Unit 1 – Chemical fundamentals: structure, properties and reactions

Unit description

Chemists design and produce a vast range of materials for many purposes, including for fuels, cosmetics, building materials and pharmaceuticals. As the science of chemistry has developed over time, there has been an increasing realisation that the properties of a material depend on, and can be explained by, the material's structure. A range of models at the atomic and molecular scale enable explanation and prediction of the structure of materials and how this structure influences properties and reactions. In this unit, students relate matter and energy in chemical reactions as they consider the breaking and reforming of bonds as new substances are produced. Students can use materials that they encounter in their lives as a context for investigating the relationships between structure and properties.

Through the investigation of appropriate contexts, students explore how evidence from multiple disciplines and individuals have contributed to developing understanding of atomic structure and chemical bonding. They explore how scientific knowledge is used to offer reliable explanations and predictions, and the ways in which it interacts with social, economic and ethical factors.

Students use science inquiry skills to develop their understanding of patterns in the properties and composition of materials. They investigate the structure of materials by describing physical and chemical properties at the macroscopic scale, and use models of structure and primary bonding at the atomic and sub-atomic scale to explain these properties. They are introduced to the mole concept as a means of quantifying matter in chemical reactions.

Learning outcomes

By the end of this unit, students:

- understand how the atomic model and models of bonding explain the structure and properties of elements and compounds
- understand the concept of enthalpy, and apply this to qualitatively and quantitatively describe and explain energy changes in chemical reactions
- understand how models and theories have developed based on evidence from a range of sources, and the uses and limitations of chemical knowledge in a range of contexts
- use science inquiry skills to design, conduct, evaluate and communicate investigations into the properties of elements, compounds and mixtures and the energy changes involved in chemical reactions
- evaluate, with reference to empirical evidence, claims about chemical properties, structures and reactions
- communicate, predict and explain chemical phenomena using qualitative and quantitative representations in appropriate modes and genres.

Unit content

This unit includes the knowledge, understandings and skills described below.

Science Inquiry Skills

- identify, research and refine questions for investigation; propose hypotheses; and predict possible outcomes
- design investigations, including the procedure(s) to be followed, the materials required, and the type and amount of primary and/or secondary data to be collected; conduct risk assessments; and consider research ethics
- conduct investigations safely, competently and methodically for the collection of valid and reliable data, including: the use of devices to accurately measure temperature change and mass, flame tests, separation techniques and heat of reaction
- represent data in meaningful and useful ways, including using appropriate graphic representations and correct units and symbols; organise and process data to identify trends, patterns and relationships; identify sources of random and systematic error and estimate their effect on measurement results; and select, synthesise and use evidence to make and justify conclusions
- interpret a range of scientific and media texts, and evaluate processes, claims and conclusions by considering the quality of available evidence; and use reasoning to construct scientific arguments
- communicate to specific audiences and for specific purposes using appropriate language, nomenclature and formats, including scientific reports

2010 Q 38c, d, e, f, g

2011 Q 37b, c, d

2013 Q 38a-g, p

Properties and structure of atoms

Science as a Human Endeavour

Findings from a range of scientific experiments contributed to the understanding of the atom, enabling scientists, including Dalton, Thomson, Rutherford, Bohr and Chadwick to develop models of atomic structure and make reliable predictions about the mass, charge and location of the sub-atomic particles.

Science Understanding

- **elements are represented by symbols**

  Each element has a unique symbol comprised of one or two letters. The first letter is a capital followed by a lower case. Most symbols use letters that match the English name, though there are a number that are based on Latin or Greek names e.g. Fe (iron) comes from ferrum.

- **atoms can be modelled as a nucleus, surrounded by electrons in distinct energy levels, held together by electrostatic forces of attraction between the nucleus and electrons; the location of electrons within atoms can be represented using electron configurations**

  Nucleus – positive protons determine element, neutral neutrons add to mass.

  Electron cloud – negative electrons in energy levels around the nucleus – determine size and chemical properties of the element

  Energy level (shell) n can hold $2n^2$ electrons, i.e. –

  Level 1 – 2 electrons

  Level 2 – 8 electrons

  Level 3 – 18 electrons

  Level 4 – 32 electrons

  Electrons fill from the lowest level first, though K and Ca start filling level 4 after only 8 electrons go into level 3 (due to levels split into sublevels)

  For example - give the electron configuration of –

  Carbon 2 6

  Chlorine 2 8 7

  Sodium 2 8 1

  Potassium 2 8 8 1

WS1 Electron Configurations

2011 Q2

2012 Q1

- **the ability of atoms to form chemical bonds can be explained by the arrangement of electrons in the atom and in particular by the stability of the valence electron shell**

  Elements with a filled valence electron shell are stable. This is best illustrated by the inert gases in group 18 which are unreactive due to their filled shell. Other atoms gain this stable configuration by losing electrons (to form positive ions), gaining electrons (to form negative ions) or sharing electrons (to form a covalent bond).

- **the structure of the periodic table is based on the atomic number and the properties of the elements**

  Atoms are placed in order of atomic number (number of protons). They are then grouped according to properties such that elements in the same group have the same (or similar) properties. E.g. group 1 is the alkali metals which are all soft and very reactive, whereas group 18 is the inert gases which are unreactive.

2010 Q8, 12

- **the elements of the periodic table show trends across periods and down main groups, including in atomic radii, valencies, 1st ionisation energy and electronegativity as exemplified by groups 1, 2, 13–18 and period 3**

  The electron configurations of a group have the same number of valence (outer shell) electrons. E.g. group 1

H	1
Li	2 1
Na	2 8 1
K	2 8 8 1

The loss or gain of electrons leads to each element of a group having the same valence – group 1 is +1, group 2 is +2, group 13 is +3, group 14 is +/-4, group 15 is -3, group 16 is -2, group 17 is -1 and group 18 is 0.

Ionisation energy is the energy to remove the most loosely bound electron from an atom. Successive ionisation energies increase due to increased attraction to the relatively more positive nucleus. There is a big jump in IE when an electron is removed from a new energy level – which is closer to the nucleus and filled.

Electronegativity is the ability of an atom to attract electrons.

