hydrocarbons used to make many useful chemicals and materials.

4.7.1.3 Properties of Hydrocarbons

1. Boiling point increases as hydrocarbon molecule size increases.
2. Viscosity increases as hydrocarbon molecule size increases.
3. Flammability decreases as hydrocarbon molecule size increases.
4. Hydrocarbons are used as fuels because they release energy when burned.
5. Complete combustion of a hydrocarbon produces carbon dioxide and water.
6. $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$.

4.7.1.4 Cracking and Alkenes

1. Cracking is the process of breaking long-chain hydrocarbons into shorter-chain hydrocarbons.
2. Hydrocarbons are cracked to produce more useful shorter hydrocarbons and alkenes.
3. Catalytic cracking uses a high temperature, a catalyst and often a high pressure.
4. Alkenes are produced alongside alkanes during cracking.
5. Bromine water changes from orange/brown to colourless when it reacts with an alkene.
6. $C_{10}H_{22} \rightarrow C_8H_{18} + C_2H_4$.

4.7.2.1 Structure and Formulae of Alkenes

1. The functional group of alkenes is the carbon-carbon double bond (C=C).
2. The general formula for alkenes is C_nH_{2n}.
3. Alkenes are unsaturated because they contain a carbon-carbon double bond.
4. The first four alkenes are ethene, propene, butene and pentene.
5. An alkene can be identified from its formula because it follows the C_nH_{2n} pattern.
6. C_4H_8 could be an alkene because it follows the general formula C_nH_{2n}.

4.7.2.2 Reactions of Alkenes

1. Alkenes are more reactive than alkanes because the double bond can break and allow atoms to add.
2. An alkene reacts with hydrogen to form an alkane.
3. Hydrogen reacts with alkenes using a nickel catalyst and heat.
4. An alkene reacts with a halogen to form a dihaloalkane.
5. The carbon-carbon double bond breaks during addition reactions.
6. $C_2H_4 + H_2 \rightarrow C_2H_6$.

4.7.2.3 Alcohols

1. The functional group of alcohols is the hydroxyl group (-OH).
2. The first four alcohols are methanol, ethanol, propanol and butanol.
3. Ethanol burns in oxygen to produce carbon dioxide and water.
4. Alcohols react with sodium to produce hydrogen gas and a salt.
5. Fermentation requires yeast, a sugar solution and a warm temperature without oxygen.
6. $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$.

4.7.2.4 Carboxylic Acids

1. The functional group of carboxylic acids is -COOH.
2. The first four carboxylic acids are methanoic acid, ethanoic acid, propanoic acid and butanoic acid.
3. Carboxylic acids react with carbonates to produce salt, water and carbon dioxide.
4. Carboxylic acids dissolve in water because they can form hydrogen bonds with water molecules.
5. A carboxylic acid reacting with an alcohol is an esterification reaction.
6. Carboxylic acids are weak acids because they only partially ionise in water.

4.7.3.1 Addition Polymerisation

1. A polymer is a large molecule made from many repeating units.
2. A monomer is a small molecule that joins with other monomers to form a polymer.
3. Addition polymerisation uses monomers containing carbon-carbon double bonds.
4. The carbon-carbon double bond opens and monomers join together.
5. Addition polymers contain the same atoms as their monomers because no atoms are lost.
6. Ethene forms poly(ethene) with the repeating unit $–CH_2–CH_2–$.

4.7.3.2 Condensation Polymerisation (HT Only)

1. Condensation polymerisation requires monomers with two functional groups.
2. Condensation polymerisation produces a small molecule, usually water, unlike addition polymerisation.
3. Water is usually produced during condensation polymerisation.
4. Propanediol and butanedioic acid form a polyester by joining through ester links.
5. Polyesters are formed when monomers join and release small molecules such as water.
6. Ethanediol and hexanedioic acid form a polyester through ester bonds formed between hydroxyl and carboxyl groups.

4.7.3.3 Amino Acids (HT Only)

1. Amino acids contain an amine group ($-NH_2$) and a carboxylic acid group (-COOH).
2. Amino acids form proteins.
3. Amino acids join together by condensation polymerisation.
4. Water is produced when amino acids polymerise.
5. The formula of glycine is NH_2CH_2COOH.
6. Different amino acids combine in different sequences to form different proteins.

