5.1	Atomic structure and the periodic table
5.1.1	A simple model of the atom, symbols, relative atomic mass, electric charge and isotopes
5.1.1.1	Atoms, elements and compounds

1. What are all substance made from? **Atoms.**
2. What is an atom? **The smallest part of an element that can exist.**
3. Approximately how many elements are there? **100**
4. What is the periodic table? **An arrangement of all the elements based on their atomic number**
5. How are compounds formed? **Compounds are formed from elements by chemical reactions**
6. What do chemical reactions involve? **Chemical reactions always involve the formation of one or more new substances and often involve a detectable change in energy (e.g. temperature change)**
7. What is a compound? **Compounds contain two or more elements chemically combined in fixed proportions**
8. How can the elements in a compound be separated? **Compounds can only be separated into elements by chemical reactions.**

5.1.1.2	Mixtures

1. What is a mixture? **A mixture consists of two or more elements or compounds not chemically combined together.**
2. Name five physical processes which can be used to separate mixtures. **Filtration, crystallisation, simple distillation, fractional distillation and chromatography.**
3. What physical process would be used to separate a mixture of liquids with different boiling points? **Fractional distillation**
4. What physical process would be used to separate an insoluble salt from a solution? **Filtration**
5. What physical process would be used to separate a solvent from a solution? **Simple distillation**
6. What physical processes would be used to separate copper sulfate crystals from a mixture of copper sulfate solution and copper oxide? **Filtration, evaporation and crystallisation.**

5.1.1.3	The development of the model of the atom

1. Why might a scientific model be changed or replaced? **New experimental evidence.**
2. How did Democritus describe the atom? **Tiny spheres that could not be divided.**
3. What did J.J. Thomson discover and what model did he suggest as a result? **The electron.**
4. Describe Thomson's model. **Thomson described the "plum pudding" model where a ball of positive charge is embedded with negatively charged electrons.**
5. Describe the alpha particle scattering experiment. **Alpha particles (helium nuclei) were fired at a thin gold sheet. If the "plum pudding" was correct the alpha particles would pass straight through.**
6. What were the results of the alpha particle scattering experiment. **Although most of the alpha particles passed through some were deflected and a few even bounced back.**
7. What conclusion did Rutherford make from the results of the alpha particle scattering experiment. **Rutherford concluded that the mass of the atom was concentrated at the centre in a nucleus.**
8. How did Bohr adapt Rutherford's model. **Bohr suggested that the electrons orbit the nucleus at specific distances.**
9. What sub-atomic particle did Rutherford discover in 1920 to explain the positive charge in an atom? **The proton.**
10. What sub-atomic particle did James Chadwick discover which explained isotopes? **The neutron.**

5.1.1.4 Relative electrical charges of subatomic particles

1. What is the relative charge of a proton? +1
2. What is the relative charge of a neutron? 0
3. What is the relative charge of an electron? -1
4. In all atoms how many electrons are there compared to protons? **The number of electrons in an atom is equal to the number of protons.**
5. What electrical charge do atoms have? **Atoms are neutral.**
6. What is the atomic number of an element? **The number of protons in the atom.**

5.1.1.5 Size and mass of atoms

1. What is the radius of an atom in m? 1×10^{-10} m
2. What is the radius of a nucleus of an atom in m? 1×10^{-14} m (1/10000 the size of the atom)
3. How is the mass of an atom distributed in the atom? **Almost all the mass is in the nucleus.**
4. What is the relative mass of a proton? 1
5. What is the relative mass of a neutron? 1
6. What is the relative mass of an electron? 0
7. What is the mass number of an element? **The sum of the protons and the neutrons.**
8. What is an isotope? **Atoms of the same element which have the same number of protons but different numbers of neutrons.**
9. Calculate the number of protons, electrons and neutrons for the following atoms

	${}^{1}_{1}H$	${}^{7}_{3}Li$	${}^{12}_{6}C$	${}^{39}_{19}K$	${}^{238}_{92}U$
Protons	1	3	6	19	92
Neutrons	0	4	6	20	146
Electrons	1	3	6	19	92

5.1.1.6 Relative atomic mass

1. What is the relative atomic mass of an element? **It is the average atomic mass that takes into account the abundance of isotopes of the element.**
2. How would you calculate the relative atomic mass of an atom? (Abundance of isotope 1 x atomic mass of isotope 1) + (Abundance of isotope 2 x atomic mass of isotope 2) / 100.

5.1.1.7	Electronic structure

1. What is the maximum number of electrons that can fit in the first shell? **2**
2. What is the maximum number of electrons that can fit in the second shell? **8**
3. What is the maximum number of electrons that can fit in the third shell? **8**
4. Which electron shell has the lowest energy level? **The innermost shell.**
5. In what order are the electron shells filled with electrons? **The electrons occupy the lowest available energy shells (inside shells to outside shells).**
6. What is the electronic structure of sodium? **2,8,1**
7. What is the electronic structure of fluorine? **2,7**
8. What is the electronic structure of sulfur? **2,8,6**
9. What is the electronic structure of hydrogen? **1**
10. What is the electronic structure of neon? **2,8**
11. Describe how the position of an element on a periodic table can be found from the electronic structure. Use chlorine as an example. **The number of electrons in the outer shell gives the group number. The number of shells is the period. Chlorine has 7 electrons in its outer shell so it is found in group 7. Chlorine has electrons in 3 shells so it is found in period 3.**

5.1.2	The periodic table
5.1.2.1	The periodic table

1. How are the elements in the periodic table arranged? **They are arranged by atomic number.**
2. How are elements with similar properties arranged? **Elements with similar properties are arranged in groups.**
3. Why is it called a "periodic" table? **Because similar properties occur at regular intervals.**
4. Explain why elements in the same group have similar chemical properties? **Because they have the same number of electrons in their outer shell.**

5.1.2.2	Development of the periodic table

1. Before the discovery of sub-atomic particles, how did scientists arrange elements in the periodic table? **By atomic weight.**
2. How did Mendeleev organise the elements in his periodic table? **He organised the elements by atomic weight and by similar properties.**
3. Why did Mendeleev leave gaps in his periodic table? **For undiscovered elements.**
4. What discovery showed why the order based on atomic weight was not always correct? **The discovery of isotopes.**

5.1.2.3	Metals and non-metals

1. Which elements react to form positive ions? **Metals**
2. Which elements do not form positive ions when they react? **Non-metals**
3. What are the majority of elements in the periodic table? **Metals**
4. Where are metals found in the periodic table. **To the left and towards the bottom.**
5. Where are non-metals found in the periodic table? **To the right and towards the top**
6. What does malleable mean? **Malleable is the ability the bend and hammer metal.**
7. Malleable is a property of metals. Name four other properties of metals. **Conductors of heat, conductors of electricity, High melting and boiling points, shiny, usually solid at room temperature.**
8. What does brittle mean? **The substance will shatter if struck.**
9. Brittle is a property of non-metals. Name four other properties of non-metals. **Generally do not conduct electricity, Are not always solid at room temperature, often have a lower density, dull.**
10. Explain why metals form positive ions in reactions? **Metals usually have only a few electrons in the outer shell. It is easier for them to lose electrons to get a full outer shell.**
11. Explain why non metals do not form positive ions in reactions. **Non metals usually have more electrons in their outer shell. It is easier for them to gain electrons or share electrons to get a full outer shell.**

5.1.2.4	Group 0

1. What is another name for group 0? **The noble gases.**
2. Describe the reactivity of group 0 elements. **Elements in group 0 are unreactive.**
3. How many electrons do elements of group 0 have in their outer shell? **Elements in group 0 have 8 electrons in their outer shell except he;Liam which has 2 electrons.**
4. Describe the change in boiling point as you go down (increase relative atomic mass) group 0. **The boiling point of group 0 increases as you go down the group (increase atomic mass).**

5.1.2.5	Group 1

1. What is another name for group 1? **The alkali metals.**
2. How many electrons do elements in group 1 have in their outer shell? **1**
3. What are the products of the reaction between lithium and water? **Lithium hydroxide + hydrogen.**
4. What are the products of the reaction between sodium and chlorine? **Sodium chloride.**
5. What are the products of the reaction between potassium and oxygen? **Potassium oxide.**
6. How does the reactivity of elements in group 1 change as you go down the group? **The reactivity increases as you go down the group (increase atomic mass).**
7. Explain, in terms of electrons, why does reactivity change as you go down the group. **Reactions occur because of the electrons in the outer shell. As you go down the group the electron shells increase. The further away the outer electron is from the nucleus, the easier it is to lose.**

5.1.2.6	Group 7

1. What is another name for group 7? **The halogens**
2. Describe the general properties of group 7 elements. **They are non-metals and they form molecules made from two atoms e.g. Cl_2.**
3. Name two gases found group 7. **Fluorine and chlorine.**
4. Name a liquid found in group 7. **Bromine.**
5. Name two solids found in group 7. **Iodine and Astatine (technically you could say tennessine, but only six atoms have ever existed)!**
6. Describe the nature of the compounds formed when halogens react with metals. **Halogens form ionic compounds when they react with metals.**
7. Describe the nature of the compounds formed when halogens react with other non-metals. **Halogens form simple covalent compounds when they react with non-metals.**
8. Describe how the melting points and boiling points change in group 7. **As you go down the group (increase the atomic mass) the melting point and boiling point increases.**
9. Describe how the reactivity of the elements in group 7 change as you go down the group. **The reactivity decrease as you go down the group (increase atomic mass).**
10. Explain, in terms of electrons, why does reactivity change as you go down the group. **Reactions occur because of the electrons in the outer shell. As you go down the group the electron shells increase. The further away the outer shell is from the nucleus, the harder it is to attract an extra electron.**
11. Describe the reaction between a more reactive halogen and a less reactive halogen which is in an aqueous solution of its salt. **A displacement reaction will occur where the more reactive halogen will displace the least reactive halogen and bind with the salt.**
12. Complete the following equation and balance: $Cl_2 + 2KI \rightarrow 2KCl + I_2$
13. Complete the following equation and balance: $I_2 + KBr \rightarrow I_2 + KBr$ **(no reaction because Iodine is less reactive than bromine)**
14. Complete the following equation and balance: $F_2 + 2KCl \rightarrow 2KF + Cl_2$

