GCSE (9-1)

Separate Chemistry 1

### Topics common to Paper 3 and Paper 4

#### Formulae, equations and hazards

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| **Students should:** | **Maths skills** |
| 0.1 | Recall the formulae of elements, simple compounds and ions |  |
| 0.2 | Write word equations |  |
| 0.3 | Write balanced chemical equations, including the use of the state symbols (s), (l), (g) and (aq) | 1c |
| 0.4 | **Write balanced ionic equations** | 1c |
| 0.5 | Describe the use of hazard symbols on containers:a to indicate the dangers associated with the contents b to inform people about safe-working precautions withthese substances in the laboratory |  |
| 0.6 | Evaluate the risks in a practical procedure and suggest suitable precautions for a range of practicals including those mentioned in the specification |  |

#### Topic 1 – Key concepts in chemistry Atomic structure

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| **Students should: Maths skills** |
| * 1. Describe how the Dalton model of an atom has changed over time because of the discovery of subatomic particles
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| 1.2 Describe the structure of an atom as a nucleus containing protons and neutrons, surrounded by electrons in shells |  |
| * 1. Recall the relative charge and relative mass of: a a proton. *Mass = 1 Charge = +1*
		1. a neutron. *Mass = 1 Charge = 0*
		2. an electron. *Mass = 1/1835 Charge = -1*
 |  |
| * 1. Explain why atoms contain equal numbers of protons and electrons
* *Atoms are neutral so charges cancel each other out.*
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| * 1. Describe the nucleus of an atom as very small compared to the overall size of the atom
* *Atom = 1 x 10-10*
* *Nucleus = 1 x 10-15 (this is smaller)*
 | 1d |
| * 1. Recall that most of the mass of an atom is concentrated in the nucleus
* *Total number of protons and neutrons*
 |  |
| 1.7 Recall the meaning of the term mass number of an atom |  |
| 1.8 Describe atoms of a given element as having the same number of protons in the nucleus and that this number is unique to that element |  |
| 1.9 Describe isotopes as different atoms of the same element containing the same number of protons but different numbers of neutrons in their nuclei - *Same atomic number, different mass number.*  |  |
| 1.10 Calculate the numbers of protons, neutrons and electrons in atoms given the atomic number and mass number | 3b |
| 1.11 Explain how the existence of isotopes results in relative atomic masses of some elements not being whole numbers*- If an element has more than one isotope, its Ar is the average of the mass numbers of all the isotopes.* | 1a, 1c |
| 1.12 **Calculate the relative atomic mass of an element from the relative masses and abundances of its isotopes** | 1a, 1c3a, 3c |

#### The periodic table

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| **Students should: Maths skills** |
| 1.13 Describe how Mendeleev arranged the elements, known at that time, in a periodic table by using properties of these elements and their compounds - *Sorted into groups based on their properties. If he put them in order of atomic mass, a pattern appeared. He could put elements with similar chemical properties in columns* | -  |
| 1.14 Describe how Mendeleev used his table to predict the existence and properties of some elements not then discovered- *To keep elements with similar chemical properties together, he left gaps – he used the properties of the other elements in the columns to predict the properties of undiscovered elements.* |  |
| 1.15 Explain that 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*- the atomic mass used was wrong due to the presence of isotopes* |  |
| 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*- order of ascending atomic no.* |  |
| * 1. Describe that in the periodic table
		1. elements are arranged in order of increasing atomic number, in rows called periods
		2. elements with similar properties are placed in the same vertical columns called groups
 |  |
| 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 |  |
| 1.19 Predict the electronic configurations of the first 20 elements in the periodic table as diagrams and in the form, for example 2.8.1 | 4a 5b |
| 1.20 Explain how the electronic configuration of an element is related to its position in the periodic table- *atomic = no. electrons**- Number of shells = the same as the period.**- Group no. = how many electrons in outer shell* | 4a |

