AS Level Chemistry Core Qs 1

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| 1 | 2.1.1 | What are isotopes? | Isotopes are atoms of the same element with different numbers of neutrons and different masses |
| 2 | 2.1.1 | Define relative atomic mass (Ar) | RAM is the weighted mean mass of an atom of the element relative to one twelfth the mass of one atom of the C-12 isotope, which has a mass of exactly 12 |
| 3 | 2.1.1 | Define relative isotopic mass | The relative isotopic mass is the mass of an atom of a particular isotope relative to one twelfth the mass of one atom of the C-12 isotope, which has a mass of exactly 12 |
| 4 | 2.1.1 | How do you calculate relative atomic mass? | (% abundance x mass) + (% abundance x mass)  100 |
| 5 | 2.1.1 | What are the two main uses of mass spectrometry? | (i) The determination of relative isotopic masses and relative abundances of the isotopes  (ii) Calculation of the relative atomic mass of an element from the relative abundances of its isotopes |
| 6 | 2.1.1 | How does the mass spectrometer show this? | The mass spectrometer provides a trace that shows the mass of each isotope and its relative abundance. |
| 7 | 2.1.1 | What are the five stages, in order, of mass spectrometry? | 1) Vaporization  2) Ionisation  3) Acceleration  4) Deflection through a magnetic field  5) Detection |
| 8 | 2.1.1 | How are the particles deflected in mass spectrometry? | Electrons are attracted to a positively charged plate. They have a large deflection because of their small mass. Neutrons do not change path in a magnetic field. Protons attracted to a negatively charged plate, but their deflection is smaller because their mass is larger. |
| 9 | 2.1.1 | What happens when an organic compound is placed in the mass spectrometer? | The organic compound loses an electron and forms a positive ion, called the molecular ion. |
| 10 | 2.1.1 | What does the mass spectrometer detect? | A mass spectrometer detects the mass-to-charge ratio (m/z) of the molecular ion, which gives the molecular mass of the ion. |
| 11 | 2.1.1 | How do you calculate a mass-to-charge ratio? | Relative mass of ion  Relative charge on ion |
| 12 | 2.1.1 | What relationship exists between the time of flight of the sample (t) and the mass (m) of the sample? | The time of flight of the sample (t) is proportional to the square root of the mass (m) of the sample. |
| 13 | 2.1.3 | What is the mole number in terms of the C-12 isotope? | The mole contains the same number of particles as exactly 12 grams of carbon-12. 1 mole of any substance contains the same number of particles. |
| 14 | 2.1.3 | What does Avogadro’s constant (Na) show? | Avogadro's constant is a measure of the number of particles in 1 mole of any substance  6.02 x 1023 mol-1 |
| 15 | 2.1.3 | What is the molar mass of an element?  What are the units? | The molar mass of an element is its atomic mass in grams per mole.  Units: g mol-1 |
| 16 | 2.1.3 | What is the equation for calculating the number of particles there are in a sample? | **Moles = Na**  Mass  Mr x Mole |
| 17 | 2.1.3 | How is the number of moles (n) calculated? | n= mass in g (m)  molar mass in g mol-1 (M)  Molar mass is equal to formula mass |
| 18 | 2.1.3 | What is the molar gas volume?  What are the units? | The molar gas volume is the gas volume per mole.  Units: dm3mol-1 |
| 19 | 2.1.3 | What volume does one mole of gas at standard conditions occupy? | One mole of any gas at room temperature (240C) and 1 atmospheric pressure has a volume of 24dm3 |
| 20 | 2.1.3 | How do you calculate the molar volume of a gas? | **n= v/V or v=nV**  Where n = number of moles  v = volume of gas  V= molar volume (24dm3) |
| 21 | 2.1.3 | What is the empirical formula? | The simplest whole number ratio of atoms of each element present in a compound. |
| 22 | 2.1.3 | What is the molecular formula? | The number and type of atoms of each element in a molecule. |
| 23 | 2.1.3 | If a salt is hydrated, what does it contain? | Water |
