C1H

AQA GCSE Combined Science (9-1)

Topic 1: Atomic Structure and the Periodic Table

4.1.1 Atoms, elements and compounds			
4.1.1.1 Atoms, elements and compounds	☺	😐	☹
All substances are made of **atoms**. An atom is the smallest 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 about **100 different elements**. Elements are shown in the **periodic table**.			
Compounds are formed from elements by **chemical reactions**. Chemical reactions always involve the formation of one or more **new substances**, and often involve a detectable **energy change**.			
Compounds contain two or more elements **chemically combined** in fixed proportions and 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**.			
Students will be supplied with a periodic table for the exam and should be able to: ★ Use the names and symbols of the **first 20 elements** in the periodic table, the elements in **Groups 1 and 7**, and other elements in this specification.			
★ Name **compounds** of these elements from given formulae or symbol equations.			
★ Write **word equations** for the reactions in this specification.			
★ Write **formulae** and **balanced chemical equations** for the reactions in this specification.			
★ (HT only) Write balanced **half equations** and **ionic equations** where appropriate.			
4.1.1.2 Mixtures	☺	😐	☹
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.			
Mixtures can be **separated** by **physical processes** such as **filtration**, **crystallisation**, **simple distillation**, **fractional distillation** and **chromatography**. These physical processes do not involve chemical reactions.			
Students should be able to: ★ **Describe**, explain and give examples of the specified processes of separation.			
★ **Suggest** suitable separation and purification techniques for mixtures when given appropriate information.			
4.1.1.3 Scientific models of the atom (*common content with physics*)	☺	😐	☹
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 was a ball of positive charge with negative electrons embedded in it.			
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 **nuclear model** replaced the plum pudding model.			
Niels Bohr adapted the nuclear model by suggesting that **electrons orbit** the nucleus at specific distances. The theoretical calculations of Bohr agreed with experimental observations.			

Later experiments led to the idea that the **positive charge** of any nucleus could be subdivided into a whole number of smaller particles, each particle having the same amount of positive charge. The name **proton** was given to these particles.

The experimental work of **James Chadwick** provided the evidence to show the existence of **neutrons** within the nucleus. This was about 20 years after the nucleus became an accepted scientific idea.

<u>Students should be able to:</u>

★ **Describe** the **difference** between the **plum pudding model** of the atom and the **nuclear model** of the atom.

★ **Describe** why the **new evidence** from the scattering experiment led to a **change** in the atomic model.

Details of experimental work supporting the Bohr model are not required. Details of these experiments are not required. Details of Chadwick's experimental work are not required.

4.1.1.4 Relative electrical charges of subatomic particles ☺ 😐 ☹

The **relative electrical charges** of the particles in atoms are:

Name of particle	Relative charge
Proton	+1
Neutron	0
Electron	–1

In an atom, the number of electrons is equal to the number of protons in the nucleus. Atoms have no overall electrical charge (they are neutral).

The number of **protons** in an atom of an element is its **atomic number**. All atoms of a particular element have the same number of protons. Atoms of different elements have different numbers of protons.

<u>Students should be able to:</u>

★ Use the atomic model to describe atoms.

4.1.1.5 Size and mass of atoms ☺ 😐 ☹

Atoms are **very small**, having a radius of **about 0.1 nm** (1×10^{-10} m).

The radius of a nucleus is less than 1/10 000 of that of the atom (about 1×10^{-14} m).

Almost all the **mass** of an atom is the **nucleus**.

The **relative masses** of protons, neutrons and electrons are:

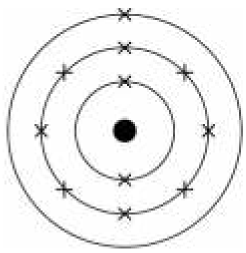
Name of particle	Relative mass
Proton	1
Neutron	1
Electron	Very small

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; these atoms are called **isotopes** of that element.

Atoms can be represented as shown in this example:

