

Year 11 Chemistry Program 2022

Week	Learning Objectives	Text References & Activities	Assessments
Term 1 Week 1	Topic 1: PROPERTIES & ATOMIC STRUCTURE <u>Essential Chemistry Skills</u> <ol style="list-style-type: none"> Represent ions using formulae which include the number of constituent atoms and the charge on the ion. Recognise, name and write the formulae of a range of ions (see Appendix 1). Determine the formulae of ionic compounds from the formulae of their constituent ions (see Appendix 1). Use molecular formulae to represent the number and type of atoms in a covalent molecule, and recognise some common covalent molecular substances (See Appendix 1). Write and use balanced chemical equations to represent chemical reactions. Define the mole as a precisely defined quantity of matter, equal to Avogadro's number of particles. Use Avogadro's number (see Data Booklet) to convert between number of particles and moles, and vice versa. Use the relationship between mass, molar mass and number of moles to convert between moles and mass, and vice versa 	<u>Essential Chemistry Skills</u> <ul style="list-style-type: none"> Lucarelli Ch 1.7 – 1.10: Atoms, isotopes, ions, A_r Lucarelli Ch 4.1 – 4.3: Symbols and formulae Lucarelli 6.1 – 6.2: Chemical equations Lucarelli Ch 10.1 – 10.5: Moles & molar mass Set 16 Qs 1-5 & 12 Pearson Ch 9.1 – 9.3: The mole 	CAP 1: Term 1 Week 5 <u>Separation Techniques</u> Practical validation based on Exp 1 – 3 (spec points 13 – 16) 5% of year
Term 1 Weeks 2 – 3	<u>Properties of Matter</u> <ol style="list-style-type: none"> Define what is meant by a pure substance, and describe their common properties, including melting and boiling points, hardness and density. State and compare the characteristics of homogeneous and heterogeneous substances. Distinguish between pure substances (elements and compounds), homogeneous mixtures and heterogeneous mixtures. Explain the relationship between the identity and relative amounts of substances in a solution, and its properties. Explain how differences in physical properties of substances in a mixture can be used to separate them. Describe a range of separating techniques, and explain which physical property is utilised in the method of separation, including: filtration, decanting, use of separating funnel, recrystallization, distillation and fractional distillation. Use knowledge of separating techniques to devise methods to separate mixtures of substances. <i>Conduct experimental separation techniques safely, competently and methodically. (SIS)</i> 	STAWA Sets 7 – Elements & symbols 13 – Compounds & formulae 22 – Molar mass 23 – Moles, particles and mass 24 – Interpretation of formulae <u>Properties of Matter</u> Theory <ul style="list-style-type: none"> Lucarelli Ch 3: Mixtures & Separation, Set 4 Pearson Ch 1: Materials in our world Practicals <ul style="list-style-type: none"> STAWA Exp 1 Mixtures STAWA Exp 2 Separation by filtration & crystallisation STAWA Exp 3 Distillation (demo?) STAWA Sets 4 – Mixtures	<u>Structure of the Atom & Nanomaterials</u> Extended response report set in week 3 Validation in Week 5 5% of year

Term 1 Week 3	<u>Nanomaterials</u> 17. Describe what nanomaterials are. 18. State and describe how the properties of nanomaterials may differ from the properties of the bulk material, and give examples. 19. Identify uses of nanomaterials and suggest advantages and disadvantages to their use. 20. <i>Describe the potential health risks associated with the use of nanoparticles. (SHE)</i> 21. <i>Describe and explain examples of environmental concerns associated with the use of nanomaterials. (SHE)</i> 22. <i>Provide a balanced discussion of the advantages and disadvantages of nanomaterials. (SHE)</i>	<u>Nanomaterials</u> Theory <ul style="list-style-type: none"> Lucarelli Ch 8.10 – 8.14, Set 14 Qs 11 – 14 Pearson Ch 1: Materials in our world Pearson Ch 4.4 (later part): Metallic nanomaterials Pearson Ch 7.2: Carbon nanomaterials
Term 1 Weeks 4 – 7	<u>Atomic Structure</u> <u>Development of the Atomic Model</u> 23. <i>Construct a timeline of the development of the atomic model. (SHE)</i> 24. <i>Describe and explain the role of the following scientists in the development of our atomic model: (SHE)</i> <ol style="list-style-type: none"> Dalton Thomson Rutherford Bohr Chadwick 25. <i>Explain how developments in technology have contributed to our understanding of the model of the atom. (SHE)</i> 26. <i>Research and communicate scientific ideas effectively in the format of a scientific report, including correct referencing of sources. (SIS)</i> 27. Describe the structure of the atom, including the locations of sub-atomic particles, holding the electrons to the nucleus. 