Chemistry Learning Guide

Year 12

Student Learning Guide

Samoa Chemistry Curriculum Year 12

Prepared by: Dr David Salter
Department of Chemistry
University of Auckland
Auckland
NEW ZEALAND

Produced by: Mahobe Resources (NZ) Ltd

Funding provided by NZAID under the Samoa Official Development Assistance Programme

Project Contractor: Auckland UniServices Ltd.

CONTENTS

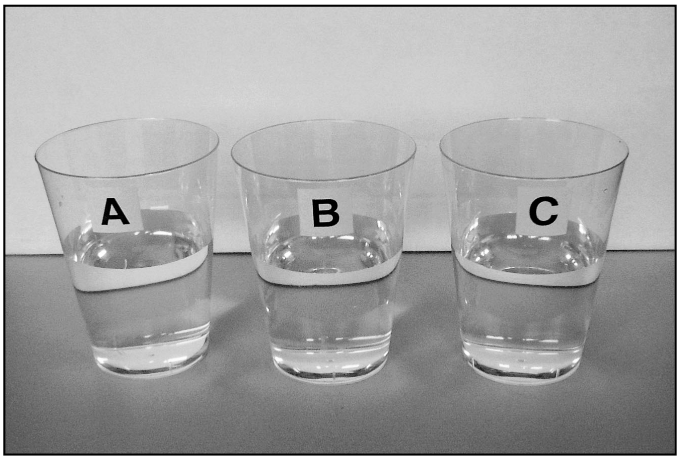
INTRODUCTION

Why *know* chemistry? (and NOT why *no* chemistry?)

THINK: Why is it useful to learn chemistry?

Suppose you were thirsty and know that water is safe to drink. How would you determine if a particular liquid is water?

There are three glasses (labelled "A", "B" and "C") in front of you. Each glass has a clear, colourless liquid in it but only one of them contains pure water. Applying chemistry can help you in deciding which glass contains the water.



When some powdered limestone (calcium carbonate) is added to each glass, liquid "A" starts to fizz but liquids "B" and "C" do not. When a lighted match is placed on a few drops of each liquid, liquid "B" starts to burn with an orange flame, but liquids "A" and "C" extinguish the match. From these experiments, chemists would classify liquid "A" as an acid and liquid "B" as a flammable substance – and both would be hazardous to drink!

Further analysis would enable chemical symbols to be given to each liquid. Liquid "A" is labeled as "dilute HCl"; liquid "B" as "CH_3CH_2OH" and liquid "C" as "H_2O". From knowing the language of chemistry and how to interpret chemical symbols, liquid "A" is identified as hydrochloric acid, liquid "B" as ethanol (an alcohol), and liquid "C" as water.

From studying the properties of these substances, liquid "C" is water and the only pure liquid that would be safe to drink. Using chemical knowledge allows you to better understand the world around you and make better informed decisions.

UNDERSTANDING CHEMISTRY

Chemistry is the study of matter and its properties and how many substances can be beneficial to our lives. Sometimes people find chemistry a difficult subject to learn. However, once the way chemistry fits together is understood, chemistry is straightforward and provides a very useful way for people to explain the properties of every substance and predict how substances may change in different conditions. Knowing some of the "big ideas" of chemistry, such as the difference between a chemical change (such as when metal rusts) and a physical change (such as when ice turns into liquid water), can be valuable for interpreting what happens in the world around us.

People have studied and recorded the characteristics of an enormous number of different substances over the last several hundred years. This has provided a large amount of information that has enabled chemists to identify patterns and group substances into classes and families based on similar observable properties. These families are given names, such as "elements", "compounds", "metals", "non-metals", and so on. Learning chemistry involves learning some of these patterns and family properties. For example, all metals have a shiny appearance and conduct electricity.

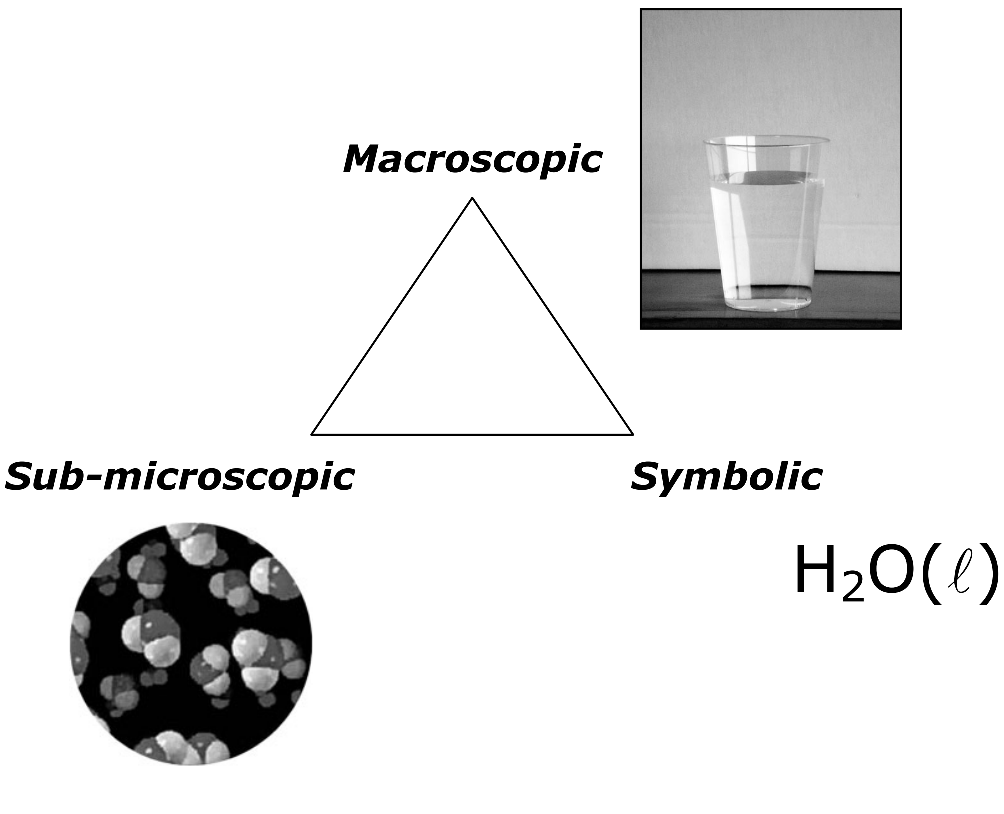
Learning chemistry also involves learning the special language and symbols that chemists use to describe substances and changes they undergo. One of the most useful aids to learning chemistry is the periodic table, as it summaries a lot of information about the substances known as elements. Learning how to correctly interpret the information shown on the periodic table is an important part of learning chemistry.

Chemistry can be regarded as having several interrelated components. Chemists classify or describe a substance according to observable properties in samples large enough to see (and sometimes taste and smell), such as colour, density, melting point and boiling point, and reactivity. This is the *macroscopic* aspect of chemistry.

In order to explain many of these observations, chemists have had to create images or mind pictures of what happens at a level that is too small to be seen. This involves thinking about substances being made up of particles such as atoms, molecules, and ions. Since atoms cannot be seen, chemists have to imagine what atoms and molecules look like. This is the "sub-microscopic" view of chemistry as "seeing" this level of detail is beyond the power of a microscope. These particles are often represented by pictures showing coloured spheres in certain arrangements. However, these are only models used to help chemists explain observations made of the macroscopic world.

Chemists also use symbols, formulae, equations, diagrams, and graphs to convey information. For example, atoms of an element are represented by one- or two-letter symbols based on the alphabet. A series of these symbols can be combined into formulae to represent other substances. Another key to chemistry is learning which symbol (or formula) stands for a particular substance. Also, rather than writing a long description of the many changes that substances undergo, chemical equations are used as a shorthand way of showing these changes. For example, an arrow ($\rightarrow$) in an equation is read as "turns into". Chemists have also created special words and language that have a particular meaning, such as the word "strong". This component of chemistry is referred to as the "symbolic" world.

The relationships between these components of chemistry are shown in the figure below. This shows how a glass of water has characteristics that are visible with the naked eye (macroscopic), as a clear, colourless liquid at room temperature. To explain water's observable properties, chemists rely on a mental model of water molecules which exist at a sub-microscopic level. As a shorthand way of referring to water, the symbol "H_2O" is used to represent water.



Experienced chemists often think about a substance in terms of these three levels all at the same time. For example, a chemist seeing a piece of black coal would call it "carbon", write the symbol "C" and imagine in their mind how spheres representing the carbon atoms are arranged to form the lump of coal. Everyone can learn to do this with practice and become an experienced chemist.

USING THE LEARNING GUIDE

Chemistry is involved in some way in almost everything in our daily lives, from the food we eat, to the air we breathe, to the clothes we wear and to the cars we drive. The concepts introduced in this course should help you develop a greater appreciation of how chemistry impacts our lives. On completing this course, you should be able to describe the properties of substances and explain how substances change using appropriate chemical ideas and language.

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 objectives set for you are listed in each section. The main points in each section and the page references in textbooks are also included. More extensive notes are provided here for those concepts that are not covered in detail by the textbooks. The key to learning chemistry is to always think about a substance in terms of its macroscopic properties, the particles making it up, and the symbol used to describe it.

The following textbooks are referred to in the notes below:

- *The Material World: Level Six Science* by S. Haigh (1995), Auckland: New House Publishers Ltd. (abbreviated as **MatW**).
- *Chemistry Year 12 – Pathfinder Series* by M. Croucher and P. Croucher (2000), Auckland: New House Publishers Ltd. (abbreviated as **Chem12**).

S T R A N D 1: THE STRUCTURE OF ATOMS

ATOMIC STRUCTURE

Students will investigate and develop their scientific understanding of **atomic structure** *when they:*

- Describe the model of an atom and the structures that make up an atom.
- Identify atomic structure (protons, neutrons, electrons, isotopes) of a given atom or ion.
- Explain how an atom changes to an ion and the combining power of an atom that causes that change.
- Describe the formation of an ion.
- Draw Lewis structures of simple molecules.
- Describe the shape of simple molecules.
- Describe the formation of ionic bonding, covalent (both polar and non-polar), and metallic bonding.

The observed properties of a macroscopic sample of a substance can be explained using the idea that matter is made up of atoms. Atoms and the particles which form them are too small to see. Therefore, chemists have created a model (a mental image or picture in the mind) of what an atom "looks like" based on properties of the atom.

KEY POINTS:

- Atoms are composed of smaller particles: protons, neutrons and electrons. The protons and neutrons are located in a small region at the centre of the atom, called the nucleus. The *atomic number* of an atom is the number of protons, as well as the number of electrons in that atom. (**MatW** p24-25; **Chem12** p4)
- In one model of the atom, electrons are in circular orbits, while another more advanced model has electrons in a cloud surrounding the nucleus (**MatW** p26-27; **Chem12** p4-5)
- Each element consists of atoms unique to that element (**Chem12** p8). Elements are arranged in the periodic table in order of increasing atomic number. An element can have *isotopes* (**MatW** p25; **Chem12** p36)
- Atoms have no electrical charge. They can gain or lose electrons in their reactions, forming electrically charged ions (**MatW** p30-31; **Chem12** p4,8)
- Chemical bonds link atoms or ions together. Depending on the nature of the atoms, the bonds are classified as covalent, ionic, or metallic (**MatW** p32-33; **Chem12** p7-8)
- Molecules are separate groups of 2 or more atoms joined by chemical bonds. A compound occurs when the atoms are different (**Chem12** p11)
- 2-Dimensional diagrams can be drawn showing how atoms are bonded together in molecules. These are called Lewis structures. The 3-dimensional shape of a molecule can be deduced from its Lewis structure (**Chem12** p7,9,12-13)

- The strength of attractions between molecules in a substance determines many of its physical properties (**Chem12** p15-18).

QUESTIONS

1. What is the key difference between an element and a compound?
2. What is the key difference between a mixture and a compound? (See **MatW** p19).
3. How can compounds have no overall charge if they consist of positive and negative ions?
4. Why is it useful to arrange the elements in the form of the periodic table? (See **MatW** p36-37).

Storyline

The toxicity of copper ions: Death to fungi

Fungi can have a devastating effect on crops. This was evidenced in 1990 in Samoa when high humidity levels contributed to a blight of leaf fungus (called "taro wilt") on taro crops. Within one year, most taro crops had been affected and crop production almost stopped and few taros were available at the markets. Prior to 1990, taro (a plant very high in carbohydrates) was a major export to New Zealand, American Samoa and other countries in the Pacific. Since 1990, the use of a fungicide on taro crops has been needed to control and reduce the spread of the fungus. As a result, taro farming is now more labour intensive and expensive because extensive monitoring for the presence of taro wilt is required as well as the purchase of a fungicide.

A common fungicide used on taro crops is "Cusol". One litre of this product contains the equivalent of 400 grams of dissolved copper sulfate, which is sprayed onto the leaves of a variety of vegetable and fruit crops in Samoa, including taro, lettuce, avocadoes and passionfruit. This controls fungal diseases such as mildew and taro wilt. The copper ions are absorbed by the fungi and build up in their cells. Copper ions are toxic as they interfere with important chemical processes that occur in living cells and long-term exposure to them eventually results in an organism's death. This is why extreme care is needed when applying fungicides, including the use of gloves and breathing apparatus, so that people do not inhale the poison or absorb it through the skin.

