

**Combined Science**

Biology, Chemistry and Physics – Paper 1 – 1 hour 15 minute paper each

**Triple Science**

Biology, Chemistry and Physics – Paper 1 – 1 hour 45 minute paper each

**Topics for both Combined and Triple Science**

**Biology**

Cell Biology

Organisation

Infection and Response

Bioenergetics

**Chemistry**

Atomic Structure & the Periodic Table

Bonding, Structure & Properties of Matter

Quantitative Chemistry

Chemical Changes

Energy Changes

**Physics**

Energy

Electricity

Particle Model of Matter

Atomic Structure

**Useful Revision Resources –**

- Fact sheets for recall of factual content (behind this summary)
  - <https://www.aqa.org.uk/subjects/science> - Syllabus information & past papers with mark schemes
  - <https://www.youtube.com/@Freesciencelessons> – excellent topic summaries presented as short videos for all Science content.
  - <https://www.physicsandmathstutor.com/> - revision resources & past paper questions and mark schemes – past paper questions are arranged by topic which is useful for revision. Covers all science content.
  - <https://www.kerboodle.com/users/login> - all students have an individual log in – can view an electronic copy of the textbook and various revision resources.
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Atomic Structure	Elements, Compounds and mixtures	<ol style="list-style-type: none"> <li>1. An element is a substance which contains only one TYPE of atom.</li> <li>2. A compound is a substance which contains two or more types of atom bonded together.</li> <li>3. Mixtures contain different elements or compounds that can be separated as they are not chemically bonded together.</li> <li>4. In chemical reactions the starting materials are called reactants and new products are made.</li> <li>5. There are 4 state symbols;           (s) = solid                                                          (l) = liquid                                                          (g) = gas                                                          (ag) = aqueous (dissolved in solution)</li> <li>6. You separate an insoluble solid from a solution by filtering it out. The solid can then be washed and dried to remove any impurities.</li> <li>7. To separate a salt from a solution you evaporate the water to produce crystals of salt.</li> <li>8. To separate and collect a liquid from a mixture you use distillation. You can use distillation to separate a mixture of liquids.</li> </ol>
	Atomic Structure	<ol style="list-style-type: none"> <li>9. There are 3 subatomic particles; protons, electrons and neutrons.</li> <li>10. Protons are positive (relative charge is +1) and have a relative mass of 1.</li> <li>11. Electrons are negative (relative charge is -1) and their relative mass is very small.</li> <li>12. Neutrons are neutral (relative charge is 0) and have a relative mass of 1.</li> <li>13. Protons and neutrons are found in the nucleus and electrons orbit the nucleus.</li> <li>14. Atoms have no overall charge because the number of positive protons is equal to the number of negative electrons.</li> <li>15. The atomic number is the number of protons</li> <li>16. The mass number is the number of protons and neutrons in total.</li> <li>17. Atoms of the same element can have different numbers of neutrons; these atoms are called isotopes.</li> <li>18. Electrons are arranged in energy levels (shells). The lowest energy level (shell) can hold a maximum of 2 electrons, the second can hold 8 and the third can hold 8.</li> <li>19. Electrons occupy the lowest available energy level.</li> <li>20. The electronic structure can be shown as a diagram or as numbers. Eg for sodium that has 11 electrons, the electronic structure is 2,8,1</li> </ol>
	Development of model of atom	<ol style="list-style-type: none"> <li>21. Dalton thought atoms were hard spheres and that elements had only one type of atom.</li> <li>22. J.J. Thompson discovered the electron (tiny negatively charged particle) and described atoms like 'plum puddings' with negative charges embedded in a cloud of positive charge.</li> <li>23. Geiger and Marsden's did experiments firing positive alpha particles at gold foil which showed atoms could not be solid.</li> <li>24. Rutherford proposed that the positive charge of an atom is in a small centre which he called the nucleus and electrons orbit this nucleus. ( This is the nuclear model)</li> <li>25. In 1914 Niels Bohr suggested electrons orbit at set distances in energy levels ( or shells).</li> <li>26. In 1932 James Chadwick discovered the neutron which has no charge and has the same mass as a proton</li> </ol>

