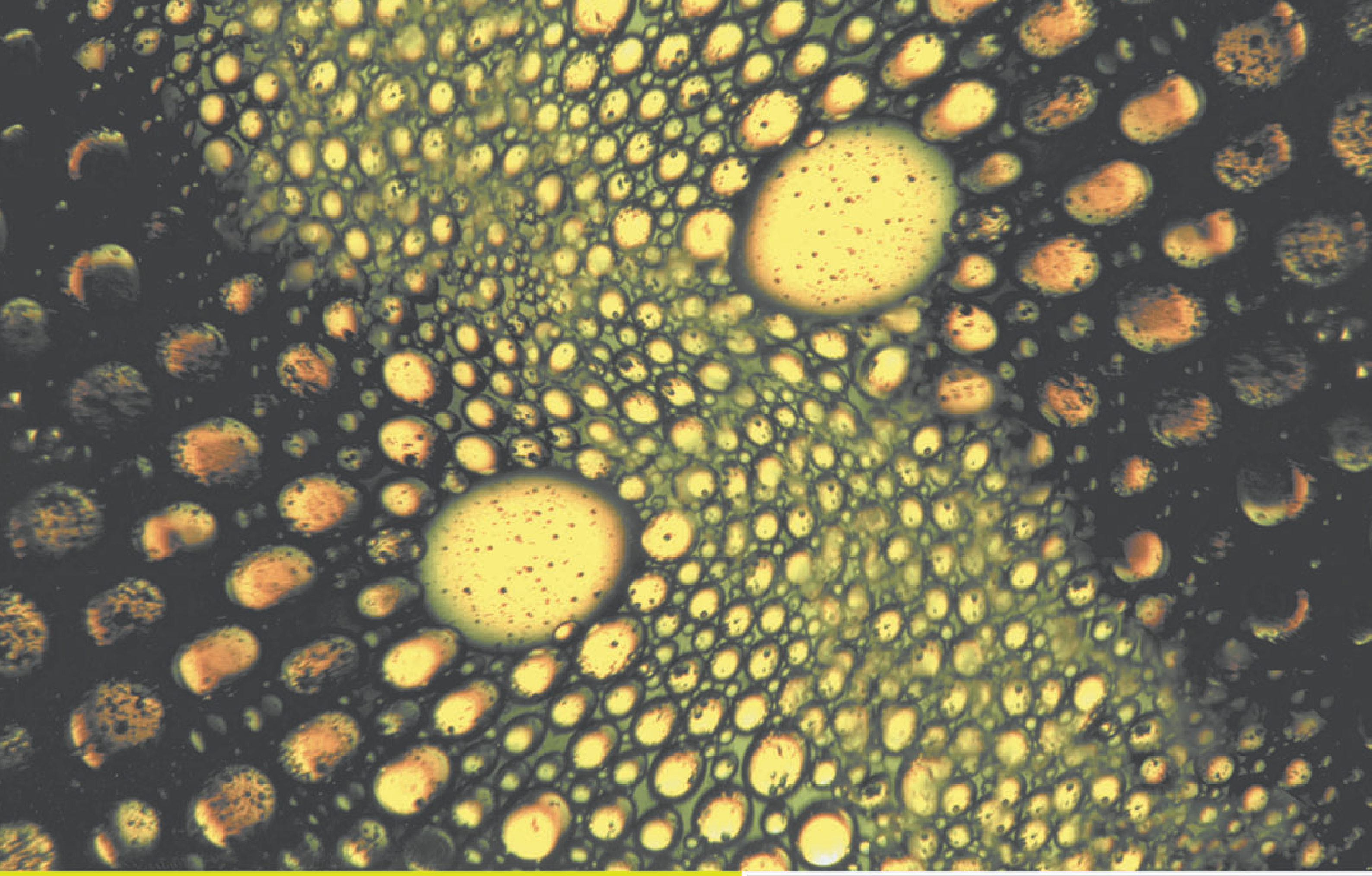

HEINEMANN
Queensland
Science Project

Chemistry

A Contextual Approach second edition

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Introduction

Heinemann Queensland Science Project—***Chemistry: A Contextual Approach second edition*** has been fully updated to support the Queensland 2007 Chemistry Syllabus.

The single textbook covers Years 11 and 12 and presents chemistry in a variety of real-world contexts. This provides for flexibility of planning, enabling topics to be selected and studied in an order that suits the teachers and students.

The organisation of the new Queensland Chemistry 2007 Syllabus involves planning units of work that focus on general objectives categorised **Knowledge and Conceptual Understanding**, **Investigative Processes** and **Evaluating and Concluding**. Seven **Key Concepts** are the foundation stones of the syllabus, two in the area of Structure and five in the area of Reactions. Any course of study must allow students to develop understandings of these key concepts embedded in at least two real-world **contextualised units of work**. Also listed in the syllabus are expansions of the **Key Concepts** identified as **Key Ideas**. These are there to assist teachers understand the breadth and depth to which the **Key Concepts** should be studied.

Chemistry: A Contextual Approach is structured according to seven Areas of Study corresponding to the Key Concepts. For each Area of Study there are:
- Contexts—real-world examples of the importance of Chemistry
- Chapters—developing in detail the Key Concepts and Key Ideas of the syllabus.

The Contexts

Contexts are linked to the Key Ideas within the chapters via clear references. Experiments and other activities within a context are aimed at providing practice or actual assessment items according to the assessment categories of the syllabus. As students engage with these activities they will develop a greater understanding of the **Key Concepts** and **Key Ideas** from the syllabus.

Teaching contextualised units of work means that assessment should also be in context. In the 2007 syllabus three categories of assessment are identified:
- Extended experimental investigation
- Supervised assessments
- Extended response task.

Activities within the contexts have the potential to be used for assessment and are categorised as:
'Consider This'—Using the information of the context as a stimulus, students are asked to consider consequences of the Chemistry involved.
'Further Investigations'—Areas of further research with suggested media of presentation.
'Experimental Investigations'—Areas of extended experimental investigation. These are often based on the experiments encountered in the contexts and related chapters.

The Chapters

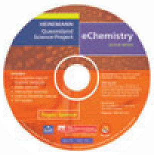
Chapters contain the chemical concepts and ideas of the syllabus and are essential to understanding the contexts. They include easily accessed summary points, experimental investigations and clearly worked examples. Each section of the chapters has related Questions, and at the end of each chapter there is an extensive range of questions in the Chapter Review. Answers to questions are included in the text.

Support material

eChemistry second edition

The *eChemsitry* Student CD-ROM includes an electronic version of the textbook, extra contexts and fully Interactive Tutorials that model and simulate key chemistry concepts. It also includes a direct link to ChemCAL.com.au

ChemCAL.com.au

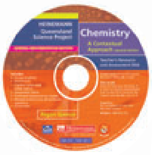


ChemCAL is a collaboration between the Chemistry Education Association, the School of Chemistry at The University of Melbourne and Heinemann. This interactive website allows students to revise key ideas and basic skills and extend their knowledge of applications and examples. Students can monitor their progress through interactive quizzes throughout, which provide instant feedback on progress. To access ChemCAL follow the instructions on the registration card contained in the back of this book.

Teacher's Resource and Assessment Disk second edition

The ***Teacher's Resource and Assessment Disk*** includes worksheets for each chapter, PowerPoint presentations, fully worked answers to all questions, experimental notes and safety advice, sample course outlines and sample assessment items.

Importantly **Sample Course Outlines** demonstrate ways of incorporating the requirements of the syllabus into a course of study presented over the Years 11 and 12.



Chemistry: A Contextual Approach can be used as a resource for teachers and students as they plan for and work through the selected contexts and chapters. Sample assessment tasks are provided for the categories — Extended experimental investigation, Supervised assessments and Extended response task. Sample criteria sheets for each of these tasks demonstrate how the criteria can be individualised for each assessment task but still link closely with the standards associated with the exit levels of achievement in the syllabus.

About the authors

Regan Spence—Coordinating author
Regan has over twenty years experience teaching science, in particular chemistry. He has been head of a science department, a panelist on district and state panels, and a member of the chemistry subcommittee throughout the trial pilot, extended trial pilot and development of the current Chemistry Syllabus. Regan is now employed by the QSA as a senior education officer. He is committed to making chemistry relevant and engaging for students and sees the use of meaningful and stimulating contexts as the key to achieving this. He is an experienced author and has been involved in the creation of award winning textbooks and CDs.

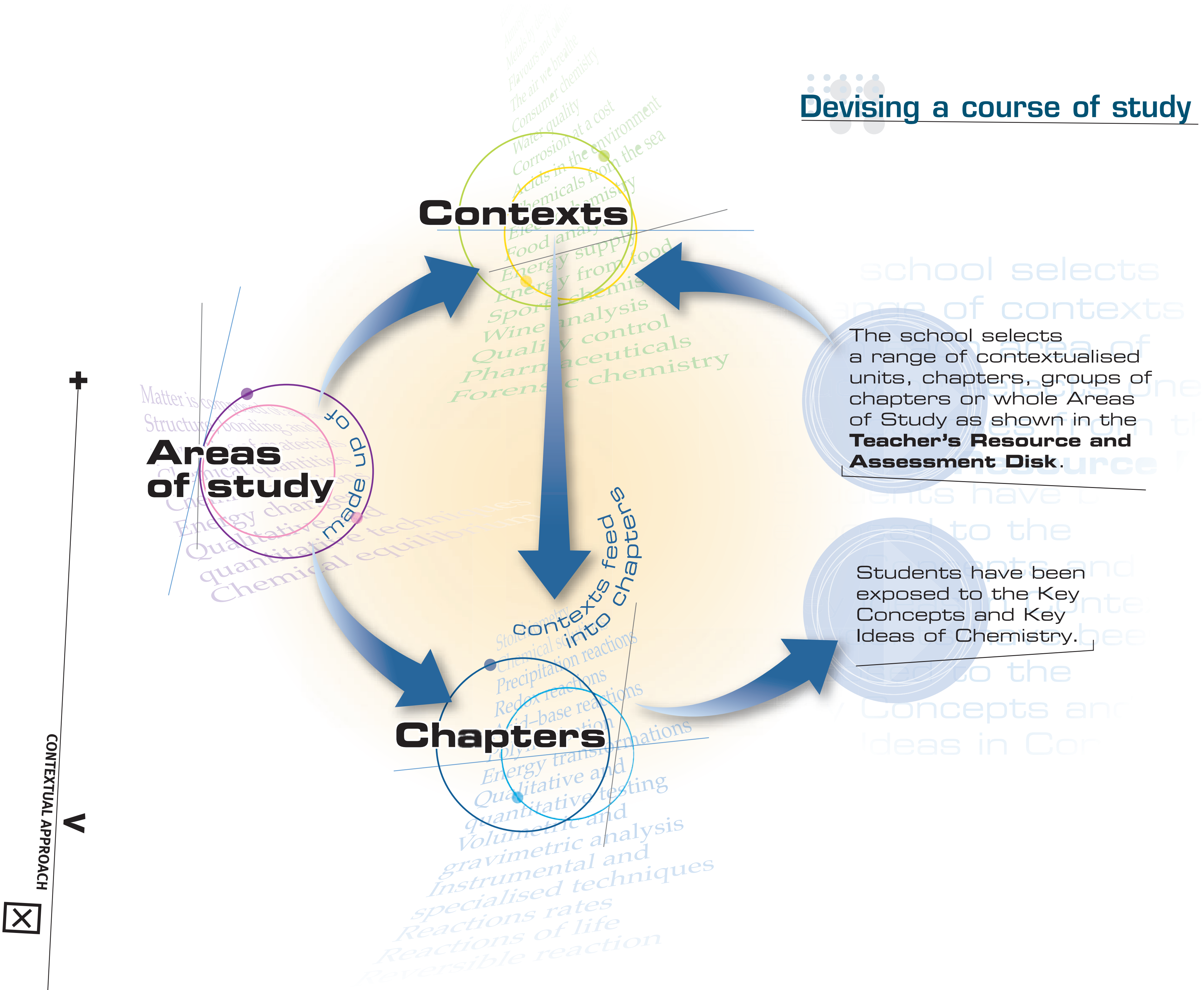
Lee-Anne Bramley
Lee-Anne has been teaching senior chemistry for the last nine years and is currently head of department at Palm Beach-Currumbin State High School. She is a member of the District Chemistry Panel.

Teresa Gemellaro
Teresa has over twenty years teaching experience in chemistry and physics at both state and private schools in rural Queensland. She is currently teaching at Ayr State High School where she has a particular focus on course structures and assessment. She has developed assessment items that have been published by the Townsville and District Science Teachers Association and other publishers. She is currently a member of the State Panel for Chemistry.

Donna Wilson
Donna has over twenty years teaching experience including schools in far-north Queensland, Brisbane and the Gold Coast. She currently teaches chemistry at Palm Beach–Currumbin State High School. She has been a level coordinator and for many years was involved in SCINET, a local Gold Coast science teachers association, where her responsibilities included judging the local science competition. She has a particular interest in biochemistry and organic chemistry.

Sam Wiseman
Sam has over twelve years experience teaching senior chemistry and biology in Queensland schools. He has a particular interest in biochemistry, and sees this as one of the many areas in which stimulating, relevant, contextual chemistry teaching and learning can be achieved. Sam is a demonstrator for RACI in Queensland primary schools, and has been head coach of the Soccer Sports Excellence program at Palm Beach–Currumbin State High School for the last six years (School National Champions 2001).

Devising a course of study



The school selects a range of contextualised units, chapters, groups of chapters or whole Areas of Study as shown in the **Teacher's Resource and Assessment Disk**.

Students have been exposed to the Key Concepts and Key Ideas of Chemistry.

Areas of Study

AOS 1 Matter is composed of atoms
Chapter 1 Atomic structure
The beginnings of chemistry
Where do elements come from?
Elements of life*

AOS 2 Structure, bonding and properties of materials
Chapter 2 Elements and compounds
Atmospheres
Metals by design
Flavours and odours
Crystals*
Chemisty and the home*

AOS 3 Chemical quantities
Chapter 3 Stoichiometry
Chapter 4 Chemical solutions
Chapter 5 Gases
The air we breathe
Consumer chemistry
Water quality
Soil chemistry*

AOS 4 Chemical reactions
Chapter 6 Precipitation reactions
Chapter 7 Redox reactions
Chapter 8 Acid–base reactions
Chapter 9 Organic chemistry
Chapter 10 Polymerisation
Corrosion at a cost
Polymers
Soaps and detergents
Mining in the North
Chemicals from the sea
Electrochemistry
Food analysis
Acids in the environment*
Metals and the modern lifestyle*
Chemistry in archeology*

AOS 5 Energy changes
Chapter 11 Energy transformations
Energy sources
Energy from food
Sports chemistry*

AOS 6 Qualitative and quantitative techniques
Chapter 12 Qualitative and quantitative testing
Chapter 13 Volumetric and gravimetric analysis
Chapter 14 Instrumental and specialised techniques
Wine analysis
Quality control
Pharmaceuticals
Forensic chemistry

AOS 7 Chemical equilibrium
Chapter 15 Reaction rates
Chapter 16 Reactions of life
Chapter 17 Reversible reactions
The chemical industry
Food
Marine chemistry
Biotechnology *

*These contexts are on the *eChemistry* second edition Student CD.

Contents

Contexts

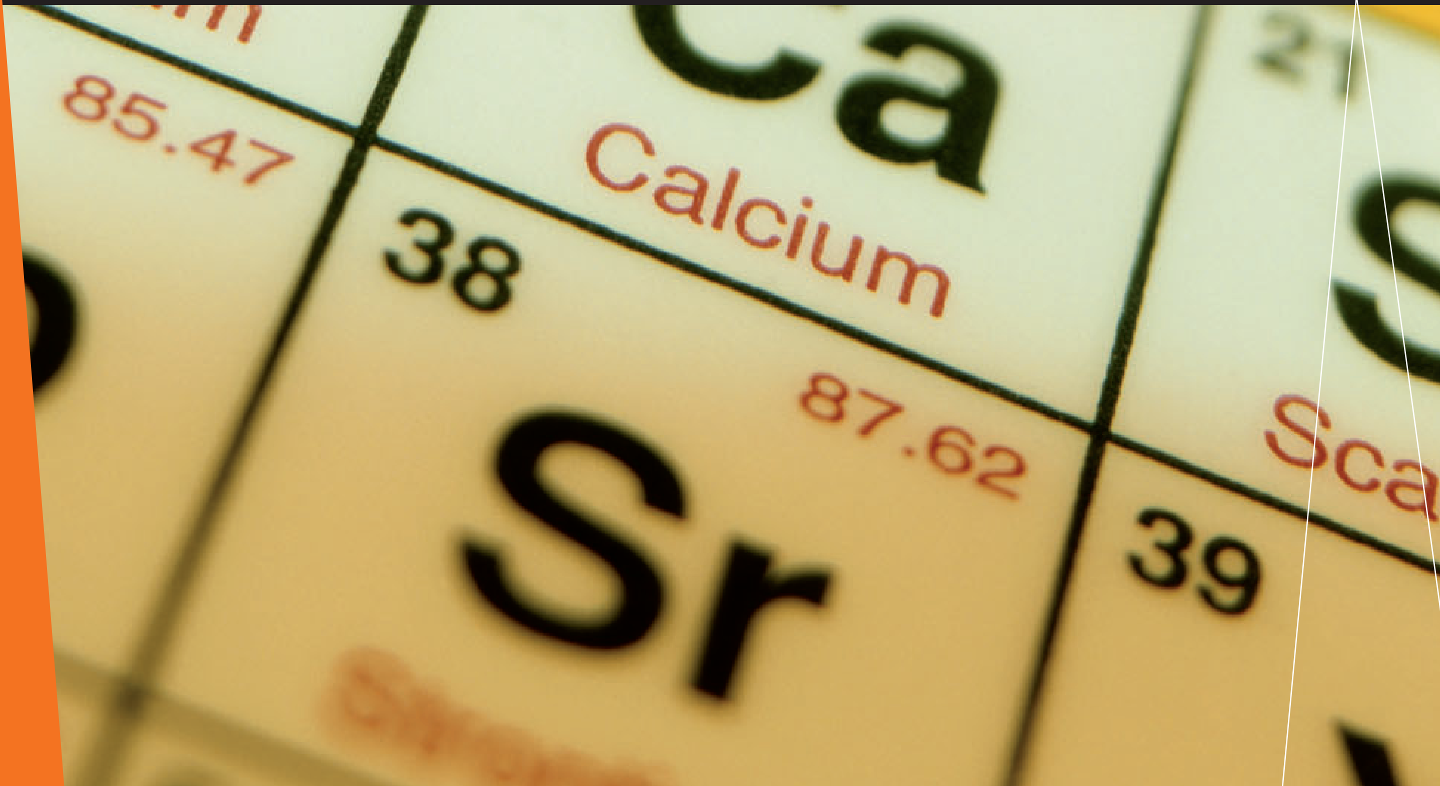
Chapters

Contexts
The beginnings of chemistry
Where do elements come from?

Chapter
1 Atomic structure

area of Study 1

Matter is composed of atoms



the beginnings of chemistry

When did the science of chemistry begin? The answer to this question is not a simple one. Many people would answer that chemistry began with Robert Boyle in the 1600s or perhaps with Joseph Priestley and Antoine Lavoisier in the 1700s. However, if you consider that the study of chemistry is concerned with the composition, properties and reactions of substances, you will realise we need to look much further back for the beginnings of chemistry.

Chemistry in ancient times

Since the beginnings of recorded history, humans have tried to explain or describe natural phenomena. Early humans were fascinated by fire, thunder and lightning, and common substances such as soil, water and air. When humans first discovered fire, they began learning to control and use chemical changes. From this time onwards, they became practising chemists, even though they could not explain the changes they were using.

figure bc.1 Early humans unknowingly used fire to make chemical changes.

figure bc.2 The development of the alloy bronze was a major event in human history.

figure bc.3 Ancient civilisations who used steel to create superior weapons dominated their less well-equipped rivals.

Early civilisations first utilised chemical reactions in about 4000 BC when the Egyptians developed a glaze for their pottery. Today, we use the same ingredients and in the same quantities to make window glass.