Going across a period the electrons in the same energy level are attracted by an increasing number of protons in the nucleus – this causes

- ionisation energy to increase
- electronegativity to increase (except inert gases which have zero electronegativity)
- atomic radius to decrease

Going down a group electrons are in energy levels further from the nucleus and are shielded from the increasingly positive nucleus due to filled energy levels – this causes –

- ionisation energy to decrease
- electronegativity to decrease
- atomic radius to increase

WS2 Periodic Table

2011 Q3, 35c

2013 Q8

Stage 3 WACE Q 28/2a, 38

- **flame tests and atomic absorption spectroscopy (AAS) are analytical techniques that can be used to identify elements; these methods rely on electron transfer between atomic energy levels and are shown by line spectra**

  When energy is added to an atom the electrons absorb this energy and go into higher energy levels. This can be seen as lines on an absorption spectrum which are the same for a given element and allow for identification of the element. When the electrons return to their original level light is given off which is seen as a specific colour for each element, e.g. sodium gives off yellow light and potassium pale purple light.

- **isotopes are atoms of an element with the same number of protons but different numbers of neutrons and are represented in the form ^{A}X (IUPAC) or X-A**

  Atomic number (Z) is the number of protons in the nucleus. This determines the atom – i.e. all carbon atoms have 6 protons.

  Mass number (A) is the number of protons and neutrons in the nucleus. This determines the mass of the atom and as the number of neutrons can vary, so does the mass of the atom. E.g. carbon-12 has 6 neutrons while carbon-14 has 8 neutrons.

  The IUPAC representation for carbon-12 is ^{12}C.

  The number of electrons in a neutral atom is the same as the proton number.

  For example - what information is available for – ^{23}Na (Z=11)

  Atomic number = 11

  Mass number = 23

  11 protons

  12 neutrons

  11 electrons

WS3 Atomic Structure

WS4 Atomic and Mass Numbers

2010 Q7, 27

2011 Q1, 26a

2012 Q26

2013 Q9, 26

- **isotopes of an element have the same electron configuration and possess similar chemical properties but have different physical properties**

  Because the number of protons does not vary for a given element, the number of electrons in a neural atom is the same giving the same electron configuration and chemical properties.

  The different physical properties such as density, melting and boiling points are all due to the slight differences in mass.

- **the relative atomic mass (atomic weight), A_r is the ratio of the average mass of the atom to 1/12 the mass of an atom of ^{12}C; relative atomic masses of the elements are calculated from their isotopic composition**

  Carbon-12 is chosen as the reference value for the relative atomic masses of the other elements. This can be calculated from the average masses of the isotopes.

  For example - what is the relative atomic mass of chlorine if there is 25% Cl-37 and 75% Cl-35?

  A_r (Cl) = 0.25 x 37 + 0.75 x 35

  = 35.5

- **mass spectrometry involves the ionisation of substances and the separation and detection of the resulting ions; the spectra which are generated can be analysed to determine the isotopic composition of elements and interpreted to determine relative atomic mass**

  As different isotopes have different masses these are differently affected in mass spectrometry. The percentages of each isotope can be determined and used to calculate the relative atomic mass of the element (as in the previous dot point.

Properties and structure of materials

Science as a Human Endeavour

Matter at the nanoscale can be manipulated to create new materials, composites and devices; the different characteristics of nanomaterials can be used to provide commercially available products. As products are designed on the basis of properties which are different from the bulk material, their use can be associated with potential risks to health, safety and the environment and this has led to regulations being developed to address new and existing nanoform materials.

Science Understanding

- **materials are pure substances with distinct measurable properties, including melting and boiling points, reactivity, hardness and density; or mixtures with properties dependent on the identity and relative amounts of the substances that make up the mixture**

  A pure substance contains one type of substance only – and hence has specific and identifiable physical and chemical properties. Also known as homogeneous – uniform composition.

  Mixtures are made up of two or more substances and their properties are going to vary depending on the composition of the mixture. Known as heterogeneous – non-uniform composition.

  Solutions are special examples of mixtures that are uniform in composition – homogeneous mixtures.

2010 Q41d

2011 Q8

2012 Q5, 24

2013 Q1, 31a,b

- **pure substances may be elements or compounds which consist of atoms of two or more elements chemically combined; the formulae of compounds indicate the relative numbers of atoms of each element in the compound**

  Elements are made up of one type of atom only – defined by atoms with the same number of protons.

  Compounds consist of two or more types of atoms chemically combined in a fixed ratio.

  Formulas are used to indicate the relative number of atoms present

  For example – how many atoms of each type are present in –?

  H_2 – 2 x H

  NaCl – 1 x Na, 1 x Cl

  CO_2 – 1 x C, 2 x O

- **nanomaterials are substances that contain particles in the size range 1–100 nm and have specific properties relating to the size of these particles which may differ from those of the bulk material**

  One of the reasons for the variation in properties is due to the increased relative surface area such that a greater proportion of atoms are on the surface compared to those on the inside. This can affect melting point, reactivity, strength and electrical characteristics.

- **differences in the physical properties of substances in a mixture, including particle size, solubility, density, and boiling point, can be used to separate them**

  Common separation techniques include –

    - filtration – insoluble substance is trapped in the filter paper while the soluble substance passes through
    - evaporation – liquid is evaporated away leaving the solid behind
    - distillation – different liquids boil at different temperatures and are collected separately

  For example - how would you separate -?

  Salt and sand – add water, salt dissolves, filter out sand, and then evaporate the water to get the salt.

  Salt from salt water – evaporate the water. If water is needed as well use distillation.

  Water and ethanol – distillation with ethanol coming off before the water as it has a lower boiling point.

- **the type of bonding within ionic, metallic and covalent substances explains their physical properties, including melting and boiling points, conductivity of both electricity and heat and hardness**

  Melting and boiling points, and hardness indicate the strength of attraction between the particles.

  Electrical conductivity indicates the presence of mobile charged particles.

  Heat conductivity indicates the presence of mobile particles.

Tutorial 1 Bonding

2010 Q9, 10, 42

2011 Q4, 28b

2013 Q32a,b

- **chemical bonds are caused by electrostatic attractions that arise because of the sharing or transfer of electrons between participating atoms; the valency is a measure of the bonding capacity of an atom**

  Electrostatic attraction is between something positive and something negative. This is different for each different type of bond.

  Valency is determined by the number of electrons lost, gained or shared by the atoms in a bond.

2011 Q35a

2012 Q2

- **ions are atoms or groups of atoms that are electrically charged due to a loss or gain of electrons; ions are represented by formulae which include the number of constituent atoms and the charge of the ion**
  (for example, O^{2-}, SO_4^{2-})

  Simple ions are those that have one atom only. Many of these have a charge determined by position in the Periodic Table and whether electrons are lost or gained to obtain inert gas configuration –

  Group 1 +1

  Group 2 +2

  Group 13 +3

  Group 15 -3

  Group 16 -2

  Group 17 -1

  Metals in the middle of the table (Transition Metals) are mostly +2 except Ag (+1) and Cr (+3).

