Topic Name	Term	Skills Developed	Link to NC Subject Content	Next link in curriculum	Other Notes
3.2.5 Transition Metals	Autumn 1	• AT d and k PS 1.2 Students could carry out test-tube reactions of complexes with monodentate, bidentate and multidentate ligands to compare ease of substitution. AT d and k PS 2.2 Students could carry out test-tube reactions of solutions of metal aqua ions with ammonia or concentrated hydrochloric acid. • PS 3.1 and 3.2 Students could determine the concentration of a solution of copper(II) ions by colorimetry.	• Define what a transition metals is and describe their physical and chemical properties. • Define the terms ligand, complex and coordination number • Define the terms monodentate, bidentate and multidentate ligand and be able to give examples of each • Explain the cheating effect in terms of the balance between entropy and enthalpy. • Explain how haemoglobin is used to transport oxygen around the blood and why carbon monoxide is toxic • Write equations for complete and incomplete ligand substitution. • Draw complexes with coordinate numbers of 2, 4 and 6 and work out the charges of the ions formed. • Name the shapes of different complexs and explain why complexes with Cl- ligand form tetrahedral complexes. • State the types of isomerism shown by complexes with monodentate and bidentate ligands.	*3.2.6 Reactions of Ions in Aqueous Solution*	

		MS 3.1 and 3.2 Students determine the concentration of a solution from a graph of absorption versus concentration. AT a, e and k Students could determine the concentration of a coloured complex ion by colorimetry. • AT d and k PS 1.2 Students could reduce vanadate(V) with zinc in acidic solution. AT b, d and k PS 4.1 Students could carry out test-tube reactions of Tollens' reagent to distinguish aldehydes and ketones. AT a, d, e and k PS 2.3, 3.2 and 3.3 Students could carry out redox titrations. • AT d and k PS 4.1 Students could investigate Mn2+ as	• Explain why transition metal ions display colour. • Calculate the energy difference between ground state and excited state. • Explain what factors effect the colour of a compound • Explain how colourimetry can be used to determine to concentration of coloured ions in solution. • Write the chemical formula of vanadium species in oxidation states V, IV, III, II and state how vanadate(V) can be reduced through each of these oxidation states • Explain how the reduction of $[Ag(NH_3)]^+$ can be used t distinguish between aldehydes and ketones. • Perform calculations for redox titrations between Fe^{2+} and MnO_4^- and $C_2O_4^{2-}$ and MnO_4^- • Define homogeneous and heterogeneous catalyst. • Write equations to show how V2O5 acts as a catalyst in the contact process • State the catalysts used for the Haber Process • Explain how catalysts can be poisoned.		

		the autocatalyst in the reaction between ethanedioic acid and acidified potassium manganate(VII).	• Explain, with the aid of equations, how Fe2+ ions catalyse the reaction between I– and S2 O8 2– • Explain, with the aid of equations, how Mn2+ ions autocatalyse the reaction between C2 O4 2– and MnO4 –		
3.2.6 Reactions of Ions in Aqueous Solution	Autumn 2	• AT d and K PS 1.2 Students could carry out test-tube reactions of metal-aqua ions with NaOH, NH3 and Na2 CO3 AT d and k PS 2.2 Students could carry out test-tube reactions to identify the positive and negative ions in this specification. PS 1.1 Students could identify unknown substances using reagents.	• Explain, in terms of the charge/size ratio of the metal ion, why the acidity of [M(H2 O)6] 3+ is greater than that of [M(H2 O)6] 2+ • 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 CO3 2–		Required Practical 11 Simple test tube reactions to Identify metal ions in aqueous solution
3.1.8 Thermodynamics	Autumn 2	• AT a, b and k PS 3.2 Students could be asked to find ΔS for vaporization of water	• Define enthalpy of formation, ionisation energy, enthalpy of atomisation, bond enthalpy, electron affinity and lattice enthalpy	•	

		using a kettle. MS 2.2, 2.3 and 2.4 Students rearrange the equation ΔG = ΔH – TΔS to find unknown values. MS 3.3 Students determine ΔS and ΔH from a graph of ΔG versus T.	• 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. • Define the term enthalpy of hydration • Perform calculations of an enthalpy change using these cycles. • Calculate entropy changes from absolute entropy values • Use the relationship ΔG = ΔH – TΔS to determine how ΔG varies with temperature • Use the relationship ΔG = ΔH – TΔS to determine the temperature at which a reaction becomes feasible.		

3.1.9 Rate Equations	Spring 1	• MS 0.0 and 2.4 Students use given rate data and deduce a rate equation, then use some of the data to calculate the rate constant including units. Rate equations could be given and students asked to calculate rate constant or rate. MS 3.3 and 3.4 Students use a graph of concentration–time and calculate the rate constant of a zero-order reaction by determination of the gradient. • AT a, b, k and l PS 2.4 and 3.1 Students could determine the order of reaction for a reactant in the iodine clock reaction. MS 3.1 Students could be given data to plot and interpret	• 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 • 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	•	Required Practical 7 Measuring the rate of reaction by initial rate method and Continuous monitoring

		in terms of order with respect to a reactant. Alternatively, students could just be given appropriate graphs and asked to derive order(s). MS 3.3 and 3.4 Students calculate the rate constant of a zero-order reaction by determining the gradient of a concentration–time graph. MS 3.5 Students plot concentration–time graphs from collected or supplied data and draw an appropriate best-fit curve. Students draw tangents to such curves to deduce rates at different times.	• 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.		

3.1.11 Electrode Potentials	Spring 2	• AT j and k PS 1.1 Students could make simple cells and use them to measure unknown electrode potentials. AT a, b, j and k PS 2.1 and 2.4 Students could be asked to plan and carry out an experiment to investigate the effect of changing conditions, such as concentration or temperature, in a voltaic cell such as Zn|Zn2+||Cu2+|Cu AT j and k PS 2.2 Students could use EΘ values to predict the direction of simple redox reactions, then test these predictions by simple test-tube reactions.	• Use EΘ values to predict the direction of simple redox reactions • Calculate the EMF of a cell • Write and apply the conventional representation of a cell • 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.	•	Required Practical 8 Measuring the EMF of an electrochemical cell