(Mass number) 23
(Atomic number) 11 $^{23}_{11}\mathrm{Na}$

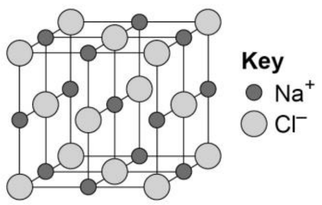
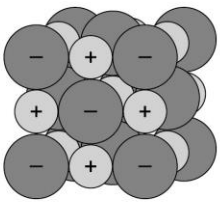
<u>Students should be able to:</u> ★ **Calculate** the numbers of protons, neutrons and electrons in an atom or ion, given its atomic number and mass number.			
★ Relate size and scale of atoms to objects in the physical world. [MS 1d]			
WS 4.3 Use SI units and the prefix **nano**.			
MS 1b Recognise expressions in **standard form**.			
4.1.1.6 Relative atomic mass	☺	😐	☹
The relative atomic mass of an element is an average value that takes account of the abundance of the isotopes of the element.			
<u>Students should be able to:</u> ★ **Calculate** the relative atomic mass of an element given the percentage abundance of its isotopes.			
4.1.1.7 Electronic structure	☺	😐	☹
The electrons in an atom occupy the lowest available **energy levels** (innermost available **shells**). The electronic structure of an atom can be represented by numbers or by a diagram. For example, the electronic structure of **sodium** is **2,8,1** or  showing two electrons in the lowest energy level, eight in the second energy level and one in the third energy level. *Students may answer questions in terms of either energy levels or shells.*			
<u>Students should be able to:</u> ★ Represent the electronic structures of the **first twenty elements** of the periodic table in both forms.			
4.1.2 The periodic table			
4.1.2.1 The periodic table	☺	😐	☹
The elements in the **periodic table** are arranged in order of atomic (proton) number and so that elements with similar properties are in columns, known as **groups**. The table is called a periodic table because **similar properties** occur at regular intervals.			
Elements in the **same group** in the periodic table have the same number of electrons in their outer shell (outer electrons) and this gives them **similar chemical properties**.			
<u>Students should be able to:</u> ★ **Explain** how the position of an element in the periodic table is related to the arrangement of electrons in its atoms and hence to its atomic number.			
★ **Predict** possible reactions and probable reactivity of elements from their positions in the periodic table.			
4.1.2.2 Development of the periodic table	☺	😐	☹
Before the discovery of protons, neutrons and electrons, scientists attempted to classify the elements by arranging them in order of their **atomic weights**.			
The early periodic tables were incomplete and some elements were placed in inappropriate groups if the strict order of atomic weights was followed.			

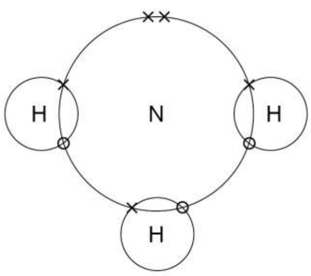
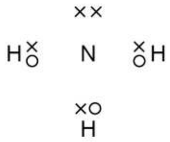
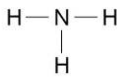
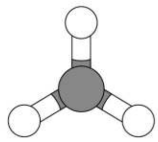
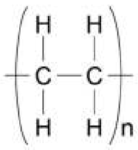
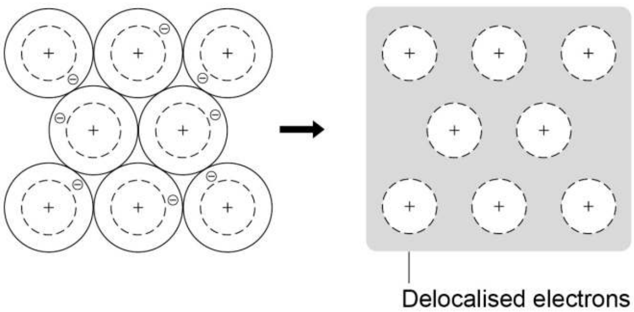
Mendeleev overcame some of the problems by **leaving gaps** for elements that he thought had not been discovered and in some places changed the order based on atomic weights.			
Elements with properties **predicted** by Mendeleev were discovered and filled the gaps. Knowledge of isotopes made it possible to explain why the order based on atomic weights was not always correct.			
Students should be able to: ★ **Describe** these steps in the development of the periodic table.			
WS 1.1+1.6 Explain how testing a prediction can support or refute a new scientific idea.			
4.1.2.3 Metals and non-metals	☺	😐	☹
Elements that react to form **positive ions** are **metals**.			
Elements that do not form positive ions are non-metals.			
The majority of elements are metals. **Metals** are found to the **left** and towards the **bottom** of the periodic table. Non-metals are found towards the right and top of the periodic table.			
Students should be able to: ★ **Explain** the differences between **metals** and **non-metals** on the basis of their characteristic **physical** and **chemical properties**. *Links with 'Group 0', 'Group 1', 'Group 7' and 'Bonding, structure and the properties of matter'.*			
★ **Explain** how the **atomic structure** of metals and non-metals relates to their **position** in the periodic table.			
★ **Explain** how the **reactions of elements** are related to the **arrangement of electrons** in their atoms and hence to their **atomic number**.			
4.1.2.4 Group 0	☺	😐	☹
The elements in **Group 0** of the periodic table are called the **noble gases**. They are **unreactive** and do not easily form molecules because their atoms have **stable arrangements** of electrons. The noble gases have **eight electrons** in their outer energy level, except for helium, which has only two electrons.			
The **boiling points** of the noble gases increase with increasing relative atomic mass (going down the group).			
Students should be able to: ★ **Explain** how properties of the elements in Group 0 depend on the outer shell of electrons of the atoms.			
★ **Predict** properties from given trends down the group.			
4.1.2.5 Group 1	☺	😐	☹
The elements in **Group 1** of the periodic table, known as the **alkali metals** and have characteristic properties because of the **single electron** in their outer shell. E.g. • are metals with **low density** (the first three elements in the group are less dense than water). • react with non-metals to form **ionic compounds** in which the metal ion carries a charge of **+1**. The compounds are white solids that dissolve in water to form colourless solutions. • react with water, releasing **hydrogen**. • form **hydroxides** that dissolve in water to give **alkaline solutions**.			
In Group 1, the **reactivity** of the elements increases going down the group.			
Students should be able to: ★ **Describe** the reactions of the first three alkali metals with oxygen, chlorine and water.			
★ **Explain** how properties of the elements in Group 1 depend on the outer shell of electrons of the atoms. [WS 1.2]			
★ **Predict** properties from given trends down the group. [WS 1.2]			

