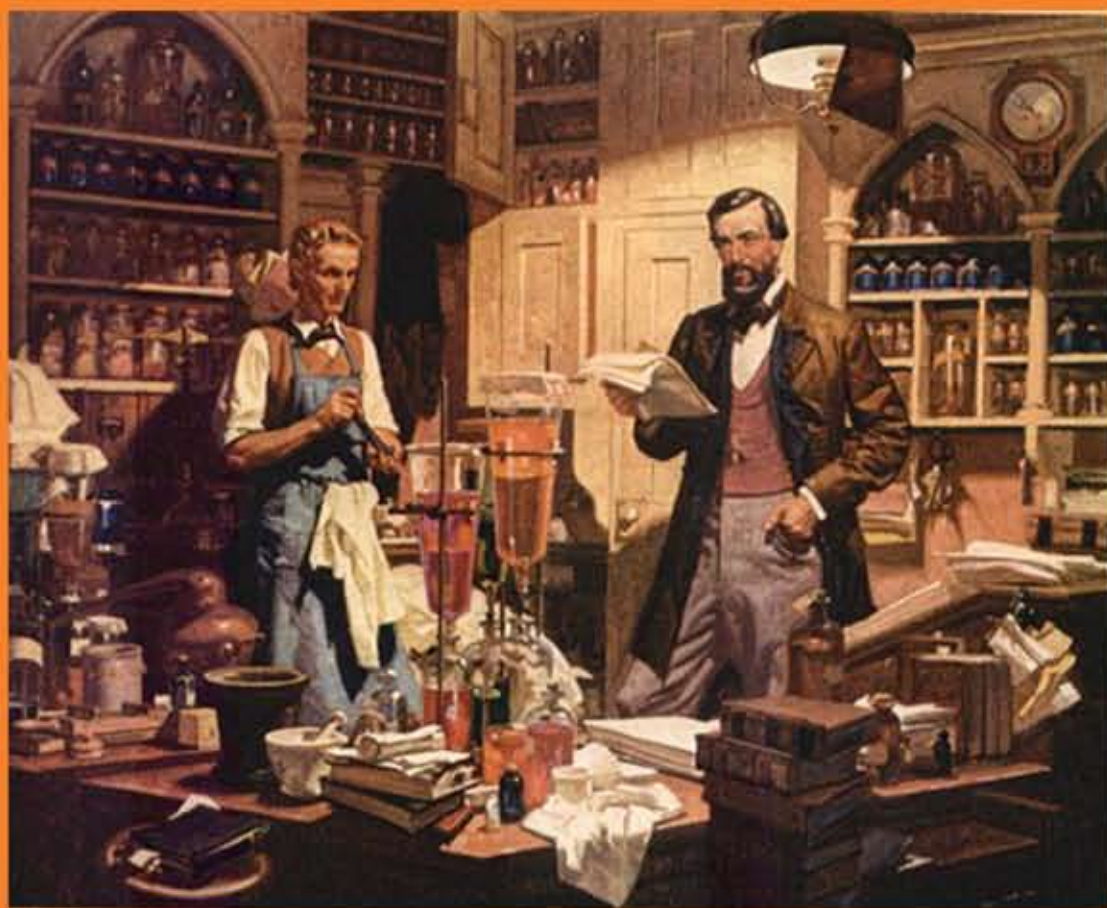# Chemistry Learning Guide

Year 13

Student Learning Guide

Sāmoa Chemistry Curriculum Year 13

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INTRODUCTION

Why study chemistry?

"Knowing" chemistry involves understanding such aspects as the language of chemistry, appropriate strategies to solve problems and the use of models and theories in chemistry. Why could some knowledge of chemistry be useful for you?

- To provide you with a level of chemical understanding that allows you to make well-informed decisions on science-related issues (particularly those involving aspects of chemistry), such as the location of rubbish tips and the effects of pollution or the role of essential foods in a healthy diet and the benefits of herbal remedies.
- To give you practice at critically analyzing problems and carrying out investigations using "scientific" approaches that allow valid comparisons (i.e. fair-testing).
- To prepare you for professional careers that require a knowledge of chemistry in areas such as:
  - Chemistry research and the synthesis of new materials.
  - Medicine and pharmacy.
  - Chemical-based manufacturing and processing industries.
  - Teaching science.
  - Biotechnology and genetics.
- To enable you to make connections between the explanations of chemists about the world and your own ways of thinking.
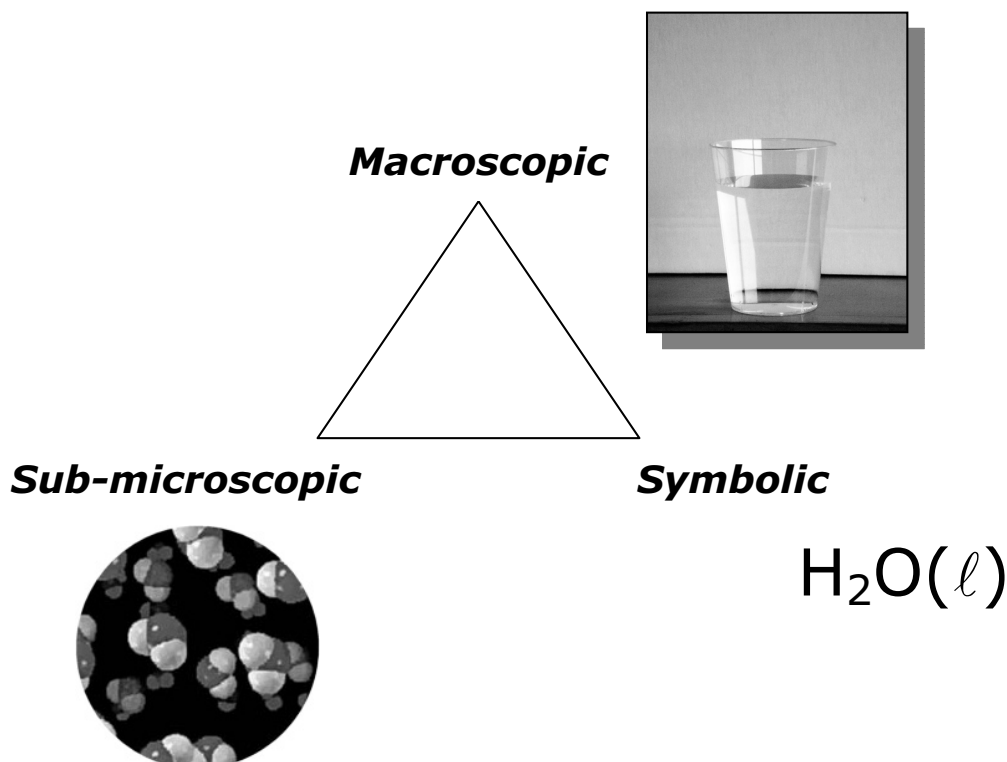
LEARNING CHEMISTRY

Chemistry is something that is carried out by people rather than just being a collection of isolated "facts". Over a period of time, chemists identified patterns and similarities in the observable properties of various substances and agreed upon criteria in which substances are grouped into classes and families based on their properties, such as "elements", "compounds", "metals", and "non-metals". Chemists have also adopted particular language and ways of communicating, such as by symbols and chemical equations. Therefore learning chemistry involves understanding chemists' accepted ideas and ways of communicating.

However, you cannot understand chemistry merely by memorising a lot of information. Instead, you must try to make sense of chemistry and remember it in ways that are meaningful to you. Learning chemistry is therefore not a passive activity in which you can simply absorb all the "facts" told to you by the teacher. Rather, learning chemistry requires you to take an active role by developing a real understanding of the concepts of chemistry. This way, you learn to "think like a chemist".

The key to learning chemistry is to always think about matter in the ways chemists do, in terms of a substance's macroscopic properties (the characteristics that are visible with the naked eye), the mental model of the particles making it up at a sub-microscopic level, and the symbol used to describe it. For example, the three levels of representation used by chemists describe water are shown below.

Macroscopic

Sub-microscopic   *Symbolic*

$H_2O(\ell)$

Chemistry often relies on numbers to convey information about the quantities or the properties of substances involved in chemical reactions. Being able to solve problems involving numbers is an important part of learning chemistry and this requires some competence in maths.

However, it is important to remember that chemistry is much more than just solving mathematical puzzles or "plugging numbers into a calculator". It is vital that whenever you solve a numerical problem in chemistry, you think about what the numbers actually represent on the molecular level in terms of atoms or molecules. This way, you can solve problems using chemical concepts rather than having to memorise lots of methods or steps and it is also consistent with how experts in chemistry solve problems.

There are several things you can do to assist in learning chemistry.
- Note the key concepts of chemistry when the teacher gives an overview of the topic. Understand why a concept is important and how it can be applied.
- Carefully read several times the detailed information contained in the appropriate section of textbook. Make a note of the key words and definitions.
- Work through solving a variety of practice problems as homework to develop and refine your problem-solving skills.
- Ask your teacher to explain concepts that are unclear to you. Study with other people and ask each other questions.

General Problem-solving Hints

There are no magic recipes that work for solving all problems in chemistry but the following general steps may assist you in solving problems.

1. Read the whole problem carefully and identify what the problem is asking. Also identify the units of the answer.
2. Identify what information is given in the problem and what are the important terms. There can also be extra information given that is not needed to solve the problem.
3. Make a list of the "knowns" and "unknowns" to help you see the information you have and the information you want to obtain.
4. Identify steps or paths to get from what is given to what is asked. Look at want you want and find a way to combine what you are given and what you know into the answer of what you want.
5. Always check your answer and method, and evaluate whether your answer is reasonable.

USING THE LEARNING GUIDE

These notes are intended to be an aid to learning the chemistry concepts that have been designated as the theory part of this course. The main points in each section and the page references in textbooks are given.

The objectives set for you in the Year 13 Samoa Chemistry Curriculum are also listed at the start of each section. Note that this Learning Guide is be used in conjunction with any examination prescriptions for chemistry, such as the Pacific Senior Secondary Certificate. *Additional* objectives also required for the PSSC Chemistry examination (1999 prescription) are shown by an asterisk (*).

The following textbooks are referred to in the notes below:
- *Chemistry Year 12 – Pathfinder Series* by M. Croucher and P. Croucher (2000), Auckland: New House Publishers Ltd. (abbreviated as **Chem12**).
- *Chemistry Year 13 – Pathfinder Series* by M. Croucher and P. Croucher (2002), Auckland: New House Publishers Ltd. (abbreviated as **Chem13**).

The aims of this chemistry course are to:

- Promote an awareness of the importance of chemistry in our everyday lives;
- Develop understanding of the important concepts and principles in chemistry and how these principles can be applied to the world around us.
- Demonstrate the use of appropriate scientific methods for carrying out studies and investigations of substances in the environment;
- Foster an interest in chemistry as an area for further learning.

S T R A N D 1: ATOMIC STRUCTURE AND BONDING

ATOMIC STRUCTURE

Students will investigate and develop their scientific understanding of the
structure of atoms *and the format of* ***the periodic table*** *when they:*
- describe the structure of an atom in terms of the nuclear atom model consisting of protons, neutrons and electrons.
- describe isotopes as atoms of an element with differing number of neutrons and mass numbers.
- recall that each isotope has a mass relative to the C-12 mass standard.
- define relative atomic mass (A_r), and relative molecular mass (M_r).
- write the electron configurations for the first 20 elements and their respective ions in terms of the distribution of electrons within orbitals of atoms.
- explain the arrangement of atoms in Periods (rows) according to their principal energy levels and Groups (columns) according to their valence electrons.
- Identify the relationship of the charge on monatomic ions to their positions on the periodic table.

Chemists have studied matter and how it changes for the last several hundred years. Several fundamental observations that they made were that the total mass remains the same during a chemical reaction and that a particular compound always has the same elements present in the same fractions by mass. These observations can be explained by an atomic model. A model is a simplified conceptual picture based on experiments that helps explain how the observed phenomenon occurs. The atomic model states that all matter consists of indivisible, unchangeable atoms of fixed, unique mass.

KEY POINTS:
- All matter is made up of tiny particles called atoms. An atom is the smallest body that retains the unique identity of an element. Atoms of one element cannot be converted into atoms of another element in a chemical reaction.
- In the nuclear atom model, each atom has a central nucleus, which contains positively charged protons and uncharged neutrons and is surrounded by a "cloud" of negatively charged electrons moving at high speed. Because the total mass of the electrons is very small in comparison to that of the protons and neutrons, almost all of the mass of an atom is in the nucleus (due to the protons and neutrons). The volume of the atom is determined by the size of the electron cloud (**Chem12** p4; **Chem13** p4).
- An atom is represented by the notation $_Z^A X$, where Z is the atomic number (number of protons), A is the mass number (number of protons and neutrons) and X is the atomic symbol (**Chem12** p4).

- An atom is electrically neutral because the number of electrons equals the number of protons. Since it is the number of electrons that gives an atom its chemical properties, all atoms with the same atomic number will react in the same way. The atomic number therefore characterises the atoms of a particular element.
- An element occurs naturally as a mixture of isotopes – atoms with the same number of protons but different numbers of neutrons (and therefore different mass numbers) (**Chem12** p4; **Chem13** p4).
- The mass of each isotope is compared to the mass of the ^{12}C isotope, which is defined has having a value of exactly 12.00 atomic mass units.
- The relative atomic mass, A_r, of an element is the average of the masses of its isotopes taking into account their natural abundances. Therefore, the relative atomic mass, A_r, of an element may not be a whole number. The relative molecular mass, M_r, of a compound is the sum of the relative atomic masses in the formula (**Chem12** p36).
- Negatively-charged electrons further away from the positively-charged nucleus have higher amounts of energy. In an atom, groups of electrons occur in volumes of space at certain allowable distances from the nucleus and therefore are found at certain energy levels (**Chem12** p5).
- The maximum number of electrons that can occupy each level varies: 2 electrons in the 1st level; 8 electrons in the 2nd level; 18 electrons in the 3rd level (although only 8 electrons occupy this level before electrons start entering the 4th level) (**Chem13** p8).
- The electrons at a particular energy level occupy atomic orbitals. These are regions of space in which there is a high probability of finding an electron. Each orbital has a particular shape and associated energy value (**Chem13** p8).
- The *electron configuration* of an element shows the distribution of electrons in its atoms (**Chem13** p9).
  - A number indicates the energy level (shell) an electron occupies (smaller numbers indicate lower energy levels and a greater probability of the electron being closer to the nucleus). Electrons occupy the lowest energy level nearest the nucleus first.
  - Each energy level has sublevels (subshells) which indicate the shape of the orbital. Each sublevel has a letter designation: *s, p, d,* f.). The 1st level has only 1 sublevel: 1*s*; the 2nd level has 2 sublevels: 2*s* and 2*p*: the 3rd level has 3 sublevels - 3*s*, 3*p* and 3*d*.
  - Within each level and sub-level, electrons occupy orbitals of specific size (energy), shape and orientation. Each orbital can accommodate a maximum of 2 electrons. The

number of orbitals in any particular sublevel is always the same. *s* sublevels have one orbital and a maximum of 2 electrons; *p* sublevels have 3 orbitals and contain a maximum of 6 electrons; *d* sublevels have 5 orbitals and contain a maximum of 10 electrons. The order of energies of sublevels is 1*s* 2*s* 2*p* 3*s* 3*p* 4*s* 3*d* 4*p*. As the 4*s* orbital is at a lower energy level than the 3*d* sublevel, it is filled before any electrons occupy the 3*d* orbitals.

- The arrangement of elements in the periodic table is related to their electron configuration: elements of a group have similar outer (valence) electron configurations and similar chemical behaviour. Atoms in Groups 1 or 2 have either 1 or 2 valence electrons in an *s* subshell. Atoms in Groups 3 to 12, called the transition metals, have 1 to 10 electrons in the *d* sublevel. Atoms in Groups 13 to 18 have from 1 to 6 valence electrons in a *p* subshell. The period number in the periodic table is the value of the highest energy level (**Chem13** p10).

- Valence electrons are those involved in chemical reactions. For the main-group elements, valence electrons are in the outer level (highest energy level furthest from the nucleus) only. For transition metals, inner *d* electrons are often involved in reactions also. Atoms generally react by losing, gaining or sharing electrons in order to achieve a full energy level (shell) (**Chem13** p12).

- Many main-group elements form electrically charged ions that have the same electron configuration as the nearest noble gas element by gaining or losing electrons in chemical reactions. An ion has completely different properties and behaviour to the original atom.

- Atoms of element in Groups 13 – 17 gain electrons in the outer *p* subshell and form anions (negatively charged ions). Atoms of elements in Groups 1 and 2 lose valence electrons from the sublevel furthest from the nucleus (outer s orbital) and form either M^+ or M^{2+} cations (positively charged ions). For atoms of the transition metals in Period 3, both 3*d* and 4*s* electrons can be removed to form transition metal cations and ions and commonly form more than one ion. Although the 4*s* sublevel is filled before the 3*d*, the transition metals lose the 4*s* electrons **before** losing electrons from the 3*d* sublevel when forming a cation.

TRENDS ACROSS THE PERIODIC TABLE

*Students will investigate and develop their scientific understanding of **periodic trends** when they:*
- explain the meaning of the following terms:
   (i).   first ionization energy
  (ii).   atomic radius
  (iii).   ionic radius
  (iv).   electronegativity
- describe the trends of (i) – (iv) above, across Periods 2 and 3 and within Groups.
- explain how atomic properties are related to the tendency to form ions.
- explain how the change in bonding from ionic to polar covalent to nonpolar covalent across a period left to right is related to differences in electronegativities of bonded atoms.