S T R A N D 2: QUANTITATIVE AND PHYSICAL CHEMISTRY

QUANTITATIVE CHEMISTRY

Students will investigate and develop their scientific understanding of **quantitative chemistry** *when they:*

- Define the terms mole, relative atomic mass, relative molecular mass and molar mass.
- Make a connection between mass and moles.
- Study the "law of conservation of mass" which leads to the balancing of simple equations.
- Carry out simple calculations based on above connections.
- Calculate percentage composition of elements in compounds.
- Deduce empirical formula.

Chemical equations show the relative amounts of atoms involved in a chemical change. Chemists need to know how many atoms, molecules or ions are present in a certain mass of a substance so that, for instance, when carrying out a reaction there is a sufficient number of particles present to react completely. Because atoms and molecules are far too small to count individually, chemists have devised a way to determine the number of particles present in a pure substance once its mass and chemical formula is known.

KEY POINTS:

- The amounts of substances are measured in chemistry by the "mole".
- **1 mole** is the amount of substance that contains the same number of particles as there are atoms in exactly 12.00 g of carbon-12 (experimentally found to be 6.023×10^{23} – an enormous number!) (**Chem12** p35)
- The mass of any atom is defined relative to the mass of an atom of carbon-12 (which is 12.00). This is referred to as its "relative atomic mass". The "relative molecular mass" of a molecule is the sum of the relative atomic masses of its atoms (**Chem12** p36)
- The *molar mass* is the mass in grams of one mole of particles and has the unit grams per mole (g mol^{-1}). The mass in grams of one mole of any atom is the **same number** as the relative atomic mass of that atom (usually shown beneath the symbol on the periodic table) (**Chem12** p35)
- For a particular substance, the amount in moles and mass in grams are related by the molar mass. Since mass can be measured using a balance, the amount of substance in moles and the number of particles present can be calculated (**Chem12** p36)

- Chemical equations show the reactants (starting materials) and products of a chemical change. Since no atoms are lost or gained in a chemical reaction, the same number and type of atoms must appear on both sides of the chemical equation for it to be balanced (**MatW** p34-35; **Chem12** p38-39)
- The *percentage composition* of a compound shows how much each element contributes to the total mass of the compound (**Chem12** p41)
- The *empirical formula* shows the simplest whole number ratio of atoms in a compound. The *molecular formula* shows the actual number of atoms present in a molecule (**Chem12** p41-42).

QUESTIONS:

1. What amount in moles of C atoms is in 1 mole of sucrose, $C_{12}H_{22}O_{11}$? How many C atoms are in 1 mole of sucrose?
2. How does a balanced equation apply the law of conservation of mass? (See **MatW** p34).
3. Calculate the percentage composition of propane, C_3H_8.

PHYSICAL CHEMISTRY

Students will investigate and develop their scientific understanding of **physical chemistry** *when they:*

- Describe exothermic and endothermic reactions.
- Investigate the factors affecting the rate of reactions.
- Discuss a simple introduction to equilibrium with reference to reversible reactions.

Chemical and physical change is accompanied by changes in the energy content of the substances involved. For example, when wood burns in air, energy is released as heat and light. Energy from the sun is absorbed when ice melts in sunlight. The energy changes accompanying a process, the speed at which the process occurs, and the extent to which reactants are converted to products are all important aspects of chemical reactivity.

KEY POINTS:

- For a chemical reaction to occur, particles must collide. Chemical reactions are faster if the particles collide more frequently or with greater force (**MatW** p48-49; **Chem12** p45-46)
- Reaction rate measures the decrease in amounts of reactants or the increase in amounts of products per unit time (**MatW** p50-51; **Chem12** p38-39)

- Reaction rate can be increased by:
  - increasing the surface area of solid reactants (**MatW** p44-45);
  - increasing the concentration of reactants (**MatW** p46);
  - increasing the temperature of the reaction mixture (**MatW** p47);
  - adding a catalyst (**MatW** p52-52; **Chem12** p46-47).
- Changes in the temperature of the surroundings indicate that the energy content of reactants and products differs.
- Reactions that release energy to the surroundings (which gets warmer) are called "exothermic" reactions and the products have lower energy content than the reactants (**Chem12** p49).
- Reactions that absorb energy from the surroundings (which gets colder) are called "endothermic" reactions and the products have higher energy content than the reactants (**Chem12** p50).
- Every reaction is reversible if all the reactants and products are kept in contact with one another. 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. The relative amounts of reactants and products at this point depend on the particular reaction taking place (**Chem12** p53).

QUESTIONS:

1. What aspects of a chemical reaction can be measured over time to obtain the reaction rate?
2. A reaction is carried out in water as the solvent. Explain how the addition of more water affects the rate of reaction.
3. What is a catalyst?
4. Classify the reaction occurring when wood is burned as an exothermic or endothermic process.
5. When a chemical company starts to use a new reaction to manufacture a product, the chemists consider the rate of the reaction and the yield of product at equilibrium. How does each of these factors affect the usefulness of the manufacturing process?

S T R A N D 3: CHEMICAL REACTIONS

An enormous number of chemical reactions occur in the world around us. Chemists have found that these are easier to understand if reactions having a common underlying chemical process are grouped together. Three important groups of reactions are redox reactions, acid-base reactions and precipitation reactions.

REDOX REACTIONS.

Students will investigate and develop their scientific understanding of **redox reactions** *when they:*

- Define oxidation and reduction in terms of transfer of oxygen, electrons and change in oxidation state.
- Work out the oxidation state of atoms and ions.
- Write half equations to describe redox reactions, which lead to balanced equations.
- Describe the action of oxidising agents and reducing agents.

Any chemical reaction that has an element as a reactant or a product is a redox (oxidation-reduction) reaction. For example, the reactions that generate electricity in batteries are redox reactions.

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.
- Both oxidation and reduction occur in every redox reaction.
- A useful method for identifying the atom (or ion) that has lost electrons and the atom (or ion) that has gained electrons in a redox reaction is to assign *oxidation numbers* to atoms using a set of rules (**Chem12** p71-72). Oxidation has occurred if the oxidation number of an atom increases. Reduction has occurred if the oxidation number of an atom decreases.
- Oxidising agents (or oxidants) are substances that **gain** electrons in redox reactions (**Chem12** p71,74).
- Reducing agents (or reductants) are substances that **lose** electrons in redox reactions (**Chem12** p71,74).
- The number of electrons lost by the reducing agent in a redox process must equal the number of electrons gained by the oxidising agent (**Chem12** p73-74).
- The addition of oxygen to an element, a compound or an ion is also referred to as oxidation (**MatW** p87; **Chem12** p71).
- The removal of oxygen from a compound or polyatomic ion is also referred to as reduction (**MatW** p87; **Chem12** p71).

QUESTIONS:

1. Explain why every redox reaction must involve an oxidising agent as well as a reducing agent.
2. Determine the oxidation number of sulfur in the following:
   (i) SO_2 (ii) SO_3 (iii) H_2SO_4 (iv) H_2S
3. Explain whether burning wood in air is a redox reaction or not.

ACID-BASE REACTIONS

Students will investigate and develop their scientific understanding of **acid-base reactions** *when they:*

- Define an acid as a proton donor and a base as a proton acceptor.
- Outline the chemical and physical properties of acids and bases.
- Carry out an investigation of the chemical reactions of acids with metals, bases and carbonates.
- Carry out simple tests for acids and bases using common indicators.
- Use the pH scale to classify substances as acidic, neutral, or basic.
- Explain the difference between strong and weak acids and bases.

Acid-base reactions occurring in aqueous solutions are important reactions in living organisms as well as in many industrial processes. 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.

KEY POINTS:

- An acid is a substance that reacts with water to produce $H^+(aq)$ ions. More generally, an acid is a proton (H^+) donor. (**MatW** p39; **Chem12** p57)
- A base is a substance that produces $OH^-(aq)$ ions when dissolved in water. More generally, a base is a proton (H^+) acceptor. (**MatW** p40; **Chem12** p57)
- The colour of substances known as indicators depends on the acidity or basicity of the solution in which they are dissolved. For example, acids turn blue litmus red and bases turn red litmus blue. (**MatW** p42)
- The pH scale is a convenient way of indicating the quantity of $H^+(aq)$ ions present in a solution. pH is defined as the negative logarithm (log) of the concentration of $H^+(aq)$. A neutral solution has a pH value of 7. The pH of an acidic solution is less than 7. The pH of a basic solution is greater than 7. (**MatW** p42; **Chem12** p59)
- Strong acids and strong bases react to the extent of 100% with water. (**Chem12** p58)
- Weak acids and weak bases react with water to less than 10% (and most of their molecules remain intact). (**Chem12** p58)

- A strong acid and a strong base react to form water and an ionic metal compound (called a "salt"). (**MatW** p41; **Chem12** p58)

- Acids react with certain metals to form hydrogen gas as one product and react with metal carbonate compounds to form carbon dioxide gas as one product. (**MatW** p38; **Chem12** p31)

QUESTIONS:

1. Explain why the water insoluble compound calcium carbonate (limestone) dissolves when mixed with hydrochloric acid solution.
2. What is meant by the words "strong" and "weak" as applied to acids and bases?
3. What does pH measure?

PRECIPITATION REACTIONS

Students will investigate and develop their scientific understanding of **precipitation reactions** *when they:*

- Demonstrate what happens to a compound that dissolves in water.
- Investigate solubility properties of chlorides, sulfates, nitrates, carbonates and hydroxides of metals.
- Use the deduced solubility rules to predict the name of the formed precipitate and its colour.
- Carry out the tests for chosen anions.
- Carry out the tests for cations.

In precipitation reactions in water, two soluble ionic compounds react to form an insoluble ionic product, known as a *precipitate*. Coral reefs and mineral deposits are the result of this type of chemical reaction.

KEY POINTS:

- Water soluble ionic solids exist in aqueous solution as hydrated positive and negative ions. Each ion is surrounded by several polar water molecules. (**Chem12** p15,50)
- Precipitation of an ionic solid occurs on mixing solutions containing hydrated ions if the attraction between the positive and negative ions is stronger than the attraction of either of the ions for water molecules. (**MatW** p84)
- Precipitation reactions can be represented by equations showing only those ions that react to form a precipitate. Other ions that remain dissolved in the solution and not react (called "spectator ions") are not shown. (**Chem12** p33)
- Chemists have studied many precipitation reactions and have generated a set of solubility rules that help in predicting whether a precipitate will be observed on mixing two solutions

of ionic compounds as well as in identifying the precipitate. (**MatW** p85)

- Test-tube scale precipitation reactions can be used to confirm the presence of particular ions in aqueous solution. (**Chem12** p31-33)

QUESTIONS:

1. What types of substances are likely to be soluble in water?
2. Why are some ionic compounds soluble in water and others are not?
3. Which ions do not appear in an equation representing a precipitation reaction?

Storyline

Coral: Chemistry in the ocean

Shallow, warm tropical seas do not support much marine life. However, corals can cover large areas close to the seashore. Coral reefs are formed by tiny invertebrate animals that inhabit a hard coral structure. These animals, called polyps, resemble anemones or jellyfish and have short tentacles with stinging cells used for catching microscopic animals in the seawater. Coral animals also host single-celled algae (called zooxanthellae) in their tissues which produce simple sugars by photosynthesis as food for themselves and for the coral. Therefore, corals do not have to catch much of their own food and can survive in tropical waters that contain little food supplies.

Coral is a colony of thousands of polyps that inhabit the thin outer layer of coral, while a hard inner skeleton, called the corallite, gives structural support and protection for the animals to withdraw into when disturbed or when the coral is exposed at low tide. The hard skeleton consists of almost pure water-insoluble calcium carbonate, $CaCO_3$. While precipitation of calcium carbonate in seawater is rare, this substance is used by many marine animals for shells using biologically-mediated calcification.

The coral polyps absorb calcium ions, $Ca^{2+}(aq)$, from seawater and convert carbon dioxide from cell respiration into hydrogen carbonate ions, $HCO_3^-(aq)$. These ions are secreted by the coral animals under the skin around the lower half of their bodies in a substance that hardens to form layers of calcium carbonate.

$$Ca^{2+}(aq) + 2HCO_3^-(aq) \rightarrow CaCO_3(s) + CO_2(aq) + H_2O(\ell)$$

The coral continually adds calcium carbonate to its skeleton so that it extends upward, ensuring that light reaches its zooxanthellae. When old polyps die, they become part of the skeleton formation and new polyps continue to build on the existing permanently attached coral structures as a base. This results in coral forming a variety of structures, including branches, tubes, and fan shapes. Calcification is a slow process and, although the rate of coral growth varies depending upon conditions, it can be around 0.5 cm per year. Because coral is colonial, huge amounts of calcium carbonate may be deposited onto a reef which may exist for thousands of years as a base for new coral to grow on over time.

Coral reefs become bleached and turn pure white when clorox bleach is dumped onto the reef by fishermen to kill fish. However, everything in that area is also killed, including the coral, so that only the insoluble calcium carbonate frame of the coral remains and the reef dies.