4.1.2.6 Group 7	☺	😐	☹
The elements in **Group 7** of the periodic table, known as the **halogens** and have similar reactions because they all have **seven electrons** in their outer shell. The halogens are **non-metals** and consist of molecules made of **pairs of atoms**.			
Halogens: • react with **metals** to form **ionic compounds** in which the halide ion carries a charge of **–1** • form **molecular compounds** with other **non-metallic** elements.			
In Group 7, the further down the group an element is the higher its relative molecular mass, **melting point** and **boiling point**.			
In Group 7, the **reactivity** of the elements decreases going down the group.			
A more reactive halogen can **displace** a less reactive halogen from an aqueous solution of its salt.			
Students should be able to: ★ **Describe** the nature of the compounds formed when chlorine, bromine and iodine react with metals and non-metals.			
★ **Explain** how properties of the elements in Group 7 depend on the outer shell of electrons of the atoms.			
★ **Predict** properties from given trends down the group.			
4.1.3 Properties of transition metals (Chemistry only)			
4.1.3.1 Comparison with Group 1 elements	☺	😐	☹
The **transition elements** are **metals** with **similar properties** which are different from those of the elements in Group 1.			
Students should be able to: ★ **Describe** the difference compared with Group 1 in melting points, densities, strength, hardness and reactivity with oxygen, water and halogens.			
★ Exemplify these general properties by reference to Cr, Mn, Fe, Co, Ni, Cu.			
4.1.3.2 Typical properties	☺	😐	☹
Many transition elements have ions with different charges, form **coloured compounds** and are useful as **catalysts**.			
Students should be able to: ★ Exemplify these general properties by reference to compounds of Cr, Mn, Fe, Co, Ni, Cu.			

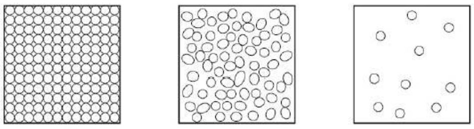
Topic 2: Bonding, structure, and the properties of matter

4.2.1 Chemical bonds, ionic, covalent and metallic			
4.2.1.1 Chemical bonds	☺	😐	☹
There are three types of strong chemical bonds: ionic, covalent and metallic. For **ionic bonding** the particles are oppositely **charged ions**. For **covalent bonding** the particles are atoms which **share pairs of electrons**. For **metallic bonding** the particles are atoms which share **delocalised electrons**.			
Ionic bonding occurs in compounds formed from **metals** combined with **non-metals**.			
Covalent bonding occurs in non-metallic elements and in compounds of **non-metals**.			
Metallic bonding occurs in **metallic elements** and alloys.			

<u>Students should be able to:</u> ★ **Explain** chemical bonding in terms of **electrostatic forces** and the **transfer or sharing of electrons**.			
4.2.1.2 Ionic bonding	☺	😐	☹
When a metal atom reacts with a non-metal atom, **electrons** in the outer shell of the metal atom are **transferred**. **Metal atoms lose electrons** to become **positively charged ions**. **Non-metal atoms gain electrons** to become **negatively charged ions**. The ions produced by metals in Groups 1 and 2 and by non-metals in Groups 6 and 7 have the electronic structure of a **noble gas** (Group 0).			
The electron transfer during the formation of an ionic compound can be represented by a **dot and cross** diagram e.g. for sodium chloride: 			
The **charge** on the ions produced by metals in Groups 1 and 2 and by non-metals in Groups 6 and 7 relates to the **group number** of the element in the periodic table.			
<u>Students should be able to:</u> ★ Draw **dot and cross diagrams** for ionic compounds formed by metals in Groups 1 and 2 with non-metals in Groups 6 and 7.			
★ Work out the **charge** on the ions of metals and non-metals from the group number of the element, limited to the metals in Groups 1 and 2, and non-metals in Groups 6 and 7.			
4.2.1.3 Ionic compounds	☺	😐	☹
An ionic compound is a **giant structure** of ions. Ionic compounds are held together by **strong electrostatic forces** of attraction between **oppositely charged ions**. These forces act in all directions in the lattice and this is called **ionic bonding**.			
The structure of **sodium chloride** can be represented in the following forms:  *Students should be familiar with the structure of **sodium chloride** but do not need to know the structures of other ionic compounds.*			
<u>Students should be able to:</u> ★ **Deduce** that a **compound is ionic** from a diagram of its structure in one of the specified forms			
★ **Describe** the **limitations** of using dot and cross, ball and stick, two and three dimensional diagrams to represent a **giant ionic structure**			
★ Work out the **empirical formula** of an ionic compound from a given model or diagram that shows the ions in the structure.			
4.2.1.4 Covalent bonding	☺	😐	☹
When atoms **share** pairs of **electrons**, they form **covalent bonds**. These bonds between atoms are **strong**.			
Covalently bonded substances may consist of **small molecules** (such as H_2, Cl_2, O_2, N_2, HCl, H_2O, NH_3 and CH_4).			

