Edexcel GCSE Triple Chem

Topic 1 – Key concepts in chemistry

Atomic structure

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

1. A Dalton atom was a solid, indivisible sphere with no internal structure.
2. The electron.
3. It showed that atoms contain smaller negatively charged particles, so atoms are divisible.
4. Protons and neutrons.
5. It showed that most of the atom is empty space surrounding a small, dense, positively charged nucleus.
6. New experimental evidence led to the discovery of subatomic particles and a more detailed model of the atom.

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

1. Protons, neutrons and electrons.
2. Protons and neutrons.
3. In shells around the nucleus.
4. Positive.
5. The positive charge of the protons balances the negative charge of the electrons.
6. Electrons occupy shells at different distances from the nucleus.

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

1. +1
2. 1
3. 0
4. 1
5. -1
6. 1/1840

1.4 Explain why atoms contain equal numbers of protons and electrons

1. A neutral atom has no overall charge, so the positive and negative charges must balance.
2. +1
3. -1
4. It becomes negatively charged.
5. It becomes positively charged.
6. Each proton contributes +1 and each electron contributes -1, so equal numbers give an overall charge of 0.

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

1. The nucleus is extremely small compared with the whole atom.
2. Only a tiny fraction of the atom's overall volume.
3. The nucleus occupies very little of the atom's volume.
4. The electron shells.
5. The nucleus is much smaller than the distances of the electrons from it.
6. It shows that most of the atom is empty space.

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

1. In the nucleus.
2. Protons and neutrons.
3. An electron has a very small relative mass.
4. A proton has a mass of about 1840 times that of an electron.
5. A neutron has a mass of about 1840 times that of an electron.
6. The nucleus contains the much more massive protons and neutrons.

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

1. The total number of protons and neutrons in the nucleus.
2. Mass number = number of protons + number of neutrons.
3. Protons and neutrons.
4. Electrons have negligible mass compared with protons and neutrons.
5. Number of neutrons = mass number – atomic number.
6. It gives the total number of protons and neutrons in the nucleus.

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

1. The number of protons in its nucleus.
2. Each element has a unique number of protons.
3. The number of protons defines the element.
4. The atomic number gives the number of protons.

5. It would become a different element.
6. Different elements have different numbers of protons.

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. Atoms of the same element with the same number of protons but different numbers of neutrons.
2. They contain the same number of protons.
3. The number of protons.
4. The number of neutrons.
5. It changes the mass number.
6. They contain different numbers of neutrons.

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

1. Number of protons = atomic number.
2. Number of electrons = atomic number for a neutral atom.
3. Number of neutrons = mass number − atomic number.
4. 8 protons, 8 neutrons, 8 electrons.
5. 11 protons, 12 neutrons, 11 electrons.
6. 17 protons, 18 neutrons, 17 electrons.

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

1. Most elements exist as mixtures of isotopes with different masses.
2. The relative atomic mass is a weighted mean of the masses of the isotopes.
3. Different isotopes occur in different abundances, so the average is weighted according to abundance.
4. More abundant isotopes have a greater effect on the relative atomic mass.
5. Chlorine is a mixture of isotopes with different masses, so its relative atomic mass is between the mass numbers of its common isotopes.
6. The relative atomic mass is an average for all naturally occurring isotopes, whereas the mass number refers to one atom.

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

1. Relative atomic mass = Σ(isotope mass × abundance) / Σ(abundances)
2. The mass of each isotope and its relative abundance.
3. To account for how common each isotope is.
4. Multiply each isotope mass by its abundance, add the results, then divide by the total abundance.

5. (10 × 20 + 11 × 80) / 100 = 10.8
6. (35 × 75 + 37 × 25) / 100 = 35.5

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. He arranged elements in order of increasing relative atomic mass while grouping elements with similar properties.
2. The properties of the elements and their compounds.
3. To identify recurring patterns in chemical behaviour.
4. Properties repeated periodically.
5. Similar elements were placed in the same groups.
6. It grouped elements according to repeating chemical patterns and properties.

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

1. He left gaps where he predicted undiscovered elements should occur.
2. To allow for elements that had not yet been discovered.
3. By using the properties and trends of neighbouring elements.
4. Strong evidence that his periodic table was valid.
5. The repeating patterns allowed properties to be predicted from surrounding elements.
6. The discoveries of elements with properties matching his predictions supported his table.

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

1. He generally observed that relative atomic mass increased as he arranged the elements.
2. Isotopes can give an element a relative atomic mass that does not fit the simple atomic-mass order.
3. Different isotopes have different masses and abundances.
4. The element with the higher atomic number can have isotopes that give it a lower average relative atomic mass.
5. Isotopes explained why some elements had to be placed out of strict relative-mass order to preserve similar properties.
6. Atomic number corresponds to proton number and gives the correct periodic 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. The number of protons in the nucleus.
2. Atomic number = number of protons.
3. It determines the element's position in increasing atomic number.
4. Each element has a unique number of protons.
5. 12
6. Match the atomic number to the element 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 order of increasing atomic number.
2. Periods.
3. Groups.
4. They have similar outer-electron arrangements and therefore similar chemical properties.
5. It gives the number of occupied electron shells.
6. Elements in the same group have similar outer-shell electron arrangements and therefore similar properties.

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

1. Mainly on the left and centre.
2. Mainly on the right.
3. Metals generally have fewer outer electrons and tend to lose electrons; non-metals generally have more outer electrons and tend to gain or share electrons.
4. Their electron arrangements are different.
5. Use its position relative to the metal/non-metal boundary.
6. Their outer-shell electron arrangements determine how they gain, lose or share electrons.

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. Electrons fill shells starting with the shell closest to the nucleus.
2. 2
3. 8
4. 2.8.1
5. 2.8.7

6. 2.8.8.2

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

1. The number of occupied shells gives the period.
2. For the first 20 elements, the number of electrons in the outer shell corresponds to the group number for Groups 1–7.
3. They have the same number of outer-shell electrons.
4. Count the occupied shells for the period and the outer-shell electrons for the group.
5. Group 1, period 3.
6. Group 7, period 3.

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. By transfer of electrons from one atom to another, forming oppositely charged ions that attract.
2. Metals.
3. Non-metals.
4. A positively charged ion formed when an atom loses electrons.
5. A negatively charged ion formed when an atom gains electrons.
6. They show the outer electrons of the atoms and the transferred electron(s) using different symbols.

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

1. An atom or group of atoms with an overall positive or negative charge.
2. By losing one or more electrons.
3. By gaining one or more electrons.
4. A cation is positive; an anion is negative.
5. The number of protons and electrons is unequal.
6. Yes, a polyatomic ion can contain more than one atom.

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

1. Number of protons = atomic number.
2. Number of neutrons = mass number − atomic number.
3. Electrons = atomic number − positive charge.

4. Electrons = atomic number + magnitude of negative charge.
5. 11 protons, 12 neutrons, 10 electrons.
6. 8 protons, 8 neutrons, 10 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. Lose 1 electron to form 1+ ions.
2. Lose 2 electrons to form 2+ ions.
3. Gain 2 electrons to form 2− ions.
4. Gain 1 electron to form 1− ions.
5. They have 1 or 2 outer electrons and lose them to obtain a full outer shell.
6. They need 2 or 1 electrons to complete their outer shell.

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

1. Usually a compound containing a simple anion with no oxygen in the name, such as chloride or oxide.
2. Usually a compound containing a polyatomic oxygen-containing ion, such as nitrate, carbonate or sulfate.
3. Oxide contains O^{2-}; nitrate contains $NO3^-$.
4. -ide
5. -ate
6. It can indicate whether the negative ion is a simple ion or an oxygen-containing polyatomic ion.

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

1. Combine the ions in the simplest whole-number ratio that gives overall charge 0.
2. Positive and negative charges must balance.
3. NaCl
4. MgCl2
5. Ca(OH)2
6. Al2(SO4)3

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

1. A giant ionic lattice.
2. A regular three-dimensional arrangement of ions.
3. Positive and negative ions alternate in a regular arrangement.

4. Strong electrostatic attractions between oppositely charged ions.
5. Electrical neutrality requires equal total positive and negative charge.
6. The ions have full charges and strong electrostatic attraction acts between neighbouring oppositely charged ions.

Covalent bonding

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

1. When two atoms share a pair of electrons.
2. A pair of electrons.
3. To obtain a more stable electron arrangement.
4. Non-metal atoms.
5. Sharing electrons helps fill their outer electron shells.
6. Show one shared pair of electrons between the bonded atoms.

1.29 Recall that covalent bonding results in the formation of molecules

1. Covalent bonding.
2. A group of atoms joined by covalent bonds.
3. By shared pairs of electrons.
4. Atoms share electrons and form discrete groups of bonded atoms.
5. Molecules.
6. Ionic bonding forms a giant ionic lattice; covalent bonding can form molecules.

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

1. About 10^-10 m.
2. About 10^-10 m.
3. The power of ten that gives the approximate size of the particle.
4. They are of a similar order of magnitude.
5. They are far too small to resolve with the naked eye.
6. For example, 1 × 10^-10 m.

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. Each hydrogen atom shares one electron to form one shared pair.
2. Hydrogen and chlorine share one pair of electrons.
3. Oxygen shares one pair with each hydrogen atom, forming two O-H bonds.
4. Carbon shares one pair of electrons with each of four hydrogen atoms, forming four C-H bonds.

5. The two oxygen atoms share two pairs of electrons, forming a double bond.
6. Carbon shares two pairs of electrons with each oxygen atom, forming two C=O double bonds.

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. Ionic, simple molecular covalent, giant covalent and metallic.
2. Strong electrostatic forces between ions require a lot of energy to overcome, giving high melting and boiling points.
3. Only weak intermolecular forces need to be overcome, giving low melting and boiling points.
4. Strong covalent bonds throughout the giant structure give high melting points and generally poor electrical conductivity, apart from structures such as graphite and graphene.
5. Delocalised electrons can move through the metal lattice.
6. Different bonding and structures determine the strength of forces, particle mobility and therefore melting point, solubility and electrical conductivity.

