Edexcel GCSE Comb Sci Chem

Topic 1 – Key concepts in chemistry

Atomic structure

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

1. How did Dalton's original model describe the structure of an atom?
2. What discovery showed that atoms are not indivisible solid spheres as Dalton proposed?
3. How did the discovery of the electron change the Dalton model of the atom?
4. What discovery led to the nuclear model of the atom?
5. How did the discovery of the nucleus change the model of the atom?
6. How has the modern model of the atom developed from Dalton's original model as a result of discoveries of subatomic particles?

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

1. What particles are found in the nucleus of an atom?
2. Where are electrons found in an atom?
3. What is meant by an electron shell?
4. What is the structure of the nucleus of an atom?
5. Where is most of the mass of an atom located?
6. Describe the overall structure of an atom in terms of its nucleus, protons, neutrons and electrons.

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

1. What are the relative charge and relative mass of a proton?
2. What are the relative charge and relative mass of a neutron?
3. What are the relative charge and relative mass of an electron?
4. Which subatomic particle has a relative charge of −1 and a relative mass of approximately 1/1840?
5. Which two subatomic particles have approximately the same relative mass?

6. What are the relative charges of a proton, neutron and 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 contained more protons than electrons?
5. What would happen to the overall charge of an atom if it contained more electrons than protons?
6. Explain why an atom with equal numbers of protons and electrons has no overall charge.

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. Where are the electrons located relative to the tiny nucleus?
4. Why can the nucleus be described as very small compared with the overall size of an atom?
5. What does the small size of the nucleus indicate about the amount of empty space within an atom?
6. How does the size of the nucleus compare with the distance over which the electrons are found in 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 do electrons contribute very little to the mass of an atom?
4. How does the mass of the nucleus compare with the mass of the electrons in an atom?
5. What relationship does the mass of an atom have with the number of protons and neutrons in its nucleus?
6. Why is the nucleus described as containing most of the mass of an atom?

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

1. What is meant by the mass number of an atom?
2. How is the mass number of an atom calculated from its protons and neutrons?

3. Which subatomic particles are included in the mass number?
4. Why are electrons not included when calculating the mass number?
5. An atom contains 11 protons and 12 neutrons. What is its mass number?
6. An atom has a mass number of 23 and contains 11 protons. How many neutrons does it contain?

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. Why do all atoms of the same element have the same number of protons?
3. What is meant by the atomic number of an element?
4. Why is the number of protons unique to each element?
5. An atom contains 17 protons. Which element does it represent?
6. Can two different elements have atoms containing the same number of protons? Explain your answer.

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 about the nuclei of isotopes of the same element?
4. What is different about the nuclei of isotopes of the same element?
5. How does the number of neutrons affect the mass number of an isotope?
6. Explain why carbon-12 and carbon-14 are isotopes of the same element.

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 neutrons in an atom be calculated from its mass number and atomic number?
3. How can the number of electrons in a neutral atom be determined from its atomic number?
4. An atom has atomic number 13 and mass number 27. How many protons, neutrons and electrons does it contain?
5. An atom has atomic number 17 and mass number 35. Calculate the numbers of protons, neutrons and electrons in the atom.
6. An atom has 19 protons and a mass number of 39. Calculate the numbers of neutrons and electrons in the atom.

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

1. Why can the relative atomic mass of an element be a non-whole number?
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 usually a weighted mean rather than the mass of one particular atom?
4. Why is the relative atomic mass of chlorine approximately 35.5 rather than a whole number?
5. How does the abundance of each isotope affect the relative atomic mass of an element?
6. Explain why an element consisting of two isotopes with different masses can have a relative atomic mass between the two isotope masses.

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 the masses and abundances of its isotopes?
2. An element has two isotopes with masses 10 and 11, present in abundances of 20% and 80%. Calculate its relative atomic mass.
3. An element has two isotopes with masses 63 and 65, present in abundances of 70% and 30%. Calculate its relative atomic mass.
4. An element has two isotopes with masses 24 and 26, present in abundances of 75% and 25%. Calculate its relative atomic mass.
5. An element has isotopes of relative masses 35 and 37 with abundances of 75% and 25%. Calculate its relative atomic mass.
6. An element has three isotopes with relative masses 20, 21 and 22 and abundances of 90%, 5% and 5%. Calculate 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 known at the time in his periodic table?
2. What properties did Mendeleev use to help arrange the elements in his periodic table?
3. Why did Mendeleev leave gaps in his periodic table?
4. How did Mendeleev use the properties of compounds to help identify patterns between elements?
5. Why was Mendeleev's arrangement of elements an improvement on earlier attempts to classify the elements?
6. How did Mendeleev's periodic table allow elements with similar properties to be grouped together?

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 to predict the existence of undiscovered elements?
2. What properties of an undiscovered element could Mendeleev predict from its position in his periodic table?
3. How could Mendeleev predict the relative atomic mass of an undiscovered element?
4. Why was the later discovery of elements with properties predicted by Mendeleev strong evidence supporting his periodic table?
5. How did the discovery of elements that matched Mendeleev's predictions demonstrate the usefulness of his periodic table?
6. Why did Mendeleev leave gaps rather than forcing all known elements into an incorrect pattern?

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 increasing relative atomic mass not always the same as increasing atomic number?
3. How can the relative abundance of isotopes cause one element to have a greater relative atomic mass than the next element in atomic-number order?
4. Why does the existence of isotopes explain some apparent anomalies in Mendeleev's ordering?
5. Why is atomic number now used instead of relative atomic mass to arrange elements in the modern periodic table?
6. How does the relative abundance of isotopes affect the relative atomic mass of an element and explain why some pairs of elements appear out of mass order?

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

1. What is meant by the atomic number of an element?
2. How is the atomic number related to the number of protons in the nucleus of an atom?
3. How does the atomic number determine the position of an element in the periodic table?
4. Why does every element have a unique atomic number?

5. An element has atomic number 12. How many protons are in the nucleus of each of its atoms?
6. An atom contains 16 protons. What is its atomic number and what does this tell you about its position in the periodic table?

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

1. In what order are elements arranged in the modern periodic table?
2. What is a period in the periodic table?
3. What is a group in the periodic table?
4. Why are elements with similar chemical properties placed in the same group?
5. How does moving from left to right across a period affect the atomic number of the elements?
6. What is the difference between a period and a group in the periodic table?

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. What feature of atomic structure generally distinguishes metals from non-metals?
4. Why do metals generally tend to lose electrons when they react?
5. Why do non-metals generally tend to gain or share electrons when they react?
6. How can the position of an element in the periodic table be used to predict whether it is a metal or a non-metal?

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

1. What is meant by the electronic configuration of an atom?
2. How many electrons can occupy the first electron shell?
3. How many electrons can occupy the second electron shell for the first 20 elements?
4. What is the electronic configuration of sodium?
5. What is the electronic configuration of chlorine?
6. What is the electronic configuration 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 the period number of an element?
2. How is the number of electrons in the outer shell related to the group number for the main groups?
3. What can the electronic configuration of an element tell you about its position in the periodic table?
4. An element has an electronic configuration of 2.8.1. What period and group is the element in?
5. An element has an electronic configuration of 2.8.7. What period and group is the element in?
6. An element is in Group 2 and Period 3. What electronic configuration would you predict for this element?

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. What happens to electrons when an ionic bond forms between a metal atom and a non-metal atom?
2. How does an atom become a cation during ionic bonding?
3. How does an atom become an anion during ionic bonding?
4. Why does electron transfer allow the atoms involved in ionic bonding to achieve more stable electronic configurations?
5. How can a dot and cross diagram show the transfer of electrons during ionic bonding?
6. Using a dot and cross diagram, how would you show the formation of sodium chloride from sodium and chlorine atoms?

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 positive ion?
3. How does an atom become a negative ion?
4. What is a cation?
5. What is an anion?
6. Can an ion consist of a group of atoms rather than a single 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 an ion be determined from its atomic number?
2. How can the number of neutrons in an 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. A sodium ion, Na^+, has atomic number 11 and mass number 23. Calculate its numbers of protons, neutrons and electrons.
6. A chloride ion, Cl^-, has atomic number 17 and mass number 35. Calculate its numbers of protons, neutrons and electrons.