4.7.3.4 DNA and Other Naturally Occurring Polymers

1. DNA is a polymer that carries genetic information.
2. The monomers of DNA are nucleotides.
3. DNA has a double helix structure.
4. DNA encodes genetic information used to make proteins.
5. The monomers of proteins are amino acids.
6. Starch and cellulose are made from glucose monomers.

Topic 7 Review

1. Alkanes are saturated hydrocarbons with single bonds, while alkenes are unsaturated hydrocarbons containing a C=C double bond.
2. Fractional distillation separates crude oil because hydrocarbons have different boiling points.
3. Hydrocarbon combustion: hydrocarbon + oxygen → carbon dioxide + water.
4. Cracking produces more useful hydrocarbons by converting long-chain molecules into shorter alkanes and alkenes.
5. Addition polymerisation uses one type of alkene monomer and produces no small molecule; condensation polymerisation uses two monomers and releases a small molecule.
6. Carbon atoms can form many organic compounds because they can form four strong covalent bonds and bond with other carbon atoms to make chains and rings.

Topic 8 – Chemical Analysis

4.8.1.1 Pure Substances

1. A pure substance contains only one element or one compound.
2. A pure substance has fixed composition, while a mixture contains two or more substances not chemically bonded.
3. Pure substances have specific melting and boiling points because they contain particles with the same strength of attraction.
4. An impure substance melts over a range of temperatures and usually has a lower melting point.
5. Boiling point data can show whether a substance is pure because a pure substance boils at a fixed temperature.
6. A substance melting over a range of temperatures suggests it is impure.

4.8.1.2 Formulations

1. A formulation is a mixture designed as a useful product with specific quantities of components.
2. Formulations are made with carefully measured quantities so they have the required properties.
3. Examples include fuels, medicines and paints.
4. Each chemical in a formulation has a specific purpose, such as improving colour, stability or effectiveness.
5. A formulation can be identified because it contains several substances with specific functions.

6. Paint is a formulation because it contains pigments, solvents and other substances mixed in specific amounts.

4.8.1.3 Chromatography

1. Chromatography is used to separate and identify substances in a mixture.
2. The stationary phase is the paper and the mobile phase is the solvent.
3. Paper chromatography separates substances because they travel at different speeds depending on their solubility and attraction to the paper.
4. Rf value = distance travelled by substance ÷ distance travelled by solvent.
5. Rf value = 6 ÷ 12 = 0.5.
6. Chromatography can distinguish pure and impure substances because a pure substance produces one spot, while an impure substance produces multiple spots.

Required Practical 6: Investigating Paper Chromatography

1. The aim is to separate and identify substances in a mixture using chromatography.
2. The solvent level must be below the spots so the samples do not dissolve directly into the solvent.
3. Pencil is used because graphite does not dissolve in the solvent.
4. Different solvents are used because different substances dissolve better in different solvents.
5. Rf values can be compared with known substances to identify unknown compounds.
6. A pure compound produces one spot because it contains only one substance.

4.8.2.1 Test for Hydrogen

1. A lit splint is used to test for hydrogen gas.
2. Hydrogen burns with a squeaky pop sound when a burning splint is placed into it.
3. A pop sound is produced.
4. The burning splint is placed at the mouth of the container holding the gas.
5. A squeaky pop confirms the gas is hydrogen.
6. Hydrogen produces a pop because it reacts quickly with oxygen in the air to form water.

4.8.2.2 Test for Oxygen

1. A glowing splint is used to identify oxygen gas.

2. The glowing splint relights when placed into oxygen.
3. Relighting confirms that the gas contains oxygen.
4. Oxygen supports combustion, causing the glowing splint to reignite.
5. A glowing splint is used for the oxygen test.
6. Oxygen supports combustion because it allows burning reactions to continue.

4.8.2.3 Test for Carbon Dioxide

1. Limewater is used to test for carbon dioxide.
2. Carbon dioxide turns limewater cloudy.
3. Cloudiness confirms the presence of carbon dioxide.
4. Limewater is calcium hydroxide solution.
5. Limewater turns cloudy because calcium carbonate precipitate is formed.
6. Carbon dioxide + calcium hydroxide → calcium carbonate + water.

