

## Biddulph High School Curriculum Intent

To deliver a broad and enriching curriculum through engaging and challenging lessons that provide a wide range of opportunities for all students to achieve their potential.

Students will all be prepared to take their next steps in a diverse and ever changing future ready to make a positive contribution to society.

Through a broad programme of extracurricular activities students will have the opportunities to showcase their talents and experience new challenges.

We value individuals and all that they can offer as well as supporting each other with kindness and empathy.

### Curriculum Intent for Chemistry: Year 12

*. Our overarching aim is to develop well rounded chemists who can not only explain complex theoretical concepts, but can investigate them practically for themselves. Investigative skills and techniques are at the heart of chemistry at BHS and we have embedded frequent opportunities for our learners to develop these at all stages of the curriculum. Our curriculum reflects the vast number of our students for whom chemistry will form part or all of their further studies and to this end, from year 9 to year 13, we deliver our lessons with the expectation that our students will take the skills and knowledge they have gained here on into their future careers*

All teachers will follow the schemes of work provided by the department. This will ensure that all students receive the same high-quality provision. All units of work will provide a clear outline of the knowledge and skills required and assessments will ensure that this knowledge has been retained and that skills can be evidenced.

Teachers will ensure that gaps are closed through regular monitoring within the classroom. DINT activities will allow for interleaving and recap of previous learning. Misconceptions will be identified through effective questioning and the regular inspection of student work.

Chemistry Long Term Overview						
Year Group	Autumn 1	Autumn 2	Spring 1	Spring 2	Summer 1	Summer 2
9	Particle theory, equations and the periodic table	Methods of separation	Maths skills and calculating relative formula mass	pH, acids and indicators	Green Chemistry	Energetics and atomic theory
10	Atoms, isotopes and periodicity	Bonding and molar calculations	Acid, base reactions and electrolysis	Energetics and electrochemical cells. Rates of reaction.	Mock examinations and intervention	Equilibrium
11	Equilibria and extracting metals	Corrosion, Alloys and Mock examinations	Organic Chemistry	Interpreting and interacting with Earth systems.	Examinations	Examinations
12	Foundations in Chemistry. Nomenclature and alkanes	Acids and base chemistry. Periodicity. Alkene reactions and mechanisms.	Energetics. Alcohols and Haloalkanes	Hess's Law and Rates. Synthetic pathways.	Equilibria and revision. Analysis – Mass spec and IR.	Examinations. Intervention. begin Module 5 – Rates and Module 6 – Benzene.
13	Equilibria – Kc and Kp. Acids and bases. Arenes and carbonyls.	Buffers. Energetics. Carboxylic acids and derivatives. Amines and amides. C-C bond formation.	Revision Mock examinations.	Redox titration. Electrochemistry. Transition metals. Synthesis. NMR and gas chromatography.	Examinations	Examinations

<b>Chemistry: Medium Term Overview</b>			
<b>Year 13</b>	<b>Autumn Term 1</b>	<b>Unit Title: How far, How fast.</b>	<b>No of Lessons: 33</b>
<b>Overview/Intent</b>	<p>The largely qualitative treatment of reaction rates and equilibria encountered in Module 3 is developed within a quantitative and graphical context. This section also allows learners to develop practical quantitative techniques involved in the determination of reaction rates and pH. There are many opportunities for developing mathematical skills, including use of logarithms and exponents, when studying the content of this section and when carrying out quantitative practical work.</p> <p>Born–Haber cycles are used as a theoretical model to illustrate the energy changes associated with ionic bonding. Entropy and free energy are then introduced as concepts used to predict quantitatively the feasibility of chemical change. Redox chemistry permeates chemistry and the introductory work in Module 2 is developed further within this section, including use of volumetric analysis for redox titrations and an introduction of electrochemistry in the context of electrode potentials.</p>		
<b>Assessment</b>			