| 24 | 2.1.3 | If a salt is anhydrous, what has been removed? | Water |
| 25 | 2.1.3 | Define the term 'water of crystallisation' | Water molecules form an essential part of the crystal structure of the compound. % water of crystallisation shows how much of the hydrated salt is water. |
| 26 | 2.1.3 | How do you determine the % water of crystallisation? | A known mass of the hydrated salt is heated gently in a crucible until the mass becomes constant despite further heating. The mass of the anhydrous salt is measured, and from this the percentage of water present in the hydrated salt can be calculated. |
| 27 | 2.1.3 | What is the Ideal Gas equation? | **pV=nRT** |
| 28 | 2.1.3 | What is the standard measure for pressure in the Ideal Gas equation? | The standard unit for pressure is the Pascal (Pa).  In the Ideal Gas equation, the standard form is 1 atmospheric pressure, which equals 100 x 103 Pa. |
| 29 | 2.1.3 | What is the standard measure for volume in the Ideal Gas equation? | The standard unit for volume is m3.  m3 = 1000dm3 = 106cm3 |
| 30 | 2.1.3 | In the Ideal Gas equation, what does n stand for? | The number of moles |
| 31 | 2.1.3 | What is the standard measure for temperature in the Ideal Gas equation? | Standard: 298 Kelvins |
| 32 | 2.1.3 | How do you convert Celsius to Kelvin? | Celsius + 273 = Kelvins |
| 33 | 2.1.3 | In the Ideal Gas equation, what does R stand for? | R is the gas constant.  It has a value of 8.31441 J K-1 mol-1 |
| 34 | 2.1.3 | How do you calculate the molecular formula from the empirical formula? | The molecular formula shows how many times the empirical formula goes into the molar mass.  molar mass  empirical formula  Multiply the empirical formula by the answer to find the molecular formula. |
| 35 | 2.1.3 | What do ionic equations show? | Ionic equations show only the reacting particles (and the products they form.) Ions that don’t get involved in the reaction - the spectator ions - aren't written. |
| 36 | 2.1.2 | What is the formula of the nitrate ion? | NO3- |
| 37 | 2.1.2 | What is the formula of the carbonate ion? | CO32- |
| 38 | 2.1.2 | What is the formula of the sulphate ion? | SO42- |
| 39 | 2.1.2 | What is the formula of the hydroxide ion? | OH- |
| 40 | 2.1.2 | What is the formula of the ammonium ion? | NH4+ |
| 41 | 2.1.2 | What is the formula of the zinc ion? | Zn2+ |
| 42 | 2.1.2 | What is the formula of the silver ion? | Ag+ |
| 43 | 2.1.4 | What is an acid? | An acid is a **proton donor**. When mixed with water, they **release hydrogen ions - H+.** |
| 44 | 2.1.4 | What is a base? | A base is a **proton acceptor.** They want to grab H+ ions. |
| 45 | 2.1.4 | What is an alkali? | An alkali is a base that is **soluble** in water. Alkalis **produce OH- ions** in an aqueous solution. |
| 46 | 2.1.4 | Name three common acids and give their formula | Hydrochloric acid - HCl  Sulfuric acid - H2SO4  Nitric acid - HNO3  Ethanoic acid - CH3COOH |
| 47 | 2.1.4 | Name three common bases and give their formula | Sodium hydroxide - NaOH  Potassium hydroxide - KOH  Ammonia - NH3 |
| 48 | 2.1.4 | What kind of reaction occurs between an acid + water or a base + water? | Reversible |
| 49 | 2.1.4 | What happens when strong acids are reacted with water? Which arrow is used for the reaction equation? | For strong acids, nearly all of the acid will **dissociate** in water, and **lots of the H+ ions are released.** As the equilibrium lies very far to the right, a forward arrow is used. |
| 50 | 2.1.4 | What happens when strong bases are reacted with water? Which arrow is used for the reaction equation? | For strong bases, nearly all of the base will **dissociate** in water, and **lots of the OH-ions are released.** As the equilibrium lies very far to the right, a forward arrow is used. |
| 51 | 2.1.4 | Which arrow is used when weak acids and bases react with water? Why? | A reversible arrow is used, as the backward reaction is favoured and only a few H+ or OH- ions are released. |