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. Strong electrostatic forces between oppositely charged ions require a lot of energy to overcome.
2. Strong electrostatic forces of attraction.
3. The ions are fixed in position and cannot move.
4. The ions are free to move and carry charge.
5. The ions are separated and can move through the solution.
6. The ions become mobile when molten or dissolved.

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. Only weak intermolecular forces need to be overcome.
2. Intermolecular forces.
3. Covalent bonds are much stronger than intermolecular forces.
4. They contain no mobile charged particles.

5. The electrons are held in covalent bonds and are not free to move through the substance.
6. Melting does not produce enough mobile charged particles to conduct electricity.

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

1. Giant covalent forms of carbon.
2. They are allotropes containing only carbon atoms arranged differently.
3. Giant covalent substances.
4. A different structural form of the same element.
5. Diamond and graphite.
6. Their atoms are arranged and bonded differently.

1.36 Describe the structures of graphite and diamond

1. Each carbon atom is covalently bonded to four other carbon atoms in a giant three-dimensional structure.
2. Four.
3. Layers of carbon atoms arranged in hexagons.
4. Carbon atoms form hexagonal layers.
5. Three.
6. Weak forces between the layers.

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. Each carbon atom has one delocalised electron that can move through the structure.
2. It conducts electricity and is stable at high temperatures.
3. Weak forces between layers allow the layers to slide over one another.
4. Only weak forces act between the layers.
5. It has a rigid giant structure with strong covalent bonds in all directions.
6. Diamond has strong covalent bonds throughout its three-dimensional structure, whereas graphite has weak forces between layers.

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

1. Carbon structures made from molecules containing carbon atoms arranged in hollow cages or tubes.
2. A spherical cage containing 60 carbon atoms.
3. A single layer of carbon atoms arranged in a hexagonal lattice.
4. It contains delocalised electrons that can move through the structure.

5. Strong covalent bonds connect the carbon atoms throughout the sheet.
6. Graphene is a single flat layer; C60 is a spherical molecular structure.

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

1. A large molecule made from many repeating units.
2. Long chains of carbon atoms with hydrogen atoms attached.
3. Carbon atoms.
4. They contain many repeating units joined together.
5. They form long, continuous chains.
6. Ethene.

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

1. The ability to be hammered or bent into shape without breaking.
2. Layers of metal ions can slide while metallic bonding is maintained.
3. Delocalised electrons can move through the structure.
4. Delocalised electrons are free to move.
5. Metallic bonding.
6. Strong metallic bonding explains strength, while mobile delocalised electrons explain electrical conductivity and the ability of layers to slide explains malleability.

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. It may not show the actual three-dimensional shape or relative sizes accurately.
2. The balls and sticks do not represent the real sizes and distances between atoms accurately.
3. The atoms appear as separate spheres with exaggerated gaps.
4. It cannot fully show the three-dimensional arrangement.
5. They simplify complex structures and make them easier to visualise.
6. Different models represent different aspects and therefore each has limitations.

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. Shiny solids, high melting points, high density and good electrical conductivity.

2. They contain delocalised electrons that can move through the metal.
3. Generally lower boiling points and poor electrical conductivity.
4. Most metals have higher melting points.
5. Most metals have higher densities.
6. Metals are generally good conductors; non-metals are generally poor conductors.

Calculations involving masses

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

1. Add the relative atomic masses of all atoms in the formula.
2. Percentage by mass = (mass of element in one formula unit / Mr of compound) × 100
3. Mr(H_2O) = 2(1) + 16 = 18
4. Mr($CaCO_3$) = 40 + 12 + 3(16) = 100
5. Percentage oxygen = (16 / 18) × 100 = 88.9%
6. Percentage carbon = (12 / 44) × 100 = 27.3%

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

1. The simplest whole-number ratio of atoms of each element in a compound.
2. Convert each reacting mass into moles, then find the simplest whole-number ratio.
3. Treat the percentages as masses out of 100 g, convert each to moles, then simplify the ratio.
4. Moles allow the amounts of different elements to be compared directly.
5. Divide all mole amounts by the smallest and multiply if necessary to obtain whole numbers.
6. The simplest ratio of the atoms of each element.

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

1. Divide all subscripts in the molecular formula by their highest common factor.
2. Multiplier = Mr of molecular formula / Mr of empirical formula.
3. The empirical formula and relative molecular mass.
4. Divide the molecular Mr by the empirical formula Mr.
5. CH_2O

6. Calculate the empirical formula Mr, divide the molecular Mr by it, then multiply every empirical-formula subscript by the resulting whole number.

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

1. Weigh magnesium, heat it strongly in oxygen, allow it to cool, then reweigh the magnesium oxide and calculate the masses of magnesium and oxygen before converting them to moles and finding the simplest ratio.
2. To determine the starting mass of magnesium.
3. To react the magnesium with oxygen and form magnesium oxide.
4. Mass of magnesium before heating and mass of magnesium oxide after heating.
5. To ensure the reaction is complete and all possible oxygen has reacted.
6. Calculate moles using n = mass / Ar, then divide by the smallest number of moles.

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

1. Mass is neither created nor destroyed in a chemical reaction.
2. No reactants or products can escape or enter the system.
3. The total mass before and after the precipitation reaction remains the same.
4. Gas can enter or leave the system.
5. The gas escapes from the flask, so the measured mass decreases.
6. Gas from the surroundings enters the flask, so the measured mass increases.

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

1. Convert the known mass to moles, use the mole ratio from the balanced equation, then convert the required moles back to mass.
2. The coefficients give the correct mole ratios between reactants and products.
3. From the coefficients in the balanced equation.
4. The mass of the known substance, its Ar/Mr, the balanced equation and the Ar/Mr of the unknown substance.
5. Mass = moles × Mr, or mass = moles × Ar for atoms.
6. Use the 2:2 mole ratio for Mg:MgO, calculate moles of Mg from the known mass, then calculate the same number of moles of MgO and convert to mass.

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

1. The mass of solute dissolved per cubic decimetre of solution.
2. Concentration = mass / volume
3. Mass in g and volume in dm^3.
4. concentration (g dm^{-3}) = mass of solute (g) / volume of solution (dm^3)
5. mass = concentration × volume
6. volume = mass / concentration

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

1. The amount of substance containing 6.02×10^{23} particles.
2. 6.02×10^{23}
3. 6.02×10^{23} mol^{-1}
4. Atoms, molecules, formula units or ions.
5. Its relative particle mass in grams.
6. The mass of one mole in grams is numerically equal to the 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. moles = mass / Mr
2. mass = moles × Mr
3. particles = moles × 6.02×10^{23}
4. moles = particles / (6.02×10^{23})
5. Calculate moles from mass, then multiply by 6.02×10^{23}.
6. Calculate moles from particles, then multiply by Mr.

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. The reactant that is completely used up first.
2. Once it is used up, no further product can form.
3. It is completely used up.
4. It remains unreacted.
5. Calculate the moles of each reactant and compare them with the stoichiometric ratio in the balanced equation.
6. The limiting reactant has already been completely consumed, so extra excess reactant cannot react.

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

1. 135 g Al = 5.0 mol.
   $4Al \rightarrow 2Al_2O_3$, so Al_2O_3 = 2.5 mol.
   Mr(Al_2O_3) = 102.
   Mass = 2.5 × 102 = 255 g.
2. 20 g Ca = 20 / 40 = 0.50 mol.
   CaO = 0.50 mol.
   Mr(CaO) = 56.
   Mass = 0.50 × 56 = 28 g.
3. 12 g Mg = 12 / 24 = 0.50 mol.
   MgO = 0.50 mol.
   Mr(MgO) = 40.
   Mass = 0.50 × 40 = 20 g.
4. 6 g C = 6 / 12 = 0.50 mol.
   CO_2 = 0.50 mol.
   Mr(CO_2) = 44.
   Mass = 0.50 × 44 = 22 g.
5. 28 g Fe = 28 / 56 = 0.50 mol.
   FeS = 0.50 mol.
   Mr(FeS) = 88.
   Mass = 0.50 × 88 = 44 g.
6. 12 g C = 1 mol and 44 g CO_2 = 1 mol.
   Simplest ratio = 1:1.

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. Closely packed in a fixed, regular arrangement.
2. Vibrate about fixed positions.
3. Close together but irregularly arranged.
4. Move around and past one another.
5. Far apart and randomly arranged.
6. Particle energy generally increases from solid to liquid to gas.

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. Melting.

2. Boiling or evaporation.
3. Condensation.
4. Freezing.
5. Sublimation.
6. Deposition.

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

1. Particles become less ordered, move more freely and gain energy.
2. The particles gain energy.
3. Particles move closer together, become more ordered and lose energy.
4. The particles lose energy.
5. Particles gain energy and move directly from a solid arrangement to widely separated gas particles.
6. Gas particles have greater energy and move more freely, overcoming the forces keeping them close together.

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

1. Below the melting point it is solid; above the melting point it is liquid unless the boiling point has been exceeded.
2. Below the boiling point it is liquid; above the boiling point it is gas, provided it is above the melting point.
3. Solid.
4. Liquid.
5. Gas.
6. Solid.

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. A substance containing only one element or one compound.
2. Everyday use can mean clean, uncontaminated or not mixed with unwanted material; in chemistry it has a precise meaning.
3. A pure substance contains only one substance; a mixture contains two or more substances.

4. Because only one substance is present.
5. Different elements and compounds can be physically mixed without reacting.
6. They have not necessarily chemically reacted and can retain their own properties.

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

1. A pure substance has a sharp melting point; a mixture usually melts over a range.
2. It suggests the substance is pure.
3. It suggests the substance is a mixture or impure.
4. It is likely to be a pure substance.
5. It is likely to be a mixture or impure.
6. Compare whether melting occurs sharply at one temperature or over a range.