4.8.2.4 Test for Chlorine

1. Damp litmus paper is used to identify chlorine gas.
2. Chlorine bleaches damp litmus paper.
3. Bleached litmus paper becomes white.
4. Litmus paper must be damp because chlorine reacts with water to form bleaching substances.
5. A bleached damp litmus paper confirms chlorine.
6. Chlorine is used as a bleaching agent because it destroys coloured compounds.

4.8.3.1 Flame Tests

1. Flame tests are used to identify metal ions.
2. Lithium ions produce a crimson flame.
3. Sodium ions produce a yellow flame.
4. Potassium ions produce a lilac flame.
5. Copper ions produce a green flame.
6. An orange-red flame indicates calcium ions.

4.8.3.2 Metal Hydroxides

1. Sodium hydroxide solution is added to identify metal ions by forming precipitates.
2. Copper(II) ions produce a blue precipitate.
3. Iron(II) ions produce a green precipitate.
4. Iron(III) ions produce a brown precipitate.
5. Aluminium hydroxide dissolves in excess sodium hydroxide solution.
6. $Mg^{2+} + 2OH^- \rightarrow Mg(OH)_2$.

4.8.3.3 Carbonates

1. Carbonates react with dilute acids to produce carbon dioxide gas.
2. Carbon dioxide is identified using limewater, which turns cloudy.
3. Carbon dioxide turns limewater cloudy because calcium carbonate is produced.
4. A chemical test using dilute acid and limewater identifies carbonate ions.
5. Carbonate ions can be identified because they produce carbon dioxide gas.
6. Carbonate + acid → salt + water + carbon dioxide.

4.8.3.4 Halides

1. Silver nitrate solution is used to test for halide ions.
2. Dilute nitric acid is added to remove interfering ions.
3. Chloride ions produce a white precipitate.
4. Bromide ions produce a cream precipitate.
5. Iodide ions produce a yellow precipitate.
6. A cream precipitate indicates bromide ions.

4.8.3.5 Sulfates

1. Barium chloride solution is used to test for sulfate ions.
2. Dilute hydrochloric acid is added to remove carbonate ions.
3. A white precipitate forms when sulfate ions react with barium chloride.
4. A white precipitate confirms sulfate ions are present.
5. Sulfate ions + barium ions → barium sulfate.
6. A white precipitate forms because insoluble barium sulfate is produced.

Required Practical 7: Identifying Ions in Unknown Ionic Compounds

1. The aim is to identify unknown ions in an ionic compound.
2. Metal ions can be identified using flame tests and sodium hydroxide tests.
3. Negative ions can be identified using tests for halides, sulfates and carbonates.
4. Several tests are needed because different ions may give similar observations.
5. Observations such as colours, precipitates and gases are used to identify ions.
6. Using more than one test increases confidence because results can be confirmed.

4.8.3.6 Instrumental Methods

1. Instrumental methods are used to identify substances and measure their quantities.
2. They are useful when only a small sample is available because they are very sensitive.
3. Advantages include greater accuracy, faster analysis and use of smaller samples.
4. Forensic scientists use instrumental methods to identify unknown substances.
5. Instrumental methods use machines and data analysis, while chemical tests rely on observations.
6. Instrumental methods are more accurate because they provide precise measurements.

4.8.3.7 Flame Emission Spectroscopy

1. Flame emission spectroscopy is used to analyse metal ions.
2. It identifies metal ions by their unique line spectra.
3. A line spectrum is produced.
4. Concentration is measured by comparing the intensity of emitted light with reference samples.
5. Reference spectra are needed to match unknown samples with known ions.
6. A sample with the same spectrum as lithium contains lithium ions.

Topic 8 Review

1. Melting point and boiling point data identify pure substances because pure substances have fixed values.
2. Paper chromatography separates mixtures by differences in solubility and attraction to the stationary phase.
3. Hydrogen gives a squeaky pop, oxygen relights a glowing splint, carbon dioxide turns limewater cloudy and chlorine bleaches damp litmus paper.
4. Flame tests identify metal ions by flame colour and precipitation reactions identify ions by precipitate formation.
5. Instrumental methods are usually faster, more accurate and more sensitive than chemical tests.
6. Analytical tests provide evidence about unknown substances by comparing observations with known results.