| 52 | 2.1.4 | What is the general equation for a neutralisation reaction between an acid and a base? | Acid + Base  Salt + Water  The hydrogen ions released by the acid and the hydroxide ions released by the alkali combine to form water:  H+(aq) + OH-(aq) ≈ H2O(l)  The salt is produced when hydrogen ions in the acid are replaced by metal ions from the alkali. |
| 53 | 2.1.4 | Metal + Acid = ? | Metal Salt + Hydrogen |
| 54 | 2.1.4 | Metal Oxide + Acid = ? | Salt + Water |
| 55 | 2.1.4 | Metal Hydroxide + Acid = ? | Salt + Water |
| 56 | 2.1.4 | Metal Carbonate + Acid = ? | Metal Salt + Carbon Dioxide + Water |
| 57 | 2.1.4 | When ammonia reacts with acid, what is produced? | An ammonium salt (aqueous) |
| 58 | 2.1.4 | Which indicator is used when adding acid to alkali and what colour does it turn? | Methyl orange turns from **yellow** to **red** when adding acid to alkali |
| 59 | 2.1.4 | Which indicator is used when adding alkali to acid and what colour does it turn? | Phenolphthalein indicator turns from **pink** to **colourless** when adding acid to alkali. |
| 60 | 2.1.4 | What is a diprotic acid? | A diprotic acid can donate two protons (i.e. H2SO4). |
| 61 | 2.1.4 | If you are completing a titration calculation, what do you have to consider when using a diprotic acid? | You need to double the number of moles of a base to neutralise a diprotic acid. |
| 62 | 2.1.4 | What is the molar volume equation? | Moles = concentration x vol (cm3)  1000 |
| 63 | 2.1.3 | How do you calculate percentage yield? | Percentage yield = Actual yield x100  Theoretical yield |
| 64 | 2.1.3 | Define 'atom economy.' | Atom economy is a measure of the proportion of reactant atoms that become part of the desired product. |
| 65 | 2.1.3 | How do you calculate atom economy? | Molecular mass of desired products x100  Sum of molecular mass |
| 66 | 2.1.3 | What is an addition reaction? | In an addition reaction, the reactants combine to form a single product. The atom economy for addition reactions is always 100% since no atoms are wasted. |
| 67 | 2.1.3 | What is a substitution reaction? | A substitution reaction is one where some atoms from one reactant are swapped with atoms from another reactant. This type of reaction always results in at least two products. |
| 68 | 2.1.3 | Give two environmental benefits of having a high atom economy. | 1. Few waste products: Waste products need to be disposed of and can be harmful to the environment.  2. More sustainable: Many raw materials are in limited supply, so using them efficiently makes them last as long as possible. |
| 69 | 2.1.3 | Give two economical benefits of having a high atom economy. | 1. It costs money to separate and dispose of waste products. Less waste = less cost.  2. Reactant chemicals are costly. The more which can be transferred to useful products, the better. |
| 70 | 2.1.5 | What is an oxidation number? | Oxidation numbers tell you how many **electrons** an atom has **donated** or **accepted** to form an ion, or to form part of a compound. |
| 71 | 2.1.5 | What is the oxidation number of all uncombined elements? | 0 |
| 72 | 2.1.5 | The oxidation number of a monatomic ion is always the same as its \_\_\_\_\_\_\_\_. | Charge |
| 73 | 2.1.5 | For molecular ions, the sum of the oxidation ions is always equal to the \_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_. | Overall charge |
| 74 | 2.1.5 | Oxygen nearly always has an oxidation number of \_\_\_\_  (Exceptions: peroxides and molecular oxygen.) | -2 |
| 75 | 2.1.5 | In peroxides, oxygen has an oxidation number of \_\_\_\_. | -1 |
| 76 | 2.1.5 | Hydrogen usually has an oxidation number of \_\_\_\_  (Exceptions: metal hydrides and molecular hydrogen.) | +1 |
| 77 | 2.1.5 | In metal hydrides, hydrogen has an oxidation number of \_\_\_. | -1 |
| 78 | 2.1.5 | Roman numerals on an ion tell you the oxidation number. What oxidation number would be assigned to iron in iron (II) sulfate? | +2 |
