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Atomic Structure	Elements, Compounds and mixtures	1. An element is a substance which contains only one TYPE of atom. 2. A compound is a substance which contains two or more types of atom bonded together. 3. Mixtures contain different elements or compounds that can be separated as they are not chemically bonded together. 4. In chemical reactions the starting materials are called reactants and new products are made. 5. There are 4 state symbols; (s) = solid (l) = liquid (g) = gas (ag) = aqueous (dissolved in solution) 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. 7. To separate a salt from a solution you evaporate the water to produce crystals of salt. 8. To separate and collect a liquid from a mixture you use distillation. You can use distillation to separate a mixture of liquids.

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	Atomic Structure	9. There are 3 subatomic particles; protons, electrons and neutrons. 10. Protons are positive (relative charge is +1) and have a relative mass of 1. 11. Electrons are negative (relative charge is -1) and their relative mass is very small. 12. Neutrons are neutral (relative charge is 0) and have a relative mass of 1. 13. Protons and neutrons are found in the nucleus and electrons orbit the nucleus. 14. Atoms have no overall charge because the number of positive protons is equal to the number of negative electrons. 15. The atomic number is the number of protons 16. The mass number is the number of protons and neutrons in total. 17. Atoms of the same element can have different numbers of neutrons; these atoms are called isotopes. 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. 19. Electrons occupy the lowest available energy level. 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
	Development of model of atom	21. Dalton thought atoms were hard spheres and that elements had only one type of atom. 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. 23. Geiger and Marsden’s did experiments firing positive alpha particles at gold foil which showed atoms could not be solid. 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) 25. In 1914 Niels Bohr suggested electrons orbit at set distances in energy levels (or shells). 26. In 1932 James Chadwick discovered the neutron which has no charge and has the same mass as a proton

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The Periodic Table	Development of periodic table	27. Before the discovery of protons, electrons and neutrons, scientists ordered the elements by their atomic weight. 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. 29. Elements with properties Mendeleev predicted were discovered and filled the gaps. 30. The elements on the modern periodic table are arranged in order of increasing atomic number. 31. Elements are arranged in columns called groups (going down) on the periodic table. 32. The number of electrons in the outermost shell of an atom is the same as its group on the periodic table.
	Group 0	33. Elements in the same group have similar properties. 34. Group 0 are called the noble gases. They are unreactive because of their very stable electron arrangement (full out shell) 35. The boiling point of the noble gases increases with increasing relative atomic mass (as you go down the group). 36. Boiling point and condensing point are the same temperature.
	Group 1	37. Group 1 are the alkali metals. 38. All group 1 atoms have 1 electron in their outer energy level. 39. Group 1 metals are very reactive. 40. When they react they lose their outer electron to form a full outer energy level and become stable. 41. Group 1 metals are stored in oil to stop the oxidising. 42. They have low densities and float in water. 43. Group 1 metals react vigorously with water producing an alkaline solution of the metal hydroxide plus hydrogen gas. 44. Group 1 metals react with halogen to produce metal halides which are white soluble solids. 45. The reactivity of group 1 increases as you go down the group.

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	Group 7	46. Group 7 elements are known as the halogens. 47. Group 7 elements all react in a similar way as they all have 7 electrons in their outer shell. 48. Group 7 elements are non- metals and consist of molecules made of pairs of atoms eg Br_2 49. The melting point and boiling point of group 7 increase as you go down the group. 50. In group 7 the reactivity decreases as you go down the group. 51. The halogens form ions with a charge of 1- by gaining an electron when reacting with metals to form ionic compounds. 52. A more reactive halogen will displace a less reactive halogen from a compound. 53. The halogens form covalent molecules by sharing electrons with other non- metals.
	Metals	54. Metals are found in the centre of the periodic table and to the left and bottom of the periodic table. 55. Metals react to form positive ions. 56. The atoms in metals are closely packed together in regular layers 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. 58. Most metals have high melting and boiling points. 59. Metals can be bent and shaped as the layers of atoms can slide over each other. 60. Alloys are mixtures of a metal with other elements. 61. Alloys are harder than pure metals as the layers are distorted and cannot slide. 62. Metals are good conductors of electricity because the delocalised electrons move through the structure. 63. Metals are good conductors of heat as the delocalised electrons can transfer the thermal energy. 64. **Transition metals have higher melting points than group 1 metals and are also stronger, harder and more dense.** 65. **Transition metals are less reactive than group 1 metals.** 66. **Transition metal elements have ions with different charges, form coloured compounds and are useful as cataylsts.**