The coral found on beaches is old, dead pieces of coral that are washed ashore, as shown in the photo on the right. Some people use these pieces to spread outside their homes as decoration. Larger blocks of coral were used as bricks of calcium carbonate for buildings. Early churches in Samoa, such as the Congregational Christian Church of Samoa building across road from the Ioane Viliamu building in Apia (photo, below left), are made of coral bricks. Coral blocks were also used to build other structures such as the wharf shown in the photo (below, right).

When solid calcium carbonate is strongly heated, solid calcium oxide, CaO, (also called "lime") is produced. A mixture containing calcium oxide can be purchased as an industrial waste product from BOC Gases (Samoa) Ltd, who use it to dry gases. When mixed with water to form a slurry, it is used by many people as a decorative whitewash for rocks that line the road in villages and stone walls. The calcium oxide firstly reacts with water to form calcium hydroxide, $Ca(OH)_2$, which then reacts with carbon dioxide in the air to produce calcium carbonate. An insoluble white coating of calcium carbonate is remains on the rocks once the water has evaporated.

S T R A N D 4: 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.

METALS

Students will investigate and develop their scientific understanding of **metals** *when they:*

- Carry out an investigation to identify and compare the physical properties of metals and non metals.
- Investigate the properties of metals through their reactions with air (oxygen), water, and dilute acids.
- Deduce the activity series of metals using reactions between metals and air, water and dilute acid and the displacement in metallic salts.
- Investigate metal corrosion and how to prevent it.
- Define the term 'alloy'.
- Compare the physical properties of the pure elements in an alloy with the properties of the alloy.

About three-quarters of elements are metals and have many common properties. Non-metal elements show properties that are very different from the properties of metals.

KEY POINTS:

- Metals are generally shiny solids at room temperature (mercury is the only liquid), are good conductors of heat and electricity, and can be hammered into thin sheets (malleable) and pulled into wires (ductile). (**MatW** p62-63)
- Non-metals are generally gases or dull, brittle solids at room temperature (bromine is the only liquid). Non-metals conduct heat and electricity poorly.
- Metals react as reducing agents and lose electrons in chemical reactions to form positive ions.
- Non-metals react with metals and hydrogen as oxidising agents and gain electrons in chemical reactions to become more negatively charged.
- Reactive metals undergo a chemical reaction with:
  - acids to produce an ionic metal compound and hydrogen gas (**MatW** p70-71).
  - water (or steam) to produce a metal hydroxide compound in solution and hydrogen gas (**MatW** p68-69).
  - oxygen (in the air) to form a metal oxide compound (**MatW** p66-67).

    - the dissolved ions of another less reactive metal (**MatW** p73).

- Chemists list metals in order of their reactivity with water, steam, acid and ability to reduce aqueous ions of another metal. This is known as the "activity series" of metals. The most reactive metal is listed at the top of the list and the least reactive metal at the bottom. This list can be used to predict the reactivity of a particular metal. (**MatW** p72)

- Corrosion is a natural redox reaction that causes metals to be oxidised to their metal oxide compound. Oxygen (in air) **and** water are needed for this reaction to take place. The most common form of corrosion is the rusting of iron (the main component of steel). Corrosion damages metal structures by weakened them. (**MatW** p74-75)

- Rusting of iron can be prevented (or limited) by protecting the surface of the metal with an unreactive coating (such as paint, oil, plastic or chromium) or by attaching a **more** reactive metal (such as zinc or magnesium) to it. (**MatW** p76-77)

- Alloys, such as brass and solder, are solid mixtures of metals with different properties to their elements. Many metallic materials are alloys, such as coins, jewellery, machine parts, and construction materials (**MatW** p64-65).

QUESTIONS:

1. Describe the physical properties of metals and compare them with those of non-metals.
2. What is the "activity series"?
3. By what ways can rusting of iron be prevented?
4. What is an "alloy"?
5. Metals act as reducing agents. What charge resides on the resulting metal ion?

Storyline

Rusting: The downfall of iron

Rusting is a redox reaction. Iron that comes into contact with water and oxygen corrodes by rusting as it forms a hydrated iron oxide compound that flakes off to expose fresh metal for further reaction. Eventually an iron object may completely rust away if not protected from water and air. Rusting can be controlled in several ways: by protecting the surface of the iron with an unreactive surface coating, such as paint, oil or plastic, which acts as a barrier to water and oxygen, or by attaching a more reactive metal to the iron, such as the zinc covering on galvanised iron, and which corrodes in preference to the iron (and referred to as “sacrificial protection”).

The wind in Samoa has a high content of sea-salt (sodium chloride) which accelerates rusting. This happens to unpainted iron roofs, cars with dents and scratches, and wire fencing around plantations. Corrugated iron that is galvanised with a coating of zinc and used on most houses for roofing quickly develops rust if left unpainted (as shown in the photo by the rusting corrugated iron of the Fale Samoa in the Apia area). Therefore, it is important to paint iron roofs. “Coloursteel” roofing is imported from New Zealand and is iron roofing that has already been covered with a protective layer of paint. Although it is more expensive that corrugated iron, it does not require continual repainting as corrugated iron does. Other products, such as “Rust Stop”, that are also painted onto iron objects to limit rusting can be purchased from the hardware store.

Fuel for vehicles and planes arrives in Samoa in tankers that moor off-shore in Apia Harbour. The fuel is transferred to steel storage tanks on land using a steel pipe that is submerged in the harbour and runs along the sea-floor to the tanker’s mooring point. After the seawater in the pipe has been pumped out, jet fuel, then petrol and then diesel from the tanker are pumped through this pipe into the storage tanks. However, as this pipe fills with seawater when not used, it is vital to protect the iron pipe (as well as the inside of the fuel storage tanks) from seawater to prevent rusting occurring. This happens in several ways. Firstly, the pipe is painted on both the inside and the outside with an epoxy paint containing zinc. The epoxy in the paint forms a plastic-like coating over the iron that acts a physical barrier to water and oxygen, while the zinc present in the paint provides sacrificial protection. Therefore, rusting on both the inside and outside of the pipe is prevented. Several pieces of zinc metal are also attached to the pipe in various places that act as "sacrificial anodes" and corrode in preference to the pipe. The steel storage tanks are also painted on the inside with an epoxy paint containing zinc.

OXYGEN AND HYDROGEN

Students will investigate and develop their scientific understanding of **oxygen and hydrogen and their compounds** *when they:*

- Carry out the laboratory preparation of oxygen and hydrogen.
- Investigate the occurrence and properties of molecular oxygen and ozone.
- Discuss the ozone layer and its importance to the earth.
- Describe the effect of CFCs on the ozone layer.
- Describe the industrial preparation of oxygen and hydrogen.
- List the uses of hydrogen.

Oxygen

Symbol: O *Atomic number:* 8 *Electron arrangement:* 2, 6.

- The most abundant element on Earth's surface.
- Present as elemental oxygen, O_2, in 21% of the volume of the atmosphere.
- Contributes about 50% of the mass of Earth's crust as compounds with metals and silicon and 89% of the mass of water, H_2O.
- Essential constituent of all living matter (together with C, H and N).
- Required for respiration (the process by which organisms obtain energy needed for life).

Respiration: $\underset{\text{glucose}}{C_6H_{12}O_6} + 6O_2 \rightarrow 6CO_2 + 6H_2O + \text{energy}$

- Only slightly soluble in water but enough oxygen dissolves to support aquatic life.

Physical properties

- Occurs as diatomic molecules, O_2, with a double bond (O=O) linking two oxygen atoms.
- Colourless and odourless gas; liquefies at -183°C and freezes at -218°C.

Chemical properties

- Reacts directly with most other elements and forms oxide compounds with all elements except He, Ne, Ar and Kr.

  Metals form metal oxides

  e.g. $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

Non-metals form non-metal oxides:

e.g. $S(s) + O_2(g) \rightarrow SO_2(g)$

Oxides are classified based on their acid/base properties.

- Basic oxides react with acids (most metal oxides) e.g. MgO
- Acidic oxides react with bases (most non-metal oxides) e.g. SO_2
- Some oxides are "amphoteric" and *react* with **both** acids and bases e.g. Al_2O_3.
- Some oxides do not react with either acids or bases e.g. CO.

- Combustion reactions: Many substances burn vigorously (combust) in pure oxygen to produce oxides. Fuels (such as methane, CH_4) give off much energy in combustion reactions as well as products of carbon dioxide, CO_2, and water, H_2O.

Uses of oxygen

- In the manufacture of steel: Iron that has a large amount of carbon impurity is brittle and weak. Steel consists mostly of iron. Oxygen is reacted with carbon to reduce the amount of impurities in steel, making it harder.
- For cutting and welding metals: "Oxy-acetylene" torches burn ethyne (acetylene, C_2H_2) gas in oxygen and give a very hot flame (3200°C) that is hot enough to melt metals.
- In breathing apparatus: Hospitals use oxygen for patients with breathing difficulties and to revive accident victims. Also used in air-tanks for deep-sea diving and high-altitude mountaineering.
- In space travel: Used as a fuel in rockets.

Industrial production of oxygen: Obtained from the atmosphere by fractional distillation of a mixture of N_2 and O_2 from liquid air that has had dust, CO_2 and water vapour removed. Nitrogen boils off first at -195°C leaving liquid oxygen, which is then stored under pressure in metal cylinders.

Laboratory preparation of oxygen: Several methods can be used in a laboratory to produce oxygen. A test for oxygen gas is that a stick with a glowing ember on its end reignites when oxygen is present.

- Heating purple crystals of potassium permanganate, $KMnO_4$, causes decomposition to potassium manganate, manganese dioxide and oxygen gas.

$$2KMnO_4(s) \rightarrow K_2MnO_4(s) + MnO_2(s) + O_2(g)$$

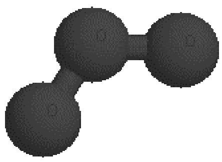
- Hydrogen peroxide, H_2O_2, is readily decomposed into water and oxygen by a catalyst (such as powdered manganese dioxide, MnO_2).

$$2H_2O_2(aq) \xrightarrow{\text{catalyst}} 2H_2O(\ell) + O_2(g)$$

- Passing electricity through water (in a process called *electrolysis*) converts water molecules into oxygen and hydrogen gases (see **MatW** p116).

Ozone, O_3

- Oxygen occurs in two forms: oxygen gas, O_2, and ozone, O_3. Different forms of an element in the same state are called **allotropes** (see **Chem12** p22). Allotropes have different physical properties.



Ozone, O_3, has two bonds linking three O atoms. Each bond is intermediate in length between a single O-O bond and a double O=O bond. The molecule is angular with a bond angle of about 117°.

- Ozone is a pale blue gas and causes headaches and difficulties in breathing in concentrations above about 100 parts per million.
- Ozone is the less stable allotrope and is very reactive, spontaneously decomposing to O_2 at sea level.
- Ozone is a powerful oxidising agent, and used commercially as a bleach and to destroy disease-causing organisms in drinking water.

The ozone layer

- Ozone is formed in the upper atmosphere (about 20-50 km above Earth's surface). High-energy ultraviolet (UV) radiation from outer space has enough energy to cause a molecule of oxygen, O_2, to split into two reactive oxygen atoms. These atoms then react with another oxygen molecule to produce ozone. The ozone that is formed can absorb more UV radiation and dissociate back into oxygen molecules.

  $O_2(g) + UV \rightarrow 2O(g)$ then $2O(g) + O_2(g) \rightarrow O_3(g)$

  and:

  $O_3(g) + UV \rightarrow O_2(g) + O(g)$ then $O_3(g) + O(g) \rightarrow 2O_2(g)$

  Therefore ozone in the upper atmosphere absorbs radiation that would be harmful if it penetrated to the Earth's surface.

Changes in the ozone layer

- Scientists have discovered a decrease in the concentration of ozone over Antarctica and the Arctic. This is linked to the presence of compounds known as CFCs (an abbreviation for chlorofluoroalkanes) that are widely used in refrigerators and aerosol cans.
- Although CFCs are chemically unreactive in the lower atmosphere, UV radiation causes these compounds to break down to chlorine atoms and chlorine oxide, ClO, in the upper atmosphere. These react with ozone molecules, causing the amount of ozone in the upper atmosphere to decrease.
- It is predicted that a decrease in the ozone layer will allow an increase in the amount of harmful radiation reaching the Earth's surface, and consequently contribute to an increase in the incidence of skin cancer.

QUESTIONS:

1. Which of the elements that occurs in air is always a reactant in combustion reactions?
2. Passing electricity through water liberates its elements. What are the formulae of the two products of this reaction?
3. Explain why the presence of ozone in the upper atmosphere is desirable.

Hydrogen

Symbol: H *Atomic number:* 1 *Electron arrangement:* 1.

- The most abundant element, making up about 90% of the universe.
- The lightest element and occurs on Earth combined in water (H_2O), natural gas (methane, CH_4) and living matter (in proteins, carbohydrates and fats and oils).

Physical properties

- Occurs as diatomic molecules, H_2, with a single bond (H-H) linking two hydrogen atoms.
- Colourless and odourless gas; liquefies at -253°C and solidifies at -259°C. Not soluble in water.

Chemical properties

- Chemistry dominated by tendency to share its electron with other non-metal atoms.
- Burns in air with a colourless flame and can react explosively with oxygen.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell)$$

Uses of hydrogen

- Production of ammonia, NH_3, in the "Haber process":

  $$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

  Ammonia is then used in the manufacture of fertilisers and nitric acid, HNO_3.

