Edexcel GCSE Triple 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. How did Dalton's original model describe the structure of an atom?
2. Which subatomic particle was discovered that showed atoms were not indivisible?
3. How did the discovery of the electron change Dalton's model of the atom?
4. Which subatomic particles were discovered after the electron?
5. How did the discovery of the nucleus change the model of the atom?
6. Why is the modern atomic model different from Dalton's original model?

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

1. What three types of subatomic particle make up an atom?
2. Which subatomic particles are found in the nucleus of an atom?
3. Where are electrons located in an atom?
4. What charge does the nucleus of an atom have?
5. Why does an atom have no overall electrical charge?
6. How are electrons arranged around the nucleus of an atom?

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

1. What is the relative charge of a proton?
2. What is the relative mass of a proton?
3. What is the relative charge of a neutron?
4. What is the relative mass of a neutron?
5. What is the relative charge of an electron?
6. What is the relative mass of an electron?

1.4 Explain why atoms contain equal numbers of protons and electrons

1. Why does a neutral atom contain equal numbers of protons and electrons?
2. What charge does a proton have?
3. What charge does an electron have?
4. What would happen to the overall charge of an atom if it gained an electron?
5. What would happen to the overall charge of an atom if it lost an electron?
6. How do equal numbers of protons and electrons make an atom electrically neutral?

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

1. How does the size of an atomic nucleus compare with the overall size of an atom?
2. What proportion of the overall size of an atom is occupied by its nucleus?
3. Why is most of an atom considered to be empty space?
4. Which part of an atom occupies most of its overall volume?
5. How does the size of the nucleus compare with the distance of the electrons from the nucleus?
6. What does the small size of the nucleus tell us about the structure of an atom?

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

1. Where is most of the mass of an atom concentrated?
2. Which two subatomic particles account for almost all of an atom's mass?
3. Why does the electron contribute very little to the mass of an atom?
4. How does the mass of a proton compare with the mass of an electron?
5. How does the mass of a neutron compare with the mass of an electron?
6. Why is the nucleus much more massive than the surrounding electron shells?

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

1. What does the mass number of an atom represent?
2. How is the mass number of an atom calculated from its subatomic particles?
3. Which two types of subatomic particle are included in the mass number?
4. Why are electrons not included when calculating the mass number?
5. How can the number of neutrons in an atom be calculated from its mass number and atomic number?
6. What information does the mass number give about the nucleus of an atom?

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. What determines which element an atom belongs to?
2. What is unique about the number of protons in atoms of each element?
3. Why do all atoms of the same element have the same atomic number?
4. How can the number of protons identify an element?
5. What would happen to an atom's identity if the number of protons in its nucleus changed?
6. How does the number of protons distinguish one element from another?

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. What is an isotope?
2. Why do isotopes of the same element have the same chemical identity?
3. What is the same in the nuclei of all isotopes of an element?
4. What is different in the nuclei of isotopes of the same element?
5. How does changing the number of neutrons affect the mass number of an isotope?
6. Why can isotopes of the same element have different masses?

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

1. How can the number of protons in an atom be determined from its atomic number?
2. How can the number of electrons in a neutral atom be determined from its atomic number?
3. How can the number of neutrons in an atom be calculated from its mass number and atomic number?
4. How many protons, neutrons and electrons are in an atom with an atomic number of 8 and a mass number of 16?
5. How many protons, neutrons and electrons are in an atom with an atomic number of 11 and a mass number of 23?
6. How many protons, neutrons and electrons are in an atom with an atomic number of 17 and a mass number of 35?

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

1. Why are the relative atomic masses of some elements not whole numbers?
2. How does the existence of isotopes affect the relative atomic mass of an element?
3. Why is the relative atomic mass of an element a weighted average?

4. How does the abundance of each isotope affect the relative atomic mass?
5. Why does the relative atomic mass of chlorine differ from the mass number of either of its common isotopes?
6. Why is the relative atomic mass of an element usually different from the mass number of an individual atom?

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

1. How is the relative atomic mass of an element calculated from its isotopes?
2. What information about isotopes is needed to calculate relative atomic mass?
3. Why must isotope masses be multiplied by their relative abundances when calculating relative atomic mass?
4. How would you calculate the relative atomic mass of an element containing two isotopes with known masses and abundances?
5. An element has two isotopes with masses of 10 and 11 in abundances of 20% and 80%; what is its relative atomic mass?
6. An element has isotopes of masses 35 and 37 with abundances of 75% and 25%; what is its relative atomic mass?

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. How did Dmitri Mendeleev arrange the elements in his periodic table?
2. Which properties did Mendeleev use when arranging the elements?
3. Why did Mendeleev consider the properties of compounds when developing his periodic table?
4. What did Mendeleev notice about the properties of elements when arranging them?
5. How did Mendeleev's arrangement allow elements with similar properties to be grouped together?
6. Why was Mendeleev's periodic table different from simply arranging elements alphabetically?

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

1. How did Mendeleev use gaps in his periodic table?
2. Why did Mendeleev leave gaps in his periodic table?
3. How could Mendeleev predict the properties of undiscovered elements?

4. What did the successful discovery of predicted elements provide evidence for?
5. Why was Mendeleev able to predict the properties of elements that had not yet been discovered?
6. How did the discovery of elements that matched Mendeleev's predictions support his periodic table?

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. Why did Mendeleev initially think he had arranged elements in order of increasing relative atomic mass?
2. Why is the order of some elements in the periodic table not exactly the order of their relative atomic masses?
3. How can isotopes affect the relative atomic mass of an element?
4. Why can an element with a higher atomic number have a lower relative atomic mass than the element before it?
5. How did the existence of isotopes explain anomalies in Mendeleev's arrangement?
6. Why is atomic number now used instead of relative atomic mass to order the modern periodic table?

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. What does the atomic number of an element represent?
2. How is an element's atomic number related to the number of protons in its nucleus?
3. How does atomic number determine the position of an element in the periodic table?
4. Why does each element have a unique atomic number?
5. What is the atomic number of an element containing 12 protons?
6. How can the atomic number be used to identify an element?

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. How are elements arranged across the modern periodic table?
2. What are the horizontal rows of the periodic table called?
3. What are the vertical columns of the periodic table called?
4. Why are elements with similar properties placed in the same group?
5. What does an element's period tell you about its electron arrangement?

6. What does an element's group tell you about the elements with which it has 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. Where are metals generally found in the periodic table?
2. Where are non-metals generally found in the periodic table?
3. How does the atomic structure of metals differ from that of non-metals?
4. Why are elements on opposite sides of the periodic table generally different in their properties?
5. How can an element's position in the periodic table be used to predict whether it is a metal or non-metal?
6. How is the distinction between metals and non-metals related to their electron arrangements?

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. How are electrons arranged in shells around the nucleus of an atom?
2. What is the maximum number of electrons that can occupy the first electron shell?
3. What is the maximum number of electrons that can occupy the second electron shell for the first 20 elements?
4. What is the electronic configuration of an atom of sodium?
5. What is the electronic configuration of an atom of chlorine?
6. What is the electronic configuration of an atom of calcium?

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

1. How is the number of occupied electron shells related to an element's period?
2. How is the number of electrons in the outer shell related to an element's group?
3. Why do elements in the same group have similar chemical properties?
4. How can an element's electronic configuration be used to determine its position in the periodic table?
5. What does the electronic configuration 2.8.1 tell you about an element's position in the periodic table?
6. What does the electronic configuration 2.8.7 tell you about an element's group and period?

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. How is an ionic bond formed between two atoms?
2. Which type of atom loses electrons when forming an ionic bond?
3. Which type of atom gains electrons when forming an ionic bond?
4. What is a cation and how is it formed?
5. What is an anion and how is it formed?
6. How can a dot-and-cross diagram show the formation of an ionic bond?

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

1. What is an ion?
2. How does an atom become a positively charged ion?
3. How does an atom become a negatively charged ion?
4. What is the difference between a cation and an anion?
5. Why does an ion have an overall electrical charge?
6. Can an ion consist of more than one atom?

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

1. How can the number of protons in a simple ion be determined from its atomic number?
2. How can the number of neutrons in a simple ion be calculated from its mass number and atomic number?
3. How can the number of electrons in a positive ion be calculated from its atomic number and charge?
4. How can the number of electrons in a negative ion be calculated from its atomic number and charge?
5. How many protons, neutrons and electrons are in Na^+ with an atomic number of 11 and a mass number of 23?
6. How many protons, neutrons and electrons are in O^{2-} with an atomic number of 8 and a mass number of 16?

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. How do Group 1 atoms form ions?
2. How do Group 2 atoms form ions?
3. How do Group 6 atoms form ions?

4. How do Group 7 atoms form ions?
5. Why do Group 1 and Group 2 atoms lose electrons when forming ions?
6. Why do Group 6 and Group 7 atoms gain electrons when forming ions?

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

1. What does the ending "-ide" generally indicate in the name of a compound?
2. What does the ending "-ate" generally indicate in the name of a compound?
3. What is the difference between an oxide and a nitrate in terms of their names?
4. Which ending is used for compounds containing a simple non-metal ion such as chloride?
5. Which ending is used in compounds containing ions such as nitrate, carbonate and sulfate?
6. How can the ending of an ionic compound's name help identify 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. How can the formula of an ionic compound be deduced from the charges of its constituent ions?
2. Why must an ionic compound have an overall charge of zero?
3. What is the formula of the ionic compound formed from Na^+ and Cl^-?
4. What is the formula of the ionic compound formed from Mg^{2+} and Cl^-?
5. What is the formula of the ionic compound formed from Ca^{2+} and OH^-?
6. What is the formula of the ionic compound formed from Al^{3+} and SO_4^{2-}?

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. What is the structure of an ionic compound?
2. What is meant by a lattice structure in an ionic compound?
3. How are the ions arranged within an ionic lattice?
4. What holds oppositely charged ions together in an ionic lattice?
5. Why does an ionic lattice contain both positive and negative ions?
6. Why are the electrostatic forces between ions in an ionic lattice strong?

Covalent bonding

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

1. How is a covalent bond formed?
2. What is shared between two atoms in a covalent bond?
3. Why do atoms form covalent bonds?
4. Which types of atoms commonly form covalent bonds?
5. How does sharing electrons help atoms achieve a more stable electron arrangement?
6. How can a dot-and-cross diagram represent a covalent bond?

1.29 Recall that covalent bonding results in the formation of molecules

1. What type of bonding results in the formation of molecules?
2. What is a molecule?
3. How are atoms held together within a covalent molecule?
4. Why are covalent substances able to form molecules?
5. What type of particles make up a simple molecular covalent substance?
6. How does covalent bonding differ from ionic bonding in terms of the particles formed?

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

1. What is the typical order of magnitude of the size of an atom?
2. What is the approximate size of an atom in metres?
3. What is meant by the order of magnitude of a particle's size?
4. How does the size of a small molecule compare with the size of an atom?
5. Why are atoms and small molecules too small to see with the naked eye?
6. How can standard form be used to represent the size of atoms and small molecules?

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. How is the covalent bond in a hydrogen molecule formed?
2. How is the covalent bond in a hydrogen chloride molecule formed?
3. How are the covalent bonds in a water molecule formed?
4. How are the covalent bonds in a methane molecule formed?
5. How is the double covalent bond in an oxygen molecule represented?
6. How are the covalent bonds in a carbon dioxide molecule represented?

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. What are the four main types of substance based on their structure and bonding?
2. How does the structure and bonding of ionic substances affect their melting and boiling points?
3. How does the structure and bonding of simple molecular covalent substances affect their melting and boiling points?
4. How does the structure and bonding of giant covalent substances affect their physical properties?
5. How does metallic bonding allow metals to conduct electricity as solids?
6. How do the structures and bonding of different types of substances affect their solubility in water and ability 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. Why do ionic compounds generally have high melting and boiling points?
2. What type of forces hold the ions together in an ionic compound?
3. Why do solid ionic compounds not conduct electricity?
4. Why do molten ionic compounds conduct electricity?
5. Why do ionic compounds conduct electricity when dissolved in water?
6. What happens to the ions in an ionic compound when the compound is 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. Why do simple molecular covalent substances generally have low melting and boiling points?
2. What forces have to be overcome when a simple molecular substance melts or boils?
3. Why are intermolecular forces weaker than the covalent bonds within molecules?
4. Why do simple molecular covalent substances generally not conduct electricity?
5. Why does having covalent bonds not allow simple molecular substances to conduct electricity?

6. Why does a simple molecular covalent substance remain a poor conductor when melted?

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

1. What are diamond and graphite?
2. Why are diamond and graphite described as different forms of carbon?
3. What type of substances are diamond and graphite?
4. What is meant by an allotrope of carbon?
5. Which two giant covalent allotropes of carbon are specified in the Edexcel GCSE Chemistry specification?
6. Why can diamond and graphite have different properties even though both contain only carbon atoms?

1.36 Describe the structures of graphite and diamond

1. What is the structure of diamond?
2. How many covalent bonds does each carbon atom form in diamond?
3. What is the structure of graphite?
4. How are carbon atoms arranged in the layers of graphite?
5. How many covalent bonds does each carbon atom form in graphite?
6. How are the layers in graphite held together?

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. Why can graphite conduct electricity?
2. Why is graphite suitable for use as an electrode?
3. Why can graphite be used as a lubricant?
4. Why can graphite layers slide over one another?
5. Why is diamond suitable for use in cutting tools?
6. Why is diamond harder than graphite?

