Edexcel GCSE Comb Sci Chem

Topic 1 – Key concepts in chemistry

Atomic structure

1.1 Describe how the Dalton model of an atom has changed over time because of the discovery of subatomic particles

1. Dalton's original model described atoms as indivisible, solid spheres.
2. The discovery of electrons showed that atoms contain smaller subatomic particles.
3. Thomson proposed a model in which electrons were embedded in a sphere of positive charge.
4. Rutherford's alpha-scattering experiment provided evidence for a small, dense, positively charged nucleus.
5. The nuclear model showed that most of the atom is empty space, with electrons surrounding a tiny nucleus.
6. The modern model developed from Dalton's model through the discovery of electrons, the nucleus, protons and neutrons.

1.2 Describe the structure of an atom as a nucleus containing protons and neutrons, surrounded by electrons in shells

1. Protons and neutrons are found in the nucleus.
2. Electrons are found in shells surrounding the nucleus.
3. An electron shell is a region around the nucleus occupied by electrons.
4. The nucleus contains protons and neutrons.
5. Most of the mass of an atom is located in the nucleus.
6. An atom consists of a tiny nucleus containing protons and neutrons, surrounded by electrons arranged in shells.

1.3 Recall the relative charge and relative mass of: a a proton b a neutron c an electron

1. A proton has a relative charge of +1 and a relative mass of 1.
2. A neutron has a relative charge of 0 and a relative mass of 1.
3. **An electron has a relative charge of −1 and a relative mass of approximately 1/1840.**

4. **The subatomic particle with a relative charge of −1 and relative mass of approximately 1/1840 is an electron.**
5. Protons and neutrons have approximately the same relative mass.
6. **The relative charges of a proton, neutron and electron are +1, 0 and −1 respectively.**

1.4 Explain why atoms contain equal numbers of protons and electrons

1. A neutral atom contains equal numbers of protons and electrons because their charges are equal in magnitude and opposite in sign.
2. A proton has a relative charge of +1.
3. **An electron has a relative charge of −1.**
4. More protons than electrons would give the atom an overall positive charge.
5. More electrons than protons would give the atom an overall negative charge.
6. **Equal numbers of +1 protons and −1 electrons cancel each other's charges, giving the atom no overall charge.**

1.5 Describe the nucleus of an atom as very small compared to the overall size of the atom

1. The nucleus is extremely small compared with the overall size of the atom.
2. The nucleus occupies only a very small part of the overall size of an atom.
3. Electrons are found in shells surrounding the tiny nucleus.
4. The nucleus is very small compared with the region occupied by the electrons.
5. The small size of the nucleus indicates that most of an atom is empty space.
6. The nucleus is tiny compared with the much larger region over which the electrons are found.

1.6 Recall that most of the mass of an atom is concentrated in the nucleus

1. Most of the mass of an atom is concentrated in the nucleus.
2. Protons and neutrons account for almost all of an atom's mass.
3. Electrons contribute very little to the mass of an atom because their relative mass is approximately 1/1840.
4. The nucleus has almost all the mass of the atom, while the electrons have very little mass.
5. The mass of an atom is approximately determined by the total number of protons and neutrons in its nucleus.
6. The nucleus contains most of the mass because it contains the relatively massive protons and neutrons.

1.7 Recall the meaning of the term mass number of an atom

1. The mass number is the total number of protons and neutrons in an atom.
2. Mass number = number of protons + number of neutrons.
3. Protons and neutrons are included in the mass number.
4. Electrons are not included because their mass is negligible compared with the masses of protons and neutrons.
5. An atom containing 11 protons and 12 neutrons has a mass number of 23.
6. An atom with mass number 23 and 11 protons contains 12 neutrons.

1.8 Describe atoms of a given element as having the same number of protons in the nucleus and that this number is unique to that element

1. The number of protons in the nucleus determines which element an atom belongs to.
2. All atoms of the same element have the same number of protons.
3. The atomic number is the number of protons in the nucleus of an atom.
4. Each element has a unique number of protons.
5. An atom containing 17 protons is chlorine.
6. No two different elements can have atoms with the same number of protons because the proton number defines the element.

1.9 Describe isotopes as different atoms of the same element containing the same number of protons but different numbers of neutrons in their nuclei

1. An isotope is a different atom of the same element with the same number of protons but a different number of neutrons.
2. Isotopes have the same number of protons and therefore have the same chemical identity.
3. The nuclei of isotopes of the same element contain the same number of protons.
4. The nuclei of isotopes contain different numbers of neutrons.
5. Increasing the number of neutrons increases the mass number of an isotope.
6. Carbon-12 and carbon-14 are isotopes because both have 6 protons but carbon-12 has 6 neutrons whereas carbon-14 has 8 neutrons.

1.10 Calculate the numbers of protons, neutrons and electrons in atoms given the atomic number and mass number

1. The number of protons is equal to the atomic number.
2. **Number of neutrons = mass number – atomic number.**
3. In a neutral atom, the number of electrons equals the number of protons.
4. Atomic number 13, mass number 27: 13 protons, 14 neutrons and 13 electrons.

5. Atomic number 17, mass number 35: 17 protons, 18 neutrons and 17 electrons.
6. An atom with 19 protons and mass number 39 has 20 neutrons and 19 electrons.

1.11 Explain how the existence of isotopes results in relative atomic masses of some elements not being whole numbers

1. Relative atomic mass can be a non-whole number because an element can exist as a mixture of isotopes with different masses.
2. The different isotopes of an element contribute to its overall relative atomic mass.
3. Relative atomic mass is a weighted mean of the masses of the naturally occurring isotopes.
4. Chlorine has isotopes with different masses and abundances, giving chlorine a relative atomic mass of approximately 35.5.
5. More abundant isotopes have a greater effect on the relative atomic mass than less abundant isotopes.
6. The relative atomic mass lies between the masses of the isotopes when an element contains isotopes with different masses.

1.12 Calculate the relative atomic mass of an element from the relative masses and abundances of its isotopes

1. Relative atomic mass = (isotope mass × abundance + isotope mass × abundance + ...) ÷ total abundance.
2. (10 × 20 + 11 × 80) ÷ 100 = 10.8
3. (63 × 70 + 65 × 30) ÷ 100 = 63.6
4. (24 × 75 + 26 × 25) ÷ 100 = 24.5
5. (35 × 75 + 37 × 25) ÷ 100 = 35.5
6. (20 × 90 + 21 × 5 + 22 × 5) ÷ 100 = 20.15

The periodic table

1.13 Describe how Dmitri Mendeleev arranged the elements, known at that time, in a periodic table by using properties of these elements and their compounds

1. Mendeleev arranged the known elements into a periodic table based largely on their relative atomic masses and chemical properties.
2. He used the properties of elements to identify patterns and similarities.
3. Mendeleev left gaps where no known element fitted the pattern.
4. He considered the properties of compounds when identifying similarities between elements.
5. His arrangement improved classification by placing elements with similar properties together and revealing repeating patterns.

6. Elements with similar properties were placed into the same groups.

1.14 Describe how Dmitri Mendeleev used his table to predict the existence and properties of some elements not then discovered

1. Mendeleev left gaps in his periodic table for elements that he predicted had not yet been discovered.
2. He used the position of a gap to predict the existence of an undiscovered element.
3. He could predict properties such as relative atomic mass and chemical properties from an element's position.
4. The later discovery of elements with properties matching Mendeleev's predictions supported his periodic table.
5. These discoveries demonstrated that his periodic table could successfully predict previously unknown elements.
6. Mendeleev left gaps because he prioritised the pattern of chemical properties rather than forcing all known elements into incorrect positions.

1.15 Explain that Dmitri Mendeleev thought he had arranged elements in order of increasing relative atomic mass but this was not always true because of the relative abundance of isotopes of some pairs of elements in the periodic table

1. Mendeleev thought he had arranged elements in order of increasing relative atomic mass because he used relative atomic mass as an important basis for his arrangement.
2. Relative atomic mass and atomic number do not always increase in exactly the same order.
3. Differences in the relative abundances of isotopes can give one element a greater relative atomic mass than another element with a higher atomic number.
4. Isotopes with different masses and abundances can produce apparent anomalies in relative atomic mass ordering.
5. The modern periodic table is arranged by increasing atomic number because atomic number directly identifies the element.
6. The relative abundance of isotopes affects the weighted mean relative atomic mass and can explain why some pairs of elements do not follow increasing mass order.

1.16 Explain the meaning of atomic number of an element in terms of position in the periodic table and number of protons in the nucleus

1. Atomic number is the number of protons in the nucleus of an atom.
2. The atomic number is equal to the number of protons in the nucleus.
3. Elements are positioned in the periodic table according to increasing atomic number.

4. Every element has a unique atomic number because every element has a unique number of protons.
5. An element with atomic number 12 has 12 protons.
6. An atom with 16 protons has atomic number 16 and occupies the position of sulfur in the periodic table.

1.17 Describe that in the periodic table a elements are arranged in order of increasing atomic number, in rows called periods b elements with similar properties are placed in the same vertical columns called groups

1. Elements in the modern periodic table are arranged in order of increasing atomic number.
2. A period is a horizontal row in the periodic table.
3. A group is a vertical column in the periodic table.
4. Elements in the same group have similar chemical properties.
5. Moving from left to right across a period, the atomic number generally increases by one.
6. A period is a horizontal row, whereas a group is a vertical column containing elements with similar properties.

1.18 Identify elements as metals or non-metals according to their position in the periodic table, explaining this division in terms of the atomic structures of the elements

1. Metals are generally found on the left and centre of the periodic table.
2. Non-metals are generally found on the right-hand side of the periodic table.
3. Metals generally have atoms that can lose electrons relatively easily, whereas non-metals generally gain or share electrons.
4. Metals tend to lose electrons because they generally have fewer outer-shell electrons.
5. Non-metals tend to gain or share electrons to achieve a more stable outer electron arrangement.
6. The position of an element in the periodic table can therefore be used to predict whether it is likely to be a metal or non-metal.

1.19 Predict the electronic configurations of the first 20 elements in the periodic table as diagrams and in the form, for example 2.8.1

1. Electronic configuration describes how the electrons in an atom are arranged in electron shells.
2. The first electron shell can contain a maximum of 2 electrons.
3. For the first 20 elements, the second shell can contain a maximum of 8 electrons.
4. Sodium has the electronic configuration 2.8.1.

5. Chlorine has the electronic configuration 2.8.7.
6. Calcium has the electronic configuration 2.8.8.2.

1.20 Explain how the electronic configuration of an element is related to its position in the periodic table

1. The number of occupied electron shells corresponds to the period number.
2. For the main groups, the number of electrons in the outer shell corresponds to the group number, apart from helium.
3. Electronic configuration can therefore be used to identify an element's position in the periodic table.
4. An element with electronic configuration 2.8.1 is in Group 1 and Period 3.
5. An element with electronic configuration 2.8.7 is in Group 7 and Period 3.
6. An element in Group 2 and Period 3 has the electronic configuration 2.8.2.