	Homework tasks. End of topic tests and formative assessments.	
<p><b><u>Essential Knowledge (what must students know):</u></b></p> <p><b>Equilibrium</b></p> <ul style="list-style-type: none"> <li>• use of the terms mole fraction and partial pressure See also 3.2.3 Chemical Equilibrium. (b) calculation of quantities present at equilibrium, given appropriate data (c) the techniques and procedures used to determine quantities present at equilibrium Not for Kp.</li> <li>• HSW4 Opportunities to carry out experimental and investigative work. (d) expressions for Kc and Kp for homogeneous and heterogeneous equilibria Note: liquid and solid concentrations are constant and are omitted in heterogeneous Kc and Kp expressions. (e) calculations of Kc and Kp, or related quantities, including determination of units</li> <li>• the qualitative effect on equilibrium constants of changing temperature for exothermic and endothermic reactions (ii) the constancy of equilibrium constants with changes in concentration, pressure or in the presence of a catalyst (g) explanation of how an equilibrium constant controls the</li> </ul>	<p><b><u>Essential Skills (what must students be able to demonstrate):</u></b></p> <p><b>Students will be able to:</b></p> <ul style="list-style-type: none"> <li>• Calculate Kc and Kp from experimental data and build ICE tables</li> <li>• Link the position of equilibria to the value of Kc</li> <li>• Make predictions on equilibria based on equations and enthalpy values.</li> <li>• Calculate pH of a strong acid</li> <li>• Link conjugate acid-base pairs</li> <li>• Calculate the pH of a weak acid using the Ka equation.</li> <li>• Use ionic product of water Kw to calculate the pH of a strong alkali.</li> <li>• Use molar calculations to determine the pH of remaining solutions following acid-base reactions.</li> <li>• State the definition of a buffer.</li> <li>• Calculate pH of a solution formed from a weak acid and its conjugate base.</li> <li>• Calculate the pH of a buffer solution from a weak acid and a strong alkali.</li> </ul>	<p><b>Lessons:</b></p> <ol style="list-style-type: none"> <li>1 Kc</li> <li>2 Kc</li> <li>3 Kp</li> <li>4 Kp</li> <li>5 Controlling equilibria</li> <li>6 Topic review</li> <li>7 TEST</li> <li>8 Intervention</li> <li>9 Conjugate pairs</li> <li>10 pH of strong acids</li> <li>11 Ka</li> <li>12 pH of weak acids</li> <li>13 pH of strong bases</li> <li>14 Revision</li> <li>15 Revision</li> <li>16 Assessment 1</li> <li>17 Intervention</li> <li>18 Buffers</li> <li>19 Buffers in the body</li> <li>20 Examination practice on buffers</li> <li>21 PAG 11.1</li> <li>22 Neutralisation and titration curves</li> <li>23 Born Haber cycles</li> <li>24 Calculating lattice enthalpy</li> </ol>