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

1. What is a fullerene?
2. What is the structure of C60?
3. What is graphene?
4. Why can graphene conduct electricity?
5. Why is graphene strong despite being only one atom thick?
6. How does the structure of graphene differ from the structure of C60?

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

1. What is a polymer?
2. What is the structure of a simple polymer such as poly(ethene)?
3. What type of atoms form the main chain in poly(ethene)?
4. Why are polymers described as large molecules?
5. How are the carbon atoms arranged in the chain of poly(ethene)?
6. What monomer is used to produce poly(ethene)?

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

1. What is meant by the malleability of a metal?
2. Why are metals malleable?
3. Why can metals conduct electricity?
4. How does the structure of metallic substances allow electrons to move?
5. What type of bonding occurs in metals?
6. How does metallic bonding explain the physical properties 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. What is one limitation of a dot-and-cross diagram when representing a chemical structure?
2. What is one limitation of a ball-and-stick model when representing a chemical structure?
3. Why can a ball-and-stick model give a misleading impression of the space occupied by atoms?
4. What is one limitation of representing a three-dimensional chemical structure in two dimensions?
5. Why are models useful when studying chemical structures despite their limitations?
6. Why can no single model provide a completely accurate representation of a chemical structure?

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. What are the typical physical properties of most metals?
2. Why are most metals good conductors of electricity?
3. What are the typical physical properties of most non-metals?
4. How do the melting points of most metals compare with those of most non-metals?
5. How does the density of most metals compare with that of most non-metals?

6. How does the electrical conductivity of most metals compare with that of most non-metals?

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. How is the relative formula mass of a compound calculated from its relative atomic masses?
2. How is the percentage by mass of an element in a compound calculated?
3. What is the relative formula mass of H_2O?
4. What is the relative formula mass of $CaCO_3$?
5. How would you calculate the percentage by mass of oxygen in H_2O?
6. How would you calculate the percentage by mass of carbon in CO_2?

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

1. What is an empirical formula?
2. How can reacting masses be used to determine the empirical formula of a compound?
3. How can percentage composition be used to determine an empirical formula?
4. Why are the masses of elements converted into amounts in moles when determining an empirical formula?
5. How are mole ratios used to obtain an empirical formula?
6. What does an empirical formula tell you about the atoms in a compound?

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. How can the empirical formula of a compound be determined from its molecular formula?
2. How can the molecular formula of a compound be determined from its empirical formula?
3. What information is needed to determine a molecular formula from an empirical formula?
4. How is the multiplier between the empirical formula and molecular formula calculated?
5. What is the empirical formula of $C_6H_{12}O_6$?
6. How can the relative molecular mass be used to determine the molecular formula of a compound?

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

1. How can an experiment be used to determine the empirical formula of magnesium oxide?
2. Why is magnesium weighed before it is heated in an experiment to determine the formula of magnesium oxide?
3. Why is magnesium heated in oxygen when determining the empirical formula of magnesium oxide?
4. What measurements are needed to calculate the empirical formula of magnesium oxide?
5. Why must magnesium oxide be heated to constant mass?
6. How can the measured masses of magnesium and oxygen be converted into a ratio of moles?

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. What does the law of conservation of mass state?
2. Why is mass conserved during a chemical reaction in a closed system?
3. How does a precipitation reaction in a closed flask demonstrate conservation of mass?
4. Why can the measured mass of a non-enclosed system change during a chemical reaction?
5. Why can a reaction in an open flask appear to lose mass when a gas is produced?
6. Why can a reaction in an open flask appear to gain mass when a gas from the surroundings is taken in?

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

1. How can the mass of an unknown reactant or product be calculated from a balanced chemical equation?
2. Why must a chemical equation be balanced before using it to calculate reacting masses?
3. How are mole ratios obtained from a balanced chemical equation?
4. What information is needed to calculate the mass of an unknown substance from a balanced equation?
5. How are moles converted into mass when calculating reacting masses?
6. How can the balanced equation $2Mg + O_2 \rightarrow 2MgO$ be used to calculate the mass of magnesium oxide produced from a known mass of magnesium?

1.49 Calculate the concentration of solutions in g dm^{-3}

1. What is meant by the concentration of a solution in g dm^{-3}?
2. How is concentration in g dm^{-3} calculated?
3. What units are used for mass and volume when calculating concentration in g dm^{-3}?
4. How can the concentration of a solution be calculated from its mass of solute and volume?
5. How can the mass of solute be calculated when the concentration and volume are known?
6. How can the volume of solution be calculated when the concentration and mass of solute are known?

1.50 Recall that one mole of particles of a substance is defined as: a the Avogadro constant number of particles (6.02×10^{23} atoms, molecules, formulae or ions) of that substance b a mass of 'relative particle mass'

1. What is a mole?
2. How many particles are present in one mole of a substance?
3. What is the Avogadro constant?
4. Which types of particles can be counted using the Avogadro constant?
5. What does the mass of one mole of a substance correspond to?
6. How is the relative particle mass related to the mass of one mole of a substance?

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. How can the number of moles in a given mass of a substance be calculated?
2. How can the mass of a substance be calculated when the number of moles is known?
3. How can the number of particles in a given number of moles be calculated using the Avogadro constant?
4. How can the number of moles be calculated when the number of particles is known?
5. How can the number of particles in a given mass of a substance be calculated?
6. How can the mass of a substance be calculated when the number of particles is known?

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. What is meant by the limiting reactant in a chemical reaction?

2. Why does the limiting reactant control the maximum amount of product that can form?
3. What happens to the limiting reactant when a reaction goes to completion?
4. What happens to a reactant that is present in excess when the limiting reactant is completely used up?
5. How can the limiting reactant be identified from the amounts of reactants used in a reaction?
6. Why does increasing the amount of an excess reactant not increase the amount of product once the limiting reactant is used up?

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

1. Aluminium reacts with oxygen according to the equation $4Al + 3O_2 \rightarrow 2Al_2O_3$. What mass of aluminium oxide can be formed when 135 g of aluminium is completely reacted with oxygen?
2. Calcium reacts with oxygen according to the equation $2Ca + O_2 \rightarrow 2CaO$. What mass of calcium oxide is formed when 20 g of calcium reacts completely with oxygen?
3. Magnesium reacts with oxygen according to the equation $2Mg + O_2 \rightarrow 2MgO$. What mass of magnesium oxide is formed when 12 g of magnesium reacts completely with oxygen?
4. Carbon reacts with oxygen according to the equation $C + O_2 \rightarrow CO_2$. What mass of carbon dioxide is formed when 6 g of carbon reacts completely with oxygen?
5. Iron reacts with sulfur according to the equation $Fe + S \rightarrow FeS$. What mass of iron sulfide is formed when 28 g of iron reacts completely with sulfur?
6. A reaction produces 44 g of carbon dioxide from 12 g of carbon. Use these masses to determine the simplest mole ratio of carbon to carbon dioxide in the reaction.

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. How are particles arranged in a solid?
2. How do particles move in a solid?
3. How are particles arranged in a liquid?
4. How do particles move in a liquid?
5. How are particles arranged in a gas?
6. How does the energy of particles compare between solids, liquids and gases?

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. What is the name of the change from a solid to a liquid?
2. What is the name of the change from a liquid to a gas?
3. What is the name of the change from a gas to a liquid?
4. What is the name of the change from a liquid to a solid?
5. What is the name of the change from a solid directly to a gas?
6. What is the name of the change from a gas directly to a solid?

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

1. What happens to the arrangement and movement of particles when a solid melts?
2. What happens to the energy of particles when a liquid evaporates?
3. What happens to the arrangement and movement of particles when a gas condenses?
4. What happens to the energy of particles when a liquid freezes?
5. What happens to the particles when a solid sublimes?
6. Why do particles become more widely separated when a substance changes from a liquid to a gas?

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

1. How can melting point data be used to determine whether a substance is solid or liquid at a given temperature?
2. How can boiling point data be used to determine whether a substance is liquid or gas at a given temperature?
3. A substance has a melting point of 20°C and a boiling point of 80°C. What state is it in at 10°C?
4. A substance has a melting point of −10°C and a boiling point of 60°C. What state is it in at 25°C?
5. A substance has a melting point of 50°C and a boiling point of 120°C. What state is it in at 150°C?
6. A substance has a melting point of −5°C and a boiling point of 95°C. What state is it in at −20°C?

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. What does the term "pure" mean in chemistry?
2. How can the everyday meaning of "pure" differ from its meaning in chemistry?
3. What is the difference between a pure substance and a mixture?
4. Why does a pure substance contain only one substance?
5. Why can a mixture contain more than one element or compound?
6. Why can the components of a mixture retain their individual chemical 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. How does the melting point of a pure substance differ from that of a mixture?
2. What does a sharp melting point indicate about the purity of a substance?
3. What does melting over a range of temperatures indicate about a substance?
4. A sample melts sharply at 78°C. What does this suggest about the sample?
5. A sample begins melting at 70°C and finishes melting at 76°C. What does this suggest about the sample?
6. How can melting point data be used to distinguish between a pure substance and a mixture?

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. What type of mixture can be separated using simple distillation?
2. What type of mixture can be separated using fractional distillation?
3. What type of mixture can be separated using filtration?
4. What type of mixture can be separated using crystallisation?
5. What type of mixture can be separated using paper chromatography?
6. How does the type of mixture determine which separation technique should be used?

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

1. How would you choose an appropriate technique for separating a mixture?

2. How would you separate an insoluble solid from a liquid?
3. How would you separate a soluble solid from a solution to obtain the solid?
4. How would you separate two miscible liquids with different boiling points?
5. How would you obtain a pure solvent from a solution containing a dissolved solid?
6. Why must the physical properties of the components be considered when choosing a separation technique?

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. What is the purpose of the solvent in paper chromatography?
2. What is meant by the mobile phase in paper chromatography?
3. What is meant by the stationary phase in paper chromatography?
4. Why do different substances move different distances during paper chromatography?
5. Why must the substances being separated be soluble in the solvent?
6. How does paper chromatography separate the components of a 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. How can a paper chromatogram be used to determine whether a substance is pure or impure?
2. How can a paper chromatogram be used to identify an unknown substance by comparison with known substances?
3. What does an Rf value represent in paper chromatography?
4. How is an Rf value calculated from a paper chromatogram?
5. A substance travels 4 cm and the solvent front travels 8 cm. What is the Rf value?
6. Why can substances with the same Rf value under identical conditions be identified as the same substance?

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

1. How can simple distillation be used to investigate the composition of an ink?
2. How can paper chromatography be used to investigate the composition of an ink?
3. What information does the chromatogram provide about the dyes present in an ink?

4. How can known dye samples be used to identify dyes in an unknown ink?
5. Why is a pencil used to mark the starting line in paper chromatography rather than ink?
6. What variables should be controlled when comparing the chromatograms of different inks?

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. What processes are used to make waste water or ground water potable?
2. What is the purpose of sedimentation when producing potable water?
3. What is the purpose of filtration when producing potable water?
4. Why is chlorine added during the production of potable water?
5. How can sea water be made potable using distillation?
6. Why must water used in chemical analysis not contain dissolved salts?

You're right. For a practical point like 3.6, every question needs to contain enough context to stand completely alone — no questions such as "What variable should be measured?" without saying what investigation it refers to.

I'll apply that standard to every question, not just practicals.

Topic 3 – Chemical changes

Acids

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

1. Which ions are produced when an acid dissolves in water?
2. Which ions are produced when an alkali dissolves in water?
3. What is the chemical formula of a hydrogen ion?
4. What is the chemical formula of a hydroxide ion?
5. Which ions are responsible for the acidic properties of aqueous acids?
6. Which ions are responsible for the alkaline properties of aqueous alkalis?

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. What pH value does a neutral solution have?
2. What pH values indicate that a solution is acidic?

3. What pH values indicate that a solution is alkaline?
4. Which is more acidic: a solution with pH 2 or a solution with pH 5?
5. Which is more alkaline: a solution with pH 9 or a solution with pH 12?
6. What does a pH of 7 indicate about a solution?

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

1. What colour does blue litmus turn when placed in an acidic solution?
2. What colour does red litmus turn when placed in an alkaline solution?
3. What colour is methyl orange in an acidic solution?
4. What colour is methyl orange in an alkaline solution?
5. What colour is phenolphthalein in an acidic solution?
6. What colour is phenolphthalein in an alkaline solution?

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. How does increasing the concentration of hydrogen ions affect the pH of an acidic solution?
2. How does decreasing the concentration of hydrogen ions affect the pH of an acidic solution?
3. How does increasing the concentration of hydroxide ions affect the pH of an alkaline solution?
4. How does decreasing the concentration of hydroxide ions affect the pH of an alkaline solution?
5. Which has the higher hydrogen ion concentration: an acidic solution with pH 2 or an acidic solution with pH 4?
6. Which has the higher hydroxide ion concentration: an alkaline solution with pH 10 or an alkaline solution with pH 12?

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. By how much does the pH decrease when the hydrogen ion concentration increases by a factor of 10?
2. By what factor does the hydrogen ion concentration increase when the pH decreases by 1?
3. By what factor does the hydrogen ion concentration increase when the pH decreases by 2?
4. An acidic solution changes from pH 5 to pH 4. By what factor has its hydrogen ion concentration increased?
5. An acidic solution changes from pH 6 to pH 3. By what factor has its hydrogen ion concentration increased?