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

1. A solution containing a soluble solid and a solvent.
2. A mixture of miscible liquids with different boiling points.
3. An insoluble solid mixed with a liquid.
4. A solution containing a dissolved solid.
5. A mixture of soluble substances such as dyes or inks.
6. Choose the method according to properties such as solubility, boiling point and physical state.

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

1. Identify the physical properties that differ, then choose the appropriate separation method.
2. Filtration.
3. Crystallisation.
4. Fractional distillation.
5. Simple distillation.
6. Separation techniques rely on differences in physical properties.

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. To dissolve and carry the soluble substances up the paper.
2. The solvent that moves through the paper.
3. The paper.
4. They have different attractions for the mobile and stationary phases.
5. They must dissolve in the mobile phase to move.
6. The solvent carries the substances at different rates, separating them into spots.

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. One spot indicates a pure substance; more than one spot indicates an impure substance or mixture.
2. Compare the positions of spots with those from known substances.
3. Rf = distance travelled by substance / distance travelled by solvent front
4. Rf = distance travelled by substance / distance travelled by solvent front
5. Rf = 4 / 8 = 0.50
6. Under the same conditions, the same substance has the same Rf value.

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

1. Heat the ink and collect the solvent that distils, leaving the dissolved substances behind.
2. Place ink on chromatography paper, place the paper in a suitable solvent and compare the separated spots.
3. The number and positions of the dye spots present.
4. Compare the spots or Rf values with known dyes.
5. Ink would dissolve and move with the solvent, whereas pencil does not.
6. Solvent, paper type, spot size, starting-line position, solvent depth and temperature.

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. Sedimentation, filtration and chlorination.
2. To allow larger suspended particles to settle.
3. To remove smaller insoluble particles.
4. To kill microorganisms.

5. Distil the seawater so that water evaporates and condenses separately from dissolved salts.
6. Dissolved salts could interfere with chemical tests and analysis.

Topic 3 – Chemical changes

Acids

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

1. H^+
2. OH^-
3. H^+
4. OH^-
5. H^+
6. OH^-

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

1. 7
2. Below 7
3. Above 7
4. pH 2
5. pH 12
6. It is neutral.

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

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

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

1. pH decreases.

2. pH increases.
3. pH increases.
4. pH decreases.
5. pH 2
6. pH 12

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

1. By 1.
2. By a factor of 10.
3. By a factor of 100.
4. A factor of 10.
5. A factor of 1000.
6. pH 2.

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. Measure a fixed volume of dilute hydrochloric acid into a container, measure the pH, add known amounts of powdered calcium hydroxide, mix and measure the pH after each addition.
2. The pH of the mixture.
3. To ensure changes in pH are due to the calcium hydroxide rather than changes in acid volume.
4. To know how much calcium hydroxide has been added and allow the pH change to be related to the amount added.
5. Stir the mixture thoroughly before measuring the pH.
6. The pH increases as more calcium hydroxide is added.

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

1. A solution containing a small amount of solute per unit volume.
2. A solution containing a large amount of solute per unit volume.
3. A concentrated solution.
4. Add solvent.
5. Remove solvent or add more solute.
6. The amount of solute present per unit volume.

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

1. It dissociates almost completely into ions in aqueous solution.
2. It only partially dissociates into ions.

3. Most acid molecules dissociate into ions.
4. Only some acid molecules dissociate into ions.
5. The splitting of acid molecules into ions in solution.
6. A larger proportion of its molecules dissociate into ions.

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

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

3.10 Recall that alkalis are soluble bases

1. A soluble base.
2. It is soluble in water.
3. Yes.
4. Yes.
5. No.
6. Copper oxide is insoluble 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. Salt + hydrogen.
2. Salt + water.
3. Salt + water.
4. Salt + water + carbon dioxide.
5. Hydrogen.
6. Carbon dioxide.

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

1. Use a lit splint; hydrogen produces a squeaky pop.
2. It burns with a squeaky pop.
3. A squeaky pop.
4. Hydrogen burns rapidly in the presence of oxygen.
5. A lit splint.
6. Carbon dioxide.

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

1. A reaction between an acid and a base.
2. An acid and a base.
3. Salt and water.
4. Yes.
5. Yes.
6. The acid is neutralised by the base.

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

1. H+
2. OH–
3. Water, H2O.
4. H+ + OH– → H2O
5. The acidic H+ ions and alkaline OH– ions are removed by forming water.
6. They react together to form H2O.

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

1. To ensure all the acid reacts.
2. To prevent unreacted reactant contaminating the salt solution.
3. By filtration.
4. The excess reactant is insoluble and remains as a solid.
5. Salt and water.
6. Any acid remaining can react with the excess solid until the acid is completely used up.

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

1. Both reactants are soluble, so excess cannot be removed by filtration.
2. It would remain dissolved in the solution.
3. The exact volume needed for complete neutralisation.
4. To ensure neither reactant is left in excess.
5. Salt and water.
6. The reactants react completely with each other, so no excess reactant remains.

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

1. Sulfuric acid.
2. To ensure all sulfuric acid reacts.
3. Filter the mixture.
4. To gently evaporate water without overheating the solution.
5. Gently heat the solution to evaporate some water, then allow it to cool so crystals form; filter and dry the crystals.
6. Strong heating can remove water of crystallisation.

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

1. Burette.
2. It allows a variable volume to be delivered accurately.
3. To identify the end point of the reaction.
4. Final reading – initial reading.
5. To obtain a reliable mean titre from closely agreeing results.
6. Use the titration to find exact reacting volumes, repeat without indicator using the same volumes, then evaporate water and crystallise the salt.

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

1. All common sodium, potassium and ammonium salts.
2. Yes.
3. Silver chloride and lead chloride.
4. Lead sulfate, barium sulfate and calcium sulfate.
5. Common carbonates are insoluble except sodium, potassium and ammonium carbonates.
6. Common hydroxides are insoluble except sodium, potassium and ammonium hydroxides.

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

1. An insoluble solid formed when two aqueous solutions are mixed.
2. Yes.
3. Silver chloride, AgCl.
4. Yes.
5. Barium sulfate, BaSO4.
6. Swap the ions, apply the solubility rules and identify any insoluble product.

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

1. Mix two suitable soluble solutions containing the required ions.
2. To provide the ions needed to form the insoluble salt.
3. They form an insoluble solid precipitate.
4. Filter the mixture.
5. To remove soluble impurities.
6. Filter, wash the precipitate with distilled water and dry it.

Electrolytic processes

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

1. An ionic compound that conducts electricity when molten or dissolved in water.
2. Ionic compounds.
3. The ions are free to move.
4. The ions are free to move through the solution.
5. No.
6. Its ions must become free to move.

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

1. The decomposition of an electrolyte using electrical energy.
2. Direct current, d.c.
3. It is decomposed into simpler substances.
4. Electrical energy.
5. Broken down into simpler substances.
6. New substances are formed.

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. Cathode.
2. Anode.
3. Opposite charges attract.
4. Opposite charges attract.
5. Negative.
6. Positive.

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

1. Cathode: copper. Anode: chlorine.
2. Cathode: hydrogen. Anode: chlorine.
3. Cathode: hydrogen. Anode: oxygen.
4. Cathode: hydrogen. Anode: oxygen.
5. Cathode: lead. Anode: bromine.
6. They do not react with the electrolyte and do not interfere with the products.

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

1. The metal forms at the cathode and the non-metal forms at the anode.
2. The metal.
3. The non-metal.
4. Sodium at the cathode and chlorine at the anode.
5. Magnesium at the cathode and bromine at the anode.
6. Identify the cation and anion, then the cation forms at the cathode and the anion forms at the anode.

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

1. It shows the electron transfer occurring at one electrode.
2. Gain of electrons.
3. Loss of electrons.
4. $Cu^{2+} + 2e^- \rightarrow Cu$
5. $2Cl^- \rightarrow Cl2 + 2e^-$
6. $Pb^{2+} + 2e^- \rightarrow Pb$

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

1. Loss of electrons.
2. Gain of electrons.
3. It loses electrons.
4. It gains electrons.
5. Oxidation.
6. Reduction.

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

1. Cathode.
2. Anode.
3. Gain of electrons.
4. Loss of electrons.
5. The cations gain electrons there.
6. The anions lose electrons there.

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. Cu^{2+} ions gain electrons and form copper.
2. Copper atoms lose electrons and enter the solution as Cu^{2+} ions.
3. It increases.
4. It decreases.
5. Cu^{2+} ions removed at the cathode are replaced by Cu^{2+} ions formed at the anode.
6. Use impure copper as the anode and pure copper as the cathode; copper dissolves from the anode and is deposited as pure copper at the cathode.

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

1. Connect copper sulfate solution to a d.c. supply using inert electrodes, pass current and observe the products at each electrode.
2. Use copper electrodes with copper sulfate solution and pass current while observing and measuring the electrodes.
3. A copper coating forms on the cathode.
4. Gas bubbles of oxygen form at the anode.
5. With inert electrodes, copper is deposited at the cathode and oxygen forms at the anode; with copper electrodes, copper transfers from anode to cathode.
6. Measure electrode masses before and after electrolysis and observe any gas formation or deposits.

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. More reactive metals react more vigorously with water.
2. More reactive metals generally react more readily with dilute acids and produce hydrogen.
3. A more reactive metal displaces a less reactive metal from its salt solution.
4. Magnesium is more reactive than copper.
5. A is more reactive than B.
6. More reactive metals react more readily and displace less reactive metals.

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

1. Electrons are transferred between the reacting species.
2. They are transferred to the less reactive metal ions.
3. They gain electrons and become metal atoms.
4. Mg is oxidised; Cu^{2+} is reduced.
5. Electron loss and electron gain occur together.
6. Mg loses electrons to form Mg^{2+}, while Cu^{2+} gains electrons to form Cu.

