

A-Level Chemistry (AQA)

Exam Dates 2018:

Chemistry A-level 7405					
7405/1	Inorganic and Physical Chemistry		2h	05 June 2018	pm
7405/2	Organic and Physical Chemistry		2h	12 June 2018	pm
7405/3	Paper 3		2h	20 June 2018	am

Note: the practical endorsement is reported as a pass or fail separately

Exam Structure:

Assessments

Paper 1	+	Paper 2	+	Paper 3
What's assessed		What's assessed		What's assessed
<ul style="list-style-type: none"> Relevant Physical chemistry topics (sections 3.1.1 to 3.1.4, 3.1.6 to 3.1.8 and 3.1.10 to 3.1.12) Inorganic chemistry (Section 3.2) Relevant practical skills 		<ul style="list-style-type: none"> Relevant Physical chemistry topics (sections 3.1.2 to 3.1.6 and 3.1.9) Organic chemistry (Section 3.3) Relevant practical skills 		<ul style="list-style-type: none"> Any content Any practical skills
How it's assessed		How it's assessed		How it's assessed
<ul style="list-style-type: none"> written exam: 2 hours 105 marks 35% of A-level 		<ul style="list-style-type: none"> written exam: 2 hours 105 marks 35% of A-level 		<ul style="list-style-type: none"> written exam: 2 hours 90 marks 30% of A-level
Questions		Questions		Questions
105 marks of short and long answer questions		105 marks of short and long answer questions		40 marks of questions on practical techniques and data analysis 20 marks of questions testing across the specification 30 marks of multiple choice questions

A-Level Chemistry AQA – Summary of Topics

Year 1 / AS Chemistry		
Topic	Syllabus Reference	Textbook Reference
Topic 1: Atomic Structure	3.1.1 and 3.2.1	1.1-1.6 and 8.1-8.4
Topic 2: Amount of Substance	3.1.2	2.1-2.6
Topic 3: Bonding	3.1.3	3.1-3.7
Topic 4: Energetics	3.1.4	4.1-4.7
Topic 5: How Far? How Fast? I	3.1.5 and 3.1.6	5.1-5.3 and 6.1-6.6
Topic 6: Redox and Group Chemistry	3.1.7, 3.2.2 and 3.2.3	7.1-7.3, 9.1 and 10.1-10.4
Topic 7: Intro to Organic Chemistry	3.3.1 and 3.3.2	11.1-11.3 and 12.1-12.5 (part of 14.1)
Topic 8: Reactions of Organic Compounds	3.3.3, 3.3.4 and 3.3.5	13.1-13.3, 14.1-14.3 and 15.1-5.3
Topic 9: Analysis of Organic Compounds	3.3.6	16.1-16.3
Year 2 A-Level Chemistry		
Topic	Syllabus Reference	Textbook Reference
Topic 10: Thermodynamics	3.1.8	17.1 – 17.4
Topic 11: How Far? How Fast? II	3.1.9 and 3.1.10	18.1 – 18.4 and 19.1
Topic 12: Acids, Bases and Buffers	3.1.12	21.1 – 21.6
Topic 13: Electrochemistry	3.1.11	20.1 – 20.3
Topic 14: Reactions of Period 3 elements and oxides	3.2.4	22.1-22.3
Topic 15: Transition metals and complex ions	3.2.5 and 3.2.6	23.1-23.5 and 24.1-24.3
Topic 16: Carbonyl Compounds and Optical Isomerism	3.3.7, 3.3.8 and 3.3.9	25.1 – 26.5
Topic 17: Amines and Polymers	3.3.11 and 3.3.12	28.1 – 29.1
Topic 18: Aromatic Chemistry	3.3.10	27.1 – 27.3
Topic 19: Amino Acids, Proteins and DNA	3.3.13	30.1-30.5
Topic 20: Organic Synthesis and Analysis	3.3.14-3.3.16	31.1-33.1

Inorganic Chemistry Year 1

Topic 2.1: Periodicity

2.1.1 Classification	An element is classified as s, p, d or f block according to its position in the Periodic Table, which is determined by its proton number	
2.1.2 Physical properties of Period 3 elements	Describe the trends in atomic radius, first ionisation energy and melting point of the elements Na-Ar	
	Explain these trends in terms of the structure and bonding of the elements	

Topic 2.2: Group 2, the Alkaline Earth Metals

2.2 Group 2; the alkaline earth metals	Describe and explain the trends in atomic radius, first ionisation energy and melting point of the Group 2 elements (Mg – Ba)	
	Describe the reactions of Group 2 elements with water	
	Describe the use of magnesium in the extraction of titanium from TiCl_4	
	Describe the relative solubilities of the hydroxides of the elements Mg – Ba in water	
	Know that magnesium hydroxide ($\text{Mg}(\text{OH})_2$) is sparingly soluble	
	Describe the use of $\text{Mg}(\text{OH})_2$ in medicine and $\text{Ca}(\text{OH})_2$ in agriculture	
	Describe the use of CaO or CaCO_3 to remove acidic SO_2 from flue gases	
	Describe the relative solubilities of the sulfates of the elements Mg–Ba in water	
	Know that barium sulfate (BaSO_4) is insoluble	
	Describe the use of acidified BaCl_2 solution to test for sulfate ions	
	Explain why the BaCl_2 solution used in this test must be acidified	
Describe the use of BaSO_4 in medicine		

Topic 2.3: Group 7 (17), the Halogens

2.3.1 Trends in properties	Describe and explain the trend in electronegativity of the halogens	
	Describe and explain the trend in boiling point of the halogens	
	Describe the trend in oxidising ability of the halogens down the group, including displacement reactions of halide ions in aqueous solution	
	Describe the trend in reducing ability of the halide ions, including the reactions of solid sodium halides with concentrated sulfuric acid	
	The use of acidified silver nitrate solution to identify and distinguish between halide ions	
	Describe the trend in solubility of the silver halides in ammonia	
	Describe how silver nitrate solution is used to identify halide ions, why it must be acidified and why ammonia solution is added	
2.3.2 Uses of chlorine and chlorate(I)	Describe the reaction of chlorine with water to form chloride ions and chlorate(I) ions	
	Describe the reaction of chlorine with water to form chloride ions and oxygen	
	Discuss the advantages and disadvantages of adding chemicals to water supplies	
	Describe the use of chlorine in water treatment	
	Discuss how the benefits to health of water treatment by chlorine outweigh its toxic effects	
	Describe the reaction of chlorine with cold, dilute, aqueous NaOH and the uses of the solution formed	
Required practical 4: Carry out simple test-tube reactions to identify cations and anions		