6. An acidic solution has a pH of 4. What will its pH become if its hydrogen ion concentration increases by a factor of 100?

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. How could you investigate the change in pH when powdered calcium hydroxide is added to a fixed volume of dilute hydrochloric acid?
2. What should be measured after each addition of powdered calcium hydroxide to dilute hydrochloric acid when investigating changes in pH?
3. Why should the volume of dilute hydrochloric acid be kept constant when investigating the effect of adding powdered calcium hydroxide on pH?
4. Why should known masses or measured amounts of powdered calcium hydroxide be added to the dilute hydrochloric acid during the investigation?
5. What should be done to the hydrochloric acid mixture before measuring its pH after each addition of calcium hydroxide?
6. How would the pH of dilute hydrochloric acid change as increasing amounts of powdered calcium hydroxide are added?

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

1. What is meant by a dilute solution in terms of the amount of solute dissolved?
2. What is meant by a concentrated solution in terms of the amount of solute dissolved?
3. Which contains more solute per unit volume: a dilute solution or a concentrated solution?
4. How can a concentrated solution be made more dilute?
5. How can a dilute solution be made more concentrated?
6. What does the term concentration describe about the amount of solute in a solution?

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

1. What is meant by a strong acid in terms of its dissociation into ions?
2. What is meant by a weak acid in terms of its dissociation into ions?
3. What happens to the particles of a strong acid when it dissolves in water?
4. What happens to the particles of a weak acid when it dissolves in water?
5. What is meant by the dissociation of an acid into ions?
6. Why does a strong acid have a greater degree of dissociation into ions than a weak acid?

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

1. What is a base according to the Edexcel GCSE definition?
2. What products are formed when a base reacts with an acid?
3. What type of reaction occurs when an acid reacts with a base to form only a salt and water?
4. Is copper oxide classified as a base because it reacts with an acid to form a salt and water?
5. Is sodium hydroxide classified as a base because it reacts with an acid to form a salt and water?
6. What products must be formed when a substance reacts with an acid for that substance to meet the Edexcel definition of a base?

3.10 Recall that alkalis are soluble bases

1. What is an alkali?
2. What property of a base makes it an alkali?
3. Is sodium hydroxide an alkali?
4. Is potassium hydroxide an alkali?
5. Is copper oxide an alkali?
6. Why is copper oxide classified as a base but not as an alkali?

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. What are the general products when an aqueous acid reacts with a metal?
2. What are the general products when an aqueous acid reacts with a metal oxide?
3. What are the general products when an aqueous acid reacts with a metal hydroxide?
4. What are the general products when an aqueous acid reacts with a metal carbonate?
5. Which gas is generally produced when an aqueous acid reacts with a metal?
6. Which gas is generally produced when an aqueous acid reacts with a metal carbonate?

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

1. How can hydrogen gas be chemically tested for?
2. What positive result is observed when a lit splint is placed in a sample of hydrogen gas?
3. How can carbon dioxide gas be chemically tested for using limewater?

4. What positive result is observed when carbon dioxide is bubbled through limewater?
5. What gas produces a squeaky pop when a lit splint is applied to it?
6. What gas turns limewater cloudy when it is bubbled through limewater?

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

1. What is a neutralisation reaction?
2. What two types of substances react in a neutralisation reaction?
3. What products are formed when an acid reacts with a base in a neutralisation reaction?
4. Is the reaction between hydrochloric acid and sodium hydroxide a neutralisation reaction?
5. Is the reaction between nitric acid and copper oxide a neutralisation reaction?
6. Why is the reaction between an acid and a base described as neutralisation?

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. Which ions from an acid react with hydroxide ions during acid-alkali neutralisation?
2. Which ions from an alkali react with hydrogen ions during acid-alkali neutralisation?
3. What substance is formed when H^+ ions react with OH^- ions?
4. What is the ionic equation for acid-alkali neutralisation?
5. Why does the reaction between H^+ ions and OH^- ions result in neutralisation?
6. What happens to the H^+ ions and OH^- ions during acid-alkali neutralisation?

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. Why is an excess of an insoluble reactant added when preparing a soluble salt from an acid?
2. Why must excess insoluble reactant be removed when preparing a soluble salt from an acid?
3. How can excess insoluble reactant be removed from a salt solution?
4. Why is filtration suitable for removing excess insoluble reactant from a salt solution?
5. What substances should remain in the solution when an acid has completely reacted with an insoluble reactant?

6. Why does adding an excess of an insoluble reactant help ensure that all of the acid has reacted?

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. Why must titration be used when preparing a soluble salt from an acid and a soluble reactant?
2. Why cannot an excess soluble reactant be removed by filtration when preparing a soluble salt?
3. What is determined during a titration when preparing a soluble salt from an acid and a soluble reactant?
4. Why must the acid and soluble reactant be mixed in the correct proportions when preparing a pure soluble salt?
5. What substances should remain in the solution after an acid and soluble reactant have reacted in the correct proportions?
6. Why does using the correct proportions of an acid and soluble reactant prevent unreacted reactants remaining in the final solution?

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. Which acid reacts with copper oxide to prepare hydrated copper sulfate crystals?
2. Why is copper oxide added in excess when preparing hydrated copper sulfate crystals from copper oxide and sulfuric acid?
3. How is excess copper oxide removed when preparing hydrated copper sulfate crystals?
4. Why is a water bath used when preparing hydrated copper sulfate crystals?
5. How is water removed from the copper sulfate solution to allow hydrated copper sulfate crystals to form?
6. Why should the hydrated copper sulfate crystals not be heated strongly until completely dry?

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. What piece of apparatus is used to deliver a measured variable volume of solution during an acid-alkali titration?
2. What piece of apparatus is used to transfer an accurately measured fixed volume of solution during an acid-alkali titration?
3. Why is a suitable indicator used during an acid-alkali titration?

4. What volume should be recorded from the burette during an acid-alkali titration?
5. Why is an acid-alkali titration repeated until concordant titres are obtained?
6. How can an acid-alkali titration be used to prepare a pure, dry soluble salt?

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. Which common salts of sodium, potassium and ammonium are soluble in water?
2. Are all nitrates soluble in water?
3. Which common chlorides are insoluble in water?
4. Which common sulfates are insoluble in water?
5. Which common carbonates are insoluble in water, apart from the stated exceptions?
6. Which common hydroxides are insoluble in water, apart from the stated exceptions?

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. What is a precipitate in a chemical reaction between two aqueous solutions?
2. Will a precipitate form when aqueous silver nitrate is mixed with aqueous sodium chloride?
3. What is the name of the precipitate formed when aqueous silver nitrate is mixed with aqueous sodium chloride?
4. Will a precipitate form when aqueous barium nitrate is mixed with aqueous sodium sulfate?
5. What is the name of the precipitate formed when aqueous barium nitrate is mixed with aqueous sodium sulfate?
6. How can solubility rules be used to predict whether mixing two named aqueous solutions will produce a precipitate?

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

1. How can an insoluble salt be prepared from two soluble solutions?
2. Why are two soluble solutions mixed when preparing an insoluble salt?

3. What happens when the ions of an insoluble salt combine in aqueous solution?
4. How is an insoluble salt separated from the aqueous solution in which it forms?
5. Why is an insoluble salt washed with distilled water after filtration?
6. How is a pure, dry sample of an insoluble salt obtained after filtration and washing?

Electrolytic processes

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

1. What is an electrolyte?
2. What type of substances can act as electrolytes when molten or dissolved in water?
3. Why can an ionic compound conduct electricity when it is molten?
4. Why can an ionic compound conduct electricity when dissolved in water?
5. Can a solid ionic compound act as an electrolyte?
6. What must happen to an ionic compound for its ions to be able to move and conduct electricity?

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

1. What is electrolysis?
2. What type of electrical supply is used during electrolysis?
3. What happens to an electrolyte during electrolysis?
4. What form of energy is supplied to an electrolyte during electrolysis?
5. What does the term decomposes mean in the context of electrolysis?
6. Why is electrolysis considered a chemical change rather than a physical change?

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. Which electrode do positively charged cations move towards during electrolysis?
2. Which electrode do negatively charged anions move towards during electrolysis?
3. Why do positively charged cations move towards the cathode during electrolysis?

4. Why do negatively charged anions move towards the anode during electrolysis?
5. What charge does the cathode have during electrolysis?
6. What charge does the anode have during electrolysis?

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. What products are formed at the electrodes when aqueous copper chloride is electrolysed using inert electrodes?
2. What products are formed at the electrodes when aqueous sodium chloride is electrolysed using inert electrodes?
3. What products are formed at the electrodes when aqueous sodium sulfate is electrolysed using inert electrodes?
4. What products are formed at the electrodes when water acidified with sulfuric acid is electrolysed using inert electrodes?
5. What products are formed at the electrodes when molten lead bromide is electrolysed?
6. Why are inert electrodes used when electrolysing aqueous copper chloride, sodium chloride, sodium sulfate, or acidified water?

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

1. What products are formed when a binary ionic compound is electrolysed in the molten state?
2. What product forms from the cation at the cathode when a molten binary ionic compound is electrolysed?
3. What product forms from the anion at the anode when a molten binary ionic compound is electrolysed?
4. What products would form when molten sodium chloride is electrolysed?
5. What products would form when molten magnesium bromide is electrolysed?
6. How can the ions present in a molten binary ionic compound be used to predict the products of electrolysis?

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

1. What does a half equation show during electrolysis?
2. What type of electron change is shown in a cathode half equation during electrolysis?

3. What type of electron change is shown in an anode half equation during electrolysis?
4. What is the cathode half equation for the formation of copper from Cu^{2+} ions?
5. What is the anode half equation for the formation of chlorine from Cl^- ions?
6. What is the cathode half equation for the formation of lead from Pb^{2+} ions?

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

1. What is oxidation in terms of electrons?
2. What is reduction in terms of electrons?
3. What happens to the number of electrons when a substance is oxidised?
4. What happens to the number of electrons when a substance is reduced?
5. Which process involves the loss of electrons: oxidation or reduction?
6. Which process involves the gain of electrons: oxidation or reduction?

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

1. At which electrode does reduction occur during electrolysis?
2. At which electrode does oxidation occur during electrolysis?
3. What electron process occurs at the cathode during electrolysis?
4. What electron process occurs at the anode during electrolysis?
5. Why is the cathode associated with reduction during electrolysis?
6. Why is the anode associated with oxidation during electrolysis?

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. What happens at the cathode when copper sulfate solution is electrolysed using copper electrodes?
2. What happens at the anode when copper sulfate solution is electrolysed using copper electrodes?
3. What happens to the mass of the cathode during the electrolysis of copper sulfate solution using copper electrodes?
4. What happens to the mass of the anode during the electrolysis of copper sulfate solution using copper electrodes?
5. Why does the concentration of copper ions in copper sulfate solution remain approximately constant when copper electrodes are used?
6. How can electrolysis of copper sulfate solution using copper electrodes be used to purify impure copper?

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

1. How could you investigate the electrolysis of copper sulfate solution using inert electrodes?
2. How could you investigate the electrolysis of copper sulfate solution using copper electrodes?
3. What observations could be made at the cathode when copper sulfate solution is electrolysed using inert electrodes?
4. What observations could be made at the anode when copper sulfate solution is electrolysed using inert electrodes?
5. How would the results differ between electrolysis of copper sulfate solution using inert electrodes and using copper electrodes?
6. What measurements could be made before and after electrolysis of copper sulfate solution to investigate changes at the electrodes?

Topic 4 – Extracting metals and equilibria

Obtaining and using metals

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

1. How can the reactions of metals with water be used to compare their relative reactivities?
2. How can the reactions of metals with dilute acids be used to establish a relative reactivity order?
3. How can a metal displacement reaction with a salt solution be used to compare the reactivity of two metals?
4. If magnesium reacts with copper sulfate solution but copper does not react with magnesium sulfate solution, which metal is more reactive?
5. If metal A displaces metal B from a solution of B's salt, what does this indicate about the relative reactivities of A and B?
6. Why can observations from reactions with water, acids and salt solutions be used to deduce the relative reactivity of metals?

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

1. Why is a metal displacement reaction classified as a redox reaction?
2. In a displacement reaction, what happens to the electrons lost by the more reactive metal?
3. What happens to the metal ions when they are displaced from a compound by a more reactive metal?

4. In the reaction $Mg + Cu^{2+} \rightarrow Mg^{2+} + Cu$, which substance is oxidised and which substance is reduced?
5. Why is the oxidation of one metal accompanied by the reduction of another species in a displacement reaction?
6. How can electron transfer be used to explain why magnesium displaces copper from copper sulfate?

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. What is the Edexcel GCSE Chemistry reactivity series from most reactive to least reactive?
2. What does a metal's position in the reactivity series indicate about its tendency to form positive ions?
3. Why do metals above hydrogen in the reactivity series generally react with dilute acids to produce hydrogen?
4. Which metals in the Edexcel reactivity series are less reactive than hydrogen?
5. Why is potassium more reactive than copper according to the reactivity series?
6. How can reactions with water and dilute acids provide evidence for the order of metals in the reactivity series?

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. Where are most metals found naturally before they are extracted?
2. What is an ore?
3. Why are most metals found as compounds rather than as uncombined elements in the Earth's crust?
4. Why can unreactive metals sometimes be found naturally as uncombined elements?
5. Which type of metals are most likely to be found in the Earth's crust as uncombined elements?
6. Why can gold occur naturally as an uncombined element?