5.2 Bonding, structure and the properties of matter

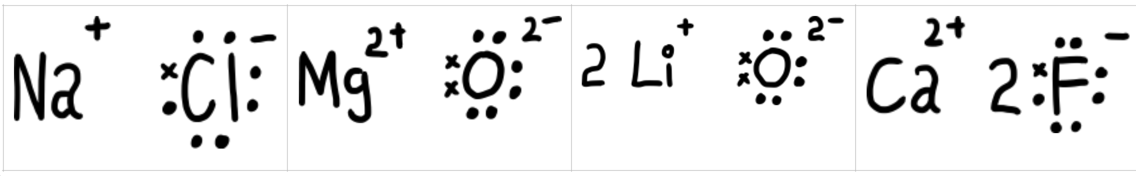
5.2.1 Chemical bonds, ionic, covalent and metallic

5.2.1.1 Chemical bonds

1. Name three types of strong chemical bonds. **Ionic, covalent and metallic.**
2. Describe the role of electrons in an ionic bond. **Electrons are either donated or received.**
3. Describe the role of electrons in a covalent bond. **Electrons are shared.**
4. Describe the role electrons in a metallic bond. **Outer shell electrons are delocalised.**
5. Which bonding occurs in compounds formed from non-metals? **Covalent.**
6. Which bonding occurs in metallic elements and alloys? **Metallic.**
7. Which bonding occurs in compounds formed from metals and non-metals? **Ionic.**

5.2.1.2 Ionic bonding

1. What happens when a metal reacts with a non-metal? **Electrons in the outer shell of the metal atom are transferred.**
2. What type of ions do metals form? **Positively charged ions.**
3. What type of ions do non-metals form? **Negatively charged ions.**
4. What is the charge of ions produce by group 1 elements? **+1.**
5. What is the charge of ions produce by group 2 elements? **+2.**
6. What is the charge of ions produce by group 6 elements? **-2.**
7. What is the charge of ions produce by group 7 elements? **-1.**
8. What electronic structure do the ions produced have? **The ions have the electronic structure of a noble gas (group 0)**
9. Draw a dot and cross diagram for sodium chloride
10. Draw a dot and cross diagram for magnesium oxide
11. Draw a dot and cross diagram for lithium oxide.
12. Draw a dot and cross diagram for calcium fluoride.

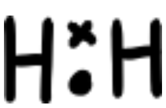
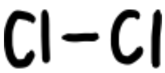
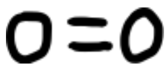
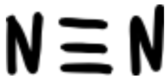
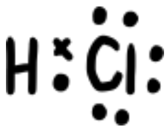
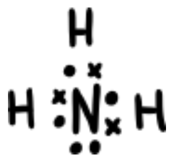
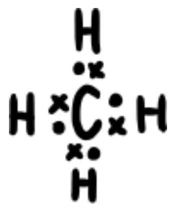


5.2.1.3 Ionic compounds

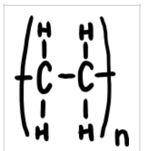
1. Describe the structure of an ionic compound. **A giant structure of ions**
2. What are forces of attraction in ionic bonding called? **Ionic structures are held in place by strong electrostatic forces of attraction between oppositely charged ions.**
3. What is the empirical formula for potassium chloride (group 1 + group 7)? **KCl**
4. What is the empirical formula for beryllium oxide (group 2 + group 6)? **BeO**
5. What is the empirical formula for sodium oxide? (group 1 + group 6)? **Na_2O**
6. What is the empirical formula for magnesium iodide (group 2 + group 7)? **MgI_2**

5.2.1.4	Covalent bonding

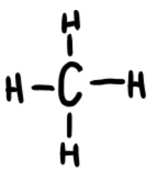
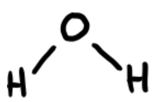
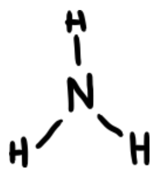
1. How are covalent bonds formed? **They are formed when atoms share electrons.**
2. Name two covalent compounds which are simple molecules. **Ammonia, water, methane or hydrogen chloride.**
3. Name a covalent compound which is a very large molecule. **Polymers**
4. Name two covalently bonded substances which form giant covalent structures. **Diamond and silicon dioxide.**
5. Draw a dot and cross diagram for hydrogen.
6. Draw a stick diagram for chlorine.
7. Draw a stick diagram for oxygen.
8. Draw a stick diagram for nitrogen.
9. Draw a dot and cross diagram for hydrogen chloride.
10. Draw a dot and cross diagram for water.
11. Draw a dot and cross diagram for ammonia.
12. Draw a dot and cross diagram for methane.

							
Q5	Q6	Q7	Q8	Q9	Q10	Q11	Q12

13. What is the chemical formula for hydrogen. H_2
14. What is the chemical formula for chlorine. Cl_2
15. What is the chemical formula for nitrogen. N_2
16. What is the chemical formula for hydrogen chloride. HCl
17. What is the chemical formula for water. H_2O
18. What is the chemical formula for ammonia. NH_3
19. What is the chemical formula for methane. CH_4
20. Draw the structure of poly(ethene).



21. Write the formula of the following structures.

		
CH_4	H_2O	NH_3

5.2.1.5	Metallic bonding

1. Describe the structure of metals. **Metals consist of giant structures of atoms arranged in a regular pattern.**
2. What are delocalised electrons? **Outer shell electrons of metal atoms which are free to move through the whole structure.**
3. How are metallic bonds formed? **The sharing of delocalised electrons gives rise to strong metallic bonds.**

5.2.2	How bonding and structure are related to the properties of substances
5.2.2.1	The three states of matter

1. Name the three states of matter. **Solid, liquid and gas.**
2. Describe the structure of solids in terms of particles. **Particles are close together in fixed positions and form a regular structure. There are strong forces of attraction between the particles. The particles vibrate in position.**
3. Describe the structure of liquids in terms of particles. **Particles are close together but can move over each other. There are weak forces of attraction between the particles. The particles are constantly moving.**
4. Describe the structure of gases in terms of particles. **Particles are far apart. There are very weak forces of attraction between the particles. Particles move constantly in straight lines.**
5. What term describes solid turning into liquid at a specific temperature. **Melting**
6. What term describes liquid turning into gas at a specific temperature. **Boiling**
7. What term describes gas turning into liquid at a specific temperature. **Condensation**
8. What term describes liquid turning into solid at a specific temperature. **Freezing**
9. What term describes solid turning into gas at a specific temperature. **Sublimation**
10. Bromine has a melting point of -7°C and a boiling point of 59°C. What state is it at 75°C? **Gas**
11. Why does a single atom not have a state of matter? **Atoms themselves do not have the bulk properties of materials**
12. What are the limitations of particle theory (HT only). **Particle theory represents particles as solid in elastic spheres which have no forces between them.**
13. Describe the energy required to change states. **The energy required to change state depends on the strength of forces between the particles.**
14. Describe how the forces between particles affects the melting points and boiling points. **The stronger the forces between the particles the higher the melting and boiling points.**

5.2.2.2	State symbols

1. What does the state symbol (aq) represent? **Aqueous (dissolved)**
2. What does the state symbol (l) represent? **Liquid.**
3. What does the state symbol (g) represent? **Gas.**
4. What does the state symbol (s) represent? **Solid.**

5.2.2.3 Properties of ionic compounds

1. Describe the structure of an ionic compound. **Ionic compounds have regular structures called giant ionic lattices. The are strong forces of attraction between oppositely charged ions.**
2. Describe the general properties of ionic compounds. **Ionic compounds have high melting points and boiling points because a large amount of energy is required to break the multitude of strong bonds.**
3. What can happen if ionic compounds are melted or dissolved in water? **The ionic compounds will conduct electricity.**

5.2.2.4 Properties of small molecules

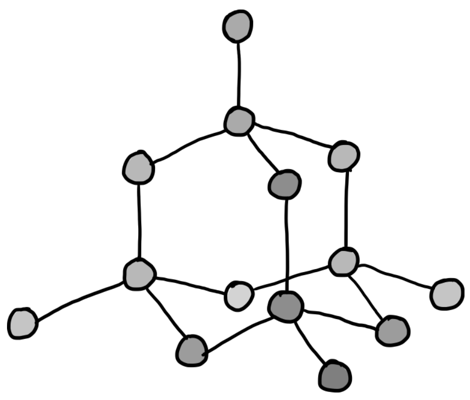
1. What type of bonding is found in small molecules? **Covalent bonding.**
2. Describe the general properties of small molecules. **Small molecules are usually gases or liquids which have low melting and boiling points.**
3. What are the forces of interaction between small molecules called? **Weak intermolecular forces.**
4. Describe what happens to small molecules when they melt or boil. **The weak intermolecular forces between the molecules are broken, not the covalent bonds between the atoms.**

5.2.2.5 Polymers

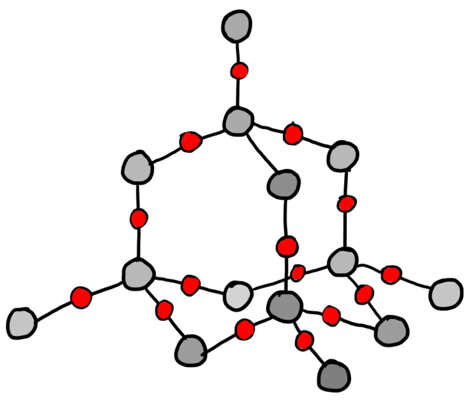
1. What is a monomer? **A single unit of a polymer.**
2. What is a polymer? **Polymers are very large molecules made of repeating units.**
3. What is the role of strong covalent bonds in a polymer? **Strong covalent bonds link the atoms together in polymers.**
4. How do intermolecular forces affect the properties of polymers? **The intermolecular forces between polymers are relatively strong so polymers are usually solids at room temperature.**

5.2.2.6 Giant covalent structures

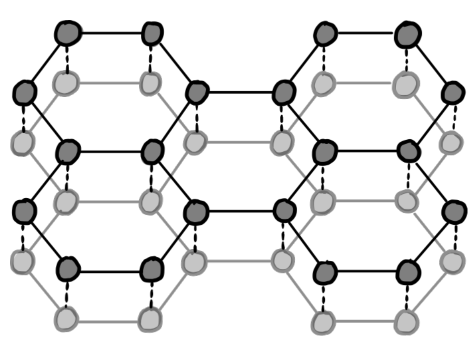
1. Describe the general structure of giant covalent structures. **The atoms are in the structures are linked to other atoms by strong covalent bonds.**
2. Describe the properties of giant covalent structures. **They are solids with high melting points.**
3. Give three examples of giant covalent structures. **Diamond, graphite and silicon dioxide (silica)**
4. Name the following giant covalent structures.