Organic Chemistry Year 1

Topic 3.1: Introduction to Organic Chemistry

3.1.1 Nomenclature	Organic compounds can be represented by: empirical formula; molecular formula; general formula; structural formula; displayed formula; skeletal formula	
	The characteristics of a homologous series are a series of compounds containing the same functional group	
	Use IUPAC rules to assign names for organic compounds, chains and rings, with up to six carbon atoms each	
	Draw structural, displayed and skeletal formulas for compounds given their IUPAC names, for chains and rings with up to six carbon atoms each	
3.1.2 Reaction mechanisms	Reactions of organic compounds can be explained using mechanisms	
	Know that unpaired electrons in radicals are represented by dots	
	Write balanced equations for each step in the free-radical substitution mechanism	
	Know that the formation of a covalent bond is shown by a curly arrow that starts from a lone pair of electrons or from another covalent bond	
	Know that the breaking of a covalent bond is shown by a curly arrow starting from the bond	
	Outline mechanisms by drawing the structures of the species involved and curly arrows to represent the movement of electron pairs	
3.1.3 Isomerism	Define the terms structural isomer and stereoisomer	
	Know that E-Z isomerism is a form of stereoisomerism and occurs as a result of restricted rotation about the planar carbon-carbon double bond in an alkene	
	Use the Cahn-Ingold-Prelog (CIP) priority rules to identify stereoisomers	
	Draw the structural formulas of E and Z isomers	
	Draw the structures of chain, position and functional group isomers	

Topic 3.2: Alkanes

3.2.1 Fractional distillation of crude oil	Know that alkanes are saturated hydrocarbons	
	Know that petroleum is a mixture consisting mainly of alkane hydrocarbons that can be separated by fractional distillation	
3.2.2 Modification of alkanes by cracking	Know that cracking involves breaking the C-C bonds in alkanes	
	Thermal cracking takes place at high pressure and high temperature and produces a high percentage of alkenes	
	Catalytic cracking takes place at a slight pressure, high temperature and in the presence of a zeolite catalyst and is used mainly to produce motor fuels and aromatic hydrocarbons	
	Understand the economic reasons for cracking alkanes	
3.2.3 Combustion of alkanes	Know that alkanes are used as fuels	
	Combustion of alkanes and other organic compounds can be complete or incomplete	
	The internal combustion engine produces a number of pollutants including NO _x , CO, carbon particles and unburned hydrocarbons	
	The gaseous pollutants from internal combustion engines can be removed using catalytic converters	
	Combustion of hydrocarbons containing sulfur leads to sulfur dioxide, which causes air pollution in the form of acid rain	

	Sulfur dioxide can be removed from flue gases using calcium oxide or calcium carbonate	
3.2.4 Chlorination of alkanes	Describe the reaction of methane with chlorine	
	Describe the mechanism steps of free-radical substitution (initiation, propagation, termination) for this reaction	

Topic 3.3: Halogenoalkanes

3.3.1 Nucleophilic substitution	Halogenoalkanes contain polar bonds	
	Halogenoalkanes undergo substitution reactions with the nucleophiles OH ⁻ , CN ⁻ and NH ₃	
	Outline the nucleophilic substitution mechanisms of these reactions	
	Explain why the carbon-halogen bond enthalpy influences the rate of reaction	
3.3.2 Elimination	Describe the concurrent substitution and elimination reactions of a halogenoalkane (e.g. 2-bromopropane with potassium hydroxide)	
	Explain the role of the reagent as both nucleophile and base	
	Outline the mechanism of these reactions	
3.3.3 Ozone depletion	Ozone, formed naturally in the upper atmosphere, is beneficial because it absorbs ultraviolet radiation	
	Chlorine atoms are formed in the upper atmosphere when ultraviolet radiation causes C-Cl bonds in CFCs to break	
	Chlorine atoms catalyse the decomposition of ozone and contribute to the hole in the ozone layer	
	Results of research by different groups in the scientific community has provided evidence for legislation to ban the use of CFCs as solvents and refrigerants. Chemists have now developed alternative, chlorine-free compounds.	
	Write balanced equations to explain how chlorine atoms (radicals) catalyse the decomposition of ozone	

Topic 3.4: Alkenes

3.4.1 Structure, bonding and reactivity	Know that alkenes are unsaturated hydrocarbons	
	Describe the bonding in alkenes as involving a double covalent bond, which is a centre of high electron density	
3.4.2 Addition reactions of alkenes	Describe the electrophilic addition reactions of alkenes with HBr, H ₂ SO ₄ and Br ₂	
	Describe the use of the bromine water test for unsaturation	
	Explain the formation of major and minor products in addition reactions of unsymmetrical alkenes by reference to the relative stabilities of the carbocation intermediates	
	Outline the mechanism of addition reactions to alkenes	
3.4.3 Addition polymers	Addition polymers are formed from alkenes and substituted alkenes	
	Recognise the repeating unit of addition polymers	
	Use IUPAC rules for naming addition polymers	
	Explain why addition polymers are unreactive	
	Explain the nature of the intermolecular forces between molecules of polyalkenes	
	Appreciate that knowledge and understanding of the production and properties of polymers has developed over time	
	Describe typical uses of poly(chloroethene) (PVC) and how its properties can be modified with a plasticiser	
	Draw the repeating unit from a monomer structure or from a section of the polymer chain	
Draw the structure of the monomer from a section of the polymer		

