West Kirby Grammar School

Curriculum Map – Year 12 – Chemistry (Teacher 1)

Topic Name	Term	Skills Developed	Link to NC Subject Content	Next link in curriculum	Other Notes
3.1.1 Atomic Structure	*Autumn 1*	• MS 1.1 Students report calculations to an appropriate number of significant figures, given raw data quoted to varying numbers of significant figures. • MS 1.2 Students calculate weighted means eg calculation of an atomic mass based on supplied isotopic abundances. • MS 3.1 Students interpret and analyse spectra.	• Appreciate that knowledge and understanding of atomic structure has evolved over time. Protons, neutrons and electrons: relative charge and relative mass. An atom consists of a nucleus containing protons and neutrons surrounded by electrons. • Mass number (A) and atomic (proton) number (Z). • Determine the number of fundamental particles in atoms and ions using mass number, atomic number and charge • Explain the existence of isotopes. • Calculate relative atomic mass from isotopic abundance, limited to mononuclear ions. • Interpret simple mass spectra of elements. • Calculate relative atomic mass from isotopic abundance, limited to mononuclear ions. • Define first ionisation energy • Write equations for first and successive ionisation energies • Explain how first and successive ionisation energies in Period 3 (Na–Ar) and in Group 2 (Be–Ba) give evidence for	• Transition Metals (Y13) Redox Titration Calculations. • Periodicity.	

			electron configuration in sub-shells and in shells.		
3.1.3 Bonding	*Autumn 1*	• PS 1.1 Students could be asked to find the type of structure of unknowns by experiment (eg to test solubility, conductivity and ease of melting). • MS 0.3 and 4.1 Students could be given familiar and unfamiliar examples of species and asked to deduce the shape according to valence shell electron pair repulsion (VSEPR) principles.	• Predict the charge on a simple ion using the position of the element in the Periodic Table. • Construct formulas for ionic compounds. • A covalent bond using a line • A co-ordinate bond using an arrow • Relate the melting point and conductivity of materials to the type of structure and the bonding present. • Explain the energy changes associated with changes of state Draw diagrams to represent these structures involving specified numbers of particles. • Explain the shapes of, and bond angles in, simple molecules and ions with up to six electron pairs (including lone pairs of electrons) surrounding the central atom. • Use partial charges to show that a bond is polar.	• Period 3 and their oxides (Y13) When discussing melting and boiling points of Period 3 and their oxides in relation to their structure. • Transition Metals (Y13) When discussing shapes of complex ions.	

		• PS 1.2 Students could try to deflect jets of various liquids from burettes to investigate the presence of different types and relative size of intermolecular forces.	• Explain why some molecules with polar bonds do not have a permanent dipole. • Explain the existence of these forces between familiar and unfamiliar molecules. • Explain how melting and boiling points are influenced by these intermolecular forces.		
3.1.4 Energetics	Autumn 2	• MS 0.0 and 1.1 Students understand that the correct units need to be used in q = mcΔT Students report calculations to an appropriate number of significant figures, given raw data quoted to varying numbers of significant figures. Students understand that calculated	• Define standard enthalpy of combustion (Δc HΘ) • Define standard enthalpy of formation (Δf HΘ). The heat change, q, in a reaction is given by the equation q = mcΔT • Use this equation to calculate the molar enthalpy change for a reaction • Use this equation in related calculations.		

		results can only be reported to the limits of the least accurate measurement. • PS 2.4, 3.1, 3.2, 3.3 and 4.1 Students could be asked to find ΔH for a reaction by calorimetry. Examples of reactions could include: • dissolution of potassium chloride • dissolution of sodium carbonate • neutralising NaOH with HCl • displacement reaction between CuSO4 + Zn • combustion of alcohols. • MS 2.4 Students carry out Hess's law calculations. • AT a and k PS 2.4, 3.2 and 4.1 Students could be asked to find ΔH for a reaction using Hess's law and calorimetry, then	• Use Hess's law to perform calculations, including calculation of enthalpy changes for reactions from enthalpies of combustion or from enthalpies of formation.		• Required practical 2 Measurement of an enthalpy change.

		present data in appropriate ways. Examples of reactions could include: • thermal decomposition of $NaHCO_3$ • hydration of $MgSO_4$ • hydration of $CuSO_4$ • MS 1.2 Students understand that bond enthalpies are mean values across a range of compounds containing that bond.	• Define the term mean bond enthalpy • Use mean bond enthalpies to calculate an approximate value of ΔH for reactions in the gaseous phase. • Explain why values from mean bond enthalpy calculations differ from those determined using Hess's law.		
3.1.5 Kinetics	Autumn 2	• AT a, b, k and l PS 2.4 and 3.1 Students could investigate the effect of temperature on the rate of reaction of sodium thiosulfate and hydrochloric acid by	• Define the term activation energy • Explain why most collisions do not lead to a reaction. • Draw and interpret distribution curves for different temperatures. • Use the Maxwell–Boltzmann distribution to explain why a small temperature increase can lead to a large increase in rate.		• Required practical 3 Investigation of how the rate of a reaction changes with temperature.

		an initial rate method. • Research opportunity Students could investigate how knowledge and understanding of the factors that affect the rate of chemical reaction have changed methods of storage and cooking of food. • AT a, e, k and i Students could investigate the effect of changing the concentration of acid on the rate of a reaction of calcium carbonate and hydrochloric acid by a continuous monitoring method.	• Explain how a change in concentration or a change in pressure influences the rate of a reaction. • Use a Maxwell–Boltzmann distribution to help explain how a catalyst increases the rate of a reaction involving a gas.		