By 2000 BC, an alloy of copper and tin, known as bronze, was widely used. Bronze is made by heating copper and tin ores together. Appropriately, this time in history is known as the Bronze Age.

Bronze was the most commonly used metal alloy until a civilisation called the Hittites discovered how to extract iron from iron ore in about 1500 BC. The addition of carbon to iron to produce steel heralded the Iron Age of history, during which empires expanded on the strength of their iron products and weapons.

At about this time, the ancient Egyptians were exploring other areas of chemistry using pigments and plant substances to preserve bodies for mummification.

By 600 BC, the Greeks had begun to question the nature of substances and in 450 BC the Greek philosopher Leucippus proposed the idea that the subdivision of matter would eventually lead to a piece so small that it could no longer be divided any further. A follower of Leucippus called Democritus expanded on this idea. He named this piece *atomos* (which means 'indivisible') and believed that the atomos of each substance had a different size and shape, and gave substances their different properties. Unfortunately this idea was not widely accepted until a few hundred years later when the Romans adopted the idea. The ancient Greeks preferred the ideas of the philosopher Aristotle, who believed that all matter was composed of four elements: earth, air, fire and water. Aristotle proposed that each element had its own properties unique to that element and that substances in nature differed due to the ratio of the four elements that they contained.

figure bc.4 Ancient Egyptians made use of chemical substances in the preservation of bodies.

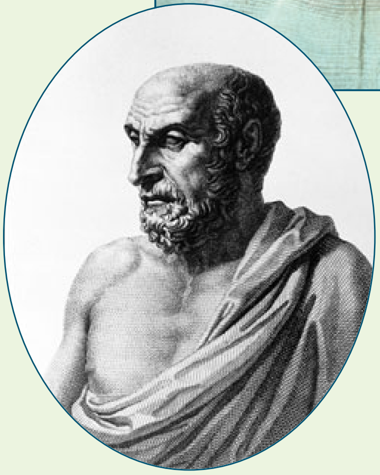


figure bc.5 Democritus proposed the name atomos for tiny indivisible particles. More than 2000 years later, Dalton used the name atoms in his famous atomic theory.

* Section 1.1 **The structure of atoms** describes the modern model of atoms (page 209).

The alchemists

Alchemy is often described as a pseudo (false) science. It was pursued by several early civilisations including the Arabs, the Chinese and the Egyptians. The best known pursuit of alchemy was the attempt to change base (common) metals into gold. Some alchemists also tried to find the 'panacea' (cure for all diseases) and 'chin tan' (a way to prolong life).

We now know that it is impossible to simply change metals such as lead into gold. It did not seem so impossible based on the observations made by alchemists of the time. They observed that by altering conditions, one substance could be changed into another; for example, water could be converted to steam and wood could become fire and air. It seemed logical, therefore, that one metal could be changed into another.

At the end of the third century, Roman emperor Diocletian banned the writings of alchemy. With the rise of Christianity in the fourth century, alchemy was considered a pagan pursuit and became an obscure art.

Alchemy was revived in the 600s by the Arabs, who drew from the ideas of ancient Greek science. At the end of the eleventh century, European Christians launched the first crusade on the world of Islam. This led to Christian occupation of the Holy Land for two centuries. During this time, Arabic science began to filter back to Europe. By the mid 1100s, English and Italian scholars had translated Arabic alchemy into Latin, making it accessible to European scholars. In this way, the Arabic alchemists contributed many important ideas to European science, such as the preparation of plaster of Paris and its role in setting broken bones. Using the knowledge gained from the Arabs, German alchemist Andreas Libau published the first chemical textbook considered to be of worth, *Alchemia*, in 1597.

figure bc.6 Although alchemy was not considered a true science and often shrouded with mysticism, the work of the alchemists gave rise to the science of chemistry.

figure bc.7 For hundreds of years, alchemists tried to turn common metals such as lead into gold. Many claimed success over the years; however, none of them became very wealthy from their supposed successes!

From alchemy to chemistry

The importance of alchemy began to decline by the late seventeenth century and by the early eighteenth century, the term alchemy was used to describe attempts to turn base metals into gold. The rise of modern chemistry had begun.

Many scientists working in the early days of chemistry were interested in the study of gases. In 1643 an Italian scientist Torricelli demonstrated that air could exert pressure on things; this led him to invent the barometer which is used to measure air pressure. This encouraged other scientists to investigate the properties of air.

The Irish chemist Robert Boyle was particularly interested in air. He made a pump that could be used to compress air. His experiments led him to propose that increasing the pressure on a sample of air decreased its volume and decreasing the pressure on a sample of air increased its volume. We still use this relationship today when describing the behaviour of gases.

Boyle believed gas was made of particles with large spaces between them because they could be compressed. In thinking this he revived the idea of substances being made up of atoms. In 1661 Boyle published *The Skeptical Chymist*. In this book he first used the term elements in the sense that we use it today, and he proposed that all substances should be considered elements until tested and proven otherwise.

Early modern chemists were preoccupied with trying to isolate elements, but it was not until the eighteenth century that a significant number of new elements were discovered.

figure bc.8 Robert Boyle proposed a relationship between pressure and volume of gases which we still use today.

The era of modern chemistry

Over 100 years after the publication of Boyle's work, scientists studying gases were to herald the era of modern chemistry with their discoveries.

Englishman Joseph Priestley began researching gases as a hobby, however he approached it in a much more systematic way than had been done before. In 1774, he isolated oxygen gas (which he called dephlogisticated air) from the heating of mercury(II) oxide. At about the same time, Swedish scientist Karl Wilhelm Scheele also prepared oxygen. Unfortunately for Scheele, his results were not published until much later than Priestley's so his achievement is often overlooked.

Priestley believed in phlogiston theory. According to this theory, combustible substances contained phlogiston, which was released when they were burned, leaving ash.

After listening to Priestley speak, Frenchman Antoine Lavoisier repeated Priestley's experiments. Lavoisier realised that the gas obtained from the experiments supported combustion and was essential to life. He called the gas oxygen. In 1789, Lavoisier published his work *Traité Élémentaire de Chimie* (Elements of Chemistry). This groundbreaking work was the first real presentation of what we call modern chemistry. The work described ideas such as:

- the Law of Conservation of Matter
- the nature of gases
- acid–base reactions
- the effect of heat on reactions
- a list of all known elements of the time.

Unfortunately for Lavoisier, personal circumstances made him a target for the guillotine in the French Revolution and he was beheaded in 1794. When considering his great contribution, many consider Lavoisier the father of modern chemistry.

Lavoisier's demonstration that oxygen gas supported combustion disproved the phlogiston theory.

Along came atoms

The work of Priestley and Lavoisier inspired other scientists to perform more accurate experiments and carefully examine the results. One of these was John Dalton, an English mathematics tutor. He repeated their experiments and performed some of his own. He noticed that elements mixed in simple whole

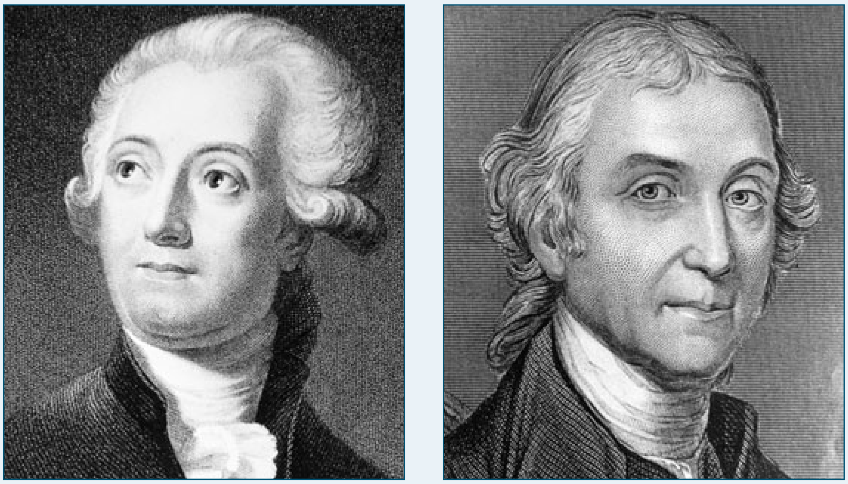


figure bc.9 Lavoisier (left) proved that the substance obtained from Priestley's experiments, which supported combustion and was essential to life, was oxygen and not phlogiston.

number ratios to form compounds. This became the Law of Multiple Proportions. Dalton then took his ideas even further by examining the reactions on the level of individual particles. Dalton recalled Isaac Newton's comments concerning 'impenetrable, moveable particles' which may 'remain by themselves or also compose bodies'. Dalton used the term atoms (from the Greek *atomos*) for individual particles and set out his atomic theory. In 1808 he published *A New System of Chemical Philosophy*. His atomic theory was revolutionary and proved to be of great value in explaining chemical reactions. This new way of thinking met with a mixed reception; some were eager to accept it, some hesitant and others ridiculed his new proposals. Despite this, describing chemistry in terms of things that could not yet be seen was a great step forward in the development of theoretical chemistry.

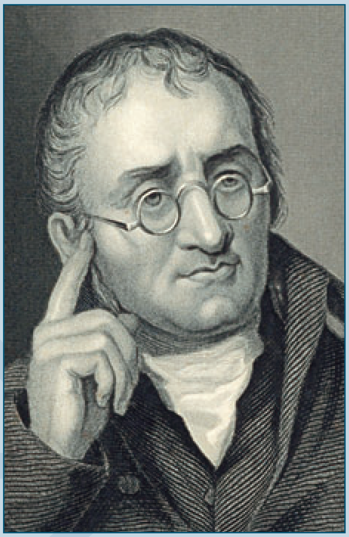


figure bc.10 Dalton's atomic theory was revolutionary in its time and explained much about chemical substances and reactions.

Organising the elements

Dalton used balls of different colours or markings to represent different atoms. A better system for representing different elements was proposed by Swedish chemist Jons Jacob Berzelius in 1819. By this time, nearly 50 elements were known. He proposed that it would be more convenient for each element to be represented by a unique letter symbol—either a single letter or a combination of two letters (the first being a capital letter).

Section 1.2 **Elements and isotopes** discusses the symbols used to represent elements and how the symbols relate to the element (page 211).

Berzelius also suggested a notation for writing the element symbols to describe compounds; for example, he wrote the formula for water as HOH. We have modified his approach; today we write H_2O as the formula for water. He also assigned elements relative atomic masses based on oxygen as 100 atomic mass units. Some of his masses were incorrect, however Berzelius' work inspired other chemists.

Section 1.3 **Atomic mass** describes the modern view of relative atomic mass (page 213).

In 1864 English chemist John Newlands prepared a table in which the elements were arranged in order of increasing atomic weight. His work was not taken seriously or published because it was believed that there were too many coincidences. In 1870 German chemist Julius Lothar Meyer noticed trends in properties of elements and organised them into groups of eight. Unfortunately, Russian chemist Dmitri Mendeleev published the definitive classification of the elements one year earlier than Meyer's proposal. Mendeleev approached the task by looking at the valence electrons of each element. His work was comparable to, but improved upon, the work of Newlands and Meyer.

* Section 1.4 **Electronic structure** describes valence electrons (page 215).

Mendeleev arranged elements with similar chemical properties into vertical groups and elements with increasing atomic mass into rows (or periods). At the time Mendeleev created his table, 62 elements were known. He distinguished his work from that of Meyer and Newlands by placing some elements out of order of established atomic weights. He also left gaps to represent undiscovered elements and predicted properties for some of these undiscovered elements. Elements discovered later matched the properties predicted by Mendeleev's table. Even though the periodic table was incomplete in Mendeleev's lifetime, it inspired the search for other unknown elements.

* Section 1.6 **Organising information—the periodic table** describes the structure of the modern periodic table (page 218).

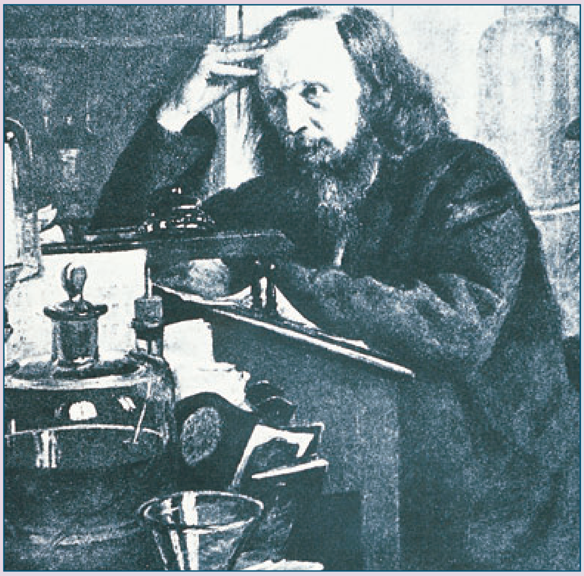


figure bc.11 Dmitri Mendeleev provided the earliest structure of the periodic table that is used today.

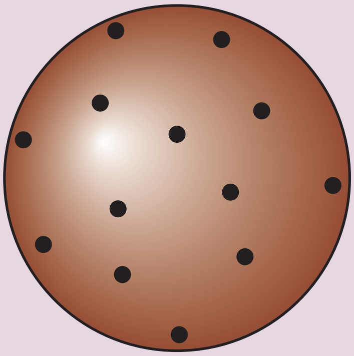


figure bc.12 Thomson's plum-pudding model proposed that electrons were scattered randomly throughout the positively charged sphere of the atom.

Later atomic theory

By the 1900s it was accepted that atoms were composed of positively charged and negatively charged particles. The negatively charged particles, called electrons, were discovered by J. J. Thomson in 1897. His plum-pudding model proposed that the electrons were spread throughout the positively charged sphere of the atom like raisins in a plum pudding.

It soon became apparent through the observations of other chemists that Thomson's model was not correct. In 1910, experiments performed by a group led by New Zealand born Ernest Rutherford changed the accepted view of the structure of the atom.

Rutherford shot a beam of radioactive positively charged alpha particles at a sheet of gold leaf (very thin gold). Most alpha particles passed straight through the gold leaf, producing flashes on the fluorescent screen on the other side. However it was also observed that some passed through but had been deflected, and others hit the gold leaf and bounced back. They concluded that atoms must be mostly empty space to have allowed the majority of alpha particles through and have a very small positively charged area to have deflected the alpha particles. They called this area the nucleus.

Using the results of these experiments, in 1911 Rutherford proposed a new model for the structure of the atom. His atom had a nucleus of positive charge which was orbited by much smaller negatively charged electrons in roughly circular orbits, similar to the way the planets orbited around the sun. This model was referred to as the solar system model.

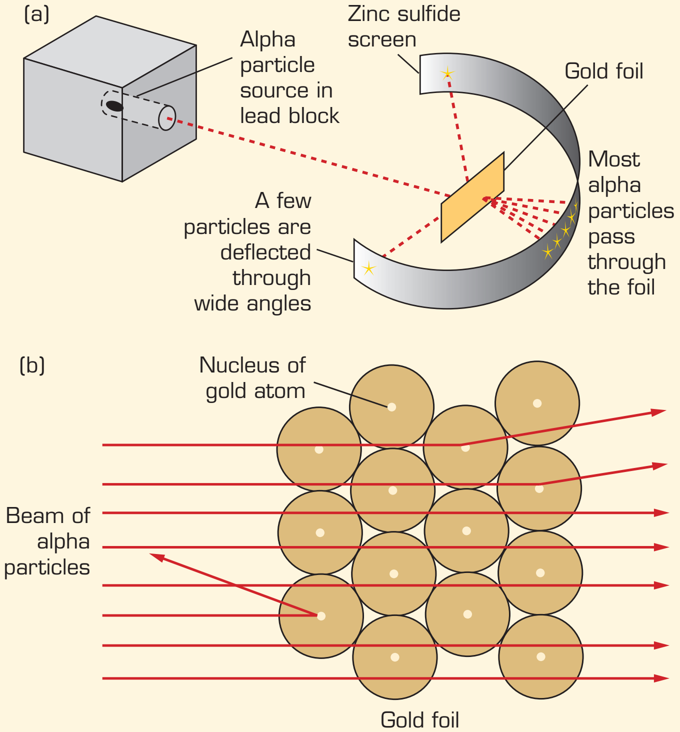


figure bc.13 Rutherford's alpha particle experiment.

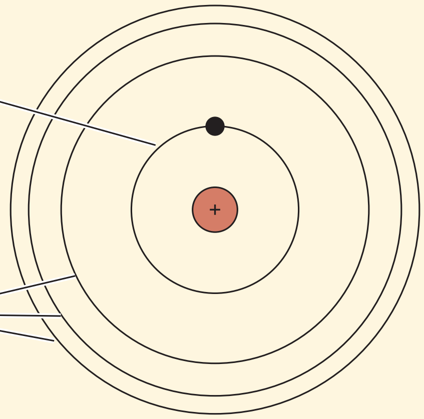


figure bc.15 Bohr's model of the atom was called the shell model. The electrons of an atom are arranged in these energy levels (shells) from lowest energy to highest energy.

Rutherford's atomic model proved to have some flaws. According to accepted classical laws of physics, electrons orbiting in circular paths should emit electromagnetic radiation; this had not been observed. More importantly, if electrons did emit this radiation they would lose energy and spiral into the nucleus. This meant that atoms would only exist for a very short time. Rutherford's model also did not explain the observation that light of specific energies is emitted when atoms are heated.

Danish physicist Niels Bohr proposed a model for the atom in 1913. This model refined Rutherford's model and accounted for some experimental observations not explained by Rutherford's model. His model explained why electrons could move in orbits without losing energy by suggesting that the orbits in which the electrons moved had a fixed energy—orbits closest to the nucleus had the lowest energy and orbits farthest from the nucleus had the highest energy. Bohr's atomic model suggested that the classical laws of physics did not hold for the orbits of electrons in atoms. Bohr's model also explained the emission of light of specific energies (emission spectra). He drew upon the work of German scientists Planck and Einstein which showed that light was emitted in packets of energy which he called quanta of energy. Bohr proposed that if an electron gained the right amount of energy it would be excited to a higher energy level. It would then very quickly return back to the lower energy level, releasing the absorbed energy and giving off a quantum of electromagnetic radiation (such as light). This explained why light of specific energies is given off when atoms are heated.

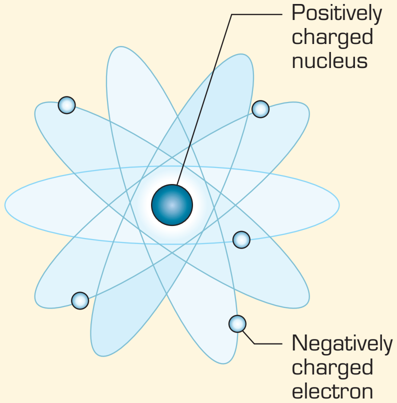


figure bc.14 Rutherford proposed that the electrons orbited around the tiny, positive nucleus in a similar fashion to planets orbiting the sun.

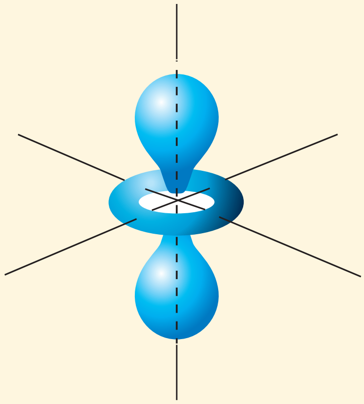


figure bc.16 A representation of electron orbitals as described by quantum mechanics.

*Section 1.4 **Electronic structure** describes the energy levels of electrons (page 215).