Topic 3.5: Alcohols

3.5.1 Alcohol production	Describe how alcohols are produced industrially by hydration of alkenes in the presence of an acid catalyst	
	Describe how ethanol is produced industrially by fermentation of glucose	
	Justify the conditions used in the production of ethanol by fermentation of glucose	
	Know that the ethanol produced industrially by fermentation is separated by fractional distillation and can then be used as a biofuel	
	Explain the meaning of the term biofuel	
	Write equations to support the statement that ethanol produced by fermentation is a carbon-neutral fuel and give reasons why this statement is not valid	
	Outline the mechanism for the formation of an alcohol by the reaction of an alkene with steam in the presence of an acid catalyst	
	Discuss the environmental and ethical issues involving biofuel use	
3.5.2 Oxidation of alcohols	Classify alcohols as primary, secondary or tertiary	
	Know that primary alcohols can be oxidised to aldehydes, which can be further oxidised to carboxylic acids	
	Describe how the method used to oxidise a primary alcohol determines whether an aldehyde or a carboxylic acid is obtained	
	Know that secondary alcohols can be oxidised to ketones	
	Know that tertiary alcohols are not easily oxidised	
	Acidified potassium dichromate(VI) is a suitable oxidising agent	
	Write equations for the oxidation reactions, using [O] as the oxidant	
	Use chemical tests to distinguish between aldehydes and ketones, including Fehling's solution and Tollens' reagent	
3.5.3 Elimination	Alkenes can be formed from alcohols by acid-catalysed elimination reactions	
	Alkenes produced by this method can be used to produce addition polymers without using monomers derived from crude oil	
	Outline the mechanism for the elimination of water from alcohols	
Required practical 5: Distillation of a product from a reaction		

Topic 3.6: Organic Analysis

3.6.1 Identification of functional groups	Identify and describe the reactions of the functional groups covered in this module; alcohols, aldehydes, alkenes, carboxylic acids using test-tube reactions	
3.6.2 Mass spectrometry	Mass spectrometry can be used to determine the molecular formula of a compound	
	Use precise atomic masses and the precise molecular mass to determine the molecular formula of a compound	
3.6.3 Infrared spectroscopy	Know that bonds in a molecule absorb infrared radiation at characteristic wavenumbers	
	'Fingerprinting' allows identification of a molecule by comparison of spectra	
	Use data provided to identify particular bonds, functional groups and impurities present in infrared spectra	
	Describe the link between absorption of infrared radiation by bonds in CO ₂ , methane and water vapour to global warming	
Required practical 6: Tests for alcohol, aldehyde, alkene and carboxylic acid groups		

Physical Chemistry Year 1

Topic 1.1: Atomic Structure

1.1.1 Fundamental Particles	How models of atomic structure have changed over time	
	The relative charges, relative masses and location of: protons, neutrons and electrons in an atom	
1.1.2 Mass Number and Isotopes	How to determine mass number and proton number for an element	
	How to determine the number of fundamental particles in atoms and ions using mass number, atomic number and charge data	
	Explain the definition and existence of isotopes	
	Describe the principles of a simple time of flight (TOF) mass spectrometer and how the data is analysed	
	Interpret and analyse mass spectrometry data to give information about relative isotopic mass and relative abundance of isotopes	
	Use mass spectrometry data to identify elements and determine relative molecular mass	
1.1.3 Electron Configuration	Write electron configurations for atoms and ions up to $Z = 36$ in terms of shells and sub-shells (orbitals) s , p and d .	
	Define first ionisation energy	
	Write equations for first and successive ionisation energies	
	Explain how ionisation energy data for Period 3 (Na – Ar) and Group 2 (Be – B) give evidence for electron configuration in shells and sub-shells	

Topic 1.2: Amount of Substance

1.2.1 Relative atomic and molecular mass	How the ^{12}C isotope is used to determine relative molecular mass	
	Definitions of relative atomic mass (A_r) and relative molecular mass (M_r)	
1.2.2 The mole and the Avogadro constant	The use of the Avogadro constant (6.02×10^{23}) as the number of particles in a mole (<i>you do not have to recall the value of the constant</i>)	
	The definition of a mole in relation to electrons, atoms, molecules, ions, formulas and equations	
	How to determine the concentration of a solution in mol dm^{-3}	
	Calculating moles using mass and M_r data	
	Calculating moles using concentration and volume of solution data	
	Calculating moles using waters of crystallisation of hydrated salts data	
1.2.3 The ideal gas equation	Use of the ideal gas equation ($pV = nRT$) and correct SI units	
	Calculating moles using the ideal gas equation	
1.2.4 Empirical and molecular formula	Understand that an empirical formula is the simplest whole number ratio of atoms of each element in a compound	
	Understand that a molecular formula is the actual number of atoms of each element in a compound	
	Understand the relationship between empirical formula and molecular formula	
	Calculate empirical and molecular formulae	
1.2.5 Balanced equations and associated calculations	Write balanced chemical equations (full and ionic) for reactions	
	Calculate the percentage atom economy of reactions	
	Discuss the economic, ethical and environmental implications of developing processes with a high atom economy	
	Use balanced equations to calculate: masses; volumes of gases; percentage yields; concentrations and volumes of solution reactions	
Required Practical 1: Make up a volumetric solution and carry out a simple acid-base titration		

Topic 1.3: Bonding

1.3.1 Ionic Bonding	Understand that ionic bonding involves electrostatic attraction between oppositely charged ions in a lattice	
	Write formulas of compound ions, including: sulfate, hydroxide, nitrate, carbonate and ammonium	
	Predict the charge on a simple ion using the position of the element in the Periodic Table	
	Use the charges on the component ions to write formulae for ionic compounds	
1.3.2 Nature of covalent and dative covalent bonds	Know that a single covalent bond contains a shared pair of electrons	
	Know that multiple bonds contain multiple pairs of electrons	
	Know that a coordinate (dative covalent) bond contains a shared pair of electrons with both electrons supplied by one atom	
	Use single lines to represent covalent bonds	
	Use arrows to represent coordinate bonds	
1.3.3 Metallic Bonding	Understand that metallic bonding involves attraction between delocalised electrons and positive ions arranged in a lattice	
1.3.4 Bonding and physical properties	Describe the structure and properties of the four types of crystal structure: ionic, metallic, macromolecular (giant covalent) and molecular (simple covalent)	
	Describe the structures and properties of these substances: diamond; graphite; ice; iodine; magnesium; sodium chloride	
	Explain how the physical properties of substances (melting point, solubility, conductivity) are related to their structures	
	Explain the energy changes associated with changing state	
	Draw diagrams to represent the particles in the different structure types	
1.3.5 Shapes of simple molecules and ions	Understand that bonding pairs and lone (non-bonding) pairs of electrons are charge clouds which repel each other	
	Pairs of electrons in the outer shell of atoms arrange themselves as far apart as possible to minimise repulsion	
	Lone pair-lone pair repulsion > lone pair-bond pair repulsion > bond pair-bond pair repulsion	
	The number of pairs of electrons and the repulsion between them determines the shape and bond angle within a molecule (VSEPR theory)	
	Deduce the shape and bond angle of molecules and ions with up to six pairs of electrons surrounding the central atom	
1.3.6 Bond polarity	Define electronegativity as the power of an atom to attract the pair of electrons in a covalent bond	
	The electron distribution in a covalent bond between elements with different electronegativities will be unsymmetrical, producing a polar covalent bond and may give the molecule a permanent dipole	
	Use partial charges (δ^+/δ^-) to show that a bond is polar	
	Explain why some molecules with polar bonds do not have a permanent dipole	
1.3.7 Forces between molecules	Describe the three different types of inter-molecular forces: permanent dipole-dipole forces; induced dipole-dipole (van der Waals, dispersion, London) forces; hydrogen bonding	
	Explain how the melting and boiling points of substances are influenced by the strength of these forces	
	The importance of hydrogen bonding in the low density of ice and the anomalous boiling points of compounds	
	Recognise which types of intermolecular forces are present in molecules	