Ionic bonding

1.21 Explain how ionic bonds are formed by the transfer of electrons between atoms to produce cations and anions, including the use of dot and cross diagrams

1. Ionic bonding involves the transfer of electrons from one atom to another.
2. A metal atom loses electrons to form a positively charged cation.
3. A non-metal atom gains electrons to form a negatively charged anion.
4. Electron transfer allows the resulting ions to achieve more stable electron configurations.
5. Dot and cross diagrams show the outer electrons from each atom and the transfer of electrons between them.
6. In sodium chloride, sodium transfers one electron to chlorine, forming Na^+ and Cl^- ions with stable outer electron arrangements.

1.22 Recall that an ion is an atom or group of atoms with a positive or negative charge

1. An ion is an atom or group of atoms with an overall positive or negative charge.
2. An atom becomes a positive ion when it loses one or more electrons.
3. An atom becomes a negative ion when it gains one or more electrons.
4. A cation is a positively charged ion.
5. An anion is a negatively charged ion.
6. An ion can consist of a group of atoms, such as sulfate, SO_4^{2-}.

1.23 Calculate the numbers of protons, neutrons and electrons in simple ions given the atomic number and mass number

1. The number of protons in an ion is equal to its atomic number.
2. **Number of neutrons = mass number – atomic number.**
3. The number of electrons in a positive ion is found by subtracting the positive charge from the atomic number.
4. The number of electrons in a negative ion is found by adding the magnitude of the negative charge to the atomic number.
5. Na^+: 11 protons, 12 neutrons and 10 electrons.
6. Cl^-: 17 protons, 18 neutrons and 18 electrons.

1.24 Explain the formation of ions in ionic compounds from their atoms, limited to compounds of elements in groups 1, 2, 6 and 7

1. Group 1 atoms lose one electron to form +1 ions.
2. Group 2 atoms lose two electrons to form +2 ions.
3. **Group 6 atoms gain two electrons to form –2 ions.**
4. **Group 7 atoms gain one electron to form –1 ions.**
5. Magnesium loses two electrons to form Mg^{2+}, while oxygen gains two electrons to form O^{2-}.
6. Calcium forms Ca^{2+} and chlorine forms Cl^-, so two chloride ions are required for each calcium ion, giving $CaCl_2$.

1.25 Explain the use of the endings –ide and –ate in the names of compounds

1. The ending –ide generally indicates a compound containing a simple negative ion.
2. The ending –ate generally indicates a compound containing a negative ion containing oxygen.
3. Chloride is a simple negative ion, whereas chlorate contains oxygen.
4. Sodium chloride contains sodium ions and chloride ions.
5. Sodium sulfate contains sulfate ions, SO_4^{2-}.
6. The ending –ide or –ate can therefore provide information about the type of negative ion present.

1.26 Deduce the formulae of ionic compounds (including oxides, hydroxides, halides, nitrates, carbonates and sulfates) given the formulae of the constituent ions

1. Combine the ions in the simplest whole-number ratio that gives an overall neutral charge.
2. Na^+ and O^{2-} form Na_2O.
3. Mg^{2+} and Cl^- form $MgCl_2$.
4. Al^{3+} and SO_4^{2-} form $Al_2(SO_4)_3$.
5. Ca^{2+} and NO_3^- form $Ca(NO_3)_2$.

6. Al^{3+} and CO_3^{2-} form $Al_2(CO_3)_3$.

1.27 Explain the structure of an ionic compound as a lattice structure a consisting of a regular arrangement of ions b held together by strong electrostatic forces (ionic bonds) between oppositely-charged ions

1. An ionic lattice is a regular, repeating arrangement of ions.
2. Positive and negative ions are arranged in a regular pattern throughout the lattice.
3. Oppositely charged ions are held together by strong electrostatic forces.
4. Electrostatic forces are forces of attraction between oppositely charged particles.
5. The regular arrangement contains positive and negative ions because each ion is attracted to ions of opposite charge.
6. Strong electrostatic attractions between oppositely charged ions hold the ionic lattice together.

Covalent bonding

1.28 Explain how a covalent bond is formed when a pair of electrons is shared between two atoms

1. A covalent bond is a strong bond formed when two atoms share a pair of electrons.
2. A covalent bond forms when two atoms share electrons.
3. A single covalent bond contains one shared pair of electrons.
4. Sharing electrons allows atoms to obtain a more stable electron configuration.
5. The shared electrons are attracted to the nuclei of both bonded atoms.
6. Two chlorine atoms each contribute one electron to form one shared pair, producing a covalent bond in Cl_2.

1.29 Recall that covalent bonding results in the formation of molecules

1. Covalent bonding between atoms can form simple molecular substances.
2. A molecule is a group of atoms held together by covalent bonds.
3. The atoms within a molecule are held together by covalent bonds.
4. Covalent bonds hold the atoms together within a molecule.
5. A molecule contains two or more atoms chemically bonded together, whereas an individual atom consists of a single atom.
6. Examples include hydrogen, H_2, and water, H_2O.

1.30 Recall the typical size (order of magnitude) of atoms and small molecules

1. The typical order of magnitude of the size of an atom is 10^{-10} m.

2. The typical order of magnitude of a small molecule is also around 10^{-10} m.
3. A small molecule is generally larger than an individual atom because it contains multiple atoms.
4. A typical atom has a diameter of approximately 10^{-10} m.
5. Individual atoms are far too small to be observed with the naked eye.
6. A football is approximately 10^{8} times larger in diameter than a typical atom.

1.31 Explain the formation of simple molecular, covalent substances, using dot and cross diagrams, including: a hydrogen b hydrogen chloride c water d methane e oxygen f carbon dioxide

1. In H_2, each hydrogen atom contributes one electron to form one shared pair.
2. In HCl, hydrogen and chlorine share one pair of electrons, giving both atoms a stable outer arrangement.
3. In H_2O, oxygen forms two covalent bonds with two hydrogen atoms and has two lone pairs of electrons.
4. In CH_4, carbon forms four covalent bonds with four hydrogen atoms.
5. In O_2, the two oxygen atoms share two pairs of electrons, forming a double covalent bond.
6. In CO_2, carbon forms two double covalent bonds, one with each oxygen atom.

Types of substance

1.32 Explain why elements and compounds can be classified as: a ionic b simple molecular (covalent) c giant covalent d metallic and how the structure and bonding of these types of substances results in different physical properties, including relative melting point and boiling point, relative solubility in water and ability to conduct electricity (as solids and in solution)

1. The four main types are ionic, simple molecular covalent, giant covalent and metallic.
2. Ionic substances contain a lattice of oppositely charged ions, whereas simple molecular covalent substances contain separate molecules.
3. Giant covalent substances contain a giant network of atoms joined by strong covalent bonds.
4. Metals conduct electricity as solids because they contain delocalised electrons that can move through the structure.
5. Different structures and bonding produce different strengths of attraction between particles, resulting in different melting and boiling points.
6. The type of structure and bonding affects whether a substance dissolves in water and whether charged particles or electrons are available to conduct electricity.

1.33 Explain the properties of ionic compounds limited to: a high melting points and boiling points, in terms of forces between ions b whether or not they conduct electricity as solids, when molten and in aqueous solution

1. Ionic compounds have high melting points because strong electrostatic attractions between oppositely charged ions require a lot of energy to overcome.
2. Ionic compounds have high boiling points because the strong electrostatic attractions between ions require a lot of energy to overcome.
3. Solid ionic compounds do not conduct electricity because their ions are fixed in position and cannot move.
4. Molten ionic compounds conduct electricity because their ions are free to move.
5. Ionic compounds in aqueous solution conduct electricity because their ions are free to move through the solution.
6. Ionic compounds conduct electricity only when their ions are able to move, such as when molten or dissolved in water.

1.34 Explain the properties of typical covalent, simple molecular compounds limited to: a low melting points and boiling points, in terms of forces between molecules (intermolecular forces) b poor conduction of electricity

1. Simple molecular covalent substances generally have low melting points because the intermolecular forces between molecules are weak.
2. They generally have low boiling points because only weak intermolecular forces need to be overcome.
3. Intermolecular forces are forces of attraction between molecules.
4. Melting a simple molecular substance involves overcoming intermolecular forces rather than breaking the strong covalent bonds within the molecules.
5. Simple molecular covalent substances generally conduct electricity poorly because they do not contain freely moving charged particles.
6. Weak intermolecular forces account for the relatively low melting and boiling points of simple molecular covalent substances.

1.35 Recall that graphite and diamond are different forms of carbon and that they are examples of giant covalent substances

1. Graphite and diamond are both made from carbon.
2. They are different forms, or allotropes, of the element carbon.
3. Graphite and diamond are giant covalent substances because their carbon atoms form large structures joined by covalent bonds.
4. Carbon atoms in graphite are joined by strong covalent bonds.
5. Carbon atoms in diamond are joined by strong covalent bonds.

6. Their different structures and bonding arrangements give graphite and diamond different physical properties.

1.36 Describe the structures of graphite and diamond

1. Each carbon atom in diamond is covalently bonded to four other carbon atoms.
2. Diamond has a giant three-dimensional structure of carbon atoms.
3. Each carbon atom in graphite is covalently bonded to three other carbon atoms within a layer.
4. Graphite consists of layers of carbon atoms arranged in hexagonal structures.
5. Weak forces act between the layers of graphite.
6. Diamond has a rigid three-dimensional network, whereas graphite has layers that can slide over one another.

1.37 Explain, in terms of structure and bonding, why graphite is used to make electrodes and as a lubricant, whereas diamond is used in cutting tools

1. Graphite contains delocalised electrons that can move through its layers, allowing it to conduct electricity.
2. Graphite can conduct electricity, so it can be used as an electrode.
3. Graphite can act as a lubricant because its layers can slide over one another.
4. Diamond is suitable for cutting tools because it is extremely hard.
5. The layers of graphite are held together by weak forces, allowing them to move relative to one another.
6. Diamond has strong covalent bonds throughout its rigid three-dimensional structure, making it very hard, whereas graphite's layered structure allows it to conduct electricity and act as a lubricant.

1.38 Explain the properties of fullerenes including C60 and graphene in terms of their structures and bonding

1. A fullerene is a form of carbon in which carbon atoms form hollow structures such as cages or tubes.
2. C_{60} consists of 60 carbon atoms arranged in a hollow spherical structure.
3. Some fullerenes can act as lubricants because their structures allow layers or molecules to slide over one another.
4. Graphene is a single layer of carbon atoms arranged in a hexagonal lattice.
5. Graphene can conduct electricity because it contains delocalised electrons that can move through the structure.

6. The structures and bonding of fullerenes and graphene give them properties such as low density, strength, conductivity and useful behaviour as lubricants.

1.39 Describe, using poly(ethene) as the example, that simple polymers consist of large molecules containing chains of carbon atoms

1. A polymer is a substance made from very large molecules formed from many repeating units.
2. Poly(ethene) consists of long chains of carbon atoms with hydrogen atoms attached.
3. The main chain of poly(ethene) contains carbon atoms.
4. Poly(ethene) is described as having large molecules because each molecule contains a very large number of atoms.
5. The carbon atoms form a long continuous chain in a poly(ethene) molecule.
6. Poly(ethene) is formed from ethene molecules joining together to form long polymer chains.