Diamond



Silicon Oxide



Graphite

5.2.2.7	Properties of metal and alloys

1. What is an alloy? **An alloy is a mixture of metals e.g. 18 carat Gold is a mixture of gold and silver.**
2. Describe the general structure of metals and alloys. **Giant structures of atoms with strong metallic bonding.**
3. What are the general properties of metals? **Metals are usually solid at room temperature and have high melting points and boiling points**
4. How are the atoms arranged in pure metals? **Atoms in metals are arranged in layers. These layers can slide over each other.**
5. What properties of pure metals are a result of the arrangement of atoms? **Metals can be bent and shaped.**
6. Explain why alloys are harder than pure metals. **Alloys are harder because the mixture of metal atoms causes a distortion of the layers which prevents them sliding easily over each other.**

5.2.2.8	Metals as conductors

1. Why are metals good conductors of electricity? **Delocalised electrons carry charge through the metal.**
2. Why are metals good conductors of thermal energy? **Delocalised electrons transfer energy through the metal.**

5.2.3	Structure and bonding of carbon
5.2.3.1	Diamond

1. How many covalent bonds does each carbon atom form in diamond? **4.**
2. What type of structure is diamond? **A giant covalent structure**
3. Describe three properties of diamond. **Very hard, a very high melting point and it does not conduct electricity.**

5.2.3.2	Graphite

1. How many covalent bonds does each carbon atom form in graphite? **3.**
2. How many delocalised electrons does each carbon atom in graphite have? **1.**
3. Describe the structure of graphite. **Layers of hexagonal rings formed by covalent bonds between the carbon atoms. There are no covalent bonds between the layers which means they are free to slide over each other.**
4. Why does graphite have similar properties to metals? **Delocalised electrons.**

5.2.3.3	Graphene and fullerenes

1. Describe the structure of graphene. **A single layer of graphite.**
2. What is graphene useful for? **Electronics and composites.**
3. How many delocalised electrons does each carbon atom in graphene have? **1.**
4. What is a fullerene? **Molecules of carbon atoms with hollow shapes.**
5. What is the structure of fullerene based on? **Hexagonal rings of six carbon atoms, but they may also have rings of 5 or carbon atoms.**
6. What was the first fullerene to be discovered? **Buckiminsterfullerene (C_{60})**
7. What are cylindrical fullerenes called? **Nanotubes**
8. What is the ratio between the length and diameter in cylindrical fullerenes. **High length to diameter ratios (long and thin)**
9. What are cylindrical fullerenes useful for? **Their properties make them useful for nanotechnology, electronics and materials.**

5.3 Quantitative chemistry

5.3.1 Chemical measurements, conservation of mass and the quantitative interpretation of chemical reactions

5.3.1.1 Conservation of mass and balanced chemical equations

1. What is the law of conservation of mass? **As no atoms are lost or made in a chemical reaction, the mass of the products will equal the mass of the reactants.**
2. State the numbers of atoms for each element in H_2O. **Hydrogen = 2, Oxygen = 1.**
3. State the number of atoms for each element in NH_3. **Nitrogen = 1, Hydrogen = 3.**
4. State the number of atoms in $Ca(OH)_2$. **5**
5. State the number of atoms for each element in $Ca(OH)_2$. **Calcium = 1, Oxygen = 2, Hydrogen = 2.**
6. Balance $H_2 + O_2 \rightarrow H_2O$. **$2H_2 + O_2 \rightarrow 2H_2O$**
7. Balance $Cl_2 + KI \rightarrow KCl + I_2$. **$Cl_2 + 2KI \rightarrow 2KCl + I_2$**

5.3.1.2 Relative formula mass

1. What is the relative formula mass of a compound? **The sum of the relative atomic masses of the atoms in the compound.**
2. What is the relative formula mass of H_2SO_4? **H = 1 x 2 = 2; S = 32; O = 16 x 4 = 64; = 98**
3. What is the relative formula mass of Na_2CO_3? **Na = 23 x 2 = 46; C = 12; O = 16 x 3 = 48; =106**
4. What is the relative formula mass of $Ca(OH)_2$? **Ca = 40; O = 16 x 2 = 32; H = 1 x 2 = 2; = 74**

5.3.1.3 Mass changes when a reactant or product is a gas

1. Explain why some reactions may seem to involve a change in mass? **In a non-enclosed system one of the reactants or products may be a gas and its mass has not been measured.**
2. Describe the mass changes that occur when a metal reacts with oxygen in a non-enclosed system. **When a metal reacts with oxygen the mass of the metal oxide will be more than the mass of the metal because of the addition of oxygen gas.**
3. Describe the mass changes that occur during the thermal decomposition of metal carbonates in a non-enclosed system. **When a metal carbonate decomposes the mass of the products will appear less than the mass of the reactants because carbon dioxide gas is given off.**

5.3.1.4 Chemical measurements

1. Define uncertainty. **Uncertainty is the amount of error in your measurements.**
2. How do you calculate the range of a set of measurements? **The range is the highest repeat value minus the lowest repeat value.**
3. What does a large range of a set of measurements about the mean signify? **A large range suggest the measurements are imprecise and there is a large uncertainty about the results.**
4. What is the formula to calculate the uncertainty about the mean. **Uncertainty = range / 2.**
5. Calculate the uncertainty of the following repeat values.

Experiment	1	2	3	Mean
Volume of CO_2 (mL)	20.0	20.1	19.8	20.0

Range = 20.1 - 19.8 = 0.3mL; Uncertainty = 0.3 / 2 = 0.15mL. Uncertainty of the mean = 20.0 ± 0.15

5.3.2	Use of amount of substance in relation to masses of pure substances
5.3.2.1	Moles (HT)

1. What are chemical amounts measured using? **Moles**
2. What is the symbol for the mole? **Mol**
3. How is the relative formula mass of a substance linked to the mole? **The mass of 1 mole of a substance is equal to its relative formula mass in grams. E.g. M_r of carbon = 12; therefore 1 mole of carbon has a mass of 12g.**
4. Compare the number of particles in one mole of carbon (C) with the number of particles in one mole of carbon dioxide (CO_2). **The number of particles in one mole of carbon is equal to the number of particles in one mole of carbon dioxide.**
5. What is the value of the Avogadro constant? **6.02×10^{23} per mole**
6. What is the formula that links the number of moles, relative formula (or atomic) mass and mass in grams? **Number of moles = mass in grams / M_r of the substance**
7. How many moles are there in 44g of H_2O? **Number of moles = 44 / (1 x 2) + 16; Number of moles = 44 / 18; Number of moles = 2.4 mol**
8. Calculate the mass of 0.4mol of CO_2. **Rearrange the equation; mass = number of moles x M_r of the substance; mass = 0.4 x (12 + (16 x2); mass = 0.4 x 44; mass = 17.6g**

5.3.2.2	Amounts of substances in equations (HT)

1. Describe the following equation in terms of moles: $Mg + 2HCl \rightarrow MgCl_2 + H_2$. **1 mole of magnesium reacts with 2 moles of hydrochloric acid to form 1 mole of magnesium chloride and 1 mole of hydrogen.**
2. What is the formula to calculate the percentage mass of an element in a compound? **Percentage mass of an element in a compound = (A_r x number of atoms of the element / M_r of the compound) x 100**

5.3.2.3	Using moles to balance equations (HT)

1. Describe how you would balance an equation using the masses of the products and reactants. **Divide the mass of each substance by its relative formula mass to find the number of moles of each substance. Divide the number of moles of each substance by the smallest number of moles in the reaction. If the any of the numbers are not whole numbers, multiply all the numbers so that they become whole numbers.**
2. 12g of magnesium (Mg) react with 8g of oxygen (O_2) to produce 20g of magnesium oxide (MgO). Write a balanced equation for the reaction. **Number of moles of magnesium = 12 / 24 = 0.5 moles ; number of moles of oxygen = 8 / (16 x 2); number of moles of oxygen = 8 / 32 = 0.25 moles. Number of moles of MgO = 20 / 40 = 0.5 moles. Divide each substance by the smallest number of moles in the reaction (oxygen = 0.25); Mg = 0.5 / 0.25 = 2; O_2 = 0.25 /0.25 = 1; MgO = 0.5 / 0.25 = 2. The balanced equation for the reaction is: $2Mg + O_2 \rightarrow 2MgO$**