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

1. What ion does a Group 1 metal atom form when it loses one electron?
2. What ion does a Group 2 metal atom form when it loses two electrons?
3. What ion does a Group 6 non-metal atom form when it gains two electrons?
4. What ion does a Group 7 non-metal atom form when it gains one electron?
5. Explain how magnesium and oxygen atoms form Mg^{2+} and O^{2-} ions.
6. Explain how calcium and chlorine atoms form Ca^{2+} and Cl^- ions in calcium chloride.

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 an ionic compound?
2. What does the ending –ate generally indicate in the name of an ionic compound?
3. What is the difference between a chloride ion and a chlorate ion in terms of the naming convention?
4. What does the name sodium chloride tell you about the ions present in the compound?
5. What does the name sodium sulfate tell you about the type of negative ion present?
6. How can the ending of the name of an ionic compound help you identify the type of anion 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 you deduce the formula of an ionic compound from the charges of its constituent ions?

2. What is the formula of the ionic compound formed from Na^+ and O^{2-} ions?
3. What is the formula of the ionic compound formed from Mg^{2+} and Cl^- ions?
4. What is the formula of the ionic compound formed from Al^{3+} and SO_4^{2-} ions?
5. What is the formula of the ionic compound formed from Ca^{2+} and NO_3^- ions?
6. What is the formula of the ionic compound formed from Al^{3+} and CO_3^{2-} ions?

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

Covalent bonding

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

1. What is a covalent bond?
2. How is a covalent bond formed between two atoms?
3. How many electrons are involved in a single covalent bond?
4. Why does sharing a pair of electrons allow atoms to achieve a more stable electron configuration?
5. What happens to the shared pair of electrons in a covalent bond?
6. Explain how two chlorine atoms form a covalent bond by sharing a pair of electrons.

1.29 Recall that covalent bonding results in the formation of molecules

1. What type of structure is formed when atoms are joined by covalent bonds in simple molecular substances?
2. What is a molecule?
3. How are the atoms within a molecule held together?
4. What type of bonding holds the atoms together within a molecule?
5. What is the difference between a molecule and an individual atom?

6. Give two examples of substances that exist as simple covalent molecules.

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 typical order of magnitude of the size of a small molecule?
3. Which is smaller, an atom or a small molecule?
4. Approximately how many metres is the diameter of a typical atom?
5. Why is it difficult to observe individual atoms using the naked eye?
6. How does the size of an atom compare with the size of an everyday object such as a football?

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 does a dot and cross diagram show the formation of a hydrogen molecule, H_2?
2. How does a dot and cross diagram show the formation of a hydrogen chloride molecule, HCl?
3. How does a dot and cross diagram show the formation of a water molecule, H_2O?
4. How does a dot and cross diagram show the formation of a methane molecule, CH_4?
5. How does a dot and cross diagram show the formation of an oxygen molecule, O_2?
6. How does a dot and cross diagram show the formation of a carbon dioxide molecule, CO_2?

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 classified according to their structure and bonding?
2. How does the structure of an ionic substance differ from that of a simple molecular covalent substance?
3. How does the structure of a giant covalent substance differ from that of a simple molecular covalent substance?

4. How does metallic bonding result in metals being able to conduct electricity as solids?
5. Why do substances with different structures and bonding have different melting points and boiling points?
6. How can the structure and bonding of a substance be used to predict its 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 points?
2. Why do ionic compounds generally have high boiling points?
3. Why do solid ionic compounds not conduct electricity?
4. Why can molten ionic compounds conduct electricity?
5. Why can ionic compounds dissolved in water conduct electricity?
6. Explain how the movement of ions determines whether an ionic compound can conduct electricity as a solid, molten liquid or aqueous solution.

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 points?
2. Why do simple molecular covalent substances generally have low boiling points?
3. What are intermolecular forces?
4. Why does melting a simple molecular substance not require breaking the covalent bonds within its molecules?
5. Why do simple molecular covalent substances generally conduct electricity poorly?
6. Explain how weak intermolecular forces account for the relatively low melting and boiling points of simple molecular covalent substances.

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

1. What element are graphite and diamond both made from?
2. What is meant by different forms of the same element?
3. Why are graphite and diamond classified as giant covalent substances?
4. What type of bonding holds the carbon atoms together in graphite?
5. What type of bonding holds the carbon atoms together in diamond?
6. Why can graphite and diamond have different properties despite both being made only from carbon?

1.36 Describe the structures of graphite and diamond

1. How are carbon atoms arranged in the structure of diamond?
2. How many covalent bonds does each carbon atom form in diamond?
3. How are carbon atoms arranged in the structure of graphite?
4. How many covalent bonds does each carbon atom form in a layer of graphite?
5. What is the structure of graphite made up of between its layers?
6. What structural difference between graphite and diamond explains their different physical properties?

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 can graphite be used as an electrode?
3. Why can graphite be used as a lubricant?
4. Why is diamond suitable for use in cutting tools?
5. How does the structure of graphite allow its layers to slide over one another?
6. Explain why the strong covalent bonding throughout diamond makes it suitable for cutting tools, whereas graphite is suitable as a lubricant and electrode.

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 C_{60}?
3. Why are some fullerenes useful as lubricants?
4. What is graphene?
5. Why can graphene conduct electricity?
6. Explain how the structure and bonding of graphene and fullerenes give them useful properties.

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 poly(ethene) molecule?
3. What type of atoms form the main chain in poly(ethene)?
4. Why is poly(ethene) described as having large molecules?
5. How are the carbon atoms arranged in the chain of a poly(ethene) molecule?
6. What is the relationship between ethene and 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 can metals be bent and shaped without breaking?
3. Why can metals conduct electricity?
4. What particles carry electrical charge through a metal?
5. How does metallic bonding allow the layers of metal atoms to move while keeping the structure together?
6. Explain how the structure and bonding of metals account for their malleability and ability to conduct electricity.

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 using a dot and cross diagram to represent bonding?
2. What is one limitation of using a ball and stick model to represent a molecule?
3. Why can a ball and stick model give a misleading impression of the relative sizes of atoms?
4. Why can a two-dimensional representation fail to show the true three-dimensional arrangement of atoms?
5. Why can a three-dimensional model still be an imperfect representation of a real chemical structure?
6. Why is it useful to use different representations and models when describing chemical structures?

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 is the typical appearance of most metals?
4. What are the typical physical properties of most non-metals?
5. How do the typical melting points and densities of metals compare with those of non-metals?
6. How do the electrical conductivities of most metals compare with those 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. Calculate the relative formula mass of $CaCO_3$ using Ar values Ca = 40, C = 12 and O = 16.
3. Calculate the relative formula mass of $Al_2(SO_4)_3$ using Ar values Al = 27, S = 32 and O = 16.
4. What calculation is used to determine the percentage by mass of an element in a compound?
5. Calculate the percentage by mass of oxygen in CO_2 using Ar values C = 12 and O = 16.
6. Calculate the percentage by mass of calcium in $CaCO_3$ using Ar values Ca = 40, C = 12 and O = 16.

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. A compound contains 24 g of carbon and 4 g of hydrogen. Calculate its empirical formula.
4. A compound contains 14 g of nitrogen and 16 g of oxygen. Calculate its empirical formula.
5. A compound contains 40.0% carbon, 6.7% hydrogen and 53.3% oxygen by mass. Calculate its empirical formula.
6. Why must the masses or percentage compositions of elements be converted into moles before determining an empirical formula?

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. What is the empirical formula of $C_6H_{12}O_6$?
3. What is the empirical formula of C_2H_4?
4. How can the molecular formula of a compound be determined from its empirical formula and relative molecular mass?
5. A compound has empirical formula CH_2O and relative molecular mass 180. Calculate its molecular formula.
6. A compound has empirical formula NO_2 and relative molecular mass 92. Calculate its molecular formula.