- As a fuel
  - In rockets: Reaction of hydrogen with oxygen generates water and a large amount of heat, providing a rocket with enough thrust needed to escape Earth's gravity.
  - In fuel cells: Electricity is produced in a controlled reaction of hydrogen with oxygen (in the opposite process to electrolysis of water). Used as a source of electricity in spacecraft. Considered as an alternative to using non-renewable petrol as a fuel for cars.

- Production of methanol, CH_3OH. Methanol (an alcohol) is produced from carbon monoxide and hydrogen (at 300°C and using a zinc catalyst) and is used as a fuel.

  $$CO(g) + 2H_2(g) \rightarrow CH_3OH(g)$$

- For hydrogenation of oils. Used for converting liquid vegetable oils into margarine, a soft solid fat that is easier to spread using a knife. Vegetable oils (with low melting points) are long-chain organic compounds that contain a lot of carbon-carbon double bonds (C=C) and are said to be polyunsaturated (poly: many; unsaturated: contains C=C bonds). Oils solidify to fats when some of the C=C bonds in the oil molecules are converted into single C-C bonds by adding a hydrogen atom to each C atom (and each C atom now has the maximum number of atoms attached to it). This occurs when the vegetable oil is pressurised with hydrogen and heated to 200°C then exposed to a nickel catalyst that aids the reaction where some of the C=C bonds convert into HC-CH bonds.

$$\text{vegetable oil}(\ell) + H_2(g) \xrightarrow{\text{Ni catalyst}} \text{margarine}(s)$$

  Eating fat derived from plants is considered to be healthier than eating animal fat (**MatW** p99).

Industrial production of hydrogen: Hydrogen can be obtained by reacting natural gas from an underground source (mainly methane) with steam (water) in a process called "steam reforming". High pressure and temperature (700°C) and a nickel catalyst are used to ensure that large amounts of products (hydrogen and carbon monoxide, CO) are produced.

$$CH_4(g) + H_2O(g) \xrightarrow{\text{Ni catalyst}} CO(g) + 3H_2(g)$$

The carbon monoxide gas produced can be reacted with more steam at 400°C using an iron catalyst to form more hydrogen as well as carbon dioxide in a further reaction.

Electricity can be used to manufacture hydrogen from sources other than non-renewable hydrocarbons. Hydrogen is produced from the electrolysis of water as well as being produced during the electrolysis of brine (a sodium chloride solution) for the manufacture of sodium hydroxide.

Laboratory preparation of hydrogen: The most convenient method for preparing hydrogen in the laboratory is the reaction of zinc (or another reactive metal) with a dilute acid (hydrochloric acid or sulfuric acid but NOT nitric acid). The hydrogen gas produced can be collected by displacing water from an inverted glass jar.

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

To test if hydrogen gas is given off in a reaction, a stick with a glowing ember on its end is placed quickly into a test-tube containing hydrogen gas that was stopped and a "squeak" or "pop" is heard (if performed quickly before the hydrogen escapes).

QUESTIONS:

1. Hydrogen reacts directly with the elements that are two major components of air.
    a. Give the names and formulae of these elements.
    b. Give the formulae of the products of the two reactions.
    c. The heat from one of these reactions is used to propel rockets. Which one is it?
    d. What is the name of the product of the other reaction, and how is it used commercially?
2. Elemental hydrogen is very common in the upper atmosphere; however on Earth hydrogen occurs combined with other elements. Two simple compounds of hydrogen are used as a commercial source of elemental hydrogen.
    a. Name of these compounds.
    b. The two compounds react together in the presence of a catalyst a high temperature and pressure to give elemental hydrogen. What is the second product of the reaction?

WATER

Students will investigate and develop their scientific understanding of **the compound water** *when they:*

- Describe the following properties of water such as melting point and boiling point and solvent properties.
- Explain the purification of water by the process of filtration and chlorination.
- Define the terms "soft" and "hard" water.
- Investigate the action of soap in water.
- Define the terms: deliquescence, hygroscopic, and efflorescence.

Water is all around us. It is a unique substance that has some unusual properties, such as great solvent power, that are vital for life on Earth. These properties arise from the nature of the water molecule that is made up from two H atoms and an O atom.

Formula: H_2O (**MatW** p 116-117)

- Most abundant liquid on Earth (about 70% of Earth's surface covered by water). Living organisms are composed mostly of water.
- Burning H_2 and O_2 gases in 2:1 ratio results in only water remaining.
- Water is 88.89% oxygen by mass and 11.11% hydrogen by mass.
- Water molecule has angular or bent structure with a H-O-H angle of 105° (as 2 bonding electron pairs and 2 non-bonding electron pairs are arranged approximately tetrahedrally around the oxygen atom for maximum repulsion. (**Chem12** p13)
- Unsymmetrical distribution of electrons causes water molecule to be highly polar (H atoms have partial positive charges and O atom has a partial negative charge).
- An attraction occurs between the hydrogen atom in one water molecule and the oxygen atom in another water molecule (this fairly strong attraction between water molecules is called *hydrogen bonding*).
- Hydrogen bonding results in many unusual properties for water.

Physical properties:

- Colourless and odourless liquid.
- Freezes at 0°C and boils at 100°C.
- Density defined as 1.000 g mL^{-1}.
- Very poor conductor of electricity (unless it contains large amounts of dissolved ions).

Solvent properties: Water molecules are attracted to the charged ions in an ionic solid and cause the crystal structure to break down as each ion becomes surrounded by water molecules and separates from the crystal, which is accompanied by the release of energy. This explains why water is a good solvent for ionic compounds. Other polar substances dissolve readily in water as they form hydrogen bonds with water. Ionic and polar solids become more soluble in water as the temperature of water increases. (See diagram **Chem12** p50).

"*Water of crystallisation*": Some ionic solids are referred to as being "hydrated" as they contain water molecules as part of their crystal structures. This water is called "water of crystallisation". Different hydrated substances may contain different amounts of water of crystallisation. Since the relative amount of water in a particular hydrated solid is always the same, the names and formulae of hydrated solids also show the water of crystallisation which appears after the dot. For example, the formula for cobalt chloride hexahydrate is $CoCl_2{\bullet}6H_2O$ (**MatW** p 121; **Chem12** p42).

When hydrated solids are heated and all the water of crystallisation is driven off, the resulting solid is said to be "anhydrous". Some hydrated solids lose some of their water of crystallisation naturally to the air as water vapour. These substances are said to be "efflorescent". For example, sodium carbonate decahydrate, which is known as "washing soda", is efflorescent. Sodium carbonate crystallises from solution as sodium carbonate decahydrate, $Na_2CO_3{\bullet}10H_2O$. These crystals give off water vapour (effloresce) to form sodium carbonate monohydrate, $Na_2CO_3{\bullet}H_2O$ on standing in air. Both of these hydrated solids are converted to anhydrous sodium carbonate by heating at 100°C.

Anhydrous cobalt chloride, $CoCl_2$, can gain water from the atmosphere (which contains water vapour) but not enough to form a solution. Such substances are referred to as being "hygroscopic". Some anhydrous solids, such as potassium carbonate, K_2CO_3, absorb enough water from the atmosphere to form a solution and are said to be "deliquescent" (**MatW** p 121).

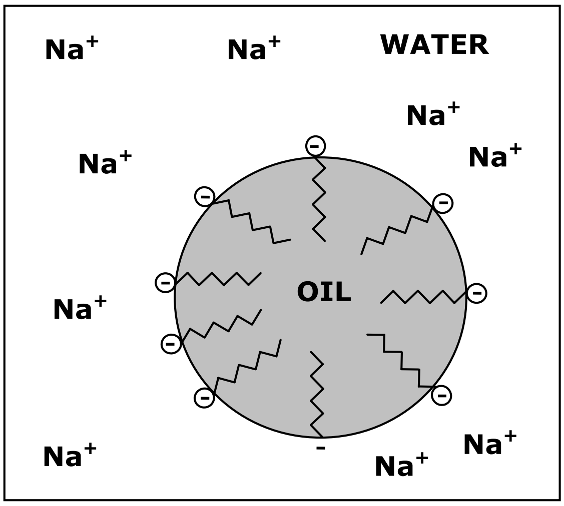
Unusual properties:
- Water is very stable towards heat.
- Water's melting point and boiling point are much higher than predicted for a small molecule: Significant energy is required to break the hydrogen bonds between water molecules before ice melts completely to liquid water or liquid water evaporates completely to a vapour.

- Water has a high *heat capacity*: Large volumes of water (such as lakes and oceans) can absorb a lot of heat energy without their overall temperature changing much. Earth's oceans help prevent vast changes in the climate when extreme variations in temperatures occur.
- Ice (solid) floats on water (liquid). Water has a higher density than ice, unlike most other substances where their solid form is denser than their liquid form.

  When water freezes, hydrogen bonding extends throughout the whole structure of ice and causes the water molecules to become arranged in a regular pattern (in a pattern similar to that of diamond). The structure spaces the molecules further apart in ice than they are in liquid water, and therefore the volume of water increases by 9% when it freezes (which is why glass bottles crack if used when freezing water). When ice melts, most of the hydrogen bonds are broken by heat and the water molecules pack closer together causing liquid water to be denser than ice.

  This can explain why lakes and oceans freeze from the surface downwards in winter in cold climates. As cold air above the lake cools water at the surface, the surface water becomes less dense and remains at the top where it freezes. The layer of ice on the surface helps to insulate the water underneath from further heat loss and fish and plants can surface under the ice.

Soap: Animal fat heated in a solution of sodium hydroxide, NaOH(aq), is converted into sodium stearate, $C_{17}H_{35}COO^{-}Na^{+}$, (also called sodium octadecanoate) as the main product. Because this is the cleaning agent in soap, this reaction is referred to as "saponification". This ionic compound releases sodium ions and stearate ions when dissolved in water.



Oils are not soluble in water and can only be washed away in water when soap is present. This is because 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 stays in water. This way, the stearate ions completely surround the oil, forming tiny droplets suspended in water that can be rinsed away.

Hard water (**MatW** p 121): Water is described as being "hard" if it forms a solid "scum" that cannot be washed away when soap is added. The presence of moderate amounts of calcium ions, Ca^{2+}, and magnesium ions, Mg^{2+}, makes water "hard". These react with stearate ions to produce insoluble precipitates in water. Soap will not produce a lather and clean in hard water until all the calcium and magnesium ions have been precipitated. Hard water results when rain water flows over rocks containing calcium or magnesium minerals, which dissolve in the water. It is difficult to wash with soap in seawater since seawater also is "hard" as it contains significant quantities of $Ca^{2+}(aq)$ and $Mg^{2+}(aq)$ ions. Water that contains almost no calcium or magnesium ions is referred to as "soft" water.

Purification of water for drinking (**MatW** p 119)*:* Water for drinking can come from the collection of rainwater, streams and rivers, and underground wells. Many sources do not give pure water and contain impurities and bacteria that make the water unsafe to drink. It is often necessary to purify water before it can be used for drinking.

The initial step in water treatment is to filter the water through beds of fine sand so that floating debris and particles of suspended matter are removed. Particles may include silt, natural organic matter, and microorganisms. A chemical called "alum" may also be added to the water to cause particles of dirt to clump together so that they settle out of the water as sediment due to gravity. Filtration clarifies water and improves the effectiveness of disinfection.

Water must be disinfected to kill or inactivate disease-causing microorganisms. This is the most important step in drinking water treatment and involves "chlorination" of the water. Chlorine is added to the water as chlorine gas, Cl_2, sodium hypochlorite solution, $NaOCl(aq)$, or solid calcium hypochlorite, $Ca(OCl)_2$. When dissolved in water, these substances provide dissolved chlorine, which destroys harmful organisms. The chlorine that is added to water also gives residual protection against biological contamination in the water distribution system.

QUESTIONS

1. Why is water referred to as a polar molecule?
2. The force of attraction between water molecules is fairly strong. What is the name given to this attractive force?
3. Choose the word in brackets that correctly completes the statement: Anhydrous solids, hydrogroscopic solids and deliquescent solids all have a tendency to (*gain* / *lose*) water (*to* / *from*) the atmosphere.
4. Give the formula for the two ions that are present in water described as being hard.

NITROGEN

Students will investigate and develop their scientific understanding of **the occurrence, properties, preparation and uses of nitrogen and its compounds** *when they:*

- Investigate the occurrence, properties and uses of nitrogen.
- Describe the properties of nitric oxide, NO_2 and HNO_3.
- Carry out the laboratory preparation of nitrogen dioxide and ammonia.
- Discuss the nitrogen cycle including the use of bacteria and fertilisers.
- Outline the industrial preparation of nitric acid and ammonia.
- Outline the properties and uses of ammonia.

Nitrogen:

Symbol: N *Atomic number:* 7 *Electron arrangement:* 2, 5.

- Nitrogen, N_2, is the major constituent of air by volume (78%).