Some covalently bonded substances have very **large molecules**, such as **polymers**.			
Some covalently bonded substances have **giant covalent structures**, such as **diamond** and **silicon dioxide**.			
The covalent bonds in **molecules** and **giant structures** can be represented in the following forms: For ammonia (NH_3)  and/or  and/or  and/or 			
Polymers can be represented in the form:  poly(ethene) …where **n** is a large number.			
Students should be able to: ★ Draw **dot and cross** diagrams for the molecules of hydrogen, chlorine, oxygen, nitrogen, hydrogen chloride, water, ammonia and methane.			
★ Represent the covalent bonds in small molecules, in the repeating units of polymers and in part of giant covalent structures, using **a line to represent a single bond**.			
★ **Describe** the **limitations** of using dot and cross, ball and stick, two and three dimensional diagrams to represent molecules or giant structures.			
★ **Deduce** the **molecular formula** of a substance from a given model or diagram in these forms showing the atoms and bonds in the molecule.			
★ Recognise substances as small molecules, polymers or giant structures from diagrams showing their bonding.			
4.2.1.5 Metallic bonding	☺	😐	☹
Metals consist of **giant structures** of atoms arranged in a **regular pattern**.			
The electrons in the outer shell of metal atoms are **delocalised** and so are **free to move through the whole structure**. The sharing of delocalised electrons gives rise to **strong metallic bonds**.			
The bonding in metals may be represented in the following form:  			
WS 1.2 Recognise substances as metallic giant structures from diagrams showing their bonding.			

4.2.2 How bonding and structure are related to the properties of substances

4.2.2.1 The three states of matter	☺	😐	☹
The three states of matter are **solid**, **liquid** and **gas**. **Melting** and **freezing** take place at the **melting point**, **boiling** and **condensing** take place at the **boiling point**.			
The three states of matter can be represented by a **simple model**. In this model, particles are represented by small solid spheres. **Particle theory** can help to explain melting, boiling, freezing and condensing.  Solid Liquid Gas			
The amount of **energy needed** to change state from solid to liquid and from liquid to gas depends on the **strength of the forces** between the particles of the substance. The nature of the particles involved depends on the type of bonding and the structure of the substance. The stronger the forces between the particles the higher the **melting point** and **boiling point** of the substance.			
(HT only) **Limitations** of the simple model include that there are **no forces** between the spheres, that all particles are **represented as spheres** and that the **spheres are solid**.			
<u>Students should be able to:</u> ★ **Predict the states** of substances at different temperatures given appropriate data			
★ **Explain** the different temperatures at which changes of state occur in terms of **energy transfers** and **types of bonding**			
★ Recognise that atoms themselves do not have the bulk properties of materials			
★ (HT only) Explain the **limitations** of the particle theory in relation to changes of state when particles are represented by solid inelastic spheres which have no forces between them.			
4.2.2.2 State symbols	☺	😐	☹
In chemical equations, the three states of matter are shown as (s), (l) and (g), with (**aq**) for **aqueous solutions**.			
<u>Students should be able to:</u> ★ Include appropriate state symbols in chemical equations for the reactions in this specification.			
4.2.2.3 Properties of ionic compounds	☺	😐	☹
Ionic compounds have **regular structures** (**giant ionic lattices**) in which there are **strong electrostatic forces** of attraction in all directions between oppositely charged ions. *Knowledge of the structures of specific ionic compounds other than **sodium chloride** is not required*			
These compounds have **high melting points** and **high boiling points** because of the large amounts of **energy** needed to break the many **strong bonds**.			
When melted or dissolved in water, ionic compounds **conduct electricity** because the **ions are free to move** and so charge can flow.			
4.2.2.4 Properties of small molecules	☺	😐	☹
Substances that consist of **small molecules** are usually **gases** or **liquids** that have relatively **low melting points** and **boiling points**.			
These substances have only **weak forces** between the molecules (**intermolecular forces**). It is these intermolecular forces that are overcome, not the covalent bonds, when the substance melts or boils.			