  There are also a number of atoms that take more than one charge – which is indicated in the ion name –

  Cu (+1 or +2) – copper I or copper II

  Sn and Pb (+2 or +4)

  Fe (+2 or +3)

  Polyatomic ions have more than one atom and need to be learned.

2013 Q31c

- **ionic bonding can be modelled as a regular arrangement of positively and negatively charged ions in a crystalline lattice with electrostatic forces of attraction between oppositely charged ions**

2012 Q12, 22

- **the ionic bonding model can be used to explain the properties of ionic compounds, including high melting point, brittleness and non-conductivity in the solid state; the ability of ionic compounds to conduct electricity when molten or in aqueous solution can be explained by the breaking of the bonds in the lattice to give mobile ions**

See Tutorial 1 Bonding

- **the formulae of ionic compounds can be determined from the charges on the relevant ions**

  The total charge on any compound must be zero. This means that the same amount of positive charge from the metal ions and negative charge from the non-metal ions is required.

  For example –

  Sodium chloride – Na +1 and Cl -1 so charges are equal NaCl

  Magnesium chloride – Mg +2 and Cl -1 so need 2 Cl to make charges equal $MgCl_2$

WS5 Ionic Formulas 1

WS6 Ionic Formulas 2

WS7 Ionic Formulas and Names

2010 Q26

2011 Q27

2012 Q4, 27

2013 Q10, 27

- **metallic bonding can be modelled as a regular arrangement of atoms with electrostatic forces of attraction between the nuclei of these atoms and their delocalised electrons that are able to move within the three dimensional lattice**

Stage 3 WACE Q 28/2b, 28/3

- **the metallic bonding model can be used to explain the properties of metals, including malleability, thermal conductivity, generally high melting point and electrical conductivity; covalent bonding can be modelled as the sharing of pairs of electrons resulting in electrostatic forces of attraction between the shared electrons and the nuclei of adjacent atoms**

See Tutorial 1 Bonding

2011 Q5

2012 Q23, 32

Stage 3 WACE Q 26/1

- **the properties of covalent network substances, including high melting point, hardness and electrical conductivity, are explained by modelling covalent networks as three-dimensional structures that comprise covalently bonded atoms**

See Tutorial 1 Bonding

- **elemental carbon exists as a range of allotropes, including graphite, diamond and fullerenes, with significantly different structures and physical properties**

  Allotropes are different forms of the same element. As the structures are different, their properties will also be different.

  Diamond has each carbon attached to 4 other carbons in a continuous lattice. This makes the diamond extremely hard with high melting point due to having to break many strong covalent bonds.

  Graphite is layers of carbon atoms with each carbon attached to 3 others. This leaves one electron unbonded and delocalised making graphite able to conduct electricity. The layers also make graphite soft and a good lubricant.

  Fullerenes are small molecules of carbon including buckey balls and nanotubes. Their small size mean very different properties to diamond and graphite.

- **the properties of covalent molecular substances, including low melting point, can be explained by their structure and the weak intermolecular forces between molecules; their non-conductivity in the solid and liquid/molten states can be explained by the absence of mobile charged particles in their molecular structure**

See Tutorial 1 Bonding

Stage 3 WACE Q 33

- **molecular formulae represent the number and type of atoms present in the molecules**

  Covalent compounds are named using the following –

    - first element name
    - second element name ends in "ide"
    - prefixes are used to indicate the number of atoms. Mono (one) is only used when there is one atom of the second element.

  Reversing this process gives the formula!

WS8 Covalent Formulas

Tutorial 2 Formulas (includes element formulas)

2010 Q26

2011 Q27

2012 Q27

2013 Q11, 27

- **percentage composition of a compound can be calculated from the relative atomic masses of the elements in the compound and the formula of the compound**

Percentage by mass is found from the formula of the compound, using relative atomic masses.

For example

H_2O	H	2x1.008	%H = 2.016/18.016 x 100 =11.19%
	O	16.00	%O = 16.00/18.016 x 100 = 88.81%
	Total	18.016	

2010 Q32b

2011 Q38a

2012 Q30

2013 Q37a

- **hydrocarbons, including alkanes, alkenes and benzene, have different chemical properties that are determined by the nature of the bonding within the molecules**

Hydrocarbons contain carbon and hydrogen only, covalently bonded.

Alkanes have single bonds only between the carbons and are generally unreactive.

Alkenes have one or more double bonds between the carbons, these bonds are very reactive.

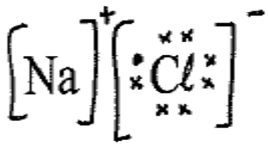
Benzene is a 6 carbon ring, with one hydrogen per carbon atom. The remaining electrons are found in a delocalised ring above and below the ring. It looks as though there are three double bonds but the unreactive nature of benzene indicates that these are not double bonds.

2010 Q40

- **molecular structural formulae (condensed or showing bonds) can be used to show the arrangement of atoms and bonding in covalent molecular substances**

Lewis diagrams are used to show the bonding in both ionic compounds and covalent molecules. These show the electrons that are in each bond.

For example NaCl can be represented by –

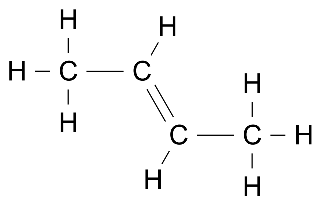


And hydrogen gas by –

H:H or H-H

Carbon compounds are not usually drawn using the electrons, but have two main forms – condensed and all bonds.

For example –



Is the same as – $CH_3CH{=}CHCH_3$ and $CH_3CHCHCH_3$

Tutorial 3 Representing Bonding

2010 Q28, 40f,g

2011 Q21, 22, 28a, 35b

2012 Q28

2013 Q28, 32c, 33

- **IUPAC nomenclature is used to name straight and simple branched alkanes and alkenes from C_1- C_8**

Some information about carbon chemistry –

- Carbon is unique in its chemistry, being the only element that forms long chains and rings of the same atom.
- Each carbon atom has 4 valence electrons, and can hence form 4 bonds.
- The general arrangement of the bonds around a carbon atom is **tetrahedral**, as pairs of electrons push each other away.

- Carbon compounds are named using the IUPAC system which is a standard system for nomenclature (naming). The name and the formula are interchangeable, which makes carbon compounds the easiest of all, providing you learn a few rules.
- Structural isomers have the same number of atoms of each type but different arrangements. These will always have different names.