1.40 Explain the properties of metals, including malleability and the ability to conduct electricity

1. Malleability is the ability of a metal to be hammered or bent into different shapes without breaking.
2. Metals can be bent and shaped because layers of metal atoms can move relative to one another while the metallic bonding remains.
3. Metals conduct electricity because they contain delocalised electrons that can move through the structure.
4. Delocalised electrons carry electrical charge through a metal.
5. Metallic bonding allows layers of atoms to move while the attraction between positive metal ions and delocalised electrons maintains the structure.
6. The arrangement of metal ions and delocalised electrons accounts for the malleability and electrical conductivity of metals.

1.41 Describe the limitations of particular representations and models, to include dot and cross, ball and stick models and two- and three-dimensional representations

1. Dot and cross diagrams do not show the actual sizes or three-dimensional shapes of atoms and molecules.
2. Ball and stick models do not accurately represent the relative sizes of atoms and the distances between them.
3. Ball and stick models can make atoms appear larger or smaller relative to their actual sizes.

4. A two-dimensional representation cannot fully show the three-dimensional arrangement of atoms.
5. A three-dimensional model is still a simplified representation and cannot show every feature of a real chemical structure accurately.
6. Different representations are useful because each model highlights some features while having limitations in what it can show.

1.42 Describe most metals as shiny solids which have high melting points, high density and are good conductors of electricity whereas most non-metals have low boiling points and are poor conductors of electricity

1. Most metals are shiny solids with high melting points, high density and good electrical conductivity.
2. Most metals conduct electricity well because they contain delocalised electrons.
3. Metals generally have a shiny appearance.
4. Most non-metals have relatively low boiling points and are poor conductors of electricity.
5. Metals generally have higher melting points and densities than non-metals.
6. Most metals are good electrical conductors, whereas most non-metals are poor electrical conductors.

Calculations involving masses

1.43 Calculate: a relative formula mass given relative atomic masses b percentage by mass of an element in a compound given relative atomic masses

1. Relative formula mass is calculated by adding the relative atomic masses of all the atoms shown in the formula.
2. $Mr(CaCO_3) = 40 + 12 + (3 \times 16) = 100$.
3. $Mr(Al_2(SO_4)_3) = (2 \times 27) + (3 \times 32) + (12 \times 16) = 342$.
4. Percentage by mass = (mass of element in the formula ÷ relative formula mass) × 100.
5. Percentage oxygen in $CO_2 = (32 \div 44) \times 100 = 72.7\%$.
6. Percentage calcium in $CaCO_3 = (40 \div 100) \times 100 = 40.0\%$.

1.44 Calculate the formulae of simple compounds from reacting masses or percentage composition and understand that these are empirical formulae

1. An empirical formula shows the simplest whole-number ratio of atoms of each element in a compound.
2. Convert the reacting masses into moles and divide the mole values by the smallest value to obtain the simplest whole-number ratio.

3. C: 24 ÷ 12 = 2; H: 4 ÷ 1 = 4 → ratio 1:2 → CH_2.
4. N: 14 ÷ 14 = 1; O: 16 ÷ 16 = 1 → NO.
5. Assume 100 g: C = 40/12 = 3.33; H = 6.7/1 = 6.7; O = 53.3/16 = 3.33 → ratio 1:2:1 → CH_2O.
6. Masses or percentages must be converted to moles because chemical formulae represent ratios of numbers of atoms, and moles give the number of particles.

1.45 Deduce: a the empirical formula of a compound from the formula of its molecule b the molecular formula of a compound from its empirical formula and its relative molecular mass

1. Divide all the subscripts in the molecular formula by their greatest common factor to obtain the simplest whole-number ratio.
2. The empirical formula of $C_6H_{12}O_6$ is CH_2O.
3. The empirical formula of C_2H_4 is CH_2.
4. First calculate the empirical formula mass, then divide the molecular relative mass by the empirical formula mass to find the multiplier.
5. CH_2O has empirical formula mass 30; 180 ÷ 30 = 6 ⟶ molecular formula $C_6H_{12}O_6$.
6. NO_2 has empirical formula mass 46; 92 ÷ 46 = 2 ⟶ molecular formula N_2O_4.

1.46 Describe an experiment to determine the empirical formula of a simple compound such as magnesium oxide

1. Weigh a clean, dry crucible and lid, add magnesium, then weigh the crucible, lid and magnesium before heating strongly in air.
2. Magnesium is heated strongly in oxygen/air so that it reacts completely to form magnesium oxide.
3. The crucible and lid must be weighed first so that the mass of the magnesium can be determined accurately.
4. Heating, cooling and reweighing until a constant mass ensures that the magnesium has reacted completely and no further mass increase occurs.
5. Calculate the mass of magnesium and the mass of oxygen gained, convert each to moles and determine the simplest whole-number ratio.
6. Wear safety goggles and use appropriate heat-resistant equipment; do not look directly at burning magnesium.

1.47 Explain the law of conservation of mass applied to: a a closed system including a precipitation reaction in a closed flask b a non-enclosed system including a reaction in an open flask that takes in or gives out a gas

1. The law of conservation of mass states that mass is neither created nor destroyed during a chemical reaction.

2. In a closed system, no substances can enter or leave, so the total mass remains constant.
3. In a sealed flask, all substances involved in the precipitation reaction remain inside the system, so the total mass is unchanged.
4. If a gas is produced and escapes from an open flask, the measured mass of the apparatus and contents decreases.
5. If an open system takes in a gas from the surroundings, the measured mass of the apparatus and contents increases.
6. Mass is still conserved in an open system because the gas has simply moved into or out of the system rather than being created or destroyed.

1.48 Calculate masses of reactants and products from balanced equations, given the mass of one substance

1. Use the balanced equation to determine the mole ratio, convert the known mass to moles, use the ratio to find the unknown moles, then convert to mass.
2. 24 g Mg = 1 mol Mg; ratio Mg:MgO = 1:1; Mr(MgO) = 40 → 40 g MgO.
3. 4 g H_2 = 2 mol H_2; ratio H_2:H_2O = 1:1 ⟶ 2 mol H_2O; 2 × 18 = 36 g H_2O.
4. 100 g $CaCO_3$ = 1 mol; ratio $CaCO_3$:CO_2 = 1:1; Mr(CO_2) = 44 ⟶ 44 g CO_2.
5. 54 g Al = 2 mol; ratio Al:$AlCl_3$ = 1:1; Mr($AlCl_3$) = 133.5 ⟶ 267 g $AlCl_3$.
6. 160 g Fe_2O_3 = 1 mol; ratio Fe_2O_3:Fe = 1:2 ⟶ 2 mol Fe; mass = 112 g Fe.

1.49 Calculate the concentration of solutions in g dm–3

1. Concentration (g dm^{-3}) = mass of solute (g) ÷ volume of solution (dm^3).
2. 10 ÷ 2 = 5 g dm^{-3}.
3. 500 cm^3 = 0.500 dm^3; 15 ÷ 0.500 = 30 g dm^{-3}.
4. Mass = concentration × volume = 20 × 0.5 = 10 g.
5. Volume = mass ÷ concentration = 12 ÷ 8 = 1.5 dm^3.
6. 250 cm^3 = 0.250 dm^3; 25 ÷ 0.250 = 100 g dm^{-3}.

1.50 Recall that one mole of particles of a substance is defined as: a the Avogadro constant number of particles (6.02 × 1023 atoms, molecules, formulae or ions) of that substance b a mass of 'relative particle mass'

1. One mole is an amount of substance containing the Avogadro constant number of particles.
2. The Avogadro constant is 6.02 × 10^{23} mol^{-1}.
3. One mole contains 6.02 × 10^{23} particles.
4. The mole can count atoms, molecules, formula units and ions.
5. Relative particle mass is the relative mass of a particle compared with the chosen standard.

6. One mole of a substance has a mass in grams numerically equal to its relative particle mass.

1.51 Calculate the number of: a moles of particles of a substance in a given mass of that substance and vice versa b particles of a substance in a given number of moles of that substance and vice versa c particles of a substance in a given mass of that substance and vice versa

1. Number of moles = mass ÷ relative particle mass.
2. Moles of H_2O = 18 ÷ 18 = 1.00 mol.
3. Mass = moles × relative particle mass = 0.50 × 44 = 22 g CO_2.
4. Number of particles = number of moles × Avogadro constant.
5. 0.25 × 6.02 × 10^{23} = 1.505 × 10^{23} molecules of O_2.
6. 9 g H_2O = 9 ÷ 18 = 0.5 mol; particles = 0.5 × 6.02 × 10^{23} = 3.01 × 10^{23} molecules.

1.52 Explain why, in a reaction, the mass of product formed is controlled by the mass of the reactant which is not in excess

1. A reactant in excess is present in a greater amount than is required to react completely with the other reactants.
2. A limiting reactant is the reactant that is completely used up first.
3. The limiting reactant determines the maximum amount of product because the reaction cannot continue once it has been used up.
4. Reactant B controls the amount of product because it is completely used up.
5. Adding more of a reactant already in excess does not increase the product because the limiting reactant still becomes completely used up first.
6. The reactant that is not in excess is the limiting reactant, so its amount determines the maximum mass of product formed.

1.53 Deduce the stoichiometry of a reaction from the masses of the reactants and products

1. Stoichiometry is the quantitative relationship between the amounts of reactants and products in a chemical reaction.
2. Convert the masses of the reactants and products to moles, then compare the mole amounts to obtain the simplest whole-number ratio.
3. 135 g Al = 5 mol; ratio Al:Al_2O_3 = 2:1 ⟶ 2.5 mol Al_2O_3; Mr = 102 ⟶ 255 g Al_2O_3.
4. 48 g Mg = 2 mol; ratio Mg:MgO = 1:1 → 2 mol MgO; Mr = 40 → 80 g MgO.
5. 250 g $CaCO_3$ = 2.5 mol; ratio $CaCO_3$:CaO = 1:1 ⟶ 2.5 mol CaO; Mr = 56 ⟶ 140 g CaO.
6. 11.2 g Fe = 0.2 mol; 17.6 g FeS = 0.2 mol. Therefore Fe:S = 1:1.

Topic 2 – States of matter and mixtures

States of matter

2.1 Describe the arrangement, movement and the relative energy of particles in each of the three states of matter: solid, liquid and gas

1. In a solid, particles are closely packed in a fixed, regular arrangement.
2. In a liquid, particles are close together but can move past one another.
3. In a gas, particles are far apart and move freely in random directions.
4. Particles in a solid vibrate about fixed positions.
5. Particles in a liquid have more movement than particles in a solid while remaining relatively close together.
6. Particles in a gas have the greatest movement and generally the greatest kinetic energy of the three states.

2.2 Recall the names used for the interconversions between the three states of matter, recognising that these are physical changes: contrasted with chemical reactions that result in chemical changes

1. Solid → liquid: melting.
2. Liquid → gas: boiling/evaporation.
3. Gas → liquid: condensation.
4. Liquid → solid: freezing.
5. Solid → gas: sublimation.
6. Gas → solid: deposition. These changes of state are physical changes because no new substance is formed, whereas chemical reactions produce chemical changes.