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

1. Potassium, sodium, calcium, magnesium, aluminium, carbon, zinc, iron, hydrogen, copper, silver, gold.
2. A higher position means a greater tendency to form cations.
3. They are above hydrogen and readily lose electrons to form ions while H+ ions are reduced to H2.
4. Copper, silver and gold.
5. Potassium loses electrons more readily and forms K+ ions more easily.
6. More reactive metals react more readily with water or acids, providing evidence for their tendency to form cations.

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. Ores in the Earth's crust.
2. A rock containing enough of a metal compound for the metal to be extracted economically.
3. They readily react with other substances and form compounds.
4. They are too unreactive to react readily with other elements.
5. Unreactive metals.
6. Gold is very unreactive, so it can remain as the uncombined element.

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

1. Gain of oxygen.
2. Loss of oxygen.
3. It gains oxygen.
4. It loses oxygen.
5. Copper oxide is reduced to copper.
6. Copper is oxidised to copper oxide.

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

1. The metal compound must be reduced to obtain the metal.
2. The metal ion is converted into the metal.
3. Oxygen is removed from the metal oxide.
4. Metal ions must gain electrons to become neutral metal atoms.
5. Oxygen is removed from the metal oxide.
6. The metal oxide loses oxygen and forms 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. They are less reactive than carbon, so their oxides can be reduced by carbon.
2. Aluminium is more reactive than carbon, so carbon cannot reduce aluminium oxide.
3. Aluminium ions require electrical energy to be reduced.
4. Electrolysis requires a large amount of electrical energy.
5. Iron is below carbon in the reactivity series.
6. Metals below carbon can generally be extracted by carbon reduction; metals above carbon require electrolysis.

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

1. Using bacteria to help extract metal compounds from ores.
2. Bacteria cause chemical reactions that convert metal compounds into soluble compounds that can be recovered.
3. Using plants to absorb metal ions from soil.
4. Grow plants that absorb metal ions, harvest them and recover the metals from the plant material.
5. Lower energy use and potentially less environmental damage.
6. They can be slow and may produce lower yields or require large areas of land.

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

1. Less reactive metals are more resistant to oxidation.
2. They have a lower tendency to lose electrons and react with oxygen.
3. Gold, because it is less reactive than potassium.
4. Potassium readily loses electrons and is easily oxidised.
5. A greater tendency to form a cation means a greater tendency to be oxidised.
6. The reactivity series ranks metals according to how readily they react, including their 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. It reduces the need to extract new ores.
2. It reduces mining, habitat destruction, pollution and waste.
3. Recycling generally uses less energy than extracting metals from ores.
4. It reduces raw-material and energy costs and can reduce waste-disposal costs.
5. Less new ore needs to be mined.
6. It conserves raw materials, reduces environmental damage and can save energy and money.

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

1. An assessment of the environmental impact of a product throughout its life.
2. Obtaining raw materials, manufacture, use and disposal.
3. Mining, extraction, transport and habitat damage.
4. Energy use, emissions, waste and resource use.
5. The product may use energy or release pollutants during its lifetime.
6. Disposal can produce waste, require energy or release pollutants.

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

1. The environmental impacts associated with each stage of the product's life.
2. Compare the impacts across all stages and consider the overall effect.
3. A product may have low impact at one stage but high impact at another.
4. The lower manufacturing impact could be outweighed by the greater disposal impact.

5. Add or compare the relevant environmental impacts for all stages.
6. Environmental impact must be considered over the whole life cycle, not one stage alone.

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. A reaction that can proceed in both directions.
2. ⇋
3. The reaction can proceed forwards and backwards.
4. Products can react to reform reactants.
5. Changing temperature, pressure or concentration can shift the equilibrium position.
6. A reversible reaction can proceed in both directions, whereas an irreversible reaction proceeds effectively in one direction.

4.14 — Explain what is meant by dynamic equilibrium

1. A state in a closed system where the forward and backward reactions occur at equal rates.
2. It becomes equal to the backward reaction rate.
3. It becomes equal to the forward reaction rate.
4. Reactants are being converted into products at the same rate that products are converted back into reactants.
5. Both reactions continue to occur.
6. The forward and backward reaction rates must be equal.

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

1. Nitrogen and hydrogen.
2. From the air.
3. From natural gas.
4. $N_2 + 3H_2 \rightleftharpoons 2NH_3$
5. The reaction is reversible and both forward and backward reactions occur.
6. Their concentrations remain constant.

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

1. 450 °C

2. 200 atmospheres
3. Iron.
4. 450 °C, 200 atmospheres and an iron catalyst.
5. It increases the rate of reaction.
6. 450 °C, 200 atmospheres and an iron catalyst.

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

1. Shifts to the right, increasing ammonia production.
2. Shifts to the right to replace some of the ammonia removed.
3. Shifts to the right because there are fewer moles of gas on the right.
4. Shifts to the left because there are more moles of gas on the left.
5. Shifts to the left because the forward reaction is exothermic.
6. Shifts to the right because the forward reaction releases heat.

Topic 5 – Separate chemistry 1

Transition metals, alloys and corrosion

5.1C — Recall that most metals are transition metals and that their typical properties include: a high melting point b high density c the formation of coloured compounds d catalytic activity of the metals and their compounds as exemplified by iron

1. High melting points, high densities, coloured compounds and catalytic activity.
2. Strong metallic bonding requires a large amount of energy to overcome.
3. Their atoms are relatively closely packed and have high relative atomic masses.
4. Many transition metal compounds are brightly coloured.
5. Catalytic activity is the ability to speed up chemical reactions without being used up; transition metals have suitable electronic structures that allow them to act as catalysts.
6. Iron is used as a catalyst in some industrial reactions, for example the Haber process.

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

1. Corrosion is the gradual destruction of a metal by chemical reactions with its surroundings.
2. The metal loses electrons and is converted into oxidised compounds.
3. Oxidation changes the metal into compounds that weaken and damage the original metal.
4. Corrosion of metals is caused by their oxidation.
5. The metal is gradually converted into corrosion products, weakening the object.
6. Oxidation of the metal.

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

1. Oxygen and water.
2. Without oxygen, the oxidation of iron cannot occur.
3. Without water, the electrochemical processes needed for rusting cannot occur.
4. Paint forms a barrier that prevents oxygen and water reaching the iron.
5. Sacrificial protection is using a more reactive metal to protect iron by allowing the more reactive metal to oxidise instead.
6. The more reactive metal is oxidised preferentially, preventing the iron from being oxidised.

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

1. Electroplating is coating an object with a thin layer of another metal using electrolysis.
2. It gives the object a smooth, attractive surface of the plating metal.
3. The coating prevents oxygen and water reaching the underlying metal.
4. A less reactive metal is less easily oxidised.
5. Metal ions gain electrons at the cathode and form a layer of metal on the object.
6. It can make the object more attractive while also protecting it from corrosion.

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

1. An alloy is a mixture containing a metal and one or more other elements.
2. Different-sized atoms distort the regular arrangement of metal atoms.
3. In a pure metal, identical layers of atoms can slide over each other more easily.
4. Different-sized atoms disrupt the regular arrangement of atoms.

5. The distorted layers cannot slide as easily past one another.
6. Alloying generally makes a metal stronger.

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

1. To change and improve its properties.
2. It can increase strength, hardness and other useful properties.
3. Alloy steels can be stronger, harder and more resistant to corrosion than pure iron.
4. Different metals produce different combinations of properties.
5. Different-sized atoms disrupt the regular arrangement and make the layers harder to slide.
6. They provide properties better suited to specific applications.

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

1. Aluminium has a low density, making it useful for aircraft.
2. Copper has high electrical conductivity and is ductile.
3. Gold is unreactive, attractive and malleable.
4. Magnalium combines relatively low density with increased strength.
5. Brass is strong, hard, corrosion-resistant and attractive.
6. Their uses depend on properties such as density, conductivity, strength, malleability and resistance to corrosion.

Quantitative analysis

5.8C — Calculate the concentration of solutions in mol dm^{-3} and convert concentration in g dm^{-3} into mol dm^{-3} and vice versa

1. Concentration = moles ÷ volume.
   $c = n / V$
2. 0.50 ÷ 2.0 = 0.25 mol dm^{-3}
3. Concentration in mol dm^{-3} = concentration in g dm^{-3} ÷ Mr.
4. 5.85 ÷ 58.5 = 0.100 mol dm^{-3}
5. Concentration in g dm^{-3} = concentration in mol dm^{-3} × Mr.
6. 0.250 × 40.0 = 10.0 g dm^{-3}

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

1. A pipette.
2. A burette allows an accurate variable volume of acid to be delivered.
3. To show when the neutralisation endpoint has been reached.
4. To estimate the approximate titre and identify a suitable range for accurate titres.
5. To show that the results are consistent and reliable.
6. Titre = final burette reading – initial burette reading.

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

1. Use the titre to calculate moles of the known solution, use the balanced equation to find the mole ratio, then calculate the unknown concentration.
   concentration = moles ÷ volume
2. Use the same method, using the known acid or alkali concentration and the balanced equation.
3. The balanced equation, known concentration, known volume and titre of the unknown solution.
4. $HCl + NaOH \rightarrow NaCl + H_2O$
   moles HCl = 0.100 × 25.0/1000 = 0.00250 mol
   1:1 ratio, so moles NaOH = 0.00250 mol
   concentration NaOH = 0.00250 ÷ 20.0/1000 = 0.125 mol dm^{-3}
5. It gives the mole ratio between the reacting substances.
6. Use the mole ratio from the balanced equation to calculate the required moles, then use:
   volume = moles ÷ concentration

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

1. Percentage yield = (actual yield ÷ theoretical yield) × 100
2. The maximum amount of product that could be produced.
3. The amount of product actually obtained.
4. (40 ÷ 50) × 100 = 80%
5. The actual yield equals the theoretical yield.
6. Reactions may be incomplete, products may be lost, or side reactions may occur.

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

1. Reactions may be incomplete, product may be lost, or unwanted side reactions may occur.
2. Some reactants remain unreacted, so less product is formed.
3. Product can be lost during filtration, transfer, washing or purification.
4. Some reactants form unwanted products instead of the desired product.
5. Some product remains in the original container or equipment.
6. Incomplete reactions, practical losses and side reactions.