- Find the longest carbon chain.
- Use the prefixes (see below) to indicate this longest chain.
- End the name in "**-ane**" if single bonds only are present, "**-ene**" if a double bond is present
- Prefix the entire name with "**cyclo-**" if the carbon atoms are present in a ring.
- Look for any attached groups on the carbon chain or ring, and name these as "**-yl**" if a carbon group, or chloride, bromide etc if halogens.
- Number the carbon atoms to give the **lowest** possible numbers to the carbon atoms with attached groups.
- Numbers can also distinguish where double or triple bonds are found – these should have the lowest possible number, regardless of attached group.
- Indicate in the name, the position (carbon number), the type of bonds (single, double or triple) and the number and type of groups (e.g. dimethyl).

PREFIXES FOR CARBON NUMBER

meth-1	hex-6
eth-2	hept-7
prop-3	oct-8
but-4	non-9
pent- 5	dec-10

WS9 Alkanes and Alkenes

WS10 Cyclic Hydrocarbons

WS11 Hydrocarbons

2010 Q17, 18

2011 Q23

2012 Q17, 33

2013 Q 21, 23, 34a

- **alkanes, alkenes and benzene undergo characteristic reactions such as combustion, addition reactions for alkenes and substitution reactions for alkanes and benzene**

Halogens can react with alkanes in the presence of light causing a hydrogen atom to be substituted with a halogen atom. This reaction is hard to control and a mixture of haloalkanes is usually produced. The other product is a hydrogen halide. Benzene also undergoes substitution, usually with a catalyst.

For example -methane and chlorine reactions –

$CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$

$CH_3Cl + Cl_2 \rightarrow CH_2Cl_2 + HCl$

Etc!

The double bond in alkenes allows small molecules such as hydrogen, halogens, hydrogen halides and water to be added. The double bond breaks leaving a single bond behind.

For example -

Ethene + hydrogen (metal catalyst such as Ni, Pt or Pd)

$CH_2{=}CH_2 + H_2 \rightarrow CH_3CH_3$

Ethene + bromine

$CH_2{=}CH_2 + Br_2 \rightarrow CH_2BrCH_2Br$

Ethene + hydrogen chloride

$CH_2{=}CH_2 + HCl \rightarrow CH_2ClCH_3$

Ethene + water (sulfuric acid catalyst)

$CH_2{=}CH_2 + H_2O \rightarrow CH_3CH_2OH$

In the presence of sufficient oxygen, hydrocarbons react to produce carbon dioxide and water. If the oxygen is limited, carbon monoxide may be produced.

WS12 Reactions of Hydrocarbons

2010 Q19, 40h,i,j

2011 Q25, 34

2012 Q33d, 36

2013 Q22, 24, 34b

Chemical reactions: reactants, products and energy change

Science as a Human Endeavour

There are differences in the energy output and carbon emissions of fossil fuels (including coal, oil, petroleum and natural gas) and biofuels (including biogas, biodiesel and bioethanol). These differences, together with social, economic, cultural and political values, determine how widely these fuels are used.

Science Understanding

- **chemical reactions can be represented by chemical equations; balanced chemical equations indicate the relative numbers of particles (atoms, molecules or ions) that are involved in the reaction**

  A chemical reaction indicates the following –

  - reactants (left hand side)
  - products (right hand side)
  - state of reactants and products
  - number of each reactant and product

  As matter cannot be created nor destroyed the mass of reactants must equal the mass of products. This means that the number of each type of atom must be same on both sides of the equation in order for it to be balanced.

  For example

  $$2C\ (s) + O_2\ (g) \rightarrow 2CO\ (g)$$

  The reactants are solid carbon and gaseous oxygen producing (indicated by the →) gaseous carbon monoxide. There are 2 carbon atoms and 2 oxygen atoms on each side of the equation so it is balanced.

Tutorial 4 Writing Equations

WS13 Balancing Equations 1

WS14 Balancing Equations 2

2010 Q6

2013 Q13, 38q

- **chemical reactions and phase changes involve enthalpy changes, commonly observable as changes in the temperature of the surroundings and/or the emission of light**

  Enthalpy is a measure of the heat energy present in the chemical species involved in the reaction. During a chemical reaction bonds are broken (which takes energy) and new bonds are formed (which releases energy) which leads to a change in the enthalpy present. The difference is released as heat or light, or absorbed as heat, leading to changes in temperature.

2011 Q32a,b

- **endothermic and exothermic reactions can be explained in terms of the Law of Conservation of Energy and the breaking of existing bonds and forming of new bonds; heat energy released or absorbed by the system to or from the surroundings, can be represented in thermochemical equations**

  Change in enthalpy is defined as product enthalpy – reactant enthalpy.

  Endothermic reactions require additional energy to occur. This is supplied by the environment and is often observed as a decrease in temperature. As energy is added to the system the products have greater enthalpy than the reactants. This means that more energy is needed to break the bonds than is released by the formation of new bonds. The energy change is written as a reactant and is positive.

  Exothermic reactions give off energy. This is generally seen as an increase in temperature. As energy is released the reactants have more enthalpy than the products. More energy is released by the formation of new bonds than is required to break bonds. The energy change is written as a product and is negative.

WS15 Heats of Reaction

2010 Q24

2011 Q17

2013 Q14, 15

- **fossil fuels (including coal, oil, petroleum and natural gas) and biofuels (including biogas, biodiesel and bioethanol) can be compared in terms of their energy output, suitability for purpose, and the nature of products of combustion**

  Two important aspects of fuels are fuel value (energy output) and carbon emissions.

  Fossil fuels produce carbon dioxide which contributes to global warming by increasing the greenhouse effect.

  Biofuels such as ethanol are considered carbon neutral as the amount of carbon dioxide released equals the amount needed to produce the fuel. These are also renewable sources of energy as opposed to non-renewable fossil fuels. Australia only uses about 4% renewable energy as of 2011.

2011 Q37a

- **the mole is a precisely defined quantity of matter equal to Avogadro's number of particles**

In order to compare amounts of a substance (such as in an equation) we need to know the number of particles involved. The mole is a large counting number equal to the number of particles in 12.00g of carbon. This is also known as Avogadro's number and is equal to 6.022×10^{23}.

