# <u>**Year 10 Triple Science Chemistry Paper 1 Checklist**</u>

## Atomic Structure and the Periodic Table

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| • All substances are made of atoms. An atom is the smallest whole part of an element that can exist. |
| • Atoms of each element are represented by a chemical symbol, eg O represents an atom of oxygen, Na represents an atom of sodium. |
| • There are over 100 different elements. |
| • Elements are shown in the periodic table. |
| • An element is made up of atoms of one type, eg gold contains only gold atoms. |
| • Compounds are formed from elements by chemical reactions. |
| • Chemical reactions always involve the formation of one or more new substances. |
| • Compounds contain two or more difference elements chemically combined in fixed proportions. |
| • Compounds can be represented by formulae using the symbols of the atoms from which they were formed. |
| • Compounds can only be separated into elements by chemical reactions. |
| • Chemical reactions can be represented by word equations or equations using symbols and formulae. |
| • A molecule is made up of 2 or more atoms and can be an element or a compound. It is the smallest identifiable unit that retains the same properties. |
| • A mixture consists of two or more elements or compounds not chemically combined together. |
| • The chemical properties of each substance in the mixture are unchanged. |
| • A solvent is something that dissolves something else |
| • A solute is the substance that is dissolved in a solvent to make a solution. |
| • Mixtures can be separated by physical processes such as: filtration, crystallisation, simple distillation, fractional distillation, chromatography. |
| • These physical processes do not involve chemical reactions and no new substances are made. |
| • Filtration works by separating an insoluble solid from a liquid |
| • Distillation works by separating mixtures of liquids that have different boiling points |
| • Crystallisation separates a soluble salt from a solution. |
| • Chromatography separates mixtures based on their solubility |
| • Rf = distance travelled by sample/distance travelled by solvent |
| • New experimental evidence may lead to a scientific model being changed or replaced |
| • Before the discovery of the electron, atoms were thought to be tiny spheres that could not be divided |
| • The discovery of the electron led to the plum pudding model of the atom. The plum pudding model suggested that the atom is a ball of positive charge with negative electrons embedded in it |
| • Ernest Rutherford did an experiment called the gold foil experiment. He fired alpha particles at a thin sheet of gold foil. Some bounced back, some were deflected and some went straight through. |
| • The results from the alpha particle scattering experiment led to the conclusion that the mass of an atom was concentrated at the centre (nucleus) and that the nucleus was charged. This model was called the nuclear model and it replaced the plum pudding model. |
| • Niels Bohr adapted the nuclear model, suggesting that electrons orbit the nucleus at specific distances. |
| • Later experiments led to the idea that the positive charge of any nucleus could be subdivided into a whole number of smaller particles, with each particle having the same amount of positive charge. These particles were called protons |
| • James Chadwick discovered the existence of neutrons in the nucleus of the atom about 20 years after the nucleus was accepted |
| • The nucleus of an atom contains two types of subatomic particles; the proton and the neutron |
| • Protons have a mass of one and a charge of +1. |
| • Neutrons have a mass of 1 and a charge of 0. |
| • The third type of subatomic particle is called the electron. These are found around the outside of the nucleus. |
| • The electron has a mass of 0 and a charge of -1. |
| • In an atom, the number of electrons is equal to the number of protons in the nucleus. |
| • The number of protons in an atom is its atomic number. All atoms of the same element have the same number of protons. |
| • Atoms of different elements have different numbers of protons. |
| • Atoms are very small, having a radius of about 0.1nm ($1x10^{10}$m) |
| • The radius of a nucleus is less than 1/10000 of that of an atom (about $1x10^{14}$) |
| • The sum of the protons and neutrons in an atom is its mass number |
| • Atoms of the same element can have different numbers of neutrons. |
| • An atom with the same number of protons but different numbers of neutrons are called isotopes. |
| • The relative atomic mass of an element is an average value that takes into account the abundance of the isotopes of the element |
| • The electrons in an atom occupy the lowest available energy levels (the innermost available shells) |
| • The electronic structure can be represented by numbers or by a diagram |
| • The lowest energy level can hold a maximum of 2 electrons |
| • The second and third energy level can hold a maximum of 8 electrons<br>• For example, the electronic structure of sodium is 2, 8, 1 |
| • Before the discovery of subatomic particles, scientists attempted to classify the elements by arranging them in order of their atomic weights. |