Chemists discovered that the elements show a periodic recurrence of similar properties, forming the basis for the periodic table. In understanding the chemistry of elements, all physical and chemical behaviour of the elements can be based ultimately on the electron configurations of their atoms, where recurring electron configurations lead to trends in atomic properties, which in turn lead to trends in chemical behaviour.

KEY POINTS:
- Explanations for trends in atomic properties need to take into account the relative size of the electrostatic attraction between the protons in the nucleus and the valence electrons. This electrostatic attraction depends on:
  - the number of protons in the nucleus (electrostatic attraction increases with the size of the effective nuclear charge);
  - the distance of the valence electrons from the nucleus (electrostatic attraction decreases as the distance between the positive and negative charges increases);
  - the shielding of the inner shells of electrons (increased shielding decreases the electrostatic attraction between the protons and valence electrons) (**Chem12** p27; **Chem13** p23).
- As a general rule, when going across a period from left to right the effect of the increased nuclear charge is the most important factor whereas when going down a group the increased number of electron shells (increased radius) is more important (**Chem12** p20,21; **Chem13** p23).
- The first ionisation energy is the energy required to remove the outermost electron from one mole of gaseous atoms (or ions). It is inversely related to atomic size and decreases down a main group (as electrons are being added to energy levels at higher energies further from the nucleus with additional shielding from inner shells of electrons) and increases across a period from left to right (as electrons are added into the same energy level with

no additional shielding). Successive ionisation energies show an exceptionally large increase when inner (core) electrons start to be removed. Ionisation energies are all positive values as energy must be added to the system to overcome the electrostatic attraction between the valence electron and the nucleus (**Chem13** p24).

- As a result of the relative sizes of ionisation energies, in their ionic compounds, Group 1 and 2 elements tend to form cations and Group 16 and 17 elements tend to form anions. Also the decrease in ionisation energy going down a Group can explain the increase in reactivity of Group 1 and 2 elements going down the Group.

- The atomic radius of a metal is half the distance between the nuclei in two adjacent atoms. For elements that exist as diatomic molecules, the atomic radius is half the distance between the nuclei of the two atoms in a molecule.

- Atomic size increases down a main group (as valence electrons occupy orbitals at higher energy levels further from the nucleus) and decreases across a period left to right (as electrons are added to the same energy level with an increase in the nuclear charge, the electrons are attracted more closely to the nucleus). Across a transition metal series, size remains relatively constant (**Chem12** p21, 29).

- Cations are smaller than their parent atoms as removing one or more electrons reduces electron-electron repulsion and the same nuclear charge attracts a smaller number of electrons closer to the nucleus. Also, the formation of cations often results in the loss of all the electrons in the outer valence shell. Anions are larger than their parent atoms since the nuclear charge remains the same but the extra repulsion resulting from the additional electron increases the size of the electron cloud. Ionic radius increases down a group.

- Across a period from left to right, cation and anion radii decrease, but a large increase occurs from cation to anion (**Chem12** p21).

- Electronegativity is the relative ability of a bonded atom to attract towards itself electrons shared in a chemical bond. Electronegativity is based on a scale assigning the highest value of 4.0 to fluorine and increases across a period left to right and decreases down a group. This trend is the reverse to that in atomic size. The elements of Group 18 are not given electronegativity values as they do not tend to form bonds with other elements (**Chem12** p8; **Chem13** p25).

CHEMICAL BONDING

Compounds consist of atoms or ions held together by chemical bonds (ionic or covalent). Chemical bonding results in the lowering of energy of atoms and ions in their compounds and the formation of a bond releases energy. Atoms may gain, lose or share electrons in bond formation. The type and strength of attractive forces both within and between particles of a substance determine the properties of the substance. When the particles are packed in an orderly array the solid is crystalline and is classified according to the nature of the attractive forces that hold the atoms, ions or molecules in place. There are four classes of solids: metals (metallic solids), ionic solids, covalent network solids, molecular solids.

IONIC BONDING

*Students will investigate and develop their scientific understanding of **ionic bonding** when they:*
- describe ionic bonding as resulting from a metal atom transferring electrons to a non-metal atom and the resulting ions become attracted to each other to form an ionic solid.
- describe the properties of ionic substances (electrical conductivity, melting point, rigidity, brittleness) and relate the physical properties of ionic substances to their structures and bonding.

KEY POINTS:
- Ionic compounds exist as 3-dimensional crystal lattices with anions surrounded by cations and *vice versa*. The formula of an ionic compound shows only the ratio of positive and negative ions in the crystal lattice.
- Although ion formation requires energy, much more energy is released when ions come together to form a crystal lattice.
- The strong electrostatic attractions between the oppositely charged positive and negative ions in the lattice are referred to as ionic bonds (**Chem12** p8; **Chem13** p12, 20, 21).
- As much energy is needed to separate the ions from each other, ionic substances have high melting points and boiling points and are solid at room temperature.
- Ionic solids are hard but brittle as any distortion to the crystal results in ions of like charge being forced closer together which results in strong repulsive forces that push the crystal apart.
- Ionic compounds only conduct electricity when melted or dissolved in water since the ions are immobile in a solid (**Chem12** p16).

COVALENT BONDING

*Students will investigate and develop their scientific understanding of **covalent bonding** when they:*
- describe a covalent bond as a shared pair of valence electrons attracted by the nuclei of two non-metal atoms.
- determine the bond polarity of covalent bonds (polar or non-polar) and the direction of polarity arising from differences in electronegativities of bonded atoms for polar bonds.
- identify the continuum in bond character from non-polar through polar to ionic bonding arising from the distribution of valence electrons in the molecule.
- describe the polarity of the water molecule and the process of hydration of ionic substances.
- describe the properties of discrete molecular substances and infinitely extended covalent networks and relate these physical properties to their structures and bonding (electrical conductivity, melting point, boiling point, solubility, hardness).
- compare the strength of inter-molecular and intra-molecular forces and the role played by the hydrogen bond.

KEY POINTS:
- A covalent bond is a pair of valence electrons shared between the nuclei of two adjacent atoms (**Chem12** p7; **Chem13** p12).
- Molecules consist of discrete groups of covalently bonded atoms.
- The shared electrons in the covalent bond are called bonding electrons, and may or may not be shared equally, depending on the relative electronegativities of the two atoms involved.
- When two non-metal atoms having the same (or very similar) electronegativity combine, the bonding electrons are shared equally between the two nuclei and a non-polar covalent bond occurs. The bond in H_2 is a non-polar covalent bond as the atoms are identical.
- When two atoms with different electronegativities form a covalent bond, the bonding electron pair is shared unequally and attracted more towards the more electronegative atom (which has a greater electron attracting power). This results in a polar covalent bond, with partial negative and positive charges on the atoms in the bond. The atom which has the greater electronegativity will form the partially negative end of the polar bond (**Chem12** p7; **Chem13** p13).
- The greater the difference in electronegativity values for the two atoms in the bond, the more polar the bond is and greater its ionic character.
- A continuum from non-polar covalent bonds to polar bonds to ionic bonds occurs as the ionic character of the bond increases. The dividing line between a strongly polar bond and an ionic bond is arbitrary and based on observable physical properties such as melting point and conductivity. When the difference in electronegativity values is greater than 1.7, the bond can be

considered as mostly ionic. The majority of compounds can be thought of having bonds that are partially ionic and partially covalent.

- A gradation of bond type occurs across a period from left to right, from ionic to polar covalent to non-polar covalent, as shown by the compounds formed when a halogen is combined with a Period 3 element.

- The arrangement of polar bonds in a molecule results in one side of the molecule being more positively-charged than the other side and the molecule is said to have a dipole moment (**Chem12** p11).

- In water, H_2O, the nucleus of the O atom attracts the shared bonding electrons more strongly than the nuclei of the H atoms, creating an uneven charge distribution in each O-H bond. The combined effects of its polar bonds and the bent shape of the water molecule make water a polar molecule with a partially negative O end and a partially positive end located between the two H atoms.

- When an ionic compound dissolves in water, the H_2O molecules separate, surround, and disperse the ions into the liquid, forming a solution containing dissolved ions that are kept apart from each other by the attractions to the water molecules between them. The negative ends of the H_2O molecules are attracted to the positive ions, while the positive ends of the H_2O molecules are attracted to the negative ions. For many ionic compounds, these attractions are strong enough to overcome the strong ionic bonds holding the ionic crystal together and it will dissolve. Many other ionic compounds do not dissolve in water as the electrostatic attraction between ions in the compound is greater than the attraction between ions and water molecules (**Chem12** p15).

- In a molecular substance, there are strong covalent bonds between atoms *within* the molecules (intra-molecular forces) but only very weak attractive forces *between* individual molecules (inter-molecular forces) in the solid and liquid states. These weak intermolecular attractive forces, called van der Waals forces, are electrostatic attractions between the slightly positive end of one molecule and the slightly negative end of another. Not much energy is needed to overcome these weak attractions and to separate the molecules from each other. The temperature at which a substance forms a solid or a liquid depends on the strength of the attractive forces between the particles. Therefore, molecular compounds have low melting points and boiling points and often exist as soft solids, liquids or gases at room temperature (**Chem12** p15, 16; **Chem13** p19).

- There are three types of van der Waals intermolecular forces – dipole-dipole attractions, dispersion forces (also called London forces or attractions between instantaneous dipoles) and

hydrogen bonding. The total intermolecular force acting between two molecules is the sum of all the separate types of forces. Ionic and covalent bonds are much stronger bonds than intermolecular forces (**Chem13** p19).

- The melting point and boiling point of a particular substance depend on the mass of the molecules (molecules with higher masses have larger clouds of electrons that result in stronger instantaneous dipole-dipole attractions and a higher melting point and boiling point) and the polarity of the molecule (the more polar the molecule, the stronger intermolecular attractive forces are and higher the melting point and boiling point) (**Chem12** p16).

- Molecular substances do not conduct electricity as molecules do not have an electrical charge.

- Polar molecular substances (those with a dipole moment) dissolve in polar solvents whereas non-polar substances dissolve in non-polar solvents. This solubility behaviour is referred to as "like dissolves like".

- Covalent network solids are made up of linear chains, 2-dimensional layers or 3-dimensional networks which are not easily broken as strong intramolecular covalent bonding link atoms.
    - Linear chains (as in rubber, plastic and some allotropes of sulfur) have weak attractions between the chains, so that the substances are flexible and have low to medium melting points.
    - Graphite is an allotrope of carbon made up of 2-dimensional layers. Each layer is formed by C atoms linked by covalent bonds, while layers are held together by weak attractive forces. The weak attractions allow the layers to easily slide over each other and graphite is used as the "lead" in pencils. Graphite is the only non-metal to conduct electricity in the solid. Graphite has a high melting point and is insoluble in polar solvents such as water (**Chem12** p17; **Chem13** p21).
    - Carbon (as diamond), silicon and silica (silicon dioxide, SiO_2) exist as 3-dimensional covalent network solids. All atoms in the network structure are linked by covalent bonds so that these substances are all rigid, strong and have very high melting points (as the covalent bonds have to be broken before the substance can melt). They do not conduct heat or electricity either in the solid or liquid states as they have no charged particles that are free to move. They are very insoluble (**Chem12** p17, 18; **Chem13** p21).

- Water has an unusually high boiling point and melting point for its molar mass and is less dense in the solid form (ice) than in the liquid form (ice floats on water). These properties are due

to the intermolecular attractions between water molecules known as *hydrogen bonding*. This is a specific case of permanent dipole-dipole attraction occurs between molecules having a hydrogen atom bonded to a small, highly electronegative atom that has non-bonding pairs of electrons - O, N and F. The small size of these atoms also allows the molecules to get much closer together resulting in a stronger attractive force (**Chem13** p20).

METALLIC BONDING

*Students will investigate and develop their scientific understanding of **metallic bonding** when they:*
- describe the electron-sea model of metallic bonding involving de-localised valence electrons of metal atoms.
- describe the physical properties of metallic substances and relate these properties to their structures and bonding (electrical conductivity, malleability, ductility).

KEY POINTS:
- A metal consists of metal atoms stacked in layers and held together by strong metallic bonds. Metal atoms have loosely held valence electrons which readily move between atoms. The metallic bond is the electrostatic attraction between a "sea" of mobile valence electrons and the positively charged metal ion (**Chem12** p16; **Chem13** p21).
- Metals have moderate melting points as the attractions between metal ions and the electron sea are not disrupted on melting.
- Metals have high boiling points as boiling requires separating the metal atoms.
- Metals can be deformed because the metal ions can move past each other through the surrounding sea of electrons (which prevents repulsions among the cations). This means that metals are malleable (can be flattened into sheets) and ductile (drawn into wires).
- Metals conduct electricity and heat because their electrons are mobile. The mobility of the electrons also accounts for the lustre of metals (metals are shiny when the oxide or other coating is removed).

SHAPES OF MOLECULES

*Students will investigate and develop their scientific understanding of **shapes of molecules** when they:*
- state the Octet rule that applies to small atoms and use it to draw Lewis structures for simple covalent molecules which have only four valence shell electron pairs.
- deduce the shape of simple covalent molecules using the valence shell electron pair repulsion (VSEPR) theory.
- name and draw molecules with the following shapes:
  (i). linear e.g. CO_2
  (ii). bent (v-shaped) e.g. H_2O
  (iii). trigonal (triangular) pyramid e.g. NH_3
  (iv). tetrahedral e.g. CH_4
- use molecular shape and bond polarity to determine whether a molecule is polar (and possesses a dipole moment) or is non-polar.

KEY POINTS:
- A Lewis diagram (or Lewis structure) is a representation of a molecule (or ion) that shows how atoms are linked in a molecule and the distribution of all valence electron as bonding and lone pairs (**Chem12** p7; **Chem13** p13).
- Lewis diagrams for molecules containing atoms from the first two periods of the periodic table have 8 (4 pairs; an octet) of electrons are located around each atom (except for hydrogen which only has 2 electrons).
- For some molecules, the sharing of one pair of electrons will not result in 8 electrons around both bonded atoms and therefore two pairs of electrons (4 electrons) are shared between two adjacent atoms resulting in a double covalent bond. Similarly, three pairs of electrons shared between two adjacent atoms results in a triple covalent bond.
- Non-bonding valence electrons are shown on atoms in Lewis diagrams as lone pairs of electrons.
- A stepwise process is used to convert a molecular formula into a Lewis diagram. The first step in drawing a Lewis diagram is to link atoms so that they are all connected by a single covalent bond in the appropriate arrangement. If there are not enough electrons to complete the octet for each atom (except H) then non-bonding pairs need to be shifted to form a double covalent bond (or triple bond).
- The VSEPR (valence-shell electron-pair repulsion) model is used to predict the three-dimensional shape of a molecule. It states that each group of electrons (i.e. single bond, multiple bond, or non-bonding pair of electrons) around a central atom remains as far away from the others as possible to reduce repulsions. Therefore, the shape a molecule adopts is determined by groups of electron pairs that mutually repel each other to maximise the angles between them (**Chem12** p12, 13; (**Chem13** p13).

- The shape of the molecule depends on *both* the relative positions of the atoms and the number of regions of electron density (bonded and non-bonded electron pairs) around the central atom. Different molecular shapes occur depending on whether an electron pair is bonding or non-bonding (**Chem13** p14).
- The bond angle is the angle between two bonds to the same atom.
- A *linear* molecular shape occurs when two electron groups are attached to the central atom and are orientated as far apart as possible in opposite directions (giving a bond angle of 180°). Examples include HCN and commonly found with molecules having two double bonds such as CO_2.
- Three atoms linked to a central atom repel each other into the corners of a equilateral triangle to give a molecule with a trigonal planar shape. Commonly found with carbon linked by a double covalent bond to one atom, and 2 single covalent bonds to two other atoms, such as H_2CO.
- All molecules or ions with four electron groups around a central atom adopt a tetrahedral electron-group arrangement (**Chem12** p12).
- When all four electron groups are bonding groups to other atoms (4 atoms linked to a central atom by 4 single covalent bonds and no non-bonding pairs on the central atom), the molecular shape is *tetrahedral* and the bond angles are all 109.5°. Examples include CH_4 and $SiCl_4$.
- When one of the four electron groups in the tetrahedral arrangement is a non-bonding pair of electrons (the central atom is linked to only three other atoms and also has a non-bonding pair of electrons), the molecular shape is a *trigonal pyramid* (a tetrahedral with one vertex "missing"). The bond angle between the atoms is 107.3° as the bonding pairs are forced closer by the non-bonding electron pair. Common for compounds of N and P such as NH_3 and PF_3.
- When the four electron groups around a central atom include two bonding groups and two non-bonding groups, the molecular shape is *bent* (or *V-shaped*) and the bond angle is 104.5°. Commonly found in molecules of O and S such as H_2S. Water is the most important bent-shaped molecule.
- Other molecular shapes can occur, particularly when the central atom has more than four electron groups around it.
- In molecules with more than two atoms, *both* shape and bond polarity determine molecular polarity, which is measured as a dipole moment (and indicates the magnitude of the partial charges on the ends of a molecule).
- *Polar* molecules have at least one polar bond (a covalent bond between two atoms of different electronegativity) and an asymmetric (not symmetrical) arrangement of polar bonds.