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Bonding and Properties of Matter	Ionic Bonding	67. Ionic bonding occurs between metals and non- metals. 68. Metal atoms lose electrons to form positive ions. 69. Non- metal atoms tend to gain electrons to form negative ions. 70. Ionic compounds are held together by strong electrostatic forces of attraction between oppositely charged ions. 71. Ionic compounds form giant lattices. 72. Ionic compounds have high melting and boiling points. 73. Ionic compounds do not conduct electricity when solid as the ions are not free to move. 74. When melted or dissolved, ionic compounds conduct electricity as the ions are free to move.
	Covalent Bonding	75. Covalent bonds are formed when non-metal atoms share pairs of electrons. 76. Covalent bonds are strong. 77. Many covalent compounds consist of small, simple molecules e.g., oxygen, chlorine and water. 78. They are usually gases or liquids with low melting and boiling points. 79. Small covalent molecules have weak intermolecular forces between the molecules. 80. When these substances are heated it is the intermolecular forces that break not the covalent bonds. 81. Simple covalent molecules do not conduct electricity because the molecules do not have an overall charge. 82. Polymers are very large molecules held together by strong covalent bonds. 83. Some covalently bonded substances have giant structures eg silicon dioxide , diamond and graphite. 84. Giant covalent structures have very high melting and boiling points.

	Structure and bonding of carbon	85. The element carbon has different forms in which the atoms are arranged in different ways. These are diamond, graphite, graphene and fullerenes. 86. In diamond each carbon atom forms 4 covalent bonds with other carbon atoms. 87. Diamond is very hard as it is held together by strong covalent bonds. 88. Diamond does not conduct electricity as it has no free (delocalised) electrons. 89. 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. 90. The layers of atoms in graphite are not bonded together and can slide over each other which makes graphite soft and slippery. 91. Graphite can conduct electricity as one electron from each carbon atom is delocalised. 92. Graphite is a good thermal conductor because of the delocalised electrons which can move through the layers. 93. Graphene is a single layer of graphite so is just one atom thick. 94. Graphene conducts electricity and is very strong. 95. Fullerenes are molecules of carbon atoms with hollow shapes. 96. Buckminster fullerene (bucky ball) has a formula of C_{60} 97. Carbon nanotubes have very high length to diameter ration. 98. Fullerenes are finding uses eg to deliver drugs to specific body parts, as catalysts and as reinforcement for composite materials. 99. Graphene will help create new developments in the electronics industry in the future.

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Bulk and surface properties of matter	**Sizes of particles and their uses.**	100. **Nanoscience is the study of small particles that are between 1 -100 nanometres.** 101. **A nanometre (1nm) = 1×10^{-9}m.** 102. **A micrometre (1um) is 1×10^{-6}m** 103. **The particles in the air eg pollutants and pollen are known as particulate matter (PM).** 104. **Coarse particles (often referred to as dust) have diameters 10 times bigger than particulate matter = PM_{10}** **105. Fine particles have diameters between 0.1um (100nm or 1×10^{-7}m) and 2.5um.** **106. As a cube decreases its side by a factor of 10, its surface area to volume ratio decreases by a factor of 10.** **107. Nanoparticles have a high surface area to volume ratio and have different properties than the same materials in bulk.** **108. Smaller quantities of material are needed when they are in nanoparticles to be as effective as normal sized particles.** **109. Nanoparticles have many applications in medicines, electronics, cosmetics, sun creams, deodorants and catalysts.**