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| • | The early periodic tables were incomplete, and some elements were placed in inappropriate groups if the strict order of atomic weights was followed. |
| • | Mendeleev overcame some of the problems by leaving gaps for elements that he thought had not been discovered and in some places changed the order based on atomic weights. |
| • | Elements with properties predicted by Mendeleev were discovered and filled the gaps. |
| • | Knowledge of isotopes made it possible to explain why the order based on atomic weights was not always correct |
| • | The modern periodic table has elements ordered by atomic number. |
| • | The group number shows the number of electrons in the outer shell. |
| • | The rows are called periods. Each period shows another full shell of electrons |
| • | Metals are found on the left of the periodic table and non-metals are found on the right |
| • | Metals are malleable, electronic conductors, thermal conductors |
| • | Non-metals are brittle, electronic insulators and thermal insulators |
| • | Elements that react to form positive ions are metals |
| • | Elements that react to form negative ions are non-metals |
| • | Elements in group 0 are called Noble gases. They are unreactive and do not easily form molecules. They have stable arrangements of electrons |
| • | The boiling points of the noble gases increase down the group |
| • | Group 1 metals are also known as alkali metals. They form an alkaline solution when reacting with water. |
| • | Group 1 metals form +1 ions and have low densities. |
| • | The reactivity of group 1 metals increases down the group. This is because it is easier to lose an electron when the outer shell is further from the nucleus. |
| • | The elements in group 7 are also known as the halogens. They have seven electrons in the outer shell. |
| • | Group 7 elements are non-metals and consist of molecules of pairs of atoms. |
| • | The reactivity of group 7 elements decreases as you go down the group. This is because it is more difficult to gain an electron when the outer shell is further from the nucleus. |
| • | A more reactive halogen can displace a less reactive halogen from an aqueous solution of its salt. |
| • | The transitions metals are found in the centre block of the periodic table. They form ions with different charges, form coloured compounds and are useful as catalysts |
| • | They have a higher density than group 1 metals and are less reactive |

## Bonding, Structure and the Properties of Matter

| **By the end of this topic you should know and understand:** | |
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| • | There are 3 types of strong bond: Ionic, Covalent and metallic |
| • | Ionic bonding is when oppositely charged ions are held together by electrostatic forces of attraction |
| • | Covalent bonding is when atoms share pairs of electrons |
| • | Metallic bonding is when atoms share delocalised electrons |
| • | Ionic bonding is the transfer of outer shell electrons between metals and non-metals |
| • | Metals lose electrons to become positively charged ions; Non-metals gain electrons to become negatively charged |
| • | Groups 1&2 and 6&7 transfer electrons to obtain a noble gas electronic structure |
| • | Dot-and-Cross diagrams are used to show the transfer of electrons to produce ions |
| • | Ionic compounds are giant structures of ions, held together by electrostatic forces of attraction |
| • | Electrostatic forces act in all directions to form a lattice structure |
| • | Limitations of the 3D models representing ionic compounds, including forces being represented as lines |
| • | Covalent bonding is between mostly non-metal atoms sharing electrons to form elements or compounds |
| • | Covalent substances can consist of small molecules, these can be deduced from their chemical formula; Some covalent substances have large molecules, such as polymers; Some covalent substances are giant structures such as diamond and silicon dioxide |
| • | These can be represented by dot-and-cross diagrams, ball and stick models and as repeating unit diagrams |
| • | Covalent substances can form single bonds, sharing one pair of electrons and represented by a single line on diagrams; can form double bonds, sharing two pairs of electrons and represented by double lines on diagrams |
| • | Metallic bonding occurs in metal elements and alloys |
| • | Metals consist of giant structures of atoms arranged in a regular pattern |
| • | The electrons in the outer shell of metal atoms are delocalised, so are free to move through the structure |
| - | The sharing of delocalised electrons gives rise to strong metallic bonds, the more electrons delocalised from the outer shell, the stronger the bond and the better the metal is at conducting electricity. This occurs due to the electrostatic attraction between the positive metal ions and the negative electrons |
| • | Metallic bonding can be represented by the diagram of a positive nucleus surrounded by a sea of delocalised electrons |
| • | The three states of matter are solid, liquid and gas and state symbols are used to represent these in symbol equations (s) (l) (g) respectively, with (aq) for aqueous solutions |
| • | Melting and freezing happen at the melting point of a substance; evaporation and condensation happen at the boiling point of a substance |
| • | States of matter can be represented by a simple particle model, representing the atoms as solid spheres |
| • | Particle theory can help explain changes in state |