| 79 | 2.1.5 | What occurs in a redox reaction? | Electrons are transferred. |
| 80 | 2.1.5 | A loss of electrons is called \_\_\_\_\_\_\_. | Oxidation |
| 81 | 2.1.5 | A gain of electrons is called \_\_\_\_\_\_\_. | Reduction |
| 82 | 2.1.5 | What happens to an oxidising agent? | It **accepts electrons** and gets **reduced.** The oxidation number decreases by 1 for each electron gained. |
| 83 | 2.1.5 | What happens to a reducing agent? | It **loses electrons** and gets **oxidised.** The oxidation number increases by 1 for each electron lost. |
| 84 | 2.1.5 | Metals are \_\_\_\_\_\_\_ when they react with acids. | Oxidised  This produces positive metal ions (in salts), and their oxidation number increases. |
| 85 | 2.2.1 | What is the relationship between the quantum number and the number of electrons the shell can hold? | 2n2 |
| 86 | 2.2.1 | How many electrons can the first four shells hold? | 1st: 2  2nd: 8  3rd: 18  4th: 32 |
| 87 | 2.2.1 | How do you define an atomic orbital? | A region in space around the nucleus that can hold up to two electrons with opposite spin. |
| 88 | 2.2.1 | What is the shape of an s orbital? | Spherical |
| 89 | 2.2.1 | In which quantum shells do s orbitals appear? | There is an s orbital in every quantum shell. |
| 90 | 2.2.1 | What is the shape of a p orbital? | Dumbbell shaped. There are three p orbitals and they are at right angles to one another. |
| 91 | 2.2.1 | In which quantum shells do p orbitals appear? | All quantum shells apart from the first. |
| 92 | 2.2.1 | How many electrons in total can the p orbitals in one shell hold? | 6 |
| 93 | 2.2.1 | How many electrons in total can d orbitals in one shell hold? | 10 |
| 94 | 2.2.1 | Which three rules apply when filling energy levels? | 1. Electrons enter the lowest energy level available  2. Two electrons in the same orbital must have opposite spin  3. Electrons will try and remain unpaired to reduce electrostatic repulsion and increase stability. |
| 95 | 2.2.2 | What is an ionic bond? | An electrostatic attraction between positive and negative ions in a giant lattice. |
| 96 | 2.2.2 | Describe a giant ionic lattice structure. | Oppositely charged ions are strongly attracted in all directions. |
| 97 | 2.2.2 | Do ionic compounds conduct electricity? | Only when molten or in solution, as the charged ions are free to move. |
| 98 | 2.2.2 | How is the ionic structure linked to boiling point? | Ionic compounds have high melting and boiling points. It takes a lot of energy to overcome the strong electrostatic forces of attraction. |
| 99 | 2.2.2 | Are ionic substances soluble in water? | Most ionic compounds will dissolve in water, as water molecules are **polar** and are attracted to the charged ions. An ionic substance only dissolves if more energy is released by separating the bonds then breaking the bonds. |
| 100 | 2.2.2 | What is a covalent bond? | A covalent bond is the strong electrostatic attraction between a shared pair of electrons and the nuclei of the bonded atoms. |
| 101 | 2.2.2 | What is a dative (coordinate) bond? | A dative bond occurs when both of the electrons in a single bond come from one of the atoms. The bond length and strength is identical to any other covalent bond. |
| 102 | 2.2.2 | What does average bond enthalpy measure? | Average bond enthalpy measures the energy required to break a covalent bond. The stronger a bond = the greater the average bond enthalpy. |
| 103 | 3.1.1 | What is a metallic bond? | The electrostatic attraction between positive metal ions and a sea of delocalised electrons. |
| 104 | 2.2.2 | What is electronegativity? | The power of an atom to attract the electrons in a covalent bond. |
| 105 | 2.2.2 | What are the three factors which affect electronegativity? | 1. Atomic radius (distance from nucleus.)  2. Shielding  3. Nuclear charge |
| 106 | 2.2.2 | Which element is the most electronegative? | Fluorine - it has an electronegativity of 4.0. |