Topic 1.4: Energetics

1.4.1 Enthalpy change	Reactions can be endothermic or exothermic	
	Know that enthalpy change (ΔH) is the heat energy change measured under conditions of constant pressure	
	Define standard conditions	
	Define standard enthalpy of combustion ($\Delta_c H^\ominus$) and standard enthalpy of formation ($\Delta_f H^\ominus$)	
1.4.2 Calorimetry	Know that the heat change (q), in a reaction is given by the equation $q = mc\Delta T$ and understand what each term in this equation represents	
	Use this equation to calculate the molar enthalpy change for a reaction	
	Describe how to find the ΔH of a reaction using calorimetry, including: neutralisation; dissolution of a solid; combustion of a fuel	
1.4.3 Hess's Law	Describe the purpose and application of Hess's law	
	Use Hess's law to perform calculations of enthalpy changes using enthalpy of combustion data and enthalpy of formation data	
1.4.4 Bond enthalpies	Define the term mean bond enthalpy	
	Use mean bond enthalpies to calculate approximate values of ΔH for reactions in the gas phase	
	Explain why values calculated from mean bond enthalpy calculations differ from those determined using Hess's law	
Required Practical 2: Measurement of an enthalpy change using calorimetry		

Topic 1.5: Kinetics

1.5.1 Collision theory	Reactions can only occur when collisions take place between particles having sufficient energy, which is called the activation energy	
	Define the term activation energy	
	Explain why most collisions between particles do not lead to a reaction	
1.5.2 Maxwell-Boltzmann distribution	Describe the Maxwell-Boltzmann distribution of molecular energies in gases	
	Draw and interpret distribution curves for different temperatures of reaction	
1.5.3 Effect of temperature on reaction rate	Define the term rate of reaction	
	Describe different methods for monitoring the rate of a reaction	
	Describe the qualitative effect of temperature changes on the rate of a reaction	
	Use the Maxwell-Boltzmann distribution to explain why a small temperature increase can lead to a large increase in reaction rate	
1.5.4 Effect of concentration and pressure	Describe the qualitative effect of changes in concentration on collision frequency and therefore, reaction rate	
	Describe the qualitative effect of a change in pressure of a gas on collision frequency and therefore, reaction rate	
1.5.5 Catalysts	Understand that a catalyst is a substance that increases the rate of a chemical reaction without being changed in chemical composition or amount	
	Catalysts work by providing an alternative reaction pathway with a lower activation energy	
	Use the Maxwell-Boltzmann distribution to explain how a catalyst increases the rate of a reaction involving a gas	
Required Practical 3: Investigation of how the rate of a reaction changes with temperature		

Topic 1.6: Equilibria

1.6.1 Chemical equilibria and Le Chatelier's principle	Know that many chemical reactions are reversible (give examples)	
	A reversible reaction is one in which the forward and reverse reactions proceed at equal rates and the concentrations of reactants and products remain constant	
	Le Chatelier's principle can be used to predict the effects of changes in temperature, pressure and concentration on the position of equilibrium in homogeneous reactions	
	Understand that a catalyst does not affect the position of equilibrium	
	Use Le Chatelier's principle to predict qualitatively the effect of changes in temperature, pressure and concentration on the position of equilibrium of a reversible reaction	
1.6.2 Equilibrium constant, K_c , for homogeneous systems	Explain why, for a reversible reaction used in an industrial process, a compromise temperature and pressure may be used	
	Know that the equilibrium constant, K_c , is deduced from the balanced chemical equation of a reversible reaction	
	The concentration (in mol dm^{-3}) of a species X involved in the expression for K_c is represented by [X]	
	The value of K_c is not affected by changes in concentration or by the addition of a catalyst	
	Construct an expression for K_c for a homogeneous system in equilibrium	
	Use the expression to calculate a value for K_c at constant temperature	
	Perform calculations involving K_c	
Predict the qualitative effects of changes in temperature on K_c values		

Topic 1.7: Oxidation, reduction and redox equations

1.7.1 Oxidation, reduction and redox equations	Know that oxidation is the process of electron loss and oxidising agents are electron acceptors	
	Know that reduction is the process of electron gain and reducing agents are electron donors	
	Know the rules for assigning oxidation states	
	Work out the oxidation state of an element in a compound or ion from its formula	
	Write half-equations identifying the oxidation and reduction processes in redox reactions	
	Combine half-equations to give an overall redox equation	

Inorganic Chemistry Year 2

Topic 2.4: Properties of Period 3 Elements and their Oxides (Periodicity)

3.2.4 Properties of Period 3 Elements and their Oxides	Describe the reactions of Na and Mg with water	
	Describe the trend in the reactions of the elements Na, Mg, Al, Si, P and S with oxygen; limited to the formation of Na ₂ O, MgO, Al ₂ O ₃ , SiO ₂ , P ₄ O ₁₀ , SO ₂ and SO ₃	
	Describe and explain the trend in melting point of the oxides of the elements Na – S in terms of their structure and bonding	
	Describe the reactions of the oxides of the elements Na – S with water; limited to Na ₂ O, MgO, Al ₂ O ₃ , SiO ₂ , P ₄ O ₁₀ , SO ₂ and SO ₃ and the pH of the solutions formed	
	Know the structures of the acids and the anions formed when P ₄ O ₁₀ , SO ₂ and SO ₃ react with water	
	Explain the trends in the reactions of the oxides with water in terms of the type of bonding present in each oxide	
	Write equations for the reactions that occur between oxides of the elements Na – S and given acids and bases	