Since the bond polarities are not counterbalanced, the molecule has an overall dipole moment. Examples of polar molecules include H_2O, NH_3, HCN (as the two bonds are of different polarity so do not cancel out) and $CHCl_3$ (as tetrahedral molecule does not have a symmetric arrangement of polar bonds that counterbalance each other) (**Chem12** p13).

- Polar substances include water, ethanol, methanol, ethanoic acid, ammonia, and hydrogen chloride.
- *Nonpolar* molecules occur when molecules have no polar bonds such as O_2, or when polar bonds which are arranged symmetrically so that bond polarities counterbalance each other (as in CF_4 which has four polar C-F bonds in a symmetric tetrahedral arrangement).
- Non-polar substances include cyclohexane, hydrocarbons (e.g. petrol, wax, turps), and tetrachloromethane, CCl_4.
- Molecular shape and polarity can affect physical properties, such as boiling point where more energy is needed to overcome the weak attractive forces between polar molecules than with non-polar molecules of the same size (**Chem13** p19).

S T R A N D 2: QUANTITATIVE CHEMISTRY

THE MOLE

The concepts of quantitative chemistry enable a chemist to relate the observable amount of a substance to the number of atoms, molecules or ions that make it up.

*Students will investigate and develop their scientific understanding of **the mole** and **molar mass** when they:*
  * define amount of substance in moles and molar mass.
  * recall that the molar mass of an element is numerically the same as its relative atomic mass.
  * carry out simple calculations involving mass in grams, molar mass in g mol^{-1}, and amount in moles.
  * define Avogadro's number*

KEY POINTS:
  * A mole is the SI unit for the amount of a substance that contains a number of objects (atoms, molecules, ions) equal to the number of atoms in exactly 12 g of ^{12}C (**Chem12** p35).
  * The number of atoms in exactly 12 g of ^{12}C has been determined to be 6.02×10^{23} and is called Avogadro's number.
  * The molar mass is the mass in grams of 1 mole of objects (atoms, molecules, ions) of a substance and has units of g mol^{-1} (**Chem12** p35).
  * The mass in grams of 1 mole of atoms of an element has the same *numerical* value as the atomic mass of the element.
  * The molar mass of a compound is the sum of the atomic masses of the atoms in the formula.
  * By using the molar mass in g mol^{-1} of an element (or compound), amount in moles can be converted in mass in grams and *vice versa* (**Chem12** p36, 39).

EMPIRICAL AND MOLECULAR FORMULAE

*Students will investigate and develop their scientific understanding of **empirical** and **molecular formulae** when they:*
  * calculate
    (i). percentage composition; e.g. water of crystallisation, elements in a compound.
    (ii). empirical formula from mass percentage (and *vice versa*).
    (iii). molecular formula from empirical formula and molar mass (and *vice versa*).

KEY POINTS:
  * The percentage composition of a compound is determined by dividing the mass of each element by the total mass of the compound and expressing as a percentage (**Chem12** p41; **Chem13** p46, 47).

- The masses of elements in an unknown compound can be determined experimentally, giving the relative amounts in moles and the empirical formula.
- The empirical formula is the chemical formula that shows the lowest relative number of atoms of each element in a compound (**Chem12** p41).
- The molecular formula shows the actual number of atoms of each element in a molecule and can be determined if the molar mass and empirical formula are known (**Chem12** p42).

STOICHIOMETRY IN CHEMICAL REACTIONS

*Students will investigate and develop their scientific understanding of **stoichiometry in chemical reactions** when they:*
- write balanced chemical equations given the names and/or formulae of the reactants and products.
- recognise the symbols used to describe states of substances in chemical equations [(*g*) gas, (*l*) liquid, (*s*) solid, (*aq*) aqueous solution].
- perform stoichiometric calculations based on mass, amount in moles, concentration and volume.

KEY POINTS:
- A chemical change can be represented by a chemical equation.
- A chemical equation shows the formulae of reactants on the left of an arrow and products on the right and must be balanced in terms of the number and type of each atom (**Chem12** p38).
- Coefficients are numbers appearing in front of formulae and multiply all the atoms in the formula that follows it. Coefficients apply to the individual particle or to amount in moles of the particle.
- The amount in moles of one substance is related to the amount in moles of any other substance by the stoichiometrically equivalent molar ratios shown in the balanced chemical equation (**Chem12** p39).
- Abbreviations are used in chemical equations to indicate the physical state of each substance or whether it is dissolved in water These are solid (*s*), liquid (*l*), gas (*g*), and aqueous solution (*aq*) when dissolved in water.
- The amount in moles of a solute that is dissolved in a certain volume of solvent can be determined from the concentration of the solution (**Chem12** p61, **Chem13** p42).

ACID-BASE TITRATIONS

*Students will investigate and develop their scientific understanding of **acid-base titrations** when they:*
- define standard solution, titre, aliquot, equivalence point, end point.
- express concentration in terms of g L^{-1} and mol L^{-1}.
- calculate the:
  - (i). concentration to specified dilutions.
  - (ii). mass required to make a standard solution given the volume of a volumetric flask and the molar mass of the standard.
  - (iii). unknown concentration based on titration data.
- describe all the equipment and methodology for preparing a standard solution and carrying out a dilution and a titration.
- describe the colour changes of the common indicators methyl orange and phenolphthalein when the end point is reached.

KEY POINTS:
- The concentration of a solution is the amount of solute dissolved in a given volume of solution. Concentration is usually expressed as amount in moles present in a given volume of solution has units of mol L^{-1} but can also be expressed as the mass in grams of substance completely dissolved in a given volume having units of g L^{-1} (**Chem12** p62; **Chem13** p42, 42).
- When a solution is diluted, only solvent is added so that the solute is dispersed in a larger volume. The final solution contains fewer particles in a given volume and has a lower concentration.
- Titration is an experimental method of determining the concentration of a solution by monitoring its reaction with a solution of known concentration (**Chem12** p62, 63; **Chem13** p105).
- A standard solution is a solution of accurately known concentration prepared by precisely weighing out a known mass of pure solid and dissolving this in a known volume of deionised water in a volumetric flask before mixing thoroughly. Substances suitable as primary standards must be pure solids that do not absorb water and are stable in solution (**Chem12** p61; **Chem13** p107, 108).
- The neutralisation reaction between an acid and base is used in titration to determine the concentration of one of the reactants when the concentration of the other is known accurately.
- In an acid-base titration, a volume (referred to as an *aliquot*) of an acid of unknown concentration is measured using a pipette and placed in a flask beneath a burette that contains a base on known concentration.
- An indicator that changes colour once the reaction is complete is also added to the flask. The base is added slowly to the acid until the indicator permanently changes colour (referred to as the **end-point**).

- The volume of base required to completely react with the acid is recorded. The titration is repeated several times until the burette volumes (referred to as *titres*) agree within acceptable limits.
- The point when the amount of added base is exactly equal to the initial amount of acid in the solution is known as the **equivalence point** of the titration (and is the same as the end-point if an appropriate indicator is used).
- The pH of the solution at the equivalence point depends on the nature of the acid and the base that were titrated (**Chem12** p60).
- The titre volume is used to calculate the concentration of the unknown acid or base (**Chem12** p63).
- Common indicators that show whether a solution is acidic or basic include litmus (red in acidic solutions; blue in basic solutions), methyl orange (red in acidic solutions; yellow in basic solutions) and phenolphthalein (colourless in acidic solutions; pink-red in basic solutions).

S T R A N D 3: INORGANIC CHEMISTRY

Inorganic chemistry is an area of chemistry that focuses mainly on the compounds formed by elements other than carbon. Based on their properties, elements can be classified as metals, non-metals or metalloids. Metals are found on the left-hand side of the periodic table while non-metals are found on the upper right-hand side of the periodic table.

THE PROPERTIES OF SELECTED PERIOD 3 COMPOUNDS

*Students will investigate and develop their scientific understanding of **oxides** (illustrated by **Na$_2$O, MgO, Al$_2$O$_3$, SiO$_2$, SO$_3$**) when they:*
- deduce the formulae of the oxides of sodium, magnesium and aluminium and recall the formulae for the oxides of silicon and sulfur.
- explain whether an element of Period 3 forms an ionic or covalent oxide based on its position in the periodic table.
- describe the trend of oxides in water from basic oxides to acidic oxides as the bond between the Period 3 element and oxygen becomes more covalent.
- write balanced equations which show the acidic, amphoteric or basic property of an oxide.

KEY POINTS:
- Metallic behaviour decreases from left to right across a period and increases down a group in the periodic table (**Chem12** p20).
- Metals tend to lose electrons and exist as positively-charged ions in their compounds.
- The acid-base behaviour of metal oxides and non-metal oxides is different in water.
- As elements become less metallic across a period from left to right, their oxides become more acidic (**Chem12** p24, 25).
- Most main-group metals form ionic oxides, such as Na$_2$O and MgO. Metal oxides act as bases because if they are insoluble in water, they react with acids to produce a solution containing a metal salt. Metal oxides that are soluble in water produce OH$^-$ ions when dissolved.
- Some metal oxides are amphoteric (react with **both** acid and base), such as Al$_2$O$_3$ which is insoluble in water but will dissolve in both acid and base.
- Non-metals form covalent oxides that are either molecular substances, such as SO$_3$, or network solids, such as SiO$_2$. Non-metal oxides (oxides of elements in Groups 14 - 17) act as acids as they react with bases. In water, many non-metal oxides react to form acids that dissociate in water producing H$_3$O$^+$ ions.

*Students will investigate and develop their scientific understanding of **chlorides** (illustrated by NaCl, MgCl$_2$, AlCl$_3$, PCl$_3$, HCl)* when they:
- write the formulae for the chlorides of elements from Na - P in Period 3.
- relate the melting points and the electrical conductivities of Period 3 chlorides to their structure and bonding.
- write balanced equations which demonstrate the reaction of Period 3 chlorides with water.

KEY POINTS:
- A change in bond type from ionic through polar covalent to non-polar covalent is observed when Cl$_2$, a non-metal, is combined with each of the other elements in Period 3.
- Trends in the properties of the compounds formed from each Period 3 element with chlorine are observed across the period (**Chem12** p25, 26).
- As the difference in electronegativity between the element and chlorine becomes smaller in the Period 3 chlorides, the bond becomes more covalent and the properties of the chloride compounds change from a solid consisting of ions to those of a gas consisting of individual molecules.
- The melting point of Period 3 chlorides (as well as the electrical conductivity at melting point) reduces from left to right across the period as the structure changes from ionic through polar covalent to non-polar molecules. Molecular Period 3 chlorides have strong covalent bonding between atoms in the molecules but weak attractive forces between the molecules.
- The properties of aluminium chloride, AlCl$_3$, indicate a highly polar covalent bond between Al and Cl atoms rather than ions, resulting in a solid consisting of extended layers of linked Al and Cl atoms.
- Only the ionic Period 3 chlorides dissolve in water to produce a solution containing metal cations and chloride ions. Molecular Period 3 chlorides do not dissolve in water but instead react with water (in a hydrolysis reaction) to produce an acidic solution containing H$_3$O$^+$ ions.

TESTING FOR UNKNOWN IONIC SPECIES

Qualitative analysis of solutions containing mixtures of ions involves adding ions that result in precipitation and complex ion formation as ways to separate and identify the unknown ions that are present.

Students will investigate and develop their scientific understanding of **testing for unknown ionic species** *when they:*
- recall the solubilities of nitrates, sulfates, chlorides, carbonates, hydroxides, salts of Group 1 metals and ammonium salts.
- apply knowledge of solubilities and complex ions to carry out tests for the presence of Mg^{2+}, Ag^+, Fe^{2+}, Fe^{3+}, Cu^{2+}, Al^{3+}, CO_3^{2-}, Cl^-, SO_4^{2-} and state observations for positive tests.
- write balanced equations for the reactions for positive tests for the ions listed above.
- recognise the formulae and the appearance in solution of the following complex ions: $[Cu(NH_3)_4]^{2+}$, $[Ag(NH_3)_2]^+$, $[FeSCN]^{2+}$, $[Al(OH)_4]^-$.

KEY POINTS:
- Many ionic compounds dissolve in water to produce a clear solution containing ions.
- Ionic compounds which do not readily dissolve in water are referred to as insoluble. A precipitate forms when two solutions containing soluble ionic compounds are mixed and react to from an insoluble product that causes the mixture to become cloudy.
- To predict whether a precipitate will form, note the ions present in the reactants and consider the possible cation-anion combinations. Consult the list of solubility rules to decide whether any combination of ions gives an insoluble product.
- A set of solubility rules for ionic compounds in water has been generated from many observations. The formation of a precipitate and its colour may assist with identifying the presence of a certain ion (**Chem12** p31-33; **Chem13** p39, 40).
- Net ionic equations represent precipitation reactions and show the precipitate formed as the product and the two ions that it is made from as reactants. Spectator ions, which are not involved in any reaction, are excluded from the net ionic equation.
- Precipitates of amphoteric metal hydroxides dissolve on addition of an acid or a base due to acid-base reactions that form soluble complex ions.
- $Fe^{3+}(aq)$ ions react with thiocyanate ions, $SCN^-(aq)$, to form a dark blood-red solution containing the complex ion $FeSCN^{2+}(aq)$.

THE CHEMISTRY OF WATER

Water possesses unique physical and chemical properties that arise from the nature of the H and O atoms that make up the water molecule. Because the water molecule has two O-H bonds (with a large electronegativity difference between the two atoms) and two non-bonding electron pairs on the O atom, the water molecule is bent and highly polar. The properties of water are vital for life on Earth.

Students will investigate and develop their scientific understanding of the **properties of water** *when they:*
- describe the following properties of water: boiling point and melting point, heat capacity, electrical conductivity, density of ice compared with liquid water, solvent properties.
- use anhydrous copper sulfate and cobalt chloride paper to test for water and state the colour changes which occur.

KEY POINTS:
- Hydrogen bonding, the moderately strong attraction between the hydrogen atom in one water molecule and the oxygen atom in another water molecule, results in many unusual properties for water. A water molecule can engage in as many as four hydrogen bonds to surrounding water molecules (**Chem13** p20).
- Hydrogen bonding between water molecules means that extra energy is required to first break the hydrogen bonds before the molecules can separate and become a gas, compared to other small molecules. This results in an unexpectedly high boiling point for water.
- Hydrogen bonding also occurs between water molecules in the solid state (ice) which must be broken before ice completely melts to a liquid.
- Water has great solvent power as a result of its polarity and the ability to from hydrogen bonds.
- Water dissolves many ionic compounds as the polar molecules of water are attracted to the ions. When an ionic compound dissolves in water, the ions dissociate from each other and become surrounded by water molecules (are hydrated). This is accompanied by the release of energy. Ionic solutions conduct electricity as the hydrated ions are free to move towards an electrode (**Chem12** p50; **Chem13** p66).
- Water can also dissolve many covalent substances that also contain O-H bonds, such as sugar, by forming hydrogen bonds to them.
- Water dissolves non-polar gases from the air, such as O_2, CO_2 and N_2, to a small extent. Aquatic animals require dissolved O_2 and aquatic plants need dissolved CO_2 to survive.
- Water has an exceptionally high heat capacity (a measure of the heat absorbed by a substance when a rise in temperature occurs). When liquid water is heated, some of the energy is

used to overcome the many strong hydrogen bonds present in water. The oceans are able to absorb a lot of daytime heat from the Sun, ensuring only relatively small changes in temperature occur on Earth and support life.

- Water has a high heat of vapourisation as energy is also needed to break the hydrogen bonding between molecules before liquid water evaporates into a gas. This regulates human body temperature as the evaporation of a small amount of sweat requires much heat from the body.
- Pure water is a poor conductor of electricity as no ions are present and electrons are localised on molecules and not free to move.
- Water expands on freezing as the formation of four hydrogen bonds for each water molecule results in the molecules arranging themselves in a crystal structure with large spaces within ice. As the volume increases when liquid water freezes, solid water (ice) has a lower density than the liquid state (unlike almost every other substance). This means that ice floats on liquid water. In winter, the ice on the surface of a lake acts as an insulating layer for the water below and prevents the lake from becoming entirely solid. This way, aquatic life in lakes can survive winter in cold climates.
- As the water in cracks in rocks repeatedly freezes and thaws, the stress due to expanding and contracting breaks the rock into pebbles and soil.
- The presence of water can be detected using anhydrous copper sulfate and cobalt chloride paper: grey anhydrous copper sulfate turns blue as copper sulfate pentahydrate, $CuSO_4 \cdot 5H_2O$ is formed; pink cobalt chloride paper turns blue when it absorbs water to form cobalt chloride hexahydrate is $CoCl_2 \cdot 6H_2O$.

Students will investigate and develop their scientific understanding of **some reactions of water** *when they:*
- write balanced ionic equations (including state), to show the dissociation of ionic substances in water.
- describe the reaction of water with basic and acidic oxides and some of the implications of these reactions. (e.g. acid rain, engine pollution, formation of alkaline solutions).

KEY POINTS:
- Many ionic compounds are soluble in water as the attractions between water molecules and the ions are greater than the attractions between the positive and negative ions in the crystal lattice. Once dissolved, ions are kept apart from each other by the water molecules surrounding them and they become randomly distributed throughout the solution.
- The solubility of an ionic compound in water depends on the specific ionic compound, the amount of water present as the

solvent and the temperature of the water (substances are usually more soluble in hot water).