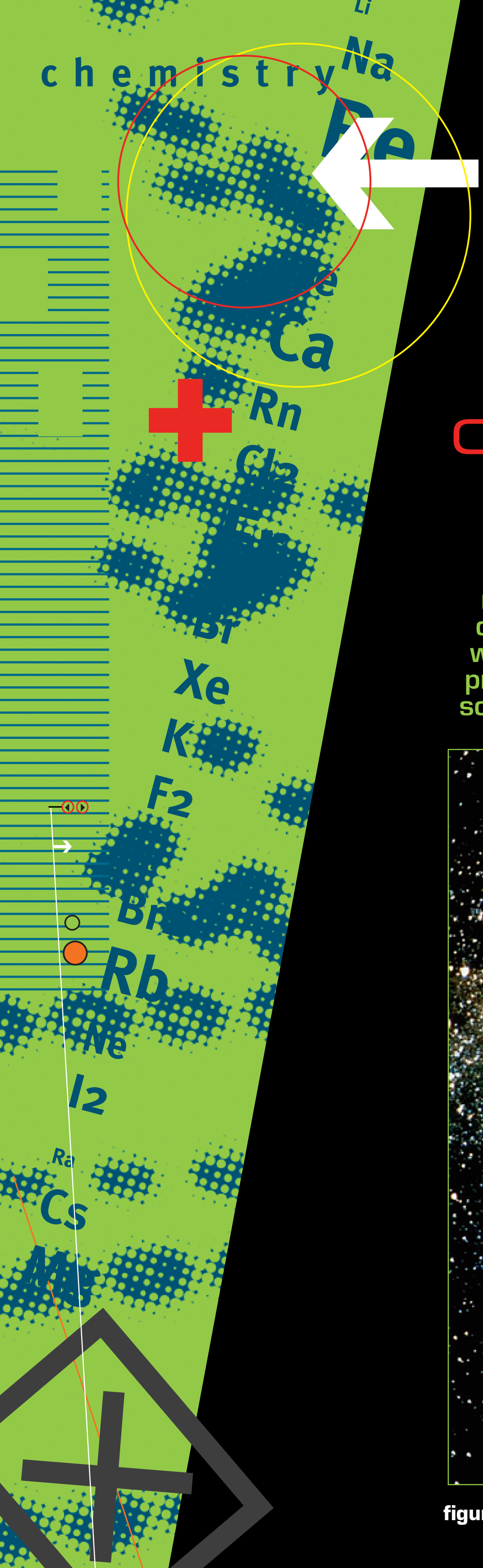
The work of Louis de Broglie in 1923 showed that particles (for example electrons) could show wave-like behaviour. Erwin Schrödinger used this idea to propose that electrons behave in a wave-like fashion around the nucleus. This was the beginning of quantum mechanics which is still in use today.

Further investigations

The history of chemistry is an enormous field that could easily occupy a book of its own. This context has mentioned a few of the people who made important contributions in our chemical history. For every contribution included, dozens were omitted.

1. Dalton's atomic theory was a cornerstone for future chemical research. Prepare a report explaining the main points of Dalton's atomic theory, including diagrams where necessary. Discuss why his theory was controversial for its time and how it paved the way for future discoveries.
2. Prepare a time line showing major events in the development of chemistry from ancient times to the present.
3. Often a brief description of a chemist's contribution does not reflect the essence of their life or work. Some chemists are renowned for being unusual characters, such as Priestley, or were taken before their time, such as Lavoisier. Some, such as Scheele, are not remembered for their work as they came in second.

   Choose a chemist who made a significant contribution to chemical history and prepare a short biography reflecting their life, their work and how they influenced others.
4. A major part of chemical history was the discovery and isolation of the chemical elements. Choose a chemical element and research its history, including areas such as discovery, naming, properties, uses and important reactions. Use this information to prepare a fact card for your element that could be used by other students as a resource when researching the element.
5. The elements of the periodic table derive their names from many different sources. For example, some are named after places, some are named after colours and some are named after important people. Some symbols are logically related to the element such as O for oxygen, whereas others seem to have no relation to their name, for example Na for sodium. Prepare a booklet describing the origin of the names of the first twenty elements. This booklet is to be used as a resource for Year 10 students researching one of the first twenty elements.
6. The discovery of the subatomic particles (protons, neutrons and electrons) was an important stepping stone in the development of atomic theory. Prepare a report on the discovery of these subatomic particles, including diagrams where possible.
7. By the 1920s, scientists had begun to propose alternate theories that made up for the deficiencies in Bohr's shell model. This field of study was called quantum mechanics. Prepare a PowerPoint or other verbal or visual presentation outlining the basic principles of this field.
8. The Solvay conference of 1927 was a major event in the development of quantum mechanics. A number of the participants went on to win a Nobel Prize. Research the Solvay conference and prepare a chapter for a chemical reference book describing its importance to the development of quantum mechanics.



where do elements come from?

Most scientists agree that our universe was created about fifteen billion years ago in an explosion called the Big Bang. At this exact moment, everything was pure energy. A tiny fraction of a second later, a seething mass of exotic particles were formed. These soon began to combine to form protons and neutrons. About three minutes later, some of these began to form small atomic nuclei.

figure oe.1 The Big Bang is considered to be the beginning of the universe.

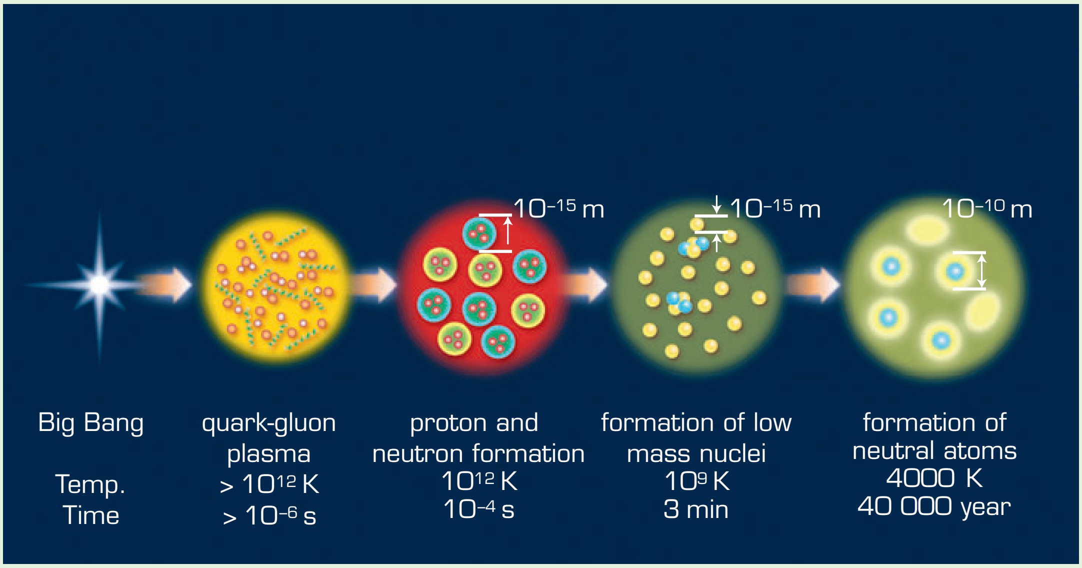


figure oe.2 The beginning of the universe, starting with the Big Bang.

At the time when these small nuclei began to form, the protons and neutrons were very hot and moving very quickly, and could easily combine. However as the universe rapidly expanded, the density and temperature dropped. This was so rapid that only simple nuclei had time to form. Most of the matter formed was hydrogen-1 (simple hydrogen nuclei with one proton and no neutrons). A small and significant amount of helium nuclei formed and trace amounts of the heavier nuclei lithium, beryllium and boron were also created.

It was so hot that these nuclei remained positively charged ions for about 400 000 years; only then was it cool enough for the nuclei to capture electrons and form neutral atoms. Figure oe.2 shows this sequence of events.

Only the simplest elements formed in the time soon after the Big Bang. There are more than ninety naturally occurring elements on earth, but elements with larger, more complex nuclei did not form in this primordial process. So, you may ask, where did the elements that exist today come from? The answer is that they have been created in stars.

* Section 1.1 The structure of atoms describes the structure of atoms, and the particles that make up atoms (page 209).

Questions

1 Name three important elements in your body that were not present in the early universe.

2 Find out which elements are most common in the Earth's crust. Were any of these present in the early universe?

Early life of stars

By observing a huge number of stars, astronomers developed a picture of the life of a star.

In the early universe, there were no stars. The first generation of stars began to form about one billion years after the Big Bang. Clouds of the original gas mixture formed, like the gas cloud in figure oe.3. The force of gravity in such large gas clouds was enough to draw gas towards regions of greater density. The gas heated up as it fell inwards and the result was a large, dense ball of hot gas called a protostar.

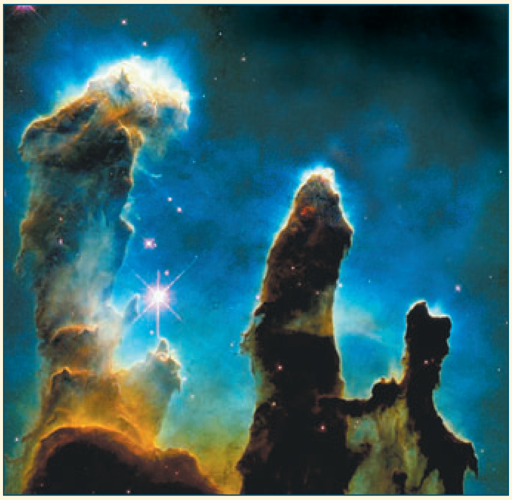


figure oe.3 Stars form in dense clouds of gas like this one, called the Eagle nebula.

If the mass of the protostar is high enough, the temperature and pressure at its centre are extremely high—the temperature can be many thousands of degrees. Under these conditions, gas does not resemble any state found on Earth. It is known as plasma, which is the fourth state of matter. At these temperatures, hydrogen nuclei begin to fuse (join together) to form helium nuclei. Note that in this and later discussion, the atoms are generally in the form of ions, stripped of their electrons. It is these nuclei that fuse together, not neutral atoms. The word atom is used interchangeably with the word nucleus.

* Section 1.2 **Elements and isotopes** describes isotopes of atoms, and how to write the symbols for them (page 211).

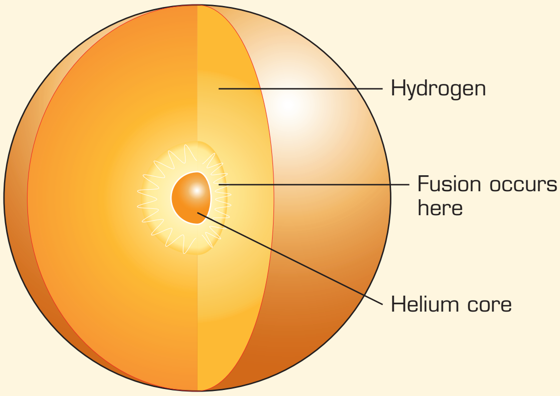


figure oe.4 Fusion of hydrogen to helium occurs in the centre of a star.

The processes and equations that follow are unlike chemical reactions and equations. In a nuclear reaction, you end up with different elements which were not present at the start of the reaction. However, the total mass number and the total charge number are conserved (stay the same). In all equations the mass numbers and charge (atomic) numbers will be shown like this: ${}^{4}_{2}\mathrm{He}$. Sometimes, unusual particles like positrons, ${}^{0}_{1}\mathrm{e}$ (positively charged electrons), will be formed to balance the charge number.

The most basic reaction occurring in all stars is the proton–proton chain (P–P chain). The first step is two hydrogen-1 atoms fusing together to form a hydrogen-2 (or deuterium) nucleus.

Step 1a: ${}^{1}_{1}\mathrm{H} + {}^{1}_{1}\mathrm{H} \rightarrow {}^{2}_{1}\mathrm{H}$

The total mass number on the left is 2, and this is conserved on the right. The hydrogen atoms on the left have a total charge of +2, while the hydrogen atom on the right has a total charge of +1—another particle is needed to balance the equation. It must have a charge of +1, but no mass number. A positron is needed to complete the equation. The balanced equation is:

Step 1b: ${}^{1}_{1}\mathrm{H} + {}^{1}_{1}\mathrm{H} \rightarrow {}^{2}_{1}\mathrm{H} + {}^{0}_{1}\mathrm{e}$

This process is shown in figure oe.5. On the left, the isotopes of hydrogen had no neutrons, but on the right a neutron has been created. One of the original protons formed a neutron by emitting the positron. This removes a positive charge from the newly formed atom. The positron will combine with an electron and form energy.

The other steps in the proton–proton chain are:

Step 2: ${}^{2}_{1}\mathrm{H} + {}^{1}_{1}\mathrm{H} \rightarrow {}^{3}_{2}\mathrm{He}$

Step 3: ${}^{3}_{2}\mathrm{He} + {}^{3}_{2}\mathrm{He} \rightarrow {}^{4}_{2}\mathrm{He} + 2{}^{1}_{1}\mathrm{H}$

These are illustrated in figures oe.6 and oe.7.

Questions

3 Why do you think fusion occurs at the centre of a star and not in the outer parts?

4 Show that steps 2 and 3 above are balanced for
   a mass number
   b charge number.

5 Step 3 involves two helium-3 atoms. How many times must steps 1 and 2 occur, to enable step 3 to occur once?

6 'The net result of these three steps is that four hydrogen-1 atoms combine to make one helium-4 atom.' Use the equations for the three steps to show that this statement is true.

The origin of starlight

Section 1.3 **Atomic mass** describes how relative atomic mass is calculated as an average of the masses of the isotopes (page 213).

We considered the simplest nuclear fusion reactions in stars. Stars, like the Sun, produce enormous amounts of energy, causing them to shine across the reaches of space. How is such a huge amount of energy produced?

Nuclear reactions are able to release far more energy than chemical reactions because some of the mass is converted into energy. The energy released (in joules) is given by the equation $E = mc^2$ (Einstein's famous equation), where m is mass lost in kg and c is the speed of light, 3×10^8 m s^{-1}. As you can predict from Einstein's equation, a slight loss of mass causes a huge release of energy.

The masses of some isotopes are tabulated in table oe.1. Masses in u can be converted to kg by multiplying by 1.66×10^{-27}. (Masses for other isotopes can be found in a reference such as the *Handbook of Chemistry and Physics*.)

Einstein's equation $E = mc^2$ is used to calculate the energy released when a given mass is lost in a nuclear reaction. Einstein explained that mass is simply one form of energy. Energy can be converted from one form to another, including the form called mass. Burning 4 g of hydrogen in oxygen, releases approximately 242 kJ of energy—enough energy to supply a household with electricity for just a few seconds. Nuclear fusion of 4 g of hydrogen releases

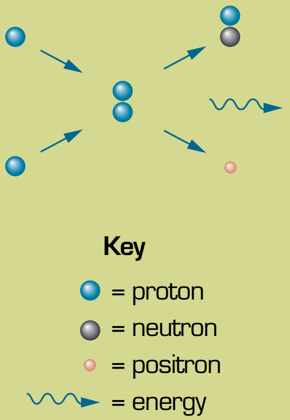


figure oe.5 Two hydrogen-1 (${}^{1}_{1}\mathrm{H}$) nuclei fuse to form a hydrogen-2 (${}^{2}_{1}\mathrm{H}$ or deuterium) nucleus and a positron.

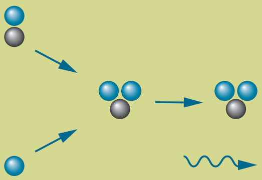


figure oe.6 This fusion reaction produces ${}^{3}_{2}\mathrm{He}$.

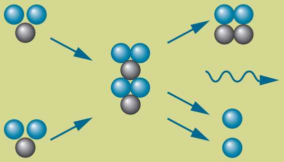


figure oe.7 The fusion of two helium-3 (${}^{3}_{2}\mathrm{He}$) nuclei produces helium-4 (${}^{4}_{2}\mathrm{He}$) and re-generates two hydrogen-1 atoms.

table oe.1 The masses of some common stellar isotopes.

Particle	Symbol	Mass (u)
proton	${}^{1}_{1}\mathrm{p}$	1.00728
neutron	${}^{1}_{0}\mathrm{n}$	1.00728
positron	${}^{0}_{1}\mathrm{e}$	0.00055
hydrogen-1	${}^{1}_{1}\mathrm{H}$	1.00782
hydrogen-2	${}^{2}_{1}\mathrm{H}$	2.0140
helium-3	${}^{3}_{2}\mathrm{He}$	3.0160
helium-4	${}^{4}_{2}\mathrm{He}$	4.0026
beryllium-8	${}^{8}_{4}\mathrm{Be}$	8.0053
carbon-12	${}^{12}_{6}\mathrm{C}$	12.00000

figure oe.8 Stars glow for billions of years, powered by fusion reactions.

more than ten million times as much energy as combustion of the same amount of hydrogen.

What happens to the energy released when these reactions occur in stars? Initially it is in the form of gamma rays, which are more energetic and powerful than X-rays. If the Sun emitted mainly gamma rays, life on Earth would not survive. As the gamma rays radiate out from the hottest regions in the centre of the star, they are absorbed and re-emitted many times. By the time this energy reaches the surface, it is emitted by stars such as the Sun as mainly visible light, with some UV rays and infrared radiation. The exact colour of the light emitted by a star depends on its surface temperature.

Binding energy

The lighter elements

For each ^4_2He nucleus created by fusion, the calculated mass loss represents large amounts of energy released. To do this in reverse and split the helium nucleus into separate protons and neutrons would require energy. This energy requirement has the effect of binding larger nuclei together, so that the electrical repulsion between the protons does not cause the nucleus to fly apart.

The binding energy per nucleon (proton or neutron) is defined as the energy required to separate a nucleus into individual protons and neutrons divided by the total number of nucleons. Figure oe.9 shows a graph of binding energies.

The graph shows that iron Fe has the greatest binding energy per nucleon. Formation of elements up to Fe by fusion releases energy. Formation of elements heavier than Fe requires energy.

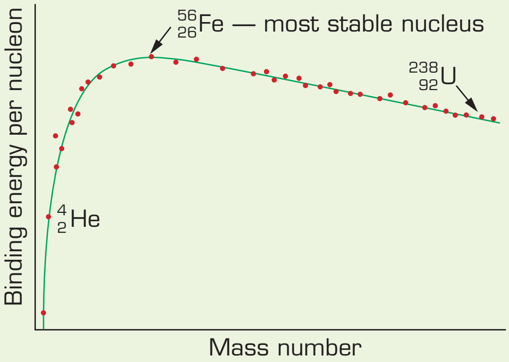


figure oe.9 Binding energy per nucleon versus atomic number for stable isotopes.

Elements heavier than iron

* Sections 1.4 **Electronic structure** (page 215) and 1.5 **Electronic structure in more detail** (page 218) describe atoms, electrons and the periodic table.

Knowing that energy is required to produce heavy elements by fusion, you might wonder how heavier elements form. The high temperatures in a large star provide the necessary energy. Elements heavier than iron also form when free neutrons are absorbed. Elements up to bismuth can form this way.

Large amounts of energy are unleashed in a supernova explosion, and the atoms already formed inside a large star are slammed into each other and bombarded with a stream of neutrons. These atoms fuse, creating most of the elements in the periodic table. Elements heavier than iron are much less abundant in the universe than the lighter elements. This is predicted by our knowledge of the life cycles of stars.

If you think about it, the larger atoms in your body are dust from ancient stars.

Older stars

Eventually the hydrogen in the core of a star is all converted to helium. The core then becomes unstable and contracts and the outer shell of hydrogen expands and cools. It begins to glow red rather than yellow or blue. The star becomes a red giant.

As the core contracts, fusion reactions convert helium to carbon, producing a further burst of energy. The fusion reactions are:

$^{4}_{2}\text{He} + ^{4}_{2}\text{He} \rightarrow ^{8}_{4}\text{Be} + \text{energy}$

$^{4}_{2}\text{He} + ^{8}_{4}\text{Be} \rightarrow ^{12}_{6}\text{C} + \text{energy}$

As the $^{4}_{2}\text{He}$ nucleus is also called an alpha-particle, the above reactions are called the triple alpha reaction. In effect

$3\,^{4}_{2}\text{He} \rightarrow ^{12}_{6}\text{C} + \text{energy}$

As time passes, some of the carbon formed mixes with the outer layers. The next step depends on the mass of the star.

For smaller stars up to about 1.4 times the mass of the Sun, fusion stops. The core collapses again. Shock waves cause the outer layers to be expelled. This expanding shell of gas glows brightly and is called a planetary nebula. It is mostly hydrogen, with some helium and carbon. The hot core, called a white dwarf, glows for some time. Eventually it cools and becomes a black dwarf. Figure oe.10 shows the life cycle of this type of star.