Topic 2.5: Transition Metals

3.2.5.1 General properties of transition metals	Transition metal characteristics of elements Ti-Cu arise from an incomplete d sub-level in atoms or ions	
	The characteristic properties include: complex formation; formation of coloured ions; variable oxidation states; catalytic activity	
	A ligand is a molecule or ion that forms a coordinate bond with a transition metal by donating a pair of electrons	
	A complex is a central metal atom or ion surrounded by ligands	
	Coordination number is a the number of coordinate bonds to the central metal atom or ion	
3.2.5.2 Substitution reactions	H ₂ O, NH ₃ and Cl ⁻ can act as monodentate ligands	
	The ligands NH ₃ and H ₂ O are similar in size and are uncharged	
	Exchange of the ligands NH ₃ and H ₂ O occurs without change of coordination number (e.g. Co ²⁺ and Cu ²⁺)	
	Substitution may be incomplete (e.g. formation of [Cu(NH ₃) ₄ (H ₂ O) ₂] ²⁺)	
	The Cl ⁻ ligand is larger than the uncharged NH ₃ and H ₂ O ligands	
	Exchange of the ligand H ₂ O by Cl ⁻ can involve a change of coordination number (e.g. Co ²⁺ , Cu ²⁺ and Fe ³⁺)	
	Ligands can be bidentate (e.g. H ₂ NCH ₂ CH ₂ NH ₂ and C ₂ O ₄ ²⁻)	
	Ligands can be multidentate (e.g. EDTA ⁴⁻)	
	Haem is an iron(II) complex with a multidentate ligand	
	Oxygen forms a coordinate bond to Fe(II) in haemoglobin, enabling oxygen to be transported in the blood	
	Carbon monoxide is toxic because it replaces oxygen coordinately bonded to Fe(II) in haemoglobin	
	Bidentate and multidentate ligands replace monodentate ligands from complexes, this is called the chelate effect	
Explain the chelate effect in terms of the balance between entropy and enthalpy changes in these reactions		
3.2.5.3	Transition metal ions commonly form octahedral complexes with small ligands (e.g. NH ₃ and H ₂ O)	

Shapes of complex ions	Octahedral complexes can display cis-trans isomerism (a special case of E-Z isomerism) with monodentate ligands and optical isomerism with bidentate ligands	
	Transition metal ions commonly form tetrahedral complexes with larger ligands (e.g. Cl ⁻)	
	Square planar complexes are also formed and can display cis-trans isomerism (e.g. the drug cisplatin is the cis isomer)	
	Ag ⁺ forms the linear complex [Ag(NH ₃) ₂] ⁺ as used in Tollens' reagent	
3.2.5.4 Formation of coloured ions	Transition metal ions can be identified by their colour	
	Colour arises when some of the wavelengths of visible light are absorbed and the remaining wavelengths of light are transmitted or reflected	
	d electrons move from the ground state to an excited state when light is absorbed	
	The energy difference between the ground state and the excited state of the d electron is given by: $\Delta E = h\nu = hc/\lambda$	
	Changes in oxidation state, coordination number and ligand alter ΔE and this leads to a change in colour	
	The absorption of visible light is used in spectroscopy	
	A simple colorimeter can be used to determine the concentration of coloured ions in solution	
3.2.5.5 Variable oxidation states	Transition elements show variable oxidation states	
	Vanadium species in oxidation states IV, III and II are formed by the reduction of vanadate(V) ions by zinc in acidic solution	
	The redox potential for a transition metal ion changing from a higher to a lower oxidation state is influenced by pH and the ligand	
	The reduction of [Ag(NH ₃) ₂] ⁺ (Tollens' reagent) to metallic silver is used to distinguish between aldehydes and ketones	
	The redox titrations of Fe ²⁺ and C ₂ O ₄ ²⁻ with MnO ₄ ⁻	
	Using redox titration data to perform calculations	
3.2.5.6 Catalysts	Transition metals and their compounds can act as heterogeneous and homogeneous catalysts	
	A heterogeneous catalyst is in a different phase from the reactants and the reaction occurs at active sites on the surface	
	Describe the use of a support medium to maximise the surface area of a heterogeneous catalyst and minimise the cost	
	V ₂ O ₅ acts as heterogeneous catalyst in the Contact process	
	Fe acts as heterogeneous catalyst in the Haber process	
	Heterogeneous catalysts can become poisoned by impurities that block the active sites and consequently have reduced efficiency; this has a cost implication	
	A homogeneous catalyst is in the same phase as the reactants	
	When catalysts and reactants are in the same phase, the reaction proceeds through an intermediate species	
	Explain the importance of variable oxidation states in catalysis	
	Explain, with equations, how V ₂ O ₅ acts as a catalyst in the Contact process	
	Explain, with equations, how Fe ²⁺ ions catalyse the reaction between I ⁻ and S ₂ O ₈ ²⁻	
	Explain, with equations, how Mn ²⁺ ions autocatalyse the reaction between C ₂ O ₄ ²⁻ and MnO ₄ ⁻	

Topic 2.6: Reactions of Ions in Aqueous Solution

3.2.6 Reactions of Ions in Aqueous Solution	Know that in aqueous solution, the following metal-aqua ions are formed: $[M(H_2O)_6]^{2+}$ (where M = Fe and Cu); $[M(H_2O)_6]^{3+}$ (where M = Al and Fe);	
	Some metal hydroxides show amphoteric character by dissolving in both acids and bases (e.g. hydroxides of Al^{3+})	
	Explain in terms of the charge/size ratio of the metal ion, why the acidity of $[M(H_2O)_6]^{3+}$ is greater than that of $[M(H_2O)_6]^{2+}$	
	Describe and explain the simple test-tube reactions of: M^{2+} (aq) ions (where M = Fe and Cu) and M^{3+} (aq) ions (where M = Al and Fe), with the bases OH^- , NH_3 and CO_3^{2-}	
Required practical 11: Test-tube reactions to identify transition metal ions in (aq) solution		