<p>position of equilibrium on changing concentration, pressure and temperature (h) application of the above principles in 5.1.2 How far? for <math>K_c</math>, <math>K_p</math> to other equilibrium constants, where appropriate</p> <p><b>Brønsted–Lowry acids and bases</b></p> <ul style="list-style-type: none"> <li>(i) a Brønsted–Lowry acid as a species that donates a proton and a Brønsted–Lowry base as a species that accepts a proton (ii) use of the term conjugate acid–base pairs (iii) monobasic, dibasic and tribasic acids Learners should be able to identify acid–base pairs in equations for acid–base equilibria. HSW1,7 Development of different models over time to explain acid–base behaviour. (b) the role of <math>H^+</math> in the reactions of acids with metals and bases (including carbonates, metal oxides and alkalis), using ionic equations (c) (i) the acid dissociation constant, <math>K_a</math>, for the extent of acid dissociation (see also 2.1.4 b) (ii) the relationship between <math>K_a</math> and <math>pK_a</math></li> </ul> <p><b>pH and <math>[H^+(aq)]</math></b></p> <ul style="list-style-type: none"> <li>use of the expression for pH as: <math>pH = -\log[H^+]</math> <math>[H^+] = 10^{-pH}</math> (e) use of the expression for the ionic product of water, <math>K_w</math> M0.1, M0.4, M2.2, M2.3, M2.4 (f) calculations of pH, or related quantities, for: (i) strong monobasic acids (ii) strong bases, using <math>K_w</math> M0.1, M0.4, M2.2, M2.3, M2.4, M2.5 (g) calculations of pH, <math>K_a</math> or related quantities, for a weak monobasic acid using approximations</li> </ul>	<ul style="list-style-type: none"> <li>Explain the use of buffers in the blood</li> <li>Use titration curves to identify the correct indicators to use in an experiment.</li> <li>Be able to build Born Haber cycles for a range of ionic compounds. This is also to include cycles for enthalpies of solution.</li> <li>Calculate lattice enthalpies.</li> <li>State and explain the factors that affect lattice enthalpy and enthalpies of hydration.</li> <li>Apply rules of entropy to determine if a reaction will have an increase or decrease in entropy.</li> <li>Calculate entropy.</li> <li>Calculate feasibility of a reaction using Gibbs equation and explain whether a reaction is feasible or not..</li> <li>Calculate EMF for cells.</li> <li>Draw and fully label cell diagrams</li> <li>Use redox rules to build equations.</li> </ul>	<p>25 Enthalpies of solution practical</p> <p>26 Unknown enthalpy cycles</p> <p>27 Factors affecting LE</p> <p>28 Entropy</p> <p>29 Gibbs</p> <p>30 TEST</p> <p>31 Intervention</p> <p>32 REDOX and electrode potentials</p> <p>33 REDOX practicals</p> <p>34 REDOX practicals</p>
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<ul style="list-style-type: none"><li>• (h) limitations of using approximations to <math>K_a</math> related calculations for 'stronger' weak acids Including reasons why <math>[HA]_{\text{equilibrium}} \sim [HA]_{\text{undissociated}}</math> may no longer be valid.</li><li>• HSW6 Understanding of the circumstances under which <math>K_a</math> approximations break down.</li></ul> <p><b>Buffers: action, uses and calculations</b></p> <ul style="list-style-type: none"><li>• (i) a buffer solution as a system that minimises pH changes on addition of small amounts of an acid or a base (j) formation of a buffer solution from: (i) a weak acid and a salt of the weak acid, e.g. <math>\text{CH}_3\text{COOH}/\text{CH}_3\text{COONa}</math> (ii) excess of a weak acid and a strong alkali, e.g. excess <math>\text{CH}_3\text{COOH}/\text{NaOH}</math> (k) explanation of the role of the conjugate acid–base pair in an acid buffer solution, e.g. <math>\text{CH}_3\text{COOH}/\text{CH}_3\text{COO}^-</math>, in the control of pH (l) calculation of the pH of a buffer solution, from the <math>K_a</math> value of a weak acid and the equilibrium concentrations of the conjugate acid–base pair; calculations of related quantities 5 (m) explanation of the control of blood pH by the carbonic acid–hydrogencarbonate buffer system</li></ul> <p><b>Neutralisation</b></p> <ul style="list-style-type: none"><li>• pH titration curves for combinations of strong and weak acids with strong and weak bases, including: (i) sketch and interpretation of their shapes (ii) explanation of the choice of suitable</li></ul>		