2.3 Explain the changes in arrangement, movement and energy of particles during these interconversions

1. During melting, particles become less ordered and can move past one another; their energy increases.
2. During evaporation, particles gain energy and move further apart.
3. During condensation, particles lose energy, move closer together and become less free to move.
4. During freezing, particles lose energy and become arranged more closely in fixed positions.
5. During sublimation, particles gain energy, become much further apart and move freely.

6. During deposition, particles lose energy and form a more closely packed, ordered solid arrangement.

2.4 Predict the physical state of a substance under specified conditions, given suitable data

1. Compare the temperature with the melting point and boiling point: below the melting point is solid, between the melting and boiling points is liquid, and above the boiling point is gas.
2. At 10 °C, the substance is a solid.
3. At 50 °C, the substance is a liquid.
4. At 100 °C, the substance is a gas.
5. Below its melting point, a substance is a solid.
6. Above its boiling point, a substance is a gas.

Methods of separating and purifying substances

2.5 Explain the difference between the use of 'pure' in chemistry compared with its everyday use and the differences in chemistry between a pure substance and a mixture

1. In chemistry, a pure substance contains only one substance and has no other substances mixed with it.
2. The everyday meaning of pure can mean clean, uncontaminated or not adulterated, rather than chemically containing only one substance.
3. A pure substance contains only one element or compound, whereas a mixture contains two or more substances.
4. Air contains several gases, so it is a mixture rather than a pure substance.
5. A mixture contains more than one substance, whereas a pure substance contains only one substance.
6. The components of a mixture retain their individual chemical identities and properties.

2.6 Interpret melting point data to distinguish between pure substances which have a sharp melting point and mixtures which melt over a range of temperatures

1. A pure substance generally melts sharply at a specific temperature, whereas a mixture melts over a range.
2. A sharp melting point suggests that a substance is pure.
3. Melting over a range of temperatures suggests that a substance is a mixture or impure.
4. A substance that melts sharply at 80 °C is likely to be pure.
5. A substance that melts from 70 °C to 76 °C is likely to be a mixture or impure.

6. Melting point data can distinguish purity because pure substances have sharp melting points while mixtures generally melt over a range.

2.7 Explain the types of mixtures that can be separated by using the following experimental techniques: a simple distillation b fractional distillation c filtration d crystallisation e paper chromatography

1. Simple distillation separates a soluble solid from a solution and can be used to obtain the solvent.
2. Fractional distillation separates a mixture of miscible liquids with different boiling points.
3. Filtration separates an insoluble solid from a liquid.
4. Crystallisation separates a soluble solid from a solution.
5. Paper chromatography separates a mixture of soluble substances, such as dyes.
6. The relevant property is the difference in physical properties such as solubility, particle size, boiling point or attraction to the stationary phase.

2.8 Describe an appropriate experimental technique to separate a mixture, knowing the properties of the components of the mixture

1. Use filtration to separate an insoluble solid from a liquid because the solid particles are too large to pass through the filter.
2. Use simple distillation to obtain the solvent from a solution because the solvent has a lower boiling point and can be vaporised and condensed.
3. Use fractional distillation to separate miscible liquids with different boiling points.
4. Use paper chromatography to separate soluble substances with different attractions to the stationary and mobile phases.
5. Identify the physical properties of each component, then select a technique that exploits a difference in those properties.
6. The components' physical properties determine whether they can be separated by filtration, distillation, crystallisation or chromatography.

2.9 Describe paper chromatography as the separation of mixtures of soluble substances by running a solvent (mobile phase) through the mixture on the paper (the paper contains the stationary phase), which causes the substances to move at different rates over the paper

1. The solvent dissolves the substances and carries them through the chromatography paper.
2. The mobile phase is the solvent that moves through the paper.
3. The stationary phase is the chromatography paper.
4. Different substances move at different rates because they have different attractions to the stationary and mobile phases.

5. The substances must dissolve in the solvent so that they can move with the mobile phase.
6. The solvent travels through the paper and carries the dissolved substances different distances, separating the mixture.

2.10 Interpret a paper chromatogram: a to distinguish between pure and impure substances b to identify substances by comparison with known substances c to identify substances by calculation and use of Rf values

1. A pure substance produces one spot, whereas an impure substance generally produces more than one spot.
2. Compare the position of an unknown spot with the position of a known substance under the same conditions to help identify it.
3. A single spot indicates that the sample may contain one substance.
4. More than one spot indicates that the sample contains more than one soluble substance.
5. Rf = distance travelled by substance ÷ distance travelled by solvent front.
6. An unknown substance can be identified by comparing its Rf value with the Rf value of a known substance under the same conditions.

2.11 Core Practical: Investigate the composition of inks using simple distillation and paper chromatography

1. Simple distillation can separate and collect the solvent from the ink because the solvent is vaporised and then condensed.
2. Paper chromatography separates the different soluble dyes in the ink so their composition can be investigated.
3. The ink is placed near the bottom of the paper so the solvent can carry the dyes upwards through the paper.
4. The solvent level must be below the ink spot so the ink does not dissolve directly into the solvent reservoir.
5. Multiple spots on the chromatogram show that the ink contains more than one dye.
6. Compare the positions or Rf values of the unknown ink's dyes with known dyes to identify the components.

2.12 Describe how: a waste and ground water can be made potable, including the need for sedimentation, filtration and chlorination b sea water can be made potable by using distillation c water used in analysis must not contain any dissolved salts

1. Waste water and groundwater can be treated using sedimentation, filtration and chlorination to make them potable.
2. Sedimentation allows suspended particles to settle out of the water.
3. Filtration removes remaining insoluble particles and suspended solids.

4. Chlorination kills microorganisms and makes the water safer to drink.
5. Sea water can be made potable by distillation, which separates water from dissolved salts.
6. Water used in chemical analysis must contain no dissolved salts because dissolved ions could interfere with or alter the results of the analysis.

3.1 Recall that acids in solution are sources of hydrogen ions and alkalis in solution are sources of hydroxide ions

1. Hydrogen ions, H^+.
2. Hydroxide ions, OH^-.
3. Hydrogen ions, H^+.
4. Hydroxide ions, OH^-.
5. H^+.
6. OH^-.

3.2 Recall that a neutral solution has a pH of 7 and that acidic solutions have lower pH values and alkaline solutions higher pH values

1. pH 7.
2. pH values below 7.
3. pH values above 7.
4. Acidic.
5. Alkaline.
6. The solution with pH 2.

3.3 Recall the effect of acids and alkalis on indicators, including litmus, methyl orange and phenolphthalein

1. Red.
2. Blue.
3. Red.
4. Yellow.
5. Colourless.
6. Pink.

3.4 Recall that the higher the concentration of hydrogen ions in an acidic solution, the lower the pH; and the higher the concentration of hydroxide ions in an alkaline solution, the higher the pH

1. The pH decreases.
2. The pH increases.
3. The pH increases.
4. The pH decreases.

5. $0.1\ mol\ dm^{-3}$.
6. $0.01\ mol\ dm^{-3}$.

3.5 Recall that as hydrogen ion concentration in a solution increases by a factor of 10, the pH of the solution decreases by 1

1. The pH decreases by 1.
2. The pH increases by 1.
3. pH 4.
4. pH 1.
5. pH 8.
6. A factor of 1000.

3.6 Core Practical: Investigate the change in pH on adding powdered calcium hydroxide or calcium oxide to a fixed volume of dilute hydrochloric acid

1. A pH meter/probe or universal indicator.
2. To ensure the amount and concentration of acid are controlled variables, making the investigation a fair test.
3. The pH of the solution.
4. To measure the effect of known amounts of calcium hydroxide and obtain accurate results.
5. The pH increases.
6. A graph of pH against mass/amount of calcium hydroxide added.

3.7 Explain the terms dilute and concentrated, with respect to amount of substances in solution

1. A solution containing a relatively small amount of solute per unit volume.
2. A solution containing a relatively large amount of solute per unit volume.
3. A concentrated solution.
4. Add water.
5. It decreases the concentration.
6. No. They describe concentration, not acid strength.

3.8 Explain the terms weak and strong acids, with respect to the degree of dissociation into ions

1. An acid that completely dissociates into ions in aqueous solution.
2. An acid that only partially dissociates into ions in aqueous solution.
3. Nearly all acid molecules dissociate into ions.
4. Only some acid molecules dissociate into ions.
5. A strong acid.
6. Yes. Strength and concentration are different properties.

3.9 Recall that a base is any substance that reacts with an acid to form a salt and water only

1. A substance that reacts with an acid to form a salt and water only.
2. A salt and water.
3. A base.
4. Yes.
5. Yes.
6. A salt.

3.10 Recall that alkalis are soluble bases

1. A soluble base.
2. It must be soluble in water.
3. No.
4. Sodium hydroxide is a soluble base.
5. Copper oxide is insoluble in water.
6. It dissolves and produces OH^- ions in solution.

3.11 Explain the general reactions of aqueous solutions of acids with: a metals b metal oxides c metal hydroxides d metal carbonates to produce salts

1. Salt and hydrogen.
2. Salt and water.
3. Salt and water.
4. Salt, water and carbon dioxide.
5. Hydrogen.
6. Carbon dioxide.

3.12 Describe the chemical test for: a hydrogen b carbon dioxide (using limewater)

1. Place a lighted splint at the mouth of the container holding the gas.
2. A squeaky pop is heard.
3. Bubble the gas through limewater.
4. The limewater turns milky/cloudy.
5. A squeaky pop.
6. It turns milky/cloudy.

3.13 Describe a neutralisation reaction as a reaction between an acid and a base

1. A reaction between an acid and a base.
2. An acid and a base.
3. A salt and water.
4. Yes.

5. A neutralisation reaction.
6. The acidic and alkaline properties are removed as the acid and base react.

3.14 Explain an acid-alkali neutralisation as a reaction in which hydrogen ions (H+) from the acid react with hydroxide ions (OH–) from the alkali to form water

1. H^+ and OH^- ions.
2. The acid.
3. The alkali.
4. Water.
5. $H^+ + OH^- \rightarrow H_2O$.
6. H^+ ions and OH^- ions react to form water.

3.15 Explain why, if soluble salts are prepared from an acid and an insoluble reactant: a excess of the reactant is added b the excess reactant is removed c the solution remaining is only salt and water

1. To ensure all the acid reacts.
2. By filtration.
3. The insoluble reactant remains as a solid and can be separated from the solution.
4. Salt and water.
5. To prevent the insoluble reactant contaminating the salt crystals.
6. The excess insoluble reactant can be removed by filtration.

3.16 Explain why, if soluble salts are prepared from an acid and a soluble reactant: a titration must be used b the acid and the soluble reactant are then mixed in the correct proportions c the solution remaining, after reaction, is only salt and water

1. Both reactants are soluble, so excess reactant cannot be removed by filtration.
2. It would remain dissolved in the solution.
3. The volumes of reactants needed to exactly neutralise each other.
4. To ensure neither reactant is left in excess.
5. Salt and water.
6. Both reactants are soluble, so filtration cannot remove excess reactant.