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The Periodic Table	Development of periodic table	<p>27. Before the discovery of protons, electrons and neutrons, scientists ordered the elements by their atomic weight.</p> <p>28. Mendeleev placed elements in more appropriate places, grouping elements with similar properties so patterns could be seen. This meant he sometimes left gaps or changed the order of atomic weight.</p> <p>29. Elements with properties Mendeleev predicted were discovered and filled the gaps.</p> <p>30. The elements on the modern periodic table are arranged in order of increasing atomic number.</p> <p>31. Elements are arranged in columns called groups (going down) on the periodic table.</p> <p>32. The number of electrons in the outermost shell of an atom is the same as its group on the periodic table.</p>
	Group 0	<p>33. Elements in the same group have similar properties.</p> <p>34. Group 0 are called the noble gases. They are unreactive because of their very stable electron arrangement ( full out shell)</p> <p>35. The boiling point of the noble gases increases with increasing relative atomic mass (as you go down the group).</p> <p>36. Boiling point and condensing point are the same temperature.</p>
	Group 1	<p>37. Group 1 are the alkali metals.</p> <p>38. All group 1 atoms have 1 electron in their outer energy level.</p> <p>39. Group 1 metals are very reactive.</p> <p>40. When they react they lose their outer electron to form a full outer energy level and become stable.</p> <p>41. Group 1 metals are stored in oil to stop the oxidising.</p> <p>42. They have low densities and float in water.</p> <p>43. Group 1 metals react vigorously with water producing an alkaline solution of the metal hydroxide plus hydrogen gas.</p> <p>44. Group 1 metals react with halogen to produce metal halides which are white soluble solids.</p> <p>45. The reactivity of group 1 increases as you go down the group.</p>
	Group 7	<p>46. Group 7 elements are known as the halogens.</p> <p>47. Group 7 elements all react in a similar way as they all have 7 electrons in their outer shell.</p> <p>48. Group 7 elements are non- metals and consist of molecules made of pairs of atoms eg Br<sub>2</sub></p> <p>49. The melting point and boiling point of group 7 increase as you go down the group.</p> <p>50. In group 7 the reactivity decreases as you go down the group.</p> <p>51. The halogens form ions with a charge of 1- by gaining an electron when reacting with metals to form ionic compounds.</p> <p>52. A more reactive halogen will displace a less reactive halogen from a compound.</p> <p>53. The halogens form covalent molecules by sharing electrons with other non- metals.</p>

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Bonding and Properties of Matter	Metals	<p>54. Metals are found in the centre of the periodic table and to the left and bottom of the periodic table.</p> <p>55. Metals react to form positive ions.</p> <p>56. The atoms in metals are closely packed together in regular layers</p> <p>57. The electrons in the outer shells of metals are delocalised and are free to move throughout the metallic lattice. This creates strong metallic bonding.</p> <p>58. Most metals have high melting and boiling points.</p> <p>59. Metals can be bent and shaped as the layers of atoms can slide over each other.</p> <p>60. Alloys are mixtures of a metal with other elements.</p> <p>61. Alloys are harder than pure metals as the layers are distorted and cannot slide.</p> <p>62. Metals are good conductors of electricity because the delocalised electrons move through the structure.</p> <p>63. Metals are good conductors of heat as the delocalised electrons can transfer the thermal energy.</p>
	Ionic Bonding	<p>64. Ionic bonding occurs between metals and non- metals.</p> <p>65. Metal atoms lose electrons to form positive ions.</p> <p>66. Non- metal atoms tend to gain electrons to form negative ions.</p> <p>67. Ionic compounds are held together by strong electrostatic forces of attraction between oppositely charged ions.</p> <p>68. Ionic compounds form giant lattices.</p> <p>69. Ionic compounds have high melting and boiling points.</p> <p>70. Ionic compounds do not conduct electricity when solid as the ions are not free to move.</p> <p>71. When melted or dissolved, ionic compounds conduct electricity as the ions are free to move.</p>
	Covalent Bonding	<p>72. Covalent bonds are formed when non-metal atoms share pairs of electrons.</p> <p>73. Covalent bonds are strong.</p> <p>74. Many covalent compounds consist of small, simple molecules e.g., oxygen, chlorine and water.</p> <p>75. They are usually gases or liquids with low melting and boiling points.</p> <p>76. Small covalent molecules have weak intermolecular forces between the molecules.</p> <p>77. When these substances are heated it is the intermolecular forces that break not the covalent bonds.</p> <p>78. Simple covalent molecules do not conduct electricity because the molecules do not have an overall charge.</p> <p>79. Polymers are very large molecules held together by strong covalent bonds.</p> <p>80. Some covalently bonded substances have giant structures eg silicon dioxide , diamond and graphite.</p> <p>81. Giant covalent structures have very high melting and boiling points.</p>