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<p>indicators, given the pH range of the indicator (iii) explanation of indicator colour changes in terms of equilibrium shift between the HA and A<sup>-</sup> forms of the indicator.</p> <ul style="list-style-type: none"><li>• No indicator is suitable for a weak acid/weak base titration. The indicator should be considered as a weak acid, HA.</li></ul> <p>(o) the techniques and procedures used when measuring pH with a pH meter.</p> <p><b>Energy</b></p> <ul style="list-style-type: none"><li>• explanation of the term lattice enthalpy (formation of 1 mol of ionic lattice from gaseous ions, <math>\Delta_{LEH}</math>) and use as a measure of the strength of ionic bonding in a giant ionic lattice (see also 2.2.2 b–c) Definition required. Born–Haber and related enthalpy cycles (b) use of the lattice enthalpy of a simple ionic solid (e.g. NaCl, MgCl<sub>2</sub>) and relevant energy terms for: (i) the construction of Born–Haber cycles (ii) related calculations</li><li>• explanation and use of the terms: (i) enthalpy change of solution (dissolving of 1 mol of solute, <math>\Delta_{solH}</math>) (ii) enthalpy change of hydration (dissolving of 1 mol of gaseous ions in water, <math>\Delta_{hydH}</math>) Definitions required. Details of infinite dilution not required. (d) use of the enthalpy change of solution of a simple ionic solid (e.g. NaCl, MgCl<sub>2</sub>) and relevant energy terms (enthalpy change of hydration and lattice enthalpy) for: (i) the</li></ul>		
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<p>construction of enthalpy cycles (ii) related calculations</p> <ul style="list-style-type: none"><li>• qualitative explanation of the effect of ionic charge and ionic radius on the exothermic value of a lattice enthalpy and enthalpy change of hydration.</li></ul> <p><b>Entropy</b></p> <ul style="list-style-type: none"><li>• (a) explanation that entropy is a measure of the dispersal of energy in a system which is greater, the more disordered a system HSW1 The model of entropy to explain thermodynamic stability. (b) explanation of the difference in magnitude of the entropy of a system: (i) of solids, liquids and gases (ii) for a reaction in which there is a change in the number of gaseous molecules (c) calculation of the entropy change of a system, <math>\Delta S</math>, and related quantities for a reaction given the entropies of the reactants and products</li><li>• Free energy (d) explanation that the feasibility of a process depends upon the entropy change and temperature in the system, <math>T\Delta S</math>, and the enthalpy change of the system, <math>\Delta H</math> HSW1,5,6 Use of entropy, enthalpy and temperature for predicting feasibility. (e) explanation, and related calculations, of the free energy change, <math>\Delta G</math>, as: <math>\Delta G = \Delta H - T\Delta S</math> (the Gibbs' equation) and that a process is feasible when <math>\Delta G</math> has a negative value</li></ul> <p><b>Redox</b></p>		
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<ul style="list-style-type: none"> <li>(a) explanation and use of the terms oxidising agent and reducing agent (see also 2.1.5 Redox) (b) construction of redox equations using halfequations and oxidation numbers. (c) interpretation and prediction of reactions involving electron transfer Redox titrations (d) the techniques and procedures used when carrying out redox titrations including those involving <math>\text{Fe}^{2+}/\text{MnO}_4^-</math> and <math>\text{I}_2/\text{S}_2\text{O}_3^{2-}</math> (see also 2.1.5 e–f) HSW4 Opportunities to carry out experimental and investigative work. (e) structured and non-structured titration calculations, based on experimental results of redox titrations involving: (i) <math>\text{Fe}^{2+}/\text{MnO}_4^-</math> and <math>\text{I}_2/\text{S}_2\text{O}_3^{2-}</math> (ii) non-familiar redox systems</li> </ul>		
<p><b>Careers Links:</b> Nuclear physics with particle models. Laboratory work. Analytical chemist. Medicinal chemist. Forensic Scientist.</p> <p><b>Homework</b> SENECA learning and examination style questions to support learning and improve skill. Work booklet for homework tasks. Completion of additional reading and research.</p>	<p><b>Enrichment:</b> Students will have the opportunity to attend the trip to Geneva.</p>	<p><b>MY PB</b> <b>Social Me- active listening, speaking effectively, working with others</b> Practical work will require aspects of the social me strand <b>Thinking Me – evaluating &amp; creativity</b> Evaluation will be utilised when assessing data from investigations <b>This is Me – Resilience, responsibility, self-motivation, integrity, self-management</b> Students will need to demonstrate resilience and self-management when looking at the assessed points across the lessons</p>