3.17 Core Practical: Investigate the preparation of pure, dry hydrated copper sulfate crystals starting from copper oxide including the use of a water bath

1. Sulfuric acid.
2. To ensure all the sulfuric acid reacts.
3. By filtration.

4. To gently heat the solution and evaporate some water without overheating it.
5. Concentrate the solution, then allow it to cool so crystals form.
6. To remove surface water and obtain a dry sample.

3.18 Describe how to carry out an acid-alkali titration, using burette, pipette and a suitable indicator, to prepare a pure, dry salt

1. Burette.
2. Pipette.
3. To show when the endpoint/neutralisation has been reached.
4. Record the titre and use the volume to determine the exact proportions of reactants.
5. To avoid contaminating the final salt with indicator.
6. Evaporate some water from the solution, allow the salt to crystallise, then filter and dry the crystals.

3.19 Recall the general rules which describe the solubility of common types of substances in water: a all common sodium, potassium and ammonium salts are soluble b all nitrates are soluble c common chlorides are soluble except those of silver and lead d common sulfates are soluble except those of lead, barium and calcium e common carbonates and hydroxides are insoluble except those of sodium, potassium and ammonium

1. Yes.
2. Yes.
3. Silver chloride and lead chloride.
4. Lead sulfate, barium sulfate and calcium sulfate.
5. Sodium, potassium and ammonium carbonates.
6. Sodium, potassium and ammonium hydroxides.

3.20 Predict, using solubility rules, whether or not a precipitate will be formed when named solutions are mixed together, naming the precipitate if any

1. Identify the possible products and use the solubility rules to determine whether either product is insoluble.
2. An insoluble solid formed from solutions during a reaction.
3. A precipitate forms.
4. Silver chloride, AgCl.
5. Barium sulfate, $BaSO_4$.
6. The possible products are soluble, so no precipitate forms.

3.21 Describe the method used to prepare a pure, dry sample of an insoluble salt

1. A precipitation reaction between two soluble solutions.

2. To form an insoluble salt as a precipitate.
3. By filtration.
4. To remove soluble impurities.
5. Leave it to dry or dry it using suitable drying equipment.
6. To remove soluble impurities remaining in the precipitate.

3.22 Recall that electrolytes are ionic compounds in the molten state or dissolved in water

1. An ionic compound that conducts electricity when molten or dissolved in water.
2. Ionic compounds.
3. The ions are free to move.
4. The ions are free to move through the solution.
5. Its ions are held in fixed positions in the lattice.
6. Free-moving charged ions.

3.23 Describe electrolysis as a process in which electrical energy, from a direct current supply, decomposes electrolytes

1. The decomposition of an electrolyte using electrical energy.
2. Electrical energy.
3. Direct current (DC).
4. It decomposes into simpler substances.
5. Electrical energy causes ions to move to the electrodes, where they undergo chemical reactions.
6. It provides the electrical energy and maintains a constant direction of ion movement.

3.24 Explain the movement of ions during electrolysis, in which: a positively charged cations migrate to the negatively charged cathode b negatively charged anions migrate to the positively charged anode

1. The negatively charged cathode.
2. The positively charged anode.
3. Opposite charges attract.
4. Opposite charges attract.
5. Positive; the cathode.
6. Negative; the anode.

3.25 Explain the formation of the products in the electrolysis, using inert electrodes, of some electrolytes, including: a copper chloride solution b sodium chloride solution c sodium sulfate solution d water acidified with sulfuric acid e molten lead bromide (demonstration)

1. Cathode: copper. Anode: chlorine.
2. Cathode: hydrogen. Anode: chlorine.
3. Cathode: hydrogen. Anode: oxygen.
4. Cathode: hydrogen. Anode: oxygen.
5. Cathode: lead. Anode: bromine.
6. Hydrogen is preferentially discharged from water because sodium is more reactive than hydrogen.

3.26 Predict the products of electrolysis of other binary, ionic compounds in the molten state

1. The metal and non-metal elements.
2. The metal forms at the cathode and the halogen forms at the anode.
3. The metal.
4. The non-metal.
5. Sodium and chlorine.
6. Magnesium and bromine.

3.27 Write half equations for reactions occurring at the anode and cathode in electrolysis

1. An equation showing the electron transfer in one half of a redox reaction.
2. It shows the ions, electrons and products involved.
3. Reduction; electrons are gained.
4. Oxidation; electrons are lost.
5. $Cu^{2+} + 2e^- \rightarrow Cu$.
6. $2Cl^- \rightarrow Cl_2 + 2e^-$.

3.28 Explain oxidation and reduction in terms of loss or gain of electrons

1. Loss of electrons.
2. Gain of electrons.
3. They lose electrons.
4. They gain electrons.
5. Oxidation Is Loss; Reduction Is Gain.
6. Reduction.

3.29 Recall that reduction occurs at the cathode and that oxidation occurs at the anode in electrolysis reactions

1. Cathode.
2. Anode.
3. They gain electrons.
4. They lose electrons.
5. Cathode.

6. Anode.

3.30 Explain the formation of the products in the electrolysis of copper sulfate solution, using copper electrodes, and how this electrolysis can be used to purify copper

1. They gain electrons and are deposited as copper atoms: $Cu^{2+} + 2e^- \rightarrow Cu$.
2. Copper atoms lose electrons and form Cu^{2+} ions: $Cu \rightarrow Cu^{2+} + 2e^-$.
3. It increases.
4. It decreases.
5. Cu^{2+} ions removed at the cathode are replaced by Cu^{2+} ions formed at the anode.
6. Impure copper is used as the anode and pure copper is deposited at the cathode. Insoluble impurities fall to the bottom as sludge.

3.31 Core Practical: Investigate the electrolysis of copper sulfate solution with inert electrodes and copper electrodes

1. Place copper sulfate solution in a suitable container, insert inert electrodes, connect them to a DC supply and observe the electrodes during electrolysis.
2. Copper is deposited at the cathode and bubbles of oxygen form at the anode.
3. Repeat using copper electrodes connected to a DC supply and observe and measure changes at the electrodes.
4. The cathode gains mass and the anode loses mass.
5. Measure the mass of each electrode before and after electrolysis and compare the changes.
6. Wear safety goggles, use a low-voltage DC supply, avoid contact with the solution, and work in suitable ventilation when gases are produced.

4.1 Deduce the relative reactivity of some metals, by their reactions with water, acids and salt solutions

1. A more reactive metal reacts more readily with water, often producing hydrogen.
2. A more reactive metal reacts more vigorously with dilute acids, usually producing hydrogen.
3. Place each metal in the other metal's salt solution and observe whether a displacement reaction occurs.
4. The metal doing the displacing is more reactive.
5. Magnesium is more reactive than copper.
6. Magnesium is more reactive than copper.

4.2 Explain displacement reactions as redox reactions, in terms of gain or loss of electrons

1. A reaction in which a more reactive metal displaces a less reactive metal from its compound.
2. Electrons are transferred between the metals, so oxidation and reduction occur.
3. They lose electrons and are oxidised.
4. They gain electrons and are reduced.
5. Zinc is oxidised; copper ions are reduced.
6. Loss of electrons is oxidation; gain of electrons is reduction.

4.3 Explain the reactivity series of metals (potassium, sodium, calcium, magnesium, aluminium, (carbon), zinc, iron, (hydrogen), copper, silver, gold) in terms of the reactivity of the metals with water and dilute acids and that these reactions show the relative tendency of metal atoms to form cations

1. Potassium, sodium, calcium, magnesium, aluminium, carbon, zinc, iron, hydrogen, copper, silver, gold.
2. Metals higher in the series form cations more readily than metals lower in the series.
3. They lose electrons more readily.
4. More reactive metals react more readily with water.
5. More reactive metals react more readily and vigorously with dilute acids.
6. Carbon and hydrogen are reference points used to compare the reactivity of metals and determine extraction methods.

4.4 Recall that: a most metals are extracted from ores found in the Earth's crust b unreactive metals are found in the Earth's crust as the uncombined elements

1. In ores in the Earth's crust.
2. A naturally occurring rock containing enough metal compound for the metal to be extracted economically.
3. Most metals are reactive and react with other elements to form compounds.
4. Unreactive metals.
5. They do not readily react with other substances.
6. More reactive metals are generally found as compounds in ores, while very unreactive metals may occur as uncombined elements.

4.5 Explain oxidation as the gain of oxygen and reduction as the loss of oxygen

1. Gain of oxygen.
2. Loss of oxygen.
3. It is oxidised.
4. It is reduced.
5. H_2 is oxidised.
6. CuO is reduced.

4.6 Recall that the extraction of metals involves reduction of ores

1. The metal compound must be reduced to obtain the metal.
2. The metal ions must be converted into metal atoms.
3. By reacting the metal oxide with a reducing agent such as carbon.
4. To remove oxygen from the metal compound and leave the metal.
5. They gain electrons and become metal atoms.
6. Oxygen is removed from the metal oxide, leaving the metal.

4.7 Explain why the method used to extract a metal from its ore is related to its position in the reactivity series and the cost of the extraction process, illustrated by a heating with carbon (including iron) b electrolysis (including aluminium) (knowledge of the blast furnace is not required)

1. Metals below carbon can generally be extracted by reduction with carbon; metals above carbon require electrolysis.
2. Carbon is more reactive than these metals and can remove oxygen from their oxides.
3. Aluminium is more reactive than carbon, so carbon cannot remove oxygen from aluminium oxide.
4. Aluminium oxide must be electrolysed because aluminium is more reactive than carbon.
5. Extraction methods require different amounts of energy and resources, affecting the overall cost.
6. Electrolysis requires large amounts of electrical energy, whereas carbon reduction requires less expensive energy and reducing agents.

4.8 Evaluate alternative biological methods of metal extraction (bacterial and phytoextraction)

1. Using bacteria to convert metal compounds into soluble forms that can be extracted.
2. Bacteria can carry out reactions that make metal compounds soluble, allowing the metal to be recovered.
3. Using plants to absorb metal compounds from contaminated soil.
4. Plants absorb metal ions through their roots and concentrate them in their tissues, which can then be harvested.
5. They can use less energy and produce less environmental damage than traditional mining and extraction.
6. They can be slow, may produce lower yields, and depend on suitable environmental conditions.

4.9 Explain how a metal's relative resistance to oxidation is related to its position in the reactivity series

1. Metals higher in the reactivity series are less resistant to oxidation.
2. They lose electrons readily and therefore oxidise easily.
3. They lose electrons less readily.
4. Copper, because it is lower in the reactivity series and is more resistant to oxidation.
5. Gold is very unreactive and does not readily react with oxygen.
6. A greater tendency to form cations means a greater tendency to lose electrons and become oxidised.

4.10 Evaluate the advantages of recycling metals, including economic implications and how recycling can preserve both the environment and the supply of valuable raw materials

1. It reduces mining, habitat destruction, waste and pollution.
2. It reduces the amount of new ore that needs to be extracted.
3. Less mining is required, reducing habitat destruction and pollution.
4. Recycling often requires less energy than extracting metals from ores.
5. It saves raw materials, can reduce energy costs and provides useful materials for reuse.
6. Collection, sorting and processing can be expensive and recycling may require significant energy.