Average percentage composition of air

GAS	%
Nitrogen (N_2)	78
Oxygen (O_2)	21
Argon (Ar)	0.9
Carbon dioxide (CO_2)	0.03
Others	0.07

- Occurs in the soil and ground water as ammonium (NH_4^+), nitrate (NO_3^-) and nitrite (NO_2^-) ions.
- Essential element for all living organisms as present in amino acids (that make up proteins) and nucleic acids (that make up DNA).
- Natural deposits of nitrate compounds (for example, sodium nitrate, $NaNO_3$) occur in dry areas (e.g. Chile) and are mined for use as fertilisers.
- Gaseous nitrogen is obtained from fractional distillation of liquid air (liquid nitrogen boils at -196°C).
- Used for the manufacture of ammonia, and this is then used to make fertilisers and nitric acid.
- Liquid nitrogen used as a coolant to rapidly freeze substances.

Physical properties

- Occurs as diatomic molecules, N_2, with a triple bond (N○N) linking two nitrogen atoms.
- Colourless and odourless gas; liquefies at -196°C and freezes at -210°C.

Chemical properties

- N_2 has low reactivity as it contains strong triple bond linking nitrogen atoms.

Ammonia: *Formula:* NH_3

- Easily liquefied gas with a characteristic pungent smell.
- Primarily used in the manufacture of nitric acid, HNO_3, and in the production of solid nitrogen fertilisers, e.g. ammonium nitrate, NH_4NO_3, and urea, $CO(NH_2)_2$.
- Dilute solutions of ammonia used in many household cleaners.
- Ammonia readily dissolves in water. In a 1 mol L^{-1} solution of ammonia, about 0.4% of ammonia molecules react with water to form ammonium, $NH_4^+(aq)$ and hydroxide, $OH^-(aq)$ ions. Therefore, aqueous ammonia is basic and turns red litmus blue.

$$NH_3(aq) + H_2O(\ell) \rightarrow NH_4^+(aq) + OH^-(aq)$$

- Ammonia neutralises acids, such as HCl, forming ionic ammonium compounds. For example, when ammonia gas comes into contact with hydrochloric acid fumes, tiny white solid particles of ammonium chloride form in the air which appears as white "smoke". This is a test for the presence of ammonia.

$$NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$$

- Ammonium compounds are very soluble in water. Ammonium ions are very weakly acidic and react with water to a small extent, resulting in the solution having a pH less than 7.

$$NH_4^+(aq) + H_2O(\ell) \rightarrow NH_3(aq) + H_3O^+(aq)$$

- Ammonium compounds are mainly used as soluble fertilisers to supply nitrogen to the soil for uptake by plants for the production of proteins. These are prepared by reacting ammonia with the appropriate acid. Common nitrogen fertilisers are:
  - ammonium nitrate, NH_4NO_3: made from ammonia and nitric acid.
  - ammonium sulfate, $(NH_4)_2SO_4$: made from ammonia and sulfuric acid.
  - a related substance also used as a solid fertiliser is urea, $CO(NH_2)_2$, made by heating ammonia with carbon dioxide under pressure.

Industrial preparation of ammonia: Ammonia is manufactured by the "Haber" process using nitrogen from the air and hydrogen obtained from natural gas. In this process, a mixture of hydrogen and nitrogen gases in a 3:1 ratio is pressurised and heated to about 450°C as it is passed over an iron catalyst. A reaction produces ammonia which is condensed and removed as a liquid.

$$N_2(g) + 3H_2(g) \xrightarrow{\text{Fe catalyst}} 2NH_3(g)$$

Laboratory preparation of ammonia: Ammonia gas can be made conveniently in the laboratory by heating an ammonium salt together with a strong base, such calcium hydroxide, in a test-tube.

$$2NH_4Cl(s) + Ca(OH)_2(s) \rightarrow 2NH_3(g) + 2H_2O(g) + CaCl_2(s)$$

The ammonia gas is collected by downwards displacement of air in an inverted test-tube (since ammonia is lighter than air). As water is also a product of the reaction, the gas is dried by passing it through solid anhydrous calcium oxide (this is a base as acidic drying agents, such as concentrated sulfuric acid, cannot be used).

Nitric acid: *Formula:* HNO_3

- Pure nitric acid is a colourless liquid that decomposes in light or when heated and turns yellow as dissolved nitrogen dioxide, NO_2, forms.
- In aqueous solutions is a strong acid:
    - Neutralises bases to form an ionic compound and water.
    - Reacts with carbonate compounds to give carbon dioxide gas as one of the products.
    - React with metal oxides or metal hydroxides to form metal nitrate compounds. All metal nitrates, such as sodium nitrate, $NaNO_3$, are soluble in water.
- Is a powerful oxidising agent (gains electrons from other substances such as metals):
    - Metals react with nitric acid to form the metal nitrate, water and oxides of nitrogen but **not** hydrogen gas (unlike other metals). For example, copper metal, Cu, is oxided to $Cu^{2+}(aq)$ ions by nitric acid. Nitrogen oxide, NO, tends to be formed when dilute nitric acid is used, and nitrogen dioxide, NO_2, tends to be formed when concentrated nitric acid is used in the reaction.

    Dilute:

    $$3Cu(s) + 8HNO_3(aq) \rightarrow 3Cu(NO_3)_2(aq) + 4H_2O(\ell) + 2NO(g)$$

    Concentrated:

    $$Cu(s) + 4HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2H_2O(\ell) + 2NO_2(g)$$

- Nitric acid of any concentration has no effect on very unreactive metals such as gold, Au, and platinum, Pt. Also, aluminium, Al, iron, Fe, and chromium, Cr, react initially with nitric acid to form a thin impervious oxide layer over the metal that stops any further reaction and these metals become "passive".
- Reaction with organic compounds leads to the production of explosives e.g. dynamite (nitroglycerine).
- Primarily used in the manufacture of nitrogen fertilisers, especially ammonium nitrate, NH_4NO_3.

Industrial preparation of nitric acid: Industrially, nitric acid is made by passing a mixture of ammonia and air (containing oxygen) over a platinum-rhodium catalyst at about 850°C to form nitrogen oxide, NO, and steam. The nitrogen oxide gas is cooled and then reacted with more air and then water to produce nitric acid, HNO_3. This process is called the "Ostwald" process.

$$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$$

$$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$$

$$3NO_2(g) + H_2O(\ell) \rightarrow 2HNO_3(aq) + NO(g)$$

The nitrogen oxide gas produced in the last reaction is recycled to make more nitrogen dioxide, NO_2.

Oxides of nitrogen (**MatW** p 112)**:** A number of different compounds containing nitrogen and oxygen occur. The common ones are:

- *Dinitrogen oxide* (also called nitrous oxide), N_2O: A colourless gas that is soluble in water and has a faintly sweet smell. Used as an anaesthetic and referred to as "laughing gas". Although fairly unreactive, it decomposes into nitrogen and oxygen gases when heated so will relight a glowing ember.
- *Nitrogen oxide* (also called nitrogen monoxide or nitric oxide), NO: A colourless gas that reacts rapidly with oxygen to form nitrogen dioxide, NO_2. Small quantities are formed during lightning flashes but large amounts are produced in vehicle engines as a pollutant. Nitrogen oxide gas is hazardous as it causes lung damage.
- *Nitrogen dioxide*, NO_2: A brown, poisonous gas that is heavier than air. Dissolves readily in water to form a mixture of nitric acid, HNO_3 and nitrogen oxide, NO.

$$3NO_2(g) + H_2O(\ell) \rightarrow 2HNO_3(aq) + NO(g)$$

  Under the high pressures and temperature that occur in vehicle engines, nitrogen and oxygen from the air react to produce nitrogen dioxide, which sinks to the ground, and many cities around the world have serious pollution problems with a brown haze in the air due to nitrogen dioxide. Nitrogen dioxide can be made in a laboratory by adding concentrated nitric acid to copper metal (see above).

QUESTIONS:

1. What is the major commercial use of elemental nitrogen?
2. Hydrogen, oxygen and nitrogen all occur as diatomic molecules. How does the nature of the bond between the atoms differ?
3. Ammonia is a basic compound that reacts with acids to form ammonium salts. What is the formula of the ammonium ion?

4. The commercial preparation of nitric acid begins by reacting ammonia with oxygen. What are the formulae for the two oxides of nitrogen that are intermediates in this process?
5. Which gases can be produced in the reaction of metals with nitric acid?
6. Give the formula of the oxide of nitrogen that dissolves in water to give nitric acid.
7. Which oxides of nitrogen contribute to air pollution?
8. Ammonia, nitric acid and ammonium nitrate are all water soluble compounds. In which case(s) does the solution turn blue litmus red and red litmus blue?

CARBON

Students will investigate and develop their scientific understanding of **the occurrence, properties, preparation (laboratory and industrial) and uses of carbon and its compounds** *when they:*

- Name the allotropes of carbon (diamond and graphite) and their respective properties and uses.
- Describe the industrial preparation of CO_2 and its properties and uses.
- Describe the properties of CO.
- Describe the causes of the greenhouse effect and its effect on climate.
- Describe the occurrence and properties of carbonates and their uses.

Carbon

Symbol: C *Atomic number:* 6 *Electron arrangement:* 2, 4.

- Carbon is a non-metal.
- Chemistry dominated by tendency to form chains and rings of carbon atoms that are also bonded to other atoms, particularly hydrogen.
- Earth's atmosphere has about 0.03% carbon dioxide gas by volume and this amount remains fairly constant. A smaller amount of carbon dioxide occurs dissolved in water in lakes and oceans.
- Carbon is found in all living organisms as it is part of all proteins, carbohydrates and fats.
- Plants use atmospheric CO_2 in the process of photosynthesis.
- Animals feed on plants and both animals and plants give off CO_2 gas in respiration of food.
- Decomposition of dead animals and plants, followed by geological events causing great pressures and temperatures over a long period of time resulted in deposits of coal, oil and natural gas (fossil fuels) being formed in Earth's crust.

- Large deposits of shells of sea animals on the ocean floors were compressed then subjected to geological upheaval to give rocks made of carbonate minerals, such as chalk, limestone and marble.
- Pure carbon occurs naturally in two crystalline forms – diamond and graphite. Different forms of an element are called "allotropes" and occur because the atoms are arranged in different patterns, causing each form to have different physical properties. Another allotrope of carbon was produced and characterised in 1985. This molecule, called "buckminsterfullerene", contains 60 carbon atoms arranged in a spherical pattern like a soccer ball. Many other related "fullerene" molecules have now been made.

Diamond

- Formed by intense pressure and heat deep within Earth's crust and occurs naturally in igneous rocks.
- Has a very high melting point and boiling point and is the hardest naturally occurring substance due to its structure.
- A crystal of diamond contains millions of carbon atoms, each joined by covalent bonds to four others resulting in a three-dimensional network structure (see diagram **Chem12** p17).
- Has a high refractive index and high reflectivity which makes diamonds sparkle, especially when faces are cut on the right angles, and is therefore prized as jewellery.
- Hard, abrasive character means that small diamonds are used in industry on the tips of metal drills and saws for cutting and engraving and boring through rock.

Graphite

- Some natural supplies occur but most graphite manufactured by heating coke (impure carbon) and sand in an electric furnace.
- Soft and slippery due to its structure and is used in making pencil "leads" and as a lubricant.
- Has a layered structure with each carbon atom covalently bonded to three others in the same plane forming a two-dimensional sheet of hexagonal rings. Weak attractive forces between the sheets allow them to move over one another easily and this gives graphite its slippery nature (see diagram **Chem12** p17).
- Covalent bonds linking the carbon atoms cause graphite to have a very high melting point and high boiling point.
- Graphite, unlike diamond, conducts electricity because some of its electrons are free to move in each sheet.

Charcoal
- An impure form of carbon made by heating wood in a limited supply of oxygen.
- Has the same structure as graphite but made of many very small crystals. It is very porous and has a large surface area, making charcoal chemically more reactive than graphite.
- Useful for absorbing gases (such as in gas masks) and coloured matter (such as removing the brown colour from raw sugar).

Oxides of carbon (**MatW** p 113)
- *Carbon monoxide*, CO
  - Colourless and odourless gas; insoluble in water.
  - Burns in air to form carbon dioxide.
  - Formed when carbon and hydrocarbon fuels burn incompletely due to lack of oxygen. Present in the exhaust fumes of petrol-driven vehicles and in cigarette smoke.
  - At concentrations above 0.1%, carbon monoxide is very poisonous as it combines with the haemoglobin in red blood cells more readily than oxygen does and suffocation from lack of oxygen quickly occurs.
  - Powerful reducing agent: reacts with iron ores (e.g. iron(III) oxide, Fe_3O_4) to give iron.

- *Carbon dioxide*, CO_2
  - Colourless and odourless gas.
  - Slightly soluble in water but dissolves much more readily in water when the pressure is increased and therefore used in the manufacture of flavoured fizzy drinks. Once the pressure is reduced by removing the bottle top, the carbon dioxide comes out of solution as bubbles of carbon dioxide gas.
  - Acidic gas as a very small amount of dissolved carbon dioxide (about 0.4%) reacts with water to form the weak acid, carbonic acid, H_2CO_3.

$$CO_2(aq) + H_2O(\ell) \rightarrow H_2CO_3(aq) \rightarrow H^+(aq) + HCO_3^-(aq)$$

  - Denser than air and does not support combustion. Can be easily liquefied at room temperature and stored in a cylinder under pressure. Used as a convenient fire extinguisher as it will smother fires and prevent oxygen from reaching the burning material and useful for electrical fires where water cannot be used.
  - Solid carbon dioxide is called "dry ice" and sublimes at -78°C. Since no liquid forms, dry ice is useful as an easily handled refrigerant for producing low temperatures.

