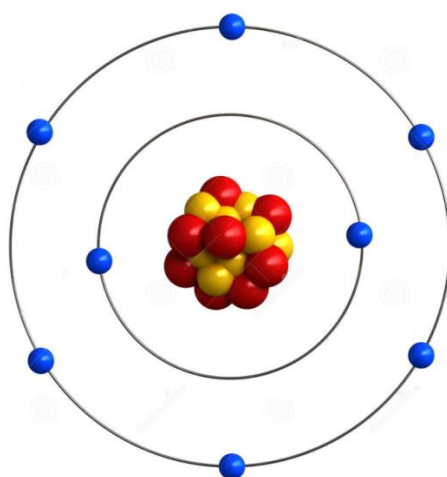# A-level Chemistry

3.1 Physical Chemistry (Y13)

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Lesson Date(s)	Topic	Learning Outcomes From Specification	Key Words & Equations	Notes (Revision tips, exam technique tips, priorities for revision etc.)
	3.1.8.1 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. Students should be able to: define each of the above terms and lattice enthalpy construct Born–Haber cycles to calculate lattice enthalpies using these enthalpy changes construct Born–Haber cycles to calculate one of the 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. Cycles are used to calculate enthalpies of solution for ionic compounds from lattice enthalpies and enthalpies of hydration. Students should be able to: define the term enthalpy of hydration perform calculations of an enthalpy change using thesecycles.		

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	3.1.8.2 Gibbs Free-Energy Change & Entropy	ΔH, whilst important, is not sufficient to explain feasible change. The concept of increasing disorder (entropy change, ΔS). ΔS accounts for the above deficiency, illustrated by physical changes and chemical changes. The balance between entropy and enthalpy determines the feasibility of a reaction given by the relationship: $\Delta G = \Delta H - T\Delta S$ (derivation not required). For a reaction to be feasible, the value of ΔG must be zero or negative. Students should be able to: calculate entropy changes from absolute entropy values use the relationship $\Delta G = \Delta H - T\Delta S$ to determine how ΔG varies with temperature use the relationship $\Delta G = \Delta H - T\Delta S$ to determine the temperature at which a reaction becomes feasible.		

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	3.1.9.1 Rate Equations	The rate of a chemical reaction is related to the concentration of reactants by a rate equation 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 and k is the rate constant. The orders m and n are restricted to the values 0, 1, and 2. 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, E_a is the activation energy and T is the temperature in K. Students should be able to: define the terms order of reaction and rate constant perform calculations using the rate equation explain the qualitative effect of changes in temperature on the rate constant k perform calculations using the equation $k = Ae^{-Ea/RT}$ understand that the equation $k = Ae^{-Ea/RT}$ can be rearranged into the form $\ln k = -Ea/RT + \ln A$ and know how to use this rearranged equation with experimental data to plot a straight line graph with slope $-Ea/R$ These equations and the gas constant, R, will be given when required.		

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	3.1.9.2 Determination of rate equation	The rate equation is an experimentally determined relationship. The orders with respect to reactants can provide information about the mechanism of a reaction. Students should be able to: 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** **by a continuous monitoring method.**		

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	3.1.10 Equilibrium Constant K_p	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. Students should be able to: 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 equilibrium predict the qualitative effects of changes in temperature on 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.		

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	3.1.11.1 Electrode potentials & cells	IUPAC convention for writing half-equations for electrode reactions. The conventional representation of cells. Cells are used to measure electrode potentials by reference to the standard hydrogen electrode. The importance of the conditions when measuring the electrode potential, E (Nernst equation not required). Standard electrode potential, $E^{\ominus}$, refers to conditions of 298 K, 100 kPa and 1.00 mol dm−3 solution of ions. Standard electrode potentials can be listed as an electrochemical series. Students should be able to: use $E^{\ominus}$ values to predict the direction of simple redox reactions calculate the EMF of a cell write and apply the conventional representation of a cell. **Required practical 8** **Measuring the EMF of an electrochemical cell.**		

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	3.1.11.2 Commercial Application of Electrochemical Cells	Content Opportunities for skills development Electrochemical cells can be used as a commercial sourceof electrical energy. The simplified electrode reactions in a lithium cell: Positive electrode: $Li^+ + CoO_2 + e^- \rightarrow Li+[CoO_2]^-$ Negative electrode: $Li \rightarrow Li^+ + e^-$ 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. The electrode reactions in an alkaline hydrogen–oxygen fuel cell. The benefits and risks to society associated with using these cells. Students should be able to: use given electrode data to deduce the reactions occurring in non-rechargeable and rechargeable cells deduce the EMF of a cell explain how the electrode reactions can be used to generate an electric current.		

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	3.1.12.1 Brosted-Lowry Acid-Base Equilibria	An acid is a proton donor. A base is a proton acceptor. Acid–base equilibria involve the transfer of protons.		
	3.1.12.2 Definition & Determination of pH	The concentration of hydrogen ions in aqueous solution covers a very wide range. Therefore, a logarithmic scale, the pH scale, is used as a measure of hydrogen ion concentration. $pH = -\log_{10}[H^+]$ Students should be able to: convert concentration of hydrogen ions into pH and vice versa calculate the pH of a solution of a strong acid from its oncentration.		

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	3.1.12.3 Ionic Product of Water K_w	Water is slightly dissociated. KW is derived from the equilibrium constant for this dissociation. $K_W = [H^+][OH^-]$ The value of K_W varies with temperature. Students should be able to use K_W to calculate the pH of a strong base from its concentration.		
	3.1.12.4 Weak Acids and base, and K_a	Weak acids and weak bases dissociate only slightly in aqueous solution. K_a is the dissociation constant for a weak acid. $pK_a = -\log_{10} K_a$ Students should be able to: construct an expression 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 pK_a and vice versa.		

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	3.1.12.5 pH Curves, Titrations & Indicators	Titrations of acids with bases. Students should be able to perform calculations for these titrations based on experimental results. Typical pH curves for acid–base titrations in all combinations of weak and strong monoprotic acids and bases. Students should be able to: sketch and explain the shapes of typical pH curves use pH curves to select an appropriate indicator. **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.**		

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		A buffer solution maintains an approximately constant pH, despite dilution or addition of small amounts of acid or base.
		
		Acidic buffer solutions contain a weak acid and the salt of that weak acid.
		
		Basic buffer solutions contain a weak base and the salt of that weak base.
		
		Applications of buffer solutions.
		
		Students should be able to:
		
		    explain qualitatively the action of acidic and basic buffers
		
		    calculate the pH of acidic buffer solutions.		