| 107 | 2.2.2 | What is the link between electronegativity and bonding? | Compounds made up of elements with a large difference in electronegativity will tend to be ionic in character. If they are similar in electronegativity they are likely to be covalent in character. |
| 108 | 2.2.2 | Molecular shape is determined by the electrostatic repulsion of electrons round the central atom. Order these angles, from smallest to largest bond angle:  - lone pair/ lone pair angles  - lone pair/bonding pair angles  - bonding pair/bonding pair angles | *Smallest*  bonding pair/bonding pair angles    lone pair/bonding pair angles    lone pair/lone pair angles  *Biggest* |
| 109 | 2.2.2 | A molecule with two bonded pairs has a \_\_\_\_\_ shape. The bond angle is \_\_\_\_\_\_\_. | Linear  180o |
| 110 | 2.2.2 | A molecule with three bonded pairs has a \_\_\_\_\_ \_\_\_\_\_\_\_\_ shape. The bond angle is \_\_\_\_\_\_\_. | Trigonal planar  120o |
| 111 | 2.2.2 | A molecule with four bonded pairs has a \_\_\_\_\_ shape. The bond angle is \_\_\_\_\_\_\_. | Tetrahedral  109.5o |
| 112 | 2.2.2 | A molecule with five bonded pairs has a \_\_\_\_\_ \_\_\_\_\_\_\_ shape. The bond angle is \_\_\_\_\_\_\_. | Trigonal bipyramidal  90o / 120o |
| 113 | 2.2.2 | A molecule with six bonded pairs has a \_\_\_\_\_ shape. The bond angle is \_\_\_\_\_\_\_. | Octahedral  90o |
| 114 | 2.2.2 | If a molecule contains lone pairs, the angle is based on the number of \_\_\_\_\_\_ \_\_\_\_\_ but the name on the number of \_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_. | The angle is based on the number of electron pairs.  The name is based on the number of bonding pairs. |
| 115 | 2.2.2 | Each lone pair reduces the bond angle by approximately \_\_\_\_o from the bond angle determined by the number of electron pairs. | 2.5o |
| 116 | 2.2.2 | How is a permanent dipole-dipole interaction caused? | In a covalent bond between two atoms of different electronegativities, the bonding electrons are pulled towards the more electronegative atom. This makes the bond polar. |
| 117 | 2.2.2 | Do diatomic gases (i.e. H2) have permanent dipoles? | No - the atoms have equal electronegativities and so the electrons are equally attracted to both nuclei. |
| 118 | 2.2.2 | How is charge arranged in a polar molecule? | The arrangement of polar bonds in a molecule determines whether or not the molecule will have an overall dipole. If charge is arranged unevenly across the molecule, there will be an overall dipole and the molecule will be polar. |
| 119 | 2.2.2 | How is the dipole moment calculated? | Multiplying the size of the charge by the distance between the charge centres. |
| 120 | 2.2.2 | In which atoms and molecules are induced dipole-dipole forces found? | All atoms and molecules |
| 121 | 2.2.2 | How are induced dipole-dipole interactions (London *[dispersion]*) forces caused? | Electrons in charge clouds are always moving. At any one moment, a **temporary dipole** is caused as electrons are on one side of the atom. This dipole can cause another temporary (induced) dipole in another atom in a chain. The overall effect is that the atoms are **attracted** to one another. |
| 122 | 2.2.2 | Do stronger induced dipole-dipole forces increase or decrease boiling point? | They increase boiling point. More energy is needed to overcome stronger intermolecular forces. |
| 123 | 2.2.2 | Which factors increase the strength of induced dipole-dipole forces? | 1. Size of electron cloud: larger molecules have larger electron clouds, meaning stronger induced dipole-dipole forces.  2. Surface area: bigger exposed electron cloud = stronger induced dipole-dipole forces |
| 124 | 2.2.2 | Order these forces from strongest to weakest:  - Induced dipole-dipole forces  - Permanent dipole-dipole interactions  - Hydrogen bonding | Induced dipole-dipole forces    Permanent dipole-dipole interactions    Hydrogen bonding |