#### Ionic bonding

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| **Students should: Maths skills** |
| 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- *Ensure you can draw sodium chloride, magnesium oxide, magnesium chloride and sodium oxide* | 5b |
| 1.22 Recall that an ion is an atom or group of atoms with a positive or negative charge |  |
| 1.23 Calculate the numbers of protons, neutrons and electrons in simple ions given the atomic number and mass number | 3b |
| 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 | 1c 5b |
| 1.25 Explain the use of the endings –ide and –ate in the names of compounds*- ide = negative ions containing only one element (apart from hydroxide OH-)**- ate = ions containing oxygen and at least one other element* |  |
| **Students should: Maths skills** |
| 1.26 Deduce the formulae of ionic compounds (including oxides (O2-), hydroxides (OH-), halides (F-, Cl-, Br-), nitrates (NO3-), carbonates (CO32-) and sulfates (SO42-) ) given the formulae of the constituent ions | 1c |
| 1.27 Explain the structure of an ionic compound as a lattice structure a consisting of a regular arrangement of ionsb held together by strong electrostatic forces (ionic bonds) between oppositely-charged ions | 5b |

#### Covalent bonding

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| **Students should: Maths skills** |
| 1.28 Explain how a covalent bond is formed when a pair of electrons is shared between two atoms |  |
| 1.29 Recall that covalent bonding results in the formation of molecules *– 1 x 10-10m* |  |
| 1.30 Recall the typical size (order of magnitude) of atoms and small molecules *– 1 x 10-9m* | 1d |
| * 1. Explain the formation of simple molecular, covalent substances, using dot and cross diagrams, including:
		1. hydrogen
		2. hydrogen chloride c water

d methane e oxygenf carbon dioxide | 5b |

#### Types of substance

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| **Students should: Maths skills** |
| 1.32 Explain why elements and compounds can be classified as: a ionic b simple molecular (covalent) c giant covalentd metallicand 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)* *Ionic = metal and non-metal. Regular lattice structure. High M.P and B.P because of strong electrostatic forces of attraction. Don’t conduct when solid. Do conduct when molten or in solution.*
* *Simple molecular = contains a few atoms. Low MP and BP. Weak intermolecular forces. Don’t conduct. Some soluble.*
* *Giant covalent = strong covalent bonds. Diamond and graphite. High MP and BP. Do not conduct except graphite and graphene. Not soluble.*
* *Metallic = giant structures. Strong electrostatic forces. Conduct because of delocalized electrons. High MP and BP. Not soluble.*
 |  |
| * 1. Explain the properties of ionic compounds limited to:
		1. high melting points and boiling points, in terms of forces between ions
		2. whether or not they conduct electricity as solids, when molten and in aqueous solution
 | 4a |
| * 1. Explain the properties of typical covalent, simple molecular compounds limited to:
		1. low melting points and boiling points, in terms of forces between molecules (intermolecular forces)
		2. poor conduction of electricity
 | 4a |
| 1.35 Recall that graphite and diamond are different forms of carbon and that they are examples of giant covalent substances |  |
| 1.36 Describe the structures of graphite and diamond *- Graphite = 3 covalent bonds. Layers. Weak IM force so layers slide. Can conduct.**- Diamond = 4 covalent bonds. High MP. Hard* | 5b |
| 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 | 5b |
| 1.38 Explain the properties of fullerenes including C60 and graphene in terms of their structures and bonding- *C60 = Sphere of carbon atoms. Used for drug delivery. Huge surface area.* *- Graphene = One layer of graphite. Nanotube are tiny cylinder of graphene. They are strong and lightweight.*  | 5b |
| 1.39 Describe, using poly(ethene) as the example, that simple polymers consist of large molecules containing chains of carbon atoms | 5b |
| 1.40 Explain the properties of metals, including malleability and the ability to conduct electricity | 5b |
| 1.41 Describe the limitations of particular representations and models to, include dot and cross, ball and stick models and two- and three-dimensional representationsSee page bottom of page 22  | 5b |
| 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 |  |