28. Describe the properties of protons, neutrons and electrons, including relative mass and charge. 29. Use the periodic table to give an element's symbol, atomic number, relative atomic mass, number of protons, neutrons and electrons (neutral atom or monatomic ion). 30. Define what isotopes are, and describe the similarities and differences between isotopes of the same element. 31. Define what the relative atomic mass of an element is, and explain how this is different from mass number. 32. Describe how mass spectrometry can be used to establish the isotopic composition of an element. 33. Analyse mass spectra of elements to calculate the isotopic composition. 34. Calculate relative atomic mass from isotopic composition data. 35. Define the mole as a precisely defined quantity of matter, equal to Avogadro's number of particles. 36. Use Avogadro's number (see Data Booklet) to convert between number of particles and moles,	<u>Atomic Model</u> Theory <ul style="list-style-type: none"> Lucarelli Ch 2: Discovering Atomic Structure Set 3 Lucarelli Ch 1.11: Mass Spectrometry Set 2 Pearson Ch 2: Atoms: structure and mass STAWA Sets 8 – Atoms & Isotopes 12 – Research Qs 21 – Relative Atomic Mass

	<p>and vice versa.</p> <p>37. Use the relationship between mass, molar mass and number of moles to convert between moles and mass, and vice versa.</p> <p><u>Electron Configuration</u></p> <p>38. Explain that electrons exist in distinct energy levels.</p> <p>39. Use the energy level / shell model to write electron configurations for the first 20 elements, (eg: 2,8,8).</p> <p>40. Explain that atomic orbitals have different energy levels, and occupy different regions of space.</p> <p>41. Use the subshell (s/p/d/f) configurations to write electron arrangements for the first 20 elements (s & p only).</p> <p>42. Explain how the transition of electrons between energy levels can produce atomic absorption and emission line spectra.</p> <p>43. Describe and explain how flame tests and atomic absorption spectroscopy (AAS) can be used to identify elements.</p> <p>44. <i>Conduct experimental work safely, competently and methodically, including flame tests. (SIS)</i></p> <p>45. Identify the valence electrons of an atom, and explain why these are the electrons involved in bonding and chemical reactions.</p> <p>46. Explain how the stability of the valence electron shell determines the reactivity of an element.</p> <p><u>The Periodic Table</u></p> <p>47. Explain how the arrangement of the periodic table is based on the atomic number of the elements.</p> <p>48. Describe the relationship between position on the periodic table and electronic arrangement, including the relationship between group and number of valence electrons, and between period and number of occupied electron shells.</p> <p>49. Use electron configuration to identify an atom or ion and its location in the periodic table, and infer its chemical properties.</p> <p>50. Describe and explain the relationship between electronic structure and chemical properties and reactivity, for elements in groups 1, 2, 13 – 18.</p> <p>51. Explain the trends in reactivity for groups 1, 2, 16 & 17.</p> <p>52. Define the terms ionisation energy and electronegativity.</p> <p>53. Describe and explain the trends in first ionisation energy, atomic radius, valency & electronegativity in the periodic table, with reference to groups 1, 2, 13 – 18 and periods 1 – 3.</p>	<p><u>Electronic Configuration</u></p> <p>Theory</p> <ul style="list-style-type: none"> • Lucarelli Ch 7: Electronic Structure & Periodic Table Set 9 • Pearson Ch 3: Electrons and the periodic table <p>Practicals</p> <ul style="list-style-type: none"> • STAWA Exp 8 – Flame Tests • STAWA Exp 9 – Emission Spectra <p>STAWA Sets</p> <p>9 – Atomic Structure & Periodic Table</p> <p><u>The Periodic Table</u></p> <p>Theory</p> <ul style="list-style-type: none"> • Lucarelli Ch 9: Periodic Trends Set 15 • Pearson Ch 3: Electrons and the periodic table <p>STAWA Sets</p> <p>10 – Ionisation Energy</p> <p>11 – Periodic Trends</p>	