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| • | The amount of energy needed to change state depends on the strength of forces between particles in the substance |
| • | The nature of the particles depends on the type of bonding and the structure of the substance |
| • | The stronger the forces between particles, the higher the melting and boiling points |
| • | **HT** Limitations of the model include: no forces shown between particles, all particles represented as solid spheres |
| • | Ionic compounds have a regular giant ionic lattice structure with strong bonds in all directions, leading to high melting and boiling points as large amounts of energy are needed to break them |
| • | Ionic compounds conduct electricity when melted (molten) or dissolved in water (aqueous), as ions are free to move and charge can flow e.g. Sodium Chloride |
| • | Covalent compounds that consist of small molecules are usually gases or liquids at room temperature as they have relatively low melting and boiling points due to weak intermolecular forces |
| • | Intermolecular forces are weak forces between molecules that are overcomes when substances melt or boil, NOT covalent bonds |
| • | Larger the molecule, the more intermolecular forces so larger molecules have higher melting and boiling points |
| • | Small molecules do not conduct electricity as there is no overall electric charge |
| • | Polymers are very large molecules, with the atoms linked by strong covalent bonds |
| • | HDPE (high density poly ethene) is made from straight chained polymers. LDPE (low density poly ethene) is made from branched chains |
| • | Thermosoftening polymers can be heated and remoulded, whereas thermosetting polymers cannot be remoulded. |
| • | Intermolecular forces between polymer molecules are relatively strong and so polymers are solid at room temperature |
| • | Giant covalent structures are solids with very high melting and boiling points |
| • | All the atoms are linked to other atoms by strong covalent bonds and these require a lot of energy to overcome them |
| • | Examples of giant covalent structures are diamond and graphite and silicon dioxide |
| • | Metals have giant structures of atoms with strong metallic bonding |
| • | Metals are good conductors of electricity because of the delocalised electrons carrying the electrical charge |
| • | Metals are good conductors of thermal energy because energy is transferred by the delocalised electrons |
| • | Most metals have high melting and boiling points |
| • | Atoms are arranged in layers in pure metals, allowing them to be bent and shaped (malleable) |
| • | Pure metals are too soft for many uses so are mixed with other metals to form alloys which are harder |
| • | Diamond is made up of carbon atoms and each carbon atom forms 4 covalent bonds with other carbon atoms in a giant covalent structure and so is very hard with a very high melting point and does not conduct electricity |
| • | Graphite is made up of carbon atoms and each carbon atom forms 3 covalent bonds with 3 other carbon atoms, forming layers of hexangonal rings with no covalent bonds between layers |
| • | The 4th electron of each carbon becomes delocalised, allowing graphite to conduct electricity |
| • | Graphene is a single layer of graphite and has properties that make it useful in electronics and composites. |
| • | Fullerenes are molecules of carbon atoms with hollow shapes. The structure of fullerenes is based on hexagonal rings of carbon atoms but they may also contain rings with five or seven carbon atoms. |
| • | The first fullerene to be discovered was Buckminsterfullerene ($C_{60}$) which has a spherical shape |
| • | Carbon nanotubes are cylindrical fullerenes with very high length to diameter ratios. Their properties make them useful for nanotechnology, electronics and materials. |
| • | Nanoscience refers to structures between 1-100nm in size (a few hundred atoms) |
| • | Nanoparticles are smaller than fine particles which are between 100-2500nm (1x10-7m and 2.5x10-6m) and coarse particles (dust) which are between 2500nm and 10000nm (2.5x10-6m and 1x10-5m) |
| • | As the side of cube decreases by a factor of 10, the surface area to volume ratio increases by a foctor of 10 |
| • | Nanoparticles may have different properties to a larger amount of the same material due to the higher surface to volume ratio |
| • | Smaller quantities may be needed for a material to be effective due to the different properties of materials with normal particle sizes |
| • | Nanoparticles have applications medicine, electronics, cosmetics and sun creams, deodorants and as catalysts |
| • | New applications of nanoparticulate materials are an important area of research |