#### Calculations involving masses

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| **Students should: Maths skills** |
| 1.43 Calculate relative formula mass given relative atomic masses*- Mr = add together Ar* | 1a, 1c |
| 1.44 Calculate the formulae of simple compounds from reacting masses or percentage composition and understand that these are empirical formulae*- the smallest whole number ration of atoms in a compound* | 1a, 1c 2a |
| * 1. Deduce:
		1. the empirical formula of a compound from the formula of its molecule - *Mass/Ar*
		2. the molecular formula of a compound from its empirical formula and its relative molecular mass
* *Divide the Mr of the compound by the Mr of the empirical formula. Multiply everything in the empirical formula by the result*
 | 1c |
| * 1. Describe an experiment to determine the empirical formula of a simple compound such as magnesium oxide
* *See your revision guide for more info*
 | 1a, 1c2a |
| * 1. Explain the law of conservation of mass applied to:
		1. a closed system including a precipitation reaction in a closed flask – *mass does not change.*
		2. a non-enclosed system including a reaction in an open flask that takes in or gives out a gas *– mass changes*
 | 1a |
| 1.48 Calculate masses of reactants and products from balanced equations, given the mass of one substance | 1a, 1c 2a |
| 1.49 Calculate the concentration of solutions in g dm–3- conc = mass / vol | 1a, 1c 2a 3b, 3c |
| * 1. Recall that one mole of particles of a substance is defined as:
		1. **the Avogadro constant number of particles**

**(6.02 × 1023 atoms, molecules, formulae or ions) of that substance*** + 1. **a mass of ‘relative particle mass’ g –** *Ar or Mr in grams*
 | 1b |
| * 1. **Calculate the number of:**
		1. **moles of particles of a substance in a given mass of that substance and vice versa –** *moles = mass/Ar*
		2. **particles of a substance in a given number of moles of that substance and vice versa** *– no. particles = moles x Avogadro.*
		3. **particles of a substance in a given mass of that substance and vice versa**
 | 1a, 1b, 1c 3a, 3b, 3c |

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| **Students should: Maths skills** |
| * 1. **Explain why, in a reaction, the mass of product formed is controlled by the mass of the reactant which is not in excess**

*-The reactant used up is called the limiting reactant. It limits the amount of product formed* | 1c |
| * 1. **Deduce the stoichiometry of a reaction from the masses of the reactants and products**
* *Calculate no. moles = mass / Ar*
* *Divide by smaller number*
* *Simplest whole number ratio*
 | 1a, 1c |

### Topics for Paper 3

#### Topic 2 – States of matter and mixtures States of matter

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| **Students should:** | **Maths skills** |
| 2.1 | Describe the arrangement, movement and the relative energy of particles in each of the three states of matter: solid, liquid and gas | 5b |
| 2.2 | Recall the names used for the interconversions *(changes)* between the three states of matter, recognising that these are physical changes: contrasted with chemical reactions that result in chemical changes* *Melting (S to L), Evaporating (L to G), Condensing (G to L), Freezing (L to S), sublimation (S to G) and Deposition (G to S)*
 |  |
| 2.3 | Explain the changes in arrangement, movement and energy of particles during these interconversions | 5b |
| 2.4 | Predict the physical state of a substance under specified conditions, given suitable data | 1d4a |

#### Methods of separating and purifying substances

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| **Students should:** | **Maths skills** |
| 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* *Pure = contain only one thing (single element/compound)*
 |  |
| 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 | 1a |
| 2.7 | Explain the types of mixtures that can be separated by using the following experimental techniques:1. simple distillation
2. fractional distillation c filtration
3. crystallisation
4. paper chromatography

*a = simple distillation = separates out solutions because of a difference in boiling point**b = separate a mixture of liquids.**c = separate insoluble solid from liquid**d = separates soluble solid from a solution**e = separates a mixture of soluble substances and identifies them. Mobile phase = solvent. Stationary phase = filter paper.* |  |
| 2.8 | Describe an appropriate experimental technique to separate a mixture, knowing the properties of the components of the mixture |  |

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| **Students should: Maths skills** |
| 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 |  |
| 2.10 Interpret a paper chromatogram:a to distinguish between pure and impure substances b to identify substances by comparison with knownsubstancesc to identify substances by calculation and use of Rf values*a = pure = one blob**b = same Rf value = likely to be the same**c =Rf value = distance travelled by solute* *distance travelled by solvent* | 3a, 3c 4a |
| 2.11 *Core Practical: Investigate the composition of inks using simple distillation and paper chromatography* |  |
| * 1. Describe how:
		1. waste and ground water can be made potable, including the need for sedimentation, filtration and chlorination
		2. sea water can be made potable by using distillation
		3. water used in analysis must not contain any dissolved salts – *can interfere with reactions and give false results.*
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#### Topic 3 – Chemical change Acids