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<p>Term 1 Weeks 8 – 10</p>	<p>Topic 2: BONDING & INTERMOLECULAR FORCES</p> <p><u>Introduction to Bonding</u></p> <p>54. Describe the valency of an atom as a measure of the atom's bonding capacity, and use the periodic table to establish the valency of an atom for periods 1 to 3.</p> <p>55. Explain how the ability of atoms to form chemical bonds is related to its electronic structure, particularly the stability of the valence shell.</p> <p>56. Explain how electrostatic attraction between oppositely charged species / regions leads to chemical bonds.</p> <p>57. Use the periodic table to establish the type of bonding present in elements and compounds.</p> <p>58. Explain why the type of bonding present in a substance defines the physical properties of that substance, including melting and boiling points, conductivity of heat and electricity, and hardness.</p> <p><u>Ionic Bonding</u></p> <p>59. Define ions as atoms or groups of atoms that are electrically charged due to loss or gain of electrons.</p> <p>60. Describe and explain how ions are held together to form an ionic lattice.</p> <p>61. Use understanding of ionic bonding to explain the physical properties of ionic compounds, including:</p> <ol style="list-style-type: none"> High melting and boiling points Hardness and brittleness Conductivity in solid, liquid and aqueous states. <p>62. Calculate percentage composition of an ionic compound from the relative atomic masses of the constituent species.</p> <p>63. Define and describe the difference between anhydrous and hydrated salts, and explain the role of water of crystallisation.</p> <p>64. Calculate the formula of an ionic compound from empirical data, including the formula of a hydrated salt.</p> <p>65. Perform stoichiometric calculations in order to establish percentage composition and empirical formula of ionic compounds.</p> <p><u>Metallic Bonding</u></p> <p>66. Describe how metallic bonding holds metal atoms together, in a lattice of positively charged ions surrounded by delocalised electrons.</p> <p>67. Explain the physical properties of metals, including:</p> <ol style="list-style-type: none"> malleability and ductility thermal and electrical conductivity generally high melting and boiling points. 	<p><u>Introduction to Bonding</u></p> <p>Theory</p> <ul style="list-style-type: none"> Lucarelli Ch 4: Chemical Formulae Set 5 Lucarelli Ch 8.1: Classes of substance <p>Practicals</p> <ul style="list-style-type: none"> STAWA Inv 12: Physical properties & structure <p>STAWA Sets 24 – Interpretation of formulae</p> <p><u>Ionic Bonding</u></p> <p>Theory</p> <ul style="list-style-type: none"> Lucarelli Ch 8.3: Ionic Compounds Set 11 Pearson Ch 5: Ionic Bonding Pearson Ch 9.4: Percentage composition <p>Practicals</p> <ul style="list-style-type: none"> STAWA Exp 16: Empirical formula of magnesium oxide STAWA Exp 17 or Inv 18: Water of crystallisation (either) <p>STAWA Sets 25 – Percentage composition</p> <p><u>Metallic Bonding</u></p> <p>Theory</p> <ul style="list-style-type: none"> Lucarelli Ch 8.2: Metallic Substances Set 10 Pearson Ch 4: Metals 	
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Term 2 Week 1	<p><u>Covalent Bonding</u></p> <p>68. Describe a covalent bond as a shared pair of electrons, resulting in electrostatic attraction between positive nuclei of the atoms and the shared electron pair.</p> <p>69. Draw Lewis structures for covalent molecules, showing both bonding pairs and lone pairs of electrons.</p> <p>70. Use the valence shell electron pair repulsion (VSEPR) theory along with Lewis structures to explain, predict and draw the shapes of molecules and polyatomic ions.</p> <p>71. Explain the physical properties of covalent molecular substances, including:</p> <ol style="list-style-type: none"> low melting and boiling points non-conductivity softness and brittleness. <p>72. Recognise some common covalent network substances, including diamond, graphite and silicon dioxide.</p> <p>73. Describe the range of allotropes of elemental carbon (diamond, graphite and fullerenes), and explain why they have significantly different physical and chemical properties.</p> <p>74. Explain the physical properties of covalent network substances, including:</p> <ol style="list-style-type: none"> high melting and boiling points hardness and brittleness thermal conductivity electrical conductivity <p><u>Intermolecular Forces</u></p> <p>75. Describe the difference between intramolecular bonding and intermolecular forces.</p> <p>76. Compare the relative strength of ionic, covalent and metallic bonding with intermolecular forces.</p> <p>77. Recognise that intermolecular forces are present in all substances, but are most significant in discussion of the properties of covalent molecular substances, due to the lack of any stronger interactions between molecules.</p> <p>78. Describe how dispersion forces arise in all substances, and explain how molecular shape, and number of atoms affect the strength of dispersion forces.</p> <p>79. Use knowledge of electronegativity to establish bond polarity.</p> <p>80. Use understanding of molecular shape, bond polarity and symmetry to establish and explain the polarity of molecules.</p> <p>81. Describe how dipole-dipole forces arise, and explain how the relative size of a molecular dipole affects the strength of dipole-dipole forces.