## Quantitative Chemistry

| **By the end of this topic you should know and understand:** | |
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| • | The law of conservation of mass states that no atoms are lost or made during a chemical reaction, so the mass of the products equals the mass of the reactants. |
| • | Chemical reactions can be represented by symbol equations which are balanced in terms of the numbers of atom of each element involved on both sides of the equation |
| • | Some reactions may appear to involve a change in mass but this can usually be explained because a reactant or product is a gas and its mass has not been taken into account. |
| • | For example, when a metal reacts with oxygen the mass of the oxide produced is greater than the mass of the metal. |
| • | In thermal decomposition of metal carbonate, carbon dioxide is produced and escapes into the atmosphere leaving the metal oxide as the only solid product. |
| • | The relative formula mass ($M_r$) of a compound is the sum of the relative atomic masses of the atoms in the numbers shown in the formula |

- In a balanced chemical equation, the sum of the relative formula masses of the reactants in the quantities shown equals the sum of the relative formula masses of the products shown.
- Percent by mass tells us how much of the mass of a compound is made by one of the elements. It is calculated by: (Mr of element/Mr of compound) x 100
- Whenever a measurement is made there is always some uncertainty about the result obtained
- A systematic error causes readings to differ from the true value by the same amount each time. This could be caused by environmental conditions, methods of observation or the equipment used.
- A random error causes readings to be spread about the true value, with results varying in an unpredictable way. The effect of random errors can be reduced by taking more measurements and calculating a mean.
- A zero error is caused when a measuring device does not read zero when nothing is measured. It could result in systematic uncertainty.
- Chemical amounts are measured in moles. The symbol for the unit mole is mol.
- The mass of one mole of a substance in grams is numerically equal to its relative formula mass
- One mole of a substance contains the same number of the stated particles, atoms, molecules or ions as one mole of any other substance.
- The number of atoms, molecules or ions in a mole of a given substance is the Avogadro constant. This is $6.02 \times 10^{23}$ per mole.
- The masses of reactants and products can be calculated from balanced symbol equations.
- Chemical equations can be interpreted in terms of moles. For example: $Mg + 2HCl \rightarrow 2MgCl_2 + H_2$ shows that one mole of magnesium reacts with two moles of hydrochloric acid to produce one mole of magnesium chloride and one mole of hydrogen gas.
- The balancing numbers in a symbol equation can be calculated from the masses of reactants and products by converting the masses in grams to amounts in moles and converting the numbers of moles to simple whole number ratios
- In a chemical reaction involving two reactants, it is common to use an excess of one of the reactants to ensure all of the other reactant is used. The reactant that is completely used up is called the limiting reactant because it limits the amount of products
- Many chemical reactions take place in solutions.
- The concentration of a solution can be measured in mass per given volume of solution eg. grams per $dm^3$ ($g/dm^3$)

- *Even though no atoms are gained or lost in a chemical reaction, it is not always possible to obtain the calculated amount of product because:*
  - *A) The reaction may not go to completion because it is reversible*
  - *B) Some of the product may be lost when it is separated from the reaction mixture*
  - *C) Some reactants may react in ways different to the expected reaction.*

- *The amount of product obtained is known as the yield. When compared with the maximum theoretical amount as a percentage, it is called the percentage yield.*
- *% yield = (mass of product actually made/maximum theoretical mass of product) x 100*
- *The atom economy (atom utilisation) is a measure of the amount of starting materials that end up as useful products.*
- *It is important for sustainable development and for economic reasons to use reactions with high atom economy*
- *The percentage atom economy of a reaction is calculated using the balanced equation for the reaction and the following equation: (relative formula mass of desired product from equation/sum of relative formula masses of all reactants from equation) x 100*
- *The concentration of a solution can be measured in* $mol/dm^3$
- *The amount in moles of solute or the mass in grams of solute in a given volume of solution can be calculated from its concentration in* $mol/dm^3$
- *If the volumes of two solutions that react completely are known and the concentration of one solution is known, the concentration of the other solution can be calculated*
- *Equal amounts in moles of gases occupy the same volume under the same conditions of temperature and pressure*
- *The volume of one mole of any gas at room temperature and pressure (20°C and 1 atmosphere pressure) is* $24dm^3$
- *The volumes of gaseous reactants and products can be calculated from the balanced equation for the reaction*