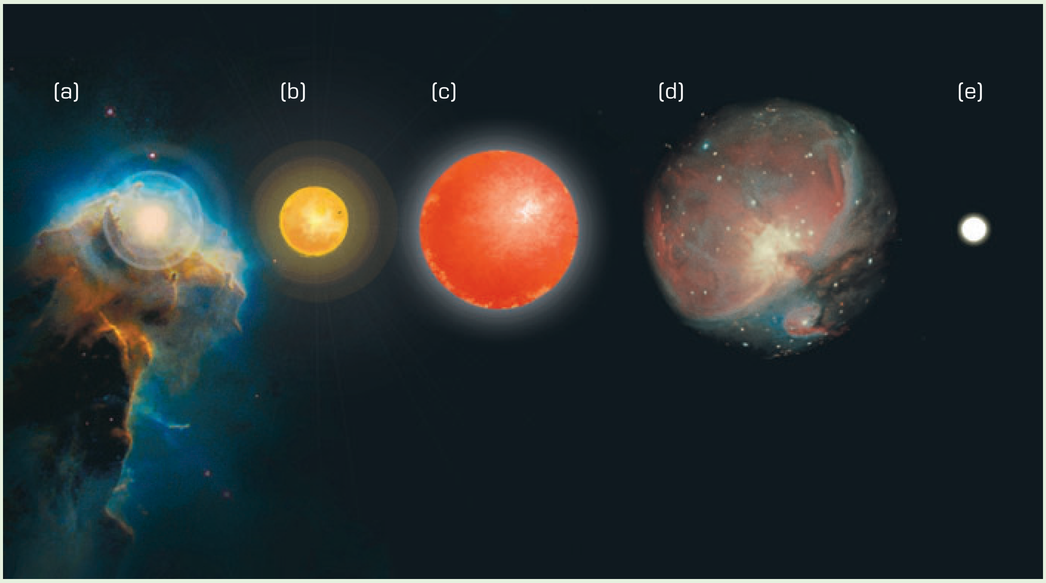


figure oe.10 Stages in the life of a lower mass star: (a) star formation in a gas cloud, (b) yellow star, (c) red giant stage, (d) planetary nebula, (e) white dwarf.

Larger stars explode in a final burst of energy called a supernova. This sudden temporary burst of light was once thought to be a new star. The remnant core collapses on itself with such force that the electrons are forced into the nuclei of atoms. They combine with protons to form neutrons and the core becomes a tiny, dense ball of neutrons called a neutron star. Some of these neutron stars spin rapidly and emit a powerful beam of radiation that can be detected as rapid pulses of energy. This is called a pulsar. Figure oe.11 shows these events.

For stars that are more than five times the mass of the Sun, the nuclear fusion process does not stop when carbon forms. The huge pressures and temperatures mean elements up to iron can form.

Shells, like the layers of an onion, form around the core. In each shell a different element is present, fusing to form a heavier element at the junction between shells. Unused hydrogen forms the outermost layer.

During a supernova explosion, these elements (many common in planets and living organisms) are thrown out into space. At the moment of explosion, the core contracts to a singularity. This is a small, dense object with such strong gravitational pull that not even light can escape from its surface—a black hole.

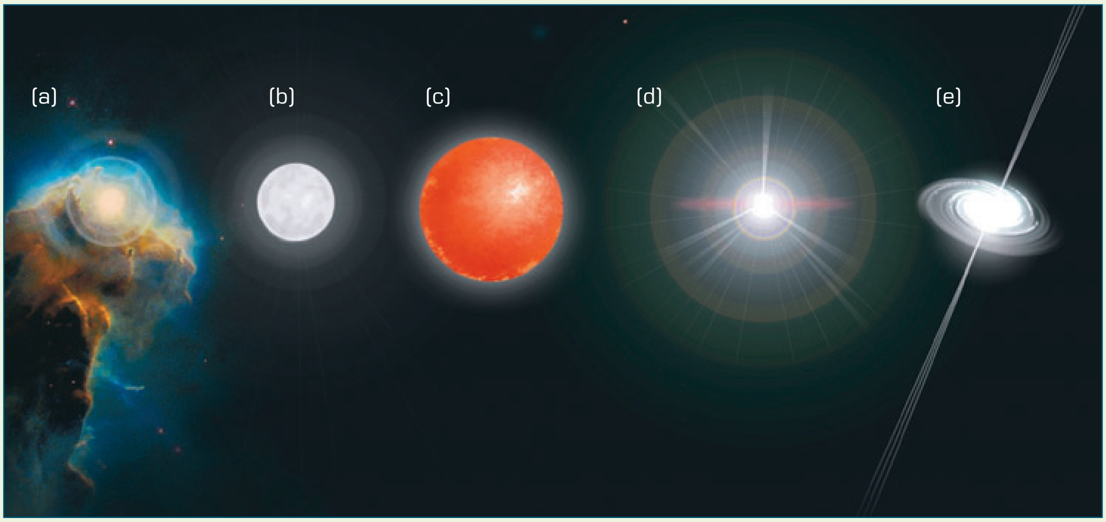


figure oe.11 The life cycle of a star several times heavier than the Sun: (a) contracting gas cloud, (b) hot white star, (c) red giant, (d) supernova, (e) pulsar (pulsating neutron star).

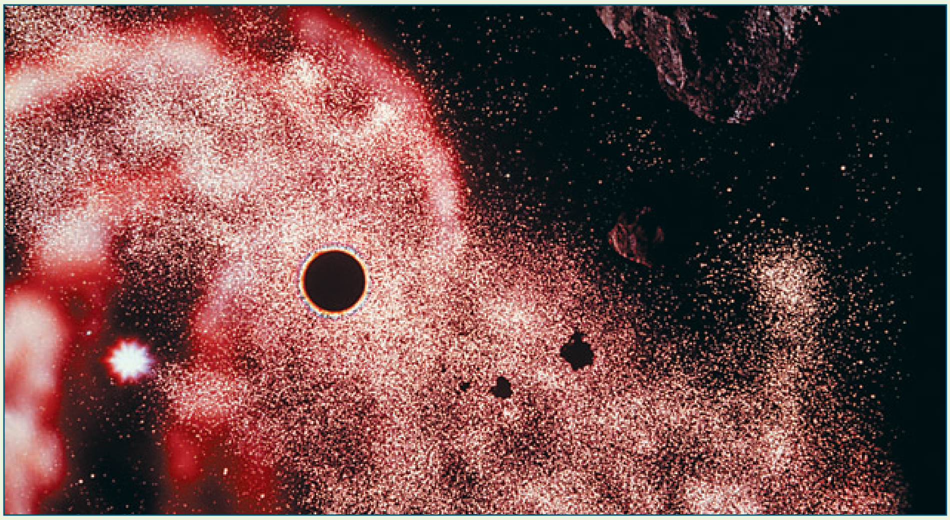


figure oe.12 A supernova like this one, can disperse large amounts of heavier elements.

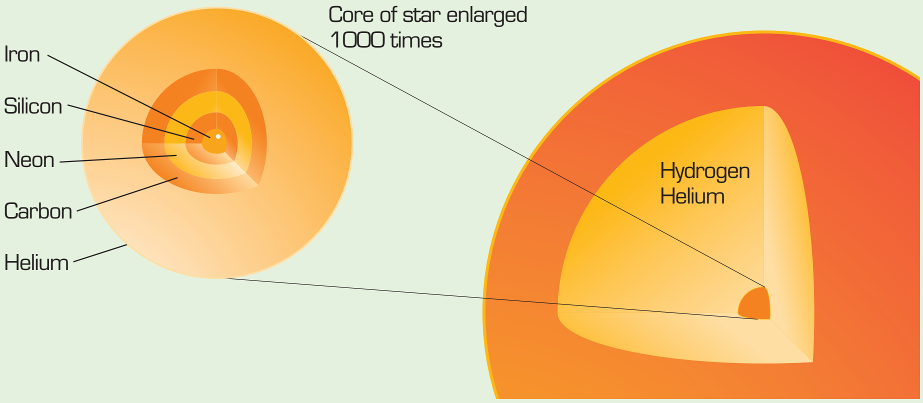


figure oe.13 The elements contained in the core of a large star, towards the end of its life cycle.

Fusion on Earth

Nuclear reactors used to generate electricity make use of fission reactions. These reactors utilise the energy stored in the nuclei of elements heavier than iron. Fission occurs as these radioactive nuclei are split apart. The result is the release of enormous amounts of energy as well as dangerous radiation and the production of radioactive wastes.

Climate change associated with the burning of fossil fuels as well as dwindling reserves of these resources has resulted in research into more sustainable ways of meeting our energy needs.

Scientists have proposed a replication of the fusion process that takes place in stars as a method of producing energy. Nuclear fusion reactors have been designed that make use of a reaction between deuterium (hydrogen-2) and tritium (hydrogen-3).

$${}^{2}_{1}\text{H} + {}^{3}_{1}\text{H} \rightarrow {}^{4}_{2}\text{He} + {}^{1}_{0}\text{n}$$

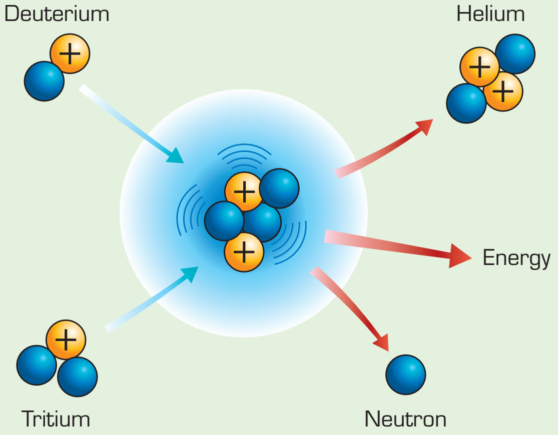


figure oe.14 The fusion of deuterium and tritium nuclei would be used in fusion reactors. Helium and a neutron are formed and a large amount of energy is released.

The heat released by the energetic neutrons could be used to heat water and drive the turbines of electric generators. The production of dangerous radiation would be avoided and the helium gas produced is unreactive and a safer product than the radioactive wastes of fission reactors.

Using fusion reactions to produce electricity has its own difficulties. These reactions require extremely high temperatures (millions of degrees) and are difficult to contain and control. However, research continues into nuclear fusion reactors in the hope that they will provide sustainable sources of energy for the future.

Consider this

How much energy is released by the fusion reactions in stars? Use $E = mc^2$ to calculate the energy released by the series of reactions that produce one atom of He. First, use the values in table oe.1 to calculate the mass loss and convert it to kg (multiply atomic mass units by 1.66×10^{-27}). Then calculate the energy released when atoms fuse in each of steps 1, 2 and 3.

Step 1: $^{1}_{1}\mathrm{H} + ^{1}_{1}\mathrm{H} \rightarrow ^{2}_{1}\mathrm{H} + ^{0}_{1}\mathrm{e}$

Step 2: $^{2}_{1}\mathrm{H} + ^{1}_{1}\mathrm{H} \rightarrow ^{3}_{2}\mathrm{He}$

Step 3: $^{3}_{2}\mathrm{He} + ^{3}_{2}\mathrm{He} \rightarrow ^{4}_{2}\mathrm{He} + 2^{1}_{1}\mathrm{H}$

How can mass be lost in these processes?

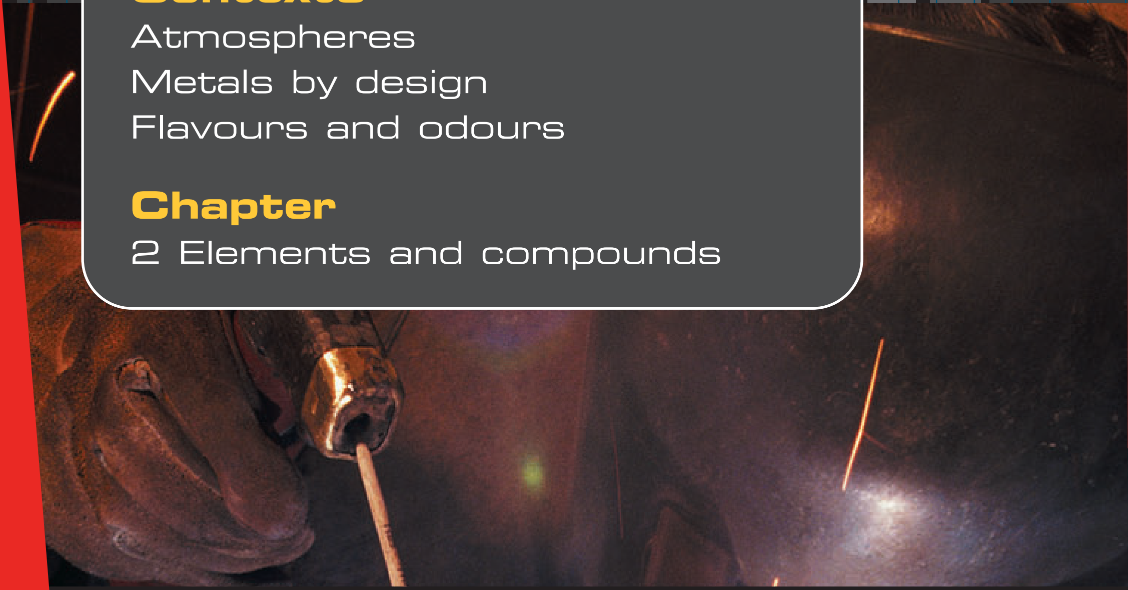
In the Sun, 700 million tons of hydrogen fuse into helium each second. How much energy would this release?

The average Australian uses 30 000 MJ (million joules) of energy each year for transport. Assume the population of Australia is approximately 20 million. If the energy produced in the Sun by fusion in one second was made available for our transport needs, how long would it last?

Further investigations

For each of the investigations below, present your results as an educational poster or as an oral presentation.

1. Fusion of hydrogen has been proposed as a clean, almost unlimited energy source for the future. Considerable research has centred on how to achieve this in a controlled, sustainable way on Earth. Compare this energy source with current energy sources. What are the advantages? What obstacles confront scientists? Are they likely to overcome these difficulties?
2. The formation of elements heavier than iron was discussed briefly. The names of the processes involved are the s-process and the r-process. Research these processes.
3. What is the future of the universe? Will stars continue to form forever or will the hydrogen run out? The universe is still expanding from the Big Bang—will this expansion continue, preventing future star formation? What form of matter will be mainly present in the future? Research one aspect of the future of the universe.
4. Research the production and use of radioactive isotopes for medical use.
5. Research the production of transuranic elements in particle accelerators.



Contexts
Atmospheres
Metals by design
Flavours and odours

Chapter
2 Elements and compounds

area of Study 2

Structure, bonding and properties of materials

atmospheres

Question: What weighs only 1.25 grams per litre (measured at sea level), is invisible, is made mostly of a substance that is extremely unreactive, and is usually unnoticed, ignored and unappreciated?
Answer: The Earth's atmosphere.

The upper atmosphere does not have a distinct boundary. Rather, it becomes thinner until it has the same density as the sparse gas in space.

figure at.1 Our atmosphere is one thing that sets the Earth apart from other planets.

The Earth's atmosphere

The solid and liquid part of the Earth has a radius of approximately 6000 km. The atmosphere extends for about 120 km above this. This represents just 2% of the Earth's total radius. An orange peel as a percentage of the size of an orange is thick by comparison.

Yet, the atmosphere is essential for life on Earth, and is used by humans as a resource for many activities and processes.

So what is the atmosphere made of? It consists mostly of a mixture of gases, with some liquid droplets and suspended solid particles. Table at.1 shows the composition of air. The gases in the atmosphere are mainly covalent molecular elements and compounds.

Section 1.2 **Elements and isotopes** describes what is meant by an element (page 211).

The concentrations of gases and the density of the atmosphere drops dramatically with altitude. At 32 km, the concentration of the air is just 1% of that at sea level. The concentrations of several major gases of the atmosphere and their temperature variations are shown in figure at.3. The temperature varies at different altitudes as different gases absorb various amounts of sunlight.

table at.1 The composition of gases in dry air measured near the Earth's surface.

Gas	Percentage of particles in the air
nitrogen	78.0
oxygen	21.0
carbon dioxide	0.0033
other gases (e.g. noble gases, methane)	< 1

figure at.2 Using the atmosphere—hot air ballooning.

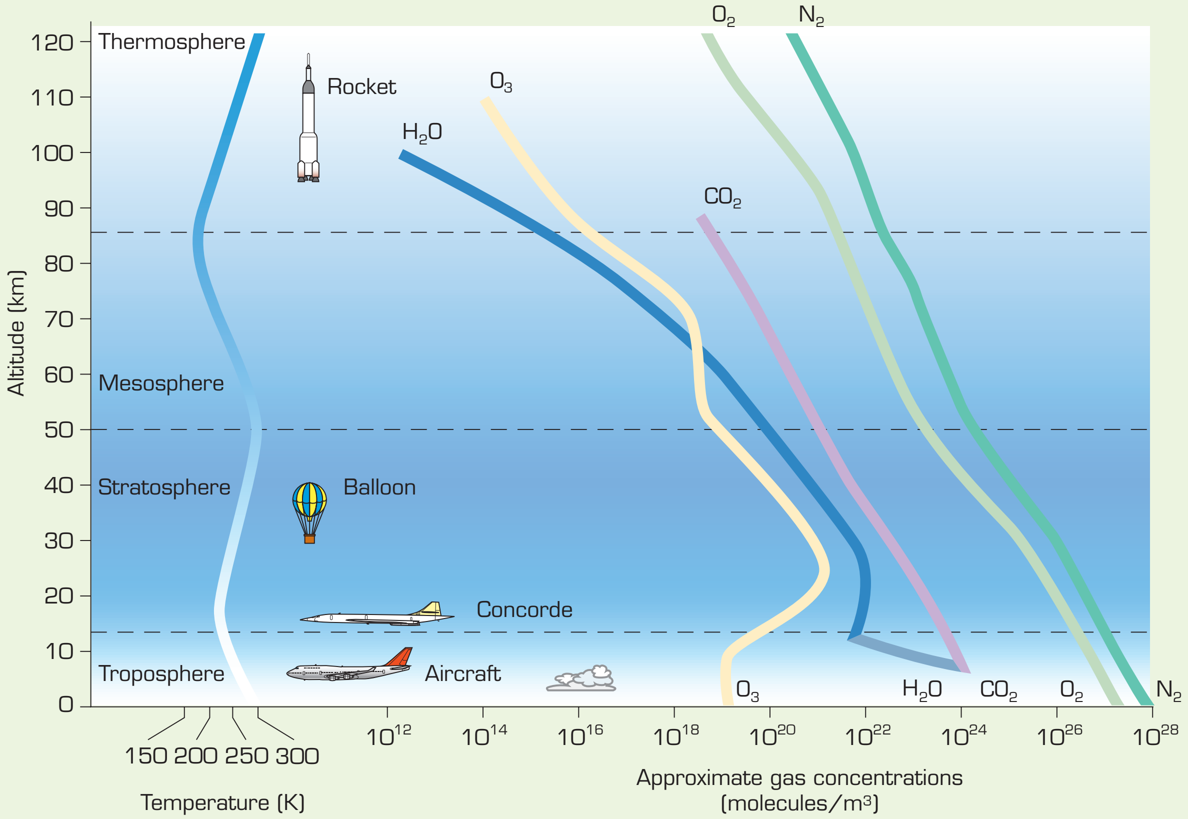


figure at.3 Variations in temperature and concentrations in the layers of the atmosphere.

Questions

1 Use figure at.3 to determine:
   - **a** the coolest altitudes in the atmosphere
   - **b** the layers or boundaries to which this corresponds.

2 **a** Use figure at.3 to determine how many times smaller the oxygen concentration is at an altitude of 40 km than at the Earth's surface.
   - **b** The concentration of most gases shown drops dramatically with altitude. Ozone, O_3, is an exception to this. At what altitude range would the ozone layer be situated?
   - **c** What is the maximum concentration of ozone in the atmosphere? Compare this to the concentration of oxygen, O_2, at the same altitude. Why do you think this region of the atmosphere is referred to as the ozone layer?