The intermolecular forces increase with the size of the molecules, so **larger molecules** have **higher melting** and **boiling points**.			
These substances do not conduct electricity because the molecules do not have an overall electric charge.			
<u>Students should be able to:</u> ★ Use the idea that intermolecular forces are weak compared with covalent bonds to explain the **bulk properties** of molecular substances.			
4.2.2.5 Polymers	☺	😐	☹
Polymers have very large molecules. The atoms in the polymer molecules are linked to other atoms by **strong covalent bonds**.			
The **intermolecular forces** between polymer molecules are **relatively strong** and so these substances are **solids at room temperature**.			
<u>Students should be able to:</u> ★ Recognise polymers from diagrams showing their bonding.			
4.2.2.6 Giant covalent structures	☺	😐	☹
Substances that consist of **giant covalent structures** are **solids** with very **high melting points**. All of the atoms in these structures are linked to other atoms by **strong covalent bonds**. These bonds must be overcome to melt or boil these substances.			
Diamond and **graphite** (forms of carbon) and **silicon dioxide** (silica) are examples of giant covalent structures.			
<u>Students should be able to:</u> ★ Recognise giant covalent structures from diagrams showing their bonding.			
4.2.2.7 Properties of metals and alloys	☺	😐	☹
Metals have **giant structures** of atoms with **strong metallic bonding**. This means that most metals have **high melting** and **boiling points**.			
In pure metals, atoms are arranged in **layers**, which allows metals to be **bent** and **shaped**. Pure metals (e.g. copper, gold, iron and aluminium) are **too soft** for many uses and so are **mixed** with other metals to make **alloys** which are **harder**.			
The **different sizes** of atoms in an alloy **distort** the layers in the structure, making it more difficult for them to slide over each other, so alloys are **harder** than pure metals.			
<u>Students should be able to:</u> ★ **Explain** why **alloys** are **harder** than pure metals in terms of **distortion** of the **layers** of atoms in the structure of a pure metal.			
4.2.2.8 Metals as conductors	☺	😐	☹
Metals are **good conductors** of **electricity** because the **delocalised electrons** in the metal carry electrical charge **through the metal**.			
Metals are good conductors of **thermal energy** because energy is transferred by the delocalised electrons.			
4.2.3 Structure and bonding of carbon			
4.2.3.1 Diamond	☺	😐	☹
In **diamond**, each carbon atom forms **four covalent bonds** with other carbon atoms in a **giant covalent structure**, so diamond is **very hard**, has a **very high melting point** and does not conduct electricity.			
<u>Students should be able to:</u> ★ **Explain** the **properties** of **diamond** in terms of its structure and bonding.			

4.2.3.2 Graphite	☺	😐	☹
In **graphite**, each carbon atom forms **three covalent bonds** with three other carbon atoms, forming layers of **hexagonal rings** which have no covalent bonds between the **layers**.			
Graphite has a **high melting point**. The layers are free to **slide** over each other because there are no covalent bonds between the layers and so graphite is **soft** and **slippery**.			
In graphite, **one electron** from each carbon atom is **delocalised**. These delocalised electrons allow graphite to **conduct** thermal energy and electricity.			
Students should be able to: ★ **Explain** the **properties** of **graphite** in terms of its structure and bonding.			
★ Know that graphite is similar to metals in that it has **delocalised electrons**.			
4.2.3.3 Graphene and fullerenes	☺	😐	☹
Graphene is a **single layer** of graphite (one atom thick) 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 (e.g. high tensile strength, high electrical conductivity and high thermal conductivity).			
Students should be able to: ★ **Explain** the **properties** of **graphene** in terms of its structure and bonding.			
★ Recognise graphene and fullerenes from diagrams and descriptions of their bonding and structure.			
★ Give examples of the uses of fullerenes, including carbon nanotubes (e.g. **drug deliver**y into the body, as **lubricants**, as **catalysts** and carbon nanotubes can be used for **reinforcing materials**, e.g. in tennis rackets).			
4.2.4 Bulk and surface properties of matter including nanoparticles (chemistry only)			
4.2.4.1 Sizes of particles and their properties	☺	😐	☹
Nanoscience refers to structures that are **1–100 nm** in size, of the order of a **few hundred atoms**.			
Nanoparticles, are smaller than **fine particles** ($PM_{2.5}$), which have diameters between 100 and 2500 nm (1×10^{-7} m and 2.5×10^{-6} m).			
Coarse particles (PM_{10}) have diameters between 1×10^{-5} m and 2.5×10^{-6} m. Coarse particles are often referred to as dust.			
As the side of cube decreases by a factor of 10 the **surface area to volume ratio** increases by a factor of 10.			
Nanoparticles may have **properties** different from those for the same materials in bulk because of their **high surface area to volume ratio**. It may also mean that smaller quantities are needed to be effective than for materials with normal particle sizes.			
Students should be able to: ★ **Compare** ‘nano’ dimensions to typical dimensions of atoms and molecules.			
MS 2h Make order of magnitude calculations.			
MS 5c Calculate areas of triangles and rectangles, surface areas and volumes of cubes.			
MS 1b Recognise and use expressions in standard form.			
MS 1c Use ratios, fractions and percentages.			