- Produced by yeast enzymes in the fermentation of a sugar (glucose) to ethanol (an alcohol), giving beer and sparkling wine their bubbles when the bottles are opened.
- Produced from the complete combustion of fuels. For example:

  $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(\ell)$

Industrial preparation of carbon dioxide: produced in industry as a by-product during the manufacture of quicklime (calcium oxide, CaO) by heating solid calcium carbonate:

$$CaCO_3(s) \xrightarrow{\text{heat}} CO_2(g) + CaO(s)$$

Also formed in the fermentation process in the brewing industry:

$$C_6H_{12}O_6(aq) \xrightarrow{\text{enzymes}} 2CO_2(g) + C_2H_5OH(aq)$$

Laboratory preparation of carbon dioxide: made conveniently by reaction of dilute hydrochloric acid or dilute nitric acid with marble chips (solid calcium carbonate, $CaCO_3$).

$$CaCO_3(s) + 2HCl(aq) \rightarrow CO_2(g) + CaCl_2(aq) + H_2O(\ell)$$

Carbon dioxide gas is given off and can be collected downwards in a test-tube (as it is heavier than air). Dilute sulfuric acid cannot be used since an insoluble layer of calcium sulfate, $CaSO_4$, forms around the marble chips, preventing further reaction of the acid with the calcium carbonate below.

- *Metal carbonates and metal hydrogen carbonates*
  - Metal carbonates, such as copper carbonate, $CuCO_3$, can be prepared easily by precipitation of the metal ion and carbonate ion by reacting together two soluble ionic compounds containing the particular ions. For example:

    $CuSO_4(aq) + Na_2CO_3(aq) \rightarrow CuCO_3(s) + Na_2SO_4(aq)$
  - Only sodium carbonate, Na_2CO_3, potassium carbonate, K_2CO_3, and ammonium carbonate, $(NH_4)_2CO_3$ are soluble in water.
  - Metal carbonate compounds react with acids to produce carbon dioxide as one product (**MatW** p78-79).
  - Carbon dioxide gas is tested for by bubbling it through *limewater* (a solution of calcium hydroxide, $Ca(OH)_2$). The solution becomes cloudy when CO_2 is present due to the formation of tiny solid particles of insoluble calcium carbonate, $CaCO_3$.

    $Ca(OH)_2(aq) + CO_2(g) \rightarrow + CaCO_3(s) + H_2O(\ell)$

Continued bubbling of CO_2 through this solution causes the cloudiness to disappear due to the reaction of calcium carbonate forming soluble calcium hydrogen carbonate.

$$CaCO_3(s) + CO_2(g) + H_2O(\ell) \rightarrow Ca(HCO_3)_2(aq)$$

- Some metal carbonate compounds, such as calcium carbonate, $CaCO_3$, decompose when heated, forming a metal oxide compound as well as giving off carbon dioxide gas (**MatW** p80-81).
- Calcium carbonate, $CaCO_3$: Occurs naturally in two crystalline mineral forms – calcite (in marble, limestone, chalk) and aragonite (in sea shells and coral). Limestone is mined for use in making sodium carbonate, Na_2CO_3, which is required for glass manufacture, water and sewage treatment, production of soaps, and paper. Calcium carbonate is also used for making cement. Limestone (calcium carbonate) is placed on soils that are naturally acidic in order to reduce their acidity and raise the pH.
- Only sodium hydrogen carbonate, $NaHCO_3$, and potassium hydrogen carbonate, $KHCO_3$, are common in solid form. These solids are soluble in water (giving basic solutions) but decompose on heating to form carbonate compounds and carbon dioxide gas.

$$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(\ell)$$

- Solid sodium hydrogen carbonate, $NaHCO_3$, is known as "baking soda" and is used in cooking to give bread and cakes their fluffy texture. Baking soda is added to the wet dough or cake mixture and then decomposes when heated in the oven to give off carbon dioxide gas, causing the mixture to rise as the gas gets trapped in the dough. Baking powder includes a solid acid that reacts with sodium hydrogen carbonate when both are dissolved in the wet dough.

The Greenhouse Effect (**MatW** p109)**:** The average temperature of the Earth's surface (about 15°C) is determined by a balance that occurs between the energy Earth receives from the Sun that warms the ground or the sea and the heat energy radiated back into space by the warm surface of Earth. Some of the energy lost from the Earth's surface is captured by water vapour, carbon dioxide, methane, CFCs and other gases in the lower atmosphere and radiated back to the ground, preventing excess heat loss from Earth and helping to maintain a temperature that makes Earth habitable for life. This warming effect of carbon dioxide and water is called the "greenhouse effect" and without it, Earth would be a frozen planet with an average temperature of about -20°C.

Most of the greenhouse heating of Earth's surface is due to the large amount of water vapour in the atmosphere. However, carbon dioxide in the air also contributes to this warming effect. Increasing the concentration of carbon dioxide in the air increases the amount of radiation absorbed and increases the warming effect of the lower atmosphere.

Over the last hundred years, there has been an increase in the concentration of carbon dioxide in the atmosphere, resulting in an increase of 0.5°C in Earth's average temperature. If this trend continues, the concentration of CO_2 in the air could be doubled by 2050, causing an increase in Earth's average temperature by 2-3°C (and even greater at the poles). An effect of global warming would be the melting of the polar ice caps which would lead to a rise in the level of the oceans and widespread flooding of low-lying countries around the world. Also, the climate would change, making some areas drier with less rain and other areas wetter. This would be disastrous for many food crops.

The combustion of fossil fuels (coal and hydrocarbons) is the source of most of the carbon dioxide that is enhancing the greenhouse effect. Also, as a result of extensive tree-cutting to make more pastures from many tropical rain forests around the world, there are fewer trees to remove carbon dioxide from the atmosphere by photosynthesis. Other gases, including CFC's, are known to add to the greenhouse effect. In order to avoid further increases in the amount of greenhouse gases released to the air and an increase in global temperature, it is essential to stop using CFC's in aerosols, to decrease the combustion of fossil fuels and to drastically reduce deforestation around the world.

QUESTIONS

1. Name the allotrope of carbon in which each carbon is bonded to three other carbons.
2. Which oxide of carbon reacts with elemental oxygen?
3. Which oxide of carbon is classed as an acidic oxide?
4. What is dry ice?
5. Combustion reactions of hydrocarbons produce CO_2. What is the second product in these combustion reactions?
6. What is the formula for carbonate ion?
7. Carbonate ion is a base, and metal carbonates characteristically react with acids. A metal compound is formed as well as two other products. Give the formulae for the other products.
8. Give the formula of the precipitate responsible for the cloudiness seen when carbon dioxide is bubbled into limewater.

SULFUR

Students will investigate and develop their scientific understanding of **the occurrence, properties, preparation and uses of sulfur and its compounds** *when they:*

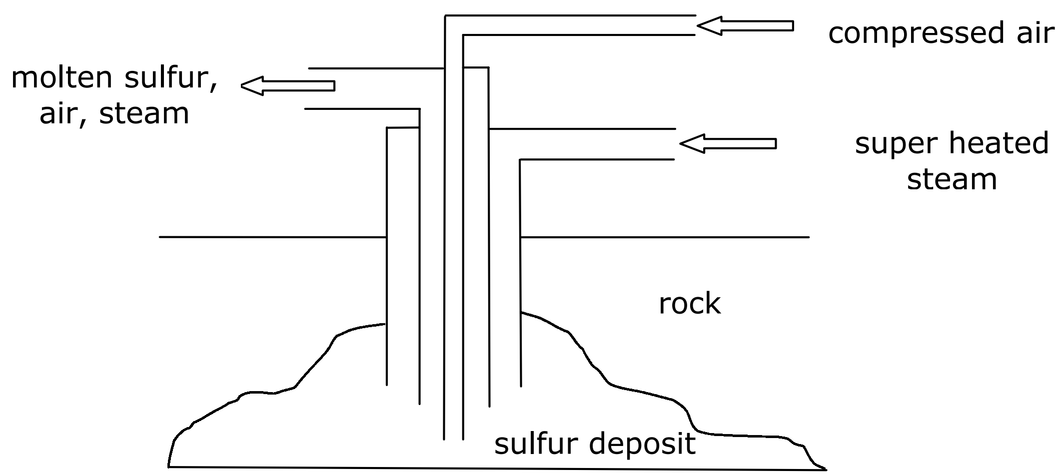
- Outline the extraction of sulfur by the Frasch process.
- Investigate the properties of sulfur including its allotropes and the effect of heat on sulfur.
- Discuss the uses of sulfur.
- Investigate the preparation and properties of SO_2 and SO_3.
- Outline the Contact process for the manufacture of sulfuric acid.
- Discuss the properties of sulfuric acid and its uses.

Sulfur:

Symbol: S *Atomic number:* 16 *Electron arrangement:* 2, 8, 6.

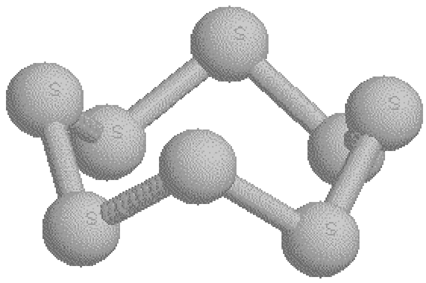
- Found in the elemental form in underground deposits (in USA, Poland and Japan) as well as around volcanic and geothermal areas.
- Occurs in combination with many metals as metal sulfide minerals (such as zinc blende, ZnS, and galena, PbS) and as metal sulfate deposits in dry areas (such as gypsum, $CaSO_4 \bullet 2H_2O$, and Epsom salts, $MgSO_4$).
- Other important sources of sulfur are crude oil and natural gas, from which it is extracted as hydrogen sulfide, H_2S.
- Used in the manufacture of sulfuric acid, H_2SO_4, and this is then used in making superphosphate fertiliser.

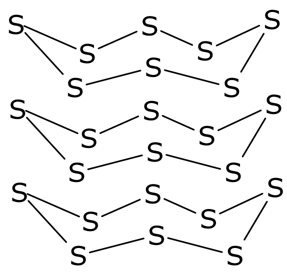
Mining sulfur from underground deposits: The common method used to extract underground sulfur is the "Frasch process", which utilises the low melting point of sulfur (119°C). Three concentric pipes are sunk deep underground into sulfur-containing rock. Superheated water at 170°C is forced down the outer pipe into the sulfur deposit which then melts. Compressed air pumped down the inner pipe forces a foam of molten sulfur, water and air up to the surface through the middle pipe and the pure sulfur is allowed to solidify.



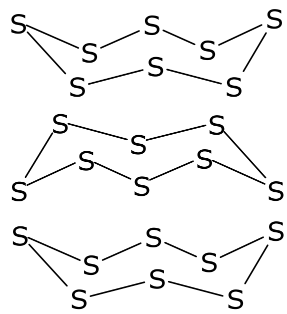
Physical properties

- Yellow solid that is insoluble in water (as contains non-polar molecules). Poor conductor of heat and electricity.
- The two main allotropes of sulfur are *rhombic* sulfur and *monoclinic* sulfur. These differ in the shapes of crystals they form due to the S_8 molecules arranging in different patterns (S_8 molecules pack more densely in rhombic form) (**Chem12** p22).
- Both allotropes contain S_8 molecules, in which single bonds link sulfur atoms into a puckered 8-membered ring.





rhombic



monoclinic

- Rhombic sulfur is he common form. It consists of large yellow crystals and is stable at temperatures below 96°C.
- Monoclinic sulfur is obtained when sulfur solidifies at temperatures above 96°C. It forms long amber-yellow needles, which change into small rhombic crystals if the temperature drops below 96°C.

Heating sulfur: Both rhombic sulfur (melting point 113°C) and monoclinic sulfur (melting point 119°C) melt to a yellow liquid of S_8 molecules. As the temperature rises, the colour of the liquid darkens from a red-brown colour until it is almost black and very viscous (very sticky) at about 200°C. This is due to the rings of sulfur atoms breaking open and joining up to form very long chains of many sulfur atoms. The long chains become easily tangled, resulting in the high viscosity.

As the liquid is heated further to about 400°C, the liquid becomes more mobile as the long chains break down and re-form into S_8 rings. Sulfur boils at 445°C giving a vapour of molecules containing from two to eight sulfur atoms.

If nearly boiling sulfur is poured into cold water, a soft and rubbery brown solid is obtained. This is another allotrope of sulfur, known as "plastic sulfur". It consists of a random arrangement of chains of sulfur atoms which, when stretched, align themselves parallel to each other. With time, it changes back into rhombic sulfur as the chains break and reform S_8 rings.

Chemical properties

- Burns in air and oxygen with a blue flame and producing sulfur dioxide gas. This is a redox reaction in which sulfur acts as the reducing agent (donating electrons).

  $S(s) + O_2(g) \rightarrow SO_2(g)$

- Combines with most metals when heated to give metal sulfide compounds. This is also a redox reaction in which sulfur acts as a oxidising agent (gaining electrons).