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

1. Atom economy is the proportion of reactant atoms incorporated into the desired product.
2. A high proportion of the reactants becomes the desired product.
3. A larger proportion of reactant atoms becomes waste or unwanted products.
4. It reduces waste and makes better use of raw materials.
5. They become part of by-products or waste.
6. Fewer atoms are wasted in unwanted products.

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

1. Percentage atom economy = (Mr of desired product ÷ total Mr of reactants) × 100
2. (80 ÷ 100) × 100 = 80%
3. Mr($CaCO_3$) = 40 + 12 + 48 = 100
   Mr(CaO) = 40 + 16 = 56
   Atom economy = (56 ÷ 100) × 100 = 56%
4. Calculate the Mr of the desired product and the total Mr of the reactants, then apply the atom economy equation.
5. The desired product determines which reactant atoms are counted as being used usefully.
6. 75% of the atoms in the reactants are incorporated into the desired product.

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

1. Less waste is produced and raw materials are used more efficiently.
2. More of the desired product is obtained from a given amount of reactants.

3. A faster reaction can reduce production time and costs.
4. It affects how much product can be obtained at equilibrium.
5. They can provide an additional useful product and reduce waste.
6. The pathway should balance environmental, economic and practical factors, including high atom economy, high yield, suitable rate, favourable equilibrium and useful by-products.

5.16C — Describe the molar volume, of any gas at room temperature and pressure, as the volume occupied by one mole of molecules of any gas at room temperature and pressure

1. The volume occupied by one mole of gas at room temperature and pressure.
2. 24 dm^3
3. 24 000 cm^3
4. 2.5 × 24 = 60 dm^3
5. Moles = volume ÷ molar volume.
6. Equal numbers of gas particles occupy the same volume at the same temperature and pressure.

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

1. 0.50 × 24 = 12 dm^3
2. Convert the mass of the solid to moles, use the equation to find the moles of gas, then multiply by 24 dm^3.
3. Moles $CaCO_3$ = 10.0 ÷ 100 = 0.100 mol
   1:1 ratio, so CO_2 = 0.100 mol
   Volume = 0.100 × 24 = 2.40 dm^3
4. Moles CO_2 = 4.80 ÷ 24 = 0.200 mol
   1:1 ratio, so $CaCO_3$ = 0.200 mol
   Mass = 0.200 × 100 = 20.0 g
5. The coefficients give the mole ratio between substances.
6. Convert gas volume to moles using the molar volume, use the mole ratio, then convert moles of solid to mass using Mr.

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

1. Equal volumes of gases at the same temperature and pressure contain equal numbers of molecules.
2. N_2 : H_2 = 1 : 3, so 30 dm^3 H_2
3. N_2 : NH_3 = 1 : 2, so 30 dm^3 NH_3

4. $H_2 : O_2 = 2 : 1$, so 20 dm^3 O_2
5. The coefficients give the ratio of reacting gas volumes.
6. At the same temperature and pressure, gas volume is proportional to the number of moles.

The Haber process and fertilisers

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

1. Nitrogen reacts reversibly with hydrogen to form ammonia.
2. Nitrogen and hydrogen.
3. $N_2 + 3H_2 \leftrightharpoons 2NH_3$
4. From the air.
5. Mainly from natural gas.
6. The forward and reverse reactions can both occur, leading to equal forward and backward rates.

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

1. Increasing temperature increases the rate of reaching equilibrium.
2. Increasing pressure generally increases the rate of reaching equilibrium for reacting gases.
3. Increasing reactant concentration increases the rate of reaching equilibrium.
4. Increasing product concentration increases the rate of the reverse reaction and allows equilibrium to be reached sooner.
5. It increases the rate of reaching equilibrium.
6. It lowers the activation energy for both forward and reverse reactions equally, without changing the equilibrium position.

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

1. Expensive or limited raw materials increase production costs.
2. High energy use increases operating costs.

3. A lower temperature would give a higher equilibrium yield but a slower rate, so a compromise temperature is used.
4. Higher pressure favours ammonia formation but very high pressures are expensive and hazardous.
5. It increases the rate without changing the equilibrium position.
6. Industry must balance yield, rate, cost, safety and energy use.

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

1. Nitrogen, phosphorus and potassium.
2. Nitrogen is needed to make proteins and other important compounds.
3. Phosphorus is needed for healthy growth and processes such as energy transfer.
4. Potassium supports healthy plant growth and enzyme activity.
5. Nitrogen, phosphorus and potassium compounds.
6. To replace essential mineral ions and increase crop growth.

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

1. Ammonium nitrate.
2. Ammonia + nitric acid → ammonium nitrate
3. $NH_3 + HNO_3 \rightarrow NH_4NO_3$
4. It provides nitrogen needed for plant growth.
5. Neutralisation.
6. NH_4^+ and NO_3^-.

5.24C — Describe and compare: a the laboratory preparation of ammonium sulfate from ammonia solution and dilute sulfuric acid on a small scale b the industrial production of ammonium sulfate

1. Ammonia solution is reacted with dilute sulfuric acid, with the correct proportions found by titration, then the solution is concentrated and crystallised.
2. Ammonia solution and dilute sulfuric acid.
3. $2NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$
4. Laboratory production is small-scale; industrial production is large-scale.
5. Raw materials must first be processed to obtain ammonia and sulfuric acid.
6. The laboratory method directly reacts prepared chemicals on a small scale, whereas industrial production involves multiple large-scale stages.

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

1. A voltage.
2. One of the reactants is eventually used up.
3. The chemical reaction can no longer continue effectively and the voltage falls to zero.
4. Chemical energy → electrical energy.
5. The amount of reactants available.
6. The reactants are consumed during the chemical reaction.

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

1. Hydrogen and oxygen.
2. Water.
3. Hydrogen + oxygen → water
4. $2H_2 + O_2 \rightarrow 2H_2O$
5. Chemical reactions in the cell transfer energy to an electrical circuit.
6. The overall reaction produces water without forming carbon dioxide.

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

1. They produce electricity with water as the direct product and no carbon dioxide at the point of use.
2. Renewable electricity can be used to produce hydrogen without directly using fossil fuels.
3. Hydrogen is difficult to store and transport because it has a low density and can require high pressure or low temperatures.
4. Producing hydrogen from fossil fuels can generate carbon dioxide.
5. They can continuously generate electricity while reactants are supplied.
6. Cost, fuel availability, storage, infrastructure, efficiency, emissions and intended use.

Topic 6 – Groups in the periodic table

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. They are the elements in Group 1.
2. They are the elements in Group 7.
3. They are the elements in Group 0.
4. Group 1.
5. Group 7.
6. Group 0.

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

1. They are soft.
2. They have relatively low melting points.
3. Their atoms are held together less strongly than in many other metals.
4. Soft and relatively low melting points.
5. They are generally lower.
6. Their metallic bonding is relatively weak.

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

1. It floats, moves slowly, fizzes and gradually disappears, forming lithium hydroxide and hydrogen.
2. It floats, moves rapidly and melts into a ball, producing sodium hydroxide and hydrogen.
3. It reacts very vigorously, usually with a lilac flame, producing potassium hydroxide and hydrogen.
4. A metal hydroxide and hydrogen.
5. Hydrogen.
6. Metal + water → metal hydroxide + hydrogen

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

1. Reactivity increases down Group 1.
2. Sodium.
3. Potassium.
4. Metals lower down the group are predicted to be more reactive.
5. Rubidium.
6. The reaction becomes more vigorous.

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

1. The outer electron becomes easier to remove down the group.
2. It becomes further from the nucleus.
3. Shielding increases, reducing the attraction between the nucleus and the outer electron.
4. Potassium has a greater distance and more shielding.
5. It decreases.
6. They each have one electron in their outer shell and tend to lose it to form +1 ions.

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

1. Yellow-green gas.
2. Red-brown liquid.
3. Grey-black solid.
4. Chlorine.
5. Bromine.
6. Iodine.

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. Gas → liquid → solid.
2. Colours become darker.
3. Melting point increases.
4. Boiling point increases.
5. A solid.
6. Use the trends in colour, melting point, boiling point and physical state down the group.

6.8 — Describe the chemical test for chlorine

1. Expose damp blue litmus paper to the gas.
2. It is bleached.
3. Blue → red → white.
4. Chlorine reacts with water to form substances that bleach the dye.
5. Damp blue litmus paper is bleached.
6. Water is needed for chlorine to produce the reactive species responsible for bleaching.

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

1. Metal halides.
2. Sodium chloride, NaCl.
3. Potassium bromide, KBr.
4. Aluminium iodide, AlI_3.
5. Metal + halogen → metal halide
6. Use the pattern that halogens react with metals to form metal halides.

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

1. Hydrogen halides.
2. Hydrochloric acid solution.
3. They release hydrogen ions when dissolved in water.
4. H^+ and Cl^- ions.
5. Hydrobromic acid solution.
6. Other halogens are predicted to react with hydrogen to form hydrogen halides that produce acidic solutions in water.

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

1. Reactivity decreases down the group.
2. Chlorine displaces bromine:
   $Cl_2 + 2KBr \rightarrow 2KCl + Br_2$
3. Chlorine displaces iodine:
   $Cl_2 + 2KI \rightarrow 2KCl + I_2$
4. Bromine displaces iodine:
   $Br_2 + 2KI \rightarrow 2KBr + I_2$
5. Chlorine is more reactive and gains electrons more readily.
6. Astatine would be expected to be less reactive than chlorine, bromine and iodine.

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

1. Electrons are transferred between the halogen and halide ion.
2. It loses electrons.
3. It gains electrons.
4. Br^- is oxidised.
5. Cl_2 is reduced.
6. Cl_2 gains electrons and Br^- loses electrons.