- A solution of an ionic compound dissolved in water is clear (transparent) but may be colourless or coloured depending on the particular ions present.
- The chemical equation representing the dissolution of an ionic compound into ions in water is shown as:

$$KBr(s) \xrightarrow{H_2O} K^+(aq) + Br^-(aq)$$
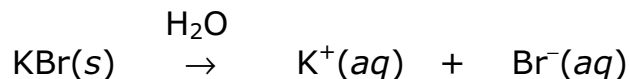
- The "H_2O" above the arrow in the equation above indicates that water is required but is not a reactant.
- The formula of an ionic compound represents the ratio of different ions in the crystal lattice and indicates the relative amounts in moles of ions that result when the compound dissolves.
- When the amount of water in a solution containing a soluble ionic compound is reduced by evaporation, the ions come together and reform the crystal lattice. This process is referred to as crystallisation.
- Oxides are known for almost every element and many main-group oxides react with water to give acidic or basic solutions.
- Ionic metal oxides are basic, such as barium oxide, BaO, and react with water to produce solutions containing hydroxide ions, $OH^-(aq)$.
- Covalent non-metal oxides are acidic, such as sulfur dioxide, SO_2. Water bonds to the atoms of the non-metal element in the oxide to form an acid (e.g. H_2SO_3), which then releases $H^+(aq)$ in water.
- Sulfur dioxide, SO_2, is formed by factories burning coal with a high sulfur content and forms sulfurous acid, H_2SO_3 when it comes it contact with water in the atmosphere. Atmospheric oxidation of SO_2 results in the formation of sulfur trioxide, SO_3. Sulfur trioxide forms sulfuric acid, H_2SO_4, on contact with atmospheric water.
- Nitrogen oxides form in the reaction of N_2 and O_2 in car engines and include NO_2, which hydrolyses to nitric acid, HNO_3, and NO in the presence of water. These acids all contribute to the problem of acid rain.

*Students will investigate and develop their scientific understanding of the **purification of water** when they:*
- explain the processes of filtration, flocculation and chlorination.*
- use a flow diagram to describe a modern water purification system.*
- construct a flow diagram which shows a local water treatment system.*
- relate water treatment to community health.*

KEY POINTS:
- Water for drinking can come from the collection of rainwater, from streams and rivers, and from underground wells. Many mineral ions, such as Fe^{3+}(*aq*) and NO_3^-(*aq*), may be present in high concentrations as well as dissolved organic compounds, which may be toxic. Fine clay particles and a range of microorganisms are also usually present. Therefore, it is often necessary to purify drinking water before it can be used for drinking.
- In purifying water for drinking, the first step involves removing large particles of debris by filtering through screens. Finer particles, including microorganisms, are removed by adding lime, CaO, and alum, $Al_2(SO_4)_3$, which react to form a gel of $Al(OH)_3$ as one product The fine particles are trapped in the gel, which settles out due to gravity as sediment at the bottom of a large tank. The water is then filtered through sand and aerated to cause oxidation of dissolved organic compounds.
- Water must then be disinfected to kill or inactivate disease-causing microorganisms. This usually occurs through the addition of chlorine gas, Cl_2, and/or aqueous solutions of the hypochlorite ion, ClO^-(*aq*). The chlorine that is added to water also gives residual protection against biological contamination in the water distribution system.
- The dissolved ions are usually left in the water but can be removed by passing the water through ion-exchange resin which is a polymer that binds ions to itself.

A lot of water for drinking comes from rivers and streams. Therefore it is important to ensure that water supplies remain clean and free from pollution that is caused by dumping rubbish into rivers or the run-off of fertilizers from farms.

S T R A N D 4: PRINCIPLES OF PHYSICAL CHEMISTRY

THERMOCHEMISTRY

All chemical and physical changes to matter are accompanied by changes in the energy content of the matter. The difference in energy before and after the change may result in energy being released to the surroundings or absorbed from the surroundings. The energy change arises from the breaking and making of bonds within molecules and between particles (atoms, ions or molecules).

*Students will investigate and develop their scientific understanding of **energy changes** when they:*
- explain that chemical reactions are accompanied by heat changes (changes in enthalpy, ΔH).
- define exothermic and endothermic reactions and recognise them in terms of the qualitative treatment of energy changes in a chemical reaction, where: $\Delta H_{reaction} = \Sigma\Delta H_{products} - \Sigma\Delta H_{reactants}$
- interpret energy profile diagrams in terms of:
    (i). activation energy.
    (ii). energy change in a reaction (ΔH).
    (iii). catalysed pathways.
- state Hess' Law and apply it to calculate the energy change for a chemical reaction by rearranging and combining two related equations.

KEY POINTS:
- Each particle in a system has potential energy (chemical potential energy stored in the bonds) and kinetic energy (energy of motion). The energy content of a system is the sum of the potential and kinetic energies of the particles.
- Thermal energy is transferred as heat when the system and surroundings are at different temperatures. This occurs in the direction from a region of high temperature to a region of low temperature (**Chem13** p50).
- As temperature is a measure of the movement of molecules (average kinetic energy), heat is transferred from a region of high particle motion to a region with lower particle motion.
- When a chemical reaction or change of state occurs, a change in the heat energy of both the system and the surroundings is observed.
- Energy is always conserved so that the total quantity of energy of the system and surroundings is constant. Therefore, a change in energy of the system is always accompanied by an opposite change in the energy of the surroundings.
- Energy coming into the system has a positive value; energy going out of the system has a negative value.
- Energy is measured in units of joules (J) or kilojoules (kJ).
- The change in energy due to heat transfer is measured using a thermodynamic quantity called enthalpy (**Chem12** p49; (**Chem13** p54).

- The change in enthalpy, ΔH, that occurs with a physical or chemical change equals the heat transferred to or from the system at constant pressure. The change in enthalpy (also called the heat of reaction) is given by ΔH, where:

$$\Delta H \;=\; H_{final} - H_{initial} \;\;=\; \Sigma H_{products} \;-\; \Sigma H_{reactants}$$

- Heat can be regarded as a "reactant" or a "product" in a physical or chemical change (**Chem12** p49, 50).
- A chemical (or physical) change that releases heat ("heat out") is *exothermic* and results in a decrease in the enthalpy of the system: $H_{final} < H_{initial}$ $\Delta H < 0$.
- The products formed in an exothermic reaction are more stable than the reactants as they have a lower enthalpy. The heat released increases the kinetic energy of all particles (reactants and products) and is also transferred to the surroundings. Therefore, an increase in temperature is observed.
- The reaction in which methane, CH_4, burns in oxygen is accompanied by a decrease in enthalpy because heat leaves the system and is exothermic. All combustion reactions are exothermic.
- A chemical (or physical) change that absorbs heat ("heat in") is *endothermic* and results in an increase in the enthalpy of the system: $H_{final} > H_{initial}$ $\Delta H > 0$.
- The products formed in an endothermic reaction are less stable than the reactants as they have a higher enthalpy. The heat is absorbed from the surroundings, which gets colder.
- Ice melts with an increase in enthalpy because heat enters the system and is endothermic.
- The heat released or absorbed during a chemical change is due to differences between the strengths of bonds or attractive forces in the products and the reactants.
- Breaking bonds requires energy; making bonds releases energy.
- In an exothermic reaction, the energy needed to break the bonds of the reactants is less than the energy released in forming the new bonds of the products.
- Bonds in fuels, such as petrol, and food are weaker (less stable, higher energy) than the bonds in the products of reactions, CO_2 and H_2O and therefore heat is given out when fuels and food react with oxygen.
- Phase changes from gas to liquid, or liquid to solid are exothermic as the kinetic energy of the particles has decreased and energy is lost to the surroundings (**Chem13** p49).
- In an endothermic reaction, the energy needed to break the bonds of the reactants is more than the energy released in forming the new bonds of the products.

- Phase changes from solid to liquid, or liquid to gas are endothermic as the particles within the system have increased kinetic energy and moving faster so must absorb energy from the surroundings.
- The sign of ΔH for a forward reaction is opposite that for the reverse reaction.
- ΔH values can be experimentally determined using a calorimeter (an insulated container) and are quoted as ΔH per mole of a reaction.
- The value of ΔH given for a reaction is related to the balanced equation for the chemical reaction.
- The magnitude of ΔH depends on the amount and physical state of the substance reacting. If the amount reacted is doubled then the magnitude of ΔH will double.
- The quantity of heat lost or gained when a given amount of substance reacts can be determined using the value for ΔH in kJ mol^{-1} and the amount of substance in moles.
- Changes in enthalpy can be represented using energy profile diagrams that show final and initial states as horizontal lines on a vertical energy axis. The difference between the two lines is the change in enthalpy, ΔH.
- The activation energy, E_a, is the minimum energy molecules must have to react and is also depicted on an energy profile diagram. The enthalpy change for a reaction, $\Delta_r H$, is independent of the size of the activation energy barrier and no correlation between the energy change and the rate of reaction occurs. For example, some exothermic reactions are very slow.
- A catalyst is a substance that increases the rate of a reaction without being consumed. It does this by providing a different pathway for the reaction to occur by that has a lower activation energy. While the reaction proceeds faster, a catalyst does not affect the overall reaction yield. The effects of a catalyst can also be depicted on an energy profile diagram.
- It is necessary to calculate the enthalpy change, $\Delta_r H$, for reactions that cannot be carried out experimentally. In these cases, the enthalpy change, $\Delta_r H$, is the sum of the enthalpy changes for reactions that add up to give the required reaction. This is referred to as *Hess's law of heat summation* (**Chem13 p55**).
- Hess's law states that the enthalpy change of an overall process is the sum of the enthalpy changes of its individual reaction steps. This means that that the enthalpy difference depends only on the difference in enthalpy between the reactants and products and not on the reaction path taken.
- Hess's law can be written as: $\Delta H_{total} = (\Delta H_1 + \Delta H_2 + \ldots. \Delta H_n)$.

RATES OF REACTION

In studying chemical change and the steps by which a reaction proceeds, it is important to establish the rate at which reactant becomes product (i.e. how the concentration of the reactants (or products) changes with time).

*Students will investigate and develop their scientific understanding of **rates of reaction** when they:*
- explain what is meant by rate of reaction.
- list the factors affecting rate of reaction (temperature, concentration, surface area of solid reactants and presence of a catalyst) and explain the effects of these factors in terms of simple collision theory.
- describe the role of a catalyst in a chemical reaction.

KEY POINTS:
- Reaction rates describe how fast or slow chemical reactions occur.
- The rate of a reaction can be determined by measuring either how fast a reactant is used up or how quickly a product is formed.
- Under a certain set of conditions, each reaction has its own characteristic rate, which is determined by the chemical nature of the reactants.
- The rate of a reaction can be controlled by altering four factors: the concentrations of the reactants, the physical state of the reactants, the temperature at which the reaction occurs, and the use of a catalyst (**Chem12** p45-47).
- For a chemical reaction to occur, particles must collide. Chemical reactions are faster if the particles collide more frequently or with greater force (**Chem12** p45; **Chem13** p50).
- A reaction occurs when particles collide with sufficient energy (the activation energy) and in the correct orientation. While not all collisions between particles result in a chemical reaction, the more frequently effective collisions occur, the faster the rate of reaction.
- Reactants remain unchanged when particles do not collide in the correct orientation or do not have sufficient kinetic energy to exceed the activation energy, E_a.
- The overall reaction rate will increase if conditions are changed to allow more frequent, effective collisions to occur.
- Concentration affects the reaction rate by influencing the frequency of collisions between reactant molecules: increasing the *concentration* of reactant particles increases the frequency of collisions as there are more particles available for collision in a given volume at any one time and this will increase the rate of the chemical reaction (**Chem13** p52).
- The rates of all chemical reactions get slower with time since the concentration of reactants decreases as the reaction progresses.

- Physical state affects reaction rate by determining the surface area per unit volume of the reactants: as the surface area of the reactant is increased, more particles are exposed at the surface to collisions with another reactant, resulting in a greater frequency of collisions that increases the rate of reaction.
- The surface area of a solid reactant can be increased by crushing or powdering one lump into smaller pieces. Dissolving a solid reactant separates the solid into individual particles and increases its surface area available for collision to the maximum (**Chem13** p52).
- Temperature affects the reaction rate by influencing the frequency and the energy of the collisions: when temperature is increased, the average kinetic energy of the particles increases and collisions with other particles are more frequent. The effectiveness of collisions is also increased as a larger proportion of the collisions will exceed the activation energy (**Chem13** p50, 51).
- A catalyst speeds up a reaction by providing a new lower energy pathway for the reaction to proceed by and providing a lower activation energy, E_a. A lower E_a results in an increased frequency of effective collisions and faster reaction rate (**Chem13** p51, 52).

EQUILIBRIUM

The rate of a reaction and the yield of products are not necessarily related in a chemical reaction. The concept of chemical equilibrium is used to indicate the extent of a reaction. This is shown by the concentration of reactants and products that no longer change with time.

*Students will investigate and develop their scientific understanding of **equilibrium** when they:*
- define dynamic equilibrium as when forward and reverse reaction rates in a system are equal and represent this by chemical equations where A + B ⇌ C + D.
- recognise a closed and an open system and infer that equilibrium can only occur in a closed system.
- write the equilibrium constant expression for a system and recall that the equilibrium constant, *K*, is a number that is equal to a particular ratio of concentration terms.
- explain how the magnitude of the equilibrium constant, *K*, is related to the extent of the reaction.
- explain the qualitative effects on a system at equilibrium by changing:
  (i). temperature.
  (ii). total pressure (as in increasing pressure by reducing volume of container and *vice versa*).
  (iii). concentration.
  (iv). adding a catalyst to a system at equilibrium.
- define Le Châtelier's principle.*
- apply equilibrium and rate consideration in industrial (commercial) production of:
  (i). ammonia (Haber process).
  (ii). sulfuric acid (Contact process).

KEY POINTS:
- Every chemical reaction is reversible if all the reactants and products are kept in contact with one another in a closed container. Therefore, as the amounts of products increase, they may react to reform the reactants.
- At "equilibrium", the rates of the forward and reverse reactions are equal and there is no further change in the composition of the mixture at a particular temperature. The relative amounts of reactants and products at this point depend on the particular reaction taking place and the temperature at equilibrium (**Chem12** p53).
- On the macroscopic scale, the reaction appears to have stopped but at the molecular level it is dynamic as molecules breakdown and reform.
- In a closed container at constant temperatures, phase changes of many substances are reversible and reach equilibrium when particles trapped in the container move from one phase to the other at equal rates (**Chem13** p61).

- A reaction will go to completion (all reactants are consumed) without reaching equilibrium if a product is removed from the system, such as a gas, or if a product exists in a form that makes it less available for reaction by lowering its concentration, such as a precipitate.
- Substances in a container open to the surroundings do not reach equilibrium between the liquid and gas phases as gas particles are lost to the surroundings and the liquid completely evaporates.
- Reactions at equilibrium are represented by equations with a pair of arrows pointing in opposite directions ( ). The forward reaction converts reactants to products and the reverse reaction converts products to reactants (**Chem13** p62).
- The equilibrium constant, K, is a number based on a particular ratio of product and reactant concentrations at equilibrium at a particular temperature. The value of K has no units (**Chem12** p53; **Chem13** p62).
- For the reaction:

$$H_2(g) \; + \; I_2(g) \; \Leftrightarrow \; 2HI(g) \qquad K \; = \; \frac{[HI]^2}{[H_2][I_2]}$$
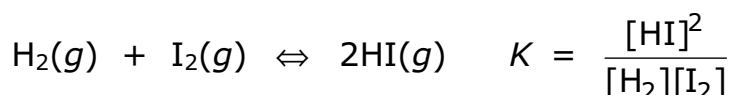
  where [HI] is the magnitude of the concentration of HI at equilibrium in mol L^{-1}.
- The magnitude of the equilibrium constant, K, is an indication of how far a reaction proceeds towards the product at a given temperature. Each reaction has its own characteristic value of K which can only be changed by changing the temperature.
- The equilibrium constant, K, is small (<<1) for reactions that reach equilibrium with a high concentration of reactants. A very low value of K ($<10^{-6}$) indicates that almost no reactants are being converted to product.
- The equilibrium constant, K, is large (>>1) for reactions that reach equilibrium with a low concentration of reactants and a very large value of K ($> 10^{6}$) indicates equilibrium is not reached until almost all reactants are converted to products and the reaction is said to go to completion.
- Pure liquids and solids do not appear in the terms of K as their concentrations are constant.
- The equilibrium constant, K, indicates whether the reaction is reactant-favoured or product-favoured and by how much but does not specify the time taken to reach equilibrium.
- When a system at equilibrium is disturbed by a change applied to it, a reaction occurs to counteract the disturbance and when equilibrium is restored, reactant and product concentrations are changed. This is referred to as Le Châtelier's principle.
- Factors that can cause the reactant and product concentrations in the reaction mixture at equilibrium to change include changes in: the concentration of reactant or product;

temperature, and pressure (for equilibrium systems with gas components) (**Chem12** p54, 55; **Chem13** p62).