4.11 Describe that a life-cycle assessment for a product involves consideration of the effect on the environment of obtaining the raw materials, manufacturing the product, using the product and disposing of the product when it is no longer useful

1. An assessment of the environmental impacts of a product throughout its life.
2. Mining, extraction, habitat destruction, energy use and pollution.
3. Energy use, emissions, waste and pollution.
4. Energy consumption, emissions and other environmental effects during use.
5. Waste produced, recycling, reuse and disposal impacts.
6. Obtaining raw materials, manufacturing, using the product and disposing of it.

4.12 Evaluate data from a life cycle assessment of a product

1. Compare the environmental impacts at each stage and the overall impact.
2. The environmental impacts and resource use at all stages of the product's life.
3. Obtaining raw materials can involve mining, energy use, habitat destruction and pollution.

4. The overall environmental impact depends on all stages of the product's life.
5. Compare the total environmental impacts shown by the data.
6. A greater impact at another stage may outweigh the lower impact at one stage.

4.13 Recall that chemical reactions are reversible, the use of the symbol ⇌ in equations and that the direction of some reversible reactions can be altered by changing the reaction conditions

1. A reaction in which the products can react to reform the original reactants.
2. ⇌.
3. The reaction can proceed in both directions.
4. Reactants form products.
5. Products react to form reactants.
6. Changing temperature, pressure or concentration can alter the position/direction of equilibrium.

4.14 Explain what is meant by dynamic equilibrium

1. A state in a closed system where the forward and reverse reactions occur at equal rates.
2. It continues to occur.
3. It continues to occur.
4. The reactants and products are being formed and used at equal rates.
5. Both reactions continue, but at equal rates.
6. The forward and reverse reaction rates are equal.

4.15 Describe the formation of ammonia as a reversible reaction between nitrogen (extracted from the air) and hydrogen (obtained from natural gas) and that it can reach a dynamic equilibrium

1. Nitrogen and hydrogen.
2. From the air.
3. From natural gas.
4. $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$.
5. The reaction is reversible, so ammonia can react to reform nitrogen and hydrogen.
6. Their concentrations remain constant because the forward and reverse reactions occur at equal rates.

4.16 Recall the conditions for the Haber process as: a temperature 450 °C b pressure 200 atmospheres c iron catalyst

1. 450 °C.

2. 200 atmospheres.
3. Iron.
4. 450 °C, 200 atmospheres and an iron catalyst.
5. It increases the rate of reaction without being used up.
6. These conditions provide a compromise between reaction rate, ammonia yield and industrial cost.

4.17 Predict how the position of a dynamic equilibrium is affected by changes in: a temperature b pressure c concentration

1. It shifts to the left, towards nitrogen and hydrogen.
2. It shifts to the right, towards ammonia.
3. It shifts to the right, towards ammonia.
4. It shifts to the left, towards nitrogen and hydrogen.
5. It shifts towards the products, using up some of the added reactant.
6. It shifts towards the products, replacing some of the product that was removed.

6.1 Explain why some elements can be classified as alkali metals (group 1), halogens (group 7) or noble gases (group 0), based on their position in the periodic table

1. Group 1.
2. Group 7.
3. Group 0.
4. Elements in the same group have similar chemical properties.
5. It tells you which group the element belongs to and therefore its classification.
6. They are all elements in Group 1.

6.2 Recall that alkali metals a are soft b have relatively low melting points

1. They are soft.
2. They have relatively low melting points.
3. They can be cut relatively easily.
4. It melts and changes from a solid to a liquid.
5. They are soft and have relatively low melting points.
6. They are softer and have lower melting points than many other metals.

6.3 Describe the reactions of lithium, sodium and potassium with water

1. Lithium floats, moves around on the surface and fizzes as it reacts, forming lithium hydroxide and hydrogen.
2. Sodium melts into a ball, moves rapidly on the surface and fizzes, forming sodium hydroxide and hydrogen.

3. Potassium reacts very vigorously, often igniting with a lilac flame, forming potassium hydroxide and hydrogen.
4. A metal hydroxide and hydrogen.
5. It floats, melts into a ball, moves around, fizzes and gradually disappears.
6. It reacts very vigorously, moves rapidly, fizzes and usually burns with a lilac flame.

6.4 Describe the pattern in reactivity of the alkali metals, lithium, sodium and potassium, with water; and use this pattern to predict the reactivity of other alkali metals

1. Reactivity increases down Group 1.
2. Sodium.
3. Potassium.
4. Rubidium would be more reactive than potassium.
5. Caesium would be even more reactive than rubidium.
6. The reaction becomes faster and more violent.

6.5 Explain this pattern in reactivity in terms of electronic configurations

1. The outer electron becomes easier to remove.
2. It is further from the nucleus.
3. Potassium has a larger atomic radius and greater electron shielding, so its outer electron is less strongly attracted to the nucleus.
4. The attraction between the nucleus and outer electron decreases.
5. Shielding reduces the attraction between the nucleus and the outer electron.
6. Group 1 metals have one outer electron, which becomes easier to lose down the group, increasing reactivity.

6.6 Recall the colours and physical states of chlorine, bromine and iodine at room temperature

1. Pale green gas.
2. Red-brown liquid.
3. Grey-black solid.
4. Pale green.
5. Red-brown.
6. Grey-black.

6.7 Describe the pattern in the physical properties of the halogens, chlorine, bromine and iodine, and use this pattern to predict the physical properties of other halogens

1. They change from gas to liquid to solid.

2. The colour becomes darker.
3. Their melting and boiling points increase.
4. Astatine is predicted to be a solid.
5. The halogen would be predicted to become darker down the group.
6. Melting point, boiling point and density generally increase, while the physical state changes from gas to liquid to solid.

6.8 Describe the chemical test for chlorine

1. Expose damp litmus paper to the gas.
2. It is bleached.
3. Blue litmus turns red and is then bleached white.
4. Chlorine reacts with water to form substances that bleach the dye.
5. Chlorine must dissolve in water for the bleaching reaction to occur.
6. Damp litmus paper is bleached.

6.9 Describe the reactions of the halogens, chlorine, bromine and iodine, with metals to form metal halides, and use this pattern to predict the reactions of other halogens

1. Metal halides.
2. It reacts with the metal to form a metal chloride.
3. It reacts with the metal to form a metal bromide.
4. It reacts with the metal to form a metal iodide.
5. **A halide ion with a −1 charge.**
6. Other halogens are predicted to react with metals to form metal halides.

6.10 Recall that the halogens, chlorine, bromine and iodine, form hydrogen halides which dissolve in water to form acidic solutions, and use this pattern to predict the reactions of other halogens

1. Hydrogen chloride, hydrogen bromide and hydrogen iodide.
2. Hydrogen halides.
3. It dissolves and forms an acidic solution.
4. They produce hydrogen ions in water.
5. Hydrobromic acid.
6. Another halogen is predicted to react with hydrogen to form a hydrogen halide that dissolves in water to form an acidic solution.

6.11 Describe the relative reactivity of the halogens chlorine, bromine and iodine, as shown by their displacement reactions with halide ions in aqueous solution, and use this pattern to predict the reactions of astatine

1. Reactivity decreases down the group: chlorine > bromine > iodine.
2. Chlorine displaces bromide ions, producing bromine.

3. Chlorine displaces iodide ions, producing iodine.
4. Bromine displaces iodide ions, producing iodine.
5. Chlorine is more reactive and gains electrons more readily than bromine and iodine.
6. Astatine would be less reactive than iodine and would not displace chloride, bromide or iodide ions.

6.12 Explain why these displacement reactions are redox reactions in terms of gain and loss of electrons, identifying which of the substances are oxidised and which are reduced

1. Electrons are transferred, so oxidation and reduction both occur.
2. They lose electrons and are oxidised.
3. They gain electrons and are reduced.
4. Br^- is oxidised.
5. Cl_2 is reduced.
6. Loss of electrons is oxidation; gain of electrons is reduction.

6.13 Explain the relative reactivity of the halogens in terms of electronic configurations

1. The outer shell is further from the nucleus and electron shielding increases, making it harder to attract an incoming electron.
2. It increases.
3. It increases.
4. Chlorine has a smaller atomic radius and less electron shielding, so its nucleus attracts an incoming electron more strongly.
5. The attraction decreases.
6. Chlorine attracts an incoming electron more strongly because its outer shell is closer to the nucleus and there is less shielding.

6.14 Explain why the noble gases are chemically inert, compared with the other elements, in terms of their electronic configurations

1. They have full outer electron shells.
2. Their outer electron shell is full.
3. A full outer shell is stable, so atoms have little tendency to gain, lose or share electrons.
4. Their outer shell is already full.
5. Their outer shell is already stable and full.
6. Their full outer shells make them very stable and therefore chemically unreactive.

6.15 Explain how the uses of noble gases depend on their inertness, low density and/or non-flammability

1. Helium has a low density, so it provides lift.
2. Helium is non-flammable, making it safer than hydrogen.
3. Argon is inert, so it prevents the hot filament from reacting.
4. It prevents reactions with the hot filament and extends the bulb's life.
5. They can be used where a gas must not readily burn or support combustion.
6. Their properties determine suitable uses: low density for balloons, inertness for protective atmospheres and non-flammability where safety is important.

6.16 Describe the pattern in the physical properties of some noble gases and use this pattern to predict the physical properties of other noble gases

1. They increase down Group 0.
2. They increase down Group 0.
3. They remain gases at room temperature.
4. Krypton would be a gas with a higher boiling point and density than argon.
5. Its density would be greater than that of lighter noble gases.
6. Boiling point and density generally increase down the group.

Topic 7 – Rates of reaction and energy changes

Rates of reaction

7.1 Core Practical: Investigate the effects of changing the conditions of a reaction on the rates of chemical reactions by: a measuring the production of a gas (in the reaction between hydrochloric acid and marble chips) b observing a colour change (in the reaction between sodium thiosulfate and hydrochloric acid)

1. Use different concentrations of hydrochloric acid, keeping the mass and surface area of marble chips, volume, temperature and apparatus constant. Measure the volume of CO_2 produced over time.
2. Measure the volume of carbon dioxide gas produced at regular time intervals.
3. Mix sodium thiosulfate with hydrochloric acid over a marked cross and measure the time taken for the cross to become obscured.
4. Sulfur precipitate forms, making the solution cloudy.
5. Compare the gradients of graphs or calculate the rate using change in measured quantity ÷ time. The greater rate is the faster reaction.

6. Keep all variables constant except the condition being investigated, for example temperature, volumes, concentrations, mass/surface area and apparatus.

7.2 Suggest practical methods for determining the rate of a given reaction

1. Collect the gas in a gas syringe or measure its volume by displacement of water over time.
2. Measure the time taken for a precipitate to form to a specified opacity or to obscure a marked cross.
3. Measure the time taken for a specified colour change to occur.
4. Measure the mass of the reaction mixture at regular time intervals using a balance.
5. Calculate the gradient of the graph at the required point. A steeper gradient represents a faster rate.
6. Consider what measurable change occurs, the required accuracy, reaction speed, suitable apparatus and whether the method gives reliable results.

