AQA A-LEVEL CHEMISTRY

SUMMARY OF EXTRA CONTENT NEEDED FOR PAPER 1

Topic 10 – Thermodynamics (Paper 1 only)

Topic 11 – How Far How Fast II (both Papers unless stated)

Topic 12 – Acids, Bases and Buffers (Paper 1 only)

Topic 13 – Electrochemistry (Paper 1 only)

Topic 14 – Reactions of Period 3 Elements and their Oxides (Paper 1 only)

Topic 15 – Transition Metals and Complex Ions (Paper 1 only)

Topic 10 – Thermodynamics

* I can define the terms enthalpy of lattice dissociation, enthalpy of lattice formation, enthalpy of formation, ionisation energy, enthalpy of atomisation, bond enthalpy and electron affinity
* I can construct Born–Haber cycles and use them to calculate lattice enthalpies
* I can 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
* I can define the term enthalpy of hydration and use cycles to calculate enthalpies of solution for ionic compounds from lattice enthalpies and enthalpies of hydration and to perform other enthalpy change calculations
* I can explain that ∆H, whilst important, is not sufficient to explain feasible change, that the concept of increasing disorder (entropy change, ∆S) accounts for the above deficiency, illustrated by physical changes and chemical changes, and I can calculate entropy changes from absolute entropy values
* I can explain that balance between entropy and enthalpy determines the feasibility of a reaction given by the relationship: ∆G = ∆H – T∆S (derivation not required), explain that for a reaction to be feasible, the value of ∆G must be zero or negative, and use the relationship ∆G = ∆H – T∆S to determine how ∆G varies with temperature and to determine the temperature at which a reaction becomes feasible

Topic 11 – How Far How Fast II

* I can construct rate equations of the form Rate = k[A]m[B]n where m and n are the orders of reaction with respect to reactants A and B (restricted to the values 0, 1 and 2) and k is the rate constant, I can appreciate that the rate equation is an experimentally derived relationship (Paper 2 only)
* I can define the terms order of reaction and rate constant (Paper 2 only)
* I can explain the rate constant k varies with temperature as shown by the equation: k = Ae(–Ea/RT) (where A is a constant, known as the Arrhenius constant, Ea is the activation energy and T is the temperature in K), perform calculations using this equation and rearrange this equation to plot a straight line graph with slope Ea/R given suitable data (Paper 2 only)
* I can use the orders with respect to reactants to deduce information about the mechanism of a reaction and its rate determining step in particular (Paper 2 only)
* I can use concentration–time graphs to deduce the rate of a reaction, including the initial rate of a reaction (Paper 2 only)
* I can use rate–concentration data or graphs to deduce the order (0, 1 or 2) with respect to a reactant and derive the rate equation for a reaction from the orders with respect to each of the reactants (Paper 2 only)
* I can carry out and describe an experiment to measure the rate of a reaction and the order of reaction with respect to a reactant using the initial rate method **(Required Practical 7a)**
* I can carry out and describe an experiment to measure the rate of a reaction with respect to a reactant using a continuous monitoring method **(Required Practical 7b)**
* I can construct an expression for the equilibrium constant Kp from the equation for a reversible reaction occurring in the gas phase, using the partial pressures for a system at constant temperature
* I can derive partial pressure from mole fraction and total pressure and perform calculations involving Kp
* I can predict the qualitative effects of changes in temperature and pressure on the position of equilibrium and the value of Kp
* I can explain that whilst a catalyst can affect the rate of attainment of an equilibrium, it does not affect the value of the equilibrium constant