MS 1d Make estimates of the results of simple calculations.			
4.2.4.2 Uses of nanoparticles	☺	😐	☹
Nanoparticles have many applications in **medicine**, in **electronics**, in **cosmetics** and **sun creams**, as **deodorants**, and as **catalysts**. New applications for nanoparticulate materials are an important area of research.			
Students should be able to: ★ Consider **advantages** and **disadvantages** of the applications of these nanoparticulate materials, but do not need to know specific examples or properties other than those specified.			
★ Given appropriate information, **evaluate** the use of nanoparticles for a specified purpose			
★ **Explain** that there are **possible risks** associated with the use of nanoparticles.			

Topic 3: Quantitative Chemistry

4.3.1 Conservation of mass and the quantitative interpretation of chemical equations			
4.3.1.1 Conservation of mass and balanced chemical equations	☺	😐	☹
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**.			
This means that chemical reactions can be represented by **symbol equations** which are **balanced** in terms of the numbers of atoms of each element involved on both sides of the equation.			
Students should: ★ **Understand** the use of the multipliers in equations in normal script before a formula and in subscript within a formula.			
4.3.1.2 Relative formula mass	☺	😐	☹
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 in the quantities shown.			
4.3.1.3 Mass changes when a reactant or product is a gas	☺	😐	☹
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 or in thermal decompositions of metal carbonates carbon dioxide is produced and escapes into the atmosphere leaving the metal oxide as the only solid product.*			
Students should be able to: ★ **Explain** any observed **changes in mass** in non-enclosed systems during a chemical reaction given the balanced symbol equation for the reaction and explain these changes in terms of the particle model.			
4.3.1.4 Chemical measurements	☺	😐	☹
Whenever a measurement is made there is always some uncertainty about the result obtained.			
Students should be able to: ★ Represent the distribution of results and make estimations of uncertainty			
★ Use the range of a set of measurements about the mean as a measure of uncertainty			

4.3.2 Use of amount of substance in relation to masses of pure substances			
4.3.2.1 Moles (HT only)	☺	😐	☹
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**. The value of the Avogadro constant is **6.02×10^{23}** per mole.			
Students should: ★ **Understand** that the measurement of amounts in moles can apply to atoms, molecules, ions, electrons, formulae and equations, for example that in one mole of carbon (C) the number of atoms is the same as the number of molecules in one mole of carbon dioxide (CO_2).			
★ Be able to use the relative formula mass of a substance to calculate the number of moles in a given mass of that substance and vice versa. [MS 1c]			
MS 1a Recognise and use expressions in decimal form.			
MS 1b Recognise and use expressions in standard form.			
MS 2a Use an appropriate number of significant figures.			
MS 3a Understand and use the symbols: =, <, <<, >>, >, ∝, ~			
MS 3b Change the subject of an equation.			
4.3.2.2 Amounts of substances in equations (HT only)	☺	😐	☹
The **masses** of **reactants** and **products** can be calculated from balanced symbol equations.			
Chemical equations can be interpreted in terms of **moles**. For example: $Mg + 2HCl \rightarrow MgCl_2 + H_2$ 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.			
Students should be able to: ★ **Calculate** the masses of substances shown in a balanced symbol equation.			
★ **Calculate** the masses of reactants and products from the balanced symbol equation and the mass of a given reactant or product.			
MS 1a Recognise and use expressions in decimal form.			
MS 1c Use ratios, fractions and percentages.			
MS 3b Change the subject of an equation.			
MS 3c Substitute numerical values into algebraic equations using appropriate units for physical quantities.			
4.3.2.3 Using moles to balance equations (HT only)	☺	😐	☹
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.			
Students should be able to: ★ **Balance an equation** given the **masses** of reactants and products.			
★ Change the subject of a mathematical equation.			
MS 3c Substitute numerical values into algebraic equations using appropriate units for physical quantities.			

4.3.2.4 Limiting reactants (HT only)	☺	😐	☹
In a chemical reaction involving **two reactants**, it is common to use an **excess** of one of the reactants to ensure that 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.			
Students should be able to: ★ **Explain** the effect of a **limiting quantity** of a reactant on the amount of products it is possible to obtain in terms of amounts in **moles** or **masses** in grams.			
4.3.2.5 Concentration of solutions	☺	😐	☹
Many chemical reactions take place in **solutions**. The **concentration** of a solution can be measured in mass per given volume of solution, e.g. grams per dm^3 (**g/dm^3**).			
Students should be able to: ★ **Calculate** the mass of **solute** in a given **volume** of solution of known **concentration** in terms of mass per given volume of solution.			
★ (HT only) **Explain** how the **mass** of a **solute** and the **volume** of a **solution** is related to the **concentration** of the **solution**.			
MS 1c Use ratios, fractions and percentages.			
MS 3b Change the subject of an equation.			