WS16 Mole and Number of Particles

- **the mole concept relates mass, moles and molar mass and, with the Law of Conservation of Mass; can be used to calculate the masses of reactants and products in a chemical reaction**

Molar mass (M) is defined as the mass (m) of one mole (i.e. Avogadro's number) of a substance. It is determined from the relative atomic mass(es) of the atom(s) involved. This leads to a simple formula to find mole (n) –

$$n = m/M$$

Tutorial 5 Chemical Calculations

WS17 Mole Concept Calculations

WS18 Mole Calculations 1

WS19 Mole Calculations 2

The Law of Conservation of Mass which was used to balance equations relates also to calculations. Mass cannot be created or destroyed so the relationship between reacting substances gives information about the masses involved.

Stoichiometry is the mole relationship between atoms in a compound and substances in a reaction. It is the same as the atom ratio or substance ratio in an equation.

Tutorial 6 Stoichiometry

WS20 Mole Calculations 3

In practice it is rare to use stoichiometric quantities in reactions. In order to ensure that one reactant is completely reacted the other reactant(s) may be in excess where more than is needed is added.

The limiting reagent is the one which is completely consumed and determines the amount of products produced. The excess reactant is the one which will have some left over. In order to determine which is which, the number of mole of each is calculated and compared to the ratio from the equation.

Tutorial 7 Limiting Reagent

WS21 Limiting Reagent

The final consideration in chemical calculations is the fact that this is experimental and hence involves measurement. The final answer to a calculation must reflect the precision of the measurements involved as it is not possible to get a more precise answer simply by calculating! This precision is indicated by significant figures, and answers will generally be stated to 3 or 4 significant figures.

Tutorial 8 Significant Figures

WS21 Significant Figures

2011 Q40b

2012 Q19

2013 Q12, 37b,c,d, 38h-o

Unit 2 – Molecular interactions and reactions

Unit description

Students develop their understanding of the physical and chemical properties of materials, including gases, water and aqueous solutions, acids and bases. Students explore the characteristic properties of water that make it essential for physical, chemical and biological processes on Earth, including the properties of aqueous solutions. They investigate and explain the solubility of substances in water, and compare and analyse a range of solutions. They learn how rates of reaction can be measured and altered to meet particular needs, and use models of energy transfer and the structure of matter to explain and predict changes to rates of reaction. Students gain an understanding of how to control the rates of chemical reactions, including through the use of a range of catalysts.

Through the investigation of appropriate contexts, students explore how evidence from multiple disciplines and individuals have contributed to developing understanding of intermolecular forces and chemical reactions. They explore how scientific knowledge is used to offer reliable explanations and predictions, and the ways in which it interacts with social, economic and ethical factors.

Students use a range of practical and research inquiry skills to investigate chemical reactions, including the prediction and identification of products and the measurement of the rate of reaction. They investigate the behaviour of gases, and use the Kinetic Theory to predict the effects of changing temperature, volume and pressure in gaseous systems.

Learning outcomes

By the end of this unit, students:

- understand how models of the shape and structure of molecules and intermolecular forces can be used to explain the properties of substances, including the solubility of substances in water
- understand how kinetic theory can be used to explain the behaviour of gaseous systems, and how collision theory can be used to explain and predict the effect of varying conditions on the rate of reaction
- understand how models and theories have developed based on evidence from a range of sources, and the uses and limitations of chemical knowledge in a range of contexts
- use science inquiry skills to design, conduct, evaluate and communicate investigations into the properties and behaviour of gases, water, aqueous solutions and acids and bases, and into the factors that affect the rate of chemical reactions
- evaluate, with reference to empirical evidence, claims about chemical properties, structures and reactions
- communicate, predict and explain chemical phenomena using qualitative and quantitative representations in appropriate modes and genres.

Unit content

This unit builds on the content covered in Unit 1.

This unit includes the knowledge, understandings and skills described below.

Science Inquiry Skills

- identify, research, construct and refine questions for investigation; propose hypotheses; and predict possible outcomes
- design investigations, including the procedure(s) to be followed, the materials required, and the type and amount of primary and/or secondary data to be collected; conduct risk assessments; and consider research ethics
- conduct investigations safely, competently and methodically for the collection of valid and reliable data, including: chromatography, measuring pH, rate of reaction, identification of the products of reactions, and determination of solubilities of ionic compounds to recognise patterns in solubility
- represent data in meaningful and useful ways, including using appropriate graphic representations and correct units and symbols; organise and process data to identify trends, patterns and relationships; identify sources of random and systematic error; identify anomalous data; estimate the effect of error on measured results; and select, synthesise and use evidence to make and justify conclusions
- interpret a range of scientific and media texts, and evaluate processes, claims and conclusions by considering the quality of available evidence; and use reasoning to construct scientific arguments
- communicate to specific audiences and for specific purposes using appropriate language, nomenclature and formats, including scientific reports

Stage 3 WACE Q 37

Intermolecular forces and gases

Science as a Human Endeavour

Chromatographic techniques, including thin layer chromatography (TLC), gas chromatography (GC), and high performance liquid chromatography (HPLC), are used to determine the components of a wide range of mixtures in various settings. The decision to use a particular chromatographic technique depends on a number of factors, including the properties of the substances being separated, the amount of substance available for analysis and the sensitivity of the equipment. Chromatographic techniques have a wide range of analytical and forensic applications, including monitoring air and water pollutants, drug testing of urine and blood samples, and testing for food additives and quality.

Science Understanding

- **observable properties, including vapour pressure, melting point, boiling point and solubility, can be explained by considering the nature and strength of intermolecular forces within a covalent molecular substance**

Melting and boiling points depend on the strength of bonding present. The stronger the bonding, the higher the melting and boiling points. For molecules of similar size, the order of strength is dispersion < dipole-dipole < hydrogen bonding.

Increasing the size of a molecule increases the number of electrons to be involved in dispersion forces, so this also increases the melting and boiling points. Note – this is due to the number of electrons, not to be stated as the molar mass of substance.

For each group there is a general increase in melting and boiling points due to the increased number of electrons as the molecules become larger. The unexpectedly high values of NH_3, H_2O and HF is due to these molecules having hydrogen bonding which is the strongest intermolecular force.

For example - which has greater BP?

- Nitrogen gas or helium gas – nitrogen as it is a larger molecule than the helium atom so it has greater dispersion forces.
- Methane or dichloromethane – dichloromethane is polar so will experience dipole-dipole forces which are greater than the dispersion forces of methane.
- Water or carbon dioxide – water as it has hydrogen bonding which is stronger than the dispersion forces of carbon dioxide.

2010 Q5

- **the valence shell electron pair repulsion (VSEPR) theory and Lewis structure diagrams can be used to explain, predict and draw the shapes of molecules**

VSEPR states that pairs of electrons repel to be as far away from each other as possible.