5.3.2.4	Limiting reactants (HT)

1. What is a limiting reactant? **The limiting reactant limits the amount of product made in a reaction.**
2. Why is it common to use an excess of one of the reactants in a chemical reaction? **To ensure that the other reactants involved are used up.**
3. What does the mass of a product formed in a chemical reaction depend upon? **The mass of the limiting reactant.**
4. What are the steps required to calculate the amount of aluminium oxide formed when 135g of aluminium is reacted with an excess of oxygen. **Write out a balanced equation. $4Al + 3O_2 \rightarrow 2Al_2O_3$. Calculate the relative formula masses of aluminium (27) and aluminium oxide (102). Calculate the number of moles from the substance you are given the mass of (aluminium). Moles of aluminium = mass / M_r: Moles of aluminium = 135 / 27 = 5. Using the balanced equation calculate the number of moles of the other substance. 4 moles of aluminium react to produce 2 moles of aluminium oxide (half the number of moles). So 5 moles of aluminium would produce 2.5 moles of aluminium oxide. Use the number of moles to calculate the mass. Mass of aluminium oxide = moles x M_r: mass of aluminium oxide = 2.5 x 102 = 255g**
5. Calculate the mass of potassium chloride when 24g of potassium iodide reacts with an excess of chlorine? **Write out a balanced equation. $Cl_2 + 2KI \rightarrow 2KCl + I_2$. Calculate the relative formula masses of potassium chloride(74.5) and potassium iodide (166). Calculate the number of moles from the substance you are given the mass of (potassium iodide). Moles of potassium iodide = mass / M_r: Moles of potassium iodide = 24 / 166 = 0.14 moles. Using the balanced equation calculate the number of moles of the other substance. 2 moles of potassium iodide react to produce 2 moles of potassium chloride. So 0.14 moles of potassium iodide would produce 0.14 moles of potassium chloride. Use the number of moles to calculate the mass. Mass of potassium chloride = moles x M_r: mass of potassium chloride = 0.14 x 74.5 = 10.4g**

5.3.2.5	Concentration of solutions

1. What is a solution? **A solution consists of a solute (solid) dissolved in a solvent (liquid).**
2. What is a solute? **The solid part of a solution which has been dissolved.**
3. What is a solvent? **The liquid part of the solution.**
4. What is the formula to calculate the concentration of a solution? **Concentration = mass of the solute (g) / volume of solvent (dm^3).**
5. How many cm^3 in 1 dm^3? **1000cm^3 = 1 dm^3**
6. What is the concentration of a salt solution when 20g of salt is dissolved in 500cm^3 of water? **Convert 500cm^3 into 0.5 dm^3. Concentration = mass of solute / volume of solvent: concentration = 20 / 0.5 = 40g/dm^3**
7. Explain how the concentration of the solution is related to the mass of the solute and the volume of the solvent (HT). **The more solute added for a given volume the higher the concentration of a solution. The more solvent added for a given mass of solute the lower the concentration of the solution.**

5.4	Chemical changes
5.4.1	Reactivity of metals
5.4.1.1	Metal oxides

1. What are the products when metals react with oxygen? **Metal oxides.**
2. What type of reactions occur when metals react with oxygen? **Oxidation.**
3. Define oxidation with reference to oxygen. **Oxidation occurs when a substance gains oxygen.**
4. Define reduction with reference to oxygen. **Reduction occurs when a substance loses oxygen.**

5.4.1.2	The reactivity series

1. What do metals form when they react with other substances? **Metals form positive ions.**
2. What is the reactivity of a metal related to? **The reactivity of a metal is related to the ease by which it can form ions (lose electrons).**
3. Put zinc, lithium, potassium, copper, iron, calcium, sodium and magnesium in order of reactivity (most reactive first). **Potassium, sodium, lithium, calcium, magnesium, zinc, iron and copper.**
4. Which two non-metals are often placed in the reactivity series. **Carbon and hydrogen.**
5. Write out the reactivity series including the two non-metals. **Potassium, sodium, lithium, calcium, magnesium, <u>carbon</u>, zinc, iron, <u>hydrogen</u> and copper**
6. What type of reaction occurs between a reactive metal and a less reactive metal compound? **A displacement reaction.**
7. What is the general equation to show the reaction of metal with water. **Metal + water → metal hydroxide + hydrogen.**
8. Write out a balanced equation to show the reaction between sodium and water. **$2Na + H_2O \rightarrow Na_2O + H_2$.**
9. What is the general equation to show the reaction of metal with acid. **Metal + acid → salt + water.**
10. Write out a balanced equation to show the reaction between calcium and sulfuric acid. **$Ca + H_2SO_4 \rightarrow CaSO_4 + H_2O$.**

5.4.1.3	Extraction of metals and reduction

1. How are most metals found in the Earth? **Metals are usually found as compounds in the Earth.**
2. What is a metal ore? **A metal compound which is mined.**
3. Why is gold found as a metal in the Earth? **Because it is unreactive and does not form compounds.**
4. How can iron, zinc and copper be extracted from their oxides? **These metals are extracted from their oxides by reduction with carbon. This works because these metals are less reactive than carbon.**
5. Identify which substances have been oxidised and which substances have been reduced in the following equation: $2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$. **Iron oxide ($Fe_2O_3$) is reduced (loses oxygen). Carbon (C) is oxidised (gains oxygen).**

5.4.1.4	Oxidation and reduction in terms of electrons (HT)

1. Define oxidation with reference to electrons. **Oxidation is electron loss.**
2. Define reduction with reference to electrons. **Reduction is electron gain.**
3. Write out the ionic equation for the following displacement reaction: $Fe + CuSO_4 \rightarrow FeSO_4 + Cu$. **$Fe_{(s)} + Cu^{2+}{}_{(aq)} \rightarrow Cu_{(s)} + Fe^{2+}{}_{(aq)}$.**
4. Define the term "spectator ion". **An ion that takes no part in the reaction.**
5. What are the spectator ions in the following equation: $Mg + FeSO_4 \rightarrow MgSO_4 + Fe$. **$SO_4{}^{2-}$.**
6. Write out the half equations for the following reaction $Fe + 2HCl \rightarrow FeCl_2 + H_2$. **$Fe \rightarrow Fe^{2+} + 2e^-$; $2H+ + 2e^- \rightarrow H_2$.**

5.4.2	Reactions of acids

5.4.2.1	Reactions of acids with metals

1. What are the products of a reaction between some metals and acid? **Salt and hydrogen.**
2. Describe the type of reaction that occurs between metals and acids. **Redox reactions.**
3. Which substances are oxidised in the reaction between metals and acids? **Metals**
4. Which substances are reduced in the reaction between metals and acids? **Hydrogen**
5. State which species are oxidised and which are reduced in the following reactions:

Reaction	Oxidised	Reduced
magnesium + hydrochloric acid	**Magnesium**	**Hydrogen**
zinc + hydrochloric acid	**Zinc**	**Hydrogen**
iron + hydrochloric acid	**Iron**	**Hydrogen**
magnesium + sulfuric acid	**Magnesium**	**Hydrogen**
zinc + sulfuric acid	**Zinc**	**Hydrogen**
iron + sulfuric acid	**Iron**	**Hydrogen**

5.4.2.2	Neutralisation of acids and salt production

1. Define alkali. **Soluble metal hydroxides.**
2. Define base. **Insoluble metal hydroxides and metal oxides.**
3. How are acids neutralised? **By reacting with alkalis or bases.**
4. What is the general equation to show the reaction between acids and metal oxides? **Metal oxide + acid → salt + water.**
5. What is the general equation to show the reaction between acids and metal hydroxides? **Metal hydroxide + acid → salt + water.**
6. What is the general equation to show the reaction between acids and metal carbonates? **Metal carbonates + acid → salt + water + carbon dioxide (bubbles!)**
7. The production of a particular salt depends on which two factors? **The acid used and the positive (metallic) ions in the alkali or base.**
8. Predict the products from a reaction between nitric acid and potassium hydroxide. **Potassium nitrate.**
9. Predict the products from a reaction between hydrochloric acid and calcium carbonate. **Calcium chloride.**
10. Predict the products from a reaction between sulfuric acid and copper oxide. **Copper sulfate.**
11. Deduce the formula of magnesium chloride using the ions Mg^{2+} and Cl^-. **$MgCl_2$.**
12. Deduce the formula of zinc sulfate using the ions Zn^{2+} and $SO_4{}^{2-}$. **$ZnSO_4$.**

5.4.2.3	Soluble salts

1. Describe how soluble salts can be made. **React an acid with an insoluble substance (e.g. metal, metal oxide, metal hydroxide or metal carbonate). Add the solid till no more reacts. The unreacted powder will be visible in the beaker. Excess acid is filtered off and the solution of the new salt is collected.**
2. How can salt solutions be used to form solid salts? **Crystallisation.**

5.4.2.4	The pH scale and neutralisation

1. What ions do acids produce in aqueous solution? **H^+.**
2. What ions do alkalis form in aqueous solution. **OH^-.**
3. What does pH mean? **per Hydrogen.**
4. What is the pH scale a measure of? **The acidity or alkalinity of a solution.**
5. What numbers on the scale represent acids? **1 - 6.**
6. What numbers on the pH scale represent alkalis? **8 - 14.**
7. What does pH 7 represent on the scale? **Neutral.**
8. Name two ways the pH of a solution can be measured. **Using a wide range indicator (e.g. universal indicator) or a pH probe.**
9. What type of reaction occurs between an acid and an alkali. **Neutralisation.**
10. Write an ionic equation for the reaction between an acid and an alkali. **$H^+ + OH^- \rightarrow H_2O$.**

5.4.2.5	Strong and weak acids (HT)

1. What is a strong acid? **A strong acid is completely ionised in solution.**
2. Give three examples of strong acids. **Hydrochloric, nitric and sulfuric acids.**
3. What is a weak acid? **A weak acid is only partially ionised in solution.**
4. Give three examples of weak acids. **Ethanol, citric and carbonic acids.**
5. Explain the difference between the terms concentrated and dilute acids and weak and strong acids. **Dilute and concentrated refer to the amount of substance in a solution (e.g. the number of moles). Weak and strong refer to the amount of ionisation that has occurred.**
6. For a given concentration of an aqueous solution, what is the relationship between the strength of acid and the pH. **The stronger the acid, the lower the pH.**
7. Describe how the pH scale is linked to the hydrogen ion concentration. **As pH concentration decreases by one unit (e.g. from 5 to 4), the hydrogen ion concentration increases by by a factor of ten (e.g. from 10 to 100).**

5.4.3	Electrolysis
5.4.3.1	The process of electrolysis

1. What type of compounds can be electrolysed? **Ionic compounds.**
2. What state must the compound be in for electrolysis to take place? **Molten (liquid) or in an aqueous solution.**
3. What is an electrolyte? **Liquids or solutions that are able to conduct electricity.**
4. What is the negative electrode called? **Cathode.**
5. What is the positive electrode called? **Anode.**
6. What happens when an electric current is passed through an electrolyte? **Positive charged ions move to the negative electrode (cathode). Negatively charged ions move towards the positive electrode (anode). The ions are gain or lose electrons at the electrodes producing elements.**