- Whenever the concentration of a component changes, the system at equilibrium reacts to consume some of the added substance or produce more of the removed substance (**Chem12** p54; **Chem13** p62).
  - The reaction occurs in the forward direction if a reactant is added or a product is removed.
  - The reaction occurs in the reverse direction if a reactant is removed or a product is added.
- When the volume of a vessel containing gases at equilibrium changes, a change in the concentration of the gases as well as gas pressure occurs: a decrease in container volume raises the concentration and gas pressure, and an increase in volume lowers the concentration and gas pressure for reactions having different amounts in moles of gaseous reactants and products (**Chem12** p54; **Chem13** p62).
- When the volume becomes smaller (resulting in a higher pressure), a reaction occurs so that the total number of gas molecules decreases (gives the smaller amount in moles of gases).
- When the volume becomes larger (resulting in a lower pressure), a reaction occurs so that the total number of gas molecules increases (gives the larger amount in moles of gases).
- Although the equilibrium concentrations that result from concentration and volume changes are different from those of the original equilibrium mixture, the magnitude of the equilibrium constant, K, remains the same at a particular temperature.
- A temperature change does change the magnitude of the equilibrium constant, K, as heat can be considered as a component of the equilibrium system. (**Chem12** p55; **Chem13** p62)
- A temperature increase (adding heat) results in the endothermic (heat-absorbing) direction.
- A temperature decrease (removing heat) results in the exothermic (heat-releasing) direction.
- A higher temperature increases the equilibrium constant, K, for an endothermic reaction (positive ΔH) and decreases K for an exothermic reaction (negative ΔH). This differs from the effect of an increase in temperature on reaction rate which is to increase the rate of reaction equally in both directions.
- Addition of a catalyst causes the system to reach equilibrium faster by speeding up the forward and reverse reactions equally, but does not affect the concentrations of reactants and products in the reaction mixture. Catalysts have the effect of

increasing the rate of the reaction so that the reaction will come to equilibrium more quickly (**Chem12** p55; **Chem13** p62).

- Ammonia, NH_3, is used to make nitric acid, fertilisers, dyes and explosives and is produced from N_2 and H_2 through the Haber process:

$$N_2(g) \; + \; 3H_2(g) \; \Leftrightarrow \; 2NH_3(g) \quad \Delta H \; = \; \text{-91.8 kJ}$$
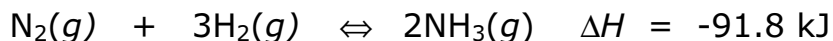
- Nitrogen and hydrogen gases are mixed in the ratio 1:3, pressurised to around 200 atmospheres and passed over a catalyst of powdered iron at a temperature between 350 and 500^oC to produce ammonia gas.
- The use of high pressure (decrease in volume), a low temperature and continual removal of product maximises the yield of ammonia.
- A low temperature results in a slow reaction, therefore an intermediate temperature and a catalyst are used to make the process economical.
- Huge amounts of sulfuric acid, H_2SO_4, are manufactured industrially each year, mostly by the Contact process. This process is based on the catalysed oxidation of $SO_2(g)$ to give $SO_3(g)$.

$$SO_2(g) \; + \; \tfrac{1}{2}O_2(g) \; \Leftrightarrow \; SO_3(g) \quad \Delta H \; = \; \text{-99 kJ}$$
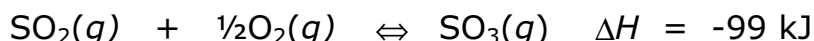
- $SO_2(g)$ is formed from burning sulfur in dry air:

$$S(l) \; + \; O_2(g) \; \rightarrow \; SO_2(g)$$
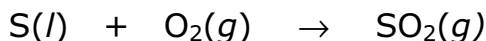
- Sulfur dioxide is mixed with more air and is passed through a catalyst (vanadium pentoxide, V_2O_5) to produce sulfur trioxide. The catalyst is active between $400 - 600^0C$.
- The yield of $SO_3(g)$ can be increased by removing heat as this reaction is exothermic. However, the reaction is very slow at room temperature and therefore a catalyst is used to optimise rate and yield.
- Increasing the pressure will result in a the production of more products and an increased yield of $SO_3(g)$. However, the pressure effect is small and not economical to exploit.
- Adjusting concentrations by adding excess $O_2(g)$ and removing $SO_3(g)$ will yield more $SO_3(g)$.
- The SO_3 gas is cooled and dissolved in concentrated sulfuric acid to form a thick fuming liquid called "oleum", $H_2S_2O_7$.

$$SO_3(g) \; + \; H_2SO_4(l) \; \rightarrow \; H_2S_2O_7(l)$$
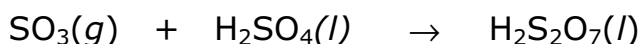
- The oleum is cooled and diluted with water to produce sulfuric acid.

$$H_2S_2O_7(l) \; + \; H_2O(l) \; \rightarrow \; 2H_2SO_4(l)$$
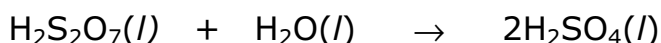
- Sulfuric acid is used to make fertilisers such as ammonium sulfate and superphosphate (which includes a soluble phosphate compound, $Ca(H_2PO_4)_2$, that is required by plants).

$$Ca_3(PO_4)_2(s) + 2H_2SO_4(aq) \rightarrow Ca(H_2PO_4)_2(s) + 2CaSO_4(s)$$
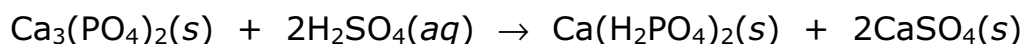
ACID-BASE EQUILIBRIA

Many important reactions in living organisms involve acid-base reactions occurring in aqueous solutions. The principles of chemical equilibrium apply to acids and bases in water. For acid-base reactions in water, water can be a reactant or a product as well as the solvent. Chemists have devised definitions for acids and bases as a way to better understand acid-base reactions.

*Students will investigate and develop their scientific understanding of **aqueous acid and base equilibria** when they:*
- recall the Brønsted-Lowry acid-base definitions of acids as proton donors and bases as proton acceptors.
- differentiate between strong and weak acids in water and strong and weak bases in water and write relevant equations to show this (i.e. how the strength of an acid or base is related to the extent of its dissociation into ions in water).
- explain how water can act as a proton donor and a proton acceptor (i.e. water is amphiprotic).
- describe the equilibrium reaction for water (autoionization) and its ion-product constant, K_w, by the expression $[H_3O^+][OH^-] = K_w = 1 \times 10^{-14}$ for pure water at 25°C.
- explain why $[H_3O^+]$ is inversely related to $[OH^-]$ in any aqueous solution.
- define pH = $-\log[H_3O^+]$.
- calculate pH given $[H_3O^+]$ and *vice versa*.
- use the pH scale to differentiate acids, bases, and neutral substances.
- describe the basic properties of hydroxides and write equations to indicate such properties.*
- define pOH = $-\log[OH^-]$ for the hydroxide ion concentration in solution.*
- recall that pH + pOH = 14 in any aqueous solution at 25°C.*

KEY POINTS:
- Classical definitions state that an acid is a substance that reacts with water to produce $H^+(aq)$ ions and a base is a substance that produces $OH^-(aq)$ ions when it is dissolved in water (**Chem12** p57).
- By these definitions, an acid-base reaction is the reaction of $H^+(aq)$ and $OH^-(aq)$ to form H_2O.
- Because the products of an acid-base reaction have lower acidity and basicity than the reactants, these reactions are also referred to as neutralisation reactions.
- A broader definition of an acid and a base that includes acids and bases in aqueous solutions is the Brønsted-Lowry acid-base definition (**Chem12** p57).

- The Brønsted-Lowry definition of an acid is a substance that donates a proton, H^+.
- Proton donation to a water molecule forms H_3O^+ (hydronium) ions.
- The Brønsted-Lowry definition of a base is a substance that accepts a proton, H^+.
- Proton donation by water to a base produces OH^- ions.
- Substances that behave as both acids and bases are referred to as amphiprotic. Water is an amphiprotic substance.
- Acids react with bases by proton transfer.
- In an acid-base reaction, acids and bases form their conjugates (**Chem12** p58).
- When an acid reacts, it forms a conjugate base with one fewer H and one more negative charge than the acid.
- When a base reacts, it forms a conjugate acid with one more H and one fewer negative charge than the base.
- The strength of an acid refers to the extent to which it donates protons to water and is shown by the ratio of the concentration of H_3O^+(*aq*) relative to the concentration of the acid before it reacts with water (**Chem12** p58; **Chem13** p70).
- Strong acids (and strong bases) react to the extent of 100% with water. (**Chem12** p58; **Chem13** p70).
- Weak acids (and weak bases) react with water to less than 10% (**Chem12** p58; **Chem13** p70, 87).
- Strong acids have weak conjugate bases that do not react with water. Weak acids have strong conjugate bases.
- The *strength* of an acid cannot be determined using universal indicator since indicators only show the concentration of H_3O^+ ions in solution, irrespective of whether they originate from a strong or a weak acid.
- A strong acid and a strong base react to form water and a metal salt (**Chem12** p58).
- The extent of dissociation of an acid is expressed by the acid-dissociation constant, K_a (**Chem12** p59; **Chem13** p72).
- Strong acids have large K_a values; weak acids have K_a values ranging from about 10^{-1} to 10^{-12}.
- Water reacts with itself to a very small extent to produce ions in a process called autoionisation. Water is a very weak conductor of electricity because of this process (**Chem12** p58).

$$H_2O \ + \ H_2O \ \Leftrightarrow \ H_3O^+ \ + \ OH^-$$
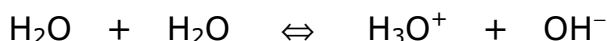
- This equilibrium is very reactant-favoured so that most of the water is in the molecular form.
- One hydronium ion and one hydroxide ion are produced for each H_2O that dissociates. Therefore, in pure water $[H_3O^+] = [OH^-]$.

- The equilibrium constant for this reaction is K_w. This is given by the equation:

$$K_w = [H_3O^+][OH^-]$$

  and this applies in any aqueous solution.
- From this equation, $[H_3O^+]$ and $[OH^-]$ in any aqueous solution are inversely related.
- At 25 ^{o}C, in pure water $K_w = 1 \times 10^{-14}$ and therefore the concentration at equilibrium of *both* H_3O^+ and OH^- ions is 1×10^{-7} mol L^{-1}.
- K_w equals 1.0×10^{-14} in all aqueous solutions at 25°C. Therefore, both H_3O^+ and OH^- ions are present in all aqueous solutions.
- When an acid is added to pure water, the $[H_3O^+]$ increases and the $[OH^-]$ decreases since the ion-product, K_w, remains constant.
- In acidic solutions, $[H_3O^+] > [OH^-]$; in basic solutions, $[H_3O^+] < [OH^-]$.
- In neutral solutions, $[H_3O^+] = [OH^-] = 1 \times 10^{-7}$ mol L^{-1}.
- To express $[H_3O^+]$ more simply, the pH scale is used, where pH = -log $[H_3O^+]$ (**Chem12** p59; **Chem13** p71).
- The pH of a solution is a measure of the concentration of the H_3O^+ (hydronium) ion.
- A high pH represents a low $[H_3O^+]$, therefore an acidic solution has a lower pH than a basic solution.
- In acidic solutions, pH < 7.00; in basic solutions, pH > 7.00; and in neutral solutions, pH = 7.00 (as $[H_3O^+]$ = 1×10^{-7} mol L^{-1}).
- Main-group metal oxides are ionic and produce OH$^-$ ions if they dissolve in water.
- pOH = -log$[OH^-]$; acidic solutions have a higher pOH than basic solutions.
- pH + pOH = 14 (= pK_w) at 25$\circ$C.

S T R A N D 5: OXIDATION AND REDUCTION

Any chemical reaction that has an element as a reactant or a product is a redox (oxidation-reduction) reaction. A redox reaction involves the transfer of electrons from one reactant to another.

REDOX REACTIONS

*Students will investigate and develop their scientific understanding of **oxidation state** when they:*
- determine the oxidation state (number) of each atom in a given element, molecule or ion.
- b) recognise and explain oxidation and reduction reactions in terms of:
  (i). changes in oxidation states (numbers).
  (ii). transfer of electrons.
  (iii). transfer of oxygen.

KEY POINTS:
- The driving force for redox reactions is the *transfer of electrons* from one atom to another atom than has a greater attraction for electrons (**Chem12** p71).
- Oxidation is the *loss* of electrons.
- Reduction is the *gain* of electrons.
- The addition of oxygen to an element, a compound or an ion is also referred to as oxidation (**Chem12** p71).
- The removal of oxygen from a compound or polyatomic ion is also referred to as reduction (**Chem12** p71).
- In every redox reaction, electron loss (oxidation) is always accompanied by electron gain (reduction).
- A useful way of tracking the ownership of electrons in a redox reaction involves assigning oxidation numbers to each atom in an element, compound or ion using a set of rules (**Chem12** p71, 72; **Chem13** p90).
- The oxidation number of an atom in an element is zero.
- The oxidation number of an atom in a monatomic ion is the same as the charge on the ion.
- Hydrogen is assigned an oxidation number of +1 (except in metal hydrides where the oxidation number is -1).
- Oxygen is assigned an oxidation number of -2 (except in peroxides where the oxidation number is -1).
- The sum of the oxidation numbers of all atoms in a molecule is zero.
- In polyatomic ions, the sum of the oxidation numbers of all the atoms is equal to the charge on the ion.
- Oxidation has occurred if the oxidation number of an atom increases. Reduction has occurred if the oxidation number of an atom decreases.

REDOX EQUATIONS

*Students will investigate and develop their scientific understanding of **oxidising and reducing agents** when they:*
- define oxidising agents (oxidants) as substances that accept electrons in a reaction and undergo a decrease in oxidation number.
- define reducing agents (reductants) as substances that donate electrons in a reaction and undergo an increase in oxidation number.
- identify the common oxidising agents (oxygen, chlorine, metals with dilute acids, hydrogen peroxide, permanganate, dichromate).
- identify the common reducing agents: metals (e.g. zinc, magnesium and iron), carbon, sulfur dioxide, carbon monoxide.
- write balanced half equations for a given oxidation or reduction process using the ion-electron method.
- write fully balanced equations for the redox reactions by combining half equations for oxidation and reduction.

KEY POINTS:
- Oxidising agents (or oxidants) are substances that **gain** electrons in redox reactions (**Chem12** p71, 74; **Chem13** p92).
- Oxidising agents contain an atom that decreases in oxidation number in a redox reaction.
- Reducing agents (or reductants) are substances that **lose** electrons in redox reactions (**Chem12** p71, 74; **Chem13** p92).
- Reducing agents contain an atom that increases in oxidation number in a redox reaction.
- The total number of electrons lost by the reducing agent in a redox process must equal the total number of electrons gained by the oxidising agent (**Chem12** p73, 74).
- Oxidising and reducing ability are related to an element's position in the periodic table: elements in Groups 1 and 2 are strong reducing agents; elements in Group 17 and oxygen (Group 16) are strong oxidising agents.
- Common oxidising agents include (**Chem12** p74):
    - Oxygen, O_2, produces the oxide ion, O^{2-} in redox reactions. All combustion reactions are redox reactions.
    - Chlorine, Cl_2 (a yellow-green gas), produces the colourless chloride ion, Cl^-. Chlorine is used as a disinfectant and to sterilise swimming pools on account of its ability to oxidise other substances.
    - Hydrogen ions, $H^+(aq)$, in dilute acids produce hydrogen gas, H_2, in redox reactions with metals more reactive than hydrogen. These are potassium, K, sodium, Na, calcium, Ca, magnesium, Mg, aluminium, Al, zinc, Zn, iron, Fe, tin, Sn, and lead, Pb. Most of these reactive metals also react violently with water and are not found naturally in elemental form.
    - Hydrogen peroxide, H_2O_2 (a colourless liquid), produces water.
    $$H_2O_2(l) + 2H^+(aq) + 2e^- \rightarrow 2H_2O(l)$$
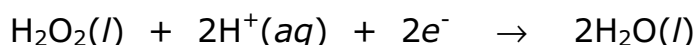
- Permanganate ion, MnO_4^-, (a purple ion) produces colourless Mn^{2+} ion when the reaction is carried out in an acidic solution.

  $$MnO_4^-(aq) \ + \ 8H^+(aq) \ + \ 5e^- \ \rightarrow \ Mn^{2+}(aq) \ + \ 4H_2O(l)$$

  When the reaction is carried out in neutral or slightly basic conditions, brown solid manganese dioxide, MnO_2, is produced.

  $$MnO_4^-(aq) + 2H_2O(l) + 3e^- \ \rightarrow \ MnO_2(s) + 4OH^-(aq)$$
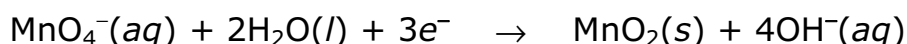
- Dichromate ion, $Cr_2O_7^{2-}$ (an orange ion) produces the green Cr^{3+} ion. The reaction requires acidic conditions.

  $$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \ \rightarrow \ 2Cr^{3+}(aq) + 7H_2O(l)$$

- Metal ions in aqueous solution react to form the corresponding metal in a redox reaction with a more reactive metal (a metal above it in the activity series). For example, zinc will react with a copper sulfate solution and produce copper metal and a zinc sulfate solution.
- Common reducing agents include (**Chem12** p74):
  - Reactive metals (those high on the activity series) produce the appropriate metal ion. For example, Zn forms the colourless Zn^{2+} ion, Mg forms the colourless Mg^{2+} ion and Fe forms the pale-green Fe^{2+} ion. The pale-green iron(II) ion, Fe^{2+}, is also easily oxidised to the orange Fe^{3+} ion.
  - Carbon, C, (as black solid graphite or coal) produces carbon monoxide, CO (a colourless, poisonous gas) when a limited amount of oxygen is present or at high temperatures. In a plentiful supply of air, carbon and carbon monoxide are both oxidised to CO_2 (a colourless gas).
  - Sulfur dioxide, SO_2, (a colourless gas) produces the colourless sulfate ion, SO_4^{2-}. Sulfur dioxide is commonly used as a food preservative in wine and dried fruits as it delays oxidation reactions by oxygen in the air.
- Balancing redox equations is commonly carried out using the ion-electron half-equation method which divides the overall redox reaction into oxidation and reduction half-reactions (**Chem12** p74; **Chem13** p90).
  - **Step 1** Write the oxidation and reduction processes as separate half-reactions that each contain the oxidised and reduced forms of one of the reactants.
  - **Step 2** Balance atoms other than O and H atoms.
  - **Step 3** Balance the O atoms by adding the appropriate number of water molecules.
  - **Step 4** Balance the H atoms by adding H^+ ions.