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

1. What is oxidation in terms of oxygen?
2. What is reduction in terms of oxygen?
3. What happens to a substance when it is oxidised in an oxygen-based definition?

4. What happens to a substance when it is reduced in an oxygen-based definition?
5. In the reaction copper oxide + carbon → copper + carbon dioxide, which substance is reduced?
6. In the reaction copper + oxygen → copper oxide, which substance is oxidised?

4.6 — Recall that the extraction of metals involves reduction of ores

1. Why does extracting a metal from its ore generally involve reduction?
2. What happens to the metal compound in an ore during the reduction stage of metal extraction?
3. What does reduction mean when extracting a metal from a metal oxide?
4. Why must a metal ion in an ore gain electrons to become the metal?
5. What happens to the oxygen in a metal oxide when the metal oxide is reduced?
6. Why is the extraction of a metal from its oxide described as a reduction process?

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. Why can metals below carbon in the reactivity series be extracted by heating their ores with carbon?
2. Why cannot aluminium be extracted from its ore by heating it with carbon?
3. Why is electrolysis used to extract aluminium from its ore?
4. Why is extracting a very reactive metal by electrolysis generally more expensive than extracting a less reactive metal using carbon?
5. Why can iron be extracted from its ore by heating the ore with carbon?
6. How does a metal's position relative to carbon in the reactivity series determine whether carbon reduction or electrolysis is required for extraction?

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

1. What is bacterial extraction of metals?
2. How can bacteria help to extract metals from ores?
3. What is phytoextraction?
4. How can plants be used to extract metal compounds from contaminated soil?
5. What is one advantage of using biological methods to extract metals compared with traditional extraction methods?

6. What is one disadvantage of using biological methods of metal extraction?

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

1. How does a metal's position in the reactivity series affect its resistance to oxidation?
2. Why are less reactive metals generally more resistant to oxidation?
3. Which is more resistant to oxidation, gold or potassium, and why?
4. Why does potassium oxidise more readily than gold?
5. How does the tendency of a metal atom to form a cation relate to its tendency to be oxidised?
6. Why can the reactivity series be used to compare the relative resistance of metals to oxidation?

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. How does recycling metals help to conserve valuable raw materials?
2. How can recycling metals reduce damage to the environment?
3. Why can recycling metals reduce the energy needed compared with extracting metals from ores?
4. How can recycling metals have economic benefits?
5. Why can recycling metals reduce the amount of mining required?
6. What are the main environmental and economic advantages of recycling metals?

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. What is a life-cycle assessment?
2. Which four stages of a product's life are considered in a life-cycle assessment?
3. What environmental effects of obtaining raw materials are considered in a life-cycle assessment?
4. What environmental effects of manufacturing a product are considered in a life-cycle assessment?
5. Why does a life-cycle assessment consider how a product is used?
6. Why does a life-cycle assessment consider how a product is disposed of when it is no longer useful?

4.12 — Evaluate data from a life cycle assessment of a product

1. What information can be obtained from data in a life-cycle assessment?
2. How can life-cycle assessment data be used to compare the environmental impacts of two products?
3. Why should all stages of a product's life be considered when evaluating its environmental impact?
4. If one product has a lower environmental impact during manufacture but a higher impact during disposal, why must both stages be considered?
5. How could numerical data from a life-cycle assessment be used to decide which product has the lower overall environmental impact?
6. Why might life-cycle assessment data need to be evaluated rather than simply choosing the product with the lowest impact at one stage?

Reversible reactions and equilibria

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. What is meant by a reversible chemical reaction?
2. What symbol is used to represent a reversible chemical reaction?
3. What does the symbol ⇋ indicate in a chemical equation?
4. What happens to the products and reactants in a reversible reaction?
5. How can changing reaction conditions affect the direction of a reversible reaction?
6. What is the difference between a reversible reaction and a reaction that only proceeds in one direction?

4.14 — Explain what is meant by dynamic equilibrium

1. What is meant by dynamic equilibrium in a reversible reaction?
2. What happens to the forward reaction rate when a reversible reaction reaches dynamic equilibrium?
3. What happens to the backward reaction rate when a reversible reaction reaches dynamic equilibrium?
4. Why does the concentration of reactants and products remain constant at dynamic equilibrium?
5. Why does a reversible reaction continue to occur even after dynamic equilibrium has been reached?
6. What condition must the forward and backward reaction rates satisfy for dynamic equilibrium to exist?

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. What are the reactants used to manufacture ammonia in the Haber process?
2. Where is the nitrogen used in the Haber process obtained from?
3. Where is the hydrogen used in the Haber process obtained from?
4. What is the balanced equation for the reversible formation of ammonia from nitrogen and hydrogen?
5. Why can the reaction between nitrogen and hydrogen reach dynamic equilibrium?
6. What happens to the concentrations of nitrogen, hydrogen and ammonia when the Haber process reaches dynamic equilibrium?

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

1. What temperature is used in the Haber process?
2. What pressure is used in the Haber process?
3. Which catalyst is used in the Haber process?
4. What are the three industrial conditions used in the Haber process?
5. Why is an iron catalyst used in the Haber process?
6. What temperature, pressure and catalyst are required industrially for the Haber process?

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

1. For the equilibrium $N_2(g) + 3H_2(g) \leftrightharpoons 2NH_3(g)$, what happens to the position of equilibrium when the concentration of nitrogen is increased?
2. For the equilibrium $N_2(g) + 3H_2(g) \leftrightharpoons 2NH_3(g)$, what happens to the position of equilibrium when the concentration of ammonia is decreased?
3. For the equilibrium $N_2(g) + 3H_2(g) \leftrightharpoons 2NH_3(g)$, what happens to the position of equilibrium when the pressure is increased?
4. For the equilibrium $N_2(g) + 3H_2(g) \leftrightharpoons 2NH_3(g)$, what happens to the position of equilibrium when the pressure is decreased?
5. For the exothermic equilibrium $N_2(g) + 3H_2(g) \leftrightharpoons 2NH_3(g)$, what happens to the position of equilibrium when the temperature is increased?
6. For the exothermic equilibrium $N_2(g) + 3H_2(g) \leftrightharpoons 2NH_3(g)$, what happens to the position of equilibrium when the temperature is decreased?

Topic 5 – Separate chemistry 1

Transition metals, alloys and corrosion

5.1C — Recall that most metals are transition metals and that their typical properties include: a high melting point b high density c the formation of

coloured compounds d catalytic activity of the metals and their compounds as exemplified by iron

1. What are the four typical properties of transition metals required by Edexcel GCSE Chemistry?
2. Why do transition metals generally have high melting points?
3. Why do transition metals generally have high densities?
4. What property of many transition metal compounds makes them useful in applications involving colour?
5. What is catalytic activity, and why are transition metals useful as catalysts?
6. How does iron demonstrate the catalytic activity of transition metals and their compounds?

5.2C — Recall that the oxidation of metals results in corrosion

1. What is corrosion?
2. What happens to a metal when it is oxidised during corrosion?
3. Why does the oxidation of metals result in corrosion?
4. What is the relationship between oxidation and corrosion of metals?
5. Why does corrosion cause a metal object to deteriorate over time?
6. What chemical process is responsible for the corrosion of metals?

5.3C — Explain how rusting of iron can be prevented by: a exclusion of oxygen b exclusion of water c sacrificial protection

1. What two substances are required for iron to rust?
2. How does excluding oxygen prevent iron from rusting?
3. How does excluding water prevent iron from rusting?
4. How does painting an iron object help prevent rusting?
5. What is sacrificial protection?
6. Why can a more reactive metal protect iron from rusting through sacrificial protection?

5.4C — Explain how electroplating can be used to improve the appearance and/or the resistance to corrosion of metal objects

1. What is electroplating?
2. How does electroplating improve the appearance of a metal object?
3. How does electroplating protect a metal object from corrosion?
4. Why can coating an object with a less reactive metal increase its resistance to corrosion?
5. What happens to metal ions during the electroplating of an object?
6. Why is electroplating useful for objects that need both an attractive appearance and corrosion resistance?

5.5C — Explain, using models, why converting pure metals into alloys often increases the strength of the product

1. What is an alloy?
2. Why does adding different-sized atoms to a pure metal make the metal stronger?
3. Why can layers of atoms slide more easily in a pure metal than in an alloy?
4. How does the presence of different-sized atoms disrupt the regular arrangement of atoms in an alloy?
5. Why does the disrupted structure of an alloy make it harder for layers of atoms to slide?
6. How does alloying a metal usually affect its strength compared with the pure metal?

5.6C — Explain why iron is alloyed with other metals to produce alloy steels

1. Why is iron alloyed with other metals to produce alloy steels?
2. How does alloying change the properties of iron?
3. Why are alloy steels generally stronger or more useful than pure iron?
4. Why can different metals be added to iron to produce steels with different properties?
5. How does alloying iron affect the ability of layers of atoms to slide?
6. Why are alloy steels preferred to pure iron for many applications?

5.7C — Explain how the uses of metals are related to their properties (and vice versa), including aluminium, copper and gold and their alloys including magnalium and brass

1. Why is aluminium useful for aircraft despite being a metal?
2. Why is copper widely used for electrical wiring?
3. Why is gold useful for jewellery?
4. Why is magnalium useful when a strong but relatively lightweight material is required?
5. Why is brass useful for making objects such as musical instruments and fittings?
6. How are the uses of aluminium, copper and gold determined by their physical and chemical properties?

Quantitative analysis

5.8C — Calculate the concentration of solutions in mol dm–3 and convert concentration in g dm–3 into mol dm–3 and vice versa

1. What equation is used to calculate concentration in mol dm^{-3} from the number of moles and volume of solution?

2. What concentration in mol dm^{-3} is produced by dissolving 0.50 mol of solute in 2.0 dm^3 of solution?
3. How can a concentration in g dm^{-3} be converted into mol dm^{-3}?
4. What concentration in mol dm^{-3} is produced by a solution containing 5.85 g of sodium chloride (Mr = 58.5) in 1.00 dm^3?
5. How can a concentration in mol dm^{-3} be converted into g dm^{-3}?
6. What concentration in g dm^{-3} corresponds to a 0.250 mol dm^{-3} solution of sodium hydroxide (Mr = 40.0)?

5.9C — Core Practical: Carry out an accurate acid-alkali titration, using burette, pipette and a suitable indicator

1. What apparatus is used to measure a fixed volume of alkali accurately during an acid-alkali titration?
2. Why is a burette used to add the acid accurately during an acid-alkali titration?
3. Why is an indicator used during an acid-alkali titration?
4. What is the purpose of carrying out a rough titration before obtaining accurate titres?
5. Why should concordant titres be obtained in an accurate acid-alkali titration?
6. How is the titre calculated from the initial and final burette readings?

5.10C — Carry out simple calculations using the results of titrations to calculate an unknown concentration of a solution or an unknown volume of solution required

1. How can titration results be used to calculate the unknown concentration of an acid?
2. How can titration results be used to calculate the unknown concentration of an alkali?
3. What information from a titration is needed to calculate an unknown concentration?
4. If 25.0 cm^3 of 0.100 mol dm^{-3} HCl exactly neutralises 20.0 cm^3 of NaOH, what is the concentration of the NaOH?
5. How can a balanced chemical equation be used when calculating an unknown concentration from titration results?
6. How can titration results be used to calculate the volume of solution required for a reaction?

5.11C — Calculate the percentage yield of a reaction from the actual yield and the theoretical yield

1. What equation is used to calculate percentage yield?
2. What is meant by the theoretical yield of a reaction?

3. What is meant by the actual yield of a reaction?
4. What is the percentage yield when the theoretical yield is 50 g and the actual yield is 40 g?
5. What does a percentage yield of 100% mean?
6. Why is the actual yield of a reaction usually less than the theoretical yield?

5.12C — Describe that the actual yield of a reaction is usually less than the theoretical yield and that the causes of this include: a incomplete reactions b practical losses during the experiment c competing, unwanted reactions (side reactions)

1. Why is the actual yield of a reaction usually less than the theoretical yield?
2. How can an incomplete reaction reduce the actual yield of a product?
3. How can practical losses during an experiment reduce the actual yield?
4. How can side reactions reduce the actual yield of the desired product?
5. Why can transferring a product between containers cause a practical loss?
6. What are the three main causes of an actual yield being lower than the theoretical yield?

5.13C — Recall the atom economy of a reaction forming a desired product

1. What is atom economy?
2. What does a high atom economy indicate about a chemical reaction?
3. What does a low atom economy indicate about a chemical reaction?
4. Why is atom economy important when choosing a reaction pathway for manufacturing a product?
5. What happens to the atoms that are not incorporated into the desired product?
6. Why does a reaction producing fewer unwanted products generally have a higher atom economy?

5.14C — Calculate the atom economy of a reaction forming a desired product

1. What equation is used to calculate the percentage atom economy of a reaction?
2. What is the atom economy of a reaction with a desired product Mr of 80 and total Mr of reactants of 100?
3. What is the atom economy of the reaction $CaCO_3 \rightarrow CaO + CO_2$ if CaO is the desired product? (Ar: Ca = 40, C = 12, O = 16)
4. How can the relative formula masses of the reactants and desired product be used to calculate atom economy?
5. Why must the desired product be identified before calculating atom economy?
6. What does an atom economy of 75% mean?