5.4.3.2 Electrolysis of molten ionic compounds.

1. What ions do lead bromide form when melted? **Pb^{2+} and Br^-.**
2. What type of electrodes are used during the electrolysis of lead bromide? **Inert electrodes.**
3. What happens when a molten ionic compound is electrolysed? **The metal ions are attracted to the negative electrode (cathode). The non-metal ions are attracted to the positive electrode (anode).**
4. What is formed at the positive electrode (anode) during the electrolysis of lead bromide? **Bromine.**
5. What is the ionic equation for the positive electrode (anode)? **$2Br^- \rightarrow Br_2 + 2e^-$.**
6. What is formed at the negative electrode (cathode) during the electrolysis of lead bromide? **Lead.**
7. What is the ionic equation for the negative electrode (cathode)? **$Pb2^+ + 2e^- \rightarrow Pb$**
8. What ions do zinc chloride form when melted? **$Zn2^+$ and Cl^-.**
9. Predict the products of the electrolysis of zinc chloride. **Chlorine will be formed at the positive electrode (anode) and zinc will be formed at the negative electrode (cathode).**

5.4.3.3 Using electrolysis to extract metals.

1. How can metals be extracted by electrolysis? **Metals can be extracted from molten compounds.**
2. Which type of metals are extracted by electrolysis? **Metals which are too reactive to be extracted by reduction with carbon. Metals which react with carbon.**
3. Why is aluminium sometimes called "solid electricity"? **Large amounts of energy are used in the extraction process to melt the compounds and to produce the electrical current for electrolysis.**
4. Describe how aluminium is manufactured. **Aluminium is extracted by the electrolysis of a mixture of aluminium oxide and cryolite using carbon as the positive electrode (anode).**
5. What is formed at the negative electrode (cathode) during the electrolysis of aluminium oxide? **Aluminium.**
6. What is the ionic equation for the negative electrode (cathode) during the electrolysis of aluminium oxide? **$Al^{3+} + 3e^- \rightarrow Al$.**
7. What is formed at the positive electrode (anode) during the electrolysis of aluminium oxide? **Oxygen.**
8. What is the ionic equation for the positive electrode (anode) during the electrolysis of aluminium oxide? **$2O^{2-} \rightarrow O_2 + 4e^-$.**
9. Why is cryolite used in this process? **To reduce the melting point of aluminium oxide.**
10. Why does the positive electrode (anode) have to be constantly replaced? **The oxygen formed at the positive electrode (anode) reacts with the carbon in the electrode.**

5.4.3.4 Electrolysis of aqueous solutions

1. What ions are present in all aqueous solutions? **H^+ and OH^-.**
2. What affects the ions discharged when an aqueous solution is electrolysed? **The relative reactivity of the elements involved.**
3. Explain what will be discharged from a negative electrode (cathode) during the electrolysis of an aqueous solution. **If the metal ion is less reactive than hydrogen the metal will be formed. Otherwise, hydrogen will be released.**
4. Explain what will be discharged from a positive electrode (anode) during the electrolysis of an aqueous solution. **Oxygen is produced unless the solution contains halide ions in which event a halogen will be produced.**
5. What are the ions found in a solution of copper sulfate ($CuSO_{4(aq)}$)? **H^+, OH^-, Cu^{2+} and SO_4^{2-}**
6. Explain what will happen when an aqueous solution of copper sulfate is electrolysed. **Copper is less reactive than hydrogen so copper is formed at the negative electrode (cathode). As copper sulfate contains no halide ions oxygen will be formed at the positive electrode (anode).**
7. What are the ions found in a solution of sodium chloride ($NaCl_{(aq)}$)? **H^+, OH^-, Na^+ and Cl^-**
8. Explain what will happen when an aqueous solution of sodium chloride is electrolysed. **Sodium is more reactive than hydrogen so copper is formed at the negative electrode (cathode). As there are halide ions present in solution (the chloride ions) chlorine gas will be given off at the positive electrode (anode).**

5.4.3.5	**Representation of reactions at electrodes as half equations (HT).**

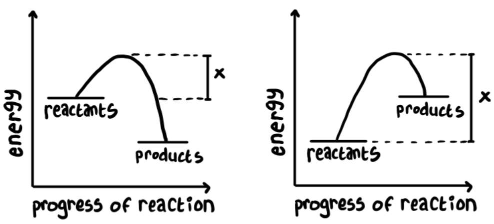
1. What happens at the cathode (negative electrode) during electrolysis? **Positive metal ions or hydrogen gain electrons.**
2. What type of reaction occurs at the cathode (negative electrode)? **Reduction.**
3. What happens at the anode (positive electrode) during electrolysis? **Negative non-metal ions lose electrons**
4. What type of reaction occurs at the anode (positive electrode)? **Oxidation.**
5. Write out the half equations for the electrolysis of copper sulfate solution. **At the negative electrode (cathode): $Cu^{2+} + 2e^- \rightarrow Cu$. At the positive electrode (anode): $4OH^- \rightarrow O_2 + 2H_2O + 4e^-$.**
6. Write out the half equations for the electrolysis of sodium chloride solution. **At the negative electrode (cathode): $2H^+ + 2e^- \rightarrow H_2$. At the positive electrode (anode): $2Cl^- \rightarrow Cl_2 + 2e^-$.**

5.5	Energy changes
5.5.1	Exothermic and endothermic reactions
5.5.1.1	Energy transfer during exothermic and endothermic reactions.

1. What happens to energy in chemical reactions? **Energy is conserved in chemical reactions.**
2. Compare the energy in the reactants and the products if energy is transferred to the surroundings during the reaction. **The product molecules have less energy than the reactant molecules.**
3. Compare the energy in the reactants and the products if energy is transferred from the surroundings during the reaction. **The reactant molecules have less energy than the reactant molecules.**
4. What is an exothermic reaction? **A reaction that transfers energy to the surroundings, increasing the temperature of the surroundings.**
5. Give three examples of exothermic reactions. **Combustion, oxidation reactions and neutralisation reactions.**
6. What is an endothermic reaction? **A reaction that transfers energy from the surroundings, decreasing the temperature of the surroundings.**
7. Give three examples of endothermic reactions. **Thermal decomposition, the reaction between citric acid and sodium hydrogencarbonate and some sports injury packs.**
8. Describe an experiment to measure the energy released by a neutralisation reaction. **Place 25mL of 0.25mol hydrochloric acid and 25mL of 0.25mol sodium hydroxide into separate beakers. Measure the temperature of both reagents and ensure they are equal. Place a polystyrene beaker in a large beaker and surround with cotton wool balls for insulation. Now add both chemicals to the polystyrene cup and cover with a lid with a thermometer through the middle. Measure the temperature of the mixture every 30 seconds and record the highest temperature reached. Repeat the experiment with 0.5mol and 1.0mol solutions. Plot the highest temperatures against concentration.**

5.5.1.2	Reaction profiles

1. Describe how chemical reactions can occur between particles. **A chemical reaction occurs when particles collide with each with sufficient energy.**
2. What is activation energy? **The minimum amount of energy the particles must have for a reaction to take place.**
3. What is a reaction profile? **A graph to show the change in energy between reactants and products over the course of a reaction.**
4. What is represented by X in the following diagrams? **Activation energy**



5.5.1.3	The energy changes of reactions (HT)

1. Describe a chemical reaction in terms of energy and bonds. **Energy needs to be supplied to break bonds in the reactants. Energy is released when bonds in the products are formed.**
2. What is the bond energy of a molecule? **The energy needed to break or form bonds between atoms.**
3. What reaction occurs when bonds are broken? **Endothermic**
4. What reaction occurs when bonds are made? **Exothermic**
5. How do you calculate the overall energy change for a reaction? **Calculate the difference between the sum of the energy needed to break the bonds of the reactants and the sum of the energy released when bonds of the products are formed.**
6. What is the overall energy change for an exothermic reaction? **The energy released from forming new bonds is greater than the energy needed to break existing bonds (e.g. the products have less energy than the reactants).**
7. What is the overall energy change for an endothermic reaction? **The energy needed to break existing bonds is greater than the energy needed to break existing bonds (e.g. the products have more energy than the reactants).**
8. Calculate the overall energy change for the following reaction: $H_2 + Cl_2 \rightarrow 2HCl$. Bond energies: H-H = 436 kJ/mol, Cl-Cl = 243 kJ/mol and H-Cl = 432 kJ/mol. State if the reaction is endothermic or exothermic. **Energy in = 436 (H-H) + 243 (Cl-Cl) = 679kJ/mol. Energy out = (2 x 432)(2 x H -Cl) = 864kJ/mol; Energy change = energy in - energy out = 679 - 864 = -185kJ/mol; the energy change is negative so the reaction is exothermic.**
9. Calculate the overall energy change for the following reaction: $2HBr \rightarrow H_2 + Br_2$. Bond energies: H-H = 436 kJ/mol, Br-Br = 193 kJ/mol and H-Br = 366 kJ/mol. State if the reaction is endothermic or exothermic. **Energy in = (2 x 366)(2 x H-Br) = 732 kJ/mol; Energy out = 436 (H-H) + 193 (Br-Br) = 629kJ/mol; Energy change = energy in - energy out = 732 - 629 = + 103kJ/mol; the energy change is positive so the reaction is endothermic.**

5.6	The rate and extent of chemical change
5.6.1	Rate of reaction
5.6.1.1	Calculating rates of reaction

1. How do you measure the rate of a chemical reaction? **By measuring the the quantity of the reactant as it is used or the quantity of the product as it is formed.**
2. What is the formula to calculate measure the mean rate of reaction from the reactants? **Mean rate of reaction = quantity of reactant used / time taken.**
3. What is the formula to calculate the mean rate of reaction from the products? **Mean rate of reaction = quantity of product formed / time taken.**
4. What three quantities can be used to measure the the quantity of the product or reactant? **Mass in grams, volume in cm^3 or moles.**
5. What are the units for rate of reaction? **g/s, cm^3/s or mol/s**
6. What is a tangent? **A straight line that touches the curve but does not cross it.**
7. What is the slope of a tangent used to calculate? **The rate of reaction at a particular point.**