4.4 Chemical Changes

4.4.1 Reactivity of metals			
4.4.1.1 Metal oxides	☺	😐	☹
Metals react with **oxygen** to produce **metal oxides**. The reactions are **oxidation** reactions because the metals **gain oxygen**.			
Students should be able to: ★ **Explain reduction** and **oxidation** in terms of loss or gain of **oxygen**.			
4.4.1.2 The reactivity series	☺	😐	☹
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. *The reactions of metals with water and acids are limited to **room temperature** and do not include reactions with steam.*			
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**.			
Students should be able to: ★ **Recall** and **describe** the **reactions**, if any, of potassium, sodium, lithium, calcium, magnesium, zinc, iron and copper with **water** or **dilute acids** and where appropriate, to place these metals in **order of reactivity**.			
★ **Explain** how the **reactivity** of metals with water or dilute acids is related to the tendency of the metal to form its **positive ion**.			
★ **Deduce** an order of reactivity of metals based on **experimental results**.			

4.4.1.3 Extraction of metals and reduction	☺	😐	☹
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. *Knowledge of the details of processes used in the extraction of metals is not required.*			
Metals **less reactive** than **carbon** can be extracted from their oxides by **reduction** with **carbon**. *Knowledge and understanding are limited to the reduction of oxides using carbon.*			
Reduction involves the **loss of oxygen**.			
Students should be able to: ★ **Interpret** or **evaluate** specific **metal extraction processes** when given appropriate information.			
★ Identify the substances which are **oxidised** or **reduced** in terms of gain or loss of oxygen.			
4.4.1.4 Oxidation and reduction in terms of electrons (HT only)	☺	😐	☹
Oxidation is the **loss of electrons** and **reduction** is the **gain of electrons**.			
Students should be able to: ★ Write **ionic equations** for **displacement reactions**.			
★ Identify in a given reaction, symbol equation or half equation which species are **oxidised** and which are **reduced**.			
4.4.2 Reactions of acids			
4.4.2.1 Reactions of acids with metals	☺	😐	☹
Acids react with some **metals** to produce **salts** and **hydrogen**.			
Knowledge of reactions limited to those of **magnesium**, **zinc** and **iron** with **hydrochloric** and **sulfuric acids**.			
Students should be able to: ★ (HT only) **Explain** in terms of gain or loss of electrons, that these are **redox reactions**.			
★ (HT only) Identify which species are **oxidised** and which are **reduced** in given chemical equations.			
4.4.2.2 Neutralisation of acids and salt production	☺	😐	☹
Acids are neutralised by **alkalis** (e.g. soluble metal hydroxides) and **bases** (e.g. 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** • the positive ions in the base, alkali or carbonate.			
Students should be able to: ★ **Predict** products from given reactants.			
★ Use the formulae of common ions to deduce the formulae of salts.			
4.4.2.3 Soluble 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**.			

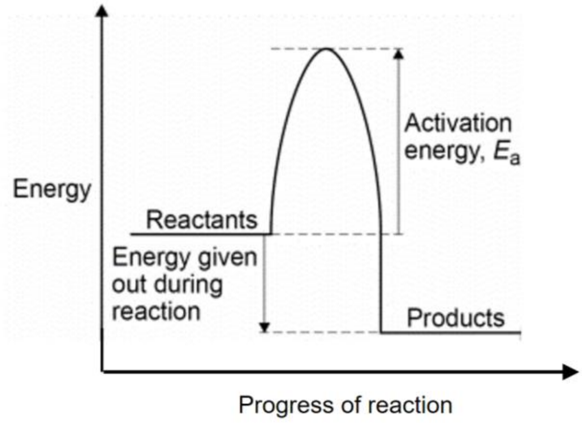
Students should be able to: ★ **Describe** how to make pure, dry samples of named soluble salts from information provided.			
REQUIRED PRACTICAL: Making salts. **AT 2, 3, 4 and 6.**			
4.4.2.4 The pH scale and neutralisation	☺	😐	☹
Acids produce **hydrogen ions** (H^+) in aqueous solutions.			
Aqueous 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**.			
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: $H^+ (aq) + OH^- (aq) \longrightarrow H_2O (l)$			
Students should be able to: ★ **Describe** the use of universal indicator or a wide range indicator to measure the approximate pH of a solution.			
★ Use the pH scale to identify acidic or alkaline solutions.			
4.4.2.5 Titration (chemistry only)			
The volumes of acid and alkali solutions that react with each other can be measured by **titration** using a suitable **indicator**.			
Students should be able to: ★ **Describe** how to carry out **titrations** using **strong acids** and **strong alkalis** only (sulfuric, hydrochloric and nitric acids only).			
★ **Calculate** the **chemical quantities** in **titrations** involving concentrations in **mol/dm^3** and in **g/dm^3**.			
4.4.2.6 Strong and weak acids (HT only)	☺	😐	☹
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**.			
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**.			
Students should be able to: ★ Use and explain the terms **dilute** and **concentrated** (in terms of amount of substance), and **weak** and **strong** (in terms of the degree of ionisation) in relation to acids.			
★ **Describe neutrality** and **relative acidity** in terms of the effect on **hydrogen ion concentration** and the numerical value of **pH** (whole numbers only).			
MS 2h Make order of magnitude calculations.			