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

1. How could you experimentally determine the empirical formula of magnesium oxide?
2. Why is magnesium heated strongly in oxygen when determining the empirical formula of magnesium oxide?
3. Why must the crucible and lid be weighed before heating magnesium?
4. Why is the magnesium oxide heated, cooled and reweighed repeatedly until a constant mass is obtained?
5. How are the masses of magnesium and oxygen used to calculate the empirical formula of magnesium oxide?
6. What safety precaution should be taken when heating magnesium strongly during an experiment to determine the empirical formula of magnesium oxide?

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 does the total mass remain constant when a reaction takes place in a closed system?
3. A precipitation reaction takes place in a sealed flask. Why does the total mass remain unchanged?
4. Why can the measured mass change when a reaction takes place in an open flask and produces a gas?
5. Why can the measured mass increase when an open system takes in a gas from the surroundings?
6. Explain why mass is conserved in an open system even when the measured mass of the apparatus and contents changes.

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

1. What information from a balanced chemical equation is needed to calculate the mass of an unknown reactant or product?
2. In the reaction $2Mg + O_2 \rightarrow 2MgO$, what mass of magnesium oxide is produced when 24 g of magnesium reacts completely? Use Ar values Mg = 24 and O = 16.
3. In the reaction $2H_2 + O_2 \rightarrow 2H_2O$, what mass of water is produced when 4 g of hydrogen reacts completely? Use Ar values H = 1 and O = 16.
4. In the reaction $CaCO_3 \rightarrow CaO + CO_2$, what mass of carbon dioxide is produced when 100 g of calcium carbonate decomposes completely? Use Ar values Ca = 40, C = 12 and O = 16.

5. In the reaction $2Al + 3Cl_2 \rightarrow 2AlCl_3$, what mass of aluminium chloride is produced when 54 g of aluminium reacts completely? Use Ar values Al = 27 and Cl = 35.5.
6. In the reaction $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$, what mass of iron is produced from 160 g of iron(III) oxide? Use Ar values Fe = 56, O = 16 and C = 12.

1.49 Calculate the concentration of solutions in g dm–3

1. How is the concentration of a solution in g dm^{-3} calculated?
2. A solution contains 10 g of sodium chloride dissolved to make 2 dm^3 of solution. Calculate its concentration in g dm^{-3}.
3. A solution contains 15 g of potassium nitrate in 500 cm^3 of solution. Calculate its concentration in g dm^{-3}.
4. A solution has a concentration of 20 g dm^{-3}. What mass of solute is present in 0.5 dm^3 of the solution?
5. A solution contains 12 g of solute and has a concentration of 8 g dm^{-3}. Calculate the volume of the solution in dm^3.
6. A solution contains 25 g of solute in 250 cm^3 of solution. Calculate its concentration in g dm^{-3}.

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

1. What is meant by one mole of a substance?
2. What is the Avogadro constant?
3. How many particles are present in one mole of a substance?
4. What types of particles can be counted using the mole?
5. What is meant by the relative particle mass of a substance?
6. What mass in grams is equal to one mole of a substance in terms of its relative particle mass?

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

1. What equation is used to calculate the number of moles from the mass and relative particle mass of a substance?
2. Calculate the number of moles in 18 g of water, H_2O, given $Mr(H_2O) = 18$.
3. Calculate the mass of 0.50 mol of carbon dioxide, CO_2, given $Mr(CO_2) = 44$.
4. How can the number of particles in a given number of moles be calculated using the Avogadro constant?
5. Calculate the number of molecules in 0.25 mol of oxygen, O_2, using the Avogadro constant 6.02×10^{23} mol^{-1}.

6. Calculate the number of molecules in 9 g of water, H_2O, given Mr(H_2O) = 18 and the Avogadro constant = 6.02×10^{23} mol^{-1}.

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 a reactant being in excess?
2. What is meant by a limiting reactant?
3. Why does the limiting reactant determine the maximum amount of product that can form?
4. If 10 g of reactant A is in excess and 5 g of reactant B is completely used up, which reactant controls the amount of product formed?
5. Why does adding more of a reactant that is already in excess not increase the amount of product formed?
6. Explain why the mass of product formed in a reaction is controlled by the reactant that is not in excess.

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

1. What is meant by the stoichiometry of a chemical reaction?
2. How can the masses of reactants and products be used to deduce the stoichiometry of a reaction?
3. Aluminium reacts with oxygen according to $4Al + 3O_2 \rightarrow 2Al_2O_3$. Calculate the mass of aluminium oxide formed when 135 g of aluminium reacts completely with excess oxygen. Use Ar values Al = 27 and O = 16.
4. Magnesium reacts with oxygen according to $2Mg + O_2 \rightarrow 2MgO$. Calculate the mass of magnesium oxide formed when 48 g of magnesium reacts completely with excess oxygen. Use Ar values Mg = 24 and O = 16.
5. Calcium carbonate decomposes according to $CaCO_3 \rightarrow CaO + CO_2$. Calculate the mass of calcium oxide produced when 250 g of calcium carbonate decomposes completely. Use Ar values Ca = 40, C = 12 and O = 16.
6. Iron reacts with sulfur according to $Fe + S \rightarrow FeS$. A reaction uses 11.2 g of iron and produces 17.6 g of iron sulfide. Use these masses and the relative atomic masses Fe = 56 and S = 32 to deduce the stoichiometric ratio of iron to sulfur in the reaction.

Topic 2 – States of matter and mixtures

States of matter

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

1. How are particles arranged in a solid, liquid and gas?
2. How does the movement of particles differ between solids, liquids and gases?
3. How does the energy of particles compare between solids, liquids and gases?
4. Why do particles in a solid remain in fixed positions rather than moving freely?
5. Why can particles in a liquid move past one another while remaining close together?
6. Why can particles in a gas move freely and spread throughout a container?

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, and how do physical changes differ from chemical changes?

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

1. What happens to the arrangement of particles when a solid melts?
2. What happens to the movement and energy of particles when a liquid evaporates?
3. What happens to the arrangement, movement and energy of particles when a gas condenses?
4. What happens to the movement and energy of particles when a liquid freezes?
5. What happens to the particles when a solid sublimes directly into a gas?
6. What happens to the particles when a gas undergoes deposition to form a solid?

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

1. How can melting point and boiling point data be used to predict whether a substance is a solid, liquid or gas at a specified temperature?
2. A substance has a melting point of 20 °C and a boiling point of 80 °C. What state is it in at 10 °C?

3. A substance has a melting point of 20 °C and a boiling point of 80 °C. What state is it in at 50 °C?
4. A substance has a melting point of 20 °C and a boiling point of 80 °C. What state is it in at 100 °C?
5. What physical state would a substance be in at a temperature below its melting point?
6. What physical state would a substance be in at a temperature above its boiling point?

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 does the scientific meaning of pure differ from its everyday meaning?
3. What is the difference between a pure substance and a mixture?
4. Why is air classified as a mixture rather than a pure substance?
5. Why can a mixture contain more than one substance while a pure substance contains only one substance?
6. Why can the components of a mixture retain their individual 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 behaviour 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 substance melts sharply at 80 °C. What does this suggest about its purity?
5. A substance begins melting at 70 °C and finishes melting at 76 °C. What does this suggest about the substance?
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. What property of the components of a mixture determines which separation technique is appropriate?

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

1. Which separation technique would be appropriate for separating an insoluble solid from a liquid, and why?
2. Which separation technique would be appropriate for obtaining a solvent from a solution containing a dissolved solid, and why?
3. Which separation technique would be appropriate for separating two miscible liquids with different boiling points, and why?
4. Which separation technique would be appropriate for separating several soluble substances with different affinities for a stationary phase, and why?
5. How would you choose an appropriate technique to separate the components of an unknown mixture?
6. Why must the physical properties of the components be considered before 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 soluble substances move different distances during paper chromatography?
5. Why must the substances being separated be soluble in the solvent used for paper chromatography?
6. How does the movement of the solvent through the paper cause a mixture to separate?

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 a known substance?
3. What does a single spot on a chromatogram indicate about a substance?
4. What does more than one spot on a chromatogram indicate about a substance?
5. How is the Rf value of a substance calculated from a paper chromatogram?
6. How can an Rf value be used to identify an unknown substance?