5.6.1.2	Factors which affect the rate of chemical reaction

1. Name five factors which affect the rate of chemical reaction? **Concentration of reactants, pressure of reacting gases, surface area of reactants, temperature of the reaction and the presence of catalysts.**
2. Describe the effect of changing these factors on the rate of chemical reaction. **Increasing concentration of reactants increases the number of colliding molecules. Increasing pressure of reacting gases increases the number of collisions between molecules. Large surface area of reactants increases number of colliding molecules. Increasing temperature of the reaction increases number of collisions. The presence of catalysts lowers the activation energy.**
3. What is turbidity? **A measure of how cloudy a solution is.**
4. Describe an experiment involving colour change or turbidity to measure the effect of concentration on the rate of reaction. **Add a set volume and concentration of sodium thiosulfate to a conical flask. Place the flask on a black cross drawn on paper. Add a known volume of hydrochloric acid to the conical flask. Start the stopwatch. Time how long it takes for the cross to disappear (due to turbidity of the reacting solution). Repeat the reaction for different concentrations of hydrochloric acid.**
5. Describe an experiment involving the collection of gas to measure the effect of concentration on the rate of reaction. **Add a set volume and concentration of hydrochloric acid to a conical flask. Add a set mass of magnesium to the flask. Place a bung with a tube attached to a gas syringe on top of the flask. Start the timer. Note the volume of gas produced every 20 seconds until the reaction stops. Repeat the experiment using different concentrations of hydrochloric acid.**

5.6.1.3	Collision theory and activation energy

1. Explain collision theory. **Chemical reactions occur when reacting particles collide with each other and with sufficient energy.**
2. What is activation energy. **The minimum energy particles must have to react.**
3. Describe the effect of increasing the concentration of the reactants on the rate of reaction using collision theory. **Increasing the concentration means there are more particles and as a result there are more collisions. This will increase the rate of reaction.**
4. Describe the effect of increasing the pressure of the reactants on the rate of reaction using collision theory. **Increasing the pressure means there are more particles near each other so they are more likely to collide. This increases the rate of reaction.**
5. Describe the effect of increasing the surface area of the reactants on the rate of reaction using collision theory. **Increasing the surface area (by breaking into smaller pieces) increases the number of particles available to collide. This increases the rate of reaction.**
6. How does the surface area to volume ratio affect the rate of reaction? **Breaking up a solid into smaller pieces increases the surface area to volume ratio which will speed up the rate of reaction.**
7. Describe the effect of increasing the temperature of the reactants on the rate of reaction using collision theory. **Increasing the temperature increases the speed of the particles. As they are moving faster they will collide more frequently and with more energy.**

5.6.1.4	Catalysts

1. What is a catalyst? **A substance that changes the rate of a chemical reaction but is not used up in the process.**
2. How do catalysts work? **Catalysts work by providing a pathway for the reaction which has a lower activation energy.**
3. What are enzymes? **Enzymes are biological catalysts found in living organisms.**
4. What is a reaction profile? **A reaction profile is a graph which shows the levels of energy required over the course of a reaction.**

5.6.2	Reversible reactions and dynamic equilibrium
5.6.2.1	Reversible reactions

1. What is a reversible reaction? **A reaction where the products of the reaction can react to produce the original reactants.**
2. Represent a reversible reaction with the reactants A and B and the products C and D. **$A + B \rightleftharpoons C + D$.**
3. How can the direction of a reversible reaction be changed? **By changing the conditions of the reaction.**
4. How can the direction of the decomposition of ammonium chloride into ammonia and hydrogen chloride be changed? **Ammonium chloride $\rightleftharpoons$ ammonia + hydrogen chloride. If heated the reaction will move to the right (produce ammonia and hydrogen chloride). If cooled the reaction will move to the left (produce ammonium chloride).**

5.6.2.2	Energy changes and reversible reactions

1. If a reaction is exothermic in one direction what will it be in the opposite direction? **Endothermic.**
2. What does hydrated mean? **The compound is associated with water molecules.**
3. What does anhydrous mean? **The compound has no water molecules associated with it.**
4. What are energy changes associated with the reversible reaction of hydrated copper sulfate changing into anhydrous copper sulfate and water. **Hydrated copper sulfate $\rightleftharpoons$ anhydrous copper sulfate + water. The reaction is endothermic to the right (formation of anhydrous copper sulfate) and exothermic to the right (formation of hydrated copper sulfate).**

5.6.2.3	Equilibrium

1. What is equilibrium? **Equilibrium occurs when the forward and reverse reactions occur at exactly the same rate (no apparent change).**
2. What is a closed system? **Apparatus which prevents the escape of the products or reactants of a reaction.**

5.6.2.4	The effect of changing conditions on equilibrium (HT)

1. What do the relative amounts of reactants and products at equilibrium depend upon? **They depend on the conditions of the reaction.**
2. Describe what happens to a system at equilibrium when a change is made to one of the conditions. **The system will respond to counteract the change.**
3. What is Le Chatelier's principle? **If the conditions of a reaction in equilibrium are changed the system will respond to counteract the change. This results in a new point of equilibrium.**
4. Where will the position of equilibrium be when ammonium chloride is heated? **To the right (ammonia and hydrogen chloride formed).**
5. Where will the position of equilibrium be when ammonia and hydrogen chloride are cooled? **To the left (ammonium chloride formed).**
6. Where will the position of equilibrium be when hydrated copper sulfate is heated? **To the right (anhydrous copper sulfate and water formed).**

5.6.2.5	The effect of changing concentration on equilibrium (HT)

1. Describe what happens if the concentration of a reactant or product in a reversible reaction is changed. **The system will no longer be in equilibrium. The concentrations of the substance will change until equilibrium is reached again.**
2. Describe the effect of increasing the concentration of a reactant in a reversible reaction. **More products will be formed until equilibrium is reached.**
3. Describe the effect of decreasing the concentration of a product in a reversible reaction. **More reactants will react until equilibrium is reached.**
4. What happens is more nitrogen is added in the following reaction: $N_2 + 3H_2 \rightleftharpoons 2NH_3$? **More is NH_3 produced.**
5. What happens is more hydrogen is added in the following reaction: $N_2 + 3H_2 \rightleftharpoons 2NH_3$? **More is NH_3 produced.**
6. What happens if the concentration of ammonia is decreased in the following reaction? $N_2 + 3H_2 \rightleftharpoons 2NH_3$? **More $N_2 + 3H_2$ will react.**

5.6.2.6	The effect of temperature changes on equilibrium (HT)

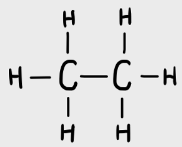
1. Describe the effect of increasing the temperature of a system at equilibrium on the relative amount of products for an exothermic reaction. **The relative amount of products at equilibrium decrease.**
2. Describe the effect of increasing the temperature of a system at equilibrium on the relative amount of products for an endothermic reaction. **The relative amount of products at equilibrium increase.**
3. Describe the effect of decreasing the temperature of a system at equilibrium on the relative amount of products for an exothermic reaction. **The relative amount of products at equilibrium increase.**
4. Describe the effect of decreasing the temperature of a system at equilibrium on the relative amount of products for an endothermic reaction. **The relative amount of products at equilibrium decrease.**
5. The following reaction is exothermic in the forward direction and endothermic in the opposite direction: $N_2 + 3H_2 \rightleftharpoons 2NH_3$. Describe what happens if the temperature is decreased. **The relative amounts of NH_3 will increase at equilibrium.**
6. The following reaction is exothermic in the forward direction and endothermic in the opposite direction: $N_2 + 3H_2 \rightleftharpoons 2NH_3$. Describe what happens if the temperature is increased. **The relative amounts of NH_3 will decrease at equilibrium.**

5.6.2.7	The effect of pressure changes on equilibrium (HT)

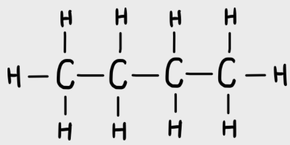
1. Describe the effect of decreasing the pressure of a gaseous reaction at equilibrium. **A decrease in pressure will cause the reaction to shift to the side with the larger number of molecules (from the balanced symbol equation).**
2. Describe the effect of increasing the pressure of a gaseous reaction at equilibrium. **An increase in pressure will cause the reaction to shift to the side with the smaller number of molecules (from the balanced symbol equation).**
3. Describe what will happen to the equilibrium position if the pressure is increased for the following reaction: $N_2 + 3H_2 \rightleftharpoons 2NH_3$. **The are less molecules on the left (2 moles compared to 4 moles) so the reaction will shift to the right (more NH_3).**
4. Describe what will happen to the equilibrium position if the pressure is decreased for the following reaction: $N_2 + 3H_2 \rightleftharpoons 2NH_3$. **The are more molecules on the right (4 moles compared to 2 moles) so the reaction will shift to the left (more $N_2 + H_2$).**

5.7	Organic chemistry
5.7.1	Carbon compounds as fuels and feedstock.
5.7.1.1	Crude oil, hydrocarbons and alkanes.

1. What is crude oil made from? **Crude oil is the remains of biomass consisting mainly of plankton buried in mud.**
2. What is crude oil? **Crude oil is a mixture of a very large number of compounds, mainly hydrocarbons.**
3. What are hydrocarbons? **Hydrocarbons are molecules made up of carbon and hydrogen only.**
4. What type of molecules are most of the hydrocarbons found in crude oil? **Alkanes.**
5. What is the general formula for alkanes? **C_nH_{2n+2}.**
6. Name the first four members of the alkanes. **Methane (C_1), ethane (C_2), propane (C_3) and butane(C_4).**
7. What is the formula of methane? **CH_4.**
8. Draw the structure of ethane.