To decide on a shape –

- draw an electron dot (Lewis structure) diagram for the substance

- count the number of bonding and non-bonding pairs around the central atom
- if there are 2 pairs – shape is LINEAR
- if there are 3 pairs – shape is TRIANGULAR PLANAR
- if there are 3 pairs where one is a non-bonding pair – shape is BENT or V-SHAPE
- if there are 4 pairs – shape is TETRAHEDRAL
- if there are 4 pairs where one pair is non-bonding – shape is TRIANGULAR PYRAMID
- If there are 4 pairs where 2 pairs are non-bonding – shape is BENT or V-SHAPE.

E.g. – Draw electron dot diagrams and decide on the shape for –

SO_2

- bent

SO_3

- triangular planar

SO_3^{2-}

- pyramidal

SO_4^{2-}

- tetrahedral

See Tutorial 3 Representing Bonding

Stage 3 WACE Q 26/3, 28/1

- **the polarity of molecules can be explained and predicted using knowledge of molecular shape, understanding of symmetry, and comparison of the electronegativity of atoms involved in the bond formation**

When atoms differ in electronegativity the electrons are found closer to the atom with greater electronegativity. This sets up a charge distribution in the bond and is considered to be polar.

All covalent bonds between different atoms are polar (except C-H bonds). Covalent bonds between the same atoms are non-polar.

A polar molecule requires polar bonds and an overall dipole or charge.

An atom will be perfectly symmetrical if all of the species around the central atom are the same. It will be non-symmetrical if there are different species (including non-bonding electron pairs) around the central atom.

When a molecular shape is perfectly symmetrical the polar bonds effectively cancel each other out leaving the molecule non-polar.

E.g. CO_2

O=C=O linear

When a molecular shape is not symmetrical the polar bonds add to give an overall charge or dipole and the molecule is polar.

E.g. H_2O

```
  ..
: O – H     bent
  |
  H
```

Stage 3 WACE Q27/2, 32

- **the shape and polarity of molecules can be used to explain and predict the nature and strength of intermolecular forces, including dispersion forces, dipole-dipole forces and hydrogen bonding**

  - dispersion forces

  Non-polar molecules have no permanent dipole. At times the electrons will be distributed in such a way that there is an overall charge distribution or temporary dipole. This can induce a dipole in neighbouring molecules leading to weak attraction between them. These are the weakest intermolecular forces.

  - dipole-dipole attractions

  Polar molecules have a permanent dipole which allows the positive end to attract to the negative end of another molecule. These are stronger forces than dispersion, but still weak.

  - hydrogen bonds

  Molecules with hydrogen atoms attached to nitrogen, oxygen or fluorine (the 3 most electronegative atoms) have a very strong dipole, with the hydrogen atom almost stripped of its electron. A weak dative/coordinate bond is formed between the electron deficient hydrogen and a non-bonding pair of electrons on another nitrogen, oxygen or fluorine forms the strongest type of intermolecular force.

Stage 3 WACE Q 26/2, 27/1, 34/1, 34/2

- **data from chromatography techniques, including thin layer chromatography (TLC), gas chromatography (GC), and high-performance liquid chromatography (HPLC), can be used to determine the composition and purity of substances; the separation of the components is caused by the variation in strength of the interactions between atoms, molecules or ions in the mobile and stationary phases**

  Chromatography works using the attraction/bonding strength between the substances present and the mobile and stationary phases. Some of the substances will be more attracted to the mobile phase and move further. Those that are more attracted to the stationary phase will not move as far.

- **the behaviour of an ideal gas, including the qualitative relationships between pressure, temperature and volume, can be explained using the Kinetic Theory**

According to the Kinetic Theory gases have particles that are far apart, moving freely without appreciable attraction between them. This accounts for the properties of gases which includes diffusion, ability to pour and variable shape and volume.

Temperature is defined as the average kinetic energy of the particles, which is related to speed.

Pressure is caused by collisions with the walls of the container and is affected by –

- temperature – the higher the temperature the higher the pressure due to increased collisions by the increased speed of the particles.
- volume – the greater the volume the lower the pressure as there will be less collisions.
- number of particles – more particles means more pressure due to more collisions.

WS23 Gas Laws

2010 Q2, 41a,b,c

2011 Q18, 19

2012 Q16, 18, 25, 40a

2013 Q2, 3, 4

- **the mole concept can be used to calculate the mass of substances and volume of gases (at standard temperature and pressure) involved in a chemical reaction**

All gases at the same temperature and pressure have the same number of mole in the same volume. This is Avogadro's hypothesis. This leads to the concept of molar volume which is the same for all gases at a given temperature and pressure.

At standard temperature and pressure (STP = 0 °C and 100 kPa) the molar volume of all gases is 22.71 L. This can be expressed as a formula –

$$n = V/22.71$$

For example –

How many mole of gas is present in 200 L at STP?

$$n = 200/22,71 = 8.807 \text{ mol}$$

What volume of gas is occupied by 0.678 mol?

$$V = 0.678 \times 22.71 = 15.4 \text{ L}$$

Tutorial 9 Volume Calculations

WS24 Volume Calculations

2010 Q 38b

2011 Q36a, 37e,f

Aqueous solutions and acidity

Science as a Human Endeavour

The supply of potable drinking water is an extremely important issue for both Australia and countries in the Asian region. Water sourced from groundwater and seawater undergoes a number of purification and treatment processes (such as desalination, chlorination, fluoridation) before it is delivered into the supply system. Chemists regularly monitor drinking water quality to ensure that it meets the regulations for safe levels of solutes. Heavy metal contamination in ground water is monitored to ensure that concentrations are at acceptable levels. Several methods can be used to reduce heavy metal contamination; the method used is influenced by economic and social factors.

Science Understanding

- **the unique physical properties of water, including melting point, boiling point, density in solid and liquid phases and surface tension, can be explained by its molecular shape and hydrogen bonding between molecules**

  Water is a small bent molecule with hydrogen bonding between the molecules.

  High MP and BP due to strong H-bonds

  Less dense in the solid due to open structure that forms with the H-bonds.

  High surface tension due to strength of H-bonds.

2011 Q39a,b

- **solutions can be classified as saturated, unsaturated or supersaturated; the concentration of a solution is defined as the quantity of solute dissolved in a quantity of solution; this can be represented in a variety of ways, including by the number of moles of the solute per litre of solution (mol L^{-1}) and the mass of the solute per litre of solution (g L^{-1}) or parts per million (ppm)**

  A saturated solution is one that has the maximum amount of solute dissolved for a given temperature.