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

1. The outer shell is further from the nucleus and shielding increases down the group.
2. To achieve a full outer electron shell.
3. It increases.
4. It increases.
5. Chlorine has less shielding and its outer shell is closer to the nucleus.
6. The attraction between the nucleus and an incoming electron becomes weaker.

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

1. They have full outer electron shells.
2. The outer shell is complete.
3. They do not need to gain, lose or share electrons readily.
4. Gaining an electron would disrupt a stable electron arrangement.
5. Losing an electron would disrupt a stable electron arrangement.
6. Their full outer shells make them chemically unreactive.

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

1. Helium has low density, so it provides lift.
2. Helium is non-flammable.
3. Argon is inert and prevents the hot filament reacting with oxygen.
4. Argon is inert, so it does not readily react during welding.
5. Its density is lower than air, so it provides lift.
6. Their lack of reactivity, low density and non-flammability make them useful in specific applications.

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. It increases down Group 0.
2. It increases down Group 0.
3. It increases down Group 0.
4. It increases down Group 0.
5. It would be expected to have a higher relative atomic mass, density, melting point and boiling point than xenon.
6. Extend the trends observed in known noble gases down the group.

Topic 7 – Rates of reaction and energy changes

Rates of reaction

7.1 Core Practical — Investigate the effects of changing the conditions of a reaction on the rates of chemical reactions

1. Place marble chips in hydrochloric acid and collect the carbon dioxide produced using a gas syringe; record gas volume against time.
2. Record the volume of carbon dioxide produced at regular time intervals.
3. Calculate the gradient of the graph; a steeper gradient means a faster reaction.
4. React sodium thiosulfate with hydrochloric acid and time how long it takes for the cross beneath the flask to disappear.
5. Sulfur precipitate forms and makes the mixture opaque, so the cross disappears as the reaction proceeds.
6. Keep marble-chip mass and size constant, vary acid concentration, measure gas volume at regular time intervals and compare reaction rates.

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

1. Collect the gas in a gas syringe or over water and measure its volume against time.
2. Measure the time taken for a precipitate to obscure a mark or use a colorimeter.
3. Measure the time taken for the colour to change.
4. Measure the mass at regular time intervals and calculate change in mass ÷ time.
5. Measure concentration at different times and calculate change in concentration ÷ time.

6. A quantity that changes with time, such as gas volume, mass, concentration or colour intensity.

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

1. Reacting particles must come into contact for bonds to break and form.
2. A collision with enough energy and the correct orientation to cause reaction.
3. More collisions per second increase the number of successful collisions per second.
4. More particles have enough energy to overcome the activation energy.
5. They may have insufficient energy or an unsuitable orientation.
6. Collision frequency and collision energy.

7.4 Explain the effects on rates of reaction of changes in temperature, concentration, surface area to volume ratio of a solid and pressure

1. Particles move faster, collide more frequently and more collisions have enough energy.
2. More particles are present in a given volume, increasing collision frequency.
3. More particles are exposed at the surface, increasing collision frequency.
4. Gas particles are closer together, increasing collision frequency.
5. Both collision frequency and the proportion of particles with sufficient energy increase.
6. Powdered calcium carbonate has a greater surface area to volume ratio, so more particles can collide with hydrochloric acid each second.

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

1. Rate of change of mass.
2. A faster rate of reaction.
3. Calculate the gradient of the graph.
4. The reaction has stopped or no more measurable product is being formed.
5. The graph with the steeper gradient represents the faster reaction.
6. $60 \div 30 = 2.0\ cm^3\ s^{-1}$

7.6 Describe a catalyst as a substance that speeds up the rate of a reaction without altering the products of the reaction

1. A substance that increases reaction rate without being chemically changed overall.
2. It increases the rate.
3. No.
4. It remains chemically unchanged overall.
5. Its mass remains unchanged overall.
6. It participates in the reaction pathway but is regenerated by the end.

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

1. The minimum energy required for a successful reaction.
2. It lowers the activation energy.
3. More particles have enough energy to react successfully.
4. It provides a pathway with a lower activation energy.
5. More collisions have enough energy to overcome the lower activation energy.
6. The peak is lower, so the activation energy is smaller.

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

1. Biological catalysts.
2. They speed up biochemical reactions without being used up overall.
3. They increase reaction rates.
4. Yeast enzymes catalyse fermentation of sugars to produce ethanol.
5. Fermentation.
6. They catalyse the reactions that convert carbohydrates into ethanol.

Heat energy changes in chemical reactions

7.9 Recall that changes in heat energy accompany the following changes

1. Depending on the salt, heat energy may be absorbed or released.
2. Energy is released or absorbed, causing a measurable temperature change.
3. A displacement reaction can release or absorb heat energy.
4. Formation of the precipitate can involve a heat energy change.
5. A temperature increase indicates heat released; a temperature decrease indicates heat absorbed.
6. Dissolving salts, neutralisation, displacement and precipitation.

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

1. A reaction that gives out heat energy.
2. They warm up.
3. Heat is transferred to the surroundings.
4. It is released to the surroundings.
5. Negative.
6. Combustion or neutralisation.

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

1. A reaction that takes in heat energy.
2. They cool down.
3. Heat is transferred from the surroundings into the reaction.
4. It is absorbed from the surroundings.
5. Positive.
6. Thermal decomposition or some dissolving reactions.

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

1. Energy is required to overcome the attractive forces holding the atoms together.
2. Endothermic.
3. Exothermic.
4. The forming atoms move into a lower-energy, more stable arrangement.
5. Energy is absorbed.
6. Energy is released.

7.13 Recall that the overall heat energy change for a reaction is exothermic/endothermic according to bond energies

1. More energy is released making bonds than is required to break bonds.
2. More energy is required to break bonds than is released making bonds.
3. Exothermic.
4. Endothermic.
5. More energy leaves the system than enters it.
6. More energy enters the system than leaves it during bond formation.

7.14 Calculate the energy change in a reaction given the energies of bonds

1. Energy change = energy required to break bonds – energy released forming bonds.
2. 500 – 700 = –200 kJ mol^{-1}
3. 900 – 650 = +250 kJ mol^{-1}
4. 1200 – 1500 = –300 kJ mol^{-1}
5. 1800 – 1400 = +400 kJ mol^{-1}
6. Energy released = 1250 kJ mol^{-1}

7.15 Explain the term activation energy

1. The minimum energy required for a reaction to occur.
2. The particles need enough energy for bonds to break and a reaction to take place.
3. The collision is unsuccessful.
4. From the energy level of the reactants to the top of the reaction profile.
5. A higher activation energy generally means a slower reaction.
6. More particles can undergo successful collisions.

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

1. Reactants, products, activation energy and energy change.
2. Reactants, products, activation energy and energy change.
3. From the reactant energy level to the peak.
4. From the reactant energy level to the peak.
5. Products are at a lower energy than reactants.
6. Products are at a higher energy than reactants.

Topic 8 – Fuels and Earth science

Fuels

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

1. A compound containing carbon and hydrogen only.
2. Carbon and hydrogen.
3. Yes.
4. No; ethanol also contains oxygen.
5. No; carbon dioxide contains oxygen.
6. Check whether the formula contains only carbon and hydrogen.

8.2 Describe crude oil as a complex mixture of hydrocarbons

1. A complex mixture of hydrocarbons.
2. It contains many different hydrocarbon molecules.
3. Carbon atoms can form chains or rings.
4. It provides fuels and feedstock for the petrochemical industry.
5. It provides raw materials for making useful chemicals and materials.
6. It is finite and cannot be replaced quickly enough to match consumption.

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

1. Fractional distillation.
2. They have different boiling points.
3. Crude oil is heated and the hydrocarbons vaporise and condense at different temperatures.
4. They rise higher in the column before condensing.
5. They condense lower down in the column.
6. Each fraction contains hydrocarbons with a similar range of boiling points.

8.4 Recall the names and uses of the following fractions

1. Domestic heating and cooking.
2. Fuel for cars.
3. Fuel for aircraft.
4. Fuel for some cars and trains.
5. Fuel for large ships and some power stations.
6. Surfacing roads and roofing.

8.5 Explain how hydrocarbons in different fractions differ from each other

1. Generally increases from lighter to heavier fractions.
2. It increases.
3. They become harder to ignite.
4. It increases.
5. Alkanes.
6. Their molecules have different sizes and therefore different intermolecular forces.

8.6 Explain a homologous series

1. A series of compounds with the same general formula, similar chemical properties and a gradual change in physical properties.
2. The same general formula.
3. They differ by CH_2.
4. They change gradually.
5. They are similar.
6. CH_2.

8.7 Describe the complete combustion of hydrocarbon fuels

1. Carbon dioxide and water.
2. Energy is released.
3. Hydrocarbon fuel and oxygen.
4. Hydrocarbon + oxygen → carbon dioxide + water
5. They are oxidised to carbon dioxide.
6. They are oxidised to water.

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

1. There is insufficient oxygen.
2. Carbon (soot).
3. Carbon monoxide, CO.
4. There is not enough oxygen for all carbon to be fully oxidised to CO_2.
5. It is limited.
6. Carbon particles and carbon monoxide.

8.9 Explain how carbon monoxide behaves as a toxic gas

1. It prevents blood from transporting enough oxygen.
2. It binds strongly to haemoglobin.
3. Haemoglobin.
4. It reduces the amount of oxygen haemoglobin can carry.
5. Cells receive less oxygen for aerobic respiration.
6. It is colourless and odourless.

8.10 Describe the problems caused by incomplete combustion producing carbon monoxide and soot

1. Carbon monoxide poisoning.
2. Carbon monoxide reduces the blood's ability to transport oxygen.
3. Soot is fine particles of carbon.
4. It can damage the respiratory system and contribute to breathing problems.

5. There may be less oxygen available, increasing incomplete combustion.
6. Carbon monoxide and soot.

8.11 Explain how impurities in some hydrocarbon fuels result in sulfur dioxide

1. The fuels contain sulfur impurities.
2. From sulfur-containing compounds in the fuel.
3. They are oxidised during combustion.
4. Sulfur dioxide, SO_2.
5. It contributes to air pollution and acid rain.
6. It must be oxidised by oxygen during combustion.