9. What is the formula of propane? **C_3H_8.**
10. Draw the structure of butane.



5.7.1.2	Fractional distillation and petrochemicals.

1. What is a crude oil fraction? **Groups of molecules with a similar number of carbon atoms.**
2. How can crude oil be separated into fractions? **Fractional distillation.**
3. How does the petrochemical industry use different fractions? **The petrochemical industry processes fractions for fuels and feedstock.**
4. What are the five main crude oil fractions? **Petrol, diesel oil, kerosene, heavy fuel oil and liquified petroleum gas.**
5. Name four useful materials produced by the petrochemical industry? **Solvents, lubricants, polymers and detergents.**
6. Why are there are vast array of natural and synthetic carbon compounds? **Carbon atoms have the ability to form families of similar compounds.**
7. Explain how fractional distillation works. **Crude oil is heated until most of it turns to gas. The gases enter the base of the fractionating column. The fractionating column has a temperature gradient. It is hot at the bottom and cool at the top. The longer hydrocarbons have high boiling points so they condense at the bottom of the column (e.g. bitumen, heavy fuel oil and diesel oil). Shorter hydrocarbons have low boiling points so they condense at the top of the column (e.g. kerosene, petrol and liquified petroleum gas).**

5.7.1.3	Properties of hydrocarbons.

1. Name three properties of hydrocarbons which depend on the size of the molecule. **Boiling point, viscosity and flammability.**
2. What is viscosity? **A measure of how a substance flows (high viscosity = low flow).**
3. Describe how these properties change with increasing molecular size. **As chain length increases, boiling point increases, viscosity increases and flammability decreases.**
4. What is released by the combustion of fuels? **Energy.**
5. What two substances are oxidised during combustion? **Carbon and hydrogen.**
6. Write a word equation for the complete combustion of a hydrocarbon. **Hydrocarbon + oxygen → carbon dioxide + water.**
7. Write a balanced chemical equation for the complete combustion of methane. $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
8. Write a balanced chemical equation for the complete combustion of pentane. $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

5.7.1.4	Cracking and alkenes.

1. What is cracking? **The breakdown of hydrocarbons into smaller, more useful molecules.**
2. Name two methods of cracking. **Catalytic cracking and steam cracking.**
3. Describe the the two methods of cracking hydrocarbons. **Catalytic cracking is when long chain hydrocarbons are heated and vapourised. The gas is passed over a catalyst (e.g. aluminium oxide powder) where the long chains are broken down into smaller chains. Steam cracking is when long chain hydrocarbons are heated, vapourised and mixed with steam. The mixture is heated to a high temperature and the long chains are broken down into short chains.**
4. What are the products of cracking? **Alkanes and alkenes.**
5. Compare the reactivity of alkanes and alkenes. **Alkenes are more reactive than alkanes.**
6. Describe the test for alkenes. **Orange bromine water is added to the hydrocarbon. The mixture is shaken. If the hydrocarbon is an alkene the orange bromine water will become colourless. If the hydrocarbon is an alkane, the bromine water will stay orange.**
7. Why is cracking required? **Because there is a high demand for fuels with small molecules.**
8. What are alkenes used for? **They are used to produce polymers and are the starting materials for many chemicals.**
9. Write a balanced equation for hexane (C_6H_{14}) being cracked into butane. $C_6H_{14} \rightarrow C_4H_{10} + C_2H_4$.

5.8	Chemical analysis
5.8.1	Purity, formulations and chromatography.
5.8.1.1	Pure substances.

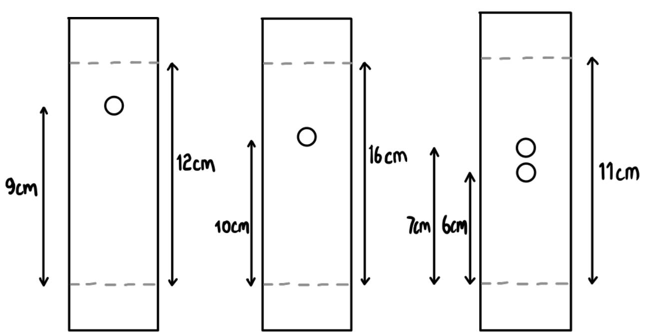
1. What is a pure substance? **A single element or compound that is not mixed with any other substance.**
2. How can pure substances be distinguished from mixtures? **Pure substances have specific melting and boiling points.**

5.8.1.2	Formulations.

1. What is a formulation? **A mixture which has been designed as a useful product.**
2. How are formulations made? **Formulations are made by mixing specific chemicals in measured quantities.**
3. Name seven examples of formulations. **Fuels, cleaning agents, paints, medicine, alloys, fertilisers and foods.**

5.7.1.3	Chromatography.

1. What is chromatography? **Chromatography is a method of separating mixtures and identifying substances.**
2. What is the mobile phase? **The solvent in which the stationary phase is placed.**
3. What is the stationary phase? **Usually a solid on which the samples are placed.**
4. Explain how samples are separated by chromatography. **Separation by chromatography depends on the distribution of substances between the mobile and stationary phases.**
5. What is the origin? **The line at the bottom of the paper where the samples are placed.**
6. What is the solvent front? **The line to denote the distance travelled by the mobile phase.**
7. What does Rf mean? **Relative front.**
8. What is the formula to calculate the Rf of a substance? **Rf = distance moved by substance / distance moved by solvent**
9. What is the Rf value of a sample which has travelled half the distance of the solvent front? **0.50**
10. What is the Rf value of a sample which has travelled a quarter of the distance of the solvent front? **0.25**
11. How can chromatography be used to determine if a substance is pure? **A pure compound will only produce a single spot on a chromatograms. A mixture will separate into two or more spots.**
12. Calculate the Rf value for each sample from the following chromatograms.



A. Rf = 9 / 12; Rf = 0.75	B. Rf = 10 / 16; Rf = 0.625	C. Rf = 6 / 11; Rf = 0.545 Rf = 7 / 11; Rf = 0.64

13. Which of the chromatograms above shows a mixture? C.

14. Describe an experiment to see if the colouring on sweets are pure or mixtures. Place different coloured sweets in a Petri dish. Add drops of water to each sweet to wash of the colour. Draw a line using a pencil about 2cm from the bottom of the paper. Carefully spot each dye onto the line. Place the paper into a beaker containing about 1cm depth of solvent. Leave the solvent to move up the paper for 20 mins or until it reaches near the top of the paper. Remove the paper and mark the level the solvent reached within a pencil line. Identify the number of spots for each sample. Calculate the Rf value for each of the spots. Pure samples will only have one spot.

5.8.2	Identification of common gases.
5.8.2.1	Test for hydrogen.

1. What is the test for hydrogen? Place a burning splint at the mouth of a test tube containing a gas. If hydrogen is present, a popping sound will be heard.

5.8.2.1	Test for oxygen.

1. What is the test for oxygen? Place a glowing splint at the mouth of a test tube containing a gas. If oxygen is present the splint will relight.

5.8.2.1	Test for carbon dioxide.

1. What is the test for carbon dioxide? Bubble a gas through an aqueous solution of calcium hydroxide (lime water). If carbon dioxide is present the lime water will turn cloudy.

5.8.2.1	Test for chlorine.

1. What is the test for chlorine? A piece of damp litmus paper is placed at the mouth of a test tube containing a gas. If the gas is chlorine the litmus paper will be bleached and turn white.

5.9	Chemistry of the atmosphere
5.9.1	The composition and evolution of the Earth's atmosphere.
5.9.1.1	The proportions of different gases in the atmosphere.

1. For approximately how long have the proportions of the gases in the atmosphere today been this way? **200 million years**
2. What is the proportion of nitrogen in the atmosphere? **Four-fifths (about 80%)**
3. What is the proportion of oxygen in the atmosphere? **One-fifth (about 20%)**
4. Name three other components that are found in small proportions in the atmosphere? **Carbon dioxide, water vapour and noble gases.**

5.9.1.2	The Earth's early atmosphere.

1. How old is the Earth thought to be? **4.6 billion years**
2. What was the atmosphere of Earth initially thought to be like? **Consisted mainly of carbon dioxide with little or no oxygen. It may have been similar to the atmospheres of Mars or Venus today.**
3. Describe how the atmosphere is thought to have developed in the first billion years. **Intense volcanic activity released gases and water vapour that formed the early atmosphere. The water vapour condensed to form the oceans.**
4. What caused the build up of nitrogen in the atmosphere? **Volcanic activity.**
5. As well as nitrogen which two other gases may have formed in small proportions? **Methane and ammonia.**
6. Describe how carbon dioxide concentrations in the atmosphere were reduced. **After the oceans were formed carbon dioxide dissolved in the water to form carbonates. These carbonates precipitated out of the water and were laid down as sediments (e.g. chalk). This reduced the levels of carbon dioxide in the atmosphere.**

5.9.1.3	How oxygen increased.

1. When was oxygen first produced? **About 2.7 billion years ago**
2. What was responsible for the production of oxygen? **Algae and plants.**
3. What process produced oxygen? **Photosynthesis.**
4. Write a word equation to show the production of oxygen. **carbon dioxide + water → glucose + oxygen**
5. Write a balanced equation to show the production of oxygen. $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$
6. Describe what happened to oxygen levels over the next billion years and the effect this had on life on Earth. **Plants evolved and the production of oxygen increased. This increase in oxygen enabled animals to evolve.**

5.9.1.3	How carbon dioxide decreased.

1. Name three ways by which carbon dioxide levels were decreased in the atmosphere. **Photosynthesis, formation of sedimentary rocks (e.g. chalk) and formation of fossil fuels (e.g. coal gas and oil) from dead animals and plant life.**
2.

5.9.2	Carbon dioxide and methane as greenhouse gases.
5.9.2.1	Greenhouse gases.

1. What is the effect of greenhouse gases on the Earth? Greenhouse gases maintain temperatures on Earth at a high enough level to support life.
2. Name three greenhouse gases? Water vapour, carbon dioxide and methane.
3. Describe the greenhouse effect. The sun emits short wave radiation which is not absorbed by greenhouse gases. The short wave radiation is absorbed by the Earth, warming it. This heat is emitted from the Earth as long wave radiation. Greenhouse gases absorb long wave radiation. The more greenhouse gases there are the more long wave radiation is absorbed causing a rise in temperature.