The impact of humans on the atmosphere

Oxides of sulfur such as SO_2 and SO_3 are formed during the combustion of coal and oil, which contain sulfur. They react with water to form acids. The equation for the reaction of SO_2 with water is

$$\begin{array}{llll} SO_2 & + \; H_2O & \rightarrow & H_2SO_3 \\ \text{sulfur dioxide} & \text{water} & & \text{sulfurous acid} \end{array}$$

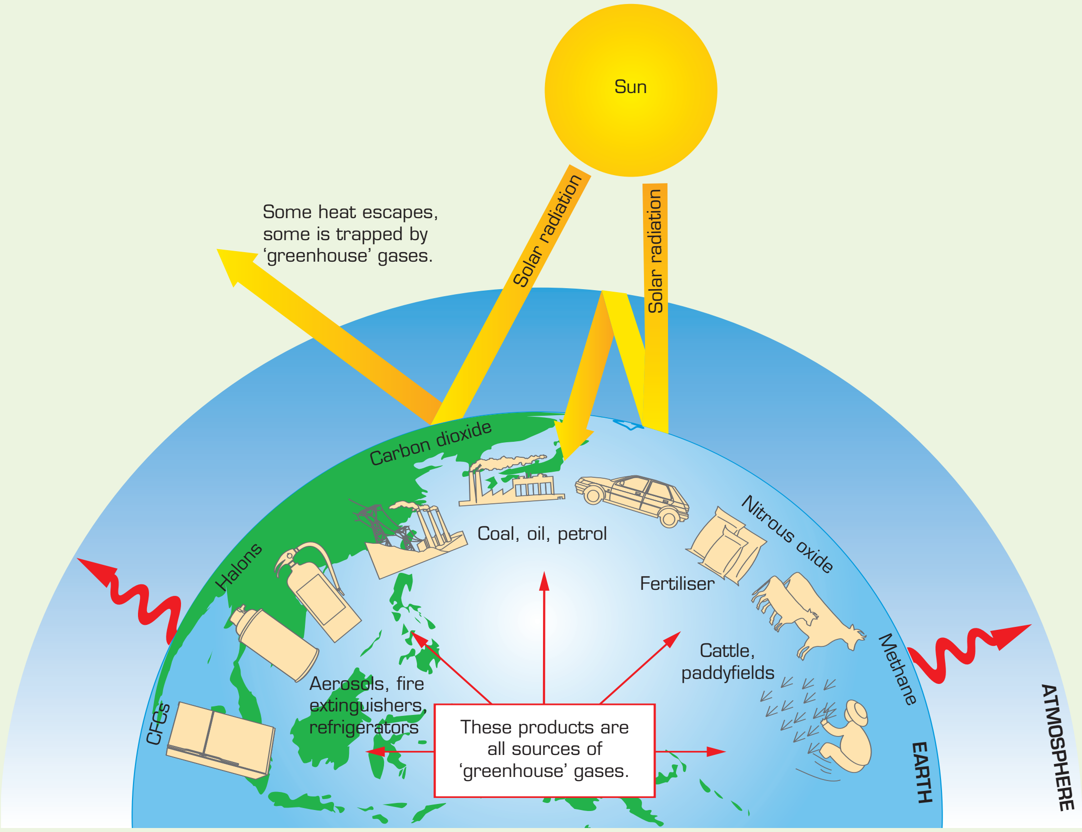


figure at.4 Light easily enters the atmosphere, but heat radiation is trapped.

Such reactions can cause acid rain when pollution in the form of sulfur oxides reacts with moisture in the atmosphere. Carbon dioxide, one of the greenhouse gases, also dissolves in rainwater, making it naturally slightly acidic. Acid rain causes many problems to ecosystems and buildings if levels of these pollutants are high.

Section 2.3 **Covalent bonding** describes covalent molecular substances such as carbon dioxide and sulfur dioxide (page 233).

Acid rain is a big problem in some areas of the world and is the result of human activity. Two other major atmospheric problems are the effects of global warming and the depletion of the ozone layer.

Global warming is caused by the enhanced greenhouse effect. The greenhouse effect occurs when sunlight reaches the Earth's surface, is absorbed and then converted to heat energy. The atmosphere cannot transmit heat as easily as it absorbed the original light, so the heat is trapped close to the Earth's surface. Heat is trapped by glass in a similar way—notice how a car gets hot when parked in sunlight. Without a heat-absorbing atmosphere, this process would not occur. The day-lit side of the Earth would be unbearably hot, while the dark side would be bitterly cold, like the Moon.

Carbon dioxide is well known for its ability to trap heat. Since the Industrial Revolution, we have been burning increasing amounts of fossil fuels and the level of carbon dioxide in the atmosphere has risen dramatically (see figure at.5).

But carbon dioxide is not the only gas responsible for the greenhouse effect. Small concentrations of certain other gases are far more efficient at trapping heat than carbon dioxide.

What are the dangers of global warming? Surely an increase of a few degrees is tolerable? The answer is both yes and no. Temperature differences in the atmosphere cause pressure fluctuations. The result is winds that circulate and drive our weather system. Increased temperatures will also cause more evaporation. This can produce higher rainfall in some areas but may also dry out the land faster after rain. Weather systems will also be more intense. Queensland can expect more cyclones of greater intensity and they will also reach areas further south. Rainfall will increase in some areas and there will be droughts in others. Some species may cope better than others with changes in carbon dioxide levels and climate.

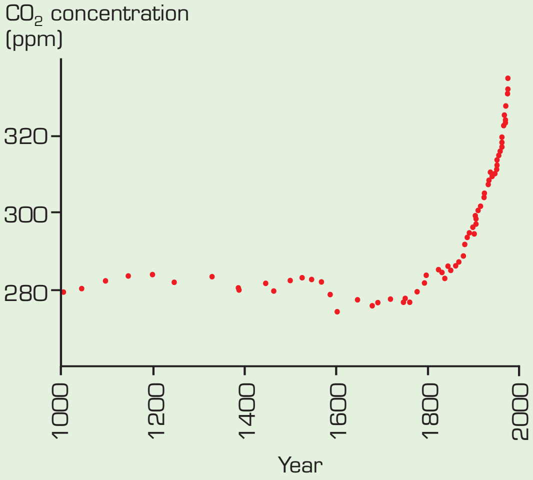


figure at.5 The level of carbon dioxide has risen as a direct result of our actions.

Questions

3 a Find out which other gases contribute to the greenhouse effect.
  b Which gases are the most efficient at trapping heat?
  c What are their sources?
  d What can be done to minimise the amounts we release?

Atmospheres of other planets

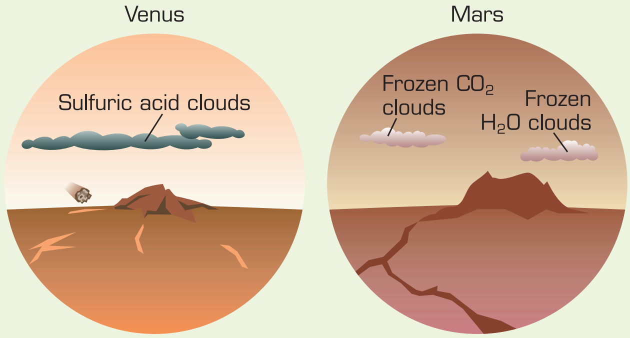


figure at.6 Most of the other planets in our solar system have atmospheres, but they are very different from Earth's.

Venus

Glowing brightly as the morning or evening star, Venus has always fascinated humans. Reflective clouds of the upper atmosphere ensure its bright golden light is noticed. As early as 1970, space probes have investigated the planet once known as our twin. But a more hostile environment could not be imagined.

Space probes descended through clouds of sulfuric acid (H_2SO_4) in the upper atmosphere with winds of up to 400 km per hour. At the surface, where active volcanoes belch out sulfur dioxide, temperatures can reach 465∞C and the atmospheric pressure is ninety-four times that on Earth. Probes have sent back valuable information and photographs, but do not survive long in the Venusian atmosphere.

The various sulfurous compounds present in the atmosphere and rocks of Venus react in a complex cycle. Sulfuric acid (H_2SO_4) dominates in the upper atmosphere. This is a strong, extremely corrosive substance, better known to us as the acid in car batteries. As we descend, sulfur trioxide (SO_3) and sulfur dioxide (SO_2) predominate. Both of these can dissolve in and react with water to form acids.

In the lower atmosphere, carbonoxysulfide (COS) and hydrogen sulfide (or rotten egg gas, H_2S) dominate. In the crust, sulfur is present as calcium sulfate ($CaSO_4$) and iron pyrites (FeS_2). Large quantities of water vapour and carbon dioxide are also present in the atmosphere, trapping heat via the greenhouse effect.

A series of reactions occurs as shown in figure at.7. In the fast cycle, sulfur dioxide reacts with carbon dioxide and water vapour, in the presence of sunlight, to form sulfuric acid. In the slow cycle, sulfuric acid is formed from hydrogen sulfide and carbonoxysulfide. Sulfuric acid then forms sulfur trioxide in the lower atmosphere. Excess sulfur dioxide in the atmosphere reacts with rocks, forming calcium sulfate, which then reacts with iron oxide and carbon dioxide to form iron pyrites. Iron pyrites from the surface of Venus reacts with carbon dioxide and water vapour, forming hydrogen sulfide and carbonoxysulfide, and the cycle continues.

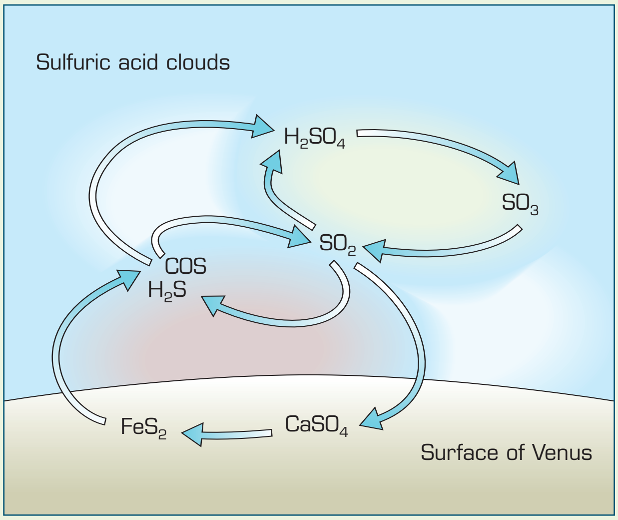


figure at.7 The sulfur cycle on Venus.

* Section 2.2 **Ionic bonding** describes the structure of ionic compounds like calcium sulfate and iron pyrites (page 229).

Mars

On Mars, the atmosphere is about one hundred times thinner than on Earth and mainly consists of carbon dioxide. Despite these levels of carbon dioxide, the atmosphere does not retain much heat, as Mars is further from the Sun. Temperatures fluctuate widely from 22∞C in the daytime at the equator to −133∞C at night at the poles.

As the temperatures are cooler, substances that we think of as gases may be liquids or solids on Mars. Large amounts of water and carbon dioxide are trapped in the rocks either in the frozen state or locked in compounds. Active volcanoes do liberate some of these but the overall amount of gas in the atmosphere remains small. Scientists speculate that early in its history Mars was warmer and may even have been suitable for life.

Properties of molecular substances such as melting point depend on the strength of intermolecular forces.

* Section 2.3 **Covalent bonding** describes covalent molecular substances and intermolecular forces (page 233).

Artificial atmospheres

On Earth

Oxygen is a very reactive gas. In certain situations this is undesirable. One example is inside a light bulb. If oxygen was present, it would rapidly react with the tungsten filament at the high temperatures reached when the bulb is on and the filament would be destroyed. To avoid this, the air inside the bulb is removed and replaced with argon, an unreactive noble gas. This creates a new atmosphere inside the bulb, which prolongs the life of the globe.

Scuba diving can involve the use of artificial atmospheres. Conventional scuba diving uses a tank of pressurised air. However this gas mixture can cause a painful, sometimes fatal, condition called the bends when a diver surfaces too quickly. Below the surface of the water, the pressure increases by 1 atm for every 10 m descended. At a depth of 30 m, the pressure is four times the pressure at the surface.

1 atm is the pressure of air at sea level.

Gases are more soluble at high pressure so during a dive high levels of gases build up in the blood stream. As the diver ascends, the gases become less soluble. If ascent happens too quickly, the gases (especially nitrogen) begin to bubble out in the blood stream, much like when the lid is removed from a bottle of soft drink.

To treat the bends the diver is placed in a hyperbaric chamber. Inside this chamber the pressure is artificially increased, causing the gases to redissolve in the bloodstream. The pressure is then gradually reduced. The artificial atmosphere created in a hyperbaric chamber is a life saver.

To avoid the bends, scuba divers can use a safer system called a rebreather. In this system, pure oxygen is used. This prevents nitrogen dissolving in the blood. The exhaled air contains a large proportion of unused oxygen. It is treated with sodium hydroxide to remove carbon dioxide and the unused oxygen is recycled.

Terraforming

Terraforming literally means 'making earth'. Some scientists have radical plans to transform the atmosphere on Mars into one that could sustain life. Releasing CFCs and other greenhouse gases to trap heat and detonating nuclear bombs on the surface could raise the temperature and revive volcanoes. These volcanoes would release more carbon dioxide into the atmosphere, further contributing to warming. Large amounts of water are believed to be trapped frozen near the surface and a warmer climate would liquefy and vaporise this water. Plants could then survive and grow, oxygen levels would increase due to photosynthesis, and eventually other organisms and even humans could survive in the new ecosystem.

figure at.8 Changing the atmosphere on Mars could allow it to be transformed from a planet devoid of life to one on which organisms from Earth could thrive.

Consider this

1 For thousands of years, humans have established themselves in new places on Earth, often altering and damaging ecosystems forever. Investigate the possibility of terraforming Mars. Do humans have the right to invade and alter other planets in this way? Present your findings and opinion as an oral presentation or essay.

2 Nobel Prize winner Paul Crutzen has proposed that small airborne particles from forest fires block and reflect a lot of light, thus minimising global warming. Is this beneficial in the short and long term? Consider the implications of this, and present your opinion as an oral presentation or a newspaper article.

3 'It is important to seek other places in the galaxy that may be suitable for life, so that humans have other options if the Earth is unsuitable for life in the future.' Consider the implications of this statement and present your opinion as a report to world governments.

4 You have won a planet! You need to design an atmosphere for your planet. Consider the life forms you want on your planet and how the gas levels will be maintained. Present your proposal as a report.

Further investigations

1 Carbon dioxide levels were much higher in past eras. Life on Earth survived through those periods, so we should not be concerned about rising levels of greenhouse gases and global warming now. Research evidence for the past history of the Earth's atmosphere and climate. Defend your opinion, with evidence, in an oral presentation.

2 Extreme temperature variations have always occurred on Earth. There have been ice ages and warmer times. Temperature variation is part of a natural cycle, so it is impossible to tell if current warming is due to natural causes or our actions. Fixing the problem will be expensive. Research temperature variation on Earth and the potential cost of proposed solutions to the problem. Is it worth the expense of trying to rectify this problem? Form an opinion and present your findings as an oral presentation.

3 James Lovelock outlined the Gaia hypothesis. In this, he proposes that the whole earth responds to change like a single organism. If excessive warming occurs, some mechanism will occur that restores the balance. For example, increased cloud cover will reflect light, and decrease warming, as well as increase rain. Can this mechanism save the Earth? Can it save humans? Research the Gaia hypothesis in order to answer the questions posed. Present your findings as an illustrated article suitable for a science magazine.

4 Find out about the atmospheres of other planets and moons in the solar system. Which would be chemically most and least suitable for life as we know it? Present your findings as a PowerPoint presentation or a web page.

metals by design

Metals feature heavily in the homes, workplaces, industries and leisure activities of people in the developed world. Copper makes great electrical wiring, aluminium foil keeps food fresh and mercury enables us to tell the temperature. Rarely are these metals found in their pure form. Most metals have to be modified to make them suitable for their various applications. This context will outline methods used to achieve the characteristics required for the metals we use.

* Section 2.1 **Physical and chemical properties** describes the general properties of elements including metals (page 226).

Alloys—mixtures of metals

Alloys are not pure substances. They are formed by melting a main metal, adding the required amounts of other components (usually other metals) and allowing the molten mixture to cool. Only small amounts need to be added to produce an alloy with vastly different properties to the original metal. Some common alloys are listed in table md.1.

The addition of other components produces alloys with specific properties. By changing the lattice structure of the original metal, the regular pattern of positive metal ions is disrupted. The alloying of metals can result in a substance with increased strength and greater resistance to corrosion. It can also produce a substance that is brittle or has a lower melting temperature than the pure metal. Generally alloys are harder and less malleable than their component metals as well as being poorer electrical conductors.

Many alloys are the result of trial and error, but an understanding of how alloying works enables the design of materials with specific properties for particular purposes.

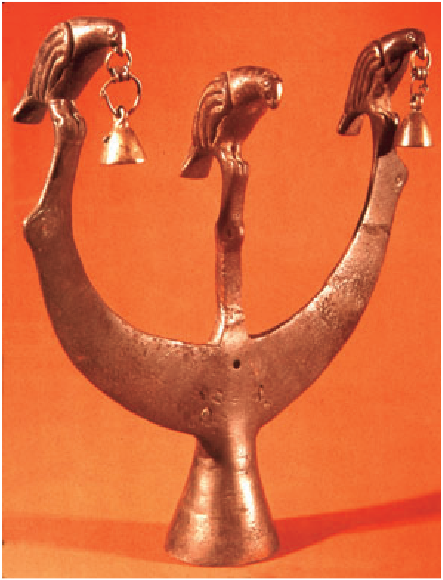


figure md.1 The discovery of bronze moved humans into an age where metal objects became far more prominent.

table md.1 Some features of common alloys.

Alloy	Composition	Properties	Uses
mild steel	iron (99.5%), carbon (0.5%)	strong, easily worked and welded	girders, car bodies
tool steel	iron (99%), carbon (1%)	strong, hard, brittle	tools, knives, cast iron
tungsten steel	iron (81%), tungsten (15%) chromium (4%)	stays hard when hot	drills, cutting tools
stainless steel	iron (74%), chromium (18%) nickel (8%)	hard, corrosion-resistant	cutlery, surgical equipment
brass	copper (65%), zinc (35%)	corrosion-resistant, strong	taps, door knobs
bronze	copper (90%), tin (10%)	corrosion-resistant, hard	ships' propellers, statues
cupro-nickel	copper (75%), nickel (25%)	corrosion-resistant	coins, turbine blades
solder	lead, tin	melts easily	joining metals
dental amalgam	mercury, tin, silver, copper	hardens slowly after mixing	dental fillings
gold alloys	gold, silver, zinc, copper	harder and less expensive than pure gold	jewellery
aluminium alloys	aluminium, magnesium, copper, zinc	light, but harder than pure aluminium	aircraft, motor vehicle fittings, boats

* Section 2.4 **Metallic bonding** describes the structure of metals (page 244).

Substitutional alloys

Some alloys are made from elements with similar chemical properties and atoms of similar size. The atoms of the metal added substitute for the main metal atoms in some positions of the lattice. Ions of both metals are attracted to the delocalised electrons so the mixture remains stable.

Bronze is an example of this kind of alloying. The Bronze Age is seen as a major step forward for the human race as it resulted in tools, implements and weapons that were harder and more resistant to corrosion than the copper ones used previously. All it took was for someone to add some tin to the copper furnace.

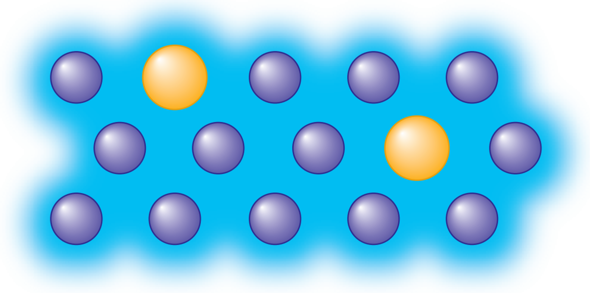


figure md.2 A substitutional alloy. Atoms of the added element have taken the place of atoms of the main element in the lattice.