Chemistry <b>Medium Term Overview</b>			
Year 13	<i>Spring Term 2</i>	Unit Title: <b>Electrochemistry and Transition metals</b>	No of Lessons: <b>40</b>
<b>Overview</b>	This section provides learners with a deeper knowledge and understanding of the periodic table within the context of the transition elements. This section includes the role of ligands in complex ions, stereochemistry, precipitation, ligand substitution and redox reactions. The colour changes and observations in these reactions increase the toolkit of qualitative inorganic tests for identifying unknown ionic compounds.		
<b>Assessment</b>	Mock examinations		
<b>Essential Knowledge (what must students know):</b>	<b>Essential Skills (what must students be able to demonstrate):</b>	<b>Lessons:</b>	
<p><b>Electrode potentials</b></p> <ul style="list-style-type: none"> <li>use of the term standard electrode (redox) potential, <math>E^{\ominus}</math>, including its measurement using a hydrogen electrode <math>E^{\ominus}</math> data will be provided on examination papers. (g) the techniques and procedures used for the measurement of cell potentials of: (i) metals or non-metals in contact with their ions in aqueous solution (ii) ions of the same element in different oxidation states in contact with a Pt electrode For measurement of standard cell potentials, ions of the same element can have concentrations of <math>1 \text{ mol dm}^{-3}</math> or be equimolar. PAG8 HSW4 Opportunities to carry out experimental and investigative</li> </ul>	<p><b>Students will be able to:</b></p> <ul style="list-style-type: none"> <li>Calculate EMF for cells.</li> <li>Draw and fully label cell diagrams</li> <li>Use redox rules to build equations.</li> <li><math>\text{Fe}^{2+}</math> can be oxidised with <math>\text{H}^+/\text{MnO}_4^-</math> and <math>\text{Fe}^{3+}</math> reduced with <math>\text{I}^-</math>, <math>\text{Cr}^{3+}</math> can be oxidised with <math>\text{H}_2\text{O}_2/\text{OH}^-</math> and <math>\text{Cr}_2\text{O}_7^{2-}</math> reduced with <math>\text{Zn}/\text{H}^+</math>, <math>\text{Cu}^{2+}</math> can be reduced with <math>\text{I}^-</math>. In aqueous conditions, <math>\text{Cu}^+</math> readily disproportionates.</li> <li>.Give examples of ligands</li> <li>Examples should include: monodentate: <math>\text{H}_2\text{O}</math>, <math>\text{Cl}^-</math> and <math>\text{NH}_3</math></li> </ul>	<ol style="list-style-type: none"> <li>Revision</li> <li>Revision</li> <li>Revision</li> <li>Revision</li> <li>Revision</li> <li>Mocks</li> <li>Mocks</li> <li>Mocks</li> <li>Mocks</li> <li>Mocks</li> <li>Intervention</li> <li>Intervention</li> <li>Intervention</li> <li>Intervention</li> <li>Intervention</li> <li>PAG 12.2</li> <li>PAG 12.2</li> </ol>	