</p> <p>82. Describe how hydrogen bonding arises, and explain why this special case of dipole-dipole bonding is so much stronger than other examples of dipole-dipole bonding.</p> <p>83. Draw diagrams to show how all types of intermolecular forces occur between neighbouring molecules.</p>	<p><u>Covalent Bonding</u></p> <p>Theory</p> <ul style="list-style-type: none"> Lucarelli Ch 8.4 – 8.9: Covalent Bonding Sets 12 – 14 Pearson Ch 6: Materials made of molecules Pearson Ch 7.1: Carbon lattices <p>STAWA Sets</p> <p>14 – Bonding & properties 15 – Uses, properties & structure 16 – Properties & structures of materials 17 – Electron dot diagrams 18 – Molecular shape</p> <p><u>Intermolecular Forces</u></p> <p>Theory</p> <ul style="list-style-type: none"> Lucarelli Ch 15: Intermolecular Forces, including shape & polarity Set 24 Shape & polarity Set 25 Intermolecular forces Pearson Ch 12: Intermolecular forces <p>STAWA Sets</p> <p>19 – Intermolecular forces 1 20 – Intermolecular forces 2</p>	<p>CAP 2: Term 2 Week 1</p> <p><u>Test 1</u> Covers aspects of Topic 1 and Topic 2</p> <p>7.5% of year</p>
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	<p>84. Compare the relative strength of the types of intermolecular forces.</p> <p>85. Use understanding of bond polarity, molecular shape and symmetry to establish the nature and strength of intermolecular forces present in a covalent molecular substance.</p> <p>86. Explain how the physical properties of covalent molecular substances depend on the strength of intermolecular forces, including:</p> <ol style="list-style-type: none"> melting and boiling point vapour pressure solubility in water and organic solvents <p>87. Use understanding of intermolecular forces to explain patterns or relationships in melting and boiling points of covalent molecular substances.</p>		
<p>Term 2 Week 2</p>	<p><u>Chromatography</u></p> <p>88. Describe the general principles of chromatography, including the role of the stationary and mobile phases.</p> <p>89. Explain the importance of intermolecular forces in the chromatography process.</p> <p>90. Explain how chromatography can be used to separate and identify the components of a mixture.</p> <p>91. Use knowledge of the polarity of the components of a mixture to be separated to select a chromatographic method for separation.</p> <p>92. Apply the general principles of chromatography apply to the following techniques:</p> <ol style="list-style-type: none"> Paper chromatography Thin-layer chromatography (TLC) Gas chromatography (GC) High-performance liquid chromatography (HPLC) <p>93. <i>Analyse data from a range of chromatographic methods to determine the composition and purity of a substance. (SHE)</i></p>	<p><u>Chromatography</u></p> <p>Theory</p> <ul style="list-style-type: none"> Lucarelli Ch 19: Chromatography Set 33 Pearson Ch 13: Chromatography <p>Practicals</p> <ul style="list-style-type: none"> STAWA Exp 4 or Inv 5: Paper chromatography (either) 	
<p>Term 2 Weeks 3 – 4</p>	<p><u>Solubility</u></p> <p>94. Explain the unique properties of water using understanding of its molecular shape and ability to form hydrogen bonds, including:</p> <ol style="list-style-type: none"> anomalous melting and boiling points relative density of solid and liquid states surface tension, viscosity, cohesive and adhesive forces. <p>95. Classify solvents as polar or non-polar, and solutes as ionic, polar and non-polar.</p> <p>96. Identify and describe interactions between solvent and solute, and use this to predict whether a solution will form.</p> <p>97. Describe and explain how ion-dipole forces form in solutions, and use this to explain why ionic solutes are often soluble in water or other polar solvents.</p> <p>98. Describe solutions using the terms saturated, unsaturated or supersaturated, and use these ideas</p>	<p><u>Solubility</u></p> <p>Theory</p> <ul style="list-style-type: none"> Lucarelli Ch 16: Solutions, including precipitation & concentration Set 26 Electrolytes Set 27 Precipitation Set 28 Concentration Lucarelli Ch 6: Chemical Equations Set 8 Pearson Ch 15: Properties and uses of water Pearson Ch 16: Aqueous solutions 	