Organic Chemistry Year 2

Topic 3.7: Optical Isomerism

3.3.7 Optical Isomerism	Optical isomerism is a form of stereoisomerism and occurs as a result of chirality in molecules, limited to molecules with a single chiral centre	
	An asymmetric carbon is chiral and gives rise to optical isomers (enantiomers), which exist as non super-imposable mirror images and differ in their effect on plane polarised light	
	A mixture of equal amounts of enantiomers is called a racemic mixture (racemate)	
	Draw structural formulas and displayed formulas of enantiomers	
	Understand how racemic mixtures are formed and why they are optically inactive	
	Draw 3D representations of chiral centres in molecules	

Topic 3.8: Aldehydes and Ketones

3.3.8 Aldehydes and ketones	Know that aldehydes are readily oxidised to carboxylic acids	
	Use chemical tests to distinguish between aldehydes and ketones (Fehling's solution and Tollens' reagent)	
	Know that aldehydes can be reduced to primary alcohols, and ketones to secondary alcohols, using $NaBH_4$ in aqueous solution. These reactions are nucleophilic addition	
	The nucleophilic addition reactions of carbonyl compounds with KCN, followed by dilute acid, produce hydroxynitriles	
	Explain why aldehydes and unsymmetrical ketones form mixtures of enantiomers when they react with KCN followed by dilute acid	
	Understand the hazards of using KCN	
	Write overall equations for reduction reactions, using $[H]$ as the reductant	
	Outline the nucleophilic addition mechanism for reduction reactions with $NaBH_4$ (showing the nucleophile as H^-)	
	Write overall equations for the formation of hydroxynitriles using HCN	
Outline the nucleophilic addition mechanism for the reaction with KCN, followed by dilute acid		

Topic 3.9: Carboxylic acids and derivatives

3.3.9.1 Carboxylic acids and derivatives	Identify the structures of carboxylic acids and esters	
	Know that carboxylic acids are weak acids, but will liberate CO ₂ from carbonates	
	Carboxylic acids and alcohols react, in the presence of an acid catalyst, to give esters	
	Esters are commonly used in solvents, plasticisers, perfumes and food flavourings	
	Vegetable oils and animal fats are esters of propane-1,2,2-triol (glycerol)	
	Esters can be hydrolysed in acid or alkaline conditions to form alcohols and carboxylic acids or salts of carboxylic acids	
	Vegetable oils and animal fats can be hydrolysed in alkaline conditions to give soap (salts of long-chain carboxylic acids) and glycerol	
	Biodiesel is a mixture of methyl esters of long-chain carboxylic acids	
3.3.9.2 Acylation	Biodiesel is produced by reacting vegetable oils with methanol in the presence of a catalyst	
	Identify the structures of acid anhydrides, acyl chlorides, amides	
	Describe the nucleophilic addition-elimination reactions of water, alcohols, ammonia and primary amines with acyl chlorides and acid anhydrides	
	Describe the industrial advantages of using ethanoic anhydride over ethanol chloride in the manufacture of the drug aspirin	
	Outline the nucleophilic addition-elimination mechanism of acyl chlorides with water, alcohols, ammonia and primary amines	
Required practical 10: Preparation of a pure organic solid and test of its purity and preparation of a pure organic liquid		

Topic 3.10: Aromatic Chemistry

3.3.10.1 Bonding	Describe the nature of the bonding in a benzene ring; its planar structure and its bond lengths which are intermediate between single and double bonds	
	Explain how the delocalisation of p electrons makes benzene more stable than the theoretical molecule cyclohexe-1,3,5-triene	
	Use thermochemical evidence from enthalpies of hydrogenation to account for the stability of benzene	
	Explain why benzene undergoes substitution reactions in preference to addition reactions	
3.3.10.2 Electrophilic substitution	Electrophilic attack on benzene rings results in substitution, limited to mono-substitutions	
	Nitration of benzene is an important step in synthesis, including the manufacture of explosives and formation of amines	
	Friedel-Crafts acylation reactions are also important steps in synthesis	
	Outline the electrophilic substitution mechanisms for nitration of benzene and acylation (using AlCl ₃ as a catalyst)	

Topic 3.11: Amines

3.3.11.1 Preparation	Primary aliphatic amines can be prepared by the reaction of ammonia with halogenoalkanes and by the reduction of nitriles	
	Aromatic amines, prepared by the reduction of nitro compounds, are used in the manufacture of dyes	
3.3.11.2 Base properties	Amines are weak bases	
	Describe and explain the difference in base strength between ammonia, primary aliphatic and primary aromatic amines in terms of the availability of the lone pair of electrons on the N atom	
3.3.11.2 Nucleophilic properties	Amines are nucleophiles	
	Outline the mechanism of the nucleophilic substitution reactions of ammonia and amines with halogenoalkanes to form primary, secondary and tertiary amines and quaternary ammonium salts	
	Describe the use of quaternary ammonium salts as cationic surfactants	
	Outline the mechanism of the nucleophilic addition-elimination reactions of ammonia and primary amines with acyl chlorides and acid anhydrides	

Topic 3.12: Polymers

3.3.12.1 Condensation polymers	Describe the formation of condensation polymers from the reactions of: dicarboxylic acids and diols; dicarboxylic acids and diamines; amino acids	
	Identify the repeating units in polyesters (e.g. Terylene) and polyamides (e.g. nylon-6,6 and Kevlar) and the linkages between units	
	Describe typical uses of polyesters and polyamides	
	Draw the repeating unit from a monomer structure or from a section of the polymer chain	
	Draw the structure of the monomer from a section of the polymer	
	Explain the nature of the intermolecular forces between molecules of condensation polymers	
3.3.12.2 Biodegradability and disposal of polymers	Polyalkenes are chemically inert and non-biodegradable	
	Polyesters and polyamides can be broken down by hydrolysis and are biodegradable	
	Explain why polyalkenes cannot be hydrolysed, but polyesters and polyamides can	
	Discuss the advantages and disadvantages of different methods of disposal of polymers, including recycling	

Topic 3.13: Amino acids, proteins and DNA

3.3.13.1 Amino acids	Amino acids have both acidic and basic properties, including the formation of zwitterions	
	Draw the structures of amino acids as zwitterions	
	Draw the structures of the ions formed from amino acids in acid solution and in alkaline solution	
3.3.13.2 Proteins	Proteins are sequences of amino acids joined by peptide links	
	Explain the importance of hydrogen bonding and sulfur-sulfur bonds in proteins	
	Describe and identify the primary, secondary (α -helix and β -pleated sheets) and tertiary structure of proteins	
	Hydrolysis of the peptide link produces the constituent amino acids	