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Quantitative chemistry	Masses in reactions	110. 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. 111. The relative formula mass of a compound is the sum of the relative atomic masses of all the atoms shown in the formula. 112. 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. *113. A mole of a substance is its relative atomic mass or relative formula mass in grams.* *114. The number of atoms, molecules or ions in a mole of a substance is always 6.02×10^{23}. This number is called the Avogadro constant.* *115. The number of moles = mass (g) / Ar or mass (g)/ Mr* *116. Balanced symbol equations show you the number of moles of each substance involved in the reaction.* 117. *The reactant that gets used up first in a reaction is the limiting reactant. This is the reactant that is NOT in excess.*
	*C*oncentration	118. The concentration of a solution can be measured in mass per volume. Grams per dm^3 (g/dm^3) 119. 1 dm^3 = 1000cm^3 120. The mass of solute in a solution = concentration x volume in dm^3 *121. The concentration of a solution can be increased by increasing the mass of solute or decreasing the volume of the solvent.* *122.* ***The concentration of a solution can be measured in mol/dm^3*** *123.* ***The amount of moles of solute = concentration x volume in dm^3***

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	Yield and atom economy	124. **% yield = $\frac{\text{mass of product actually made}}{\text{maximum theoretical mass of product}}$ x 100** 125. **It is not always possible to obtain 100% yield as the reaction may not go to completion or some of the product may be lost when it is separated from the reaction mixture.** 126. **The atom economy is a measure of the amount of starting materials that end up as useful products.** 127. **Atom economy = $\frac{\text{relative formula mass of desired product from equation}}{\text{sum of the relative formula masses of all the reactants from the equation}}$ x 100**
	gases	**128.** ***The volume of one mole of any gas at room temperature and pressure (20°C and 1 atmosphere pressure) is 24 dm^3***
Reactivity of metals	Reaction of metals	129. Metals react with oxygen to produce metal oxides. These reactions are called oxidation reactions. 130. When metals react they lose electrons to form positive ions 131. 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. 132. A more reactive metal can displace a less reactive metal from a compound. 133. Metals more reactive than hydrogen will react with acid and produce hydrogen gas. Hydrogen gas 'pops' with a lighted spill. 134. Metals less reactive than carbon can be extracted from their ores by reduction with carbon. 135. Reduction involves the loss of oxygen. 136. *Reduction is the gain of electron.* 137. *Oxidation is the loss of electrons.*

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Reactions of acids	Making salts	138. Acids react with some metals to produce salts plus hydrogen. 139. Sulphuric acid produces salts called sulphates. 140. Hydrochloric acid produces salts called chlorides. 141. Nitric acid produces salts called nitrates. 142. Alkalis are soluble metal hydroxides (eg sodium hydroxide) 143. Bases are insoluble metal hydroxides and metal oxides. 144. Acids are neutralised by alkalis and bases to produce a salt plus water. 145. Acids are neutralised by metal carbonates to produce a salt, water and carbon dioxide. 146. When reacting an insoluble base with an acid to make a soluble salt, excess solid is used then filtered off after the reaction. 147. Salt solutions can be crystallised to produce solid salts.
	pH scale	148. Acids produce hydrogen ions (H^+) in aqueous solutions. 149. Aqueous solutions of alkalis contain hydroxide ions (OH^-). 150. The pH scale. From 0-14, is a measure of the acidity or alkalinity of a solution. 151. pH can be measured using universal indicator or a pH probe. 152. A solution of pH 7 is neutral. 153. Acids have a pH less than 7. 154. Alkalis have a pH greater than 7. 155. In reactions between acids and alkalis, the hydrogen ions and the hydroxide ions neutralise each other. 156. $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$
	Titrations	157. **A titration can be used to measure the volumes of acid and alkali that react wih each other.** 158. **A pipette is used to measure a fixed volume of a solution.** 159. **A burette is used to obtain an accurate measurements of the volume of the solution added.** **160. An acid / base indicator with a sharp end point eg phenol pthalein, is used.** **161. Results are repeated until you have 2 concordant results and a mean is calculated.**