## Chemical Changes

**By the end of this topic you should know and understand:**

- Acids produce hydrogen ions (H+) in aqueous solutions.
- Aqueous solutions of alkalis contain hydroxide ions (OH–).
- The pH scale, from 0 to 14, is a measure of the acidity or alkalinity of a solution, and can be measured using universal indicator or a pH probe (wide range indicators)
- A solution with pH 7 is neutral. Aqueous solutions of acids have pH values of less than 7 and aqueous solutions of alkalis have pH values greater than 7.
- In neutralisation reactions between an acid and an alkali, hydrogen ions react with hydroxide ions to produce water. This reaction can be represented by the equation: $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$
- Acids are neutralised by alkalis (soluble metal hydroxides) and bases (insoluble metal hydroxides and metal oxides) to produce salts and water, and by metal carbonates to produce salts, water and carbon dioxide.
- The particular salt produced in any reaction between an acid and a base or alkali depends on: the acid used (hydrochloric acid produces chlorides, nitric acid produces nitrates, sulfuric acid produces sulfates) and the positive ions in the base, alkali or carbonate.
- Formulae of common ions can predict the formulae of salts
- A strong acid is completely ionised in aqueous solution. Examples of strong acids are hydrochloric, nitric and sulfuric acids.
- A weak acid is only partially ionised in aqueous solution. Examples of weak acids are ethanoic, citric and carbonic acids.

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| • | For a given concentration of aqueous solutions, the stronger an acid, the lower the pH. |
| • | As the pH decreases by one unit, the hydrogen ion concentration of the solution increases by a factor of 10. |
| • | Concentration relates to the amount of substance in a given volume. High concentration has a large amount of substance and low solution has a small amount of substance in  given volume. |
| • | Solutions can be diluted to decrease concentration. |
| • | As acids react, the hydrogen ions react to form water molecules reducing the concentration of the acid and increasing neutrality. |
| • | The volumes of acid and alkali solutions that react with each other can be measured by titration using a suitable indicator. |
| • | Metals react with oxygen to produce metal oxides. The reactions are oxidation reactions because the metals gain oxygen. |
| • | Acids react with some metals to produce salts and hydrogen. |
| • | Metals react with hydrochloric acid to form chloride salts. |
| • | Metals react with sulphuric acid to form sulphate salts. |
| • | Soluble salts can be made from acids by reacting them with solid insoluble substances, such as metals, metal oxides, hydroxides or carbonates. |
| • | The solid is added to the acid until no more reacts and the excess solid is filtered off to produce a solution of the salt. |
| • | Salt solutions can be crystallised to produce solid salts. |
| • | When metals react with other substances the metal atoms form positive ions. |
| • | The reactivity of a metal is related to its tendency to form positive ions. |
| • | Metals can be arranged in order of their reactivity in a reactivity series. |
| • | The metals potassium, sodium, lithium, calcium, magnesium, zinc, iron and copper can be put in order of their reactivity from their reactions with water and dilute acids, at room temperature. |
| • | The non-metals hydrogen and carbon are often included in the reactivity series. |
| • | A more reactive metal can displace a less reactive metal from a compound. |
| • | Unreactive metals such as gold are found in the Earth as the metal itself but most metals are found as compounds that require chemical reactions to extract the metal. |
| • | Metals less reactive than carbon can be extracted from their oxides by reduction with carbon. Reduction involves the loss of oxygen. |
| • | Oxidation is the loss of electrons and reduction is the gain of electrons. |
| • | Half equations are used to show which species has been oxidised and which species has been reduced. |
| • | Ionic equations are used to show the reacting species in a reaction. |
| • | Metals lose electrons and acids gain electrons as redox reactions (HT) Metals are oxidised and acids are reduced (HT) |
| • | When an ionic compound is melted or dissolved in water, the ions are free to move about within the liquid or solution. |
| • | These liquids and solutions are able to conduct electricity and are called electrolytes. |
| • | Passing an electric current through electrolytes causes the ions to move to the electrodes. |
| • | Positively charged ions move to the negative electrode (the cathode), and negatively charged ions move to the positive electrode (the anode). Ions are discharged at the electrodes producing elements. This process is called electrolysis. |
| • | Half equations are used to show the reactions occurring at the electrodes during electrolysis. |
| • | When a simple ionic compound (eg lead bromide) is electrolysed in the molten state using inert electrodes, the metal (lead) is produced at the cathode and the non-metal (bromine) is produced at the anode. |
| • | Metals can be extracted from molten compounds using electrolysis |
| • | Electrolysis is used if the metal is too reactive to be extracted by reduction with carbon or if the metal reacts with carbon. Large amounts of energy are used in the extraction process to melt the compounds and to produce the electrical current. |
| • | Aluminium is manufactured by the electrolysis of a molten mixture of aluminium oxide and cryolite to lower the melting point, using carbon as the positive electrode (anode). Carbon dioxide is produced as the oxygen reacts with the positive electrode and so the carbon electrode must be continually replaced. |
| • | The ions discharged when an aqueous solution is electrolysed using inert electrodes depend on the relative reactivity of the elements involved. At the negative electrode (cathode), hydrogen is produced if the metal is more reactive than hydrogen. At the positive electrode (anode), oxygen is produced unless the solution contains halide ions, when the halogen is produced. |
| • | This happens because in the aqueous solution water molecules break down producing hydrogen ions and hydroxide ions that are discharged. |
| • | During electrolysis, at the cathode (negative electrode), positively charged ions gain electrons and so the reactions are reductions. At the anode (positive electrode), negatively charged ions lose electrons and so the reactions are oxidations. |
| • | Reactions at electrodes can be represented by half equations, for example: $2H^+ + 2e^- \rightarrow H_2$<br>and $4OH^- \rightarrow O_2 + 2H_2O + 4e^-$ or $4OH^- - 4e^- \rightarrow O_2 + 2H_2O$ |