4.4.3 Electrolysis			
4.4.3.1 The process of electrolysis	☺	😐	☹
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**.			
Students should be able to: (HT only) Throughout Section 4.4.3 ★ Write **half equations** for the reactions occurring at the **electrodes** during electrolysis ★ **Complete** and **balance** supplied **half equations**.			
4.4.3.2 Electrolysis of molten ionic compounds	☺	😐	☹
When a simple **ionic compound** (e.g. 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**.			
Students should be able to: ★ **Predict** the products of the electrolysis of binary ionic compounds in the molten state.			
4.4.3.3 Using electrolysis to extract metals	☺	😐	☹
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** using carbon as the positive electrode (anode). The mixture has a **lower melting point** than pure aluminium oxide.			
Aluminium forms at the **negative** electrode (cathode) and **oxygen** at the **positive** electrode (anode).			
The **positive** electrode (anode) is made of **carbon**, which reacts with the **oxygen** to produce **carbon dioxide** and so must be continually replaced.			
Students should be able to: ★ **Explain** why a **mixture** is used as the electrolyte			
★ **Explain** why the **positive electrode** must be continually replaced.			
4.4.3.4 Electrolysis of aqueous solutions	☺	😐	☹
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.			
Students should be able to: ★ **Predict** the products of the electrolysis of aqueous solutions containing a single ionic compound.			

REQUIRED PRACTICAL: Electrolysis. **AT 3, 7 and 8.**			
4.4.3.5 Representation of reactions at electrodes as half-equations (HT only).	☺	😐	☹
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: $2H^+ + 2e^- \rightarrow H_2$ $4OH^- \rightarrow O_2 + 2H_2O + 4e^-$ $4OH^- - 4e^- \rightarrow O_2 + 2H_2O$			

Topic 5: Energy Changes

4.5.1 Exothermic and endothermic reactions			
4.5.1.1 Energy transfer during exothermic and endothermic reactions	☺	😐	☹
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. If a reaction transfers energy to the surroundings the product molecules must have less energy than the reactants, by the amount transferred.			
An **exothermic reaction** is one that transfers **energy <u>to</u> the surroundings** so the temperature of the surroundings increases.			
Exothermic reactions include **combustion**, many **oxidation reactions** and **neutralisation**.			
Everyday uses of exothermic reactions include **self-heating cans** and **hand warmers**.			
An **endothermic reaction** is one that takes in **energy <u>from</u> the surroundings** so the temperature of the surroundings decreases.			
Endothermic reactions include **thermal decompositions** and the reaction of **citric acid** and **sodium hydrogen carbonate**. Some **sports injury packs** are based on endothermic reactions.			
<u>Students should be able to:</u> ★ Distinguish between exothermic and endothermic reactions on the basis of the **temperature change** of the surroundings. *Limited to measurement of temperature change. Calculation of energy changes or ΔH is not required.*			
★ **Evaluate** uses and applications of exothermic and endothermic reactions given appropriate information.			
AT 5 An opportunity to measure temperature changes when substances react or dissolve in water.			
REQUIRED PRACTICAL: Temperature changes. **AT 1, 3, 5 and 6.**			
4.5.1.2 Reaction profiles	☺	😐	☹
Chemical reactions can occur only when reacting particles **collide** with each other and with sufficient energy. The **minimum** amount of **energy** that particles must have to react is called the **activation energy**.			
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.			

A **reaction profile** for an **exothermic reaction** can be drawn in the following form: 			
<u>Students should be able to:</u> ★ Draw simple **reaction profiles** (energy level diagrams) for **exothermic** and **endothermic** reactions showing the relative energies of reactants and products, the activation energy and the overall energy change, with a curved line to show the energy as the reaction proceeds.			
★ Use reaction profiles to identify reactions as exothermic or endothermic.			
★ **Explain** that the **activation energy** is the energy needed for a reaction to occur.			
4.5.1.3 The energy change of reactions (HT only)	☺	😐	☹
During a chemical reaction: • energy must be **supplied** to **break bonds** in the **reactants**. • energy is **released** when **bonds** in the **products** 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.			
<u>Students should be able to:</u> ★ **Calculate** the energy transferred in chemical reactions using bond energies supplied.			