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| **Students should: Maths skills** |
| 3.1 Recall that acids in solution are sources of hydrogen ions and alkalis in solution are sources of hydroxide ions |  |
| 3.2 Recall that a neutral solution has a pH of 7 and that acidic solution-s have lower *(0-6)* pH values and alkaline solutions higher *(8-14)* pH values |  |
| 3.3 Recall the effect of acids and alkalis on indicators, including litmus, methyl orange and phenolphthalein- Litmus = Red (acidic), Purple (neutral), blue (alkaline)- Methyl orange = Red (acidic), Yellow (neutral), yellow (alkaline)- Phenolphthalein = colourless (acidic), Alkaline + neutral (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** | 1c |
| 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** | 1c |
| 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* | 4a, 4c |
| 3.7 **Explain the terms dilute and concentrated, with respect to amount of substances in solution***- Dilute = small number of acid molecules compared to volume of water**- Concentrated = Large number of acid molecules compared to volume of water* |  |

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| **Students should: Maths skills** |
| 3.8 **Explain the terms weak and strong acids, with respect to the degree of dissociation into ions**- Weak = *do not fully ionize in solution* |  |
| 3.9 Recall that a base is any substance that reacts with an acid to form a salt and water only |  |
| 3.10 Recall that alkalis are soluble bases |  |
| * 1. Explain the general reactions of aqueous solutions of acids with: a metals
		1. metal oxides
		2. metal hydroxides d metal carbonates

to produce salts1. *acid + metal 🡪 salt + hydrogen*
2. *acid + metal oxide 🡪 salt + water*
3. *acid + metal hydroxide 🡪 salt + water*
4. *acid + metal carbonate 🡪 salt + water + carbon dioxide*
 |  |
| 3.12 Describe the chemical test for: a hydrogenb carbon dioxide (using limewater)*a = hydrogen = lighted splint makes a squeaky pop**b = carbon dioxide = turns limewater cloudy.* |  |
| 3.13 Describe a neutralisation reaction as a reaction between an acid and a 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*H+(aq+ +OH-(aq) 🡪 H2O(l)* |  |

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| **Students should: Maths skills** |
| 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 removedc the solution remaining is only salt and water*a – to ensure all the acid has been neutralized**b- to get solution containing only the salt + water* |  |
| * 1. Explain why, if soluble salts are prepared from an acid and a soluble reactant:
		1. titration must be used – *to work out exactly the right amount of alkali to neutralize the acid*
		2. the acid and the soluble reactant are then mixed in the correct proportions
		3. the solution remaining, after reaction, is only salt and water
 |  |
| 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* |  |
| 3.18 Describe how to carry out an acid-alkali titration, using burette, pipette and a suitable indicator, to prepare a pure, dry salt*See your revision guide for more info* |  |
| * 1. Recall the general rules which describe the solubility of common types of substances in water:
		1. all common sodium, potassium and ammonium salts are soluble
		2. all nitrates are soluble
		3. common chlorides are soluble except those of silver and lead
		4. common sulfates are soluble except those of lead, barium and calcium
		5. common carbonates and hydroxides are insoluble except those of sodium, potassium and ammonium
 |  |
| 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 |  |
| 3.21 Describe the method used to prepare a pure, dry sample of an insoluble salt – Page 46 (separate science) page 108 (combined H tier) OR page 109 (combined F tier) |  |