Quantitative chemistry	Structure and bonding of carbon	<p>82. The element carbon has different forms in which the atoms are arranged in different ways. These are diamond, graphite, graphene and fullerenes.</p> <p>83. In diamond each carbon atom forms 4 covalent bonds with other carbon atoms.</p> <p>84. Diamond is very hard as it is held together by strong covalent bonds.</p> <p>85. Diamond does not conduct electricity as it has no free (delocalised) electrons.</p> <p>86. Graphite has a structure in which the carbon atoms form layers of hexagonal rings in which each carbon atom is bonded to 3 other carbon atoms.</p> <p>87. The layers of atoms in graphite are not bonded together and can slide over each other which makes graphite soft and slippery.</p> <p>88. Graphite can conduct electricity as one electron from each carbon atom is delocalised.</p> <p>89. Graphite is a good thermal conductor because of the delocalised electrons which can move through the layers.</p> <p>90. Graphene is a single layer of graphite so is just one atom thick.</p> <p>91. Graphene conducts electricity and is very strong.</p> <p>92. Fullerenes are molecules of carbon atoms with hollow shapes.</p> <p>93. Buckminster fullerene (bucky ball) has a formula of <math>C_{60}</math></p> <p>94. Carbon nanotubes have very high length to diameter ration.</p> <p>95. Fullerenes are finding uses eg to deliver drugs to specific body parts, as catalysts and as reinforcement for composite materials.</p> <p>96. Graphene will help create new developments in the electronics industry in the future.</p>
	Masses in reactions	<p>97. The law of conservation of mass states that no atoms are lost or made during reactions so the mass of the products equals the mass of the reactants.</p> <p>98. The relative formula mass of a compound is the sum of the relative atomic masses of all the atoms shown in the formula.</p> <p>99. In a balanced equation the sum of the relative formula masses of all the reactants = the sum of the relative formula masses of all the products.</p> <p>100. <i>A mole of a substance is its relative atomic mass or relative formula mass in grams.</i></p> <p>101. <i>The number of atoms, molecules or ions in a mole of a substance is always <math>6.02 \times 10^{23}</math>. This number is called the Avogadro constant.</i></p> <p>102. <i>The number of moles = mass (g) / <math>A_r</math> or mass (g) / <math>M_r</math></i></p> <p>103. <i>Balanced symbol equations show you the number of moles of each substance involved in the reaction.</i></p> <p>104. <i>The reactant that gets used up first in a reaction is the limiting reactant. This is the reactant that is NOT in excess.</i></p>
Concentration	<p>105. The concentration of a solution can be measured in mass per volume. Grams per <math>dm^3</math> (<math>g/dm^3</math>)</p> <p>106. <math>1 dm^3 = 1000cm^3</math></p> <p>107. The mass of solute in a solution = concentration x volume in <math>dm^3</math></p> <p>108. <i>The concentration of a solution can be increased by increasing the mass of solute or decreasing the volume of the solvent.</i></p>	

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Reactivity of metals	Reaction of metals	<p>109. Metals react with oxygen to produce metal oxides. These reactions are called oxidation reactions.</p> <p>110. When metals react they lose electrons to form positive ions</p> <p>111. The reactivity series places metals in order of reactivity. From the most to the least reactive it is; potassium, sodium, lithium, calcium, magnesium, aluminium, (carbon) zinc, iron, tin, lead, (hydrogen), copper, silver, gold.</p> <p>112. A more reactive metal can displace a less reactive metal from a compound.</p> <p>113. Metals more reactive than hydrogen will react with acid and produce hydrogen gas. Hydrogen gas 'pops' with a lighted spill.</p> <p>114. Metals less reactive than carbon can be extracted from their ores by reduction with carbon.</p> <p>115. Reduction involves the loss of oxygen.</p> <p>116. <i>Reduction is the gain of electron.</i></p> <p>117. <i>Oxidation is the loss of electrons.</i></p>
Reactions of acids	Making salts	<p>118. Acids react with some metals to produce salts plus hydrogen.</p> <p>119. Sulphuric acid produces salts called sulphates.</p> <p>120. Hydrochloric acid produces salts called chlorides.</p> <p>121. Nitric acid produces salts called nitrates.</p> <p>122. Alkalis are soluble metal hydroxides ( eg sodium hydroxide)</p> <p>123. Bases are insoluble metal hydroxides and metal oxides.</p> <p>124. Acids are neutralised by alkalis and bases to produce a salt plus water.</p> <p>125. Acids are neutralised by metal carbonates to produce a salt, water and carbon dioxide.</p> <p>126. When reacting an insoluble base with an acid to make a soluble salt, excess solid is used then filtered off after the reaction.</p> <p>127. Salt solutions can be crystallised to produce solid salts.</p>
	pH scale	<p>128. Acids produce hydrogen ions (H<sup>+</sup>) in aqueous solutions.</p> <p>129. Aqueous solutions of alkalis contain hydroxide ions (OH<sup>-</sup>).</p> <p>130. The pH scale. From 0-14, is a measure of the acidity or alkalinity of a solution.</p> <p>131. pH can be measured using universal indicator or a pH probe.</p> <p>132. A solution of pH 7 is neutral.</p> <p>133. Acids have a pH less than 7.</p> <p>134. Alkalis have a pH greater than 7.</p> <p>135. In reactions between acids and alkalis, the hydrogen ions and the hydroxide ions neutralise each other.</p> <p>136. <math>\text{H}^+_{(\text{aq})} + \text{OH}^-_{(\text{aq})} \rightarrow \text{H}_2\text{O}_{(\text{l})}</math></p>
Strong and weak acids	Strong and weak acids	<p>137. <i>A strong acid is completely ionised in aqueous solutions.</i></p> <p>138. <i>Hydrochloric, nitric and sulphuric acids are strong acids.</i></p> <p>139. <i>Weak acids only partially ionise in aqueous solutions. Ethanoic, citric and carbonic acids are weak acids.</i></p> <p>140. <i>The stronger the acid the lower the pH.</i></p> <p>141. <i>As the pH decreases by one unit, the hydrogen ion concentration increases by a factor of 10.</i></p> <p>142. <i>pH 1 = hydrogen ion concentration of 0.1 mol/dm<sup>3</sup>, pH 2 = 0.01mol/dm<sup>3</sup> and so on</i></p>