5.15C — Explain why a particular reaction pathway is chosen to produce a specified product, given appropriate data such as atom economy, yield, rate, equilibrium position and usefulness of by-products

1. Why might a reaction with a high atom economy be preferred in industry?
2. Why might a reaction with a high percentage yield be preferred?
3. Why might the rate of a reaction affect the choice of industrial reaction pathway?
4. Why might the position of equilibrium affect the choice of reaction pathway?
5. Why can useful by-products make a reaction pathway more attractive?
6. How should atom economy, yield, rate, equilibrium position and useful by-products be considered when choosing an industrial reaction pathway?

5.16C — Describe the molar volume, of any gas at room temperature and pressure, as the volume occupied by one mole of molecules of any gas at room temperature and pressure (The molar volume will be provided as 24 dm3 or 24000 cm3 in calculations where it is required)

1. What is the molar volume of any gas at room temperature and pressure?
2. What volume does one mole of gas occupy at room temperature and pressure in dm^3?
3. What volume does one mole of gas occupy at room temperature and pressure in cm^3?
4. What volume would 2.5 mol of a gas occupy at room temperature and pressure?
5. How can the number of moles of a gas be calculated from its volume at room temperature and pressure?
6. Why can the molar volume be used for different gases at room temperature and pressure?

5.17C — Use the molar volume and balanced equations in calculations involving the masses of solids and volumes of gases

1. What volume of gas at room temperature and pressure is produced when 0.50 mol of gas is formed?
2. How can a balanced equation be used to calculate the volume of gas produced from a known mass of a solid?
3. What volume of CO_2 at room temperature and pressure is produced when 10.0 g of $CaCO_3$ decomposes according to $CaCO_3 \rightarrow CaO + CO_2$? (Mr $CaCO_3$ = 100)
4. What mass of calcium carbonate is required to produce 4.80 dm^3 of CO_2 at room temperature and pressure according to $CaCO_3 \rightarrow CaO + CO_2$?
5. Why must the equation be balanced before using it to calculate gas volumes?

6. How can the molar volume of a gas be combined with relative formula mass to calculate the mass of a solid involved in a reaction?

5.18C — Use Avogadro's law to calculate volumes of gases involved in a gaseous reaction, given the relevant equation

1. What does Avogadro's law state about the volumes of gases involved in a reaction at the same temperature and pressure?
2. According to $N_2 + 3H_2 \rightarrow 2NH_3$, what volume of hydrogen reacts with 10 dm^3 of nitrogen when all gases are measured at the same temperature and pressure?
3. According to $N_2 + 3H_2 \rightarrow 2NH_3$, what volume of ammonia is produced from 15 dm^3 of nitrogen when all gases are measured at the same temperature and pressure?
4. According to $2H_2 + O_2 \rightarrow 2H_2O$, what volume of oxygen reacts with 40 dm^3 of hydrogen when all gases are measured at the same temperature and pressure?
5. How can the coefficients in a balanced gaseous equation be used to calculate reacting gas volumes?
6. Why can gas volumes be compared directly using the mole ratios in a balanced equation when temperature and pressure are constant?

The Haber process and fertilisers

5.19C — Describe the Haber process as a reversible reaction between nitrogen and hydrogen to form ammonia

1. What reversible reaction occurs in the Haber process?
2. What are the two reactants used to produce ammonia in the Haber process?
3. What is the balanced symbol equation for the Haber process?
4. Where is the nitrogen used in the Haber process obtained from?
5. Where is the hydrogen used in the Haber process obtained from?
6. Why does the Haber process reach a dynamic equilibrium?

5.20C — Predict how the rate of attainment of equilibrium is affected by: a changes in temperature b changes in pressure c changes in concentration d use of a catalyst

1. How does increasing the temperature affect the rate at which equilibrium is reached in the Haber process?
2. How does increasing the pressure affect the rate at which equilibrium is reached in the Haber process?
3. How does increasing the concentration of a reactant affect the rate at which equilibrium is reached?

4. How does increasing the concentration of a product affect the rate at which equilibrium is reached?
5. How does adding a catalyst affect the rate at which equilibrium is reached?
6. Why does a catalyst allow equilibrium to be reached more quickly without changing the position of equilibrium?

5.21C — Explain how, in industrial reactions, including the Haber process, conditions used are related to: a the availability and cost of raw materials and energy supplies b the control of temperature, pressure and catalyst used produce an acceptable yield in an acceptable time

1. Why must the cost and availability of raw materials be considered when choosing industrial reaction conditions?
2. Why must the energy requirements of an industrial reaction be considered when selecting its operating conditions?
3. Why is a compromise temperature used in the Haber process rather than simply using the temperature that gives the highest equilibrium yield?
4. Why is a compromise pressure used in the Haber process rather than simply using an extremely high pressure?
5. Why is a catalyst used in the Haber process?
6. Why must industrial reaction conditions produce an acceptable yield in an acceptable time?

5.22C — Recall that fertilisers may contain nitrogen, phosphorus and potassium compounds to promote plant growth

1. Which three elements are commonly supplied by fertilisers to promote plant growth?
2. Why are nitrogen compounds included in fertilisers?
3. Why are phosphorus compounds included in fertilisers?
4. Why are potassium compounds included in fertilisers?
5. What are the three main types of nutrients supplied by fertilisers?
6. Why are fertilisers added to soil?

5.23C — Describe how ammonia reacts with nitric acid to produce a salt that is used as a fertiliser

1. What salt is produced when ammonia reacts with nitric acid?
2. What is the word equation for the reaction between ammonia and nitric acid?
3. What is the balanced symbol equation for the reaction between ammonia and nitric acid?
4. Why is ammonium nitrate used as a fertiliser?
5. What type of reaction occurs when ammonia reacts with nitric acid?

6. Which ions are present in ammonium nitrate?

5.24C — Describe and compare: a the laboratory preparation of ammonium sulfate from ammonia solution and dilute sulfuric acid on a small scale b the industrial production of ammonium sulfate, used as a fertiliser, in which several stages are required to produce ammonia and sulfuric acid from their raw materials and the production is carried out on a much larger scale (details of the industrial production of sulfuric acid are not required)

1. How is ammonium sulfate prepared in the laboratory from ammonia solution and dilute sulfuric acid?
2. What are the reactants used to prepare ammonium sulfate in the laboratory?
3. What is the balanced symbol equation for the reaction between ammonia and sulfuric acid?
4. How does the scale of laboratory production of ammonium sulfate differ from industrial production?
5. Why does industrial production of ammonium sulfate require several stages before the final salt is produced?
6. How do the laboratory and industrial methods of producing ammonium sulfate differ in scale and process?

Chemical cells and fuel cells

5.25C — Recall that a chemical cell produces a voltage until one of the reactants is used up

1. What does a chemical cell produce while it is operating?
2. Why does a chemical cell eventually stop producing a voltage?
3. What happens when one of the reactants in a chemical cell is used up?
4. What type of energy conversion occurs in a chemical cell?
5. What determines how long a chemical cell can continue producing a voltage?
6. Why cannot a chemical cell produce a voltage indefinitely?

5.26C — Recall that in a hydrogen–oxygen fuel cell hydrogen and oxygen are used to produce a voltage and water is the only product

1. Which two reactants are used in a hydrogen–oxygen fuel cell?
2. What product is formed in a hydrogen–oxygen fuel cell?
3. What is the overall word equation for a hydrogen–oxygen fuel cell?
4. What is the overall balanced symbol equation for a hydrogen–oxygen fuel cell?
5. Why does a hydrogen–oxygen fuel cell produce a voltage?
6. Why is water described as the only product of a hydrogen–oxygen fuel cell?

5.27C — Evaluate the strengths and weaknesses of fuel cells for given uses

1. What is one advantage of using a hydrogen–oxygen fuel cell compared with a fossil-fuel-powered system?
2. Why can hydrogen–oxygen fuel cells be considered environmentally beneficial when the hydrogen is produced using renewable energy?
3. What is one disadvantage of storing and transporting hydrogen for use in fuel cells?
4. Why can the production of hydrogen affect the environmental benefits of fuel cells?
5. Why are fuel cells useful for applications that require a continuous supply of electrical energy?
6. What factors should be considered when evaluating whether fuel cells are suitable for a particular use?

Topic 6 – Groups in the periodic table

Group 1

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. Why are the elements in group 1 classified as alkali metals?
2. Why are the elements in group 7 classified as halogens?
3. Why are the elements in group 0 classified as noble gases?
4. Which group of the periodic table contains the alkali metals?
5. Which group of the periodic table contains the halogens?
6. Which group of the periodic table contains the noble gases?

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

1. What is the physical texture of alkali metals?
2. How do the melting points of alkali metals compare with those of many other metals?
3. Why can alkali metals be described as soft metals?
4. Which two physical properties of alkali metals are characteristic of this group?
5. What happens to the melting point of an alkali metal compared with many transition metals?
6. Why can alkali metals be cut relatively easily?

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

1. What happens when lithium is added to water?
2. What happens when sodium is added to water?
3. What happens when potassium is added to water?
4. What products are formed when an alkali metal reacts with water?
5. What gas is produced when lithium, sodium or potassium reacts with water?
6. What is the general word equation for the reaction between an alkali metal and water?

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. How does the reactivity of lithium, sodium and potassium with water change down group 1?
2. Which is more reactive with water, lithium or sodium?
3. Which is more reactive with water, sodium or potassium?
4. How can the reactivity of an unfamiliar group 1 metal be predicted from its position in the group?
5. Which would be expected to react more vigorously with water, potassium or rubidium?
6. What happens to the vigour of the reaction with water as you move down group 1?

6.5 — Explain this pattern in reactivity in terms of electronic configurations

1. Why does the reactivity of group 1 metals increase down the group?
2. How does the distance of the outer electron from the nucleus change down group 1?
3. How does increased electron shielding affect the attraction between the nucleus and the outer electron in group 1 metals?
4. Why is the outer electron easier to remove from potassium than from lithium?
5. What happens to the first ionisation energy of group 1 metals down the group?
6. How does the electronic configuration of group 1 metals explain their tendency to form +1 ions?

Group 7

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

1. What colour and physical state is chlorine at room temperature?
2. What colour and physical state is bromine at room temperature?

3. What colour and physical state is iodine at room temperature?
4. Which of chlorine, bromine and iodine is a yellow-green gas at room temperature?
5. Which of chlorine, bromine and iodine is a red-brown liquid at room temperature?
6. Which of chlorine, bromine and iodine is a grey-black solid at room temperature?

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. How does the physical state of the halogens change down group 7?
2. How does the colour of the halogens change down group 7?
3. How does the melting point of the halogens change down group 7?
4. How does the boiling point of the halogens change down group 7?
5. What physical state would be predicted for a halogen below iodine in group 7 at room temperature?
6. How can the physical properties of an unfamiliar halogen be predicted from the trends in chlorine, bromine and iodine?

6.8 — Describe the chemical test for chlorine

1. How can chlorine gas be chemically tested?
2. What happens to damp blue litmus paper when it is exposed to chlorine?
3. What colour change occurs when chlorine is tested using damp blue litmus paper?
4. Why does chlorine eventually bleach damp blue litmus paper?
5. What observation confirms the presence of chlorine gas using damp blue litmus paper?
6. Why must the litmus paper be damp when testing for chlorine?

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. What type of compound is formed when a halogen reacts with a metal?
2. What compound is formed when chlorine reacts with sodium?
3. What compound is formed when bromine reacts with potassium?
4. What compound is formed when iodine reacts with aluminium?
5. What is the general word equation for the reaction between a metal and a halogen?
6. How can the reactions of chlorine, bromine and iodine with metals be used to predict the reactions of other halogens?

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. What type of compounds are formed when chlorine, bromine or iodine reacts with hydrogen?
2. What is formed when hydrogen chloride dissolves in water?
3. Why do hydrogen halides form acidic solutions when they dissolve in water?
4. What ions are produced when hydrogen chloride dissolves in water?
5. What acidic solution is formed when hydrogen bromide dissolves in water?
6. How can the reactions of chlorine, bromine and iodine with hydrogen be used to predict the reactions of other halogens?

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. How does the reactivity of chlorine, bromine and iodine change down group 7?
2. What happens when chlorine is added to potassium bromide solution?
3. What happens when chlorine is added to potassium iodide solution?
4. What happens when bromine is added to potassium iodide solution?
5. Why can chlorine displace bromine and iodine from their halide solutions?
6. Using the group 7 reactivity trend, which halogens would astatine be expected to be less reactive than?

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. Why is a halogen displacement reaction classified as a redox reaction?
2. What happens to a halide ion when it is oxidised during a displacement reaction?
3. What happens to a halogen molecule when it is reduced during a displacement reaction?
4. In the reaction $Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$, which species is oxidised?
5. In the reaction $Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$, which species is reduced?
6. In the reaction $Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$, which species gains electrons and which species loses electrons?

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

1. Why does the reactivity of halogens decrease down group 7?

2. Why does a halogen need to gain one electron to form a stable outer electron shell?
3. How does electron shielding change down group 7?
4. How does the distance between the nucleus and the outer shell change down group 7?
5. Why is chlorine more reactive than bromine?
6. Why does the ability of a halogen atom to attract an additional electron decrease down group 7?

Group 0

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

1. Why are noble gases chemically inert?
2. What is special about the outer electron shell of noble gases?
3. Why does a full outer electron shell make noble gases unreactive?
4. Why do noble gases have little tendency to gain electrons?
5. Why do noble gases have little tendency to lose electrons?
6. How does the electronic configuration of a noble gas explain its lack of chemical reactivity?