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

1. Sulfur dioxide dissolves in rainwater and forms acidic solutions.
2. Acidic substances are formed.
3. It can lower the pH of lakes and rivers, harming aquatic organisms.
4. It reacts with limestone and damages buildings and statues.
5. It can damage leaves, soil and forests.
6. It causes acid rain and environmental damage.

8.13 Explain why oxygen and nitrogen can react together at high temperatures

1. High temperatures provide enough energy for nitrogen and oxygen molecules to react.
2. Very high temperatures, such as those inside engines.
3. Oxides of nitrogen, NO_x.
4. They contribute to air pollution and can cause acid rain and photochemical pollution.
5. More particles have enough energy to overcome the activation energy.
6. Nitrogen and oxygen.

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

1. Hydrogen produces no carbon dioxide at the point of use in a fuel cell.
2. The fuel cell's overall product is water rather than carbon dioxide.
3. Hydrogen is difficult to store and transport and requires suitable infrastructure.
4. Hydrogen made from fossil fuels can produce carbon dioxide.

5. It must be stored at high pressure or low temperature because of its low density.
6. Emissions, production method, storage, cost, infrastructure, availability, efficiency and safety.

8.15 Recall that petrol, kerosene and diesel oil are non-renewable fossil fuels

1. They are obtained from finite crude oil reserves.
2. Crude oil.
3. Crude oil.
4. Crude oil.
5. Natural gas.
6. They are used much faster than they can be naturally replaced.

8.16 Explain how cracking involves breaking down larger saturated hydrocarbons

1. Breaking large hydrocarbon molecules into smaller, more useful molecules.
2. Large, saturated hydrocarbons called alkanes.
3. They are broken into smaller hydrocarbon molecules.
4. Smaller alkanes and alkenes.
5. They contain carbon-carbon double bonds.
6. Alkanes have only single C–C bonds; alkenes contain a C=C bond.

8.17 Explain why cracking is necessary

1. It produces more useful smaller hydrocarbons from larger ones.
2. Smaller hydrocarbons are in greater demand.
3. It produces more hydrocarbons suitable for petrol and other fuels.
4. They are used to make polymers and other chemicals.
5. It matches supply more closely to demand.
6. Crude oil contains too much of some heavy fractions and too little of some more useful smaller hydrocarbons.

Earth and atmospheric science

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

1. Volcanic activity.
2. Volcanoes released gases into the atmosphere.
3. The gases accumulated around the Earth.

4. It released large amounts of gases from Earth's interior.
5. Mainly carbon dioxide and water vapour, with smaller amounts of other gases.
6. It contained little or no oxygen and much more carbon dioxide and water vapour.

8.19 Describe the Earth's early atmosphere

1. Little or no oxygen.
2. Carbon dioxide.
3. Water vapour was released by volcanic activity.
4. Small amounts of other gases.
5. Ancient rocks can provide evidence of the atmospheric conditions when they formed.
6. Evidence from ancient rocks, fossils and geological records.

8.20 Explain how condensation of water vapour formed oceans

1. As the Earth cooled, water vapour condensed into liquid water.
2. Condensation.
3. Cooling lowers the energy of water molecules, allowing liquid water to form.
4. Condensed water collected on the Earth's surface and formed oceans.
5. It condensed into liquid water.
6. Gravity allowed liquid water to collect on the surface.

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

1. Carbon dioxide dissolved in the newly formed oceans.
2. Carbon dioxide is soluble in water.
3. It moved from the atmosphere into the oceans.
4. More carbon dioxide was removed from the atmosphere as more ocean water formed.
5. The newly formed liquid water dissolved carbon dioxide.
6. The atmospheric concentration decreased.

8.22 Explain how growth of primitive plants affected carbon dioxide and oxygen

1. They removed carbon dioxide from the atmosphere.
2. They released oxygen.
3. Carbon dioxide.
4. Oxygen.
5. More photosynthesis occurred, releasing more oxygen.

6. Photosynthesis gradually removed CO_2 and added O_2 to the atmosphere.

8.23 Describe the chemical test for oxygen

1. Insert a glowing splint into the gas.
2. The splint relights.
3. A glowing splint relights.
4. Oxygen supports combustion.
5. The glowing-splint test.
6. Oxygen.

8.24 Describe the greenhouse effect

1. The warming effect caused when atmospheric gases absorb and re-emit infrared radiation.
2. Infrared radiation.
3. Carbon dioxide, methane and water vapour.
4. They absorb infrared radiation emitted by the Earth and re-release energy in all directions.
5. They emit infrared radiation again.
6. Less heat energy would be retained in the atmosphere.

8.25 Evaluate the evidence for human activity causing climate change

1. Atmospheric CO_2 concentration and global temperature have increased over time.
2. Burning fossil fuels increases atmospheric CO_2.
3. The three quantities show correlated long-term changes consistent with human activity contributing to warming.
4. Correlation alone does not establish cause and effect.
5. Different locations can have different local conditions and measurements may not represent the whole planet equally.
6. Older measurements may be less accurate or less complete.

8.26 Describe today's atmosphere and climate change

1. Mainly nitrogen and oxygen, with smaller amounts of argon, carbon dioxide and other gases.
2. Burning fossil fuels transfers carbon from long-term stores into the atmosphere as CO_2.
3. Livestock farming produces methane, particularly from digestion and waste.

4. They increase the greenhouse effect and can contribute to global warming and climate change.
5. Renewable energy, energy efficiency, carbon capture, reforestation and changes in agriculture and transport.
6. Different methods have different costs, scales, risks and environmental impacts.

Topic 9 – Separate chemistry 2

Qualitative analysis: tests for ions

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

1. It must distinguish that ion from other possible ions.
2. A test that produces a characteristic result for one specific ion.
3. It could not reliably identify the specific ion.
4. The characteristic result can be matched to the known behaviour of the ion.
5. Different ions have different chemical properties.
6. The identity of the ion would be ambiguous.

9.2C — Describe flame tests to identify the following ions in solids

1. Red.
2. Yellow.
3. Lilac.
4. Orange-red.
5. Blue-green.
6. Heat the solid in a flame and compare the observed flame colour with known colours.

9.3C — Describe tests to identify aluminium, calcium, copper, iron(II), iron(III) and ammonium ions using sodium hydroxide solution

1. Add sodium hydroxide solution: a white precipitate forms, which dissolves in excess sodium hydroxide.
2. A white precipitate is formed.
3. A blue precipitate.
4. A green precipitate.
5. A red-brown precipitate.

6. Add sodium hydroxide and warm gently; ammonia gas is released and turns damp red litmus paper blue.

9.4C — Describe the chemical test for ammonia

1. Expose damp red litmus paper to the gas.
2. It turns blue.
3. Blue.
4. Ammonia must dissolve in water to produce alkaline conditions.
5. It indicates an alkaline gas, consistent with ammonia.
6. Damp red litmus paper turns blue.

9.5C — Describe tests to identify carbonate, sulfate, chloride, bromide and iodide ions

1. Add dilute acid; effervescence occurs as CO_2 is produced.
2. Bubble the gas through limewater; it turns cloudy.
3. Add dilute hydrochloric acid, then barium chloride solution; a white precipitate indicates sulfate.
4. A white precipitate of barium sulfate.
5. Acidify with dilute nitric acid, then add silver nitrate solution.
6. Chloride → white precipitate; bromide → cream precipitate; iodide → yellow precipitate.

9.6C — Core Practical: Identify the ions in unknown salts

1. Test separate portions for the cation and anion using the appropriate tests, then combine the results.
2. Lithium → red; sodium → yellow; potassium → lilac; calcium → orange-red; copper(II) → blue-green.
3. Add sodium hydroxide solution and observe precipitate colours or ammonia release for ammonium ions.
4. Carbonate → acid and CO_2 test; sulfate → hydrochloric acid and barium chloride; halides → nitric acid and silver nitrate.
5. To avoid one test reagent contaminating or interfering with another test.
6. Match the full set of positive results to the known ions.

9.7C — Identify the ions in unknown salts

1. Compare each test result with the known characteristic test results.
2. K^+ and Br^-.
3. Cu^{2+} and SO_4^{2-}.
4. Ca^{2+} and Cl^-.
5. Fe^{3+}.

6. One result may not uniquely identify the salt, so multiple results confirm the identification.

9.8C — Describe that instrumental methods of analysis are available

1. Methods using instruments to detect or measure substances.
2. They can detect very small concentrations.
3. They provide precise numerical measurements.
4. They can automate measurements and analyse samples rapidly.
5. The ability to detect small quantities of a substance.
6. They may provide greater sensitivity, accuracy and speed.

9.9C — Evaluate data from a flame photometer

1. Read the unknown's signal on the calibration curve and determine the corresponding concentration.
2. Ion concentration against measured signal intensity.
3. Locate the unknown reading on the graph and read across to the concentration axis.
4. Compare the measured data with reference values for known ions.
5. Match the unknown value with the closest reference value.
6. It provides the relationship needed to convert instrument readings into concentration.

Hydrocarbons

9.10C — Recall the formulae of molecules of the alkanes, methane, ethane, propane and butane

1. CH_4
2. C_2H_6
3. C_3H_8
4. C_4H_{10}
5. C_nH_{2n+2}
6. Methane: CH_4; ethane: CH_3–CH_3; propane: CH_3–CH_2–CH_3; butane: CH_3–CH_2–CH_2–CH_3, showing all covalent bonds.

9.11C — Explain why the alkanes are saturated hydrocarbons

1. A hydrocarbon containing only single bonds between carbon atoms and the maximum possible number of hydrogen atoms.
2. They contain only single C–C bonds and have the maximum number of hydrogen atoms.

3. Single covalent C–C bonds.
4. There are no multiple C–C bonds, so each carbon has the maximum number of hydrogen atoms.
5. It contains only single C–C bonds.
6. A C=C bond would mean fewer hydrogen atoms and therefore an unsaturated hydrocarbon.