	Amino acids can be separated and identified by thin-layer chromatography	
	Amino acids can be located on a chromatogram using developing agents such as ninhydrin or ultraviolet light and identified by their R_f values	
	Calculate R_f values from a chromatogram	
	Draw structures of peptides formed from up to three amino acids	
	Draw structures of amino acids formed by hydrolysis of a peptide	
3.3.13.3 Enzymes	Know that enzymes are proteins	
	Describe the action of enzymes as catalysts, including the concept of a stereospecific active site that binds to a substrate molecule	
	Explain why a stereospecific active site can only bond to one enantiomeric form of a substrate or drug	
	Describe the principle of a drug acting as an enzyme inhibitor by blocking the active site	
	Describe how computers can be used to help design such drugs	
3.3.13.4 DNA	Recognise the structures of the phosphate ion, 2-deoxyribose, the four bases (adenine, cytosine, guanine and thymine) given in the data booklet	
	A nucleotide is made up from a phosphate ion bonded to 2-deoxyribose, which is in turn bonded to one of the four bases	
	A single strand of DNA is a polymer of nucleotides linked by covalent bonds between the phosphate group of one nucleotide and the 2-deoxyribose of another nucleotide. This gives a sugar-phosphate-sugar-phosphate polymer chain with bases attached to the sugars in the chain	
	DNA exists as two complementary strands arranged in the form of a double helix	
	Explain how hydrogen bonding between base pairs leads to the two complementary strands of DNA	
3.3.13.5 Action of anti-cancer drugs	The Pt(II) complex cisplatin is used as an anti-cancer drug	
	Cisplatin prevents DNA replication in cancer cells by a ligand replacement reaction with DNA in which a bond is formed between platinum and a nitrogen atom on guanine	
	Describe the potential adverse effects of this mode of drug action	
	Discuss why society needs to assess the balance between the benefits and the adverse effects of drugs, such as cisplatin	

Topic 3.14: Organic Synthesis

3.3.14 Organic synthesis	The synthesis of an organic compound can involve several steps.	
	Explain why chemists aim to design processes that do not require a solvent and that use non-hazardous starting materials	
	Explain why chemists aim to design production methods with fewer steps that have a high percentage atom economy	
	Use the reactions covered in this course to devise a synthesis, with up to four steps, for an organic compound	

Topic 3.15: Nuclear Magnetic Resonance Spectroscopy (NMR)

3.3.15 Nuclear magnetic resonance spectroscopy	Understand that scientists have developed a range of analytical techniques, which together enable the structures of new compounds to be confirmed	
	NMR gives information about the position of ^{13}C or ^1H atoms in a molecule. ^{13}C NMR gives simpler spectra than ^1H NMR	
	Use of the δ scale for recording chemical shift	
	Know that chemical shift depends on the molecular environment	
	Use integrated data from ^1H NMR spectra to determine the relative numbers of equivalent ^1H atoms in different environments	

	¹ H NMR spectra are obtained using samples dissolved in deuterated solvents or CCl ₄	
	Explain why tetramethylsilane (TMS) is used as a standard	
	Use ¹³ C NMR and ¹ H NMR spectra and chemical shift data (provided) to suggest structures for molecules	
	Use the n+1 rule to deduce the spin-spin splitting patterns of adjacent, non-equivalent protons, limited to doublet, triplet and quartet formation in aliphatic compounds	

Topic 3.16: Chromatography

3.3.16 Chromatography	Chromatography can be used to separate and identify the components in a mixture	
	In thin-layer chromatography (TLC) a plate is coated with a solid and a solvent moves up the plate	
	In column chromatography (CC) a column is packed with a solid and a solvent moves up the plate	
	In gas chromatography (GC) a column is packed with a solid or with a solid coated by a liquid, and a gas is passed through the column under pressure at high temperature	
	Separation depends on the balance between solubility in the moving phase and retention by the stationary phase	
	Retention times and R _f values are used to identify different substances	
	Describe the use of mass spectrometry to analyse the components separated by GC	
	Calculate R _f values from a chromatogram	
	Compare retention times and R _f values with standards to identify different substances	
Required practical 12: Separation of species by thin-layer chromatography		

Physical Chemistry Year 2

Topic 1.8: Thermodynamics

1.8.1 Born-Haber cycles	Lattice enthalpy can be defined as either enthalpy of lattice dissociation or enthalpy of lattice formation	
	Born-Haber cycles are used to calculate lattice enthalpies using the following data: enthalpy of formation; ionisation energy; enthalpy of atomisation; bond enthalpy; electron affinity	
	Define each of the terms used in a Born-Haber cycle	
	Construct Born-Haber cycles and use them to calculate lattice enthalpies and other enthalpy changes	
	Compare lattice enthalpies from Born-Haber cycles with those from calculations based on a perfect ionic model to provide evidence for covalent character in ionic compounds	
	Define the terms enthalpy of solution and enthalpy of hydration	
	Construct cycles to calculate enthalpies of solution for ionic compounds from lattice enthalpies and enthalpies of hydration	
1.8.2 Gibbs free energy (ΔG) and entropy change (ΔS)	Understand that ΔH values, whilst important, are not sufficient to explain feasible reaction changes	
	Understand that entropy change (ΔS) is a concept of increasing disorder	
	Predict the sign of ΔS for physical changes and chemical changes	
	Calculate entropy changes from absolute entropy values	
	Know that the balance between entropy and enthalpy determines the feasibility of a reaction, given by: $\Delta G = \Delta H - T \Delta S$	
	Use this relationship to determine how ΔG varies with temperature	
	For a reaction to be feasible, the value of ΔG must be zero or negative.	
	Use the ΔG equation to determine the temperature at which a reaction becomes feasible	

Topic 1.9: Rate Equations (Kinetics)