	<p>to describe the concentration of a solution.</p> <p>99. Describe the concentration of a substance using moles per litre of solution (mol L^{-1}), mass per litre of solution (g L^{-1}) or parts per million (ppm).</p> <p>100. Convert concentrations between different units (listed above).</p> <p>101. Describe and explain the relationship between solubility and temperature, for both solids and gases.</p> <p>102. Draw, interpret and analyse solubility graphs for solids and gases.</p> <p>103. Use understanding of solubility, along with the data tables, to predict whether solutions will form a precipitate.</p> <p>104. Write equations to show only the reacting species and products made in a reaction, including net ionic equations for precipitation reactions.</p> <p>105. Give observations for reactions involving solutions, solids and precipitates, with colours of solutions and solids, using the Data Booklet.</p> <p>106. <i>Conduct experimental work safely, competently and methodically to determine solubility of ionic compounds. (SIS)</i></p> <p>107. <i>State that potable drinking water must undergo various treatment processes to ensure it meets regulations for safe levels of solutes, including heavy metals. (SHE)</i></p> <p>108. <i>Explain why drinking water must be monitored for safe levels on solutes. (SHE)</i></p>	<p>Practicals</p> <ul style="list-style-type: none"> STAWA Exp 14: Bonding & solubility STAWA Exp 22: Solubility rules <p>STAWA Sets</p> <p>31 – Solutions</p> <p>32 – Solution Concentrations</p>	
Term 2 Weeks 5 – 6	<p>SEMESTER 1 EXAM</p> <p>Covers Topics 1 & 2 (spec points 1 – 108)</p>		Worth 20% of year
Term 2 Weeks 7 – 9	<p>Topic 3: ACIDS & STOICHIOMETRY</p> <p><u>Acids & Bases</u></p> <p>109. Use the Arrhenius model to define and describe acids and bases in aqueous solution.</p> <p>110. Describe the common chemical and physical properties of acids and bases.</p> <p>111. State that indicators can be used to identify acids and bases in solution, and use observations of indicator colour to classify solutions as acidic, basic or neutral.</p> <p>112. Write equations to show how acids form H^+ ions and bases form OH^- ions in solution.</p> <p>113. Distinguish between the processes of ionisation, where a covalent molecule splits into ions, and dissociation, where an ionic compound dissolves in water.</p> <p>114. Describe and explain the difference between strong and weak acids and bases, using discussion of the degree of ionisation or dissociation.</p> <p>115. Write equations to show the difference between strong and weak acids and bases.</p> <p>116. Identify some common strong acids, including HCl, HNO_3, and H_2SO_4, and some common weak acids, including CH_3COOH, H_3PO_4 and other organic acids.</p>	<p><u>Acids & Bases</u></p> <p>Theory</p> <ul style="list-style-type: none"> Lucarelli Ch 17: Acids & Bases Sets 29 & 30 Pearson Ch 17: Acids and Bases <p>Practicals</p> <ul style="list-style-type: none"> STAWA Exp 29: Electrical conductivity of acids & bases STAWA Inv 30: Acid or base? Strong or weak? STAWA Exp 31: Acids & metal compounds (see assessments) STAWA Exp 32: Acids & metals (see assessments) 	<p>In Class Task: Term 2 Week 9/10</p> <p><u>Identification of Mystery Substances</u></p> <p>Practical investigation completed in lab</p> <p>5% of year</p>

Term 2 Week 10	<p>117. Identify some common strong bases, including group 1 and 2 oxides and hydroxides, and some common weak bases, including NH_3 and Na_2CO_3.</p> <p>118. Differentiate between the terms strong and concentrated when used in reference to acids and bases.</p> <p>119. Describe and give observations for the reactions of acids and bases, including the reactions between:</p> <ol style="list-style-type: none"> acids and metals acids and metal oxides acids and metal hydroxides acids and metal carbonates and metal hydrogencarbonates bases and ammonium salts bases and non-metal oxides <p>120. Write word equations, molecular equations and ionic equations for the reactions listed above.</p> <p>121. Predict observations for the reactions listed above, clearly describing the reactants and products, any colours involved (see Data Booklet) and any changes observed.</p> <p>122. Define the pH scale as a measure of hydrogen ion concentration, and state that it is an inverse logarithmic scale.</p> <p>123. Use pH values to classify a solution as acidic, basic or neutral.</p> <p>124. Describe the relationship between pH, acidity and alkalinity of a solution.</p> <p>125. Define pH as $-\log_{10}[\text{H}^+(\text{aq})]$, and use this relationship to calculate pH from hydrogen ion concentration, and vice versa.</p> <p>126. Define the ionic product of water, K_w, as the product of the hydrogen ion and hydroxide ion concentrations, and state that $K_w = 1 \times 10^{-14}$ at 298 K.</p> <p>127. State that a neutral solution is one where H^+ ion and OH^- ion concentration are equal.</p> <p>128. State that the concentrations of H^+ and OH^- are both $1 \times 10^{-7} \text{ mol L}^{-1}$ at 25 °C.</p> <p>129. Use K_w and the pH expression to calculate the pH of a solution of a strong base.</p> <p>130. Use the pH of a strong base to calculate the OH^- concentration, H^+ concentration and pH of the solution.