| 125 | 2.2.2 | Define hydrogen bonding. | Intermolecular bonding between molecules containing fluorine, nitrogen or oxygen and the H atom of -NH, -OH or HF. |
| 126 | 2.2.2 | Why does hydrogen bonding occur? | Hydrogen has a high charge density because it’s so small, and the other elements are very electronegative. The bond is so **polarised** that a bond forms between the **hydrogen** of one molecule and a **lone pair** on the other molecule. |
| 127 | 2.2.2 | Hydrogen bonding \_\_\_\_\_\_\_ boiling points and freezing points | Increases |
| 128 | 2.2.2 | What effect does hydrogen bonding have on water when it freezes? | When ice freezes, in becomes less dense. This is because it contains more hydrogen bonds then liquid water. |
| 129 | 2.2.2 | Explain two properties of simple covalent compounds. | 1. Low melting and boiling points: weak intermolecular forces, takes little energy to overcome them  2. Polar molecules are soluble  3. Don't conduct electricity: overall covalent molecules are uncharged |
| 130 | 3.1.1 | How are elements arranged in the periodic table? | 1. By increasing atomic (proton) number  2. In periods showing repeating trends in physical and chemical properties (periodicity).  3. In groups with similar chemical properties. |
| 131 | 3.1.1 | What is meant by periodicity? | The study of trends across a period |
| 132 | 3.1.1 | What is the name given to the following groups:  Group I (1)  Group II (2)  Group VII (7)  Group 8/0 | 1: Alkali metals  2: Alkaline (earth) metals  7: Halogens  8/0: Noble gases |
| 133 | 3.1.1 | Which groups are in the s block? | I, II and hydrogen |
| 134 | 3.1.1 | Which groups are in the d block? | The transition metals |
| 135 | 3.1.1 | Which groups are in the p block? | III, IV, V, VI, VII, 0 |
| 136 | 3.1.1 | Which groups are in the f block? | Lanthanides and actinides |
| 137 | 3.1.1 | Define first ionisation energy. | The energy needed to remove **1 mole** of electrons from **1 mole** of gaseous atoms. |
| 138 | 3.1.1 | Is ionisation and endothermic or exothermic reaction? | Endothermic |
| 139 | 3.1.1 | How would you write an equation for the first ionisation of oxygen? | O (g) O+ (g) + e-  The gas state symbol must be used |
| 140 | 3.1.1 | What are the factors affecting ionisation energy? | 1. Atomic radius (distance from nucleus.)  2. Shielding  3. Nuclear charge |
| 141 | 3.1.1 | Define second ionisation energy. | The energy required to remove 1 mole of electrons from 1 mole of gaseous **singly positive ions** to form 1 mole of gaseous **+2 ions.** |
| 142 | 3.1.1 | Give two reasons why the second I.E. is always greater than the first. | 1. The electron is being pulled away from a more positive species.  2. There is less electron shielding as you move down a shell. |
| 143 | 3.1.1 | Why does ionisation energy decrease down a group? | 1. Larger atomic radius  2. Inner shells provide shielding |
| 144 | 3.1.1 | Why does ionisation energy increase across a period? | 1. Nuclear charge increases as the number of protons increases. This decreases the atomic radius and the bonding electrons are more strongly attracted to the nucleus.  2. No extra shielding as the same number of shells are present |
| 145 | 3.1.2 | Group 2 elements form \_\_\_\_\_ ions. | 2+ |
| 146 | 3.1.2 | Why does reactivity increase down a group? | Ionisation energy decreases due to increasing atomic radius and shielding. The easier it is to lose electrons, the more reactive the element. |
| 147 | 3.1.2 | What is produced when group 2 elements react with water? | Metal hydroxide + hydrogen  M(OH)2 + H2 |
| 148 | 3.1.2 | What is produced when group 2 elements burn in oxygen? | They form solid white oxides.  2M(s) + O2 = 2MO(s) |
| 149 | 3.1.2 | What is produced when group 2 elements react with dilute acid? | Salt + hydrogen |
| 150 | 3.1.2 | The oxides of the group 2 metals can produce metal hydroxides when water is added. The OH- ion makes the solutions strongly alkaline. What is the trend as you go down the group? | The oxides form more strongly alkaline solutions as you go down the group, because the hydroxides get more soluble. |