#### Electrolytic processes

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| **Students should: Maths skills** |
| 3.22 Recall that electrolytes are ionic compounds in the molten state or dissolved in water |  |
| 3.23 Describe electrolysis as a process in which electrical energy, from a direct current supply, decomposes electrolytes |  |
| * 1. Explain the movement of ions during electrolysis, in which:
		1. positively charged cations migrate to the negatively charged cathode
		2. negatively charged anions migrate to the positively charged anode
 |  |
| 3.25 Explain the formation of the products in the electrolysis, using inert electrodes, of some electrolytes, including:a copper chloride solution b sodium chloride solution c sodium sulfate solution1. water acidified with sulfuric acid
2. molten lead bromide (demonstration)
3. **Cu2+ Cl-** H+ OH- (H2O left)
4. Na+ **Cl- H+** OH- (NaOH left)
5. Na+ SO42- **H+** **OH-** (Na2SO4 left)
6. **H+ OH-** SO42- (SO4 left in solution)
 |  |
| 3.26 Predict the products of electrolysis of other binary, ionic compounds in the molten state |  |
| 3.27 **Write half equations for reactions occurring at the anode and cathode in electrolysis** | 1c |
| 3.28 **Explain oxidation and reduction in terms of loss or gain of electrons –** *Oxidation is a loss of electrons. Reduction is a gain.*  |  |
| 3.29 **Recall that reduction occurs at the cathode and that oxidation occurs at the anode in electrolysis reactions** |  |
| 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*- Anode: Cu 🡪 Cu2+ + 2e-**- Cathode: Cu2+ + 2e- 🡪 Cu* |  |
| 3.31 *Core Practical: Investigate the electrolysis of copper sulfate solution with inert electrodes and copper electrodes* | 1a4a, 4b, 4c, 4d |

#### Topic 4 – Extracting metals and equilibria Obtaining and using metals

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| **Students should: Maths skills** |
| 4.1 Deduce the relative reactivity of some metals, by their reactions with water, acids and salt solutions |  |
| 4.2 **Explain displacement reactions as redox reactions, in terms of gain or loss of electrons** |  |
| 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. Recall that:
		1. most metals are extracted from ores found in the Earth’s crust
		2. unreactive metals are found in the Earth’s crust as the uncombined elements
 |  |
| 4.5 Explain oxidation as the gain of oxygen and reduction as the loss of oxygen |  |
| 4.6 Recall that the extraction of metals involves reduction of ores |  |
| 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 bya heating with carbon (including iron) b electrolysis (including aluminium)(knowledge of the blast furnace is not required) |  |
| 4.8 **Evaluate alternative biological methods of metal extraction (bacterial and phytoextraction)** |  |
| 4.9 Explain how a metal’s relative resistance to oxidation is related to its position in the reactivity series |  |
| 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 |  |
| 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 |  |
| 4.12 Evaluate data from a life cycle assessment of a product |  |

#### Reversible reactions and equilibria

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| **Students should: Maths skills** |
| 4.13 Recall that chemical reactions are reversible, the use of thesymbol ⇌ in equations and that the direction of some reversible reactions can be altered by changing the reaction conditions |  |
| 4.14 Explain what is meant by dynamic equilibrium- *the forward and backward reactions are both happening at the same time and at the same rate.* |  |
| 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. Recall the conditions for the Haber process as: a temperature 450 °C
		1. pressure 200 atmospheres
		2. iron catalyst
 |  |
| 4.17 **Predict how the position of a dynamic equilibrium is affected by changes in:****a temperature b pressure****c concentration***a Increase temp = equilibrium will shift in endothermic direction**Decrease temp =Equilibrium will shift in exothermic direction**B Increase pressure = Equilibrium will move to side with fewer molecules**Decrease pressure = Equilibrium will move to side with most moles to increase pressure* *C Increase concentration of reactants = Shift to products**Increase concentration of products = shift to reactants.* |  |

**Topic 5 – Separate chemistry 1 Transition metals, alloys and corrosion**

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| **Students should: Maths skills** |
| 5.1C Recall that most metals are transition metals and that their typical properties include:a high melting point b high density1. the formation of coloured compounds
2. catalytic activity of the metals and their compounds as exemplified by iron – *Fe2+ is light green. Fe3+ orange brown*
 |  |
| 5.2C Recall that the oxidation of metals results in corrosion |  |
| * 1. C Explain how rusting of iron can be prevented by: a exclusion of oxygen
		1. exclusion of water
		2. sacrificial protection *– placing a more reactive metal with the iron*
 |  |
| 5.4C Explain how electroplating can be used to improve the appearance and/or the resistance to corrosion of metal objects*- Applying a metal coating to an object. The cathode is the thing you want to coat* |  |
| 5.5C Explain, using models, why converting pure metals into alloys often increases the strength of the product | 5b |
| 5.6C Explain why iron is alloyed with other metals to produce alloy steels – *Harder/ steel less likely to rust* |  |
| 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 *–see green box on page 63* |  |