  $Cu(s) + S(s) \rightarrow CuS(s)$

Oxides of sulfur (**MatW** p 113; **Chem12** p25)**:**

- *Sulfur dioxide*, SO_2
  - Colourless gas with a choking smell.
  - Dissolves readily in water, forming a solution of the weak acid, sulfurous acid, H_2SO_3.

    $SO_2(g) + H_2O(\ell) \rightarrow H_2SO_3(aq)$

  - Used as a bleaching agent when chlorine is not suitable, (for example, for wood pulp in the manufacture of paper).
  - Used as a food preservative for meats and fruits as it kills micro-organisms. Also added to wine to prevent further fermentation by yeast once the wine has been bottled.
  - Major air pollutant as SO_2 is formed when sulfur-containing coal is burnt as a fuel by many industries. Once it is released into the air, it reacts with water vapour to form sulfurous acid that contributes to "acid rain" pollution.
  - Reacts slowly with oxygen to produce sulfur trioxide, SO_3.

    $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$

- *Sulfur trioxide*, SO_3
  - Exists as a gas of SO_3 molecules above 17°C.
  - Very soluble in water and reacts readily to form the strong acid, sulfuric acid, H_2SO_4.

    $SO_3(g) + H_2O(\ell) \rightarrow H_2SO_4(aq)$

  - The SO_3 which results from the oxidation of SO_2 in the atmosphere is a major pollutant due to its reaction with water forming sulfuric acid. This has contributed to "acid rain" in many industrialised countries (**MatW** p 114-115).

Sulfuric acid, H_2SO_4

- Concentrated sulfuric acid is a viscous, colourless liquid containing molecules of H_2SO_4 and is 98% pure. As few ions are present, it is a poor conductor of electricity.
- Dilute sulfuric acid is an aqueous solution containing hydronium ions, H_3O^+, and hydrogen sulfate ions, HSO_4^- and conducts electricity.
- Shows typical properties of a strong acid:
  - Reacts with water to generate hydronium, H_3O^+, ions.

    $H_2SO_4(\ell) + H_2O(\ell) \rightarrow H_3O^+(aq) + HSO_4^-(aq)$

  - Reacts with bases to form an ionic compound and water. The reaction with strong bases removes **both** hydrogens from sulfuric acid, H_2SO_4.

    $CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(\ell)$

  - Reacts with reactive metals to form a metal compound and hydrogen gas.

    $Fe(s) + H_2SO_4(aq) \rightarrow FeSO_4(aq) + H_2(g)$

  - Reacts with metal carbonates to give carbon dioxide as a product together with water and a metal compound.

    $Na_2CO_3(s) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + CO_2(g) + H_2O(\ell)$

- Acts as a dehydrating agent to remove water from moist gases as it has a high attraction for water. As much heat is given off as this occurs, dilute sulfuric acid is always prepared by pouring concentrated sulfuric acid into a large volume of water while stirring to avoid explosions.
- Prepared commercially on an enormous scale world-wide as sulfuric acid is used in the manufacture of many substances, for example, fertilisers (such as superphosphate and ammonium sulfate), synthetic fibres and plastics, detergents, medicines, dyes, car batteries and explosives.

Industrial preparation of sulfuric acid: The "Contact process" is used to commercially manufacture sufuric acid. The raw materials are sulfur, oxygen and water. Three steps are needed to prepare sulfuric acid.

1. Sulfur is burned in dry air to form sulfur dioxide.

   $S(s) + O_2(g) \rightarrow SO_2(g)$

2. Sulfur dioxide is reacted with oxygen at about 450°C to form sulfur trioxide. A catalyst, vanadium pentoxide, V_2O_5, is used to speed up the reaction, making the process more economical.

$$2SO_2(g) + O_2(g) \xrightarrow{450°C,\ V_2O_5} 2SO_3(g)$$

The sulfur trioxide formed is not absorbed directly into water as the heat generated in this reaction would vapourise the sulfuric acid formed.

2. Sulfur trioxide is reacted with sulfuric acid to form "oleum", $H_2S_2O_7$ (also called "fuming sulfuric acid") which is then reacted with water to form concentrated sulfuric acid.

$$SO_3(g) + H_2SO_4(\ell) \rightarrow H_2S_2O_7(\ell)$$

$$H_2S_2O_7(\ell) + H_2O(\ell) \rightarrow 2H_2SO_4(\ell)$$

Oleum can be stored in steel tanks without corrosion of the iron in the steel as long as no water is present to react and form corrosive sulfuric acid.

All of these reactions give out heat (are exothermic) which provides sufficient energy to operate the process.

Superphosphate fertiliser: The element phosphorus is vital for life as it forms part of the DNA molecule in all plants and animals. Soils deficient in phosphorus require application of a fertiliser containing phosphorus compounds. Deposits of calcium phosphate, $Ca_3(PO_4)_2$, (known as "rock phosphate") are easily mined, as on Nauru. However, calcium phosphate is insoluble in water and cannot be absorbed by plant roots. Reacting calcium phosphate with sulfuric acid produces a fertiliser that contains a soluble phosphate compound that plants can absorb. This fertiliser is called "superphosphate" and is a mixture of solid calcium dihydrogen phosphate, $Ca(H_2PO_4)_2$, and solid calcium sulfate, $CaSO_4•2H_2O$.

$$Ca_3(PO_4)_2(s) + 2H_2SO_4(aq) + 2H_2O \rightarrow Ca(H_2PO_4)_2(s) + 2CaSO_4•2H_2O(s)$$

QUESTIONS:

1. What are the formulae and names of two oxides of sulfur? Which oxide of sulfur reacts further with oxygen?
2. Both of the oxides of sulfur are acidic oxides. Which one reacts with water to give sulfuric acid?
3. Name the three raw materials needed for commercial preparation of sulfuric acid.

CHLORINE

Students will investigate and develop their scientific understanding of **the occurrence, properties, preparation and uses of chlorine and its compounds** *when they:*

- Discuss the properties and uses of chlorine.
- Investigate the industrial preparation of chlorine by the process of electrolysis.
- Outline the industrial preparation of hydrochloric acid.
- Discuss the properties and uses of hydrochloric acid.
- Outline the properties of chlorides.

Chlorine:

Symbol: Cl *Atomic number:* 17 *Electron arrangement:* 2,8,7.

- Chlorine is a poisonous gas with an extremely irritating smell. Has a corrosive effect on the respiratory system and used in warfare in World War I as a chemical weapon.
- Used extensively in industry. Most important use is in the manufacture of hydrochloric acid, HCl. Other uses are:
  - As a strong bleaching agent for cotton and wood pulp.
  - As a germicide in swimming pools.
  - As a sterilising agent in the purification of drinking water.
  - In the manufacture of organic chemicals such as solvents (for drying cleaning) and plastics (e.g. PVC).

Physical properties

- Occurs as non-polar diatomic molecules (Cl_2) with a single bond linking two chlorine atoms.
- Greenish-yellow pungent smelling gas; boils at -34°C and is easily liquefied under pressure for storage in cylinders.
- Denser than air and moderately soluble in water.

Chemical properties

- Cl_2 is a powerful oxidising agent (gains electrons).
  - Reacts with metals to form metal chlorides.

$$Cl_2(g) + 2Na(s) \rightarrow 2NaCl(s)$$

  - Reacts with non-metals to form covalently-bonded compounds.

$$Cl_2(g) + H_2(g) \rightarrow HCl(g)$$

  - Oxidises bromide, $Br^-(aq)$ and iodide, $I^-(aq)$ ions to the elements.

$$Cl_2(g) + 2Br^-(aq) \rightarrow 2Cl^-(aq) + Br_2(aq)$$

$$Cl_2(g) + 2I^-(aq) \rightarrow 2Cl^-(aq) + I_2(aq)$$

- Reacts with water to form hydrochloric acid, HCl and hypochlorous acid, HOCl (a powerful oxidising agent that causes bleaching).

$$Cl_2(g) + H_2O(\ell) \rightarrow HOCl(aq) + H^+(aq) + Cl^-(aq)$$

- React with dilute metal hydroxide solutions (to a greater extent than with water).

$$Cl_2(g) + 2NaOH(aq) \rightarrow NaCl(aq) + NaOCl(aq) + H_2O(\ell)$$

Solutions containing the hypochorite ion, OCl^-, (a strong oxidising agent) are used to bleach wood pulp and textiles and as disinfectants. Household bleach contains about 5% sodium hypochorite, NaOCl by mass in water and is used to kill bacteria and mould. Calcium hypochorite, $Ca(OCl)_2$, is sold as a solid and is added to swimming pools to kill bacteria in the water. Sunlight slowly decomposes the hypochorite ion so calcium hypochorite must be constantly added to swimming pools to keep them sterile.

- Test for chlorine gas is the bleaching of damp litmus paper.

Industrial preparation of chlorine (**Chem12** p78): Produced commercially by the electrolysis of a concentrated solution of sodium chloride, NaCl (referred to as "brine"). Sodium chloride (table salt) occurs in sea water and is mined from salt deposits. In solution, the dissolved $Na^+(aq)$ and $Cl^-(aq)$ ions are attracted to electrodes that are connected to a power supply. The chloride ions donate electrons to the positive electrode (called the anode) and form chlorine gas, which bubbles from the solution.

$$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$$

Although the sodium ions are attracted to the negative electrode (called the cathode), sodium ions are very poor electron acceptors and they do not react. Instead, water molecules that come in contact with the cathode gain electrons from it to form hydrogen gas and hydroxide ions.

$$2H_2O(\ell) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$$

To prevent the electrodes from reacting with the chlorine or hydrogen gases produced, graphite is used as the inert material for the electrodes. The remaining solution of sodium hydroxide is collected and used in other manufacturing processes such as soap making.

Hydrogen chloride, HCl:

- Hydrogen chloride, HCl, is a colourless gas with a sharp smell.
- Obtained industrially by burning hydrogen gas in chlorine gas.

$$Cl_2(g) + H_2(g) \rightarrow 2HCl(g)$$

- Reacts with ammonia gas to form white "smoke" in the air of tiny particles of solid ammonium chloride, NH_4Cl.
- HCl is a strong acid. It dissolves in water and is completely converted to hydronium ions, $H_3O^+(aq)$, and chloride ions, $Cl^-(aq)$. This solution is known as hydrochloric acid, HCl(aq).

$$HCl(g) + H_2O(\ell) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

- Hydrochloric acid shows typical properties of a strong acid, i.e. neutralises bases, reacts with reactive metals to give an ionic metal compound and hydrogen gas as products, and reacts with metal carbonates to give an ionic metal compound, carbon dioxide gas and water as products.
- Hydrochloric acid, HCl(aq), is used to clean metal surfaces and in the manufacture of medicines, dyes, photographic paper and plastics.

Metal chlorides (**Chem12** p25-26)**:**

- Most metal chloride compounds are readily soluble in water. One exception to this is silver chloride, AgCl.
- The presence of chloride ions, $Cl^-(aq)$, in solution can be detected by adding a solution of silver nitrate, $AgNO_3(aq)$. A white precipitate of silver chloride quickly forms in the presence of chloride ions.

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

QUESTIONS:

1. Halogens are prepared commercially from the corresponding halide ions. What is the difference in the formula of a halogen and a halide ion?
2. The reaction of chlorine with water is somewhat unusual in that chlorine acts as both an oxidising agent and a reducing agent. The two chlorine-containing products are HOCl and HCl.
   a. Identify the oxidation number of chlorine in each of these as well as in elemental chlorine.
   b. Which product is the result of elemental chlorine donating electrons?
3. A concentrated aqueous solution of sodium chloride is referred to as brine.
   a. Which ions in brine react at the electrodes when electricity is passed through the solution? What is the formula of the gases that arise from these reactions?
   b. What is the formulae of the two ions remaining in the solution?

S T R A N D 5: ORGANIC CHEMISTRY.

Organic compounds contain chains or rings of carbon atoms. The carbon atoms are usually bonded to other carbon atoms and may also be bonded to hydrogen and other elements. The vast majority of known compounds are organic compounds.

HYDROCARBONS.

Students will investigate and develop their scientific understanding of **hydrocarbons** *when they:*

- Investigate why there are numerous carbon compounds.
- Define the following terms as important concepts in organic chemistry: homologous series, isomerism in compounds of up to five carbon atoms, functional groups and saturation.
- Investigate the sources of the naturally occurring hydrocarbons.
- Outline the process of fractional distillation of petroleum.
- Outline the physical properties of alkanes.
- Investigate the combustion of alkanes.
- Outline the physical properties of alkenes.
- Investigate selected reactions of ethene: addition, polymerisation to form polythene.
- Investigate the bromine test for unsaturation.
- Discuss the properties of alkynes.
- Investigate the laboratory preparation of ethyne (acetylene).
- Investigate selected reactions of the alkynes.