7.3 Explain how reactions occur when particles collide and that rates of reaction are increased when the frequency and/or energy of collisions is increased

1. Reacting particles must collide so that bonds can break and new bonds can form.
2. A successful collision is one with sufficient energy, and the appropriate conditions/orientation, to cause a reaction.
3. They may have insufficient energy to overcome the activation energy.
4. More frequent collisions increase the number of opportunities for successful collisions, increasing the rate.
5. Higher-energy collisions are more likely to have enough energy to overcome the activation energy, increasing the rate.
6. A faster reaction has more successful collisions per unit time.

7.4 Explain the effects on rates of reaction of changes in temperature, concentration, surface area to volume ratio of a solid and pressure (on reactions involving gases) in terms of frequency and/or energy of collisions between particles

1. Particles have more kinetic energy, move faster and collide more frequently. A greater proportion also have enough energy to overcome the activation energy.
2. There are more reactant particles per unit volume, so collisions occur more frequently.
3. More particles are exposed at the surface, so collisions between reactant particles occur more frequently.

4. Gas particles are closer together, increasing the frequency of collisions.
5. Increasing temperature increases the frequency of collisions and the average energy of the particles, so more collisions have enough energy to be successful.
6. The particles are closer together or more exposed, increasing the frequency of collisions.

7.5 Interpret graphs of mass, volume or concentration of reactant or product against time

1. It represents the rate of gas production.
2. A steep gradient indicates a fast rate of reaction.
3. The reaction has stopped because the reactant concentration is no longer changing.
4. Mean rate = $60 \div 30 = 2\ cm^3\ s^{-1}$.
5. Mean rate = $90 \div 45 = 2\ cm^3\ s^{-1}$.
6. The reaction with the steeper initial gradient has the greater initial rate because the measured quantity changes more rapidly at the start.

7.6 Describe a catalyst as a substance that speeds up the rate of a reaction without altering the products of the reaction, being itself unchanged chemically and in mass at the end of the reaction

1. A catalyst is a substance that increases the rate of a reaction without being chemically changed or used up.
2. It increases the rate of reaction.
3. No. It does not alter the products formed.
4. It remains chemically unchanged overall.
5. Its mass remains unchanged.
6. Because it is not used up or chemically changed overall during the reaction.

7.7 Explain how the addition of a catalyst increases the rate of a reaction in terms of activation energy

1. Activation energy is the minimum energy particles must have for a reaction to occur.
2. A catalyst lowers the activation energy.
3. More particles have enough energy to overcome the lower activation energy, increasing the number of successful collisions.
4. It provides an alternative reaction pathway with a lower activation energy.
5. More collisions have sufficient energy to overcome the lower activation energy, so more are successful.
6. The catalysed reaction profile has a lower activation-energy peak than the uncatalysed profile.

7.8 Recall that enzymes are biological catalysts and that enzymes are used in the production of alcoholic drinks

1. Enzymes are biological catalysts.
2. They speed up reactions in living organisms without being used up.
3. They increase the rates of biochemical reactions.
4. Enzymes catalyse reactions involved in fermentation.
5. Yeast enzymes catalyse the conversion of sugars into ethanol and carbon dioxide.
6. They allow biological reactions to occur rapidly under relatively mild conditions.

Heat energy changes in chemical reactions

7.9 Recall that changes in heat energy accompany the following changes: a salts dissolving in water b neutralisation reactions c displacement reactions d precipitation reactions and that, when these reactions take place in solution, temperature changes can be measured to reflect the heat changes

1. Heat energy may be absorbed or released.
2. Chemical reactions involve energy changes, which can cause the temperature of the solution to change.
3. Energy is transferred during the reaction, causing the temperature of the solution to change.
4. Energy is transferred when the precipitate forms, causing a temperature change.
5. A temperature increase indicates heat has been released; a temperature decrease indicates heat has been absorbed.
6. Dissolving salts, neutralisation, displacement and precipitation reactions.

7.10 Describe an exothermic change or reaction as one in which heat energy is given out

1. An exothermic reaction is a reaction in which heat energy is given out.
2. Heat energy is transferred to the surroundings.
3. The temperature of the surroundings increases.
4. Heat energy is released from the reacting system into the surroundings.
5. A temperature increase shows that heat energy has been transferred to the surroundings.
6. Neutralisation of an acid with an alkali.

7.11 Describe an endothermic change or reaction as one in which heat energy is taken in

1. An endothermic reaction is a reaction in which heat energy is taken in.
2. Heat energy is absorbed from the surroundings.
3. The temperature of the surroundings decreases.
4. Heat energy is transferred from the surroundings into the reacting system.
5. A temperature decrease indicates that heat energy has been absorbed.
6. The thermal decomposition of calcium carbonate.

7.12 Recall that the breaking of bonds is endothermic and the making of bonds is exothermic

1. Energy is required to overcome the forces holding the atoms together.
2. Energy is taken in to break the bond.
3. Energy is released.
4. Forming a bond releases energy as the atoms become more stable.
5. Making bonds.
6. Breaking bonds.

7.13 Recall that the overall heat energy change for a reaction is: a exothermic if more heat energy is released in forming bonds in the products than is required in breaking bonds in the reactants b endothermic if less heat energy is released in forming bonds in the products than is required in breaking bonds in the reactants

1. More energy is released forming product bonds than is required to break reactant bonds.
2. Less energy is released forming product bonds than is required to break reactant bonds.
3. Exothermic, because 1100 kJ mol^{-1} is released and only 800 kJ mol^{-1} is required.
4. Endothermic, because 1400 kJ mol^{-1} is required and only 1000 kJ mol^{-1} is released.
5. The excess energy is transferred to the surroundings as heat.
6. More energy is required to break the reactant bonds than is released when the product bonds form, so energy is taken in overall.

7.14 Calculate the energy change in a reaction given the energies of bonds (in kJ mol–1)

1. Bonds broken = 436 + 242 = 678 kJ mol^{-1}
   Bonds formed = 2 × 431 = 862 kJ mol^{-1}
   Energy change = 678 – 862 = –184 kJ mol^{-1}.
2. Bonds broken = (2 × 436) + 498 = 1370 kJ mol^{-1}
   Bonds formed = 4 × 463 = 1852 kJ mol^{-1}
   Energy change = 1370 – 1852 = –482 kJ mol^{-1}.

3. **Energy change = 1250 − 1500 = −250 kJ mol^{-1}.**
   The negative value means the reaction is exothermic.
4. **Energy change = 980 − 760 = +220 kJ mol^{-1}.**
   The positive value means the reaction is endothermic.
5. **Energy change = 1650 − 1920 = −270 kJ mol^{-1}.**
   The negative sign means 270 kJ mol^{-1} of heat energy is released overall, so the reaction is exothermic.
6. Energy released = 1350 + 240 = 1590 kJ mol^{-1}.

7.15 Explain the term activation energy

1. The minimum energy required for a reaction to occur.
2. Particles need sufficient energy to overcome the activation energy and allow bonds to break and new bonds to form.
3. A successful collision must have at least the activation energy.
4. The rate decreases because fewer collisions have sufficient energy to be successful.
5. Particles have greater kinetic energy, so a greater proportion have energy equal to or above the activation energy.
6. A catalyst lowers the activation energy.

7.16 Draw and label reaction profiles for endothermic and exothermic reactions, identifying activation energy

1. A reaction profile shows the energy level of reactants and products during a reaction.
2. It is shown as the energy difference between the reactants and the peak of the profile.
3. The products are at a lower energy level than the reactants.
4. The products are at a higher energy level than the reactants.
5. Reactants are labelled at the initial energy level, products at the final energy level and activation energy from the reactant level to the peak.
6. Exothermic: products have lower energy than reactants.
   Endothermic: products have higher energy than reactants.

Topic 8 – Fuels and Earth science

Fuels

8.1 Recall that hydrocarbons are compounds that contain carbon and hydrogen only

1. Carbon and hydrogen only.

2. It contains only carbon and hydrogen atoms.
3. Yes. CH_4 contains only carbon and hydrogen.
4. No. Ethanol also contains oxygen.
5. No. Carbon dioxide contains carbon and oxygen, not hydrogen.
6. C_3H_8 — hydrocarbon; C_2H_4 — hydrocarbon; C_6H_6 — hydrocarbon. Each contains only carbon and hydrogen.

8.2 Describe crude oil as: a a complex mixture of hydrocarbons b containing molecules in which carbon atoms are in chains or rings (names, formulae and structures of specific ring molecules not required) c an important source of useful substances (fuels and feedstock for the petrochemical industry) d a finite resource

1. Crude oil is a complex mixture of hydrocarbons.
2. It contains many different hydrocarbon molecules with different sizes and structures.
3. Carbon atoms can be arranged in chains or rings.
4. It provides fuels and feedstock for the petrochemical industry.
5. It is a limited resource that will eventually run out.
6. It is important because it provides fuels and chemical feedstocks, but it is finite because its supply is limited and cannot be replaced on a human timescale.

8.3 Describe and explain the separation of crude oil into simpler, more useful mixtures by the process of fractional distillation

1. Fractional distillation.
2. They have different boiling points.
3. It is heated so that most of the hydrocarbons vaporise.
4. They condense at different temperatures as the vapours rise through the column.
5. The column is hotter at the bottom and cooler at the top, so hydrocarbons condense at different levels according to their boiling points.
6. Crude oil is vaporised and passed into a fractionating column, where hydrocarbons condense at different temperatures, separating them into fractions with similar boiling points.

8.4 Recall the names and uses of the following fractions: a gases, used in domestic heating and cooking b petrol, used as fuel for cars c kerosene, used as fuel for aircraft d diesel oil, used as fuel for some cars and trains e fuel oil, used as fuel for large ships and in some power stations f bitumen, used to surface roads and roofs

1. Domestic heating and cooking.
2. Fuel for cars.

3. Fuel for aircraft.
4. Fuel for some cars and trains.
5. Fuel for large ships and some power stations.
6. To surface roads and roofs.

8.5 Explain how hydrocarbons in different fractions differ from each other in: a the number of carbon and hydrogen atoms their molecules contain b boiling points c ease of ignition d viscosity and are mostly members of the alkane homologous series

1. The molecules generally contain more carbon and hydrogen atoms.
2. Boiling points increase.
3. Larger hydrocarbons are harder to ignite.
4. Viscosity increases.
5. The alkane homologous series.
6. The fraction with higher boiling points and greater viscosity contains larger molecules.

8.6 Explain an homologous series as a series of compounds which: a have the same general formula b differ by CH2 in molecular formulae from neighbouring compounds c show a gradual variation in physical properties, as exemplified by their boiling points d have similar chemical properties

1. A homologous series is a series of compounds with the same general formula and similar chemical properties, with neighbouring members differing by CH_2.
2. They differ by CH_2.
3. They share the same general formula.
4. Physical properties change gradually, for example boiling point increases with molecular size.
5. They have the same functional group and similar bonding, giving them similar chemical behaviour.
6. The gradual increase in boiling point shows a gradual change in physical properties as molecular size increases.