<p>work. (h) calculation of a standard cell potential by combining two standard electrode potentials (i) prediction of the feasibility of a reaction using standard cell potentials and the limitations of such predictions in terms of kinetics and concentration</p> <p><b>Storage and fuel cells</b></p> <ul style="list-style-type: none"> <li>• application of principles of electrode potentials to modern storage cells Details of storage cells and required equations will be provided. Relevant electrode potentials and other data will be supplied. HSW9 Benefits of electrochemical cells counteracted by risks from toxicity and fire from Li-based cells. (k) explanation that a fuel cell uses the energy from the reaction of a fuel with oxygen to create a voltage and the changes that take place at each electrode.</li> </ul> <p><b>Transition metals</b></p> <p><b>Properties</b></p> <ul style="list-style-type: none"> <li>• the electron configuration of atoms and ions of the d-block elements of Period 4 (Sc–Zn), given the atomic number and charge (see also 2.2.1 d) Learners should use sub-shell notation e.g. for Fe: <math>1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2</math>. (b) the elements Ti–Cu as transition elements i.e. d-block elements that have an ion with an incomplete d-sub-shell (c) illustration, using at least two transition elements, of: (i) the</li> </ul>	<p>bidentate: <math>\text{NH}_2\text{CH}_2\text{CH}_2\text{NH}_2</math> ('en'). In exams, other ligands could be introduced.</p> <ul style="list-style-type: none"> <li>• Learners should be able to draw 3-D diagrams to illustrate stereoisomerism.</li> <li>• For precipitation, non-complexed formulae or complexed formulae, are acceptable e.g. <math>\text{Cu}^{2+}(\text{aq})</math> or <math>[\text{Cu}(\text{H}_2\text{O})_6]^{2+}</math>; <math>\text{Cu}(\text{OH})_2(\text{s})</math> or <math>\text{Cu}(\text{OH})_2(\text{H}_2\text{O})_4</math>. With excess NaOH, only <math>\text{Cr}(\text{OH})_3</math> reacts further forming <math>[\text{Cr}(\text{OH})_6]^{3-}</math>. With excess <math>\text{NH}_3</math>, only <math>\text{Cr}(\text{OH})_3</math> and <math>\text{Cu}(\text{OH})_2</math> react forming <math>[\text{Cr}(\text{NH}_3)_6]^{3+}</math> and <math>[\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+}</math> respectively</li> </ul>	<p><b>18 Electrode potentials</b></p> <p><b>19 Predictions from electrode potentials</b></p> <p><b>20 Storage and fuel cells</b></p> <p><b>21 PAG 8.1</b></p> <p><b>22 PAG 8.1</b></p> <p><b>23 D Block elements</b></p> <p><b>24 Complex ions</b></p> <p><b>25 Stereoisomerism.</b></p> <p><b>26 Ligand substitution</b></p> <p><b>27 REDOX and QA</b></p> <p><b>28 TEST</b></p> <p><b>29 Past paper Revision</b></p> <p><b>30 Past paper Revision</b></p> <p><b>31 Past paper Revision</b></p> <p><b>32 Past paper Revision</b></p> <p><b>33 Past paper Revision</b></p> <p><b>34 Past paper Revision</b></p> <p><b>35 Past paper Revision</b></p> <p><b>36 Past paper Revision</b></p> <p><b>37 Past paper Revision</b></p> <p><b>38 Past paper Revision</b></p> <p><b>39 Past paper Revision</b></p>
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<p>existence of more than one oxidation state for each element in its compounds</p> <ul style="list-style-type: none"> <li>the formation of coloured ions</li> <li>iii) the catalytic behaviour of the elements and their compounds and their importance in the manufacture of chemicals by industry</li> </ul> <p><b>Ligands and complex ions</b></p> <ul style="list-style-type: none"> <li>explanation and use of the term ligand in terms of coordinate (dative covalent) bonding to a metal ion or metal, including bidentate ligands Examples should include: monodentate: <math>\text{H}_2\text{O}</math>, <math>\text{Cl}^-</math> and <math>\text{NH}_3</math> bidentate: <math>\text{NH}_2\text{CH}_2\text{CH}_2\text{NH}_2</math> ('en'). In exams, other ligands could be introduced. (e) use of the terms complex ion and coordination number and examples of complexes with: (i) six-fold coordination with an octahedral shape (ii) four-fold coordination with either a planar or tetrahedral shape (see also 2.2.2 g-h)</li> <li>types of stereoisomerism shown by complexes, including those associated with bidentate and multidentate ligands: (i) cis-trans isomerism e.g. <math>\text{Pt}(\text{NH}_3)_2\text{Cl}_2</math> (see also 4.1.3 c-d) (ii) optical isomerism e.g. <math>[\text{Ni}(\text{NH}_2\text{CH}_2\text{CH}_2\text{NH}_2)_3]^{2+}</math></li> <li>use of cis-platin as an anti-cancer drug and its action by binding to DNA preventing cell division HSW9 Benefits of chemotherapy; risks from unpleasant side effects. Ligand substitution (h) ligand substitution reactions and the accompanying colour changes in the formation of: (i) <math>[\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+}</math> and</li> </ul>		
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<p>[CuCl<sub>4</sub>]<sup>2-</sup> from [Cu(H<sub>2</sub>O)<sub>6</sub>]<sup>2+</sup> (ii) [Cr(NH<sub>3</sub>)<sub>6</sub>]<sup>3+</sup> from [Cr(H<sub>2</sub>O)<sub>6</sub>]<sup>3+</sup></p> <ul style="list-style-type: none"> <li>(i) explanation of the biochemical importance of iron in haemoglobin, including ligand substitution involving O<sub>2</sub> and CO Precipitation reactions (j) reactions, including ionic equations, and the accompanying colour changes of aqueous Cu<sup>2+</sup>, Fe<sup>2+</sup>, Fe<sup>3+</sup>, Mn<sup>2+</sup> and Cr<sup>3+</sup> with aqueous sodium hydroxide and aqueous ammonia, including: (i) precipitation reactions (ii) complex formation with excess aqueous sodium hydroxide and aqueous ammonia</li> <li></li> </ul> <p><b>Practical opportunities</b> PAG8 HSW4 Opportunities to carry out experimental and investigative work.</p>		
<p><b>Careers Link:</b> Forensics and police work. Environmental scientists.</p>	<p><b>Enrichment</b> Students will be taught about the role of analysis in drug production and detection.</p>	<p><b>MY PB</b> <b>Social Me- active listening, speaking effectively, working with others</b></p>