Topic 12 – Acids, Bases and Buffers

* I can describe an acid as a proton donor, a base as a proton acceptor and acid–base equilibria as reactions involving the transfer of protons
* I can appreciate that the concentration of hydrogen ions in aqueous solution covers a very wide range, and therefore a logarithmic scale, the pH scale, is used as a measure of hydrogen ion concentration: pH = –log10[H+]
* I can convert concentration of hydrogen ions into pH and vice versa
* I can calculate the pH of a solution of a strong acid from its concentration
* I understand that water is slightly dissociated, that KW is derived from the equilibrium constant for this dissociation (KW = [H+][OH–]) and that the value of KW varies with temperature
* I can use Kw to calculate the pH of a strong base from its concentration
* I can explain that weak acids and weak bases dissociate only slightly in aqueous solution, that Ka is the dissociation constant for a weak acid and that pKa = –log10Ka
* I can construct an expression for Ka and perform calculations relating the pH of a weak acid to the concentration of the acid and the dissociation constant, Ka
* I can convert Ka into pKa and vice versa
* I can describe titrations of acids with bases and I can perform calculations for these titrations based on experimental results
* I can sketch and explain the shapes of typical pH curves in all combinations of weak and strong monoprotic acids and use pH curves to select an appropriate indicator
* I can carry out an experiment to investigate how pH changes when a weak acid reacts with a strong base and when a strong acid reacts with a weak base (Required Practical 9)
* I can describe a buffer solution as maintaining an approximately constant pH, despite dilution or addition of small amounts of acid or base, that acidic buffer solutions contain a weak acid and the salt of that weak acid and that basic buffer solutions contain a weak base and the salt of that weak base, and I can explain qualitatively the action of acidic and basic buffers
* I can calculate the pH of acidic buffer solutions
* I can describe applications of buffer solutions

Topic 13 – Electrochemistry

* I can use EƟ values to calculate the EMF of a cell and to predict the direction of simple redox reactions
* I can write an apply the conventional representation of a cell
* I understand that cells are used to measure electrode potentials by reference to the standard hydrogen electrode
* I understand the importance of the conditions when measuring the electrode potential, E (Nernst equation not required) and that standard electrode potential, EƟ, refers to conditions of 298 K, 100 kPa and 1.00 mol dm−3 solution of ions.
* I can use the IUPAC convention for writing half-equations for electrode reactions
* I understand that standard electrode potentials can be listed as an electrochemical series
* I can carry out an experiment to measure the emf of a cell (Required Practical 8)
* I understand that electrochemical cells can be used as a commercial source of electrical energy, and that cells can be non-rechargeable (irreversible), rechargeable or fuel cells
* I can recall the simplified electrode reactions in a lithium cell (positive electrode Li+ + CoO2 + e– → Li+[CoO2]–, negative electrode Li → Li+ + e–, and the electrode reactions in an alkaline hydrogen–oxygen fuel cell
* I understand that fuel cells are used to generate an electric current and do not need to be electrically recharged
* I can appreciate the benefits and risks to society associated with using these cells
* I understand that fuel cells are used to generate an electric current and do not need to be electrically recharged
* I can explain how the electrode reactions can be used to generate an electric current

Topic 14 – The Reactions of Period 3 Elements and Their Oxides

* I can describe the reactions of Na and Mg with water
* I can describe the trends in the reactions of the elements Na, Mg, Al, Si, P and S with oxygen, limited to the formation of Na2O, MgO, Al2O3, SiO2, P4O10, SO2 and SO3
* I can describe the trend in the melting point of the highest oxides of the elements Na–S and explain this trend in terms of their structure and bonding
* I can describe and write equations the reactions of the oxides of the elements Na–S with water, limited to Na2O, MgO, Al2O3, SiO2, P4O10, SO2 and SO3, and the pH of the solutions formed, I can explain these trends in terms of the type of bonding present in each oxide and write equations for the reactions that occur between the oxides of the elements Na–S and given acids and base
* I can describe the structures of the acids and the anions formed when P4O10, SO2 and SO3 react with water
* I can write equations for the reactions that occur between the oxides of the elements Na–S and given acids and bases