8.7 Describe the complete combustion of hydrocarbon fuels as a reaction in which: a carbon dioxide and water are produced b energy is given out

1. Carbon dioxide and water.
2. Oxygen.
3. They are oxidised to form carbon dioxide.
4. They are oxidised to form water.
5. Exothermic.
6. Hydrocarbon + oxygen → carbon dioxide + water.

8.8 Explain why the incomplete combustion of hydrocarbons can produce carbon and carbon monoxide

1. There is insufficient oxygen for complete combustion.
2. There is insufficient oxygen to fully oxidise the carbon to carbon dioxide.
3. Some carbon is left unoxidised when there is insufficient oxygen.
4. A plentiful oxygen supply favours complete combustion; a restricted supply causes incomplete combustion.
5. Carbon and carbon monoxide.
6. There is not enough oxygen to completely oxidise all the carbon to carbon dioxide, so carbon monoxide and/or carbon can form.

8.9 Explain how carbon monoxide behaves as a toxic gas

1. It binds strongly to haemoglobin, reducing its ability to carry oxygen.
2. It binds to haemoglobin and forms carboxyhaemoglobin.
3. Carbon monoxide occupies haemoglobin binding sites that would normally carry oxygen.
4. It is colourless and odourless, so dangerous concentrations may go unnoticed.
5. Cells receive less oxygen for aerobic respiration.
6. Severe oxygen deprivation can prevent vital organs and cells from respiring properly, potentially causing death.

8.10 Describe the problems caused by incomplete combustion producing carbon monoxide and soot in appliances that use carbon compounds as fuels

1. Carbon monoxide.
2. Soot/carbon particles.
3. Carbon monoxide can build up to dangerous concentrations and reduce oxygen transport in the blood.
4. Soot particles can damage the respiratory system and contribute to breathing problems.
5. They can release toxic carbon monoxide and harmful soot.
6. Ventilation provides enough oxygen for combustion and prevents dangerous carbon monoxide from accumulating.

8.11 Explain how impurities in some hydrocarbon fuels result in the production of sulfur dioxide

1. Sulfur-containing impurities.
2. They react with oxygen during combustion.
3. Sulfur dioxide, SO_2.
4. Sulfur + oxygen → sulfur dioxide.
5. Sulfur-containing impurities are oxidised during combustion.

6. Sulfur impurities in fuels burn and react with oxygen to form sulfur dioxide, which is released into the atmosphere.

8.12 Explain some problems associated with acid rain caused when sulfur dioxide dissolves in rain water

1. It dissolves and forms an acidic solution.
2. The resulting acidic water contributes to acid rain.
3. It can lower the pH of water, harming aquatic organisms.
4. It can damage plants and trees.
5. Acids react with and dissolve carbonate-containing materials.
6. Sulfur-containing fuels produce SO_2; SO_2 dissolves in rainwater, producing acidic rain that can damage ecosystems, vegetation and buildings.

8.13 Explain why, when fuels are burned in engines, oxygen and nitrogen can react together at high temperatures to produce oxides of nitrogen, which are pollutants

1. Nitrogen and oxygen.
2. The high temperature provides enough energy for them to react.
3. Oxides of nitrogen, or nitrogen oxides.
4. They are harmful air pollutants and can contribute to problems such as acid rain and respiratory irritation.
5. More particles have sufficient energy to overcome the activation energy.
6. Burning fuel produces very high temperatures inside the engine, allowing nitrogen and oxygen from the air to react and form nitrogen oxides.

8.14 Evaluate the advantages and disadvantages of using hydrogen, rather than petrol, as a fuel in cars

1. Hydrogen produces no carbon dioxide at the point of use.
2. Water.
3. Hydrogen contains no carbon, so its reaction with oxygen produces water rather than carbon dioxide.
4. Hydrogen can be difficult and energy-intensive to produce, and it is difficult to store because it has low density.
5. If hydrogen is produced using fossil fuels, carbon dioxide can be produced during its manufacture; renewable production has lower associated emissions.
6. Hydrogen can reduce direct carbon emissions and produces water in a fuel cell, but production, storage, transport, infrastructure and availability present disadvantages. Its overall environmental benefit depends on how it is produced.

8.15 Recall that petrol, kerosene and diesel oil are non-renewable fossil fuels obtained from crude oil and methane is a nonrenewable fossil fuel found in natural gas

1. Crude oil.
2. Crude oil.
3. Crude oil.
4. Natural gas.
5. They are finite resources formed over very long periods and cannot be replaced quickly.
6. Petrol — crude oil; kerosene — crude oil; diesel oil — crude oil; methane — natural gas.

8.16 Explain why cracking involves the breaking down of larger, saturated hydrocarbon molecules (alkanes) into smaller, more useful ones, some of which are unsaturated (alkenes)

1. Large, saturated hydrocarbons called alkanes.
2. They are broken down into smaller hydrocarbon molecules.
3. Cracking breaks large molecules into smaller molecules.
4. Hydrogen atoms can be removed during cracking, leaving carbon–carbon double bonds.
5. An unsaturated hydrocarbon containing a carbon–carbon double bond.
6. Breaking large alkanes can produce smaller alkanes and alkenes, with some products having fewer hydrogen atoms and therefore containing C=C bonds.

8.17 Explain why cracking is necessary

1. To convert less useful large hydrocarbons into more useful smaller hydrocarbons.
2. There is greater demand for smaller, more easily ignited fractions than for some larger fractions.
3. It produces additional smaller hydrocarbons.
4. They are used to make polymers and other useful chemicals.
5. It converts surplus larger hydrocarbons into products with greater demand and usefulness.
6. Fractional distillation produces fractions in proportions that do not necessarily match demand, so cracking changes larger molecules into more useful products.

Earth and atmospheric science

8.18 Recall that the gases produced by volcanic activity formed the Earth's early atmosphere

1. Volcanic activity.
2. Volcanoes released gases into the atmosphere.
3. They accumulated around the Earth to form the early atmosphere.
4. Its composition reflected the gases released by volcanic activity and differed greatly from the modern atmosphere.
5. Volcanic activity.
6. Volcanic eruptions released gases such as water vapour and carbon dioxide, which accumulated to form Earth's early atmosphere.

8.19 Describe that the Earth's early atmosphere was thought to contain: a little or no oxygen b a large amount of carbon dioxide c water vapour d small amounts of other gases and interpret evidence relating to this

1. Little or no oxygen.
2. Carbon dioxide.
3. Volcanic activity released large amounts of water vapour.
4. Small amounts of other gases, such as nitrogen.
5. Evidence from rocks, fossils and geological records can be used to infer past atmospheric conditions.
6. Geological evidence can indicate that early Earth had little or no oxygen and substantial carbon dioxide, supporting the model of the early atmosphere.

8.20 Explain how condensation of water vapour formed oceans

1. The Earth cooled.
2. Water vapour changed into liquid water.
3. Cooling reduced the temperature enough for water vapour to condense.
4. Liquid water accumulated on Earth's surface, forming oceans.
5. Gas → liquid.
6. As the early Earth cooled, water vapour condensed into liquid water, which accumulated to form the oceans.

8.21 Explain how the amount of carbon dioxide in the atmosphere was decreased when carbon dioxide dissolved as the oceans formed

1. Some atmospheric carbon dioxide dissolved in the newly formed oceans.
2. Carbon dioxide was transferred from the atmosphere into the oceans.
3. It dissolves in water and can form dissolved carbon dioxide and carbonic acid.
4. The atmosphere became less rich in carbon dioxide.
5. As oceans formed, increasing amounts of CO_2 dissolved in them, reducing its atmospheric concentration.

6. Cooling formed liquid oceans; atmospheric CO_2 dissolved into the water, reducing the amount of CO_2 remaining in the atmosphere.

8.22 Explain how the growth of primitive plants used carbon dioxide and released oxygen by photosynthesis and consequently the amount of oxygen in the atmosphere gradually increased

1. Carbon dioxide.
2. Oxygen.
3. Atmospheric oxygen gradually increased.
4. Photosynthesis occurred over a long period as primitive plants grew and accumulated.
5. Photosynthesis.
6. Primitive plants removed CO_2 by photosynthesis and released O_2, causing atmospheric oxygen levels to gradually increase.

8.23 Describe the chemical test for oxygen

1. Insert a glowing splint into the gas.
2. The glowing splint relights.
3. Relighting of the glowing splint.
4. Oxygen supports combustion, so it can relight a glowing splint without the splint already being aflame.
5. Oxygen.
6. Place a glowing splint into the unknown gas. If it relights, oxygen is present.

8.24 Describe how various gases in the atmosphere, including carbon dioxide, methane and water vapour, absorb heat radiated from the Earth, subsequently releasing energy which keeps the Earth warm: this is known as the greenhouse effect

1. The greenhouse effect is the warming of Earth caused by atmospheric gases absorbing and re-emitting radiation from Earth.
2. Infrared radiation.
3. Carbon dioxide, methane and water vapour.
4. They subsequently release/re-emit the absorbed energy.
5. Energy is retained in the Earth-atmosphere system, keeping the Earth warmer.
6. Earth emits infrared radiation; greenhouse gases absorb some of this radiation and re-emit energy, helping keep the Earth warm.

8.25 Evaluate the evidence for human activity causing climate change, considering: a the correlation between the change in atmospheric carbon dioxide concentration, the consumption of fossil fuels and temperature change

b the uncertainties caused by the location where these measurements are taken and historical accuracy

1. Atmospheric CO_2 concentration and global temperature have increased over the relevant period, showing a correlation.
2. Burning fossil fuels releases CO_2, increasing atmospheric CO_2 concentration.
3. The linked increases provide evidence that human activity may be contributing to climate change.
4. Correlation shows an association but does not by itself establish that one factor causes the other.
5. Different locations can experience different temperatures and CO_2 concentrations, so measurements may not represent global conditions.
6. Increasing atmospheric CO_2, fossil fuel consumption and temperature provide evidence supporting a human contribution to climate change. However, uncertainty arises from measurement locations, limited historical data and the accuracy of older records.

8.26 Describe: a the composition of today's atmosphere b the potential effects on the climate of increased levels of carbon dioxide and methane generated by human activity, including burning fossil fuels and livestock farming c that these effects may be mitigated: consider scale, risk and environmental implications

1. Approximately 78% nitrogen, 21% oxygen and about 1% other gases, including argon and carbon dioxide.
2. Burning fossil fuels and deforestation can increase atmospheric CO_2.
3. Livestock farming produces methane, particularly from ruminant digestion and waste.
4. Increased greenhouse gas concentrations can enhance the greenhouse effect and contribute to global warming and climate change.
5. Mitigating climate change means reducing or limiting its causes and/or effects.
6. Mitigation can include reducing fossil fuel use, increasing renewable energy, improving energy efficiency and reducing methane emissions. Methods must be considered in terms of their scale, risks, costs and environmental effects.