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

1. Why is helium used to fill balloons?
2. Why is helium suitable for use where a non-flammable gas is required?
3. Why is argon used inside filament lamps?
4. Why is argon suitable for use in welding?
5. How does the low density of helium make it useful in balloons and airships?
6. How does the inertness of noble gases determine their uses?

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. How does the boiling point of noble gases change down group 0?
2. How does the melting point of noble gases change down group 0?
3. How does the density of noble gases change down group 0?
4. How does the relative atomic mass of noble gases change down group 0?
5. What physical properties would be predicted for a noble gas below xenon in group 0?
6. How can the physical properties of an unfamiliar noble gas be predicted from the trends in group 0?

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. How can the rate of reaction between hydrochloric acid and marble chips be investigated by measuring the production of a gas?
2. What measurements should be taken during the reaction between hydrochloric acid and marble chips to determine the rate at which carbon dioxide is produced?
3. How can a graph of gas volume against time be used to determine the rate of the reaction between hydrochloric acid and marble chips?
4. How can the rate of reaction between sodium thiosulfate and hydrochloric acid be investigated by observing a colour change?
5. Why does the reaction between sodium thiosulfate and hydrochloric acid allow the rate of reaction to be measured by timing how long it takes for a cross beneath the reaction mixture to disappear?
6. How could the effect of changing the concentration of hydrochloric acid on the rate of reaction with marble chips be investigated experimentally?

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

1. What practical method could be used to determine the rate of a reaction that produces a gas?
2. What practical method could be used to determine the rate of a reaction that produces a precipitate?
3. How could the rate of a reaction involving a colour change be measured experimentally?
4. How could the rate of a reaction be determined by measuring the change in mass of the reaction mixture over time?
5. How could the rate of a reaction be determined from measurements of the concentration of a reactant over time?
6. What measurements are needed to calculate the rate of a chemical reaction from experimental data?

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. Why must reacting particles collide for a chemical reaction to occur?
2. What is meant by a successful collision between reacting particles?
3. How does increasing the frequency of collisions between reacting particles affect the rate of a chemical reaction?
4. How does increasing the energy of collisions between reacting particles affect the rate of a chemical reaction?
5. Why do not all collisions between reacting particles result in a reaction?
6. What two factors involving collisions can increase the rate of a chemical reaction?

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. Why does increasing the temperature generally increase the rate of a chemical reaction?
2. How does increasing the concentration of a reactant affect the frequency of collisions between reacting particles?
3. Why does increasing the surface area to volume ratio of a solid increase the rate of reaction?
4. Why does increasing the pressure of reacting gases increase the rate of reaction?
5. How does increasing temperature affect both the frequency and energy of collisions between reacting particles?
6. Explain why powdered calcium carbonate reacts faster with hydrochloric acid than the same mass of large calcium carbonate chips.

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

1. What does the gradient of a mass-against-time graph represent during a chemical reaction?
2. What does a steeper gradient on a volume-of-gas-against-time graph indicate about the rate of reaction?
3. How can the rate of a reaction be determined from a concentration-against-time graph?
4. What does a horizontal section on a graph of product volume against time indicate about a reaction?
5. How can two mass-against-time graphs be compared to determine which reaction is faster?
6. A reaction produces 60 cm^3 of gas in 30 seconds; what is the average rate of gas production?

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. What is a catalyst?
2. How does a catalyst affect the rate of a chemical reaction?
3. Does a catalyst change the products formed in a chemical reaction?
4. What happens to the chemical composition of a catalyst during a reaction?
5. What happens to the mass of a catalyst at the end of a reaction?
6. Why can a catalyst be described as being chemically unchanged at the end of a reaction?

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

1. What is activation energy?
2. How does a catalyst affect the activation energy of a chemical reaction?
3. Why does lowering the activation energy increase the rate of a chemical reaction?
4. How does a catalyst provide an alternative reaction pathway?
5. Why does a catalyst increase the number of successful collisions between reacting particles?
6. How would the activation energy shown on a reaction profile change when a catalyst is added?

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

1. What are enzymes?
2. Why are enzymes described as biological catalysts?
3. How do enzymes affect the rate of biochemical reactions?
4. How are enzymes used in the production of alcoholic drinks?
5. Which biological process involving enzymes produces ethanol during alcoholic fermentation?
6. Why are enzymes important in the production of alcoholic drinks?

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. What happens to heat energy when a salt dissolves in water?
2. Why can a temperature change be measured during an acid-alkali neutralisation reaction?
3. Why can a temperature change be measured during a displacement reaction carried out in solution?
4. Why can a temperature change be measured when a precipitate forms in a solution?
5. How can a temperature change be used as evidence for a heat energy change during a reaction in solution?
6. Which four types of chemical change specified in GCSE Chemistry can involve measurable heat energy changes?

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

1. What is an exothermic reaction?
2. What happens to the surroundings during an exothermic reaction?
3. Why does the temperature of the surroundings usually increase during an exothermic reaction?
4. What happens to heat energy during an exothermic change?
5. Is the energy change associated with an exothermic reaction positive or negative if energy released to the surroundings is represented by a negative energy change?
6. Give one example of an exothermic chemical reaction.

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

1. What is an endothermic reaction?
2. What happens to the surroundings during an endothermic reaction?
3. Why does the temperature of the surroundings usually decrease during an endothermic reaction?
4. What happens to heat energy during an endothermic change?
5. Is the energy change associated with an endothermic reaction positive or negative when energy absorbed by the reaction is represented as a positive energy change?
6. Give one example of an endothermic chemical reaction.

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

1. Why is energy required when a chemical bond is broken?
2. Is bond breaking an endothermic or exothermic process?
3. Is bond making an endothermic or exothermic process?
4. Why is energy released when a chemical bond is formed?

5. What happens to energy when bonds in reactants are broken?
6. What happens to energy when new bonds in products are formed?

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. When is an overall chemical reaction exothermic in terms of the energy required to break bonds and the energy released when bonds form?
2. When is an overall chemical reaction endothermic in terms of the energy required to break bonds and the energy released when bonds form?
3. A reaction requires 500 kJ mol^{-1} to break bonds and releases 650 kJ mol^{-1} when new bonds form; is the reaction exothermic or endothermic?
4. A reaction requires 750 kJ mol^{-1} to break bonds and releases 600 kJ mol^{-1} when new bonds form; is the reaction exothermic or endothermic?
5. Why is a reaction exothermic when more energy is released making bonds than is required to break bonds?
6. Why is a reaction endothermic when more energy is required to break bonds than is released when new bonds form?

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

1. How is the overall energy change of a reaction calculated using bond energies?
2. Calculate the energy change for a reaction in which 500 kJ mol^{-1} is required to break bonds and 700 kJ mol^{-1} is released when bonds form.
3. Calculate the energy change for a reaction in which 900 kJ mol^{-1} is required to break bonds and 650 kJ mol^{-1} is released when bonds form.
4. The bond energies required to break all bonds in the reactants total 1200 kJ mol^{-1}, while the bond energies released when bonds form in the products total 1500 kJ mol^{-1}. Calculate the overall energy change.
5. The bond energies required to break all bonds in the reactants total 1800 kJ mol^{-1}, while the bond energies released when bonds form in the products total 1400 kJ mol^{-1}. Calculate the overall energy change.
6. A reaction has an energy change of −250 kJ mol^{-1} and requires 1000 kJ mol^{-1} to break all bonds in the reactants. How much energy is released when the bonds in the products form?

7.15 Explain the term activation energy

1. What is activation energy?

2. Why must reacting particles have at least the activation energy for a successful collision to occur?
3. What happens if colliding particles have less energy than the activation energy?
4. Where is the activation energy shown on a reaction profile?
5. How does activation energy affect the rate of a chemical reaction?
6. Why does increasing the number of particles with energy greater than the activation energy increase the rate of reaction?

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

1. What features should be labelled on a reaction profile for an exothermic reaction?
2. What features should be labelled on a reaction profile for an endothermic reaction?
3. Where should the activation energy be shown on an exothermic reaction profile?
4. Where should the activation energy be shown on an endothermic reaction profile?
5. How should the relative energy levels of the reactants and products differ on an exothermic reaction profile?
6. How should the relative energy levels of the reactants and products differ on an endothermic reaction profile?

Topic 8 – Fuels and Earth science

Fuels

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

1. What is a hydrocarbon?
2. Which two elements are present in every hydrocarbon?
3. Is methane, CH_4, a hydrocarbon?
4. Is ethanol, C_2H_5OH, a hydrocarbon?
5. Is carbon dioxide, CO_2, a hydrocarbon?
6. How can the formula of a compound be used to determine whether it is a hydrocarbon?

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. What is crude oil?
2. Why is crude oil described as a complex mixture of hydrocarbons?
3. What types of structures can the carbon atoms in crude oil molecules form?
4. Why is crude oil an important source of useful substances?
5. How is crude oil used as a feedstock for the petrochemical industry?
6. Why is crude oil described as a finite resource?

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

1. What process is used to separate crude oil into simpler, more useful mixtures?
2. Why can the hydrocarbons in crude oil be separated by fractional distillation?
3. How does fractional distillation separate the different fractions of crude oil?
4. What happens to hydrocarbons with lower boiling points during fractional distillation?
5. What happens to hydrocarbons with higher boiling points during fractional distillation?
6. Why does fractional distillation produce fractions containing hydrocarbons with similar boiling points rather than individual pure hydrocarbons?

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. What is the main use of the gases fraction obtained from crude oil?
2. What is petrol used for?
3. What is kerosene used for?
4. What is diesel oil used for?
5. What is fuel oil used for?
6. What is bitumen used for?

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. How does the number of carbon atoms in hydrocarbon molecules generally change between different crude oil fractions?
2. How does boiling point generally change as the size of hydrocarbon molecules increases?
3. How does the ease of ignition generally change as hydrocarbon molecules become larger?
4. How does viscosity generally change as hydrocarbon molecules become larger?
5. Which homologous series are most hydrocarbons in crude oil fractions members of?
6. Why do different crude oil fractions have different physical properties?

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. What is a homologous series?
2. What do members of the same homologous series have in common in terms of their general formula?
3. How do the molecular formulae of neighbouring members of a homologous series differ?
4. How do physical properties change across a homologous series?
5. How do the chemical properties of members of the same homologous series compare?
6. What molecular group is added between neighbouring members of a homologous series?

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. What products are formed when a hydrocarbon undergoes complete combustion?
2. What type of energy change occurs during the complete combustion of a hydrocarbon?
3. What two substances must be present for the complete combustion of a hydrocarbon fuel?
4. Write the word equation for the complete combustion of a hydrocarbon.
5. What happens to the carbon atoms in a hydrocarbon during complete combustion?
6. What happens to the hydrogen atoms in a hydrocarbon during complete combustion?

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

1. Why can incomplete combustion of hydrocarbons occur?
2. What solid product can be formed when a hydrocarbon undergoes incomplete combustion?
3. What poisonous gas can be produced during incomplete combustion of hydrocarbons?
4. Why does incomplete combustion produce carbon monoxide instead of only carbon dioxide?
5. What happens to the amount of oxygen available when incomplete combustion occurs?
6. What are the two carbon-containing products that can be formed during incomplete combustion of hydrocarbons?

8.9 Explain how carbon monoxide behaves as a toxic gas

1. Why is carbon monoxide a toxic gas?
2. How does carbon monoxide affect the blood?
3. Which molecule in red blood cells does carbon monoxide bind to?
4. Why does carbon monoxide reduce the ability of the blood to transport oxygen?
5. Why can carbon monoxide poisoning prevent body cells from receiving enough oxygen?
6. Why is carbon monoxide particularly dangerous because it is difficult to detect without a detector?

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

1. What health problem can carbon monoxide produced by incomplete combustion cause?
2. Why is carbon monoxide from faulty fuel-burning appliances dangerous?
3. What is soot?
4. How can soot produced by incomplete combustion affect human health?
5. Why can incomplete combustion be particularly dangerous in poorly ventilated rooms?
6. What two harmful products of incomplete combustion are particularly important in fuel-burning appliances?

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

1. Why can burning some hydrocarbon fuels produce sulfur dioxide?
2. Where does the sulfur that produces sulfur dioxide come from?

3. What happens to sulfur-containing impurities when a hydrocarbon fuel is burned?
4. What pollutant is produced when sulfur impurities in fuels are oxidised during combustion?
5. Why do sulfur-containing impurities in fuels contribute to air pollution?
6. What must happen to sulfur in a fuel for sulfur dioxide to be produced during combustion?

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

1. How does sulfur dioxide contribute to acid rain?
2. What happens when sulfur dioxide dissolves in rainwater?
3. How can acid rain damage aquatic ecosystems?
4. How can acid rain damage buildings and statues made from limestone?
5. How can acid rain affect plants and forests?
6. Why is sulfur dioxide from burning fuels an environmental problem?

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. Why can nitrogen and oxygen react together inside an engine?
2. What conditions inside an engine allow nitrogen and oxygen to react?
3. What products are formed when nitrogen and oxygen react at high temperatures in engines?
4. Why are oxides of nitrogen considered pollutants?
5. Why does the high temperature inside an engine increase the likelihood of nitrogen and oxygen reacting?
6. Which two gases from the air react together at high temperatures in engines to form oxides of nitrogen?