| 151 | 3.1.2 | Give two uses of group 2 compounds | 1. Calcium hydroxide can be used to neutralise acidic fields.  2. Magnesium hydroxide and calcium carbonate can be used to neutralise excess stomach acid - they are **antacids**. |
| 152 | 3.1.3 | Halogens always exist as \_\_\_\_\_\_\_\_ molecules. | Diatomic |
| 153 | 3.1.3 | The boiling and melting points \_\_\_\_\_\_\_\_ down the group due to the \_\_\_\_\_\_\_\_ strength of the London forces and size and relative mass of the atom. | Increase  Increasing |
| 154 | 4.1.1 | Define general formula | The simplest algebraic formula of a member of the homologous series  i.e. CnH2n+2 |
| 155 | 4.1.1 | Define structural formula | The minimal detail that shows the arrangement of atoms in a molecule  i.e. CH3CH2CH2CH3 |
| 156 | 4.1.1 | Define displayed formula | Shows the relative positioning of atoms and the number of bonds between them. |
| 157 | 4.1.1 | Define skeletal formula | The simplified organic formula, shown by removing hydrogen from the alkyl chains, leaving just a carbon skeleton and associated functional groups. |
| 158 | 4.1.1 | Define catenation | The ability to form strong covalent bonds to atoms of the same element - i.e. carbon. |
| 159 | 4.1.1 | Define homologous series | A series of organic compounds having the same functional group but with each successive member differing by CH2. |
| 160 | 4.1.1 | Define functional group | A group of atoms responsible for the characteristic reactions of a compound. Molecules with the same functional group are classified into homologous series. |
| 161 | 4.1.1 | Define aliphatic | A compound containing carbon and hydrogen joined together in straight chains, branched chains or non-aromatic rings. |
| 162 | 4.1.1 | Define alicyclic | An aliphatic compound arranged in non-aromatic rings with or without side chains. |
| 163 | 4.1.1 | Define aromatic | A compound containing benzene rings (C6H6). |
| 164 | 4.1.1 | Define saturated | Single carbon-carbon bonds only |
| 165 | 4.1.1 | Define unsaturated | Contains multiple carbon-carbon bonds |
| 166 | 4.1.1 | Alkanes have the ending \_\_\_\_. | - ene |
| 167 | 4.1.1 | Haloalkanes have the suffix's \_\_\_\_\_\_, \_\_\_\_\_\_\_, \_\_\_\_\_\_ or \_\_\_\_\_\_\_\_. | - chloro  - bromo  - fluoro  - iodo |
| 168 | 4.1.1 | Alcohols have the ending \_\_\_\_ to replace the last e if one functional group is present. If more than one functional groups are present, it begins with \_\_\_\_\_\_\_\_. | -ol  hydroxy- |
| 169 | 4.1.1 | Aldehydes have the ending \_\_\_\_ in the place of the last -e. Aldehydes must always be at the \_\_\_\_ of the chain. | -al  End |
| 170 | 4.1.1 | Ketones have the ending -one in place of the last -e. If another functional group is present, the functional group goes at the beginning as \_\_\_\_. | Oxo- |
| 171 | 4.1.1 | What is the difference between aldehydes and ketones? | The oxygen is at the end of the chain on an aldehyde, but is in the middle of the chain on a ketone. |
| 172 | 4.1.1 | Carboxylic acids have the ending -oic acid. The \_\_\_\_\_ functional chain must always be at the end of the chain. | -COOH |
| 173 | 4.1.1 | Nitriles have the ending -ile. Does the C of the CN group count as a carbon on the carbon chain? | Yes |
| 174 | 4.1.1 | Amines end in -amine, which becomes \_\_\_\_\_\_- if another functional group is present. | Amino- |
| 175 | 4.1.1 | Define structural isomerism | Compounds with the same molecular formula but a different structural formula. |
| 176 | 4.1.1 | Structural isomers have \_\_\_\_\_\_\_\_ chemical properties. | The same |
| 177 | 4.1.1 | Structural isomers have \_\_\_\_\_\_\_\_ physical properties. | Differing |
| 178 | 4.1.1 | The greater the degree of branching, the \_\_\_\_\_ the boiling point. Why? | Lower  Branching decreases the effectiveness of the intermolecular forces so less energy is needed to separate the molecules. |