  An unsaturated solution has less than the maximum amount of solid dissolved.

  A supersaturated is an unstable solution with more than the maximum amount dissolved. This is produced by cooling a hot saturated solution in a smooth clean container.

These can be distinguished from one another by adding a few crystals of solute. These will do nothing in the saturated solution, dissolve in the unsaturated one and cause a lot of solid to come out of the supersaturated one.

Concentration of solutions is commonly given as mole or grams per litre.

% by mass or parts per million are calculated using the mass of solution rather than the volume. Converting between the different concentration types requires the molar mass of the solute and the density of the solution.

$$\text{mol/L} = \frac{\text{number of mole}}{\text{One litre}}$$

$$\text{g/L} = \frac{\text{mass}}{\text{1L}}$$

$$\%\ \text{by mass} = \frac{\text{mass of solute}}{\text{100 g of solution}} \quad \text{or} \quad \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

$$\text{ppm} = \frac{\text{mass of solute}}{\text{1000000 g of solution}} \quad \text{or} \quad \frac{\text{mass of solute}}{\text{mass of solution}} \times 1000000$$

For example - $molL^{-1}$ to % by mass

Mol ------------- x by M----------------------------→ mass of solute x 100

1 litre ----------- x by density----------------------→ mass of solution

If the formula for the type of concentration is written as a ratio, then to convert mole to mass you multiply by molar mass. To convert volume of solution to mass you multiply by density.

Going from ppm or % by mass the reverse is true – convert mass of solute to mole by dividing by molar mass, and mass of solution to volume by dividing by density.

Tutorial 10 Solutions

2010 Q1, 31, 32a

2011 Q6

2012 Q20, 35

2013 Q41d

- **the presence of specific ions in solutions can be identified by observing the colour of the solution, flame tests and observing various chemical reactions, including precipitation and acid-base reactions**

The data sheet gives the key colours of ions in solution. If a colour is not given it generally means that the ion is colourless (and the solid white)

Flame tests are often used when an ion is colourless in solution and generally unreactive. For example sodium gives a yellow flame and potassium a pink/purple flame.

Precipitation reactions are those that produce a solid when ions are combined. The solubility rules are on the data sheet.

For example – silver nitrate + sodium chloride → white precipitate.

From the data sheet silver chloride is insoluble – so NaCl can be used to identify silver as a possible ion.

Acid base reactions usually identify the presence of ions such as carbonate and hydrogen carbonate.

WS25 Distinguishing Substances

2010 Q3, 4, 34

2011 Q7, 41

2012 Q37

2013 Q40

- **the solubility of substances in water, including ionic and polar and non-polar molecular substances, can be explained by the intermolecular forces, including ion-dipole interactions between species in the substances and water molecules, and is affected by changes in temperature**

For a substance to dissolve the following must occur –

- solute-solute bonds must be broken
- solvent-solvent bonds must be broken
- solute-solvent bonds need to form

This last step requires the bonds to be of similar type and strength.

There are also ion-dipole interactions such that the positive and negative ions in an ionic substance are attracted to the negative and positive "ends" of the water molecule, allowing ions to dissolve in water.

Most substances are more soluble at higher temperatures, due to increased energy for breaking solute-solute and solvent-solvent bonds.

For example - explain how water dissolves –

NaCl – ion-dipole forces occur between the positive Na ion and slightly negative oxygen in water, and between the negative Cl ion and slightly positive hydrogens in water.

NH_3 – both water and ammonia have strong hydrogen bonds so these bonds can form between the two different molecules.

For example - explain why iodine is only slightly soluble in water.

Iodine is a large molecule with strong dispersion forces. Water is a small molecule with strong hydrogen bonds and weak dispersion forces. Water will not give up strong H-bonds to form weak dispersion forces with iodine so iodine is only slightly soluble.

- **the Arrhenius model can be used to explain the behaviour of strong and weak acids and bases in aqueous solutions**

Arrhenius model states the following –

- acids ionise to form hydrogen ions in aqueous solution
- bases produce hydroxide ions in aqueous solution
- acid + base → water (+ salt) = neutralisation
- hydrogen ions are very small and positive, so a water molecule forms a dative bond with the ion to stabilise it. Called a hydronium ion.
- Strong acids and bases produce 100% ions in solution – include sulfuric, hydrochloric and nitric acids, and all metal hydroxides.
- Weak acids and bases produce less than 100% ions in solution

$HCl + H_2O \rightarrow Cl^- + H_3O^+$

$NaOH \rightarrow Na^+ + OH^-$

$HCl + NaOH \rightarrow NaCl + H_2O$

2010 Q13, 35a

2011 Q30

2012 Q10, 38

2013 Q5, 7, 41a,b

- **indicator colour and the pH scale are used to classify aqueous solutions as acidic, basic or neutral**

Indicators are chemicals that change colour depending on the acidity/basicity of the solution. The main indicator used in this way is universal indicator which is red – yellow in acid, green when neutral and blue-purple in base.

- **pH is used as a measure of the acidity of solutions and is dependent on the concentration of hydrogen ions in the solution**

  pH = - log [H^+] so is dependent on the concentration of hydrogen ions.

  Water ionises to form both hydrogen and hydroxide ions – so any aqueous solution will have both types of ions present. Hence even bases have a pH due to a small concentration of hydrogen ions.

  The pH scale is from about -1 to 15, where at 25 °C –

  pH = 7 is neutral

  pH < 7 is acidic

  pH > 7 is basic

2010 Q15, 36

2011 Q14, 30d, 40b

2012 Q8, 13

2013 Q18

- **patterns of the reactions of acids and bases, including reactions of acids with bases, metals and carbonates and the reactions of bases with acids and ammonium salts, allow products and observations to be predicted from reactants; ionic equations represent the reacting species and products in these reactions**

  Reactions of acids and bases include –

  - acid + base → salt + water
  - acid + metal → salt + hydrogen gas
  - acid + metal hydroxide/oxide → salt +water
  - acid + metal carbonate/hydrogencarbonate → salt + water + carbon dioxide
  - base + ammonium → salt + water + ammonia
  - base + non-metal oxide → salt + water

  Ionic equations are used to represent solution reactions. It is important to only include reacting species, the rest are eliminated as spectator ions.

  Use Livings Law "only strong acids and soluble ionics in solution are written as ions" to determine how the reaction should be written.