## Energy Changes

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| • | Energy is conserved in chemical reactions. |
| • | The amount of energy in the universe at the end of a chemical reaction is the same as before the reaction takes place. |
| • | An exothermic reaction is one that transfers energy to the surroundings so the temperature of the surroundings increases. |
| • | Everyday uses of exothermic reactions including self-heating cans and hand warmers. |
| • | An endothermic reaction is one that takes heat in from the surroundings so the temperature of the surroundings decreases. |
| • | Endothermic reactions include thermal decompositions and the reaction of citric acid and sodium hydrogencarbonate. Some sports injury packs also use endothermic reactions. |

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| • | Chemical reactions can occur only when reacting particles collide with each other and with sufficient energy. |
| • | Activation energy is the minimum amount of energy that particles must have to react. |
| • | Reaction profiles can be used to show the relative energies of reactants and products, the activation energy and the overall energy change of a reaction. |
| • | During a chemical reaction energy must be supplied to break bonds in the reactants. |
| • | During a chemical reaction energy is released when bonds in the product are formed. |
| • | The energy needed to break bonds and the energy released when bonds are formed can be calculated from bond energies. |
| • | The difference between the sum of the energy needed to break bonds in the reactants and the sum of the energy released when bonds in the products are formed is the overall energy change of the reaction. |
| • | In an exothermic reaction, the energy released from forming new bonds is greater than the energy needed to break existing bonds. |
| • | In an endothermic reaction, the energy needed to break existing bonds is greater than the energy released from forming new bonds. |
| • | Cells contain chemicals which react to produce electricity. |
| • | The voltage produced by a cell is dependent upon a number of factors including the type of electrode and electrolyte. |
| • | A simple cell can be made by connecting two different metals in contact with an electrolyte. |
| • | The biggest voltage occurs when the difference in the reactivity of the two metals is the largest. |
| • | Batteries consist of two or more cells connected together in series to provide a greater voltage. |
| • | In non-rechargeable cells and batteries the chemical reactions stop when one of the reactants has been used up. Alkaline batteries are non-rechargeable. |
| • | Rechargeable cells and batteries can be recharged because the chemical reactions are reversed when an external electrical current is supplied. |
| • | Fuel cells are supplied by an external source of fuel (eg hydrogen) and oxygen or air. The fuel is oxidised electrochemically within the fuel cell to produce a potential difference. |
| • | The overall reaction in a hydrogen fuel cell involves the oxidation of hydrogen to produce water. |
| • | Hydrogen fuel cells offer a potential alternative to rechargeable cells and batteries. |