* Section 2.6 **Patterns in properties** describes the trends in properties such as atomic size and metallic character (page 256).

Interstitial alloys

Interstitial alloys involve the addition of atoms that fit into the spaces between the atoms of the metal lattice. In carbon steel, for example, the smaller carbon atoms occupy the spaces in the iron lattice. Only a small number of carbon atoms are randomly fitted into the spaces between the iron ions.

figure md.3 Australian silver coins are actually made from an alloy of 75% copper and 25% nickel—a substitutional alloy.

Varying the composition of alloys varies their properties. Carbon steel containing 4% carbon is very brittle. It is sometimes known as cast iron and is used in cast iron cookware. The brittleness of cast iron makes it unsuitable for structural purposes. 0.1% carbon steel is ductile and able to be drawn into a wire for making paper clips and staples. 1% carbon steel is stronger without being brittle and is used for structural purposes such as bridge building.

figure md.5 Cast iron cookware is a good conductor of heat but its brittleness makes it unsuitable for building and engineering purposes.

Rust-free stainless steel

The British chemist Harry Brearley developed stainless steel in 1913 while investigating steel with a high chromium content for use in rifle barrels. One step in his analysis of steel involved dissolving it in acid, but Brearley found the steel would not dissolve. He also found that it stayed shiny and didn't develop a dull oxide coating.

Stainless steel develops an extremely thin coating of chromium(III) oxide (Cr_2O_3) that adheres closely to the surface of the metal, protecting it from attack by air and water while allowing the shininess of the metal to show through. If scratched, a new protective coating quickly forms.

Stainless steel is commonly used in cutlery, commercial kitchen surfaces and surgical equipment. It could be used more widely and further reduce the expense and inconvenience of corrosion, but stainless steel is expensive to produce and more cost-effective alternatives exist.

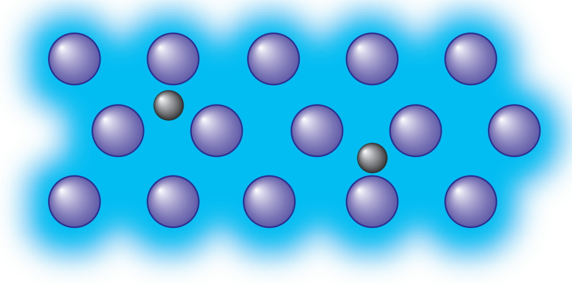


figure md.4 Carbon steel is an interstitial alloy. The smaller carbon atoms occupy some of the spaces between the iron atoms.

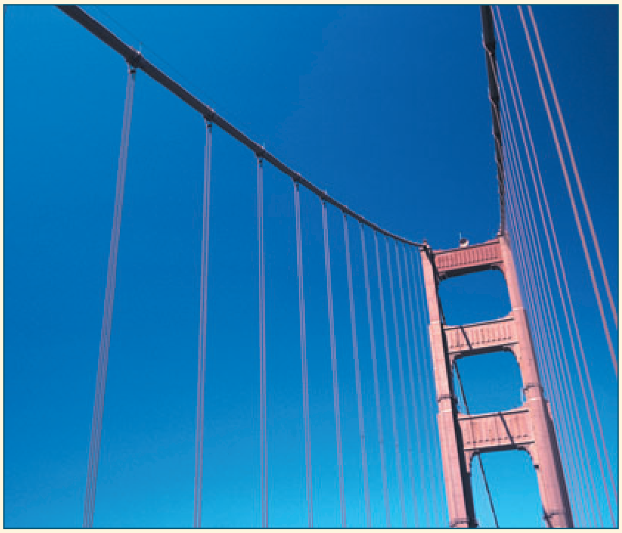


figure md.6 1% carbon steel is extremely strong and flexible enough for bridge building. Steel for this purpose may also contain other elements such as manganese, chromium and silicon to further refine its properties.

figure md.7 Stainless steel is a corrosion resistant alloy.

Experiment md.1

Alloys: preparation of solder

A pellet of solder is prepared by heating a mixture of lead and tin. Pellets of solder, lead and tin are then heated simultaneously. The solder melts before the other metals.

Method

1. To demonstrate the production of an alloy.
2. To show how the physical properties of an alloy differ from those of its constituent metals.

Procedure

Before the demonstration:

1. Prepare three moulds by pushing a dowel rod into a tray of damp sand.
2. Place 10 g of lead in a crucible. Using a fume cupboard, heat the crucible until the lead melts. Pour the molten lead into one of the moulds and allow it to cool.
3. Repeat this procedure using 10 g of tin. Use a marking pen to label the two pellets produced.

For the demonstration:

1. Show the class the pellets of lead and tin that have been prepared and repeat the procedure described above using a mixture of 5 g lead and 5 g tin.
2. When the metal has cooled, drop the three pellets of lead, tin and solder onto a hard surface from a height of 2 m. Note the relative ease with which they are dented.
3. This test must be performed in a fume cupboard. Place the three pellets on a tin lid so that they are equidistant from the centre. Using a tripod stand, pipeclay triangle and Bunsen burner, heat the centre of the lid as shown in figure md.8. Note which pellet melts first. Stop heating once one of the pellets has melted.

Discussion

1. Why is solder used in plumbing rather than just tin or lead?
2. How do the hardness and melting point of an alloy compare with those characteristics of its constituent metals?
3. Name some uses for alloys. Explain how these uses are related to their properties.
4. Explain how the physical properties of alloys reflect their bonding.

materials

- Bunsen burner and tripod
- 3 × crucibles
- pipeclay triangle
- tin lid
- shallow metal tray filled with damp sand
- tongs
- dowel rod
- marking pen
- 10 g lead
- 10 g tin
- mixture of 5 g lead and 5 g tin
- cotton thread

safety

- Wear safety glasses and a laboratory coat for this experiment.
- Lead vapour is toxic. The lead pellets must be prepared in a fume cupboard. Do not continue to heat the tin lid once the solder melts.

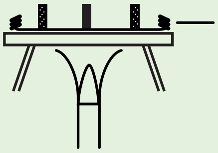


figure md.8

Altering metallic structure

Any solid piece of metal is made up of individual crystals that form as molten material cools down. Each crystal has a regular arrangement of ions within a sea of electrons. The thin individual crystals are randomly arranged. The size of the crystals and their arrangement is dependent on factors such as the rate at which the molten metal cooled. It is the arrangement of the crystals that determines properties such as malleability and ductility. Different methods can be used to change the crystal structure and make the metal suitable for specific uses.

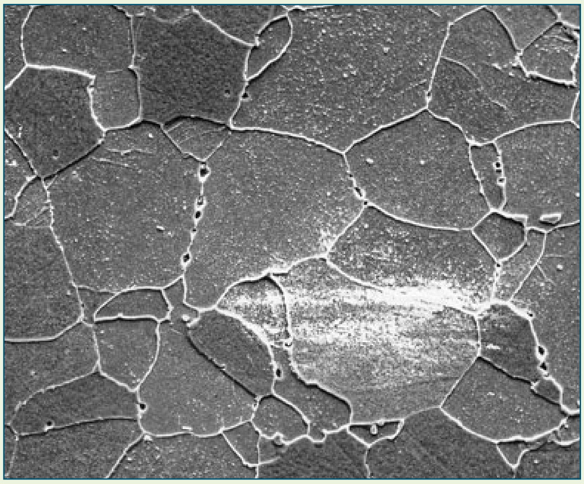


figure md.9 This piece of boiler steel (an alloy of iron, chromium and molybdenum) has individual crystals that formed as molten material cooled down.

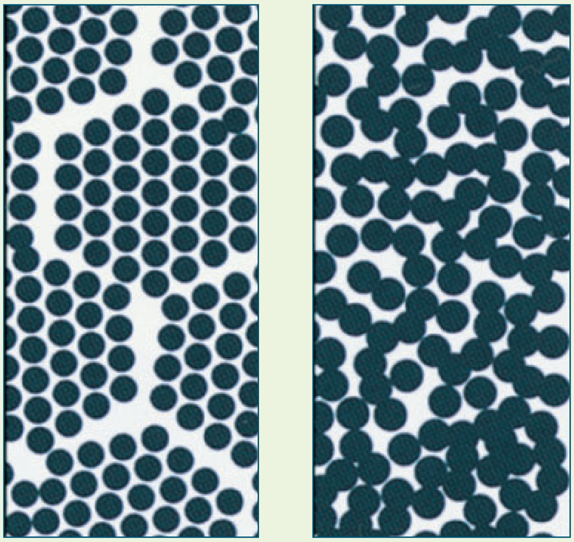


figure md.10 In these representations of a metal each circle represents an ion. On the left a solid crystalline metal is represented and on the right a molten metal with a random arrangement of ions.

figure md.11 Shaping metal by heating and hammering.

Work hardening

If you bend a paper clip once it will normally not break. To break it you need to bend it back and forth a number of times. The bending of the metal in a paper clip rearranges the crystal grains making it harder and brittle. This process is known as work hardening. Hammering (or working) a metal can be done to make it harder as well as change its shape. Sometimes working and heat treatments are used to shape a metal, making it harder and suitable for various uses.

Smithing is the skill used to produce horse shoes and other metal objects from iron. Jewellers also practice a form of smithing to shape gold and silver alloys into jewellery.

Heat treatment

The sizes of some of the crystals in a piece of metal can be altered by heating and cooling the metal at different rates. These processes allowed ancient civilisations to produce tools and weapons with sharp blades and continues to be used in the production of knives and sharpened tools.

table md.2 Three methods of heat treating metals.

Treatment	Process	Effect on metal properties	Effect on metal structure
annealing	A metal is heated to a moderate temperature and allowed to cool slowly.	The metal is softer with improved ductility.	Larger metal crystals form.
quenching	A metal is heated to a moderate temperature and cooled quickly (sometimes by plunging into water).	The metal is harder and brittle.	Tiny metal crystals form.
tempering	A quenched metal is heated (to a lower temperature than is used for quenching) and allowed to cool.	The metal is harder but less brittle.	Crystals of intermediate size form.

Since pure metals are rarely used, it is a combination of alloying and heat treatment that results in many of the objects in common use. The blanks used for making coins are work hardened by rolling. This also gives them a rim. They are then softened by heating to 850°C ready for stamping with their design.

Experiment md.2

Modifying the properties of metals

Method

To modify the properties of iron by heating.

Procedure

1 Take three needles or hair pins, keeping one as a control.

2 Strongly heat the other two needles or hair pins, one at a time, in the flame of the Bunsen burner.

3 Drop one of the hot needles or hair pins directly into the beaker of cold water. Allow the other needle or hair pin to cool very slowly by gradually moving it away from the flame of the burner and allowing it to cool in air.

4 When the needles or hair pins are quite cool, try to bend them and to break them. Record your observations.

Discussion

1 How does the rate of cooling affect the brittleness and hardness of the metal?

materials

- 3 × sewing needles or hair pins
- 250 mL beaker of cold water
- Bunsen burner
- bench mat
- tongs
- pliers (optional)

safety

- Wear safety glasses and a laboratory coat for this experiment.
- Remember not to touch the hot metal.

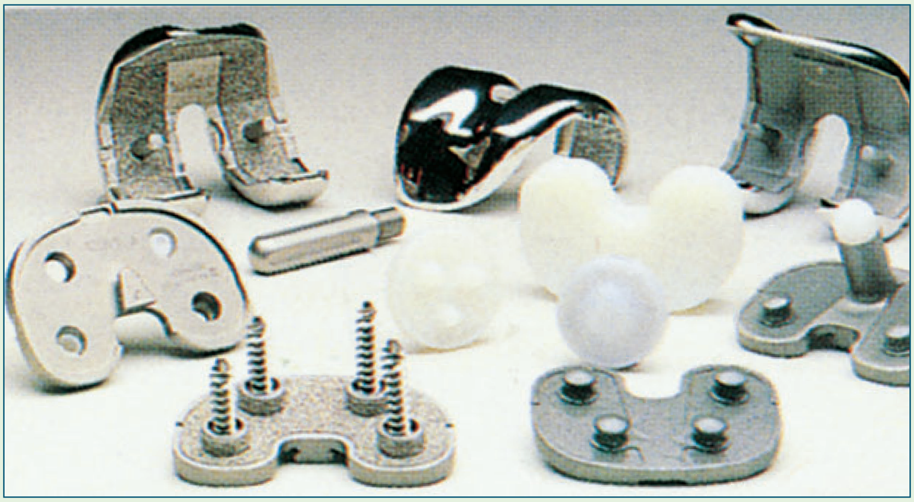


figure md.12 Body replacement parts like these knee parts are made from special alloys that will not harm the human body or break down in such a harsh environment.

figure md.13 A combination of the appropriate alloy and quenching produces the hardness of this chisel blade, which is used to shape wood.

Consider this

1 The melting temperatures and atomic radii of some elements are shown in table md.3. Using your knowledge of metallic structure and alloys, present a proposal for a possible new alloy. Design the alloy with a specific purpose in mind. Justify your choice of elements and predict the properties the alloy would have.

table md.3

Element	Melting temperature (°C)	Atomic radius (nm)	Element	Melting temperature (°C)	Atomic radius (nm)
Li	181	0.152	Mn	1244	0.124
Be	1278	0.113	Fe	1535	0.124
C	3527	0.077	Co	1495	0.125
Na	98	0.154	Ni	1453	0.125
Mg	649	0.160	Cu	1083	0.128
Al	660	0.143	Zn	420	0.133
Si	1410	0.117	Rb	39	0.248
K	64	0.227	Sr	769	0.215
Ca	839	0.197	Sn	232	0.141
Sc	1541	0.161	Pb	328	0.175
Ti	1660	0.145	Cs	28	0.265
V	1887	0.132	Ba	729	0.217
Cr	1857	0.125			

2 From your knowledge of metallic structure explain why alloying works. Why does the addition of other elements change the properties of the main metal?

3 What do you think is best—refining metal treatments and developing new alloys or producing new polymers and plastics from oil? Justify your response in terms of uses, cost and environmental factors.

4 Imagine iron didn't exist. How would this have affected the development of society? Would another metal have come into dominant use?

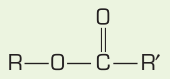
Experimental investigations

Experiment md.2 makes use of simple heat-treatment principles. Design your own experiment on the effects of smithing. You could compare different combinations of work hardening and heat treating.

flavours and odours

The flavour we associate with a particular food is due to a combination of taste and odour. Our sense of smell is extremely sensitive and is the main contributor to flavour. The flavour of a substance is due to many compounds, many of which are organic compounds such as aldehydes and alcohols. The focus here will be a group of organic compounds called esters, which give many flowers and fruits their characteristic flavour and odour. Figure fo.2 shows the general formula of an ester.

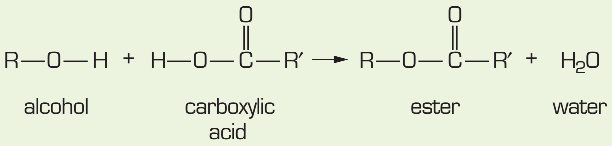
figure fo.1 The distinctive flavour of fruits is due to the production of particular esters.



$$R-O-\overset{\overset{O}{\|}}{C}-R'$$

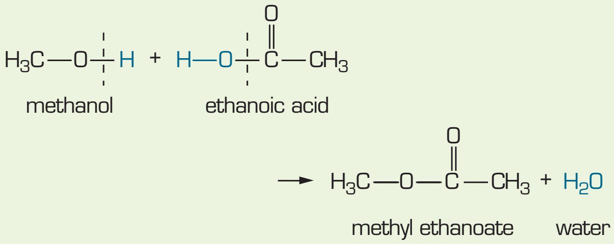
figure fo.2 The general formula for an ester, where R and R′ could be any hydrocarbon group such as $-CH_3$ or $-CH_2CH_2CH_3$. The abbreviated formula is RCOOR′.

figure fo.3 Many foods produced today contain artificial flavourings.



$$R-O-H + H-O-\overset{\overset{O}{\|}}{C}-R' \longrightarrow R-O-\overset{\overset{O}{\|}}{C}-R' + H_2O$$

figure fo.4 The reaction equation for esterification.



$$H_3C-O-H + H-O-\overset{\overset{O}{\|}}{C}-CH_3 \longrightarrow H_3C-O-\overset{\overset{O}{\|}}{C}-CH_3 + H_2O$$

figure fo.5 The formation of the ester methyl ethanoate.

Concentrated sulfuric acid is used to remove water from the reaction. It reacts with the water to form a hydrated form of sulfuric acid. This results in a higher yield during ester production.

* Section 2.5 **Organic compounds** describes some basic organic compounds and their properties (page 246).

Ester molecules are polar but they do not form hydrogen bonds as the hydrogen atoms present are not bonded directly to oxygen atoms. Esters are volatile due to the weak attractive forces between ester molecules and this is why we can readily detect their odour.

* Section 2.3 **Covalent bonding** describes types of intermolecular forces (page 233).

Artificial flavourings

Bananas produce a particular ester responsible for the odour of a banana. To reproduce the flavour and odour of a banana, this ester is produced artificially and added to processed foods. Most of the artificial flavourings used in foods, colognes and other products are esters. In fact, many of the foods we consume contain artificial flavourings.

Flavour chemistry is a large area of industrial chemistry that is responsible for the analysis of natural flavours. The task of preparing artificial flavourings to imitate natural flavours is very complex, as the overall flavour results from a mixture of different compounds, often with a single ester dominating. Advances in flavour chemistry have resulted in natural flavours being more closely imitated.

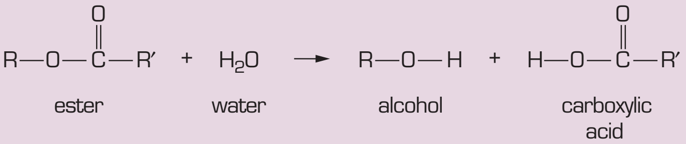
Synthesis of esters—esterification

Esterification is the manufacture of an ester using a reaction between a carboxylic acid and an alcohol. The reaction is heated in the presence of an acid catalyst, usually sulfuric acid. The products of the reaction are an ester and water. The reaction is often referred to as a condensation reaction due to the production of water. The general equation for the reaction is shown in figure fo.4.

By looking at the reaction equation, you will notice that the –OH group of the carboxylic acid bonds with the hydrogen from the alcohol group to form water. The reaction between methanol and ethanoic acid forms one of the simplest esters—methyl ethanoate (see figure fo.5).

Esters can form from any carboxylic acid and alcohol combination, so the range of possible esters is almost endless. Ester formation is now part of a new field of chemistry called combinatorial chemistry, which involves developing a selection of products resulting from all possible combinations of reactants. This field of chemistry is also used in the development of new drugs.

Esterification can also occur in the opposite direction, and is referred to as ester hydrolysis. Ester hydrolysis can be represented by the reaction below.



$$R-O-\overset{\overset{\displaystyle O}{\|}}{C}-R' + H_2O \longrightarrow R-O-H + H-O-\overset{\overset{\displaystyle O}{\|}}{C}-R'$$

figure fo.6 The reaction for ester hydrolysis. An ester reacts with water to form an alcohol and a carboxylic acid.

Ester hydrolysis occurs at the same time as esterification, so it is a challenge for manufacturers to produce the highest ester yield possible. The presence of any starting substances can affect the quality of the flavour. As a result, the final phase of ester production is purification.

figure fo.7 Esterification occurs naturally in flowers and fruits to produce particular flavours and odours. Esterification also occurs slowly during the ageing process of wine. As a wine ages, the alcohol and acids present react to produce esters which enhance the flavour.

Naming esters

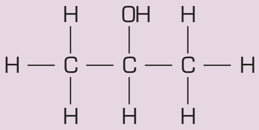
Esters are made from a reaction between an alcohol and a carboxylic acid. Esters are named using a combination of the systematic names for these reactants. To name an ester, you need to first be able to name the alcohol and the carboxylic acid used.

Systematic naming of alcohols and carboxylic acids

* Section 2.5 **Organic compounds** describes the systematic naming of hydrocarbons (page 246).

A hydrocarbon is named from its longest carbon chain. Similarly, an alcohol is named from the longest carbon chain (the parent chain) that contains the functional group –OH. In naming an alcohol, replace the parent chain ending with -ol. For most alcohols, the location of the functional group must be identified. To do this, number the carbons in the parent chain that will give the functional group, in this case the hydroxyl group –OH, the lowest number. Below are some examples of alcohols.