Topic 15 – Transition Metals and Complex Ions

* I can explain that the transition metal characteristics of elements Ti–Cu arise from an incomplete d sub-level in atoms or ions, and that the characteristic properties include complex formation, formation of coloured ions, variable oxidation states and catalytic activity
* I can describe a ligand as a molecule or ion that forms a co-ordinate bond with a transition metal by donating a pair of electrons, a complex as a central metal atom or ion surrounded by ligands and co-ordination number as the number of co-ordinate bonds to the central metal atom or ion
* I can explain that H2O, NH3 and Cl− can act as monodentate ligands, that the ligands NH3 and H2O are similar in size and uncharged, that exchange of the ligands NH3 and H2O occurs without change of co-ordination number (eg Co2+ and Cu2+), that substitution may be incomplete (eg the formation of [Cu(NH3)4(H2O)2]2+), that the Cl− ligand is larger than the uncharged ligands NH3 and H2O and that exchange of the ligand H2O by Cl– can involve a change of co-ordination number (eg Co2+, Cu2+ and Fe3+)
* I can explain that ligands can be bidentate (eg H2NCH2CH2NH2 and C2O42–) or multidentate (eg EDTA4–), that haem is an iron(II) complex with a multidentate ligand, that oxygen forms a co-ordinate bond to Fe(II) in haemoglobin, enabling oxygen to be transported in the blood and that carbon monoxide is toxic because it replaces oxygen co-ordinately bonded to Fe(II) in haemoglobin
* I can explain that bidentate and multidentate ligands replace monodentate ligands from complexes and that this is called the chelate effect, and I can explain the chelate effect in terms of the balance between the entropy and enthalpy change in these reactions
* I can explain that transition metal ions commonly form octahedral complexes with small ligands (eg H2O and NH3), that octahedral complexes can display cis–trans isomerism (a special case of E–Z isomerism) with monodentate ligands and optical isomerism with bidentate ligands, that transition metal ions commonly form tetrahedral complexes with larger ligands (eg Cl–), that square planar complexes are also formed and can display cis–trans isomerism, that cisplatin is a cis isomer and that Ag+ forms the linear complex [Ag(NH3)2]+ as used in Tollens’ reagent
* I can explain how in aqueous solution the following metal-aqua ions are formed: [M(H2O)6]2+, limited to M = Fe and Cu, and [M(H2O)6]3+, limited to M = Al and Fe, I can explain why the acidity of [M(H2O)6]3+ is greater than that of [M(H2O)6]2+ and explain this in terms of the charge/size ratio of the metal ion, I can state that some metal hydroxides show amphoteric character by dissolving in both acids and bases (eg hydroxides of Al3+), and I can describe and explain the simple test-tube reactions of: M2+(aq) ions, limited to M = Fe and Cu, and of M3+(aq) ions, limited to M = Al and Fe, with the bases OH–, NH3 and CO32–
* I can explain that transition metal ions can be identified by their colour, that this colour arises when some of the wavelengths of visible light are absorbed and the remaining wavelengths of light are transmitted or reflected, that d electrons move from the ground state to an excited state when light is absorbed, that the energy difference between the ground state and the excited state of the d electrons is given by ∆E = hν = hc/λ, that changes in oxidation state, co-ordination number and ligand alter ∆E and this leads to a change in colour, that the absorption of visible light is used in spectroscopy and that a simple colorimeter can be used to determine the concentration of coloured ions in solution
* I can explain that transition elements show variable oxidation states, that vanadium species in oxidation states IV, III and II are formed by the reduction of vanadate(V) ions by zinc in acidic solution, that the redox potential for a transition metal ion changing from a higher to a lower oxidation state is influenced by pH and by the ligand, and that the reduction of [Ag(NH3)2]+ (Tollens’ reagent) to metallic silver is used to distinguish between aldehydes and ketones
* I can describe the redox titrations of Fe2+ and C2O42– with MnO4–, and I perform calculations for these titrations and similar redox reactions
* I can explain that transition metals and their compounds can act as heterogeneous and homogeneous catalysts and that variable oxidation states are important in catalysis
* I can explain that a heterogeneous catalyst is in a different phase from the reactants and the reaction occurs at active sites on the surface, the use of a support medium to maximise the surface area of a heterogeneous catalyst and minimise the cost, that V2O5 acts as a heterogeneous catalyst in the Contact process and explain this with the aid of equations, that Fe is used as a heterogeneous catalyst in the Haber process, that heterogeneous catalysts can become poisoned by impurities that block the active sites and consequently have reduced efficiency and that this has a cost implication
* I can explain that a homogeneous catalyst is in the same phase as the reactants, that when catalysts and reactants are in the same phase the reaction proceeds through an intermediate species, write equations to show how Fe2+ ions catalyse the reaction between I− and S2O82– and explain, with the aid of equations, how Mn2+ ions autocatalyse the reaction between C2O42– and MnO4–
* I can carry out simple test-tube reactions to identify transition metal ions in aqueous solution (Required Practical 11)