3.1.6 Chemical Equilibria, Le Chatelier's Principle and Kc	Spring 1	• PS 1.1 Students could carry out test-tube equilibrium shifts to show the effect of concentration and temperature (eg Cu(H2 O)6 2+ with concentrated HCl). • MS 0.3 Students estimate the effect of changing experimental parameters on a measurable value eg how the value of Kc would change with temperature, given different specified conditions. • MS 1.1 Students report calculations to an appropriate number of significant figures, given raw data quoted to varying numbers of significant figures. Students understand that calculated results can only be reported to the limits	• Use Le Chatelier's principle to predict qualitatively the effect of changes in temperature, pressure and concentration on the position of equilibrium • Explain why, for a reversible reaction used in an industrial process, a compromise temperature and pressure may be used. • Construct an expression for Kc for a homogeneous system in equilibrium. • Calculate a value for Kc from the equilibrium concentrations for a homogeneous system at constant temperature. • Perform calculations involving Kc. • Predict the qualitative effects of changes of temperature on the value of Kc.	• Equilibrium Constant Kp for homogenous systems (Y13)	

		of the least accurate measurement. • MS 2.2 and 2.3 Students calculate the concentration of a reagent at equilibrium. Students calculate the value of an equilibrium constant Kc • PS 2.3 Students could determine the equilibrium constant, Kc , for the reaction of ethanol with ethanoic acid in the presence of a strong acid catalyst to ethyl ethanoate.			
3.1.7 Oxidation, Reduction and Redox Equations.	Spring 1		• Work out the oxidation state of an element in a compound or ion from the formula. • Write half-equations identifying the oxidation and reduction processes in redox reactions. • Combine half-equations to give an overall redox equation.	• Transition Metals (Y13) Variable Oxidation states of transition metals.	

3.2.1 Periodicity	Spring 2		• 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. • Explain the trends in atomic radius and first ionisation energy • Explain the melting point of the elements in terms of their structure and bonding.	• Period 3 and their Oxides (Y13) •	
3.2.3 Group 7- The Halogens.	Spring 2	• AT d and k PS 2.2 Students could carry out test-tube reactions of solutions of the halogens (Cl2 , Br2 , I2) with solutions containing their halide ions (eg KCl, KBr, KI). Students could record observations from reactions of NaCl, NaBr and NaI with concentrated sulfuric acid. Students could carry out tests for halide ions using acidified silver nitrate, including the use of ammonia to distinguish the silver halides formed.	• Explain the trend in electronegativity • Explain the trend in the boiling point of the elements in terms of their structure and bonding. • silver nitrate solution is used to identify halide ions • The silver nitrate solution is acidified • Ammonia solution is added.	• Periodicity.	

		• Research opportunity Students could investigate the treatment of drinking water with chlorine. Students could investigate the addition of sodium fluoride to water supplies	• The reaction of chlorine with water to form chloride ions and chlorate(I) ions. • The reaction of chlorine with water to form chloride ions and oxygen. • Appreciate that society assesses the advantages and disadvantages when deciding if chemicals should be added to water supplies. • The use of chlorine in water treatment. Appreciate that the benefits to health of water treatment by chlorine outweigh its toxic effects. • The reaction of chlorine with cold, dilute, aqueous NaOH and uses of the solution formed.		• Required practical 4 Carry out simple test-tube reactions to identify: • Cations – Group 2, NH4 + Anions – Group 7 (halide ions), OH– , CO3 2–, SO4 2–
3.2.4 Properties of Period 3 Elements and their Oxides (A-level only)	Summer 2	**AT a, c and k** **PS 2.2** Students could carry out reactions of elements with oxygen and test the pH of the resulting oxides.	• Explain the trend in the melting point of the oxides of the elements Na–S in terms of their structure and bonding • 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 the oxides of the elements Na–S and given acids and bases.		

3.1.10 Equilibrium Constant Kp for Homogeneous Systems (A-level only)	Summer 2	MS 1.1 Students report calculations to an appropriate number of significant figures, given raw data quoted to varying numbers of significant figures. Students understand that calculated results can only be reported to the limits of the least accurate measurement. MS 2.2 and 2.3 Students calculate the partial pressures of reactants and products at equilibrium. Students calculate the value of an equilibrium constant K p	• 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.		