<p>Medicinal chemistry</p> <p><b>Homework</b>            SENECA learning and examination style questions to support learning and improve skill.            Homework booklets.            Research and further reading.</p>		<p>Practical work will require aspects of the social me strand</p> <p><b>Thinking Me – evaluating &amp; creativity</b>            Evaluation will be utilised when assessing data from investigations</p> <p><b>This is Me – Resilience, responsibility, self-motivation, integrity, self-management</b>            Students will need to demonstrate resilience and self-management when looking at the assessed points across the lessons</p>
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Chemistry <b>Medium Term Overview</b>			
Year 13	<i>Summer Term 3</i>	Unit Title: End of Year examinations	No of Lessons: NA
<p><b>Overview</b></p>          <p><b>Assessment</b></p>	<p><b>Periodic table, elements and physical chemistry (Component 01)</b>            This component is worth 100 marks and is split into two sections and assesses content from teaching modules 1, 2, 3 and 5. Learners answer all questions. Section A contains multiple choice questions. This section of the paper is worth 15 marks. Section B includes short answer question styles (structured questions, problem solving, calculations, practical) and extended response questions. This section of the paper is worth 85 marks.</p> <p><b>Synthesis and analytical techniques (Component 02)</b>            This component is worth 100 marks and is split into two sections and assesses content from teaching modules 1, 2, 3 and 5. Learners answer all questions. Section A contains multiple choice questions. This section of the paper is worth 15 marks. Section B includes short answer question styles (structured questions, problem solving, calculations, practical) and extended response questions. This section of the paper is worth 85 marks.</p> <p><b>Unified chemistry (Component 03)</b>            This component assesses content from across all teaching modules 1 to 6. Learners answer all questions. This component is worth 70 marks. Question styles include short answer (structured questions, problem solving, calculations, practical) and extended response questions.</p>		

	<p><b>Practical Endorsement in chemistry (Component 04)</b></p> <p>Performance in this component is reported separately to the performance in the A level as measured through externally assessed components 01 to 03. This non-exam assessment component rewards the development of practical competency in chemistry and is teacher assessed. Learners demonstrate competence in the range of skills and techniques specified in Section 1.2 of the specification by carrying out a minimum of 12 assessed practical activities. The Practical Endorsement is teacher assessed against the Common Practical Assessment Criteria as specified in Section 5g. Learners may work in groups but must demonstrate and record independent evidence of their competency. Teachers who award a pass to their learners must be confident that each learner consistently and routinely exhibits the competencies listed in Section 5g and has demonstrated competence in all the skills detailed in section 1.2.1 and in all the apparatus and techniques detailed in Section 1.2.2 before completion of the A level course. The practical activities provided by OCR are all mapped against the specification and assessment criteria.</p>	