- **Step 5** Balance the charge by adding electrons, e^-. They will appear on the left in the reduction half-reaction and appear on the right in the oxidation half-reaction.
- **Step 6** If necessary, multiply one or both half-reactions by an integer to make the number of electrons gained in reduction equal the number of electrons lost in oxidation.
- **Step 7** Add the balanced half-reactions and check that the atoms and charges are balanced and that no electrons appear in the overall equation.

APPLICATIONS OF REDOX REACTIONS

*Students will investigate and develop their scientific understanding of **applications of chemistry in redox reactions** when they:*
- apply oxidation and reduction processes to the electrolysis of some common ionic solutions e.g. NaCl(*aq*) and molten ionic compounds e.g. NaCl(*l*).
- describe observations made at the electrodes (anode and cathode) during electrolysis.
- recognise and interpret instances of oxidation and reduction in settings commonly found in society and the environment e.g. batteries of vehicles; corrosion of metals in vehicles, buildings and bridges; galvanic protection with sacrificial electrodes; fuels; breathalyser test.

Electrochemical cells are systems that use a redox reaction either to produce electricity spontaneously or to utilise electricity to drive a non-spontaneous redox reaction. Redox reactions drive many natural electrochemical changes, such as combustion and corrosion of iron.

KEY POINTS:
- Electrochemical cells consist of two electrodes. An oxidising agent is present in one electrode and a reducing agent is present in the other electrode (**Chem12** p77).
- A conductor, such as graphite or a metal, is dipped into an electrolyte solution (an aqueous solution of ions that are involved in the reaction or carry the charge). The electrodes are connected by an external circuit which allows electrons to flow from one electrode to the other as electricity (**Chem13** p87, 94,95).
- In galvanic cells (batteries), a spontaneous redox reaction occurs between an oxidising agent and a reducing agent when the electrodes are connected through an external circuit.
- A galvanic cell operates because of the different tendencies of the reducing agent in one electrode and the oxidising agent in the other electrode to accept electrons.
- The reducing agent forms a negative electrode where electrons are released to flow through the external circuit.
- The oxidising agent forms the positive electrode where the electrons released by the reducing agent are consumed as they enter the electrode.

- The electrical energy of the electrons moving through the wire from the negative electrode to the positive electrode can be used in electrical devices such as light bulbs and motors.
- Lead-acid car batteries use a redox reaction involving lead and lead oxide immersed in an acid solution to generate an electrical current that is used to start the car. Both half-reaction produce $Pb^{2+}(aq)$ ions that then react with $HSO_4^-(aq)$ (**Chem13** p102).
  - Negative electrode:
    $Pb(s) + HSO_4^-(aq) \rightarrow PbSO_4(s)\ H^+(aq) + 2e^-$

  - Positive electrode:
$PbO_2(s) + 3H^+(aq) + HSO_4^-(aq) + 2e^- \rightarrow PbSO_4(s) + 2H_2O(l)$

  - Overall:
    $PbO_2(s) + Pb(s) + 2H_2SO_4(aq) \rightarrow 2PbSO_4(s) + 2H_2O(l)$

  The reverse reaction occurs while the car is moving and this is known as recharging the battery.
- The recharging process in a lead-acid car battery is an example of an electrolytic cell, where an external source of electricity provides electrons with the energy required to cause a non-spontaneous redox reaction to occur.
- This process involves passing an electric current through an electrolyte (either a molten ionic compound or an aqueous solution of an ionic salt) that has electrodes immersed in it. The electrodes are often inert, such as graphite or platinum) and the reaction only occurs at the surface of the electrodes (**Chem12** p77; **Chem13** p103).
- Electrons from the external power source are supplied to the negative electrode, called the cathode, and are removed from the positive electrode, called the anode.
- Cations in the electrolyte move towards the cathode (negative electrode) and are reduced (gain electrons).
- Anions in the electrolyte move towards to the anode (positive electrode) and are oxidised (lose electrons).
- Electrolysis may be used to split a compound into its elements. Many industries use electrolytic cells for preparing commercially important elements such as chlorine, aluminium and sodium.
- Metallic sodium and chlorine gas, Cl_2, are produced industrially by the electrolysis of molten sodium chloride (**Chem12** p78):

  - cathode reaction: $Na^+(l) + e^- \rightarrow Na(s)$

  - anode reaction: $2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$

  - Overall equation for the cell reaction:
    $2Cl^-(l) + 2Na^+(l) \rightarrow Cl_2(g) + 2Na(s)$

- Aluminium is prepared by the electrolysis of alumina, Al_2O_3.

- Magnesium is prepared by the electrolytic extraction of magnesium chloride where the magnesium ions are obtained from sea water.
- In electrolyte solutions that contain a mixture of substances, the stronger reducing agent reacts at the anode, and the stronger oxidising agent reacts at the cathode.
- When the electrolyte is an aqueous solution, water may undergo oxidation (to produce oxygen gas) or reduction (to produce hydrogen gas). This depends on the concentration and ease of oxidation and reduction of the ions in solution compared to water.
- The electrolysis of a concentrated aqueous solution of NaCl (referred to as brine) results in oxidation of Cl^- ions to chlorine gas, Cl_2, at the anode and reduction of water to hydrogen gas at the cathode (as Na^+ ions are not easily reduced) (**Chem12** p78).
- The electrolysis of a very dilute aqueous solution of NaCl results in the oxidation of water at the anode to producing O_2 gas rather than the formation of Cl_2 gas (**Chem12** p78).
- The activity series is a list of metals in order of their reactivity or tendency to lose electrons. In general, a metal ion can be converted into its metallic form by any metal above it in the activity series since the most reactive metal loses its electrons most easily and forms ions.
- Iron corrosion damages and weakens metal structures. This is a natural redox process that occurs in the presence of oxygen and moisture, and is increased by acidic conditions, salt water, or contact with a less reactive metal, such as copper.
- Rusting of iron involves oxidation of Fe metal and reduction of O_2 and eventually solid iron(III) oxide, Fe_2O_3, forms as rust.
- An effective method of protecting iron from rusting is to coat the iron with a thin layer of zinc metal to produce "galvanised iron". The zinc protects the iron because it is more reactive towards oxygen and water and corrodes in preference to the iron. This is called "sacrificial" protection as the iron will not corrode until all of the zinc has gone.
- Small blocks of zinc are attached to the steel (an alloy of iron and carbon) of bridges and hulls of boats to help prevent the steel corroding seawater. The zinc blocks need to be replaced regularly as they corrode away.
- The presence of ethanol, an alcohol, in a person's breathe can be detected using a breathalyser test. Ethanol is easily oxidised in a redox reaction to ethanoic acid by the dichromate ion in acid conditions ($Cr_2O_7^{2-}/H^+$). When this redox reaction occurs, the $Cr_2O_7^{2-}$ is reduced to Cr^{3+}, and the colour changes from orange to green, giving a visual indicator for the presence of ethanol.

- Combustion reactions involve substances reacting with oxygen, usually with the release of large amounts of heat and light, such as the burning of fuels in air. All combustion reactions are redox reactions as elemental oxygen is a reactant. The combustion of hydrocarbon fuels in air produces CO_2 and H_2O as well as heat and light. Other fuels that undergo combustion reactions include coal and natural gas (methane, CH_4).
- Combustion of hydrocarbons in a limited supply of oxygen results in "incomplete" combustion and the production of carbon monoxide gas, CO, and soot (carbon) particles instead of CO_2. Carbon monoxide is toxic as it prevents the transportation of oxygen in blood.

Petrol contains octane, a hydrocarbon, which burns in excess oxygen to form CO_2 and H_2O in a redox reaction. The energy produced by this reaction is used in engines to move a vehicle.

S T R A N D 6: ORGANIC CHEMISTRY

The vast majority of known compounds are organic compounds. Organic compounds contain carbon atoms that are usually bonded to other carbon atoms and may also be bonded to hydrogen and other elements. The large number of organic compounds arises as a carbon atom is able to form strong covalent bonds to itself, as well as up to four atoms of other non-metal elements. Carbon can also form multiple bonds with itself and with O and N.

Organic compounds have been found with a wide range of reactive atoms or groups of atoms attached to a chain or ring of carbon atoms. These reactive sites are called *functional groups* and react in a characteristic way. Molecules having a particular functional group have the same chemical properties.

HYDROCARBONS AND THEIR PROPERTIES

*Students will investigate and develop their scientific understanding of **alkanes, alkenes** and **alkynes** when they:*
- use the appropriate general formula to determine the molecular formulae of alkanes C_{1-8}, alkenes C_{2-5}, and alkynes C_{2-5}.
- apply the concept of structural isomerism to alkanes and alkenes by writing structural formulae.
- identify the functional groups of alkanes, alkenes and alkynes.
- use the IUPAC rules to name simple hydrocarbons.
- explain the meaning of saturation and unsaturation in hydrocarbons.

KEY POINTS:
- Organic compounds contain carbon atoms bonded to one another to form chains (linear or branched) or rings. These carbon atoms are also typically bonded to hydrogen (**Chem12** p82).
- The names of organic compounds are based on the number of carbon atoms and indicate the type of organic compound by a particular ending. Names of organic compounds have a root (middle part) for the longest carbon chain, a prefix (at the beginning) for any attached group, and a suffix (at the end) for the type of compound (**Chem12** p84).
- The simplest type of organic compounds, called hydrocarbons, contain only atoms of carbon and hydrogen.
- As hydrocarbon molecules only contain C and H atoms, their physical properties depend on the strength of their intermolecular attractions.
- Because C and H atoms have similar electronegativity values, hydrocarbons are non-polar and are insoluble in water.
- Hydrocarbons can be classified into four main groups: alkanes, alkenes, alkynes and aromatics. (**Chem12** p81).

- Alkanes, alkenes and alkynes exist as homologous series where the formula of each successive member differs from the previous one by a CH_2 unit (**Chem12** p81).
- The molecular formula of a compound shows the number and type of atoms in a molecule.
- A structural formula is a 2-dimensional representation showing how the atoms are linked. Lines are used to represent a shared pair of electrons in a covalent bond.
- Compounds with the same molecular formula (same number of each type of atom) but a different atom-to-atom bonding sequence (structural formula) are called *structural isomers*. Isomers have different physical properties, such as melting point and boiling point. If different functional groups are present, structural isomers will also have different chemical properties (**Chem12** p84).
- **Alkanes** contain only single covalent bonds linking the C atoms (**Chem12** p81-83).
- As each C atom has four single covalent bonds to other atoms, alkanes are referred to as *saturated* hydrocarbons.
- The general formula for alkanes is C_nH_{2n+2}, where molecules having n carbon atoms contain $2n+2$ hydrogen atoms.
- The geometry of bonds around each C atom in alkanes is tetrahedral with bond angles of $109.5°$ (not $90°$ as shown in structural formulae).
- The main source of alkanes is petroleum (a mixture of natural gas containing lighter hydrocarbons and crude oil containing heavier hydrocarbons). Fractional distillation is used to separate the alkane components of crude oil into pure substances, based on their differing boiling points.
- As the size of the molecule alkane increases, the strength of the attractions between the molecules increases. Therefore, the melting point and boiling point increases as the alkane molecules get larger since it takes more energy to separate molecules when changing from a liquid to a gas.
- Alkanes containing 4 or more C atoms can exist as structural isomers of "straight chain" molecules or "branched chain" molecules (**Chem12** p84).
- Alkanes are named systematically to show the longest chain of carbon atoms and the position of any substituent branches on the chain. The suffix for naming alkanes is *–ane*.
- **Alkenes** contain least two C atoms linked by a double covalent bond (C=C) (**Chem12** p87, 88).
- Because of the presence of an electron-rich reactive multiple C=C bond, alkenes are referred to as *unsaturated* hydrocarbons.
- Alkenes with one double bond have the general formula C_nH_{2n}.
- Alkenes are non-polar and insoluble in water.

- The suffix for naming alkenes is –*ene*. When the carbon chain is more than 3 atoms long, a number indicates where the double bond occurs.
- The geometry of bonds around a double bonded C is trigonal planar and the bond angles are $120°$.
- Alkenes can exist as geometrical isomers, where two different groups (or atoms) attached to the C atoms of the double bond can be arranged as *cis* (same side of C=C) or *trans* (across C=C) (**Chem12** p88; **Chem13** p111).
- **Alkynes** contain least two C atoms linked by a triple covalent bond (C∘C) (**Chem12** p92, 93).
- Because of the presence of an electron-rich reactive multiple C∘C bond, alkynes are also referred to as *unsaturated* hydrocarbons.
- Alkynes with one triple bond have the general formula C_nH_{2n-2}.
- Alkynes are non-polar and insoluble in water
- The suffix for naming alkynes is –*yne*. When the carbon chain is more than 3 atoms long, a number indicates where the triple bond occurs.
- The geometry of bonds around a triple bonded C is linear and the bond angle is $180°$.

Students will investigate and develop their scientific understanding of the
reactions of alkanes, alkenes *and* **alkynes** *when they:*
- recognise and write equations for:
  - (i). substitution reactions in terms of alkanes reacting with halogens
  - (ii). addition reactions in terms of alkenes reacting with halogens, acidified water (H_2O/H^+) and hydrogen (H_2) and alkynes reacting with halogens and hydrogen (H_2).
- describe and use the bromine and $KMnO_4$ tests for detection of unsaturated compounds.
- describe the preparation of ethene (by steam cracking and by dehydration), and ethyne (from calcium carbide), and write equations for each.
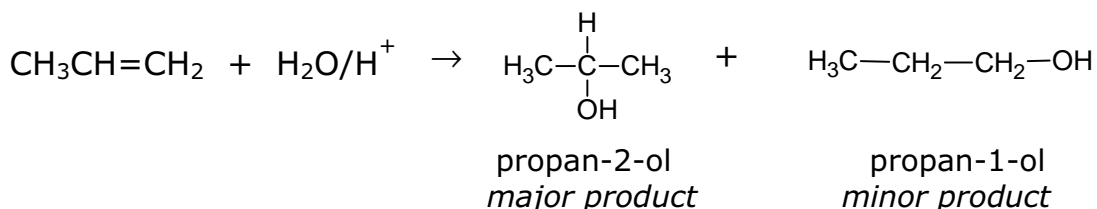
KEY POINTS:
- *Complete* combustion occurs when alkanes when burn in excess oxygen (or air) to produce carbon dioxide, CO_2, and water, H_2O and a clear blue, clean-burning flame that releases much heat (**Chem12** p83).
- *Incomplete* combustion occurs when the combustion reaction of alkanes occurs in a limited supply of air or oxygen. This reaction produces carbon monoxide, CO, or carbon, C, as well as water vapour, H_2O. The flame is sooty and yellow coloured due to the presence of glowing particles of carbon as soot. Exhaust fumes from cars contain toxic carbon monoxide gas due to incomplete combustion of petrol in the engine (**Chem12** p83).
- Alkanes undergo substitution reactions with halogens, such as Br_2. This involves one of the atoms from Br_2 (or other halogen)

substitutes for a H atom in the alkane to produce a haloalkane compound and HBr. The reaction is very slow and requires sunlight (ultraviolet light) as a catalyst. As the reaction with Br_2 progresses, the solution is slowly decolourised as orange Br_2 molecules are converted to bromoalkane and hydrogen bromide, HBr, which are both colourless (**Chem12** p83).

$$CH_3CH_2CH_2CH_3 \;+\; Br_2 \;\;\xrightarrow{\text{UV}}\;\; CH_3CH_2CH_2CH_2Br \;+\; HBr$$
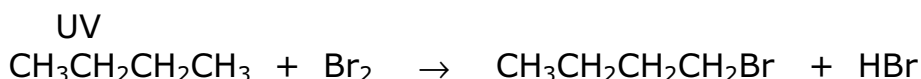
- Alkenes and alkynes also undergo combustion reactions.
- Alkenes and alkynes commonly undergo *addition* reactions that convert an unsaturated reactant into a saturated product. In these reactions, a molecule breaks down and is added across the multiple bond of the alkene (or alkyne) (**Chem12** p89, 93).
- Addition reactions of alkenes include reaction with (**Chem12** p89):
    - Halogens: addition of halogens, Br_2 and Cl_2, to alkenes are rapid and form dibromo- or dichloroalkanes.
    - Addition of water, H_2O, (as steam at a high temperature and pressure) in the presence of an acid catalyst results in two alcohols as products. The favoured product is the one in which the H atom adds to the carbon of the double bond that already carries the most H atoms (Markovnikoff's rule) (**Chem13** p117).

$$CH_3CH{=}CH_2 \;+\; H_2O/H^+ \;\rightarrow\; \underset{\substack{\text{propan-2-ol}\\ \textit{major product}}}{H_3C{-}\overset{\displaystyle H}{\underset{\displaystyle OH}{C}}{-}CH_3} \;+\; \underset{\substack{\text{propan-1-ol}\\ \textit{minor product}}}{H_3C{-}CH_2{-}CH_2{-}OH}$$

- Markovnikoff's rule also applies to the addition of other asymmetric molecules, such as HCl, to alkenes.
    - Hydrogen, H_2: only occurs at high temperatures in the presence of a catalyst (Ni or Pt) and produces an alkane.
- Addition reactions of alkynes include reaction with (**Chem12** p93):
    - Halogens: addition of one molecule of a halogen, such as Br_2 and Cl_2, forms a dihaloalkene which then adds another halogen molecule to produce a tetrahaloalkane.
    - Hydrogen, H_2: addition of two molecules of H_2 to an alkyne produces an alkane.
- The common test for an unsaturated hydrocarbon is therefore the rapid decolourisation of an orange solution of bromine to give a colourless solution when added to an alkene or alkyne.
- The redox reaction of alkenes with the permanganate ion, MnO_4^- in acid results in the purple permanganate ion being reduced to the colourless manganese ion, Mn^{2+}. This can be used to distinguish alkenes from alkanes since alkanes do not

react with potassium permanganate and the solution remains purple. In neutral solutions, the purple colour of the permanganate ion disappears as it is reduced and brown solid manganese dioxide, MnO_2, forms.