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

1. How can simple distillation be used to separate the solvent from an ink sample?
2. How can paper chromatography be used to investigate the composition of an ink sample?
3. Why must the ink sample be placed near the bottom of the chromatography paper?
4. Why should the solvent level be below the ink spot when carrying out paper chromatography?
5. How can the chromatogram be used to determine whether an ink contains more than one dye?
6. How can the dyes in an unknown ink be identified using known substances and their chromatogram results?

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 and groundwater potable?
2. What is the purpose of sedimentation when treating waste water or groundwater?
3. What is the purpose of filtration when treating waste water or groundwater?
4. Why is chlorination used when making water potable?
5. How can sea water be made potable using distillation?
6. Why must water used in chemical analysis contain no dissolved salts?

Topic 3 – Chemical change

Acids

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

1. What ion do acids produce when they dissolve in water?
2. What ion do alkalis produce when they dissolve in water?
3. Which ion is responsible for the acidic properties of an aqueous acid?
4. Which ion is responsible for the alkaline properties of an aqueous alkali?
5. What is the formula of the hydrogen ion produced by acids in aqueous solution?
6. What is the formula of the hydroxide ion produced by alkalis in aqueous solution?

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 do acidic solutions have?
3. What pH values do alkaline solutions have?
4. Is a solution with a pH of 4 acidic, neutral or alkaline?
5. Is a solution with a pH of 11 acidic, neutral or alkaline?
6. Which is more acidic: a solution with pH 2 or a solution with pH 5?

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

1. What colour does blue litmus turn in an acidic solution?
2. What colour does red litmus turn 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 hydrogen ion concentration affect the pH of an acidic solution?
2. How does decreasing the hydrogen ion concentration affect the pH of an acidic solution?
3. How does increasing the hydroxide ion concentration affect the pH of an alkaline solution?
4. How does decreasing the hydroxide ion concentration affect the pH of an alkaline solution?

5. Which solution has the lower pH: one with a hydrogen ion concentration of 0.1 mol dm^{-3} or one with 0.001 mol dm^{-3}?
6. Which alkaline solution has the higher pH: one with a hydroxide ion concentration of 0.01 mol dm^{-3} or one with 0.001 mol dm^{-3}?

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

1. What happens to the pH when the hydrogen ion concentration increases by a factor of 10?
2. What happens to the pH when the hydrogen ion concentration decreases by a factor of 10?
3. A solution has a pH of 5. What is its pH if its hydrogen ion concentration increases by a factor of 10?
4. A solution has a pH of 3. What is its pH if its hydrogen ion concentration increases by a factor of 100?
5. A solution has a pH of 6. What is its pH if its hydrogen ion concentration decreases by a factor of 100?
6. By what factor has the hydrogen ion concentration increased if the pH decreases from 6 to 3?

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. What apparatus can be used to measure the pH of hydrochloric acid during the reaction with calcium hydroxide?
2. Why should the volume and concentration of hydrochloric acid be kept constant when investigating the effect of adding calcium hydroxide?
3. What should be measured after each addition of powdered calcium hydroxide to hydrochloric acid?
4. Why should powdered calcium hydroxide be added in small measured amounts during this investigation?
5. What happens to the pH of hydrochloric acid as calcium hydroxide is added?
6. What graph could be plotted to show how the pH changes as calcium hydroxide is 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?
2. What is meant by a concentrated solution?
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 does adding water affect the concentration of a solution?
6. Do the terms dilute and concentrated describe the strength of an acid?

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?
2. What is meant by a weak acid?
3. What happens to acid molecules in water when a strong acid dissociates?
4. What happens to acid molecules in water when a weak acid dissociates?
5. Which has a greater degree of dissociation into ions: a strong acid or a weak acid?
6. Can a dilute acid be strong, and can a concentrated acid be weak?

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?
2. What two types of products are formed when a base reacts with an acid?
3. What type of substance reacts with an acid to form only a salt and water?
4. Is copper oxide a base if it reacts with hydrochloric acid to form copper chloride and water?
5. Is sodium hydroxide a base if it reacts with hydrochloric acid to form sodium chloride and water?
6. What product other than water is formed when a base reacts with an acid?

3.10 Recall that alkalis are soluble bases

1. What is an alkali?
2. What property must a base have to be classified as an alkali?
3. Are all bases alkalis?
4. Why is sodium hydroxide classified as an alkali?
5. Why is copper oxide not classified as an alkali?
6. What happens when an alkali dissolves in water?

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

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

1. How is hydrogen tested for in the laboratory?
2. What observation confirms that a gas is hydrogen?
3. How is carbon dioxide tested for using limewater?
4. What observation confirms that a gas is carbon dioxide using limewater?
5. What sound is produced when a lighted splint is placed in hydrogen?
6. What happens to limewater when carbon dioxide is bubbled through it?

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

1. What is a neutralisation reaction?
2. Which two types of substances react in a neutralisation reaction?
3. What products are formed in a neutralisation reaction between an acid and a base?
4. Is the reaction between hydrochloric acid and sodium hydroxide a neutralisation reaction?
5. What type of reaction occurs when an acid reacts with an alkali to form salt and water?
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 react together during an acid-alkali neutralisation?
2. Where do the hydrogen ions in an acid-alkali neutralisation come from?
3. Where do the hydroxide ions in an acid-alkali neutralisation come from?
4. What product is formed when H^+ ions react with OH^- ions?
5. Write the ionic equation for an acid-alkali neutralisation.
6. Why does an acid-alkali neutralisation remove the acidic and alkaline properties of the solutions?

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. How is excess insoluble reactant removed after reacting it with an acid?
3. Why can excess insoluble reactant be removed by filtration?
4. What substances should remain in the solution after an insoluble reactant has completely reacted with an acid?

5. Why must the excess insoluble reactant be removed before crystallising the soluble salt?
6. Why is an insoluble reactant suitable for preparing a soluble salt without using titration?

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 is titration used when preparing a soluble salt from an acid and a soluble reactant?
2. Why cannot an excess soluble reactant simply be removed by filtration?
3. What does titration determine when preparing a soluble salt from an acid and a soluble reactant?
4. Why are the acid and soluble reactant mixed in the correct proportions after titration?
5. What substances should remain in the solution after the acid and soluble reactant have reacted in the correct proportions?
6. Why is titration necessary when both reactants used to prepare a soluble salt are soluble?

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 copper sulfate crystals?
3. How is excess copper oxide removed from the copper sulfate solution?
4. Why is a water bath used when preparing hydrated copper sulfate crystals?
5. How are hydrated copper sulfate crystals obtained from the copper sulfate solution?
6. Why must the copper sulfate crystals be dried after they have formed?

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 volume of acid accurately during a titration?
2. What piece of apparatus is used to measure a fixed volume of alkali accurately during a titration?
3. Why is a suitable indicator added during an acid-alkali titration?

4. What is done when the indicator shows that neutralisation has been reached?
5. Why is a second titration carried out without indicator when preparing a pure salt?
6. How is a pure, dry salt obtained from the neutralised solution after titration?

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. Are all common sodium, potassium and ammonium salts 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 soluble in water?
6. Which common hydroxides are soluble in water?

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. How can solubility rules be used to predict whether mixing two solutions will produce a precipitate?
2. What is a precipitate?
3. What happens when two solutions are mixed and an insoluble product forms?
4. When sodium chloride solution is mixed with silver nitrate solution, what precipitate is formed?
5. When barium chloride solution is mixed with sodium sulfate solution, what precipitate is formed?
6. When potassium nitrate solution is mixed with sodium chloride solution, why is no precipitate formed?

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

1. What type of reaction can be used to prepare an insoluble salt?
2. Why are two soluble solutions mixed when preparing an insoluble salt?
3. How is the insoluble salt separated from the reaction mixture?
4. Why is the insoluble salt washed with distilled water?
5. How is an insoluble salt dried after it has been filtered and washed?
6. Why must the insoluble salt be washed before it is dried?

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 compounds 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 it is dissolved in water?
5. Why does solid sodium chloride not conduct electricity?
6. What must be present and free to move for an ionic substance to 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 energy is supplied during electrolysis?
3. What type of electrical supply is used for electrolysis?
4. What happens to an electrolyte during electrolysis?
5. Why does electrolysis cause an electrolyte to decompose?
6. What is the role of a direct current supply in electrolysis?