1.9.1 Rate equations	Know that the rate of a chemical reaction is related to the concentration of reactants by a rate equation of the form: $\text{rate} = k[A]^m[B]^n$, where m and n are the orders of reaction with respect to reactants A and B and k is the rate constant	
	The orders m and n are restricted to the values 0, 1, 2	
	The rate constant k varies with temperature as shown by the Arrhenius equation: $k = Ae^{-E_a/RT}$, where A is the Arrhenius constant, E_a is the activation energy and T is the temperature in Kelvin	
	Define the terms order of reaction and rate constant	
	Perform calculations involving the rate equation	
	Explain the qualitative effect of changes in temperature on the rate constant, k	
	Perform calculations involving the Arrhenius equation, $k = Ae^{-E_a/RT}$	
	Re-arrange the Arrhenius equation into the form $\ln k = -E_a/RT + \ln A$ and use this form of the equation and experimental data to plot a straight-line graph of $\ln k$ versus $1/T$, with a gradient of $-E_a/R$	
1.9.2 Determination of the rate equation	Know that the rate equation is an experimentally determined relationship	
	The orders with respect to reactants can provide information about the mechanism of a reaction	

	Use concentration-time graphs to deduce the rate of a reaction	
	Use initial concentration-time data to deduce the initial rate of a reaction	
	Use rate-concentration data or graphs to deduce the order (0, 1 or 2) with respect to a reactant	
	Derive the rate equation for a reaction from the orders with respect to each of the reactants	
	Use the orders with respect to reactants to provide information about the rate-determining (limiting) step of a reaction	
Required practical 7: Measuring the rate of reaction by an initial rate method and a continuous monitoring method		

Topic 1.10: Equilibrium constant K_p for homogeneous systems

1.10 Equilibrium constant, K_p	Know that the equilibrium constant, K_p is deduced from the equation for a reversible reaction occurring in the gas phase	
	K_p is the equilibrium constant calculated from partial pressures for a system at constant temperature	
	Derive partial pressure from mole fraction and total pressure	
	Construct an expression for K_p for a homogeneous system in equilibrium	
	Perform calculations involving K_p	
	Predict the qualitative effects of changes in temperature and pressure on the position of the equilibrium and the value of K_p	
	Understand that whilst a catalyst can affect the rate of attainment of an equilibrium, it does not affect the value of the equilibrium constant	

Topic 1.11: Electrode Potentials and Electrochemical Cells

1.11.1 Electrode potentials and cells	Know the IUPAC convention for writing half-equations for electrode reactions	
	Write and apply the conventional representation of cells	
	Understand that cells are used to measure electrode potentials by reference to the standard hydrogen electrode	
	Understand the importance of the conditions used when measuring the electrode potential (E)	
	Know that the standard electrode potential (E^\ominus), refers to conditions of 298K, 100kPa and 1.00 mol dm ⁻³ solution of ions	
	Standard electrode potentials can be listed as an electrochemical series	
	Use E^\ominus values to predict the direction of simple redox reactions	
	Calculate the EMF of a cell	
1.11.2 Commercial applications of electrochemical cells	Electrochemical cells can be used as a commercial source of electrical energy	
	Understand the simplified electrode reactions in a lithium cell: positive electrode: $\text{Li}^+ + \text{CoO}_2 + \text{e}^- \rightarrow \text{Li}^+[\text{CoO}_2]^-$ negative electrode: $\text{Li} \rightarrow \text{Li}^+ + \text{e}^-$	
	Cells can be non-rechargeable (irreversible), rechargeable or fuel cells	
	Fuel cells are used to generate an electric current and do not need to be electrically recharged	
	Describe the electrode reactions in an alkaline hydrogen-oxygen fuel cell	
	Discuss the benefits and risks to society associated with using fuel cells	

	Use given electrode data to deduce the reactions occurring in non-rechargeable and rechargeable cells	
	Deduce the EMF of a cell	
	Explain how electrode reactions can be used to generate an electric current	
Required practical 8: Measuring the EMF of an electrochemical cell		

Topic 1.12 Acids, Bases and Buffers

1.12.1 Brønsted-Lowry acid-base equilibria	Define an acid as a proton donor and a base as a proton acceptor	
	Understand that acid-base equilibria involve the transfer of protons in aqueous solution	
1.12.2 Definition and determination of pH	The concentration of hydrogen ions in aqueous solution covers a very wide range of values. Therefore, a logarithmic scale (pH) is used as a measure of hydrogen ion concentration: $\text{pH} = -\log_{10}[\text{H}^+]$	
	Use the pH relationship to convert the concentration of hydrogen ions into pH and vice versa	
	Calculate the pH of a solution of a strong acid from its concentration	
1.12.3 The ionic product of water (K_w)	Know that water (H_2O) is slightly dissociated into its ions (H^+ and OH^-)	
	K_w is derived from the equilibrium constant for this dissociation	
	$K_w = [\text{H}^+][\text{OH}^-]$	
	The value of K_w varies with temperature	
	Use the K_w expression to calculate the pH of a strong base from its concentration	
1.12.4 Weak acids and bases	Weak acids and bases dissociate only slightly in aqueous solution	
	K_a is the dissociation constant for a weak acid	
	$\text{p}K_a = -\log_{10} K_a$	
	Construct expressions for K_a	
	Perform calculations relating the pH of a weak acid to the concentration of the acid and the dissociation constant, K_a	
	Convert K_a into $\text{p}K_a$ and vice versa	
1.12.5 pH curves, titrations and indicators	Describe the method of titrations of acids and bases and perform calculations based on experimental results	
	Sketch and explain the shapes of typical pH curves for acid-base titrations of weak and strong monoprotic acids and bases	
	Use pH curves to select an appropriate indicator	
1.12.6 Buffer action	Understand that a buffer solution maintains an approximately constant pH on dilution or addition of small amounts of acid or base	
	Acidic buffer solutions contain a weak acid its salt. A basic buffer solution contains a weak base and its salt	
	Describe the applications of buffer solutions	
	Explain qualitatively the action of acidic and basic buffers	
	Calculate the pH of acidic buffer solutions	
Required practical 9: Investigate how pH changes when a weak acid reacts with a strong base and when a strong acid reacts with a weak base		

Resource links:

CGP online textbook: <https://www.cgpbooks.co.uk/bookspaceInfo>

Topic notes, videos and exam questions: <http://www.a-levelchemistry.co.uk/>

Additional theory: <https://www.chemguide.co.uk/>

AQA full specification: <http://www.aqa.org.uk/subjects/science/as-and-a-level/chemistry-7404-7405>