- Ethene, $CH_2=CH_2$, can be prepared by (**Chem12** p88):
  - Catalytic steam cracking of ethane to produce ethene and hydrogen gas.

$$CH_3CH_3 \quad \rightarrow \quad C_2H_4 \quad + \quad H_2$$

  - Dehydration of ethanol, CH_3CH_2OH, by heating with a dehydrating agent, such as concentrated sulfuric acid.

$$CH_3CH_2OH \quad \xrightarrow[\text{heat}]{\text{conc } H_2SO_4} \quad CH_2=CH_2 \quad + \quad H_2O$$

- Ethyne, C_2H_2, can be prepared by (**Chem12** p93):
  - Reacting solid calcium carbide, CaC_2, with water.

$$CaC_2(s) \quad + \quad 2H_2O(l) \rightarrow \quad C_2H_2(g) \quad + \quad Ca(OH)_2(aq)$$

Ethyne is also known as acetylene.

POLYMERISATION

Students will investigate and develop their scientific understanding of **polymerisation** *when they:*
- explain using structural formulae, the production of polyethene (polythene) from ethene and polyvinyl chloride (PVC) from vinyl chloride.
- list the uses of polythene and PVC.
- describe the environmental problems caused by polythene and PVC products.

Polymers are extremely large molecules that consist of many smaller repeating units. Synthetic polymers are used extensively in everyday life in the form of plastics, adhesives, paint and fabric.

KEY POINTS:
- Polymer molecules are very large molecules that result from the linking by covalent bonds of many smaller alkene molecules through addition reactions (**Chem12** p89, 90,94).
- The polymer is fully saturated and has no double bonds.

- Polyethene (also called polythene) is formed from many ethene monomers.
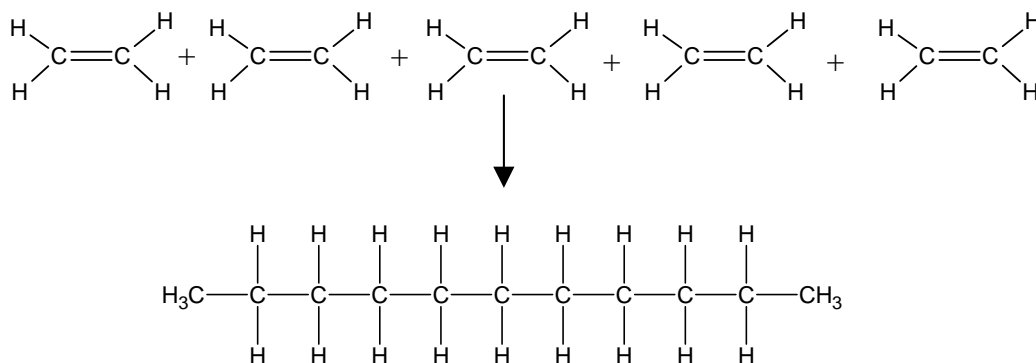
- Monomers of 1,2 dichloroethane (formed by the addition of chlorine to ethane) are used to prepare the polymer polyvinylchloride, PVC, which has a chlorine atom attached to every second carbon atom along the chain.
- Polyethene is used for making plastic bags, toys, bottles, outdoor furniture.
- Polyvinylchloride is used for making plastic wrap, garden hose, plastic plumbing pipes.
- Synthetic polymers such as polyethene and polyvinylchloride are not easily degraded and are very long-lived in the environment. This can create serious waste-disposal problems when plastic items are not disposed appropriately. Many plastic objects, such as shopping bags and food containers, can be recycled to produce other plastic items.

ALCOHOLS, KETONES, ALDEHYDES, CARBOXYLIC ACIDS, ESTERS, ALKYL HALIDES

For the vast majority of organic reactions, the organic compound reacts at the functional group. Many organic reactions change one functional group into another.

*Students will investigate and develop their scientific understanding of **alcohols, ketones** and **aldehydes** when they:*
- identify the functional groups for alcohols, ketones and aldehydes.
- write the molecular formula of C_1 - C_5 alcohols.
- use the IUPAC rules to name C_1 - C_5 alcohols, ketones and aldehydes and give examples of each.
- identify and name the three types of alcohol (primary, secondary, tertiary) and give examples of each.
- investigate the oxidation of the three types of alcohol using acidified $KMnO_4$ and acidified $K_2Cr_2O_7$.
- describe tests to differentiate between ketones and aldehydes.

KEY POINTS:

- **Alcohols** are a homologous series having a -OH functional group (**Chem12** p95; **Chem13** p115).
- The suffix for naming alcohols is –*anol*. When the carbon chain is more than 2 atoms long, a number indicates where the –OH group occurs on the carbon chain. The chain is numbered from the end giving the OH the lowest number.
- Small alcohols are liquids and highly soluble in water but become increasingly insoluble in water as carbon chain-length increases. Aqueous solutions are neutral.
- Alcohols are classified as primary, $1°$, secondary, $2°$, or tertiary, $3°$, depending on the number of C atoms that are directly attached to the C atom carrying the -OH group. Primary alcohols (apart from methanol) have the general formula RCH_2OH, secondary alcohols have the general formula R_2CHOH, tertiary alcohols have the general formula R_3COH (where R represents a carbon chain) (**Chem12** p96).
- Alcohols can undergo oxidation reactions with acidified potassium permanganate, $KMnO_4$ or acidified potassium dichromate, $K_2Cr_2O_7$. These produce aldehydes or ketones depending on whether the alcohol used in the oxidation reaction is primary or secondary (**Chem12** p96; **Chem13** p116).
- Primary alcohols (RCH_2OH) are oxidised to form aldehydes ($RCHO$), which are then easily oxidised further to form carboxylic acids (RCO_2H).
- Secondary alcohols (R_2CHOH) are oxidised to ketones (R_2CO) (**Chem13** p116).
- Tertiary alcohols do not react with oxidising agents (**Chem13** p117).
- Oxidation reactions can be used to distinguish tertiary alcohols from primary and secondary alcohols.
- When an acidified solution of potassium dichromate is used as a reagent in the oxidation reaction, the colour change observed is from the orange dichromate ion $Cr_2O_7^{2-}$ to the green Cr^{3+}.
- When an acidified solution of potassium permanganate is used as a reagent in the oxidation reaction, the colour change is from purple to colourless.
- **Aldehydes** (alkanals) are a homologous series having a functional group of a carbonyl group (C=O) bonded to a H atom at the end of a carbon chain and have the general formula RCHO (**Chem13** p118).

$$H_3C-CH_2-CH_2-C\overset{\displaystyle O}{\underset{\displaystyle H}{<}} \qquad \text{butanal}$$
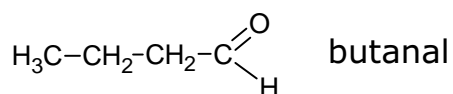
- The suffix for naming aldehydes is –*anal*. Numbering does not need to be used to show the location of the aldehyde group as it must always be on the end of the carbon chain.

- **Ketones** (alkanones) are a homologous series having a functional group of a carbonyl group (C=O) bonded to two other carbon atoms so that it occurs within the carbon chain and have the general formula RCOR' (**Chem13** p118).

$$H_3C-CH_2-CH_2-\underset{\underset{O}{\|}}{C}-CH_3$$ pentan-2-one
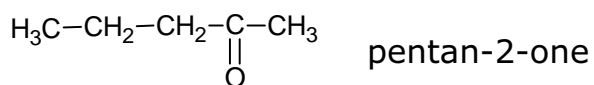
- The suffix for naming ketones is *–anone*. A number indicates the position of the carbonyl (C=O) group in the carbon chain (apart from propanone and butanone).
- Aldehydes undergo oxidation reaction but ketones do not and can be distinguished from aldehydes by observing whether a reaction with an oxidising agent occurs or not.
- Aldehydes can be distinguished from ketones by:
  - Addition of Tollens' reagent: A redox reaction occurs when a solution of $Ag(NH_3)_2^+$ (a colourless solution formed by adding excess dilute aqueous ammonia to a solution of aqueous silver nitrate) is heated with an aldehyde. This produces a layer of silver on the inside of a test tube as the aldehyde is oxidised to a carboxylic acid. No reaction occurs with a ketone or an alcohol (**Chem13** p118).
  - Addition of Benedict's reagent: A redox reaction occurs when an alkaline solution containing a copper(II) citrate complex ion (Benedict's reagent) is heated with an aldehyde and the blue Cu^{2+} complex ion is reduced to a brick red precipitate of Cu_2O. No reaction occurs with a ketone and the solution remains blue (**Chem13** p118).
  - Addition of Fehling's solution: A redox reaction occurs when an alkaline solution containing a copper(II) tartrate complex ion (Fehling's reagent) is heated with an aldehyde and the blue Cu^{2+} complex ion is reduced to a brick red precipitate of Cu_2O. No reaction occurs with a ketone and the solution remains blue (**Chem13** p118).

*Students will investigate and develop their scientific understanding of **carboxylic acids** and **esters** when they:*
- recognise the functional group of carboxylic acids and esters.
- write the molecular formula and structural formula of C_1 - C_5 carboxylic acids and use the IUPAC rules to name them.
- describe the acidic behaviour of ethanoic acid in water and write an equation to show its reaction with aqueous sodium hydroxide.
- describe a laboratory preparation of an ester and write an equation for its formation.
- list some common properties of esters.

KEY POINTS:

- **Carboxylic acids** (alkanoic acids) have a functional group of a carbonyl group (C=O) bonded to a -OH group at the end of a carbon chain and have the general formula RCOOH. The suffix for naming carboxylic acids is *–anoic acid* (**Chem12** p99; **Chem13** p118).
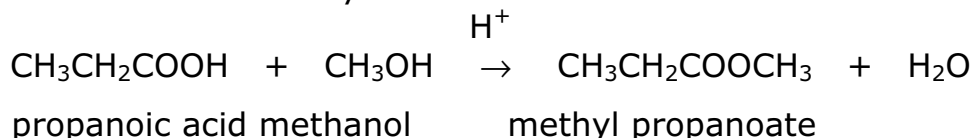
  ethanoic acid
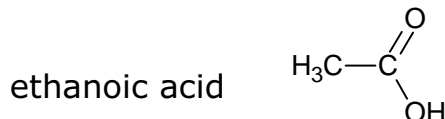
- Smaller carboxylic acids are soluble in water as the COOH group is polar, but become less soluble in water as the carbon chain increases in length and the acids become increasingly non-polar. Long-chain carboxylic acids are commonly called 'fatty acids' because of their insolubility in water and are an essential group of compounds found in all living cells.

- Carboxylic acids are weak acids in water (only slightly dissociate in water) and donate a proton to the water to produce a solution containing $H_3O^+(aq)$ ions. Aqueous solutions of carboxylic acids turn litmus red. Ethanoic acid gives vinegar its sour taste (**Chem12** p100).

- Carboxylic acids can be neutralised by reaction with a base such as aqueous NaOH.

$$CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COO^-(aq) + Na^+(aq) + H_2O(l)$$

- The anion formed when a proton is donated to water (or OH⁻) has the name ending *-anoate*.

- **Esters** are a homologous series that contain a functional group having a carbonyl group (C=O) bonded to an O atom in the middle of a carbon chain (**Chem12** p100; (**Chem13** p117).

  methyl methanoate
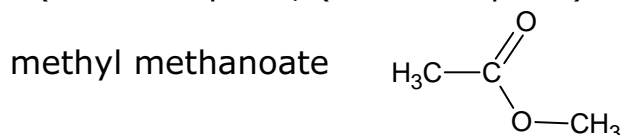
- Esters are formed by the acid catalysed reaction of an alcohol with a carboxylic acid. An acid such as concentrated H_2SO_4 can be used as the catalyst.

$$CH_3CH_2COOH + CH_3OH \xrightarrow{H^+} CH_3CH_2COOCH_3 + H_2O$$

 propanoic acid methanol   methyl propanoate

- Esters are named as *alkyl alkanoates*, where the alkyl part is derived from the alcohol, and the alkanoate from the carboxylic acid used to make the ester (**Chem12** p101; **Chem13** p122).

- Esters often have a fruity smell but this can vary and some are unpleasant. Many esters occur naturally and are used as flavourings and perfumes.

*Students will investigate and develop their scientific understanding of **alkyl halides** when they:*
- name simple primary alkyl halides and their formation from a reaction between hydrogen halides and alkenes.*

KEY POINTS:
- **Alkyl halides** (haloalkanes) are a homologous series having a halogen atom bonded to the carbon chain and having the general formula RX (X = F, Cl, Br, I) (**Chem13** p126, 127).
- Alkyl halides are named by adding the halogen as a prefix to the parent hydrocarbon name and numbering the C atom to which the halogen is attached. The chain is numbered from the end giving the halogen the lowest number.

2-chloro-4-methylpentane H_3C—HC—CH_2——CH—CH_2—CH_2—CH_3
                                                  Cl CH_3
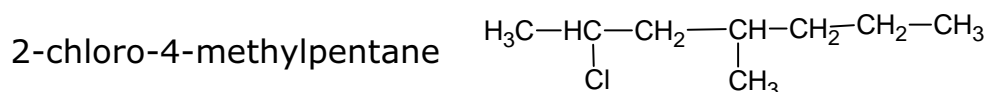
- Alkyl halides with more than 2 carbon atoms are relatively nonpolar overall and are insoluble in water as the alkane framework tends to dominate.
- Formed by the addition reaction of halogens, such as Br_2 and Cl_2, with alkenes.
- Addition of hydrogen halides, such as HCl and HBr, to alkenes results in the formation of two products, with the major product determined by Markovnikov's rule.


1-bromopropane (minor)

and

2-bromopropane (major)
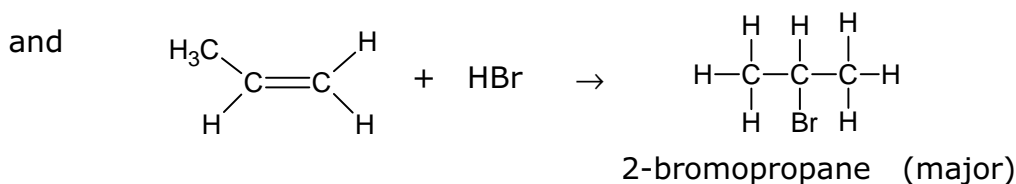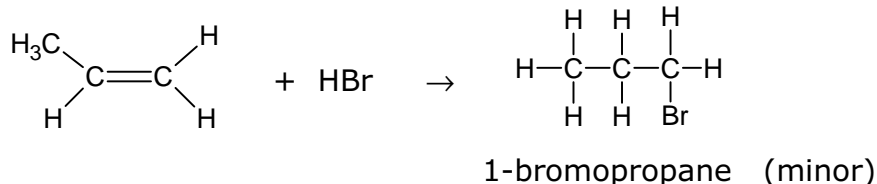
FATS AND OILS

*Students will investigate and develop their scientific understanding of **fats** and **oils** when they:*
- recognise that ester groups occur in molecules of fats and oils.
- draw the structures of triglycerides (esters composed of three fatty acids linked to glycerol) given the formulae of the constituent fatty (carboxylic) acids.
- relate the trend of rising melting points with increasing saturation of fats/oils (as utilised in the hydrogenation of oils).
- describe the extraction of fats/oils from their respective sources.
- describe how fats and oils are refined using alkali to remove free fatty acids and steam distillation to remove the undesirable volatile products.
- compare butter with margarine in terms of chain length, degree of unsaturation of fatty acids, and melting points.