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. During electrolysis, which electrode do positively charged cations migrate towards?
2. During electrolysis, which electrode do negatively charged anions migrate towards?
3. Why do cations migrate towards the cathode during electrolysis?
4. Why do anions migrate towards the anode during electrolysis?
5. What is the charge of a cation and which electrode attracts it during electrolysis?
6. What is the charge of an anion and which electrode attracts it 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 cathode and anode when aqueous copper chloride is electrolysed using inert electrodes?
2. What products are formed at the cathode and anode when aqueous sodium chloride is electrolysed using inert electrodes?
3. What products are formed at the cathode and anode when aqueous sodium sulfate is electrolysed using inert electrodes?
4. What products are formed at the cathode and anode when water acidified with sulfuric acid is electrolysed using inert electrodes?
5. What products are formed at the cathode and anode when molten lead bromide is electrolysed?
6. Why is hydrogen produced at the cathode instead of sodium during the electrolysis of aqueous sodium chloride?

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

1. What products are formed when a molten binary ionic compound is electrolysed?
2. How can the products of electrolysis of a molten metal halide be predicted?
3. What product forms from the metal cation at the cathode during electrolysis of a molten ionic compound?
4. What product forms from the non-metal anion at the anode during electrolysis of a molten ionic compound?
5. What products are formed when molten sodium chloride is electrolysed?
6. What products are formed when molten magnesium bromide is electrolysed?

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

1. What is a half equation?
2. How can a half equation show what happens to an ion during electrolysis?
3. What type of reaction occurs at the cathode during electrolysis, and how is it represented in a half equation?
4. What type of reaction occurs at the anode during electrolysis, and how is it represented in a half equation?
5. What half equation represents the reduction of Cu^{2+} ions to copper atoms at the cathode?
6. What half equation represents the oxidation of chloride ions to chlorine gas at the anode?

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 electrons when a substance is oxidised?
4. What happens to electrons when a substance is reduced?
5. How can the mnemonic OIL RIG be used to remember oxidation and reduction?
6. Is 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 happens to ions that are reduced at the cathode?
4. What happens to ions that are oxidised at the anode?
5. Which electrode is associated with electron gain during electrolysis?
6. Which electrode is associated with electron loss 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 to Cu^{2+} ions at the cathode when copper sulfate solution is electrolysed using copper electrodes?
2. What happens to copper atoms 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 sulfate solution remain approximately unchanged when copper electrodes are used?
6. How can electrolysis of copper sulfate solution using copper electrodes be used to purify copper?

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

1. How can the electrolysis of copper sulfate solution be investigated using inert electrodes?
2. What observations would be made at the electrodes when copper sulfate solution is electrolysed using inert electrodes?
3. How can the electrolysis of copper sulfate solution be investigated using copper electrodes?
4. What happens to the mass of each copper electrode during electrolysis of copper sulfate solution?

5. What measurements could be taken to compare the changes occurring at copper electrodes during electrolysis?
6. What safety precautions should be taken when investigating the electrolysis of copper sulfate solution?

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 relative reactivity of metals be deduced from their reactions with water?
2. How can the relative reactivity of metals be deduced from their reactions with dilute acids?
3. How can the relative reactivity of two metals be compared using a salt solution?
4. What does it indicate about the reactivity of a metal if it displaces another metal from its salt solution?
5. If magnesium displaces copper from copper sulfate solution, which metal is more reactive?
6. If copper does not react with dilute hydrochloric acid but magnesium does, what does this show about their relative reactivities?

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

1. What is a displacement reaction between a metal and a metal compound?
2. Why is a metal displacement reaction a redox reaction?
3. What happens to the atoms of the more reactive metal during a displacement reaction?
4. What happens to the ions of the less reactive metal during a displacement reaction?
5. In the reaction between zinc and copper sulfate, which substance is oxidised and which is reduced?
6. In a displacement reaction, how can electron transfer be used to identify oxidation and reduction?

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

1. What is the order of the metals in the reactivity series from most to least reactive?
2. How does the reactivity series show the relative tendency of metal atoms to form cations?
3. Why do metals higher in the reactivity series tend to form cations more readily?
4. How does the reaction of a metal with water provide evidence of its position in the reactivity series?
5. How does the reaction of a metal with dilute acid provide evidence of its position in the reactivity series?
6. Where are carbon and hydrogen placed in the reactivity series, and why are they useful reference points?

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 before they are extracted?
2. What is an ore?
3. Why are most metals found as compounds in ores rather than as uncombined elements?
4. Which type of metals can be found naturally in the Earth's crust as uncombined elements?
5. Why can unreactive metals such as gold be found as uncombined elements?
6. How does a metal's reactivity affect whether it is likely to be found as an ore or 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 gains oxygen?
4. What happens to a substance when it loses oxygen?
5. In the reaction $CuO + H_2 \rightarrow Cu + H_2O$, which substance is oxidised?
6. In the reaction $CuO + H_2 \rightarrow Cu + H_2O$, which substance is reduced?

4.6 Recall that the extraction of metals involves reduction of ores

1. Why does extracting a metal from its ore involve reduction?

2. What must happen to a metal compound in an ore for the metal to be extracted?
3. How is oxygen removed from a metal oxide during metal extraction?
4. Why is reduction of a metal compound necessary to obtain the metal itself?
5. What happens to the metal ions in a metal ore during extraction?
6. How does the reduction of a metal oxide produce the metal?

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

1. How does a metal's position in the reactivity series determine how it can be extracted from its ore?
2. Why can metals below carbon in the reactivity series be extracted by heating their ores with carbon?
3. Why can aluminium not be extracted from aluminium oxide by heating with carbon?
4. Why is electrolysis used to extract aluminium from its ore?
5. Why is the cost of extraction an important factor when choosing a method for extracting a metal?
6. Why is extracting a very reactive metal by electrolysis generally more expensive than extracting iron by heating its ore with carbon?

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

1. What is bacterial extraction of metals?
2. How can bacteria help extract metals from low-grade ores?
3. What is phytoextraction?
4. How can plants be used to extract metal compounds from contaminated soil?
5. What are the advantages of using biological methods to extract metals compared with traditional extraction methods?
6. What are the disadvantages of bacterial extraction and phytoextraction?

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

1. How is a metal's position in the reactivity series related to its resistance to oxidation?
2. Why are metals high in the reactivity series more easily oxidised?
3. Why are metals low in the reactivity series more resistant to oxidation?
4. Which is more resistant to oxidation, magnesium or copper, and why?

5. Why does gold have a high resistance to oxidation?
6. How does the tendency of a metal atom to form a cation relate to its tendency to be oxidised?

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

1. What are the environmental advantages of recycling metals?
2. How does recycling metals help conserve valuable raw materials?
3. How can recycling metals reduce the environmental damage caused by mining?
4. How can recycling metals reduce energy use compared with extracting metals from ores?
5. What are the economic advantages of recycling metals?
6. What factors could make recycling metals less economically or environmentally beneficial?

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. What environmental effects of obtaining raw materials are considered in a life-cycle assessment?
3. What environmental effects of manufacturing a product are considered in a life-cycle assessment?
4. What environmental effects of using a product are considered in a life-cycle assessment?
5. What environmental effects of disposing of a product are considered in a life-cycle assessment?
6. What four main stages of a product's life are considered in a life-cycle assessment?

4.12 Evaluate data from a life cycle assessment of a product

1. How can data from a life-cycle assessment be used to compare the environmental impacts of products?
2. What factors should be considered when evaluating life-cycle assessment data?
3. Why should the environmental impact of obtaining raw materials be considered when evaluating a product?
4. Why should the environmental impact of manufacturing, using and disposing of a product be considered together?

5. How can life-cycle assessment data be used to identify which product has the lower overall environmental impact?
6. Why might a product with a lower environmental impact at one stage of its life still have a greater overall environmental impact?