</p> <p>131. Calculate the pH of a solution formed by mixing solutions of acids and / or bases.</p> <p><u>Stoichiometry</u></p> <p>132. State that the molar volume of all gases is 22.71 L at STP (0°C and 100 kPa), and use this to calculate the volumes or amount in moles of gases at STP.</p> <p>133. Use the mole concept and the molar volume to calculate quantities of gases as reactants of products in a chemical reaction at STP.</p> <p>134. Perform stoichiometric calculations in a range of contexts from across year 11, including limiting reagent calculations.</p>	<p>STAWA Sets</p> <p>27 – Ionic equations</p> <p>28 – Equations & observations</p> <p>34 – The pH scale</p> <p>35 – Solutions of acids & bases</p> <p><u>Stoichiometry</u></p> <p>Theory</p> <ul style="list-style-type: none"> Lucarelli Ch 10.6: Measuring Gases Set 16 Lucarelli Ch 11: Stoichiometry Sets 17 – 19 Pearson Ch 14.2 – 14.3: Molar volume & calcs Lucarelli Ch 18: More on Stoichiometry Sets 31 & 32 <p>STAWA Sets</p> <p>26 – Gas Volumes</p> <p>29 – Stoichiometry</p> <p>30 – Stoichiometry & Gas Volumes</p> <p>33 – Reacting mass, gases and solutions</p> <p>36 – Acid & base stoichiometry</p> <p>37 – Aqueous solutions of acids & bases</p>	<p>CAP 3:</p> <p>Term 3 Week 2</p> <p><u>Reactions of Acids</u></p> <p>Practical validation based on Exp 31 & 32</p> <p>5% of year</p> <p><u>Aqueous Environment</u></p> <p>Extended response report set in Term 2 Week 9/10</p> <p>Validation in Term 3 Week 2</p> <p>5% of year</p>
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Term 3 Weeks 1 – 3	<p>Topic 4: CARBON CHEMISTRY & ENERGY</p> <p><u>Structure & Naming</u></p> <p>135. Explain why carbon is able to form such a diverse range of compounds, using knowledge of bonding capacity.</p> <p>136. Use molecular formulae, full structural formulae and condensed formulae to show the arrangement of atoms and bonds in covalent molecular substances, including hydrocarbons and haloalkanes.</p> <p>137. Define hydrocarbon, and categorise simple hydrocarbons into alkanes, alkenes and aromatic hydrocarbons (benzene).</p> <p>138. Use the general formulae for alkanes and alkenes to recognise and categorise hydrocarbons.</p> <p>139. Recognise cycloalkanes and cycloalkenes, and understand that they have different general formulae from their non-cyclic counterparts.</p> <p>140. Describe hydrocarbons as saturated or unsaturated, and explain the difference between these terms.</p> <p>141. Use IUPAC nomenclature to name straight and simple branched chain alkanes and alkenes (C₁ – C₈, one double bond per molecule).</p> <p>142. Use IUPAC nomenclature to name straight chained haloalkanes (C₁ – C₈).</p> <p>143. Draw and name structural isomers of alkanes and alkenes, specifically chain isomers and position isomers.</p> <p>144. Draw and name geometric isomers of alkenes, using cis and trans notation.</p> <p>145. Describe the nature of bonding in alkanes, alkenes and benzene, using ideas of single and multiple bonds (π bonds), and delocalisation.</p> <p>146. Explain why alkenes are able to form geometric isomers, whereas alkanes are not.</p> <p><u>Reactions of Hydrocarbons</u></p> <p>147. Explain the difference in chemical reactivity of alkanes, alkenes and benzene using discussion of bonding.</p> <p>148. Write equations, and draw and name the products for substitution reactions of alkanes and benzene with halogens.</p> <p>149. Explain why substitution reactions of alkanes require visible / UV light to take place.</p> <p>150. Recognise that substitution reactions of benzene require a catalyst.</p> <p>151. Identify, draw and name the products from single and multiple substitution reactions.</p> <p>152. Write equations, and draw and name the products of addition reactions of alkenes, with halogens, hydrogen halides and hydrogen.</p> <p>153. Describe how aqueous solutions of bromine and iodine can be used to test for unsaturation, and give observable outcomes.</p> <p>154. Write equations, and draw and name the products of complete combustion reactions of</p>	<p><u>Carbon Chemistry</u></p> <p>Theory</p> <ul style="list-style-type: none"> • Lucarelli Ch 12: Hydrocarbons Set 20 – Naming & drawing Set 21 – Reactions • Lucarelli Ch 13.8 – 13.12: Comparing Fuels Set 22 Qs 9 – 15 • Pearson Ch 8: Organic compounds <p>Practicals</p> <ul style="list-style-type: none"> • STAWA Exp 44: Molecular models (molymod) • STAWA Inv 46: Reactivity of hydrocarbons <p>STAWA Sets</p> <p>40 – Naming & drawing hydrocarbons 41 – Hydrocarbon reactions 42 – Organic chemistry</p>	