Topic 9 – Chemistry of the Atmosphere

4.9.1.1 The Proportions of Different Gases in the Atmosphere

1. Nitrogen makes up approximately 78% of the modern atmosphere.
2. Oxygen makes up approximately 21% of the modern atmosphere.
3. Two gases found in small proportions are carbon dioxide and noble gases.
4. The approximate proportions are 78% nitrogen and 21% oxygen.
5. The composition of the atmosphere is described as stable because the percentages of gases have remained similar for around 200 million years.
6. Nitrogen

4.9.1.2 The Earth's Early Atmosphere

1. The Earth's early atmosphere is thought to have contained mainly carbon dioxide, water vapour, methane and ammonia.
2. Volcanic activity released gases such as carbon dioxide and water vapour, forming the early atmosphere.
3. There was little or no oxygen because photosynthetic organisms had not yet developed.
4. Oceans formed when water vapour condensed as the Earth cooled.
5. Carbon dioxide decreased because it dissolved in oceans and became locked in carbonate rocks.
6. Evidence about the early atmosphere is limited because there are few direct records from billions of years ago.

4.9.1.3 How Oxygen Increased

1. Photosynthesis produced the oxygen in the Earth's atmosphere.
2. Algae were the first organisms to produce oxygen by photosynthesis.
3. Algae began producing oxygen approximately 2.7 billion years ago.
4. Increased oxygen allowed animals to evolve because they could use oxygen for aerobic respiration.
5. Photosynthesis word equation:
   carbon dioxide + water → glucose + oxygen.
6. Plants increased oxygen levels by removing carbon dioxide and releasing oxygen during photosynthesis.

4.9.1.4 How Carbon Dioxide Decreased

1. Plants and algae reduced carbon dioxide by absorbing it during photosynthesis.

2. Sedimentary rocks formed when carbon dioxide dissolved in oceans and formed carbonate deposits.
3. Fossil fuels formed from dead organisms containing carbon that were buried and compressed over millions of years.
4. Limestone formation reduced carbon dioxide by storing carbon in carbonate rocks.
5. Carbon-containing deposits include fossil fuels, limestone and carbonates.
6. Photosynthesis removed carbon dioxide, while fossil fuel formation stored carbon away from the atmosphere.

4.9.2.1 Greenhouse Gases

1. A greenhouse gas is a gas that absorbs infrared radiation and contributes to the greenhouse effect.
2. Three greenhouse gases are carbon dioxide, methane and water vapour.
3. Greenhouse gases are important because they keep Earth warm enough for life.
4. Short wavelength radiation from the Sun passes through the atmosphere and reaches Earth's surface.
5. Long wavelength radiation emitted from Earth is absorbed and re-emitted by greenhouse gases.
6. Greenhouse gases cause the greenhouse effect by trapping some infrared radiation and warming the atmosphere.

4.9.2.2 Human Activities Which Increase Greenhouse Gases

1. Burning fossil fuels and deforestation increase carbon dioxide levels.
2. Farming and landfill sites increase methane levels.
3. Burning fossil fuels releases carbon dioxide because carbon in fuels reacts with oxygen.
4. Farming increases methane through livestock digestion and decomposition of waste.
5. Scientists use peer review to check the reliability and accuracy of climate research.
6. Evidence about climate change is difficult to interpret because climate systems are complex and affected by many factors.

4.9.2.3 Global Climate Change

1. Global climate change is the long-term change in average global temperatures and climate patterns.
2. The main cause of increased global temperature is human activity increasing greenhouse gas levels.

3. Effects include rising sea levels, melting ice caps, extreme weather, and changes in ecosystems.
4. Rising temperatures can increase sea levels by melting ice and causing thermal expansion of oceans.
5. Climate change is difficult to predict because many environmental factors interact.
6. Scientists consider risks and impacts to understand possible consequences and develop solutions.

4.9.2.4 The Carbon Footprint and Its Reduction

1. A carbon footprint is the total amount of greenhouse gases released by an activity, product or person.
2. Carbon dioxide and methane are included when calculating carbon footprints.
3. Reducing fossil fuel use decreases carbon dioxide emissions.
4. Reducing methane emissions lowers the carbon footprint by reducing methane release.
5. Actions include using renewable energy and improving energy efficiency.
6. Some methods are limited because they can be expensive, difficult to implement or have other environmental impacts.