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 reaction?
3. What does the symbol ⇋ indicate in a chemical equation?
4. What happens in the forward reaction of a reversible reaction?
5. What happens in the reverse reaction of a reversible reaction?
6. How can changing reaction conditions alter the direction in which a reversible reaction proceeds?

4.14 Explain what is meant by dynamic equilibrium

1. What is meant by dynamic equilibrium?
2. What happens to the forward reaction at dynamic equilibrium?
3. What happens to the reverse reaction at dynamic equilibrium?
4. Why do the concentrations of reactants and products remain constant at dynamic equilibrium?
5. Why does a reversible reaction not stop when dynamic equilibrium is reached?
6. What must be true about the rates of the forward and reverse reactions at dynamic equilibrium?

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 two elements react to form 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 reaction between nitrogen and hydrogen to form ammonia?
5. Why can the reaction between nitrogen and hydrogen reach a dynamic equilibrium?
6. What happens to the concentrations of nitrogen, hydrogen and ammonia when dynamic equilibrium is reached?

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. What 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. Why is the Haber process carried out at 450 °C and 200 atmospheres rather than under extreme conditions?

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

1. For the Haber process, $N_2(g) + 3H_2(g) \leftrightharpoons 2NH_3(g)$, what happens to the position of equilibrium when the temperature is increased?
2. For the Haber process, $N_2(g) + 3H_2(g) \leftrightharpoons 2NH_3(g)$, what happens to the position of equilibrium when the temperature is decreased?
3. For the Haber process, $N_2(g) + 3H_2(g) \leftrightharpoons 2NH_3(g)$, what happens to the position of equilibrium when the pressure is increased?
4. For the Haber process, $N_2(g) + 3H_2(g) \leftrightharpoons 2NH_3(g)$, what happens to the position of equilibrium when the pressure is decreased?
5. For a reversible reaction at dynamic equilibrium, how does increasing the concentration of a reactant affect the position of equilibrium?
6. For a reversible reaction at dynamic equilibrium, how does decreasing the concentration of a product affect the position of equilibrium?

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. Which group of the periodic table contains the alkali metals?
2. Which group of the periodic table contains the halogens?
3. Which group of the periodic table contains the noble gases?
4. Why can elements in the same group of the periodic table be classified as a particular group of elements?
5. What does an element's position in the periodic table tell you about which group it belongs to?

6. Why are lithium, sodium and potassium classified as alkali metals?

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

1. What is the physical property of alkali metals that allows them to be cut relatively easily?
2. How do the melting points of alkali metals compare with those of many other metals?
3. Why are alkali metals described as soft metals?
4. What happens to the physical state of an alkali metal when it is heated above its relatively low melting point?
5. Which two physical properties are characteristic of alkali metals?
6. How do the softness and melting points of alkali metals distinguish them from many other metals?

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 observations would you make when sodium reacts with water?
6. What observations would you make when potassium reacts with 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 pattern of lithium, sodium and potassium be used to predict the reactivity of rubidium?
5. How can the reactivity pattern of Group 1 metals be used to predict the reactivity of caesium?
6. What happens to the speed and violence of the reaction with water as Group 1 metals become more reactive?

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. What happens to the outer electron of a Group 1 atom as the atomic number increases down the group?

3. Why is the outer electron easier to remove from potassium than from lithium?
4. How does increased atomic radius down Group 1 affect the attraction between the nucleus and the outer electron?
5. Why does increased electron shielding down Group 1 make the outer electron easier to remove?
6. How does the electronic configuration of Group 1 metals explain their increasing reactivity down the group?

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. What colour is chlorine at room temperature?
5. What colour is bromine at room temperature?
6. What colour is iodine 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 from chlorine to bromine to iodine?
2. How does the colour of the halogens change down Group 7?
3. What happens to the melting and boiling points of the halogens down Group 7?
4. How can the physical property pattern of chlorine, bromine and iodine be used to predict the physical state of astatine at room temperature?
5. How can the physical property pattern of chlorine, bromine and iodine be used to predict the colour of another halogen?
6. What general trend occurs in the physical properties of the halogens as you move down Group 7?

6.8 Describe the chemical test for chlorine

1. How can chlorine gas be chemically tested?
2. What happens when damp litmus paper is exposed to chlorine gas?
3. What colour change does damp blue litmus paper undergo when exposed to chlorine?
4. Why does chlorine eventually bleach damp litmus paper?

5. Why must the litmus paper be damp when testing for chlorine?
6. What observation confirms the presence of chlorine gas using damp litmus paper?

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 compounds are formed when halogens react with metals?
2. What happens when chlorine reacts with a metal?
3. What happens when bromine reacts with a metal?
4. What happens when iodine reacts with a metal?
5. What type of ion does a halogen form when it reacts with a metal?
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 compounds are formed when chlorine, bromine and iodine react with hydrogen?
2. What are hydrogen chloride, hydrogen bromide and hydrogen iodide collectively known as?
3. What happens when hydrogen chloride dissolves in water?
4. Why do hydrogen halides produce acidic solutions when they dissolve 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 reaction of another halogen?

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 compare?
2. What happens when chlorine is added to a solution containing bromide ions?
3. What happens when chlorine is added to a solution containing iodide ions?
4. What happens when bromine is added to a solution containing iodide ions?
5. Why can chlorine displace bromide ions and iodide ions from aqueous solutions?

6. Using the Group 7 reactivity pattern, how would astatine be expected to react with chloride, bromide and iodide ions?

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

1. Why is the displacement of a halide ion by a more reactive halogen a redox reaction?
2. What happens to the halide ions when they are displaced by a more reactive halogen?
3. What happens to the halogen molecules when they displace halide ions?
4. In the reaction $Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$, which substance is oxidised?
5. In the reaction $Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$, which substance is reduced?
6. In a halogen displacement reaction, how can electron gain and loss be used to identify oxidation and reduction?

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

1. Why does the reactivity of the halogens decrease down Group 7?
2. What happens to the atomic radius of halogen atoms as you move down Group 7?
3. How does electron shielding change down Group 7?
4. Why is a chlorine atom better able to attract an electron than an iodine atom?
5. How does the increasing distance between the nucleus and outer shell affect the attraction for an incoming electron?
6. How does electronic configuration explain why chlorine is more reactive than bromine and iodine?

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 gas atoms?
3. Why does a full outer electron shell make a noble gas chemically stable?
4. Why do noble gases generally not gain electrons?
5. Why do noble gases generally not lose or share electrons?
6. How does the electronic configuration of noble gases explain their 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 suitable for use in balloons?
2. Why is helium used instead of hydrogen in many balloons?
3. Why is argon used in some light bulbs?
4. Why is the inertness of argon useful when it is used in light bulbs?
5. How does the non-flammability of noble gases make them useful in particular applications?
6. How do the inertness, low density and non-flammability 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 do the boiling points of noble gases change down Group 0?
2. How does the density of noble gases change down Group 0?
3. How does the physical state of noble gases change with increasing atomic number at room temperature?
4. How can the physical property pattern of helium, neon and argon be used to predict the properties of krypton?
5. How can the physical property pattern of noble gases be used to predict the density of a heavier noble gas?
6. What general trend occurs in the physical properties of noble gases as you move down 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 could you investigate the effect of changing the concentration of hydrochloric acid on the rate of reaction with marble chips?
2. What measurement would you take when investigating the rate of reaction between hydrochloric acid and marble chips by measuring gas production?

3. How could the rate of the reaction between sodium thiosulfate and hydrochloric acid be investigated using a colour change?
4. What causes the visible colour change when sodium thiosulfate reacts with hydrochloric acid?
5. How could you use experimental results to determine which of two conditions produces the faster reaction?
6. What variables should be controlled when investigating the effect of one condition on the rate of a reaction?

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

1. How could you measure the rate of a reaction that produces a gas?
2. How could you measure the rate of a reaction that produces a precipitate?
3. How could you measure the rate of a reaction that causes a colour change?
4. How could you measure the rate of a reaction that causes a change in mass?
5. How could you determine the rate of a reaction from a graph of a measured quantity against time?
6. What factors should be considered when choosing a practical method for measuring the rate of a reaction?