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	<p>hydrocarbons.</p> <p>155. Given information about the products, write equations for the incomplete combustion reactions of hydrocarbons.</p> <p>156. Explain why incomplete combustion is undesirable when using hydrocarbons as fuels.</p> <p>157. <i>Discuss the benefits and drawbacks of using hydrocarbons as fuels, including the effect on the environment. (SHE)</i></p> <p>158. <i>Compare the effectiveness and suitability of different fuels, including hydrocarbons and biofuels (biogas, biodiesel, bioethanol), and suggest why certain fuels are used. (SHE)</i></p> <p>159. Calculate percentage composition of a hydrocarbon from the relative atomic masses of the constituent species.</p> <p>160. Use the mole concept and the law of conservation of mass to calculate quantities of reactants and products in a chemical reaction.</p> <p>161. Identify and calculate empirical formulae for hydrocarbons using percentage composition data and combustion analysis data, and relate this to the molecular formulae.</p> <p>162. Use molecular mass and empirical formulae calculations to establish molecular structure.</p>		
<p>Term 3 Weeks 4 – 6</p>	<p>Topic 5: ENERGY & RATE</p> <p><u>Energy & Enthalpy</u></p> <p>163. State that kinetic theory is used to explain the macroscopic behaviour of gases using understanding of the molecular behaviour of particles.</p> <p>164. Describe the properties of an ideal gas, and describe the differences between an ideal gas and a real gas.</p> <p>165. Use kinetic theory to explain the behaviour of ideal gases, including diffusion and compressibility.</p> <p>166. Describe the relationship between temperature, kinetic energy and velocity of particles in a qualitative fashion.</p> <p>167. Explain the origin of the Kelvin temperature scale, and state the meaning of the term absolute zero.</p> <p>168. Convert temperatures between Celsius and Kelvin.</p> <p>169. Use kinetic theory to describe and explain the relationships between pressure, temperature and volume of an ideal gas in a qualitative fashion, (Boyle's Law, Charles' Law, and Pressure Law).</p> <p>170. Draw and interpret graphs that represent the relationship between pressure, temperature and volume.</p> <p>171. State that the relationships between pressure, temperature, volume and number of moles can be combined to give the ideal gas equation, $PV = nRT$ (Use of equation not required).</p> <p>172. Describe the internal energy of a system as the sum of the kinetic energy and the potential energy of the particles.</p>	<p><u>Energy & Enthalpy</u></p> <p>Theory</p> <ul style="list-style-type: none"> • Lucarelli Ch 1.1 – 1.6: Kinetic Theory & Energy Set 1 • Lucarelli Ch 13.1 – 13.7: Energy & Chemical Change Set 22 Qs 1-8 • Pearson Ch 14.1: Kinetic theory of gases • Pearson Ch 10: Energy changes in chemical reactions • Pearson Ch 11: Fuels • Lucarelli Ch 5: Investigating in Chemistry Sets 6 & 7 • Pearson Ch 20: SIS <p>Practicals</p> <ul style="list-style-type: none"> • Enthalpy change of zinc & copper sulphate Prac (see shared folder) 	

	<p>173. Describe the enthalpy of a system as a measure of the energy stored within a system, and state that it cannot be measured directly.</p> <p>174. Describe the enthalpy change of a system as the heat energy exchange with the surroundings.</p> <p>175. State that enthalpy and enthalpy change values are given at STP conditions.</p> <p>176. Describe and explain the enthalpy changes in a system in terms of stored (chemical potential) energy of reactants and products, and energy transfer in the form of heat, using the Law of Conservation of Energy.</p> <p>177. Describe observable changes in temperature, and use this evidence to establish whether a process is exothermic or endothermic.</p> <p>178. Explain the enthalpy changes in exothermic and endothermic processes in terms of energy input for breaking of existing bonds, and energy output from the formation of new bonds.</p> <p>179. Use thermochemical equations to represent the enthalpy changes in chemical reactions.</p> <p>180. Draw and interpret enthalpy profile diagrams for exothermic and endothermic reactions using a quantitative scale for enthalpy, including accurate labels for reactants and products, activation energy, enthalpy change and transition state / activated complex.</p> <p>181. <i>Generate experimental temperature data from a reaction, to calculate the enthalpy change per mole of reaction (as stated in the chemical equation), given $\Delta H = mc\Delta T$, (c = specific heat capacity of solution, assumed to be $4.18 \text{ JK}^{-1}\text{mol}^{-1}$). (SIS)</i></p> <p>182. <i>Evaluate outcomes from experiments involving enthalpy changes, identify sources of experimental error, and discuss the effect on enthalpy change values calculated. (SIS)</i></p> <p>183. <i>Conduct an investigation into choice of fuels for a purpose; propose hypotheses, predict outcomes, collect reliable and valid data, from which conclusions can be drawn. (SIS)</i></p> <p>184. Use enthalpy change data to compare the energy content or density of fuels, including biofuels.</p>	<p>STAWA Sets</p> <p>3 – Random & systematic errors</p> <p>5 – Kinetic Theory</p> <p>6 – Properties of Matter</p> <p>38 – Energy Changes</p>	<p>CAP 4:</p> <p>Term 3 Week 9</p> <p><u>Properties of Hydrocarbons</u></p> <p>Validation based on practical aspects of carbon chemistry and energy topic</p> <p>5% of year</p> <p><u>Rates of Reaction</u></p> <p>Practicals conducted in class week 7/8</p> <p>Validation based on lab work</p> <p>5% of year</p>