**Quantitative analysis**

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| **Students should: Maths skills** |
| 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***- concentration (mol dm-3) = no of moles / volume (dm-3)**- convert mol dm-3 into g dm-3 multiply the concentration by the Mr**- convert from g dm-3 into mol dm-3 divide by the Mr* | 1a, 1b, 1c, 1d 2a3b, 3c |
| 5.9C *Core Practical: Carry out an accurate acid-alkali titration, using burette, pipette and a suitable indicator**- Use phenolphthalein or methyl orange indicator**- get concordant results* |  |
| 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****- *see your revision guide for more info*** | 1a, 1c, 1d2a, 2b 3a, 3b, 3c |
| 5.11C Calculate the percentage yield of a reaction from the actual yield and the theoretical yield- *actual / theoretical x 100* | 1a, 1c, 1d 2a3b, 3c |
| * 1. C Describe that the actual yield of a reaction is usually less than the theoretical yield and that the causes of this include:
		1. incomplete reactions
		2. practical losses during the experiment
		3. competing, unwanted reactions (side reactions)
 |  |
| 5.13C Recall the atom economy of a reaction forming a desired product-  *the % of reactants changed to useful products* |  |
| 5.14C Calculate the atom economy of a reaction forming a desired product-  *total Mr of desired products x 100* *total Mr of all products* | 1a, 1c, 1d 2a3c |
| 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** |  |
| 5.16C **Describe the molar volume, of any gas at room temperature and pressure, as the volume occupied by one mole of molecules of any gas at room temperature and pressure****(The molar volume will be provided as 24 dm3 or 24000 cm3 in calculations where it is required)*** *Moles = gas volume / molar volume*
* *Moles =gas volume / 24*
 |  |
| 5.17C **Use the molar volume and balanced equations in calculations involving the masses of solids and volumes of gases** | 1a, 1c, 2a3b, 3c |
| 5.18C **Use Avogadro’s law to calculate volumes of gases involved in a gaseous reaction, given the relevant equation** | 1a, 1c, 1d |

**Dynamic equilibria**

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| **Students should: Maths skills** |
| 5.19C Describe the Haber process as a reversible reaction between nitrogen and hydrogen to form ammonia |  |
| 5.20C **Predict how the rate of attainment of equilibrium is affected by:****a changes in temperature b changes in pressure**1. **changes in concentration**
2. **use of a catalyst –** *makes reaction faster. Reaches equilibrium faster (doesn’t affect position of equilibrium) see page 68 for more detail*
 |  |
| * 1. C **Explain how, in industrial reactions, including the Haber process, conditions used are related to:**
		1. **the availability and cost of raw materials and energy supplies**
		2. **the control of temperature, pressure and catalyst used produce an acceptable yield in an acceptable time**
 |  |
| 5.22C Recall that fertilisers may contain nitrogen, phosphorus and potassium compounds to promote plant growth |  |
| 5.23C Describe how ammonia reacts with nitric acid to produce a salt that is used as a fertilizer NH3 + HNO3 🡪 NH4NO3 |  |
| * 1. C Describe and compare:
		1. the laboratory preparation of ammonium sulfate from ammonia solution and dilute sulfuric acid on a small scale
		2. the industrial production of ammonium sulfate, used as a fertiliser, in which several stages are required to produce ammonia and sulfuric acid from their raw materials and the production is carried out on a much larger scale (details of the industrial production of sulfuric acid are not required)
* *see your revision guide for more info*
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**Chemical cells and fuel cells**

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| **Students should: Maths skills** |
| 5.25C Recall that a chemical cell produces a voltage until one of the reactants is used up |  |
| 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 |  |
| 5.27C Evaluate the strengths and weaknesses of fuel cells for given uses- *see your revision guide for more info* |  |