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?
3. Why do some collisions between reacting particles not result in a reaction?
4. How does increasing the frequency of collisions affect the rate of reaction?
5. How does increasing the energy of collisions affect the rate of reaction?
6. How does collision theory explain why a reaction becomes faster when there are more frequent successful collisions?

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. Explain why increasing the temperature increases the rate of a reaction in terms of particle collisions.
2. Explain why increasing the concentration of a solution increases the rate of a reaction.
3. Explain why increasing the surface area to volume ratio of a solid increases the rate of reaction.

4. Explain why increasing the pressure of reacting gases increases the rate of reaction.
5. How does increasing temperature affect both the frequency and energy of collisions between particles?
6. Why does increasing concentration, surface area to volume ratio or gas pressure increase the frequency of collisions?

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

1. What does the gradient of a graph of product volume against time represent?
2. What does a steep gradient on a reaction graph indicate about the rate of reaction?
3. What does a horizontal section of a graph of reactant concentration against time indicate?
4. A reaction produces 60 cm^3 of gas in 30 seconds. Calculate the mean rate of gas production.
5. A reaction produces 90 cm^3 of gas in 45 seconds. Calculate the mean rate of gas production and give the unit.
6. Two reactions produce the same final volume of gas, but one graph has a steeper initial gradient. Which reaction has the greater initial rate, and why?

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 alter the products formed in a reaction?
4. What happens to a catalyst chemically during a reaction?
5. What happens to the mass of a catalyst at the end of a reaction?
6. Why can a catalyst be used repeatedly in principle without being used up?

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

1. What is meant by the activation energy of a reaction?
2. How does a catalyst affect the activation energy of a reaction?
3. Why does lowering the activation energy increase the rate of a reaction?
4. How does a catalyst provide an alternative reaction pathway?
5. Why does lowering activation energy result in more successful collisions?
6. How would the activation energy shown on a reaction profile differ for a catalysed and uncatalysed reaction?

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 reactions in living organisms?
4. Why are enzymes involved in the production of alcoholic drinks?
5. What role do enzymes from yeast play during the production of alcoholic drinks?
6. Why are enzymes useful catalysts in biological and industrial processes?

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 some salts dissolve in water?
2. Why can a temperature change be measured during a neutralisation reaction?
3. Why can a temperature change be measured during a displacement reaction?
4. Why can a temperature change be measured during a precipitation reaction?
5. How can a temperature change be used to determine whether a reaction in solution releases or absorbs heat energy?
6. What four types of chemical change in solution can be investigated by measuring temperature 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 heat energy during an exothermic reaction?
3. What happens to the temperature of the surroundings during an exothermic reaction?
4. Why does an exothermic reaction cause the surroundings to become warmer?
5. How can a temperature increase provide evidence that a reaction is exothermic?
6. Give one example of a chemical reaction or change that could be exothermic.

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 heat energy during an endothermic reaction?
3. What happens to the temperature of the surroundings during an endothermic reaction?
4. Why does an endothermic reaction cause the surroundings to become cooler?
5. How can a temperature decrease provide evidence that a reaction is endothermic?
6. Give one example of a chemical reaction or change that could be endothermic.

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. Why is breaking a chemical bond an endothermic process?
3. What happens to energy when a new chemical bond is formed?
4. Why is making a chemical bond an exothermic process?
5. Which releases energy: breaking bonds or making bonds?
6. Which requires energy: breaking bonds or making bonds?

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

1. When is a reaction exothermic in terms of the energy required to break bonds and the energy released when bonds form?
2. When is a reaction endothermic in terms of the energy required to break bonds and the energy released when bonds form?
3. A reaction requires 800 kJ mol^{-1} to break bonds and releases 1100 kJ mol^{-1} when new bonds form. Is the reaction exothermic or endothermic?
4. A reaction requires 1400 kJ mol^{-1} to break bonds and releases 1000 kJ mol^{-1} when new bonds form. Is the reaction exothermic or endothermic?
5. Why is a reaction exothermic when more energy is released forming bonds than is required to break bonds?
6. Why is a reaction endothermic when less energy is released forming bonds than is required to break bonds?

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

1. The bond energies are H–H = 436 kJ mol^{-1}, Cl–Cl = 242 kJ mol^{-1} and H–Cl = 431 kJ mol^{-1}. Calculate the energy change for $H_2 + Cl_2 \rightarrow 2HCl$.
2. The bond energies are H–H = 436 kJ mol^{-1}, O=O = 498 kJ mol^{-1} and O–H = 463 kJ mol^{-1}. Calculate the energy change for $2H_2 + O_2 \rightarrow 2H_2O$.
3. A reaction requires 1250 kJ mol^{-1} to break the bonds in the reactants and releases 1500 kJ mol^{-1} when bonds form in the products. Calculate the overall energy change and state whether the reaction is exothermic or endothermic.
4. A reaction requires 980 kJ mol^{-1} to break bonds and releases 760 kJ mol^{-1} when new bonds form. Calculate the overall energy change and state whether the reaction is exothermic or endothermic.
5. The energy required to break bonds is 1650 kJ mol^{-1} and the energy released when new bonds form is 1920 kJ mol^{-1}. Calculate the overall energy change and explain what the sign of your answer means.
6. A reaction has an overall energy change of +240 kJ mol^{-1}. The energy required to break the reactant bonds is 1350 kJ mol^{-1}. Calculate the energy released when the product bonds form.

7.15 Explain the term activation energy

1. What is meant by the activation energy of a chemical reaction?
2. Why must reacting particles have at least the activation energy for a reaction to occur?
3. How is activation energy related to successful collisions?
4. What happens to the rate of a reaction if fewer particles have enough energy to overcome the activation energy?
5. Why does increasing temperature usually increase the number of particles able to overcome the activation energy?
6. How does a catalyst affect the activation energy of a reaction?

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

1. What does a reaction profile show?
2. How is the activation energy shown on a reaction profile?
3. How can you identify an exothermic reaction from its reaction profile?
4. How can you identify an endothermic reaction from its reaction profile?
5. Where are the reactants, products and activation energy labelled on a reaction profile?
6. How does the energy level of the products compare with the energy level of the reactants in an exothermic reaction and in an endothermic reaction?

Topic 8 – Fuels and Earth science

Fuels

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

1. What elements are present in a hydrocarbon?
2. What is the defining feature of a hydrocarbon?
3. Is methane, CH_4, a hydrocarbon? Explain why.
4. Is ethanol, C_2H_5OH, a hydrocarbon? Explain why.
5. Is carbon dioxide, CO_2, a hydrocarbon? Explain why.
6. Classify C_3H_8, C_2H_4 and C_6H_6 as hydrocarbons or not hydrocarbons, giving a reason for each.

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. How can the carbon atoms in crude oil molecules be arranged?
4. Why is crude oil an important source of useful substances?
5. What is meant by saying that crude oil is a finite resource?
6. Explain why crude oil is both an important resource and 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 fractions?
2. Why can the hydrocarbons in crude oil be separated by fractional distillation?
3. What happens to crude oil when it is heated during fractional distillation?
4. What happens to hydrocarbons with different boiling points as they pass through the fractionating column?
5. Why do hydrocarbons condense at different levels in the fractionating column?
6. Explain how fractional distillation produces simpler, more useful mixtures from crude oil.

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?
2. What is petrol used for?
3. What is kerosene used for?
4. What is diesel oil used for?
5. What are the uses of fuel oil?
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 and hydrogen atoms generally change as the hydrocarbons become larger?
2. How does boiling point change as the hydrocarbon molecules become larger?
3. How does ease of ignition change as the hydrocarbon molecules become larger?
4. How does viscosity change as the hydrocarbon molecules become larger?
5. Which homologous series do most hydrocarbons in crude oil fractions belong to?
6. A fraction contains hydrocarbons with higher boiling points and greater viscosity than another fraction. What can you predict about the relative sizes of their molecules?

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

1. What is meant by the term homologous series?
2. How do the molecular formulae of neighbouring members of a homologous series differ?
3. What formula is shared by members of an homologous series?
4. How do physical properties change across a homologous series?
5. Why do members of the same homologous series have similar chemical properties?
6. The boiling points of four members of a homologous series are 36°C, 69°C, 98°C and 126°C. What does this trend suggest about their physical properties as molecular size increases?