<p>Term 3</p> <p>Weeks</p> <p>7 – 10</p>	<p><u>Rate of Reaction</u></p> <p>185. State that the rate of a chemical reaction can be quantified by measuring the rate of change in an observable quantity that indicates amount of reactant used up or product made.</p> <p>186. Describe a range of methods to measure the rate of a reaction, including volume of gas produced, mass lost, colour change and changes to transparency.</p> <p>187. Describe how to change the rate of a chemical reaction, using changes to reactant concentration, gas pressure, temperature, surface area of solid reactant, or use of a catalyst.</p> <p>188. State that a catalyst is a substance that speeds up a chemical reaction without being consumed.</p> <p>189. <i>Conduct experimental work safely, competently and methodically in order to collect valid and reliable data to measure the rate of reaction. (SIS)</i></p> <p>190. <i>Represent rate of reaction data in tables and graphs, using correct units and symbols, in order to identify trends and any anomalous data. (SIS)</i></p> <p>191. <i>Evaluate the effect of measurement error in numerical data. (SIS)</i></p> <p>192. Explain the factors that affect the rate of reaction using collision theory.</p>	<p><u>Rate of Reaction</u></p> <p>Theory</p> <ul style="list-style-type: none"> • Lucarelli Ch 14: Reaction Rate Set 23 • Pearson Ch 18: Rates of chemical reactions • Pearson Ch 19: Catalysts <p>Practicals</p> <ul style="list-style-type: none"> • STAWA Exp 39: Concentration • STAWA Exp 40: Temperature • STAWA Exp 41: Catalysts <p>STAWA Sets</p> <p>39 – Rate of Reaction</p>	

	<p>193. Predict the outcome on the rate of reaction when conditions for a chemical reaction are altered using collision theory.</p> <p>194. State that the activation energy for a reaction is the minimum amount of energy required for particles to react on collision.</p> <p>195. Describe how the size of the activation energy affects the rate of reaction.</p> <p>196. Explain how the number and strength of bonds to be broken affects the size of the activation energy.</p> <p>197. Describe and explain the effect of using a catalyst on the rate of reaction using collision theory.</p> <p>198. Draw and interpret energy profile diagrams for catalysed and uncatalysed reactions.</p> <p>199. <i>Justify the use of catalysts in industry, to increase the rate of a reaction that would otherwise be uneconomically slow. (SHE)</i></p> <p>200. <i>Describe how catalysts can be used to reduce negative impact on the environment, for example in catalytic converters. (SHE)</i></p>		<p>CAP 5: Term 4 Week 1</p> <p><u>Test 2</u> Covers aspects of Semester 2 Topics</p> <p>7.5% of year</p>
Term 4 Weeks 3 & 4	<p>SEMESTER 2 EXAM Covers all of year 11 content.</p>		Worth 30% of year
Term 4 Weeks 5 – 8	<p>TRANSITION TO YEAR 12 See Year 12 Program</p>		

Appendix 1: Ions & Molecules

Ion name	Formula
ammonium	NH_4^+
caesium	Cs^+
hydrogen	H^+
lithium	Li^+
potassium	K^+
rubidium	Rb^+
silver	Ag^+
sodium	Na^+
barium	Ba^{2+}
calcium	Ca^{2+}
cobalt(II)	Co^{2+}
copper(II)	Cu^{2+}
iron(II)	Fe^{2+}
lead(II)	Pb^{2+}
magnesium	Mg^{2+}
manganese(II)	Mn^{2+}
nickel(II)	Ni^{2+}
strontium	Sr^{2+}
zinc	Zn^{2+}
aluminium	Al^{3+}
chromium(III)	Cr^{3+}
iron(III)	Fe^{3+}

Students should be able to recognise and write the formula of the following ions and molecules.

Ion Name	Formula
bromide	Br^-
chloride	Cl^-
cyanide	CN^-
dihydrogenphosphate	H_2PO_4^-
ethanoate (acetate)	CH_3COO^-
fluoride	F^-
hydrogencarbonate	HCO_3^-
hydrogensulfate	HSO_4^-
hydroxide	OH^-
iodide	I^-
nitrate	NO_3^-
nitrite	NO_2^-
permanganate	MnO_4^-
carbonate	CO_3^{2-}
chromate	CrO_4^{2-}
dichromate	$\text{Cr}_2\text{O}_7^{2-}$
hydrogenphosphate	HPO_4^{2-}
oxalate	$\text{C}_2\text{O}_4^{2-}$
oxide	O^{2-}
sulfate	SO_4^{2-}
sulfide	S^{2-}
sulfite	SO_3^{2-}
nitride	N^{3-}
phosphate	PO_4^{3-}

Molecular substance	Formula
ammonia	NH_3
water	H_2O
hydrogen peroxide	H_2O_2
ethanoic acid	CH_3COOH
hydrochloric acid	HCl
nitric acid	HNO_3
carbonic acid	H_2CO_3
sulfuric acid	H_2SO_4
sulfurous acid	H_2SO_3
phosphoric acid	H_3PO_4