For example – write an ionic equation for the reaction of sulfuric acid and potassium hydroxide.

H_2SO_4	+	KOH	→	K_2SO_4	+ H_2O
Strong acid		ionic solution		soluble ionic	molecular
H^+ SO_4^{2-}		K^+ OH^-	→	K^+ SO_4^{2-}	H_2O
H^+ + OH^-		→ H_2O			spectator ions removed

Observations may be requested –

To write good observations identify the following –

- Colours of solution – this is obtained from the data sheet for colours of ions. If the colour changes this must be mentioned. If the ions are white then the solution is colourless (not clear). E.g. colourless solution turns blue.
- Solids formed – describe the colour of the precipitate e.g. white precipitate, or the colour of a metal that forms by displacement e.g. grey solid.
- Reactant solids – said to dissolve, e.g. white solid dissolves
- Gases formed – describe colour, e.g. colourless odourless gas produced.

For example – $NaOH(s) + H^+(aq) \rightarrow Na^+(aq) + H_2O(l)$

White solid dissolves in a colourless solution.

Tutorial 11 Writing Ionic Equations

WS26 Ionic Equations 1

WS27 Ionic Equations 2

WS28 Ionic Equations 3

WS29 Acid Reactions 1

WS30 Acid Reactions 2

2010 Q33, 35b

2011 Q15, 29b, 38a, 40

2012 Q7, 29a, 34, 39a

2013 Q29a, 35

- **the mole concept can be used to calculate the mass of solute, and solution concentrations and volumes involved in a chemical reaction**

For calculations involving –

- mass – use n = m/M

- solutions – use $n = cV$
- gas volume – use $n = V/22.71$ at STP

A balanced equation is needed for the mole ratio.

For example - 6.00g of magnesium is completely dissolved by 200.0 mL of hydrochloric acid.

- What was the concentration of the hydrochloric acid?
- What mass of magnesium chloride was produced?
- What volume of hydrogen gas was produced at STP?

$2HCl + Mg \rightarrow MgCl_2 + H_2$

n(Mg) = 6.00/24.31 = 0.2468 mol

n(HCl) = 2 x 0.2468 = 0.4936 mol

c(HCl) = n/V = 0.4936/0.2 = 2.47 $molL^{-1}$

$n(MgCl_2)$ = 0.2468 mol

$m(MgCl_2)$ = 0.2468 x 95.21 = 23.5 g

$n(H_2)$ = 0.2468 mol

$V(H_2)$ = 0.2468 x 22.71 = 5.605 L

See Tutorial 10 Solutions

WS31 Equation Calculations 1

WS32 Equation Calculations 2

WS33 Solution Calculations

WS34 Concentration Calculations

2010 Q 32a

2011 Q20, 38c,d,e,

2012 Q29b,c, 39b,c

2013 Q29b, 36, 41c

Rates of chemical reactions

Science as a Human Endeavour

Catalysts are used in many industrial processes in order to increase the rates of reactions that would otherwise be uneconomically slow. Catalysts are also used to reduce the emission of pollutants produced by car engines. Motor vehicles have catalytic converters which are used to catalyse reactions that reduce the amount of carbon monoxide, unburnt petrol and nitrogen oxides that are emitted.

Science Understanding

- **varying the conditions under which chemical reactions occur can affect the rate of the reaction**

  Conditions that can change reaction rate include –

  - temperature – e.g. food cooks faster at a higher temperature and spoils slower at low temperature (such as in the fridge or freezer)
  - concentration – e.g. concentrated acid reacts faster than dilute acid
  - subdivision – twigs burn faster than big logs

- **the rate of chemical reactions can be quantified by measuring the rate of formation of products or the depletion of reactants**

  For example –

  $$Mg\ (s) + 2HCl\ (aq) \rightarrow MgCl_2\ (aq) + H_2\ (g)$$

  Depletion of reactants – time taken to react all of the magnesium
  - time taken for pH to change (to a set value)

  Formation of products - time taken to produce a set amount of hydrogen gas

  Some quantities are easier to measure than others – formation of gases and solids and changes such as pH work well.

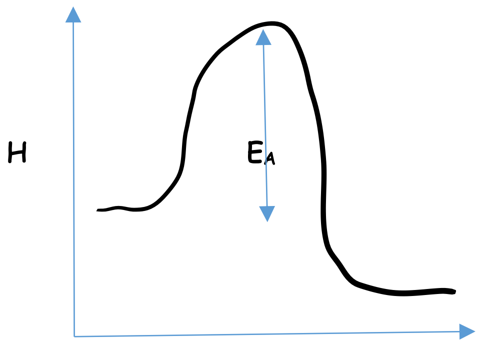
- **collision theory can be used to explain and predict the effects of concentration, temperature, pressure, the presence of catalysts and surface area on the rate of chemical reactions**

  Collision theory states that in order for a reaction to occur there must be a collision between the particles with –

  - sufficient energy (called activation energy)

- suitable orientation

Activation energy (E_A) is like an energy hill that the particles must "overcome" in order to react. The higher the hill the slower the reaction as there will be fewer particles with sufficient energy.



A catalyst is defined as a substance that increases the rate of reaction without being permanently consumed in the reaction.

Catalysts work in two ways –

1. lowering the activation energy by providing a lower energy reaction path
2. adsorbing the gas molecules onto a solid surface providing correct orientation

Temperature affects rate by –

1. speeding up the particles so they collide more frequently (minor effect)
2. increasing the number of particles with activation energy

Pressure/concentration affect rate by –

Increasing the number of collisions.

State of sub-division affects rate by –

Increasing the surface area so there can be more collisions.

We can also consider the nature of reactants – molecular reactions tend to be much slower than ionic/solution reactions as these require strong covalent bonds to be broken.

2010 Q25, 38g,h

2011 Q31

2012 Q15, 40b

2013 Q16, 17

- **the activation energy is the minimum energy required for a chemical reaction to occur and is related to the strength and number of the existing chemical bonds; the magnitude of the activation energy influences the rate of a chemical reaction**

  See previous dot point

- **energy profile diagrams, which can include the transition state and catalysed and uncatalysed pathways, can be used to represent the enthalpy changes and activation energy associated with a chemical reaction**

  See earlier dot point

2010 Q40d,e

2011 Q32c

2012 Q31

2013 Q30

- **catalysts, including enzymes and metal nanoparticles, affect the rate of certain reactions by providing an alternative reaction pathway with a reduced activation energy, hence increasing the proportion of collisions that lead to a chemical change**

  See earlier dot point

2011 Q34d