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

1. What products are formed during the complete combustion of a hydrocarbon?
2. What gas must be present for complete combustion to occur?
3. What happens to the carbon atoms during complete combustion?
4. What happens to the hydrogen atoms during complete combustion?
5. Is complete combustion exothermic or endothermic?
6. Write a word equation for the complete combustion of a hydrocarbon.

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

1. What causes incomplete combustion of a hydrocarbon?
2. Why can incomplete combustion produce carbon monoxide?
3. Why can incomplete combustion produce carbon?
4. How does the oxygen supply affect whether combustion is complete or incomplete?
5. What are the two carbon-containing products that can form during incomplete combustion?
6. Explain why a restricted oxygen supply can cause a hydrocarbon fuel to undergo incomplete combustion.

8.9 Explain how carbon monoxide behaves as a toxic gas

1. Why is carbon monoxide toxic?
2. What happens to haemoglobin when carbon monoxide enters the blood?
3. Why does carbon monoxide reduce the amount of oxygen transported around the body?
4. Why can exposure to carbon monoxide be dangerous even when the gas cannot easily be detected?
5. What effect can carbon monoxide have on cells that require oxygen for respiration?
6. Explain why carbon monoxide poisoning can be fatal.

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

1. What toxic gas can be produced by incomplete combustion?
2. What solid pollutant can be produced by incomplete combustion?
3. Why is carbon monoxide particularly dangerous in poorly ventilated rooms?
4. What health problems can be caused by breathing in soot particles?
5. Why can incomplete combustion make fuel-burning appliances dangerous?
6. Explain why adequate ventilation is important when using appliances that burn carbon compounds.

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

1. What impurity in some hydrocarbon fuels can lead to sulfur dioxide formation?
2. What happens to sulfur-containing impurities when the fuel is burned?
3. Which gas is produced when sulfur impurities are oxidised during combustion?
4. Write a word equation for the formation of sulfur dioxide when sulfur burns in oxygen.
5. Why can burning some fossil fuels release sulfur dioxide into the atmosphere?
6. Explain the link between sulfur impurities in fuels and sulfur dioxide pollution.

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

1. What happens when sulfur dioxide dissolves in rainwater?
2. Why can sulfur dioxide contribute to acid rain?
3. What effect can acid rain have on aquatic ecosystems?
4. What effect can acid rain have on plants and trees?
5. How can acid rain damage buildings and statues made from carbonate-containing materials?
6. Explain how burning sulfur-containing fuels can eventually cause environmental damage through acid rain.

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. Which two gases from the air can react in a hot engine?
2. Why can nitrogen and oxygen react together inside an engine?
3. What are the products called when nitrogen and oxygen react at high temperatures?
4. Why are oxides of nitrogen considered pollutants?
5. Why does the high temperature inside an engine promote the formation of nitrogen oxides?
6. Explain how burning fuels in engines can lead to the formation of oxides of nitrogen.

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

1. What is the main environmental advantage of using hydrogen instead of petrol in a car?
2. What product is formed when hydrogen reacts with oxygen in a fuel cell?
3. Why does hydrogen not produce carbon dioxide when it is used as a fuel?
4. What are some disadvantages associated with producing and storing hydrogen?
5. Why does the environmental impact of hydrogen depend on how the hydrogen is produced?
6. Evaluate whether hydrogen is a better car fuel than petrol, considering emissions, production, storage and availability.

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. Which fossil resource is petrol obtained from?
2. Which fossil resource is kerosene obtained from?
3. Which fossil resource is diesel oil obtained from?
4. Where is methane found naturally as a fossil fuel?
5. Why are petrol, kerosene, diesel oil and methane described as non-renewable?
6. Identify whether petrol, kerosene, diesel oil and methane are obtained from crude oil or natural gas.

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

1. What type of hydrocarbon molecules are broken down during cracking?
2. What happens to large alkane molecules during cracking?
3. Why are the products of cracking generally smaller molecules?
4. Why can cracking produce alkenes?
5. What type of hydrocarbon is an alkene?
6. Explain why cracking can produce both smaller alkanes and alkenes from larger alkanes.

8.17 Explain why cracking is necessary

1. Why is cracking necessary when processing crude oil?
2. Why is there greater demand for some fractions than others?
3. How does cracking help meet the demand for smaller hydrocarbons?
4. Why are alkenes produced by cracking useful to the chemical industry?
5. How does cracking increase the usefulness of crude oil?
6. Explain why the supply of fractions obtained directly from crude oil does not always match their demand.

Earth and atmospheric science

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

1. What was the main source of gases in the Earth's early atmosphere?
2. How did volcanic activity contribute to the formation of the early atmosphere?
3. What happened to gases released by volcanoes over time?
4. Why was the early atmosphere very different from today's atmosphere?
5. Which geological process released large amounts of gases to form the early atmosphere?
6. Explain how volcanic activity led to the formation of Earth's early atmosphere.

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

1. What was the approximate amount of oxygen thought to be present in Earth's early atmosphere?
2. Which gas was present in large amounts in the early atmosphere?
3. Why was water vapour present in the early atmosphere?
4. What other gases were thought to be present in small amounts?
5. What evidence can scientists use to infer the composition of Earth's early atmosphere?
6. Explain how evidence can be used to support the model of an early atmosphere containing little or no oxygen and large amounts of carbon dioxide.

8.20 Explain how condensation of water vapour formed oceans

1. What caused water vapour in the early atmosphere to condense?
2. What happens to water vapour when it condenses?
3. Why did cooling of Earth allow liquid water to form?
4. How did condensation contribute to the formation of the oceans?
5. What change of state is involved when water vapour becomes liquid water?
6. Explain how cooling of the early Earth led to the formation of oceans.

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

1. What happened to some atmospheric carbon dioxide when the oceans formed?
2. Why did the formation of oceans decrease atmospheric carbon dioxide?

3. What happens when carbon dioxide dissolves in ocean water?
4. How did the formation of liquid oceans affect the composition of the atmosphere?
5. Why did the amount of atmospheric carbon dioxide decrease as the oceans formed?
6. Explain the link between ocean formation and the reduction in atmospheric carbon dioxide.

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. Which gas did primitive plants remove from the atmosphere during photosynthesis?
2. Which gas did primitive plants release during photosynthesis?
3. How did the growth of primitive plants affect atmospheric oxygen levels?
4. Why did atmospheric oxygen increase gradually rather than immediately?
5. What process carried out by primitive plants caused oxygen to be released?
6. Explain how photosynthesis by primitive plants changed the composition of Earth's atmosphere.

8.23 Describe the chemical test for oxygen

1. What is the chemical test for oxygen?
2. What happens to a glowing splint when it is placed in oxygen?
3. What observation indicates a positive test for oxygen?
4. Why is a glowing splint used rather than a lit splint?
5. What gas is identified if a glowing splint relights?
6. Describe how you would test an unknown gas to determine whether it contains oxygen.

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 meant by the greenhouse effect?
2. Which type of radiation is absorbed by greenhouse gases after being emitted by Earth?
3. Name three greenhouse gases.
4. What happens to the energy absorbed by greenhouse gases?
5. How does the greenhouse effect help keep Earth's surface warm?
6. Explain how carbon dioxide, methane and water vapour contribute to the 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 exists between atmospheric carbon dioxide concentration and global temperature?
2. How is fossil fuel consumption linked to changes in atmospheric carbon dioxide concentration?
3. Why can a correlation between carbon dioxide concentration and temperature be used as evidence for human-caused climate change?
4. Why does correlation alone not prove causation?
5. How can the location of measurements introduce uncertainty into climate data?
6. Evaluate evidence for human activity causing climate change, considering carbon dioxide concentration, fossil fuel consumption, temperature records and uncertainties in historical measurements.

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. Which human activities can increase atmospheric carbon dioxide levels?
3. How can livestock farming increase atmospheric methane levels?
4. What potential effects could increased carbon dioxide and methane have on climate?
5. What is meant by mitigating climate change?
6. Evaluate possible methods of mitigating increased greenhouse gas levels, considering their scale, risks and environmental implications.