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

1. What is one environmental advantage of using hydrogen rather than petrol as a fuel in cars?
2. Why does using hydrogen in a fuel cell avoid producing carbon dioxide directly at the point of use?
3. What is one disadvantage of using hydrogen as a fuel compared with petrol?
4. Why can the method used to produce hydrogen affect its overall environmental impact?
5. What storage problem is associated with using hydrogen as a fuel in cars?

6. What factors should be considered when evaluating hydrogen against petrol as a car fuel?

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. Why are petrol, kerosene and diesel oil classified as non-renewable fossil fuels?
2. What raw material is petrol obtained from?
3. What raw material is kerosene obtained from?
4. What raw material is diesel oil obtained from?
5. Where is methane found as a non-renewable fossil fuel?
6. Why are fossil fuels described as non-renewable?

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

1. What is cracking?
2. What type of hydrocarbon molecules are broken down during cracking?
3. What happens to large alkane molecules during cracking?
4. What types of smaller hydrocarbons can be produced by cracking?
5. Why are some of the products of cracking described as unsaturated hydrocarbons?
6. What is the difference between the saturated reactants and unsaturated products that can be formed during cracking?

8.17 Explain why cracking is necessary

1. Why is cracking necessary in the petroleum industry?
2. Why is there a greater demand for some smaller hydrocarbons than the larger hydrocarbons found in crude oil?
3. How does cracking help meet the demand for petrol?
4. Why are alkenes produced by cracking useful to the chemical industry?
5. How does cracking make crude oil more useful to society?
6. Why would simply separating crude oil into fractions not be enough to meet the demand for all petroleum products?

Earth and atmospheric science

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

1. What formed the gases that made up the Earth's early atmosphere?

2. How did volcanic activity contribute to the formation of the Earth's early atmosphere?
3. What happened to gases released by volcanic activity to form the early atmosphere?
4. Why was volcanic activity important in the formation of the Earth's early atmosphere?
5. Which major gases were released by volcanic activity that contributed to the Earth's early atmosphere?
6. How did the Earth's early atmosphere differ from the atmosphere present today?

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. What was the estimated oxygen content of the Earth's early atmosphere?
2. Which gas was present in large amounts in the Earth's early atmosphere?
3. Why was water vapour present in the Earth's early atmosphere?
4. What other gases were thought to be present in small amounts in the Earth's early atmosphere?
5. How does evidence from the composition of ancient rocks help scientists understand the Earth's early atmosphere?
6. What evidence can be used to interpret the likely composition of the Earth's early atmosphere?

8.20 Explain how condensation of water vapour formed oceans

1. How did water vapour in the early atmosphere lead to the formation of oceans?
2. What process changed water vapour into liquid water as the Earth cooled?
3. Why did cooling of the Earth cause water vapour to condense?
4. How did condensation contribute to the formation of the first oceans?
5. What happened to atmospheric water vapour as the Earth's surface became cooler?
6. Why could liquid water accumulate on the Earth's surface after condensation occurred?

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

1. How did the formation of the oceans decrease the amount of carbon dioxide in the atmosphere?
2. Why did atmospheric carbon dioxide dissolve into the oceans?
3. What happened to atmospheric carbon dioxide when liquid oceans formed?

4. How did the formation of oceans contribute to the reduction of atmospheric carbon dioxide?
5. Why was carbon dioxide able to dissolve in the oceans as they formed?
6. What happened to the concentration of carbon dioxide in the atmosphere as increasing amounts dissolved in the oceans?

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. How did primitive plants affect the amount of carbon dioxide in the atmosphere?
2. How did photosynthesis by primitive plants affect the amount of oxygen in the atmosphere?
3. Which gas did primitive plants remove from the atmosphere during photosynthesis?
4. Which gas did primitive plants release during photosynthesis?
5. Why did the amount of oxygen in the atmosphere gradually increase as primitive plants became more widespread?
6. How did photosynthesis by primitive plants contribute to the change from an oxygen-poor early atmosphere to an oxygen-rich atmosphere?

8.23 Describe the chemical test for oxygen

1. How is oxygen tested for in a laboratory?
2. What happens to a glowing splint when it is placed in oxygen?
3. What positive result indicates the presence of oxygen using a glowing splint?
4. Why does a glowing splint relight when placed in oxygen?
5. What apparatus or test is used to distinguish oxygen from gases that do not support combustion?
6. A gas relights a glowing splint; which gas is indicated by this test?

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. What is the greenhouse effect?
2. Which type of radiation is emitted by the Earth and absorbed by greenhouse gases?
3. Which atmospheric gases specified in the GCSE Chemistry specification contribute to the greenhouse effect?
4. How do carbon dioxide, methane and water vapour help keep the Earth warm?

5. What happens after greenhouse gases absorb infrared radiation emitted by the Earth?
6. Why would the Earth be colder without the natural greenhouse effect?

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. What correlation is observed between atmospheric carbon dioxide concentration and global temperature change?
2. How is the consumption of fossil fuels related to changes in atmospheric carbon dioxide concentration?
3. Why does a correlation between fossil fuel consumption, atmospheric carbon dioxide concentration and temperature provide evidence for human-caused climate change?
4. Why does correlation between atmospheric carbon dioxide concentration and temperature not, by itself, prove causation?
5. How can the locations where atmospheric carbon dioxide and temperature measurements are taken introduce uncertainty into climate data?
6. Why can the historical accuracy of measurements introduce uncertainty when evaluating evidence for human-caused climate change?

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. What are the main gases that make up today's atmosphere?
2. How can burning fossil fuels increase the concentration of carbon dioxide in the atmosphere?
3. How can livestock farming increase the concentration of methane in the atmosphere?
4. How can increased atmospheric concentrations of carbon dioxide and methane affect the climate?
5. What methods could be used to mitigate climate change caused by increased carbon dioxide and methane levels?
6. Why should the scale, risks and environmental implications of proposed climate-change mitigation methods be considered?

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6. Why should the scale, risks and environmental implications of proposed climate-change mitigation methods be considered?

Topic 9 – Separate chemistry 2

Qualitative analysis: tests for ions

9.1C Explain why the test for any ion must be unique

1. Why must a chemical test for an ion produce a unique result?
2. What is meant by a unique test for an ion?
3. Why would a test that gives the same positive result for several different ions be unsuitable for identifying a specific ion?
4. How does a unique test allow an unknown ion to be identified?
5. Why is it important to distinguish between different ions when analysing an unknown substance?
6. What would be the problem if two different ions produced identical results in the same chemical test?

9.2C Describe flame tests to identify the following ions in solids: a lithium ion, Li+ (red) b sodium ion, Na+ (yellow) c potassium ion, K+ (lilac) d calcium ion, Ca2+ (orange-red) e copper ion, Cu2+ (blue-green)

1. What flame-test colour is produced by lithium ions, Li^+?
2. What flame-test colour is produced by sodium ions, Na^+?
3. What flame-test colour is produced by potassium ions, K^+?
4. What flame-test colour is produced by calcium ions, Ca^{2+}?
5. What flame-test colour is produced by copper(II) ions, Cu^{2+}?
6. How can a flame test be used to identify a metal ion present in a solid compound?

9.3C Describe tests to identify the following ions in solids or solutions as appropriate: a aluminium ion, Al3+ b calcium ion, Ca2+ c copper ion, Cu2+ d iron(II) ion, Fe2+ e iron(III) ion, Fe3+ f ammonium ion, NH4 + using sodium hydroxide solution

1. How can sodium hydroxide solution be used to test for aluminium ions, Al^{3+}?
2. How can sodium hydroxide solution be used to distinguish calcium ions, Ca^{2+}, from other specified metal ions?
3. What observation is produced when sodium hydroxide solution is added to a solution containing copper(II) ions, Cu^{2+}?
4. What observation is produced when sodium hydroxide solution is added to a solution containing iron(II) ions, Fe^{2+}?

5. What observation is produced when sodium hydroxide solution is added to a solution containing iron(III) ions, Fe^{3+}?
6. How can sodium hydroxide solution be used to test for ammonium ions, NH_4^+?

9.4C Describe the chemical test for ammonia

1. How can you test a gas to determine whether it is ammonia?
2. What happens when ammonia gas comes into contact with damp red litmus paper?
3. What colour does damp red litmus paper turn when exposed to ammonia?
4. Why must litmus paper be damp when testing for ammonia?
5. What does the colour change of damp red litmus paper show when testing for ammonia?
6. What observation would confirm that an unknown gas is ammonia?

9.5C Describe tests to identify the following ions in solids or solutions as appropriate: a carbonate ion, CO3 2–, using dilute acid and identifying the carbon dioxide evolved b sulfate ion, SO4 2–, using dilute hydrochloric acid and barium chloride solution c chloride ion, Cl–, bromide ion, Br–, iodide ion, I–, using dilute nitric acid and silver nitrate solution

1. How can carbonate ions, CO_3^{2-}, be identified using dilute acid?
2. How can the gas produced from a carbonate ion test be confirmed as carbon dioxide?
3. How can sulfate ions, SO_4^{2-}, be tested for using dilute hydrochloric acid and barium chloride solution?
4. What observation indicates the presence of sulfate ions when barium chloride solution is added?
5. How can chloride, bromide and iodide ions be tested for using dilute nitric acid and silver nitrate solution?
6. How can the results of the silver nitrate test be used to distinguish between chloride, bromide and iodide ions?

9.6C Core Practical: Identify the ions in unknown salts, using the tests for the specified cations and anions in 9.2C, 9.3C, 9.4C, 9.5C

1. How can you systematically identify the cation and anion present in an unknown salt?
2. Which flame tests can be used to identify lithium, sodium, potassium, calcium and copper(II) ions in an unknown solid?
3. Which chemical tests can be used to identify aluminium, calcium, copper(II), iron(II), iron(III) and ammonium ions?
4. Which chemical tests can be used to identify carbonate, sulfate, chloride, bromide and iodide ions?

5. Why should separate portions of an unknown salt be used for different ion tests?
6. How can a combination of positive test results be used to identify an unknown salt?

9.7C Identify the ions in unknown salts, using results of the tests above

1. How can the results of chemical tests be used to identify the ions present in an unknown salt?
2. An unknown salt gives a lilac flame and produces a cream precipitate with silver nitrate. Which ions are present?
3. An unknown salt gives a blue-green flame and produces a white precipitate with barium chloride after acidification. Which ions are present?
4. An unknown salt produces a brick-red flame and a white precipitate with silver nitrate after adding dilute nitric acid. Which ions are present?
5. An unknown salt produces a red-brown precipitate with sodium hydroxide solution. Which metal ion is present?
6. Why should more than one test result be considered when identifying an unknown salt?

9.8C Describe that instrumental methods of analysis are available and that these may improve sensitivity, accuracy and speed of tests

1. What are instrumental methods of chemical analysis?
2. Why can instrumental methods be more sensitive than simple chemical tests?
3. How can instrumental methods improve the accuracy of chemical analysis?
4. How can instrumental methods make chemical analysis faster?
5. What is meant by the sensitivity of an analytical method?
6. Why might an instrumental method be preferred over a traditional chemical test?

9.9C Evaluate data from a flame photometer: a to determine the concentration of ions in dilute solution using a calibration curve b to identify metal ions by comparing the data with reference data (no knowledge of the instrument or how it works is required)

1. How can a calibration curve from a flame photometer be used to determine the concentration of an ion in a dilute solution?
2. What information is plotted on a calibration curve used with a flame photometer?
3. How can the concentration of an unknown solution be determined from a flame-photometer calibration curve?

4. How can flame-photometer data be used to identify an unknown metal ion?
5. How can the data for an unknown metal ion be compared with reference data to identify the ion?
6. Why is a calibration curve needed when determining an unknown ion concentration using flame-photometer data?

Hydrocarbons

9.10C Recall the formulae of molecules of the alkanes, methane, ethane, propane and butane, and draw the structures of these molecules, showing all covalent bonds

1. What is the molecular formula of methane?
2. What is the molecular formula of ethane?
3. What is the molecular formula of propane?
4. What is the molecular formula of butane?
5. What is the general formula for the alkane homologous series?
6. How should the structures of methane, ethane, propane and butane be represented to show all covalent bonds?

9.11C Explain why the alkanes are saturated hydrocarbons

1. What is meant by the term saturated hydrocarbon?
2. Why are alkanes described as saturated hydrocarbons?
3. What type of carbon-carbon bond is present in alkanes?
4. Why do alkanes contain the maximum possible number of hydrogen atoms?
5. What feature of an alkane molecule distinguishes it from an unsaturated hydrocarbon?
6. Why can alkanes not contain a carbon-carbon double bond?

9.12C Recall the formulae of molecules of the alkenes, ethene, propene, butene, and draw the structures of these molecules, showing all covalent bonds (but-1-ene and but-2-ene only)

1. What is the molecular formula of ethene?
2. What is the molecular formula of propene?
3. What is the molecular formula of butene?
4. What is the molecular formula of but-1-ene?
5. What is the molecular formula of but-2-ene?
6. How should the structures of ethene, propene, but-1-ene and but-2-ene be represented to show all covalent bonds?

9.13C Explain why the alkenes are unsaturated hydrocarbons, describing that their molecules contain the functional group C=C

1. What is meant by an unsaturated hydrocarbon?
2. Why are alkenes described as unsaturated hydrocarbons?
3. What functional group is present in all alkenes?
4. What type of bond is represented by C=C?
5. Why does the C=C bond make an alkene unsaturated?
6. How can an alkene be distinguished structurally from an alkane?