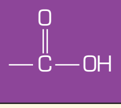
1. CH_3CH_2OH is ethanol, as it has a parent chain of two carbons and the hydroxyl group has replaced a hydrogen. In this case, the location of the hydroxyl group does not have to be identified because regardless of which carbon it is attached to it will always be attached to the first carbon.
2. $CH_3CH_2CH_2OH$ is 1-propanol. The parent chain has three carbons and the hydroxyl group is attached to the first carbon. The hydroxyl group could also be attached to the second carbon, in which case it would be named 2-propanol (see figure fo.8).



$$\begin{array}{ccccccc} & & H & & OH & & H & \\ & & | & & | & & | & \\ H & - & C & - & C & - & C & - & H \\ & & | & & | & & | & \\ & & H & & H & & H & \end{array}$$

figure fo.8 2-Propanol

The carboxyl group is commonly represented as –COOH, but you need to be aware that its structure is:



$$-\overset{\overset{\Large O}{\|}}{C}-OH$$

3 Figure fo.9 is 3-hexanol. The parent chain has six carbons and the hydroxyl group is attached to the third carbon.

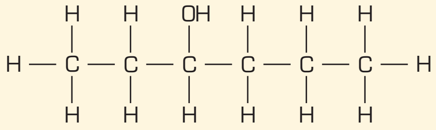


figure fo.9 3-Hexanol

The same rule applies for naming other organic compounds. Carboxylic acids contain the carboxyl functional group –COOH. In naming a carboxylic acid, the longest parent chain containing the carboxyl group must be identified and the ending replaced with -oic acid. There is no need to identify the carboxyl group with a carbon number, as it is always located at the end of the carbon chain. Below are some examples of carboxylic acids.

- CH_3CH_2COOH is called propanoic acid, as the chain has three carbons.
- CH_3COOH is called ethanoic acid, as the chain has two carbons. Ethanoic acid is its systematic name but you may know it by its common name—acetic acid.
- $CH_3CH_2CH_2CH_2CH_2CH_2CH_2COOH$ is called octanoic acid, as the chain has eight carbons.

Esters

As esters are made from a reaction between an alcohol and a carboxylic acid, there are two parts to the name of the ester. The alcohol is named first—the name of the original alcohol drops -anol and adds -yl. The second part of the name describes the carboxylic acid—the carboxylic acid drops the -ic ending and replaces it with -ate. For example, the ester formed from methanol and butanoic acid is named methyl butanoate.

figure fo.11 Methyl butanoate is the ester responsible for the odour you recognise as apple.

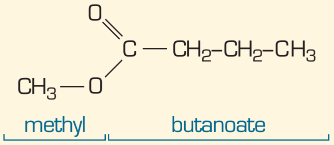


figure fo.10 The methyl part of the name comes from the alcohol methanol, and the butanoate is from the carboxylic acid butanoic acid.

table fo.1 The names and formulae of esters formed from various alcohols and carboxylic acids.

Alcohol	Carboxylic acid	Ester
ethanol CH_3CH_2OH	propanoic acid CH_3CH_2COOH	ethyl propanoate $CH_3CH_2OOCCH_2CH_3$
1-butanol $CH_3CH_2CH_2CH_2OH$	ethanoic acid CH_3COOH	butyl ethanoate $CH_3CH_2CH_2CH_2OOCCH_3$
methanol CH_3OH	ethanoic acid CH_3COOH	methyl ethanoate CH_3OOCCH_3
ethanol CH_3CH_2OH	butanoic acid $CH_3CH_2CH_2COOH$	ethyl butanoate $CH_3CH_2OOCCH_2CH_2CH_3$
1-propanol $CH_3CH_2CH_2OH$	pentanoic acid $CH_3CH_2CH_2CH_2COOH$	propyl pentanoate $CH_3CH_2CH_2OOCCH_2CH_2CH_2CH_3$

Experimental investigations

Preparing an ester

As mentioned earlier, it is possible to prepare an ester given any combination of carboxylic acid and alcohol. Regardless of the acid and alcohol used, the procedure is the same.

Method

The procedure provided is for the preparation of an ester from the reaction between 1-pentanol and ethanoic acid.

1. Half fill a 200 mL beaker with tap water and place on a heating element. Heat the water to approximately 60°C, and maintain between 60°C and 70°C.
2. Add 1 mL of ethanoic acid to a large test tube.
3. Add 2 mL of pentanol to the test tube.
4. Add 4–5 drops of sulfuric acid.
5. Gently heat the test tube in the water bath for 10–15 minutes.
6. Use tongs to remove the test tube from the water bath and cool under running tap water.
7. Add 2 mL of distilled water to the test tube. The ester should form a layer on top.
8. Add saturated sodium carbonate solution until there are minimal or no bubbles.
9. Waft the vapour produced toward your nose with your hand. Do not inhale the odour directly.

materials

- ethanoic acid (acetic acid), glacial
- 1-pentanol
- concentrated sulfuric acid (small dropping bottle), 18 M
- saturated sodium carbonate solution
- large test tube
- 5 mL measuring cylinder
- dropper
- 200 mL beaker
- heating element
- tongs
- distilled water
- thermometer

safety

- Wear safety glasses and a laboratory coat for this experiment.
- Concentrated sulfuric acid is extremely corrosive.
- Glacial acetic acid is corrosive.
- Never add water to concentrated acids.
- Alcohols, acids and esters are flammable.
- Do not allow the reaction mixture to boil or completely evaporate.

Experimental investigations

Obtain a range of alcohols and carboxylic acids, and produce as many esters as possible using different combinations. Write a reaction equation for the esters produced. Name each of the esters and draw their structure. For esters you produced, identify the odour. Are they familiar to you? Are any combinations similar in odour? Research the esters you produced and find out if you have correctly identified the odour.

Consider this

1. Your teacher may assign your group a particular odour. Use a selection of alcohols and carboxylic acids to determine the correct combination to produce the assigned odour.
2. Research the esterification process and develop your own combinations to produce particular odours. Write a reaction equation for the esters produced. Name each of the esters and draw their structure.
3. Perform a blind odour test on other class members and get them to identify the odour of the ester produced. You could even compare the odour with the actual item being imitated.

Further investigations

Research one of the following topics. Present your findings as an oral presentation, poster or PowerPoint presentation.

- Separation and purification of esters.
- Commercial preparation of a particular ester.
- Properties of esters.
- Fats and oils (which are tri-esters)—their structure, preparation and properties.
- The role of esters in pheromones of insects. Some plants, such as orchids, chemically mimic the odours of pheromones to attract pollinating insects.
- Saponification of esters. This process is used in soap production, and involves the reaction of fatty acid esters with a strong base.
- Esters are used widely in other areas apart from food. Find out other areas where esters are used, including in cosmetics, solvents and medical preparations (such as benzocaine and procaine).
- The role of esters in the plastics industry, such as polyester (many esters).
- The use of esters and flavours in animal feed.
- Combinatorial chemistry.

Contexts

The air we breathe
Consumer chemistry
Water quality

Chapters

3 Stoichiometry
4 Chemical solutions
5 Gases

area of Study 3

Chemical Quantities

the air we breathe

The air we breathe is essential to life on this planet, yet it would be true to say that most people do not appreciate its importance. Human activity has changed the nature of the atmosphere in recent decades. Scientists have only begun to realise the impact these changes in the atmosphere could have upon life on this planet.

The nature of air

The Earth is surrounded by an envelope of air we call the atmosphere. The name is derived from the Greek word ‘atmos’ which means vapour.

Besides containing the gases needed for photosynthesis and respiration, which are essential for plant and animal life, other vital functions of the atmosphere include:

- absorbing much of the damaging ultraviolet radiation from the Sun
- trapping some of the energy radiated by the surface of the Earth to provide a stable climate
- allowing the cycling of water, which provides rain and other forms of precipitation, over the land masses of Earth.

The important functions of gases in our atmosphere can be attributed in part to the unique properties of gases.

* Section 5.1 **Physical properties of gases** describes the properties of gases (page 305).

figure ab.1 Babies who are premature or have heart or lung problems rely on a supply of air enriched with oxygen to survive their early days.

The layers of the atmosphere

The Earth's atmosphere is divided into layers. The layers are separated by narrow transition zones. These layers have different chemical compositions which produce variations in temperature. As the distance from Earth increases, the chemical composition of the air changes. At higher altitudes, the atmosphere becomes enriched with lighter gases.

The troposphere

The troposphere is the layer closest to the Earth's surface. This layer is said to be about 15 km high; however, its thickness varies according to its location. It is thickest at the equator (about 18 km) and thins out at the poles (about 8 km).

The troposphere contains about 75% of the gases that make up the atmosphere and about 99% of the water vapour. This means that this layer is where nearly all weather phenomena occur, including most clouds and all precipitation. The global average surface temperature is said to be about 15°C. Air temperature drops to about –60°C at the top of the troposphere. The troposphere is separated from the next layer, the stratosphere, by a narrow zone called the tropopause. Within the tropopause, the air temperature remains relatively constant, even as the altitude increases.

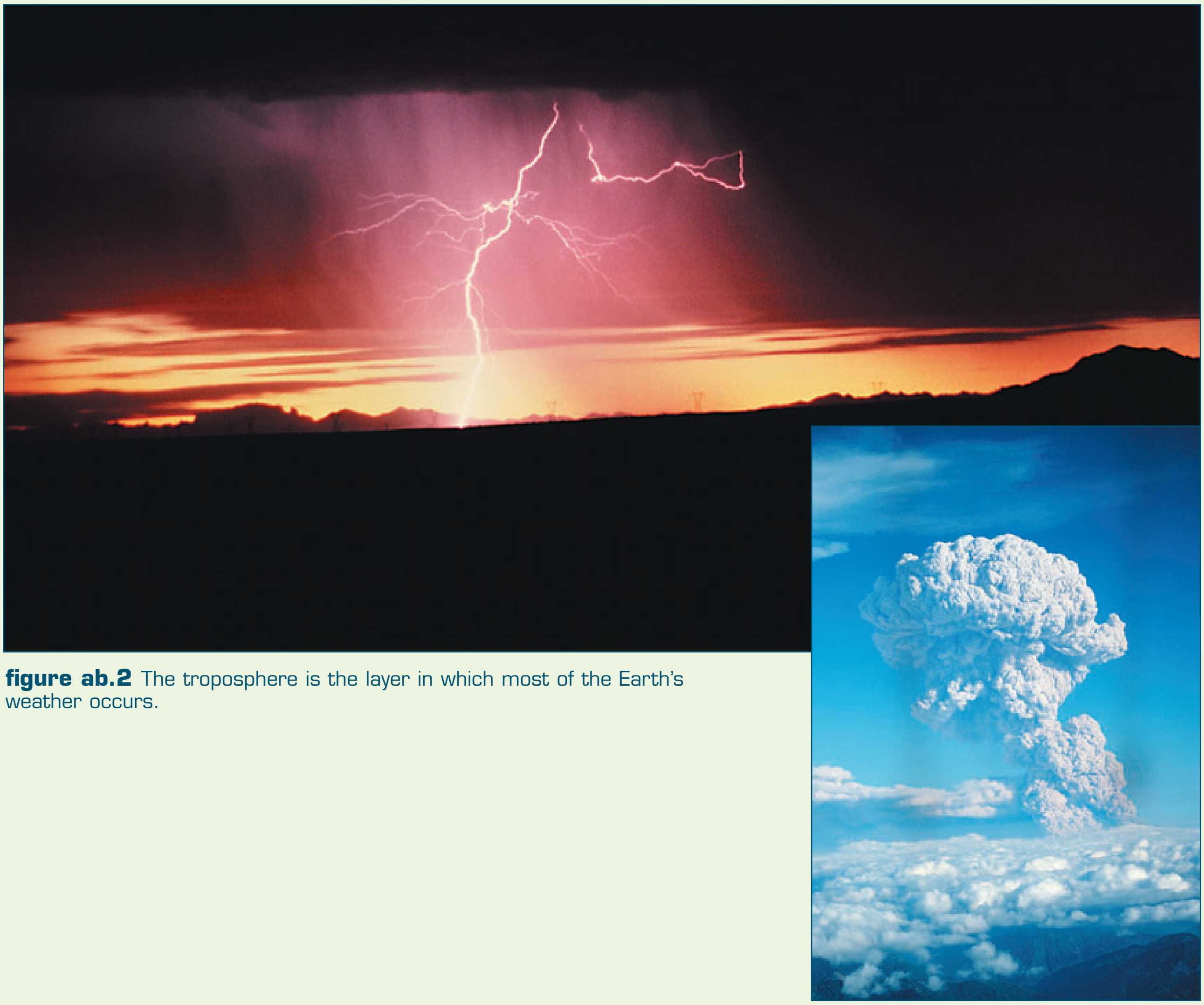


figure ab.2 The troposphere is the layer in which most of the Earth's weather occurs.

The stratosphere

The second layer of the atmosphere is called the stratosphere. It extends to approximately 50 km above the Earth's surface. The temperature remains relatively constant to about 25 km above the surface. After this, it increases with altitude. The temperature at the top of the stratosphere varies with season and position on Earth, but it averages about 25°C.

The stratosphere contains the very important ozone (O_3) layer. The ozone layer is able to absorb harmful solar ultraviolet radiation, protecting life on earth from its damaging effects. The stratosphere contains about 90% of the atmosphere's ozone. Most of the remaining ozone is in the troposphere (where its effects are not so beneficial). Most of this ozone is concentrated between altitudes of 20–30 km above the surface of the Earth.

Ozone is produced in a cyclic process by the action of UV radiation on oxygen molecules in the upper stratosphere. The ozone then travels to the lower stratosphere.

Generally, most ozone production and destruction occurs where the largest amount of UV radiation is present (the tropics). The exception to this is the destruction of ozone over the polar caps due to the action of CFCs (chlorofluorocarbons).

The production and destruction of ozone produces energy in the form of heat. This explains the higher than expected temperatures in the stratosphere.

At the top of the stratosphere is the stratopause. At this level, temperature begins to decrease with altitude.

The mesosphere

The mesosphere is the region from the stratopause to about 80 km above the Earth's surface. In this layer, temperature decreases with altitude and there is very little gas present. Temperatures at the mesopause can drop below –100°C. Most meteors that approach Earth burn up in the mesopause.

The thermosphere

The thermosphere makes up only about 0.01% of the total mass of the Earth's atmosphere. The lower parts of the thermosphere contain molecular and atomic oxygen. As the thermosphere extends outward, the chemical composition changes. The outer layer of the thermosphere contains mostly the lighter elements hydrogen and helium. The lower part of the thermosphere is known as the ionosphere. In this area, radiation bombards the gases to form ions and free electrons. This feature of the thermosphere is important for short wave radio communication.

* Section 2.2 **Ionic bonding** describes the formation of ions (page 229).

During times of increased solar flare activity, high energy particles from the flares cause gases in the thermosphere to give off visible light of various wavelengths. This produces the spectacular phenomena we call the aurora borealis (over the north pole) and aurora australis (over the south pole).

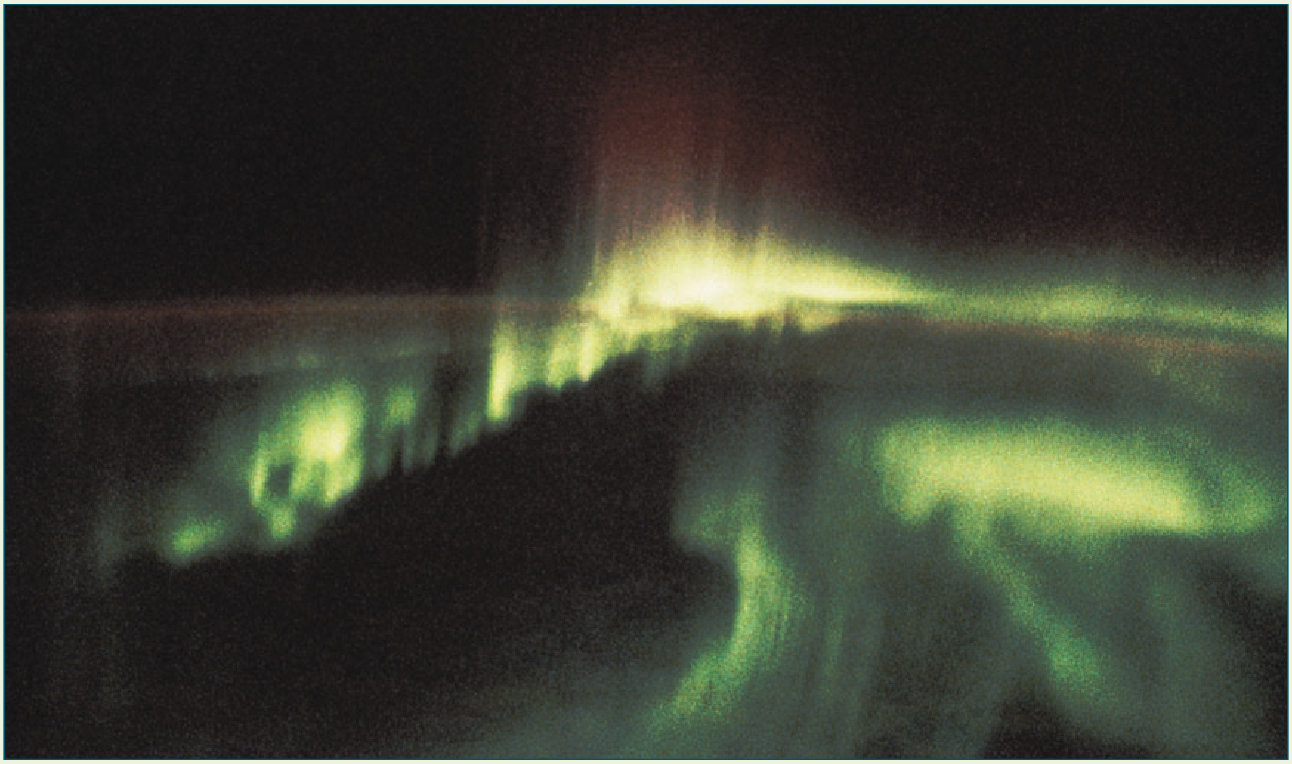


figure ab.3 The spectacular aurora can be seen in the polar regions during periods of increased solar flare activity.

Gases of the atmosphere

The atmosphere is a mixture of many different gases. However there are just two major component gases—nitrogen and oxygen. All other gases combined make up less than 1% of the atmosphere. The table below shows the average composition of dry air in the lower atmosphere.

table ab.1 The average composition of dry air near sea level.

Gas	Content by amount (mol %)	Gas	Content by amount (mol %)
nitrogen	78.08	methane	0.000 14
oxygen	20.95	krypton	0.000 1
argon	0.93	hydrogen	0.000 05
carbon dioxide	0.035	carbon monoxide	0.000 02
neon	0.001 8	ozone	0.000 004
helium	0.000 52	other gases, e.g. Xe, SO_2, NO_2	very small amounts

Generally the composition of the atmospheric mixture remains fairly constant. There are two major exceptions.

1 Ozone—most of the ozone is concentrated in the ozone layer in the stratosphere. The rest is in the troposphere.
2 Water vapour—most water vapour (about 99%) is found in the lower few kilometres of the troposphere. It is found in higher concentrations over the oceans than land masses. The concentration of water vapour is very changeable with seasons and general weather conditions.

Some of the other gases in the atmosphere will vary in composition due to changes in season and are affected by human activity, such as processes that cause air pollution.

Nitrogen

Molecular nitrogen is the most abundant gas in the atmosphere. It has a very important role in sustaining life on Earth. Atmospheric nitrogen itself is of little use to life processes. However there are a number of ways it can be changed into more usable forms such as nitrates (NO_3^-), for example by lightning and certain nitrogen-fixing bacteria.

Nitrogen in the atmosphere exists as a diatomic molecule. It is a colourless, odourless gas that is virtually insoluble in water. It is more soluble in organic liquids such as alcohols and hydrocarbons.

✱ Section 4.1 **Dissolving** describes the solubility of gases (page 286).