5.9.2.2	Human activities which contribute to an increase in greenhouse gases in the atmosphere.

1. Name two greenhouse gases which have increased as a result of human activities. Carbon dioxide and methane.
2. Name two human activities which have increased the amount of carbon dioxide in the atmosphere. Deforestation and burning fossil fuels.
3. Name two human activities which have increased the amount of methane in the atmosphere. Agriculture (cattle ranches, rice) and landfill sites.
4. Describe what the effect of human activities will be on the Earth's atmosphere? An increase in the temperature of the Earth's atmosphere which will result in global climate change.
5. How can scientists trust the data that has been collected about effect of greenhouse gases? The data published has been subjected to peer review (checked by other scientists) to ensure the data is not inaccurate or biased.
6. Explain why differing opinions are presented in the media about climate change? It is difficult to model a complex system such as the environment. In an effort to help people understand the issue, simplified models are often used. This can lead speculation and opinions presented by the media which is based only on parts of the evidence and may be biased or lack all the evidence.

5.9.2.3	Global climate change.

1. What is a major cause of global climate change? An increase in average global temperature.
2. List four potential effects of global climate change? Polar ice caps melting, changes in rainfall, changes in temperature, frequency and severity of storms, availability of water.
3. Discuss the environmental implications of global climate change? Melting of ice caps could lead to a rise in sea levels, flooding and coastal erosion. Changes in rainfall patterns could cause a change in the distribution of water and affect the ability to produce food. Severe storms will damage infrastructure and cause more people to become homeless and increase the spread of diseases such as cholera. The rise in temperature and availability of water may affect habitats and affect the distribution of wild species.

5.9.2.4	The carbon footprint and its reduction.

1. Define the term "carbon footprint". The total amount of carbon dioxide and other greenhouse gases emitted of the life cycle of a product service or event.
2. Describe how the carbon footprint can be reduced? Reduce the emissions of carbon dioxide and methane.
3. Describe actions which can be taken to reduce carbon dioxide and methane emissions. Use renewable energy or nuclear power. Use efficient processes to conserve energy and cut waste. Introduce a carbon tax to reward companies who reduce their carbon footprint. Put a cap on greenhouse emissions. Capture carbon dioxide from the atmosphere.
4. Suggest why the actions may have a limited effect. Countries do not want to sacrifice economic development. Individuals need to change habits and need to be educated how to do so.

5.9.3	Common atmospheric pollutants and their sources.
5.9.3.1	Atmospheric pollutants from fuels.

1. What is the effect of greenhouse gases on the Earth? **Greenhouse gases maintain temperatures on Earth at a high enough level to support life.**
2. Name three greenhouse gases? **Water vapour, carbon dioxide and methane.**
3. Describe the greenhouse effect. **The sun emits short wave radiation which is not absorbed by greenhouse gases. The short wave radiation is absorbed by the Earth, warming it. This heat is emitted from the Earth as long wave radiation. Greenhouse gases absorb long wave radiation. The more greenhouse gases there are the more long wave radiation is absorbed causing a rise in temperature.**

5.9.3.2	Properties and effects of atmospheric pollutants.

1. What are the properties of carbon monoxide? **A toxic colourless and odourless gas.**
2. What is the effect of carbon monoxide on the body? **Combines easily with haemoglobin in your blood and prevents oxygen from being taken up by the red blood cells.**
3. What is the effect of sulfur dioxide on the atmosphere? **Acid rain.**
4. What is the effect of oxides of nitrogen on the atmosphere? **Acid rain.**
5. What are the effects of particulates on the body? **Health problems such as aggravating asthma.**
6. What is the effect of particulates on the atmosphere? **Global dimming.**

5.10	Using resources
5.10.1	Using the Earth's resources and obtaining potable water.
5.10.1.1	Using the Earth's resources and sustainable development.

1. What do humans use the Earth's resources for? **To provide warmth, shelter, food and transport.**
2. What do natural resources, supplemented by agriculture provide? **Food, timber, clothing and fuels.**
3. What are processed finite resources from the earth, oceans and atmosphere used to provide? **Energy and materials.**
4. What is sustainable development? **Development that meets the needs of the current generations without compromising the ability of future generations to meet their own needs.**
5. Give three examples of natural products and their synthetic replacements. **Rubber (a natural product) has been replaced by man made polymers. Cotton (a natural product) has been replaced by synthetic fibres such as polyester or Lycra. Plant dyes (a natural product) have been replaced by synthetic dyes.**
6. What is a finite resource? **Resources which are not formed quickly enough to be replaceable.**
7. Give three examples of finite resources. **Fossil fuels, nuclear fuels and metal ores.**
8. What is a renewable resource? **A resources that forms at a faster or similar rate than they are being used.**
9. Give three examples of renewable resources. **Timber, fresh water and food.**

5.10.1.2	Potable water.

1. What is potable water? **Water that is safe to drink.**
2. What properties should drinking water have. **It should have only low levels of dissolved salts and microbes.**
3. Explain whether potable water is a pure substance or a mixture. **Potable water contains dissolved substances so it is not pure.**
4. What is the source of most of the potable water used in the UK. **Rain is the source of most water in the UK. It collects in the ground and in lakes and rivers.**
5. How is most of the potable water produced in the UK? **Water from a suitable source (e.g. reservoir or ground water) is passed through filter beds and then sterilised.**
6. Name three sterilising agents used to produce potable water. **Chlorine, ozone or ultraviolet light.**
7. What is the source of potable water, when fresh water supplies are limited? **Salt water or sea water.**
8. Name two methods of desalination. **Distillation or reverse osmosis.**
9. What is a drawback to using either of these methods? **These processes require large amounts of energy.**
10. Describe a method to analyse and purify water samples from different sources. **Take a sample of water and measure its pH. Filter the water using filter paper and a funnel to remove undissolved solid particles. Measure the mass of a round bottom flask. Place the water into a round bottom flask and attach to simple distillery. Collect the water that evaporates and condenses from the equipment. Retest the water for pH. Collect the round bottom flask and weigh it. Calculate the mass of dissolved solids remaining. Repeat the experiment with different samples of water.**

5.10.1.3	Waste water treatment.

1. What are the main sources of waste water? **Urban lifestyles, industrial processes and agriculture.**
2. Why does agricultural and sewage waste water require treatment? **These require treatment to remove organic matter and harmful microbes.**
3. Why does industrial waste water require treatment? **These require treatment for the removal of organic matter and harmful chemicals.**
4. Describe the four stages of sewage treatment. **Screening and grit removal; sedimentation to produce sludge and effluent; anaerobic digestion of sludge; aerobic biological treatment of effluent.**
5. Compare the relative ease of obtaining water from waste water, ground water and salt water. **Waste water requires expensive treatment to remove potentially toxic chemicals and microbes from the water. Ground water water just needs to be filtered and sterilised before it is potable. Saltwater can be purified but the processes require a lot of energy.**

5.10.1.4	Alternative methods of extracting metals (HT).

1. Why are alternative methods required for extracting metals from ores? **Copper ores are becoming scarce.**
2. Name two methods of extracting copper from low grade ores. **Phytomining and bioleaching.**
3. Describe how these new methods are different to traditional mining methods. **They do not involve digging, moving and disposing of large amounts of rocks.**
4. Describe the process of phytomining. **Phytomining uses plants to absorb metal compounds. The plants are harvested and then burned to form ash that contain metal compounds.**
5. Describe the process of bioleaching. **Bioleaching uses bacteria to produce leachate solutions that contain metal compounds.**
6. What is a leachate? **A liquid removed from the ground containing dissolved substances from the ground.**
7. How can the metal compounds extracted by alternative methods be processed to produce metal. **By displacement reactions with scrap iron or by electrolysis.**

5.10.2	Life cycle assessment and recycling.

5.10.2.1	Life cycle assessment.

1. What are life cycle assessments (LCAs). **These are assessments of the environmental impact of a product at each stage of its life.**
2. What are the four stages assessed by LCAs. **Extracting and processing raw materials; manufacturing and packaging of product; use and operation of product; disposal of product.**
3. What is easy to quantify when carrying out LCAs? **Use of water, resources, energy resources and the production of some waste.**
4. What is difficult to quantify when carrying out LCAs? **Pollutant effects.**
5. How can LCAs be misused. **Selective or abbreviated LCAs can be misused to reach pre-determined conclusions (e.g. for advertising campaigns).**
6. Carry out an LCA for shopping bags made from plastic and paper

LCA stage	Plastic bag	Paper bag
Raw materials	**Crude oil**	**Timber**
Manufacturing	**Distillation, cracking and polymerisation.**	**Pulp timber and process.**
Use	**Can be reused**	**Generally used once**
Disposal	**Not biodegradable. Can enter food chain.**	**Biodegradable and non toxic**

5.10.2.2 Ways of reducing the use of resources.

1. List three ways end users can sustainably use resources. **Reduction in use, reuse and recycling.**
2. Name four effects of reduction in the use of materials. **Reduce use of limited resources, reduce use of energy sources, reduction in waste and environmental impacts**
3. Name five materials which are produced from limited resources. **Metals, glass, building materials, clay ceramics and plastics.**
4. Name two methods of obtaining raw materials from the earth. **Quarrying and mining.**
5. What is the effect of obtaining materials by these methods? **Environmental impacts.**
6. Explain how glass products can be reused. **Glass bottles can be crushed and melted to make different glass products.**
7. What happens to other products which cannot be reused? **They are recycled for a different use.**
8. Describe how metals can be recycled. **By melting, recasting or reforming into different products.**
9. What does the amount of separation required for recycling depend upon? **Depends on the material and the properties required of the final product.**
10. How can scrap steel be recycled? **Scrap steel can be added to iron from a blast furnace to reduce the amount of iron that needs to be extracted from iron ore.**