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	Strong and weak acids	*162. A strong acid is completely ionised in aqueous solutions.* *163. Hydrochloric, nitric and sulphuric acids are strong acids.* *164. Weak acids only partially ionise in aqueous solutions. Ethanoic, citric and carbonic acids are weak acids.* *165. The stronger the acid the lower the pH.* *166. As the pH decreases by one unit, the hydrogen ion concentration increases by a factor of 10.* *167. pH 1 = hydrogen ion concentration of 0.1 mol/dm^3, pH 2 = 0.01mol/dm^3 and so on*
Electrolysis	Electrolysis of metal compounds	168. Electrolysis breaks down a substance using electricity. 169. Metals can be extracted from molten compounds using electrolysis. 170. The ions in ionic compounds are free to move when they are melted or dissolved. They can then conduct electricity. 171. Liquids and solutions that conduct electricity are called electrolytes. 172. When electricity is passed through an electrolyte, positive ions move to the negative electrode and negative ions move to the positive electrode. 173. The positive electrode is called the anode. 174. The negative electrode is called the cathode. 175. Ions are discharged at the electrodes to produce elements. 176. Metal ions are attracted to the cathode and metal is produced. 177. When extracting metals from their compounds, inert (unreactive) electrodes are used. 178. Electrolysis uses large amounts of energy to melt the compounds and to produce the electrical current used. 179. In the production of aluminium, aluminium oxide is mixed with cryolite to lower the melting point. 180. The oxygen produced at the anode when aluminium oxide is electrolysed reacts with the anode to form carbon dioxide. This electrode needs replacing.

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	Electrolysis of aqueous solutions	181. The ions discharged when an aqueous solution is electrolysed using inert electrodes depend on the relative reactivity of the elements involved. 182. The water in the solution is broken down by the electricity into hydrogen ions (H^+) and hydroxide ions (OH^-). 183. At the negative electrode (cathode), hydrogen is produced if the metal is more reactive than hydrogen. 184. At the positive electrode, oxygen is produced unless the solution contains halide ions in which case the halogen is produced. 185. 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. 186. *During electrolysis, positively charged ions gain electrons at the cathode. This is known as reduction.* 187. *At the anode, negatively charged ions lose electrons and so these reactions are oxidations.* 188. *Reactions at the electrodes can be represented by half equations, e.g.* $2H^+ + 2e^- \Rightarrow H_2$

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Energy changes in reactions	Exothermic and Endothermic reactions	189. Energy is conserved in chemical reactions. 190. An exothermic reaction is one that transfers energy to the surroundings so the temperature of the surroundings increases. 191. Exothermic reactions include combustion, many oxidation reactions and neutralisation. 192. Everyday uses of exothermic reactions include self -heating cans and hand warmers. 193. An endothermic reaction is one that takes energy in from the surroundings so the temperature of the surroundings decreases. 194. Endothermic reactions include thermal decompositions and the reaction of citric acid and sodium hydrogencarbonate. 195. An everyday use of an endothermic reaction is a sports injury pack. 196. The activation energy is the energy needed for a reaction to occur. 197. 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. 198. On a reaction profile, if the reactants have more energy than the products then energy has been released and the reaction is exothermic. 199. On a reaction profile, if the products have more energy than the reactants then energy has been taken in and the reaction is endothermic. *200. To break bonds energy is needed.* *201. Energy is released as bonds are made.* *202. In an exothermic reaction, the energy released from forming new bonds is greater than the energy needed to break existing bonds.* 203. *In an endothermic reaction, the energy needed to break existing bonds is greater than the energy released from forming new bonds.*

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Chemical cells and fuel cells	**Cells and batteries**	**204. Cells contain chemicals that react to produce electricity.** **205. A simple cell is made by connecting 2 different metals in contact with an electrolyte.** **206. Batteries consist of 2 or more cells connected in series to produce a greater voltage.** **207. In non-rechargeable cells, the chemical reaction stops when one of the rectants has been used up.** **208. Rechargeable cells can be recharged because the chemical reactions are reversed when an external electrical current is supplied.**
	Fuel cells	**209. Fuel cells are supplied by an external source of fuel (eg hydrogen) and oxygen or air.** **210. The fuel is oxidised electrochemically within the fuel cell to produce a potential difference.** **211. The overall reaction in a hydrogen fuel cell involves the oxidation of hydrogen to produce water.** **212. Hydrogen fuel cells offer a potential to rechargeable cells and batteries.**