Electrolysis	Electrolysis of metal compounds	<p>143. Electrolysis breaks down a substance using electricity.</p> <p>144. Metals can be extracted from molten compounds using electrolysis.</p> <p>145. The ions in ionic compounds are free to move when they are melted or dissolved. They can then conduct electricity.</p> <p>146. Liquids and solutions that conduct electricity are called electrolytes.</p> <p>147. When electricity is passed through an electrolyte, positive ions move to the negative electrode and negative ions move to the positive electrode.</p> <p>148. The positive electrode is called the anode.</p> <p>149. The negative electrode is called the cathode.</p> <p>150. Ions are discharged at the electrodes to produce elements.</p> <p>151. Metal ions are attracted to the cathode and metal is produced.</p> <p>152. When extracting metals from their compounds, inert (unreactive) electrodes are used.</p> <p>153. Electrolysis uses large amounts of energy to melt the compounds and to produce the electrical current used.</p> <p>154. In the production of aluminium, aluminium oxide is mixed with cryolite to lower the melting point.</p> <p>155. The oxygen produced at the anode when aluminium oxide is electrolysed reacts with the anode to form carbon dioxide. This electrode needs replacing.</p>
	Electrolysis of aqueous solutions	<p>156. The ions discharged when an aqueous solution is electrolysed using inert electrodes depend on the relative reactivity of the elements involved.</p> <p>157. The water in the solution is broken down by the electricity into hydrogen ions (H<sup>+</sup>) and hydroxide ions (OH<sup>-</sup>).</p> <p>158. At the negative electrode (cathode), hydrogen is produced if the metal is more reactive than hydrogen.</p> <p>159. At the positive electrode, oxygen is produced unless the solution contains halide ions in which case the halogen is produced.</p> <p>160. In the electrolysis of salt solution (brine), hydrogen is produced at the cathode, chlorine is produced at the anode and sodium hydroxide solution is also formed.</p> <p>161. <i>During electrolysis, positively charged ions gain electrons at the cathode. This is known as reduction.</i></p> <p>162. <i>At the anode, negatively charged ions lose electrons and so these reactions are oxidations.</i></p> <p>163. <i>Reactions at the electrodes can be represented by half equations, e.g.</i>  <math display="block">2\text{H}^+ + 2\text{e}^- \Rightarrow \text{H}_2</math> </p>

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Energy changes in reactions	Exothermic and Endothermic reactions	<p>164. Energy is conserved in chemical reactions.</p> <p>165. An exothermic reaction is one that transfers energy to the surroundings so the temperature of the surroundings increases.</p> <p>166. Exothermic reactions include combustion, many oxidation reactions and neutralisation.</p> <p>167. Everyday uses of exothermic reactions include self-heating cans and hand warmers.</p> <p>168. An endothermic reaction is one that takes energy in from the surroundings so the temperature of the surroundings decreases.</p> <p>169. An everyday use of an endothermic reaction is a sports injury pack.</p> <p>170. The activation energy is the energy needed for a reaction to occur.</p> <p>171. On a reaction profile, the activation energy is shown as the distance from the energy of the reactants to the top of the curved line showing the energy.</p> <p>172. On a reaction profile, if the reactants have more energy than the products then energy has been released and the reaction is exothermic.</p> <p>173. On a reaction profile, if the products have more energy than the reactants then energy has been taken in and the reaction is endothermic.</p> <p>174. <i>To break bonds energy is needed.</i></p> <p>175. <i>Energy is released as bonds are made.</i></p> <p>176. <i>In an exothermic reaction, the energy released from forming new bonds is greater than the energy needed to break existing bonds.</i></p> <p>177. <i>In an endothermic reaction, the energy needed to break existing bonds is greater than the energy released from forming new bonds.</i></p>
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