KEY POINTS:

- Carbon forms strong covalent bonds to itself and also other non-metal elements (e.g. H, O, N, S, P) and can form multiple bonds with itself and with O and N (**MatW** p94-95).
- 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).
- Hydrocarbons are organic compounds containing only C and H atoms. Their boiling points depend on the number of carbon atoms (**MatW** p96; **Chem12** p81).
- Mixtures of hydrocarbons occur naturally underground as natural gas and petroleum (**MatW** p110).
- Fractional distillation is used to separate hydrocarbons from the mixture known as petroleum (**MatW** p110-111).
- Alkanes are hydrocarbons having only single bonds between carbon atoms (each C atom is bonded to four other atoms) and are said to be "saturated". Alkanes have the general formula C_nH_{2n+2} (**MatW** p96-97; **Chem12** p81-82).
- Alkenes are hydrocarbons that contain at least one C=C (double) bond. Alkenes have the general formula C_nH_{2n} (**MatW** p98-99; **Chem12** p87-88).

- Alkynes are hydrocarbons that contain at least one C○C (triple) bond. Alkynes have the general formula C_nH_{2n-2} (**Chem12** p92).
- Alkenes and alkynes are said to be "unsaturated" as they contain at least one carbon-carbon multiple bond.
- 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).
- Compounds with the same molecular formula but a different atom-to-atom bonding sequence are called *structural isomers*. Isomers of hydrocarbons have different physical properties (**Chem12** p84).
- The names of organic compounds are based on the number of carbon atoms and indicate the type of organic compound by a particular ending (**MatW** p96; **Chem12** p84).
- Alkanes react with oxygen in a combustion reaction to produce CO_2, and H_2O. If insufficient oxygen is present, carbon monoxide, CO, or soot, C, are formed (**MatW** p106-107; **Chem12** p83).
- Unsaturated molecules, such as alkenes and alkynes, can undergo addition reactions in which the carbon atoms of a multiple bond become bonded to atoms from the other reactant. Bromine reacts with alkenes and alkynes by addition and the reaction mixture is decolourised (from orange) (**MatW** p99; **Chem12** p88-89, 93).
- 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).
- Ethyne, C_2H_2, (the simplest alkyne) can be easily produced by reacting solid calcium carbide, $CaC_2(s)$, with water to give ethyne gas and a solution of calcium hydroxide, $Ca(OH)_2(aq)$. Ethyne is also known as acetylene (**Chem12** p93).

QUESTIONS:

1. Which type of hydrocarbon is considered saturated?
2. What is an isomer?
3. Name the type of bond broken in an addition reaction.
4. What are the products that arise from the complete combustion of a hydrocarbon?
5. Name the reaction type that leads to the formation of polymers.

FUNCTIONAL GROUPS IN ORGANIC COMPOUNDS

Students will investigate and develop their scientific understanding of **functional groups** *when they:*

- Investigate the preparation of ethanol by fermentation.
- Outline the industrial preparation of ethanol from ethane.
- Discuss the uses of alcohols.

- Investigate the properties of carboxylic acids and reaction with alcohols to form esters.
- Outline the structure and properties of amines.

KEY POINTS:

- The functional group in an organic molecule is an atom or group of atoms at which a reaction occurs. Functional groups may contain multiply-bonded atoms (such as C=C) or bonds from carbon to oxygen or nitrogen (**Chem12** p87).
- Many reactions change one functional group in an organic compound to another.
- Alcohols have the functional group –OH (**MatW** p100; **Chem12** p95).
- Alcohols with only a few carbon atoms are soluble in water as they are polar molecules (**Chem12** p95).
- Ethanol is produced by the fermentation of glucose due to the action of enzymes from yeast (**MatW** p102-103; **Chem12** p96).
- Industrial preparation of ethanol relies on the addition reaction of water to ethene in the presence of aqueous sulfuric acid (**Chem12** p96).
- Alcohols are used as solvents, fuels, and as starting materials for other chemicals (**MatW** p100, 106).
- Organic acids are called carboxylic acids and have the functional group –COOH (**Chem12** p96, 99).
- Alcohols and carboxylic acids react to produce esters and water. Esters have the functional group $–CO_2C-$ and often have characteristic sweet or fruity smells (**MatW** p101; **Chem12** p97).
- Amines are organic derivatives of ammonia, NH_3, with at least one carbon group attached to the nitrogen. The simplest amines contain the functional group $–NH_2$, such as methylamine, CH_3NH_2.
- Small amines molecules are fishy smelling, water soluble and weakly basic (as they react with water molecules to a small extent to give hydroxide ions, $OH^-(aq)$ as one product).

QUESTIONS:

1. Draw and name the amine molecule with the formula C_3H_9N that has only one C-N bond.
2. Write the balanced equation for the complete combustion of ethanol.
3. Explain why small alcohol molecules are soluble in water.
4. Name the type of organic compound that has a sweet, fruity smell.

FOOD CHEMISTRY

Students will investigate and develop their scientific understanding of **food chemistry** *when they:*

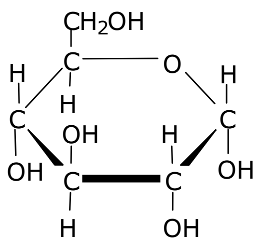
- Carry out simple food tests for carbohydrates, proteins, fats and oils.
- Compare the structure of carbohydrates, proteins, fats and oils.

To maintain a healthy body, there are several different types of food that are needed in a balanced diet. These food groups are:

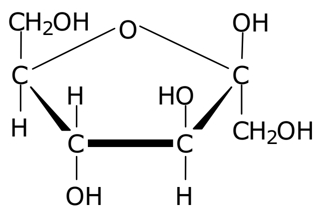
- Carbohydrates: provide an energy reserve for the body and dietary fibre.
- Proteins: needed for growth and cell repair.
- Fats and oils: store energy and for body insulation.
- Vitamins and minerals: needed in many chemical processes in the body.
- Water: for digestion, blood, sweat and chemical reactions in body (over 60% of a person is water).

CARBOHYDRATES:

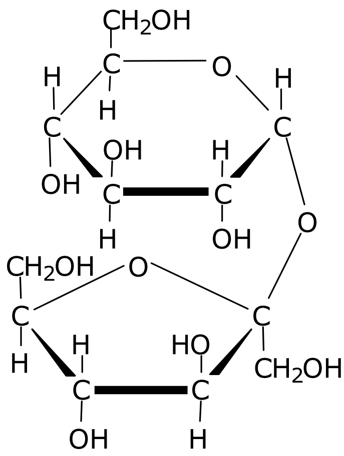
- Compounds consisting of carbon, hydrogen and oxygen atoms, and often have the general formula of $C_x(H_2O)_y$, e.g. glucose, $C_6H_{12}O_6$.
- Found in cereals, table sugar, vegetables, fruit.
- Most common carbohydrate is glucose, $C_6H_{12}O_6$, and is the primary source of energy for living organisms in respiration. Added to "high-energy" sports drinks.
- Glucose is a "simple sugar" known as a *monosaccharide*. Fructose is an isomer of glucose. Glucose and fructose occur in honey and fruit.



glucose



fructose



sucrose

- Table sugar is sucrose, which is formed from molecules of glucose and fructose joining together (with the loss of a water molecule). Sucrose is a *disaccharide* molecule and occurs in sugar cane.
- Monosaccharides and disaccharides contain several polar -OH functional groups and are water soluble.
- *Polysaccharides* are water-insoluble sugar polymers. Plants turn glucose into starch (made up of about 1500 glucose monomers) and cellulose (long chains made of about 10,000 glucose monomers). Starch is a high-energy food found in bread, potatoes, rice and pasta. Cellulose is the main structural component of plants.

Test for carbohydrates that are "reducing sugars" (such as glucose and fructose but not sucrose)*:*

1. When a few drops of Benedict's solution (a blue solution containing $Cu^{2+}(aq)$ ions) is added to a solution containing a reducing sugar and heated, an orange/red cloudy solution is observed (as a precipitate of copper(I) oxide is formed). The solution may also appear yellow or green, depending on the amount of sugar present. A similar result occurs if a few drops of Fehling's solution are added instead of Benedict's solution.
2. When a few drops of an aqueous solution of iodine are added to a colourless solution containing starch, a dark blue/black colour appears in the solution.

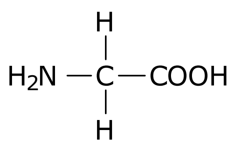
Storyline

Brewing in Samoa

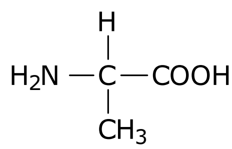
Samoa Breweries Ltd produces the popular beer, "Vailima". Beer is obtained from a process called "brewing". In this process, starch from grain is converted into smaller molecules by enzymes found in germinating grain. Yeast, a fungus, is then added which converts all the carbohydrates into glucose, $C_6H_{12}O_6$. The yeast also provides an enzyme that catalyses the fermentation of the glucose into ethanol, the alcohol present in beer. Fermentation also produces carbon dioxide gas that gives beer its bubbles. The characteristic bitter taste of beer is due to the addition of the dried flowers of the hop plant during the process. Further carbon dioxide gas is added to beer before it is bottled.

PROTEINS

- Proteins are used by the body for growth and repair of tissues, and as biological catalysts (enzymes and hormones) controlling of most of the chemical reactions in the body. Blood, hair, skin muscles and bones all contain proteins.
- Made by plants using water, carbon dioxide, and nitrogen-containing ions, such as nitrate ion, $NO_3^-(aq)$, from the soil.
- Found in meat, fish, lentils, milk, and eggs.
- Proteins are complex molecules consisting of long-chained polymers of amino acids. Insulin, the simplest protein, contains 51 amino acid units, while haemoglobin has 574 units.
- There are important 20 amino acids involved in the formation of proteins.
- Amino acids contain two reactive functional groups: an amine group, $-NH_2$, and a carboxylic acid group, -COOH. The amine part of the molecule reacts as a base, while the carboxylic acid part of the molecule reacts as an acid.
- The simplest amino acid is glycine. Different side-chains attached to the central carbon atom give rise to different amino acids, such as alanine.



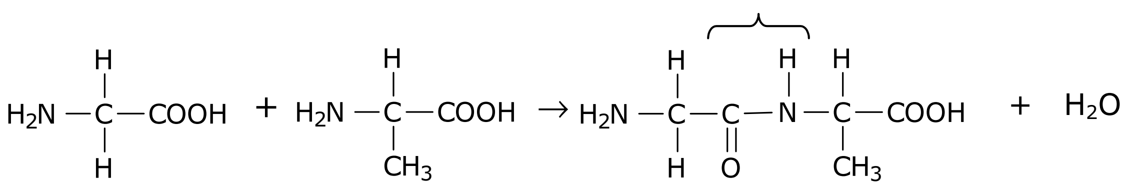
glycine



alanine

- During digestion, humans break down proteins into amino acids, which pass through the walls of the small intestine into the blood. Amino acids are then reassembled into proteins in cells.
- The $-NH_2$ group of one amino acid and the –COOH group of another amino acid react in a polymerisation reaction that eliminates water molecules and joins the amino acid units into long chains. The amino acids are linked by a *peptide* bond.

peptide bond



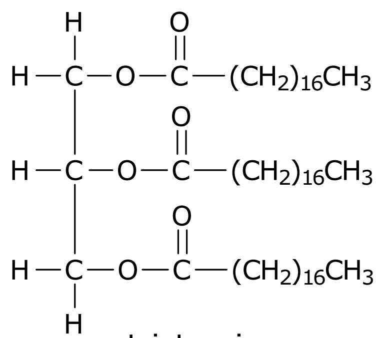
glycine alanine glycylalanine

- Dipeptides form from 2 amino acids; polypeptides form from many amino acids. A protein is a polypeptide with more than 50 amino acids.
- Amino acids are not stored in the body therefore a regular intake of protein is necessary.

Test for proteins: When a few drops of Biuret solution are added to a blue solution containing a protein, such as egg white, a pink-purple colour is observed in the solution, as the Biuret reagent (copper sulfate in a strong base) reacts with peptide bonds.

FATS AND OILS

- Provide a source of energy for the body. Fats release more than twice the energy per gram than carbohydrates. Fats are converted to glucose to release energy.
- Fats also provide insulation for the body as well as supporting and protecting delicate organs such as the kidneys.
- Occur in animals and plants. Foods high in fat include butter, cream, pork, and some nuts. Oils are found in seeds and fruit, including coconuts, olives, soya beans, and peanuts.
- Fats and oils are classified as "lipids" and belong to the class of ester compounds. Fats, such as tristearin, and oils result from the reaction of an alcohol, glycerol (also known as 1,2,3-propanetriol) with three long-chain carboxylic acids (known as "fatty acids"). The three fatty acids in a fat may be the same or different (**Chem12** p101). Fats may be saturated (carboxylic acid groups have only single bonds between carbon atoms) or polyunsaturated (carboxylic acid groups contain more than one C=C double bond).



tristearin

- Fats and oils differ in melting points: oils are liquids at room temperature while fats are solid. Generally saturated carboxylic acids solidify at room temperature as fats. Oils have unsaturated hydrocarbon chains and are liquids at room temperature.
- Fats and oils are non-polar and water insoluble.

Test for fats and oils:

1. When a few drops of cooking oil or melted fat are added to a portion of ethanol in a test-tube and the mixture is shaken, a white, cloudy emulsion forms.
2. When fat is smeared onto a piece of white paper, a translucent "stain" appears through which light can be seen.

QUESTIONS:

1. What element is always present in proteins which is not found in carbohydrates and fats?
2. Why is fructose called a monosaccharide?
3. How could the presence of a protein in a sample of food be determined?