9.14C Recall the addition reaction of ethene with bromine, showing the structures of reactants and products, and extend this to other alkenes

1. What type of reaction occurs when ethene reacts with bromine?
2. What are the products when ethene reacts with bromine?
3. What happens to the C=C bond when ethene reacts with bromine?
4. What is the structural formula of the product formed when ethene reacts with bromine?
5. What is meant by an addition reaction of an alkene?
6. How can the reaction of ethene with bromine be extended to other alkenes?

9.15C Explain how bromine water is used to distinguish between alkanes and alkenes

1. How can bromine water be used to distinguish between an alkane and an alkene?
2. What happens to the colour of bromine water when it is added to an alkene?
3. What happens to the colour of bromine water when it is added to an alkane under normal test conditions?
4. Why does an alkene decolourise bromine water?
5. What does a positive bromine-water test indicate about a hydrocarbon?
6. Which structural feature of an alkene allows it to react with bromine water?

9.16C Describe how the complete combustion of alkanes and alkenes involves the oxidation of the hydrocarbons to produce carbon dioxide and water

1. What are the products of the complete combustion of an alkane?
2. What are the products of the complete combustion of an alkene?
3. Why is the complete combustion of hydrocarbons described as an oxidation reaction?
4. What happens to the carbon atoms in a hydrocarbon during complete combustion?

5. What happens to the hydrogen atoms in a hydrocarbon during complete combustion?
6. What two products are formed when a hydrocarbon undergoes complete combustion in excess oxygen?

Polymers

9.17C Recall that a polymer is a substance of high average relative molecular mass made up of small repeating units

1. What is a polymer?
2. What does it mean for a polymer to have a high average relative molecular mass?
3. What are the small repeating units that make up a polymer called?
4. Why is a polymer described as being made up of repeating units?
5. How does the relative molecular mass of a polymer compare with that of its repeating units?
6. What structural feature distinguishes a polymer from a small molecule?

9.18C Describe: a how ethene molecules can combine together in a polymerisation reaction b that the addition polymer formed is called poly(ethene) (conditions and mechanisms not required)

1. What type of reaction occurs when ethene molecules combine to form poly(ethene)?
2. What polymer is formed when ethene molecules undergo addition polymerisation?
3. What happens to the C=C bond in an ethene molecule during addition polymerisation?
4. What is the repeating unit in poly(ethene)?
5. What is the monomer used to produce poly(ethene)?
6. Why is poly(ethene) described as an addition polymer?

9.19C Describe how other addition polymers can be made by combining together other monomer molecules containing C=C, to include poly(propene), poly(chloroethene) (PVC) and poly(tetrafluoroethene) (PTFE) (conditions and mechanisms not required)

1. What type of monomer is required to make an addition polymer?
2. Which monomer is used to make poly(propene)?
3. Which monomer is used to make poly(chloroethene), PVC?
4. Which monomer is used to make poly(tetrafluoroethene), PTFE?
5. What happens to the C=C bond when a monomer containing C=C forms an addition polymer?

6. Why can propene, chloroethene and tetrafluoroethene form addition polymers?

9.20C Deduce the structure of a monomer from the structure of an addition polymer and vice versa

1. How can the monomer of an addition polymer be identified from the polymer's repeating unit?
2. What bond in an addition polymer indicates where the original C=C bond of the monomer has reacted?
3. How can the C=C bond be reconstructed when deducing a monomer from an addition polymer?
4. How can the repeating unit of an addition polymer be deduced from its monomer?
5. What structural change occurs when a C=C-containing monomer becomes part of an addition polymer?
6. How can you determine the monomer used to make a polymer from a displayed structure of the polymer?

9.21C Explain how the uses of polymers are related to their properties and vice versa: including poly(ethene), poly(propene), poly(chloroethene) (PVC) and poly(tetrafluoroethene) (PTFE)

1. How are the uses of polymers determined by their physical properties?
2. Why is poly(ethene) suitable for making plastic bags and containers?
3. Why is poly(propene) suitable for making items such as ropes and packaging?
4. Why is PVC suitable for making products such as pipes and cable insulation?
5. Why is PTFE suitable for non-stick coatings?
6. How can the use of a polymer give information about the properties it must have?

9.22C Explain: a why polyesters are condensation polymers b how a polyester is formed when a monomer molecule containing two carboxylic acid groups is reacted with a monomer molecule containing two alcohol groups c how a molecule of water is formed each time an ester link is formed

1. Why are polyesters classified as condensation polymers?
2. What two types of functional groups react to form a polyester?
3. What type of monomer contains two carboxylic acid groups in polyester formation?
4. What type of monomer contains two alcohol groups in polyester formation?

5. What is formed when a carboxylic acid group reacts with an alcohol group during polyester formation?
6. Why is a molecule of water formed each time an ester link is formed?

9.23C Describe some problems associated with polymers including the: a availability of starting materials b persistence in landfill sites, due to non-biodegradability c gases produced during disposal by combustion d requirement to sort polymers so that they can be melted and reformed into a new product

1. Why can the availability of starting materials be a problem when producing polymers?
2. Why do many polymers persist in landfill sites?
3. Why is the non-biodegradability of many polymers an environmental problem?
4. What problem can be caused by gases released when polymers are disposed of by combustion?
5. Why must polymers be sorted before they can be melted and reformed into new products?
6. Why can the disposal of polymers create environmental problems?

9.24C Evaluate the advantages and disadvantages of recycling polymers, including economic implications, availability of starting materials and environmental impact

1. What are the environmental advantages of recycling polymers?
2. How can recycling polymers reduce the need for new starting materials?
3. How can recycling polymers affect the economic cost of producing plastic products?
4. What are the disadvantages or limitations of recycling polymers?
5. Why does sorting polymers affect the practicality of recycling?
6. How should economic, environmental and resource factors be considered when evaluating polymer recycling?

9.25C Recall that: a DNA is a polymer made from four different monomers called nucleotides (names of nucleotides not required) b starch is a polymer based on sugars c proteins are polymers based on amino acids

1. What type of polymer is DNA?
2. What are the monomers that make up DNA called?
3. How many different types of nucleotide monomer are used to make DNA?
4. What type of molecules are the repeating units in starch based on?
5. What are proteins polymers of?
6. Which monomers form the polymer chains of DNA, starch and proteins respectively?

Alcohols and carboxylic acids

9.26C Recall the formulae of molecules of the alcohols, methanol, ethanol, propanol (propan-1-ol only) and butanol (butan-1-ol only), and draw the structures of these molecules, showing all covalent bonds

1. What is the molecular formula of methanol?
2. What is the molecular formula of ethanol?
3. What is the molecular formula of propan-1-ol?
4. What is the molecular formula of butan-1-ol?
5. What is the general molecular formula of the alcohols methanol, ethanol, propan-1-ol and butan-1-ol?
6. How should the structures of methanol, ethanol, propan-1-ol and butan-1-ol be represented to show all covalent bonds?

9.27C Recall that the functional group in alcohols is –OH and that alcohols can be dehydrated to form alkenes

1. What is the functional group present in alcohols?
2. What is the name of the functional group –OH in an alcohol?
3. What happens when an alcohol is dehydrated?
4. What type of hydrocarbon is formed when an alcohol is dehydrated?
5. What structural feature is introduced when an alcohol is dehydrated to form an alkene?
6. Why does dehydration of an alcohol produce an unsaturated hydrocarbon?

9.28C Core Practical: Investigate the temperature rise produced in a known mass of water by the combustion of the alcohols ethanol, propanol, butanol and pentanol

1. How can the temperature rise produced by burning an alcohol be investigated experimentally?
2. Why should a known mass of water be used when comparing the temperature rises produced by different alcohols?
3. What measurements are needed to determine the temperature rise of water when an alcohol is burned?
4. How can the temperature rise produced by ethanol, propanol, butanol and pentanol be compared fairly?
5. What variable should be kept constant when comparing the energy released by different alcohols using this experiment?
6. Why should the alcohol burner be positioned at a fixed distance from the water container during this investigation?

9.29C Recall the formulae of molecules of the carboxylic acids, methanoic, ethanoic, propanoic and butanoic acids, and draw the structures of these molecules, showing all covalent bonds

1. What is the molecular formula of methanoic acid?
2. What is the molecular formula of ethanoic acid?
3. What is the molecular formula of propanoic acid?
4. What is the molecular formula of butanoic acid?
5. What is the general molecular formula of the carboxylic acids methanoic, ethanoic, propanoic and butanoic acid?
6. How should the structures of methanoic, ethanoic, propanoic and butanoic acid be represented to show all covalent bonds?

9.30C Recall that the functional group in carboxylic acids is –COOH and that solutions of carboxylic acids have typical acidic properties

1. What is the functional group present in carboxylic acids?
2. What is the functional group –COOH called?
3. Why do solutions of carboxylic acids have acidic properties?
4. What type of ions are produced when carboxylic acids dissolve in water?
5. How can the acidic nature of a carboxylic acid solution be demonstrated using an indicator?
6. Which structural feature is responsible for the acidic properties of carboxylic acids?

9.31C Recall that ethanol can be oxidised to produce ethanoic acid and extend this to other alcohols (reagents not required)

1. What product is formed when ethanol is oxidised?
2. What type of compound is ethanoic acid?
3. What happens to the functional group of ethanol when ethanol is oxidised?
4. What product is formed when a suitable primary alcohol is oxidised?
5. What type of alcohol can be oxidised to form a carboxylic acid?
6. How can the oxidation of ethanol be extended to other suitable alcohols?

9.32C Recall members of a given homologous series have similar reactions because their molecules contain the same functional group and use this to predict the products of other members of these series

1. Why do members of the same homologous series have similar chemical reactions?
2. What structural feature do members of the same homologous series share?

3. How does the functional group determine the chemical reactions of members of a homologous series?
4. How can the reaction of one member of a homologous series be used to predict the reaction of another member?
5. What information about a homologous series is needed to predict the products of reactions of its members?
6. Why can members of the same homologous series have similar chemical properties despite having different molecular masses?

9.33C Describe the production of ethanol by fermentation of carbohydrates in aqueous solution, using yeast to provide enzymes

1. How is ethanol produced by fermentation?
2. What type of substances are fermented to produce ethanol?
3. What role does yeast play in the production of ethanol by fermentation?
4. Why must the carbohydrates be present in aqueous solution during fermentation?
5. What products are formed when carbohydrates undergo fermentation?
6. What biological catalyst in yeast carries out the reactions involved in fermentation?

9.34C Explain how to obtain a concentrated solution of ethanol by fractional distillation of the fermentation mixture

1. Why is fractional distillation used to obtain a concentrated solution of ethanol from a fermentation mixture?
2. What property of ethanol and water allows them to be separated by fractional distillation?
3. What happens to the ethanol-water mixture during fractional distillation?
4. Why can fractional distillation produce a more concentrated ethanol solution than the original fermentation mixture?
5. Why can fractional distillation not normally produce completely pure ethanol from an ethanol-water mixture?
6. How does the boiling point of ethanol compare with that of water during fractional distillation?

Bulk and surface properties of matter including nanoparticles

9.35C Compare the size of nanoparticles with the sizes of atoms and molecules

1. What is the typical size range of nanoparticles?
2. How does the size of a nanoparticle compare with the size of an atom?
3. How does the size of a nanoparticle compare with the size of a small molecule?
4. Why are nanoparticles described as being on the nanoscale?

5. What unit is commonly used to describe the size of nanoparticles?
6. How does the size scale of nanoparticles differ from that of ordinary bulk materials?

9.36C Describe how the properties of nanoparticulate materials are related to their uses including surface area to volume ratio of the particles they contain, including sunscreens

1. Why do nanoparticles have a high surface area to volume ratio?
2. How can the high surface area to volume ratio of nanoparticles affect their properties?
3. Why are nanoparticles useful in some applications where larger particles of the same material would not be as effective?
4. Why are nanoparticles used in some sunscreens?
5. How does particle size affect the surface area to volume ratio of a nanoparticulate material?
6. How can the properties of nanoparticulate materials determine their uses?

9.37C Explain the possible risks associated with some nanoparticulate materials

1. Why can nanoparticulate materials pose potential health risks?
2. Why might nanoparticles behave differently from larger particles of the same substance?
3. How could nanoparticles enter the human body?
4. Why can the high surface area to volume ratio of nanoparticles contribute to potential risks?
5. Why is the long-term effect of some nanoparticles uncertain?
6. Why should the risks of nanoparticulate materials be considered when developing new applications?

9.38C Compare, using data, the physical properties of glass and clay ceramics, polymers, composites and metals

1. Which physical properties could be compared when evaluating glass, clay ceramics, polymers, composites and metals?
2. How can data be used to compare the physical properties of different materials?
3. Why might a metal have different physical properties from a polymer?
4. How can the physical properties of glass affect its suitability for particular uses?
5. How can the physical properties of clay ceramics affect their suitability for particular uses?
6. Why are composites often used when the properties of a single material are insufficient?

9.39C Explain why the properties of a material make it suitable for a given use and use data to select materials appropriate for specific uses

1. Why must the properties of a material be considered when selecting it for a particular use?
2. How can data be used to select the most suitable material for a specific application?
3. Which properties would be important when selecting a material for an electrical conductor?
4. Which properties would be important when selecting a material for a lightweight structural component?
5. Why might a material with a particular combination of properties be preferred over a material with one superior property?
6. How can experimental or supplied data be used to justify the choice of a material for a specific use?