4.9.3.1 Atmospheric Pollutants from Fuels

1. Combustion of fuels is a major source of atmospheric pollutants.
2. Burning fuels containing carbon and hydrogen produces carbon dioxide and water.
3. Carbon monoxide is produced by incomplete combustion.
4. Sulfur dioxide is produced when sulfur impurities in fuels burn, and nitrogen oxides form when nitrogen reacts with oxygen at high temperatures.
5. Particulates are tiny solid particles released into the atmosphere.
6. Incomplete combustion produces carbon monoxide and soot because there is insufficient oxygen for complete combustion.

4.9.3.2 Properties and Effects of Atmospheric Pollutants

1. Carbon monoxide is dangerous because it binds to haemoglobin and prevents oxygen transport in blood.
2. Carbon monoxide is difficult to detect because it is colourless and odourless.
3. Sulfur dioxide causes acid rain and can irritate the respiratory system.
4. Oxides of nitrogen contribute to acid rain and respiratory problems.
5. Particulates can damage lungs and contribute to global dimming.

6. Atmospheric pollutants affect health by causing respiratory problems and affect the environment through acid rain and climate effects.

Topic 9 Review

1. The Earth's atmosphere changed from mainly carbon dioxide and water vapour to its current composition with high nitrogen and oxygen levels.
2. Photosynthesis increased oxygen levels by releasing oxygen and reduced carbon dioxide by absorbing it.
3. Greenhouse gases absorb infrared radiation, trapping heat and warming the Earth.
4. Human activities such as burning fossil fuels and farming increase greenhouse gas levels and contribute to climate change.
5. Combustion produces pollutants such as carbon monoxide, sulfur dioxide, nitrogen oxides and particulates.
6. Reducing emissions through renewable energy, energy efficiency and pollution controls can reduce greenhouse gas emissions and atmospheric pollution.

Topic 10 – Using Resources

4.10.1.1 Using the Earth's Resources and Sustainable Development

1. Natural resources are used by humans to provide materials, fuels, food and energy.
2. A finite resource is a resource that will run out because it is not replaced quickly enough.
3. A renewable resource is a resource that can be replaced naturally.
4. Natural products can be replaced by agricultural products, for example cotton replacing some animal fibres.
5. Natural products can be replaced by synthetic products, for example plastics replacing some natural materials.
6. Sustainable development means meeting the needs of current generations without preventing future generations meeting their needs.

4.10.1.2 Potable Water

1. Potable water is water that is safe to drink.
2. Potable water contains dissolved substances but no harmful microorganisms, while pure water contains only water molecules.
3. Water is suitable as potable water when it has acceptable levels of dissolved substances and is free from harmful microbes.

4. The three main stages are filtration, sterilisation and testing.
5. Water can be sterilised using chlorine, ozone or ultraviolet light.
6. Desalination requires large amounts of energy because separating dissolved salts from seawater requires energy.

Required Practical 8: Analysis and Purification of Water Samples

1. Properties investigated include pH, dissolved solids and purity.
2. The amount of dissolved solids can be measured by evaporating water and measuring the remaining solid residue.
3. Distillation is used because it removes dissolved substances by evaporating and condensing pure water.
4. pH is measured to determine acidity or alkalinity.
5. Water from different sources contains different dissolved minerals and substances.
6. Distilled water is purer than potable water because dissolved substances have been removed.

4.10.1.3 Waste Water Treatment

1. Waste water must be treated to remove harmful substances before release into the environment.
2. Screening and grit removal remove large solids and particles.
3. Sedimentation produces sewage sludge.
4. Anaerobic digestion breaks down sewage sludge without oxygen and produces methane.
5. Aerobic biological treatment uses microorganisms and oxygen to break down organic matter.
6. Potable water is easier to obtain from fresh water because it contains fewer dissolved substances than seawater or waste water.

4.10.1.4 Alternative Methods of Extracting Metals (HT Only)

1. Alternative extraction methods are needed because traditional mining can damage the environment and high-grade ores are becoming limited.
2. Phytomining uses plants to extract metals from the ground.
3. Plants absorb metal ions from soil, then are harvested and processed to extract the metal.
4. Bioleaching uses bacteria to produce solutions containing metal compounds.
5. Copper can be extracted from a solution of copper compounds by displacement or electrolysis.
6. Biological extraction methods have the advantage of being less damaging to the environment but are slower than traditional methods.