KEY POINTS:

- Fats and oils (triesters or triglycerides) are esters that are produced from the reaction of glycerol (propane-1,2,3-triol) with three long-chain carboxylic acids (called fatty acids). Fatty acids are usually unbranched and have an even number of carbon atoms from 12-22 (**Chem12** p101; **Chem13** p121).
- The -OH group of the alcohol reacts with the acidic H of the carboxylic acid to form an ester linkage by eliminating water. Each of the three -OH groups present in glycerol can form an ester link with a different carboxylic acid.
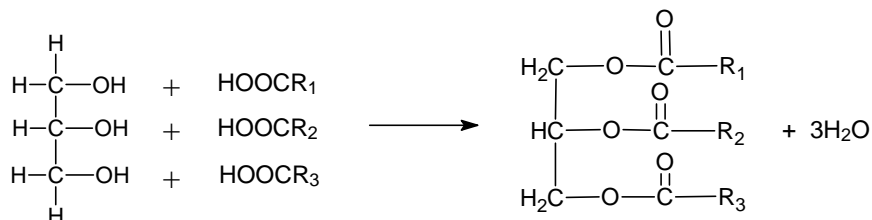
- Oils are liquids at room temperature and generally originate from plant sources. Oils are largely unsaturated (and are referred to as being polyunsaturated). Some seeds are very rich in oil, such as the coconut seed which is about 65% oil.
- Oils have a lower melting point than fats as because their fatty acids contain many *cis* C=C double bonds that cause bends and kinks in the hydrocarbon chains, making it difficult for the molecules to pack closely together to become a solid.
- Important vegetable oils include peanut, soya bean, sunflower, palm, olive, and coconut.
- Coconut oil is a colourless to pale brownish yellow oil with a melting point ranging from 23°C to 26°C. Coconut oil contains almost 95% saturated fatty acids that include approximately 50% lauric acid, $C_{11}H_{23}COOH$, as well as caprylic acid, $C_7H_{15}COOH$ (6-10%), capric acid, $C_9H_{19}COOH$ (5-10%), myristic acid, $C_{13}H_{27}COOH$ (15-23%) and palmitic acid, $C_{15}H_{31}COOH$ (6-11%).
- Oleic acid, $C_{17}H_{33}COOH$, is an unsaturated fatty acid and contains one C=C double bond. It makes up 85-95% of the acids from olive oil.
- Fats are solids at room temperature and originate from animal sources. Fats contain a high percentage of saturated fatty acids.
- Saturated fatty acids most commonly forming animal fats are palmitic acid, $C_{15}H_{31}COOH$, and stearic acid, $C_{17}H_{35}COOH$.
- Butter contains high proportions of the saturated fatty acids myrisitc acid, $C_{13}H_{27}COOH$, and stearic acid, $C_{17}H_{35}COOH$.
- Oils can be converted into solid fats by reducing the number of C=C double bonds in the fatty acid chains through hydrogenation.

- Hydrogenation involves heating the unsaturated oil at 170°C with a metal catalyst (Ni or Pd) in the presence of pressurized hydrogen gas.
- Margarine, a soft solid, is prepared so that only about two-thirds of the C=C double bonds present in the original vegetable oil (such as soya bean, sunflower or olive oil) are hydrogentated.
- Butter contains a ratio of unsaturated to saturated fats of 1:15, while margarine has a ratio of at least 2:1.
- Pure oils are colourless and odourless. The yellow colour and flavours of oils are contributed by natural materials carried along during the production of the oils from plants.
- Vegetable oils are usually extracted from crushed seeds, kernels and nuts using mechanical pressure or by solvents with low boiling points (such as ethanol) that are recovered by fractional distillation.
- Another method for extracting oil is steam distillation. As steam is passed through crushed parts of the plant, oils are released in a vaporized form. The vapour then passes through cooling tanks and condenses back to a liquid as pure oil.
- Crude fats and oils contain a number of substances that give the fat or oil an undesirable flavour or colour. These substances can be removed through a series of processing steps.
- Any free fatty acids in the fat or oil can be removed by treatment of the fats and oils with an alkali solution which converts the free fatty acids into water soluble soaps that are then washed away by water.
- The process of vacuum steam distillation is used to remove undesirable flavours and odours that mostly arise from oxidation in fats and oils. Using steam under reduced pressure, the volatile compounds can be removed from fats and oils.
- Small quantities of fats and oils are an essential part of a healthy diet as they supply nutrients, aid in the absorption of vitamins, and provide a source of energy for the body.

Coconut milk is approximately 24% fat and is made by processing grated coconut flesh with hot water which extracts the oil and compounds that give it flavour.
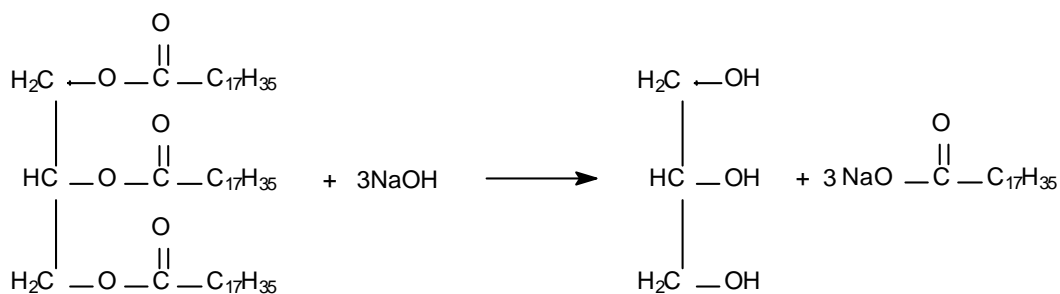
SOAP

*Students will investigate and develop their scientific understanding of **soaps** when they:*
- draw the structural formula for a soap.
- outline the process of soap manufacture.
- describe how the chemical nature of soap promotes its role as a cleaning agent.
- investigate the behaviour of soap in hard and soft water.
- describe the chemical processes involved in softening hard water.
- compare the general structure of common synthetic detergents (e.g. dishwashing liquids) with that of soap and relate this to the behaviour of each in hard water.

KEY POINTS:
- Soap is made by heating animal fat in a solution of sodium hydroxide, NaOH(*aq*). The resulting hydrolysis reaction breaks the three ester links present in the fat molecule and produces glycerol and the sodium salts of the fatty acids (typically sodium stearate, $C_{17}H_{35}COO^-Na^+$, also called sodium octadecanoate) (**Chem12** p101).

- This reaction is referred to as "saponification" because sodium stearate releases sodium ions and stearate ions when dissolved in water and is the cleaning agent in soap.
- Oils are not soluble in water and can only be washed away in water when soap is present.
- Soaps work because the two ends of a soap molecule are different. The long non-polar hydrocarbon end of the stearate ion dissolves in the thin film of oil on dirty fabric or on the skin while the ionic polar carboxylate end readily dissolves in water (which is also polar). Because of these opposing tendencies, soap molecules are attracted to both grease and water. The stearate ions are able to completely surround the oil, forming tiny droplets suspended in water that can be rinsed away.
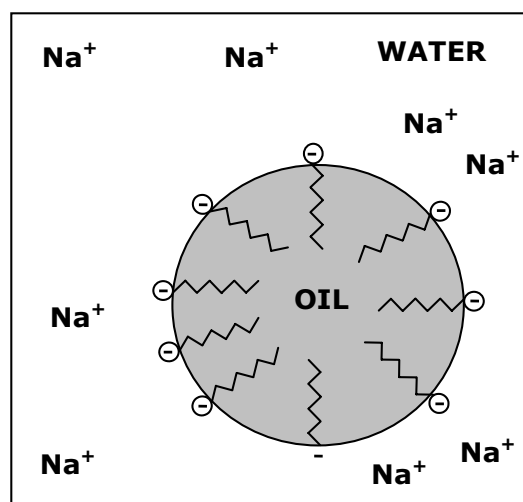
- Water is described as being "hard" if it forms a solid "scum" that cannot be washed away when soap is added. This is due to the presence of moderate amounts of calcium ions, Ca^{2+}, and magnesium ions, Mg^{2+}, in water. When rain water flows over rocks containing calcium or magnesium minerals, $Ca^{2+}(aq)$ and $Mg^{2+}(aq)$ ions dissolve in the water. These react with stearate ions to produce insoluble precipitates in water.
- Soap will not produce a lather (become frothy) and clean in hard water. The $Ca^{2+}(aq)$ and $Mg^{2+}(aq)$ ions are removed from hard water by precipitation or the use of ion-exchange resin which replaces $Ca^{2+}(aq)$ and $Mg^{2+}(aq)$ ions in the water for $Na^+(aq)$ ions.
- Water that contains almost no calcium or magnesium ions is referred to as "soft" water.
- Synthetic detergents do not form scum in hard water as the calcium and magnesium salts of detergents are water-soluble. Like soaps, synthetic detergents have water-liking (hydrophilic) heads and water-disliking (hydrophobic) tails that enable detergents to clean in the same way as soap. The hydrophilic heads may be anionic, cationic or non-ionic. Many laundry products and fabric softeners contain detergents. An example of a synthetic detergent is sodium dodecylbenzenesulfonate (shown below).

$$Na^{+} {}^{-}O-\overset{\overset{O}{\|}}{\underset{\underset{O}{\|}}{S}}-\!\!\left\langle\!\!\!\bigcirc\!\!\!\right\rangle\!\!-CH_2(CH_2)_{10}CH_3$$
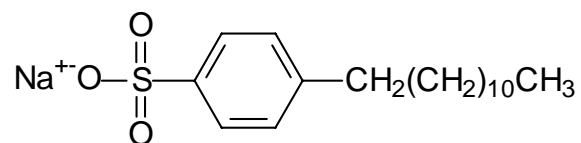
- Detergents can pose an environmental problem as some detergents are not easily decomposed by bacteria in natural water so that foam and suds build up in rivers and streams.

CARBOHYDRATES

Carbohydrates are synthesised by plants and stored in the form of starch, a polymer of simple sugars. When starch is eaten and digested, the glucose that is released provides a major source of energy for living organisms.

Students will investigate and develop their scientific understanding of **carbohydrates** *when they:*
- draw the open and cyclic structures of glucose and be aware of their inter-conversion.
- use Fehlings or Benedicts reagent to test for the presence of the aldehyde group in an open glucose structure.
- draw the structure of a disaccharide (maltose and sucrose) and a polysaccharide (e.g. starch in G form).
- identify the differences in structure and property of two types of starch (linear and branched).

KEY POINTS:
- Carbohydrates are synthesised by plants and stored in the form of starch.
- The body's major use of digestible carbohydrates is to provide energy.
- Monosaccharides (also called simple sugars) are the simplest carbohydrates and cannot be broken down into smaller units by hydrolysis. They have from 3-7 carbon atoms, and each has one aldehyde or one ketone functional group (**Chem13** p130).
- Monosaccharides react with each other to form disaccharides and polysaccharides (polymers of monosaccharides that are also called complex carbohydrates). Disaccharides and polysaccharides can be broken down (hydrolysed) to give monosaccharides.
- Monosaccharides and disaccharides are generally high-melting, white, crystalline solids that are soluble in water due to their ability to form hydrogen bonds with water.
- Monosaccharides and disaccharides are sweet-tasting and digestible.
- Glucose is the most widely-occurring monosaccharide, found in honey and fruits. It is a source of energy for living organisms. The oxidation reaction of glucose with oxygen produces CO_2 and H_2O as products as well as giving out heat that the body uses for muscle contraction, nerve transmission and tissue building and repair.
- Glucose is the monomer for starch and cellulose (a fibrous substance that provides structure for plants).
- Glucose can exist either as an open chain or as in a ring form. Crystalline glucose is entirely in the ring form but once dissolved in water, equilibrium is established between the ring and open-chain forms. The ring form is created when the oxygen on carbon number 5 links with the carbon comprising

the carbonyl group (carbon number 1) and transfers its hydrogen to the carbonyl oxygen to create a hydroxyl group (**Chem13** p130).
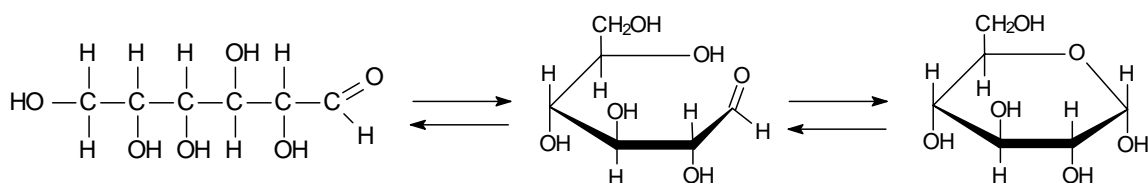

- As the open-chain form of glucose has an aldehyde group at the end of the carbon chain, Benedict's reagent or Fehling's reagent can be used as oxidising agents to convert the aldehyde to a carboxylic acid. This results in a colour change from the blue $Cu^{2+}(aq)$ ion to a brick-red precipitate of copper(I) oxide, CuO.
- Disaccharides are formed when two monosaccharide molecules are linked together. The monosaccharide units can be the same or different, resulting in a range of disaccharides.
- The disaccharide maltose is present in fermenting grains and produced in the body during starch digestion. It is used as a sweetener in foods.
- Maltose is made of two glucose rings linked together. Since the rings cannot open to chains, maltose does not react with oxidising agents.
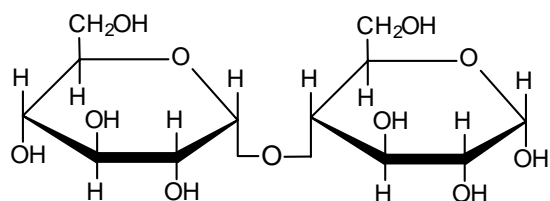
maltose

- Sucrose is ordinary table sugar that has been refined from sugar cane or sugar beets (although it is found in many plants). It is commonly used as a sweetener in foods.
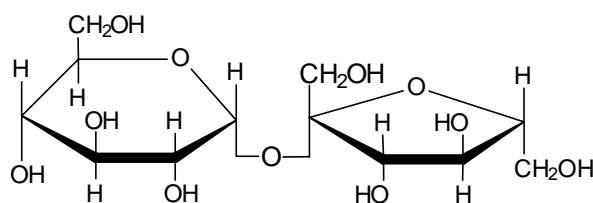

sucrose

- Sucrose is formed by one molecule of glucose linking with one molecule of fructose (a 5-carbon ketone sugar).Enzymes in the body break sucrose down into glucose and fructose molecules. Honeybees convert the sucrose in the nectar of flowers into glucose and fructose (**Chem13** p130).
- Polysaccharides are polymers of many monosaccharides (simple sugars) linked together. Polysaccharides can be formed by tens, hundreds, or thousands of monosaccharide monomers.
- Polysaccharides, unlike monosaccharides and disaccharides, are insoluble in water.
- Starch is a polymer of glucose and the major form of stored carbohydrate in plants. Starch is fully digestible and an essential part of the human diet as it is broken down to provide glucose for the body. It is found in grains, such as wheat and
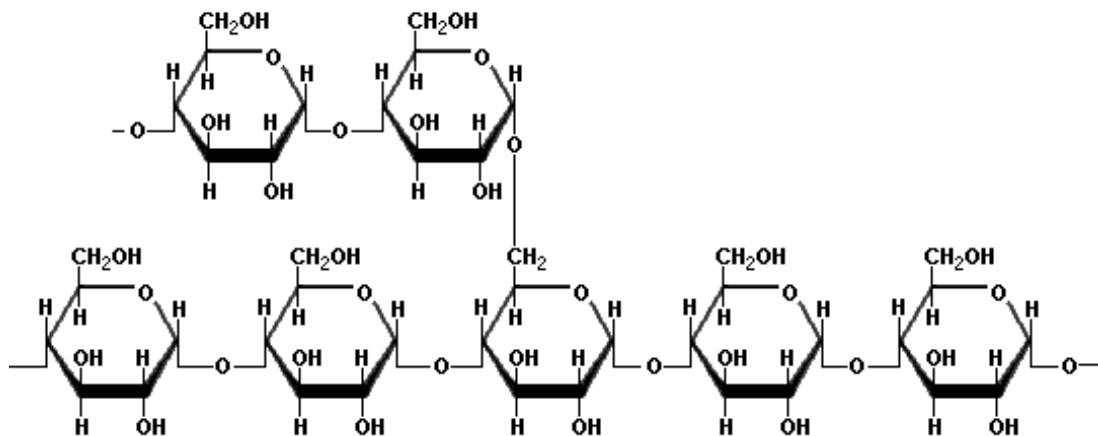
rice, root vegetables, such as potatoes, and legumes, such as peas and beans.

- There are two kinds of starch: amylose, an essentially linear polysaccharide and amylopectin, a highly branched polysaccharide. Natural starches contain 10-20% amylose and 80-90% amylopectin.
- Amylose molecules consist typically of 200 to 20,000 glucose units linked into long chains which tend to coil into helices. Amylose is partially soluble in hot water and makes the water cloudy when potatoes are boiled in water.
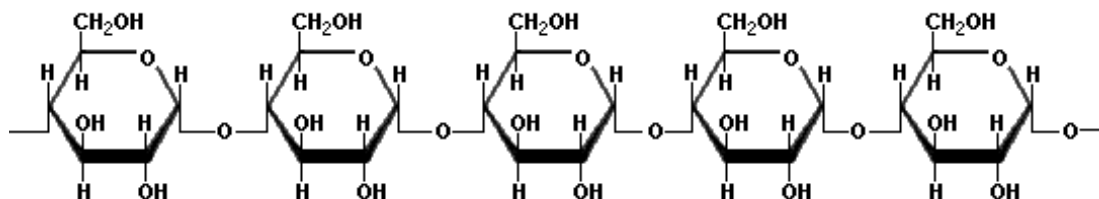
Amylose

- Amylopectin is similar to amylase but has much larger molecules and is highly branched, with short side chains of about 30 glucose units attached approximately every 25 glucose units along the chain. Amylopectin molecules may contain up to two million glucose units. Amylopectin is not water soluble.

Amylopectin