9.12C — Recall the formulae of molecules of the alkenes, ethene, propene, butene

1. C_2H_4
2. C_3H_6
3. C_4H_8
4. C_4H_8
5. C_4H_8
6. Ethene: $CH_2=CH_2$; propene: $CH_2=CH–CH_3$; but-1-ene: $CH_2=CH–CH_2–CH_3$; but-2-ene: $CH_3–CH=CH–CH_3$.

9.13C — Explain why the alkenes are unsaturated hydrocarbons

1. A hydrocarbon containing one or more carbon-carbon double or triple bonds.
2. They contain a carbon-carbon double bond and therefore do not contain the maximum number of hydrogen atoms.
3. C=C
4. A double covalent bond.
5. It means the molecule does not have the maximum possible number of hydrogen atoms.
6. Alkenes contain C=C; alkanes contain only single C–C bonds.

9.14C — Recall the addition reaction of ethene with bromine

1. Addition reaction.
2. 1,2-dibromoethane.
3. The C=C double bond becomes a C–C single bond.
4. $CH_2Br–CH_2Br$
5. A reaction in which atoms are added across a carbon-carbon double bond.
6. Bromine adds across the C=C bond of other alkenes in the same way.

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

1. Add bromine water to the hydrocarbon.

2. It decolourises.
3. It remains orange.
4. Bromine reacts by addition across the C=C bond.
5. The hydrocarbon contains a C=C double bond.
6. The C=C bond.

9.16C — Describe how the complete combustion of alkanes and alkenes involves oxidation

1. Carbon dioxide and water.
2. Carbon dioxide and water.
3. Oxygen is added to the carbon and hydrogen during oxidation.
4. They are oxidised to carbon dioxide.
5. They are oxidised to water.
6. Carbon dioxide and water.

Polymers

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

1. A substance of high average relative molecular mass made from repeating units.
2. It contains a very large number of atoms joined in long chains.
3. Monomers.
4. The same structural unit is repeated many times.
5. It is much higher.
6. A polymer contains many repeating units joined into a long chain.

9.18C — Describe how ethene molecules combine together in a polymerisation reaction

1. Addition polymerisation.
2. Poly(ethene).
3. The C=C opens and forms single C–C bonds to neighbouring monomers.
4. $-CH_2-CH_2-$
5. Ethene.
6. No small molecule is eliminated; the monomers add together.

9.19C — Describe how other addition polymers can be made

1. A monomer containing a C=C bond.
2. Propene.

3. Chloroethene.
4. Tetrafluoroethene.
5. The double bond opens to form single bonds in the polymer chain.
6. Each contains a C=C double bond that can undergo addition polymerisation.

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

1. Identify the repeating unit and restore the C=C bond.
2. The C–C single bond in the polymer backbone between the appropriate atoms.
3. Replace the relevant C–C single bond with C=C and add the appropriate hydrogen or substituent arrangement.
4. Open the C=C bond and join the monomers into a chain.
5. The C=C bond becomes C–C bonds linking neighbouring monomer units.
6. Find the repeating unit, then restore the C=C bond to obtain the monomer.

9.21C — Explain how the uses of polymers are related to their properties

1. Properties such as strength, flexibility, chemical resistance and electrical insulation determine suitable uses.
2. It is lightweight, flexible and waterproof.
3. It is strong and durable.
4. It is durable, chemically resistant and a good electrical insulator.
5. It is chemically unreactive and has a low-friction surface.
6. The required properties of the use indicate which polymer is suitable.

9.22C — Explain why polyesters are condensation polymers

1. Small molecules, including water, are eliminated during polymer formation.
2. Carboxylic acid groups and alcohol groups.
3. A di-carboxylic acid monomer.
4. A diol monomer.
5. An ester link.
6. An –OH group and an –H from the reacting groups combine to form H_2O.

9.23C — Describe some problems associated with polymers

1. Many polymers depend on finite fossil-fuel feedstocks.
2. They are often non-biodegradable.
3. They remain in landfill for long periods.
4. Toxic or harmful gases may be released.

5. Different polymers have different properties and must be separated before processing.
6. They can consume finite resources, persist in landfill and cause pollution during disposal.

9.24C — Evaluate the advantages and disadvantages of recycling polymers

1. Less polymer waste enters landfill and less new material is needed.
2. Recycled polymers replace some new raw materials.
3. It can reduce raw-material costs but collection, sorting and processing can be expensive.
4. Sorting, contamination and repeated recycling can limit their usefulness.
5. Mixed polymers may not process together successfully.
6. Consider environmental benefits, processing costs, starting-material availability and the quality of recycled products.

9.25C — Recall that DNA, starch and proteins are polymers

1. DNA is a polymer.
2. Nucleotides.
3. Four.
4. Sugars.
5. Amino acids.
6. DNA → nucleotides; starch → sugars; proteins → amino acids.

Alcohols and carboxylic acids

9.26C — Recall the formulae of molecules of the alcohols

1. CH_3OH
2. C_2H_5OH
3. C_3H_7OH
4. C_4H_9OH
5. $C_nH_{2n+1}OH$
6. Methanol: CH_3OH; ethanol: CH_3CH_2OH; propan-1-ol: $CH_3CH_2CH_2OH$; butan-1-ol: $CH_3CH_2CH_2CH_2OH$.

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

1. –OH
2. Hydroxyl group.

3. Water is removed.
4. An alkene.
5. A C=C double bond.
6. Removing H_2O leaves a carbon-carbon double bond.

9.28C — Core Practical: Investigate the temperature rise produced in water by combustion of alcohols

1. Burn a measured mass or volume of alcohol beneath a container of a known mass of water and measure the temperature rise.
2. It allows a fair comparison because the same amount of water is heated each time.
3. Initial temperature, final temperature and mass of water.
4. Use the same mass of water, same apparatus, same starting temperature and same burner position.
5. The mass of water and experimental setup should remain constant.
6. To keep the transfer of heat to the water as similar as possible between experiments.

9.29C — Recall the formulae of molecules of the carboxylic acids

1. HCOOH
2. CH_3COOH
3. CH_3CH_2COOH
4. $CH_3CH_2CH_2COOH$
5. $C_nH_{2n+1}COOH$
6. Methanoic acid: HCOOH; ethanoic acid: CH_3COOH; propanoic acid: CH_3CH_2COOH; butanoic acid: $CH_3CH_2CH_2COOH$.

9.30C — Recall that the functional group in carboxylic acids is –COOH

1. –COOH
2. Carboxyl group.
3. They partially dissociate in water to produce H^+ ions.
4. H^+ ions.
5. They turn blue litmus red or give other acidic indicator results.
6. The –COOH group.

9.31C — Recall that ethanol can be oxidised to produce ethanoic acid

1. Ethanoic acid.
2. A carboxylic acid.
3. It is oxidised to the –COOH group.

4. A carboxylic acid.
5. A primary alcohol.
6. Other primary alcohols can similarly be oxidised to the corresponding carboxylic acids.

9.32C — Recall members of a given homologous series have similar reactions

1. They contain the same functional group.
2. The same functional group.
3. It is responsible for the main chemical reactions.
4. Identify the functional group and apply the known reaction to the new member.
5. The homologous series and its functional group.
6. Their functional groups are the same.

9.33C — Describe the production of ethanol by fermentation

1. Ferment carbohydrates using yeast in aqueous solution.
2. Carbohydrates such as sugars.
3. Yeast provides enzymes that catalyse fermentation.
4. Water provides the medium in which the yeast enzymes and substrates can react.
5. Ethanol and carbon dioxide.
6. Enzymes in yeast.

9.34C — Explain how to obtain a concentrated solution of ethanol by fractional distillation

1. Ethanol and water have different boiling points.
2. Their different boiling points.
3. The mixture is heated and vapours condense at different temperatures.
4. Ethanol has a lower boiling point than water, so the distillate becomes richer in ethanol.
5. Ethanol and water cannot normally be completely separated by simple fractional distillation.
6. Ethanol has a lower boiling point than water.

Bulk and surface properties of matter including nanoparticles

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

1. About 1 nm to 100 nm.
2. Nanoparticles are much larger than individual atoms.
3. Nanoparticles are generally larger than small molecules.
4. Their dimensions are measured in nanometres.
5. Nanometres (nm).
6. Nanoparticles are much smaller than ordinary bulk materials.

9.36C — Describe how the properties of nanoparticulate materials are related to their uses

1. Their particles are very small, giving a large surface area relative to volume.
2. It can increase reactivity and alter optical or other physical properties.
3. They can have properties different from the same material in bulk form.
4. They can provide effective UV protection while remaining relatively transparent.
5. Smaller particles give a larger surface area to volume ratio.
6. Their properties determine suitable applications.

9.37C — Explain the possible risks associated with some nanoparticulate materials

1. Their very small size may allow them to enter cells or tissues.
2. Their high surface area and small size can produce different chemical and biological behaviour.
3. Through inhalation, ingestion or through damaged skin.
4. More surface area is available for chemical interactions.
5. Some nanoparticles have not been studied over sufficiently long timescales.
6. Potential health and environmental effects must be assessed before widespread use.

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

1. Strength, density, hardness, melting point, electrical conductivity, thermal conductivity and flexibility.
2. Compare numerical or measured values for the relevant properties.
3. Their different structures and bonding give different physical properties.

4. Its hardness, brittleness, transparency and chemical resistance affect its uses.
5. Their hardness, heat resistance and brittleness affect their uses.
6. They combine useful properties from different materials.

9.39C — Explain why the properties of a material make it suitable for a given use

1. The material must have properties that meet the requirements of the application.
2. Compare the supplied data with the properties required for the application.
3. High electrical conductivity, suitable strength and low resistance.
4. Low density combined with high strength.
5. The overall combination may better satisfy the application's requirements.
6. Identify the required properties, compare the data and justify the material that best meets the requirements.