4.10.2.1 Life Cycle Assessment

1. A life cycle assessment (LCA) evaluates the environmental impacts of a product throughout its life.
2. It considers raw materials, manufacture, transport, use and disposal.
3. Raw materials are considered because extraction and processing use resources and energy.
4. Pollutant effects are harder to include because they are difficult to measure accurately.
5. Selective LCAs are criticised because they may ignore important environmental impacts.
6. An LCA can compare plastic and paper bags by assessing their resource use, energy use and environmental impacts.

4.10.2.2 Ways of Reducing the Use of Resources

1. The three main ways are reduce, reuse and recycle.
2. Recycling reduces the use of limited resources by allowing materials to be used again.
3. Recycling metals reduces energy use because less energy is needed than extracting new metals from ores.
4. Glass bottles can be reused, melted down and recycled into new glass.
5. Metals are recycled by collecting, sorting, melting and reshaping them into new products.
6. Recycling reduces environmental impacts by reducing mining, quarrying and energy consumption.

4.10.3.1 Corrosion and Its Prevention

1. Corrosion is the destruction of metals by reactions with substances in the environment.
2. Rusting is the corrosion of iron.
3. Iron needs oxygen and water to rust.
4. Painting prevents rusting by creating a barrier that stops oxygen and water reaching the iron.
5. Sacrificial protection uses a more reactive metal to protect iron.
6. Zinc protects iron during galvanising because zinc reacts instead of iron and prevents iron from oxidising.

4.10.3.2 Alloys as Useful Materials

1. An alloy is a mixture of a metal with other elements.
2. Bronze contains copper and tin.
3. Brass contains copper and zinc.
4. 18 carat gold contains 75% gold.

5. Alloys are often more useful because they are harder and have improved properties compared with pure metals.
6. Stainless steel resists corrosion because it contains chromium, which forms a protective oxide layer.

4.10.3.3 Ceramics, Polymers and Composites

1. Soda-lime glass is made from sand, sodium carbonate and calcium carbonate.
2. Borosilicate glass has a higher melting point because it contains boron compounds with stronger bonding.
3. Clay ceramics are produced by shaping clay and heating it at high temperatures.
4. Thermosoftening polymers soften when heated and can be reshaped; thermosetting polymers do not soften because of strong cross-links.
5. Composite materials contain two or more different materials combined to improve properties.
6. A material's structure affects its properties, which determines its uses.

4.10.4.1 The Haber Process

1. Ammonia is used to manufacture fertilisers.
2. The raw materials are nitrogen and hydrogen.
3. Nitrogen comes from the air.
4. Hydrogen comes from natural gas.
5. Iron catalyst, high temperature (about 450°C) and a high pressure (about 200 atmospheres)
6. High temperature and pressure are used as a compromise between reaction rate, yield and cost.

4.10.4.2 Production and Uses of NPK Fertilisers

1. NPK fertilisers contain nitrogen, phosphorus and potassium.
2. NPK fertilisers are added to soil to replace nutrients needed for plant growth.
3. N represents nitrogen, P represents phosphorus and K represents potassium.
4. Phosphate rock cannot be used directly because it is insoluble and plants cannot absorb it easily.
5. Sulfuric acid, nitric acid or phosphoric acid can be used to treat phosphate rock.
6. NPK fertilisers are formulations because they contain carefully measured amounts of different substances to produce a useful product.

Topic 10 Review

1. Chemistry helps humans use resources sustainably by developing recycling methods, alternative materials and efficient processes.
2. Potable water from fresh water requires filtration and sterilisation, while seawater requires energy-intensive desalination.
3. Waste water is treated by screening, sedimentation, anaerobic digestion and aerobic biological treatment.
4. Limited resources can be reduced by reducing use, reusing materials and recycling.
5. Alloys, ceramics, polymers and composites have useful properties due to their structures and bonding.
6. The Haber process produces ammonia by reacting nitrogen and hydrogen using an iron catalyst at high pressure and temperature; conditions are a compromise between rate, yield and cost.