Molecular nitrogen is quite unreactive. This is related to its structure. Diatomic nitrogen (N_2) has a very strong triple bond between its atoms.

✱ Section 2.3 **Covalent bonding** describes the bonding of molecular substances including nitrogen (page 233).

Despite its relative unreactivity, molecular nitrogen gas has many important uses, including:

- the production of fertilisers such as ammonia, ammonium sulfate and ammonium nitrate
- as liquid nitrogen for the snap freezing of foods and storage of biological samples such as eggs, sperm and embryos
- to provide an oxygen-free environment that is useful for storing and preserving fresh fruit, etc.

figure ab.4 Human activity such as the combustion of fossil fuels in car engines can change the amount of certain gases in the atmosphere. This can lead to problems such as photochemical smog.

figure ab.5 Some plants (such as peas and beans) have nodules on their roots that contain nitrogen-fixing bacteria. These bacteria can change (fix) atmospheric nitrogen into the more useful nitrate form.

figure ab.6 The solubility of oxygen gas in water is essential to organisms such as this fish.

Oxygen

Molecular oxygen is the second most abundant gas in the atmosphere. Most humans would consider it the most important gas on the planet, as it is essential to most life on the planet for survival.

Oxygen is a colourless, odourless gas. It is denser than air and slightly soluble in water. It is extremely chemically reactive. It forms compounds with most elements except the noble gases.

The uses of oxygen vary from the saving of human lives to industrial processes and include:

- medical use for people who have heart or lung problems, diseases such as pneumonia or exposure to toxic fumes (from fires etc.)
- in combustion processes such as fossil fuels for power production
- for the treatment of sewage
- in the manufacture of steel and chemicals such as sulfuric acid.

Carbon dioxide

Carbon dioxide is often thought of as a problem gas by humans. Human activities such as burning fossil fuels and deforestation have increased the amount of CO_2 in the atmosphere. This is believed to contribute to the greenhouse effect and global warming. However carbon dioxide is as essential to the existence of life as oxygen.

figure ab.7 Oxygen is used in the manufacturing of steel. It combines with impurities to form compounds that can be easily removed.

figure ab.8 Carbon dioxide is dissolved in drinks at high pressure. When the drink is opened, the pressure is released, solubility decreases and bubbles of gas escape from the solution.

figure ab.9 As well as being a useful refrigerant, dry ice is often used to produce a mist effect in theatrical productions.

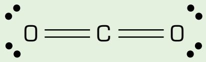


figure ab.10 The structural formula for CO_2 showing the two carbon–oxygen double bonds.

Carbon dioxide is a colourless, odourless gas that is denser than air and slightly soluble in water. At normal atmospheric pressure it does not form a liquid but forms a solid (dry ice) when cooled. Dry ice changes directly from the solid phase to the gas phase and is useful as it does not leave puddles of liquid behind. The two carbon–oxygen double bonds make it a very stable molecule. It does not support combustion under most normal circumstances.

CO_2 has many uses. It is:

- a raw material for plant photosynthesis
- used in fire extinguishers
- used in the production of carbonated soft drinks
- used as a refrigerant in its solid form (dry ice).

figure ab.11 Clearing of forests has contributed to an increase in CO_2 levels as there are fewer trees using up CO_2 during photosynthesis. Carbon dioxide is also produced when forests are burned or left to decay.

Breathing under water

The air that surrounds you exerts a pressure on your body. When you are swimming at the surface of the water, this pressure is usually about one atmosphere.

* Section 5.3 **Measuring pressure and volume** describes units for measurement of gas pressure (page 308).

The further you go below the surface of the water, the greater the pressure on your body becomes. For every ten metres you descend, the pressure on your body increases by about one atmosphere. For example, at a depth of ten metres, the pressure on your body is about two atmospheres. When you dive below the surface holding your breath, the amount of air in your body cavities (such as lungs and inner ears) stays the same. However, the increased pressure from the water causes the cavities to decrease in size. In this smaller volume the pressure of the trapped air increases (matching the external pressure). This becomes very uncomfortable, limiting how long a person can hold their breath under water.

* Section 5. 4 **The gas laws** describes the relationship between the pressure and volume of a gas (page 310).

Scuba diving equipment allows a person to dive deeper under the water for a longer period of time. Most divers use tanks of highly compressed air, which is supplied through a device called a regulator. The air inhaled is at the same pressure as the surrounding water, allowing the person to breathe comfortably.

A diver must exhale constantly as they ascend to the surface. If a diver holds their breath, the decrease in pressure as they rise causes an increase in the volume of their lungs and other cavities. This can cause serious problems.

The pressurised air breathed by divers is mostly nitrogen (as we know from the composition of air). As the pressure of the air is increased, the pressure due to the nitrogen (partial pressure) in the mixture increases accordingly.

* Section 5.2 **Kinetic molecular theory** describes partial pressure of gases (page 306).

The amount of gas that will dissolve in a liquid is proportional to the partial pressure of the gas. As the pressure of the nitrogen increases, the amount of nitrogen that dissolves in the body increases. The nitrogen is not used by the body and builds up in the tissues. As a diver rises to the surface and the pressure decreases, the nitrogen in the tissues comes out of solution.

If the diver ascends too quickly, the pressure change is rapid and the dissolved nitrogen returns to its gaseous state, forming bubbles. These bubbles can cause pain, particularly in the joints. They can also affect the nervous system, with more serious consequences such as paralysis or even death. This condition is called decompression sickness (or the bends). It can be avoided by ascending slowly so that the pressure change is gradual and the excess nitrogen is exhaled.

figure ab.12 Scuba divers must ascend at the right rate to allow their body to adjust to the change in pressure.

figure ab.13 Divers suffering from the bends can be treated in a special decompression chamber.

Further investigations

1 'The ozone layer is like sunscreen for the Earth.' Prepare a short report on the ozone layer. Describe how ozone is formed and dissociated, writing chemical reactions for these processes. Discuss whether you agree with the statement above and why. Include recent information on the depletion of the ozone layer, including any relevant chemical reactions.

2 You have been assigned the task of researching simple methods for the production and testing of oxygen and carbon dioxide gases. You have been asked to present the results of your work in the form of an experimental procedure which would be suitable for use by Year 10 students. The experiment should be written in the standard format. Make sure that you include safety and disposal information, and discussion questions.

3 Divers who suffer from decompression sickness (the bends) are usually treated using a device called a hyperbaric chamber. You have been asked to prepare a presentation for a group of navy divers explaining the chemistry involved (particularly pressure issues) with decompression sickness and treatment using the hyperbaric chamber.

4 In recent years, a substance called 'Nitrox' has become available for use for recreational SCUBA divers. You have been asked to prepare a pamphlet suitable for distribution to dive shop customers advertising the use of this substance compared to compressed air. Your pamphlet should include features such as an explanation of:

- what it is
- why they would use it
- the advantages and disadvantages of Nitrox
- the restrictions on its use.

5 Investigate the sources and uses of a few of the minor gases in the atmosphere (see table ab.1). Present your findings on a poster which would be suitable to display in your chemistry classroom.

6 The local environmental group is running a public awareness campaign on several environmental issues. You have been assigned the task of preparing an interesting and informative fact sheet suitable for distribution to the general public concerning the greenhouse effect and global warming.

Consider this

1 Prepare a visual display to illustrate the layers of the atmosphere and their major features.

2 Calculate the approximate volume of air in your classroom. Use the given table of atmospheric composition to calculate the approximate mass of nitrogen, oxygen and carbon dioxide gas in the room.

Assume SLC applies when making your calculations.

3 **a** The average growing tree absorbs approximately 20 kg per year of CO_2. The burning of 1 tonne of coal produces approximately 4 tonnes of CO_2. One coal-fired power plant uses approximately 13 million tonnes of coal per year. How many average trees would be needed to absorb the CO_2 produced assuming conditions of SLC?

figure ab.14 How many trees would have to be planted to compensate for the burning of fossil fuels?

b Estimates of the mass of the Earth's atmosphere vary, but many agree to approximately 5.2×10^{18} kg. Use table ab.1 to calculate the approximate mass of nitrogen, oxygen and carbon dioxide in the atmosphere.

c Water can be decomposed into hydrogen and oxygen gas. What volume of water would need to be decomposed to produce the amount of oxygen in part b?

consumer chemistry

Many of the substances used around the home are complex mixtures of compounds that would be difficult to identify individually. However, manufacturers are required to list the ingredients (or at least the active ingredients) of some products so that consumers can use them safely. Knowing the ingredients allows safe cleaning procedures to be used in the case of a spillage and correct first aid to be applied if the product is accidentally swallowed or makes contact with various parts of the body. Some people have allergies or an intolerance for certain substances and in this situation a knowledge of ingredients is vital.

By knowing the contents of a product we can also compare it with similar products of other manufacturers and work out whether one brand is superior to another as well as whether they are good value for money. Your growing knowledge of chemistry will open your eyes to the true nature of some common household products.

Chemistry in the supermarket

This context will give you some ideas of investigations that you may undertake. You might like to select a range of household products and decide whether they represent good value for money.

What puts the 'spring' in spring water?

This bottle of spring water (figure cc.1) has an analysis listed on the label.

> Ingredients: Natural spring water. Typical analysis mg/L: bicarbonate 146, calcium 25, chlorides 20, sodium 19, magnesium 12, potassium 3. Product of Australia.

figure cc.1 A wide variety of spring and mineral waters are available. In what ways are they similar to each other? How are they different?

* Section 2.2 **Ionic bonding** describes ions and how they are formed (page 229).

The term bicarbonate used here refers to the hydrogen carbonate ion.

This means that each litre of this brand of spring water contains:

- 146 mg of bicarbonate ions HCO_3^-
- 25 mg of calcium ions Ca^{2+}
- 20 mg of chloride ions Cl^-
- 19 mg of sodium ions Na^+
- 12 mg of magnesium ions Mg^{2+}
- 3 mg of potassium ions K^+.

The ions in spring water come from various substances that dissolve in the water as it moves through the ground. From the list above we can suggest what compounds may possibly be dissolved in the water (table cc.1).

table cc.1 Possible dissolved compounds present in bottled water.

calcium chloride	$CaCl_2$	calcium bicarbonate	$Ca(HCO_3)_2$
sodium chloride	$NaCl$	sodium bicarbonate	$NaHCO_3$
magnesium chloride	$MgCl_2$	magnesium bicarbonate	$Mg(HCO_3)_2$
potassium chloride	KCl	potassium bicarbonate	$KHCO_3$

* Section 4.1 **Dissolving** describes why these substances are soluble (page 286).

Active ingredients

A confusing number of brands of some products exist on supermarket shelves. When it comes to products like vinegar, aspirin and bleach (figure cc.2) you are likely to find at least three or four brands of each.

Vinegar

Vinegar is known for its acidic character—the sour taste. All brands of vinegar contain ethanoic acid (acetic acid), CH_3COOH (figure cc.3); however, since it is used in foods the quality of vinegar is judged on more than its ethanoic acid content. The vinegar used in a salad dressing needs to have an appealing flavour while for the one used as an environmentally friendly cleaner in the bathroom the flavour is unimportant.

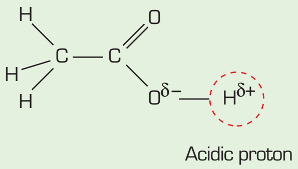


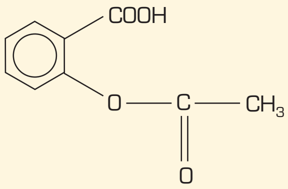
figure cc.3 The acidic properties of ethanoic acid are due to the hydrogen atom attached to the oxygen. It can detach forming a solution containing hydrogen ions.

figure cc.2 Multiple brands of these products can be found in the supermarket but the active ingredients are often identical.

* Section 2.5 **Organic compounds** describes carboxylic acids, such as ethanoic acid (acetic acid) (page 246).

Aspirin

Aspirin is a commonly purchased over-the-counter pain reliever. It is marketed under a wide variety of brand names but they all contain the same active ingredient—aspirin (figure cc.4). The brand shown in figure cc.2 has stated on the packet that each tablet contains 500 mg of aspirin. Each aspirin tablet obviously contains a large amount of something other than aspirin.



Aspirin

figure cc.4 The structural formula of aspirin.

Bleach

Bleach is a cleaning product frequently used to remove stains from white clothing and as a powerful germ and mould killer in toilets, bathrooms and kitchens. Most brands of bleach contain two main active ingredients—sodium hypochlorite ($NaOCl$) and sodium hydroxide ($NaOH$). The brand of bleach shown in figure cc.2 is labelled as containing 42 g L^{-1} sodium hypochlorite (or 4% m/v) and 9 g L^{-1} sodium hydroxide.

* Section 4.2 **Concentration** describes concentration of a solution in mol L^{-1} (page 296).

Products bought in the supermarket are likely to have units in one of the forms described in table cc.2

table cc.2 The units of concentration as often found on the labels of consumer products.

Concentration	Formula	Units
percent by mass	$\text{concentration} = \frac{\text{mass of solute (g)}}{\text{mass of solution (g)}} \times 100\%$	w/w
percent by volume	$\text{concentration} = \frac{\text{volume of solute (mL)}}{\text{volume of solution (mL)}} \times 100\%$	v/v
mass of solute per volume of solution	$\text{concentration} = \frac{\text{mass of solute (g)}}{\text{volume of solution (mL)}} \times 100\%$	w/v or m/v
mass of solute per mass of solution	$\text{concentration} = \frac{\text{mass of solute (mg)}}{\text{mass of solution (g)}}$	mg/g

The bleach described above contains 42 g of sodium hypochlorite in 1 L of solution. Using the formula for mass of solute per litre of solution:

$$\begin{aligned}\text{concentration} &= \frac{\text{mass of solute (g)}}{\text{volume of solution (mL)}} \times 100\% \\ &= \frac{42\text{ g}}{1000\text{ mL}} \\ &= 4.2\text{ m/v}\end{aligned}$$

The 4% m/v listed on the label must be a rounded value.

The reactions of stomach relief

Antacid powder (figure cc.5) is designed for the relief of an upset stomach. This product is added to water, it foams up and is then swallowed. The label on the bottle states that each 5 g contains:

- 2.29 g of sodium bicarbonate ($NaHCO_3$)
- 0.50 g of sodium carbonate (Na_2CO_3)
- 2.16 g of anhydrous citric acid.

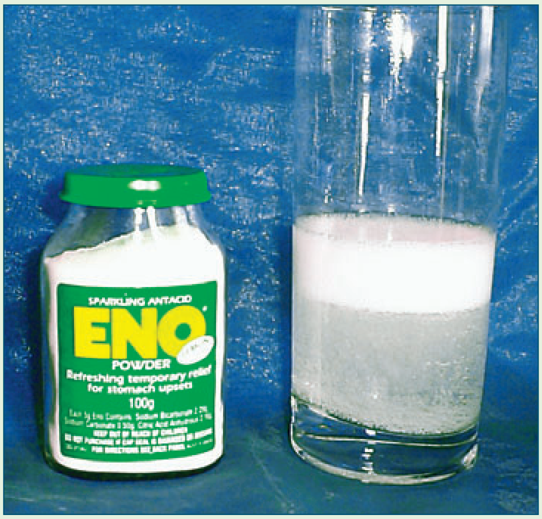


figure cc.5 Antacids are designed to neutralise stomach acids in an upset stomach.

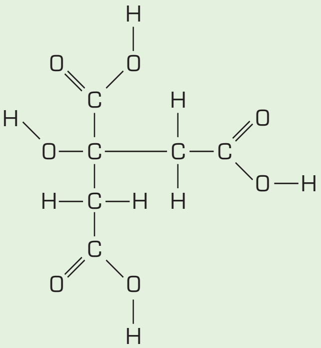


figure cc.6 The structure of citric acid. Each citric acid molecule has three hydrogen atoms available to form hydrogen ions. They are the hydrogen atoms attached to an oxygen in the –COOH group.

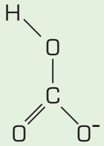


figure cc.7 The structure of a hydrogen carbonate ion.

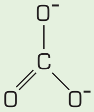


figure cc.8 The structure of a carbonate ion.

While the ingredients are dry, no chemical reaction occurs, but in water the substances dissolve and form ions which react with each other.

Citric acid (figure cc.6) is a source of hydrogen ions (H^+), sodium bicarbonate of hydrogen carbonate ions (HCO_3^-, figure cc.7) and sodium carbonate a source of carbonate ions (CO_3^{2-}, figure cc.8). These can react as shown in table cc.3.

✱ Section 3.2 **Chemical reactions** describes how to write a balanced chemical equation (page 272).

table cc.3 The reactions of antacid ingredients.

acid and carbonate	$2H^+(aq) + CO_3^{2-}(aq) \rightarrow H_2O(l) + CO_2(g)$
acid and bicarbonate	$H^+(aq) + HCO_3^-(aq) \rightarrow H_2O(l) + CO_2(g)$

The carbon dioxide gas produced forms the bubbles in the effervescent drink. Similar reactions will occur between the carbonate and hydrogen carbonate ions and stomach acids, neutralising some of the stomach acid and providing relief from stomach upset.

Sodium bicarbonate can be found on the supermarket shelf as baking soda and citric acid can also be found in the aisle containing baking ingredients. A form of sodium carbonate can be found in the cleaning section of the supermarket, labelled as washing soda.

figure cc.9 The active ingredients of antacid powder can be found in other sections of the supermarket.

Further investigations

Bottled water

1. Collect a variety of brands of spring and mineral water. Compare the ions and possible compounds that they contain. Compare the amounts of each that they contain.
2. Find out from your local council the minerals contained in your local water supply. How does this compare with bottled water?
3. Check your school's chemical stock. Would it be possible to 'manufacture' spring water? How would the cost compare with the natural product?

Active ingredients

1. Calculate the cost per millilitre of several different brands of vinegar. Compare the quality of each brand by taste. Do the more expensive vinegars taste better? Is the price difference justified?
2. Perform a similar investigation for some other products. For example, all brands of salt must be made up of sodium chloride ($NaCl$). Can you account for the different prices of different brands? Some brands contain other ingredients such as potassium or iodine. Why might they contain these ingredients?
3. Obtain some different brands of aspirin. Compare them according to their content of aspirin per tablet as well as by cents per milligram of aspirin. Which brand is the best value for money? Are there other features of certain brands (for example solubility) that make a higher cost justified?
4. Other pain relievers contain paracetamol as their active ingredient. For different brands of paracetamol perform an investigation similar to that for aspirin in the previous question, checking features such as amount of active ingredient, cost and solubility. Compare aspirin products with paracetamol products. What are the advantages and disadvantages of each?
5. Obtain several brands of household bleach from the supermarket and compare their active ingredients and concentrations. It may be necessary to convert the concentrations to the same units (perhaps $mol\ L^{-1}$). Can you account for the price difference for different brands? Are the generic brands very different to the well known ones?
6. Compare the contents of different brands of bleach with the more expensive mould and shower cleaners. Is the price difference warranted?
7. Sodium hydroxide is also known as caustic soda and is available in supermarkets. Compare the prices of different brands of sodium hydroxide. Check the contents of expensive drain cleaners and compare the price with that of caustic soda. Calculating cents per gram might be a useful way of comparing them.

Stomach relief

1 Do a survey of the supermarket shelves to find the cost of antacid powder and its individual ingredients. If you took the time to combine the ingredients in the same proportions to make your own antacid powder, how would it compare in price? (Note that washing soda may not be of a quality suitable to be taken by humans.) Do the various brands of antacid powders contain the same ingredients? Compare the prices of brands with similar ingredients.

2 Check the stoichiometric amounts of citric acid, sodium carbonate and sodium bicarbonate. How much carbon dioxide will be produced? How much sodium carbonate and sodium bicarbonate are available for neutralising stomach acid?

figure cc.10 Antacids can be used to relieve the discomfort of an upset stomach caused by over eating and other medical conditions.

* Chapter 3 **Stoichiometry** describes the relationships between reactants and products in chemical reactions (page 268).

3 What is the purpose of the antacid effervescing? Does effervescence have a role in relieving stomach upset? If neutralisation of stomach acid is the purpose of antacid products, would a person be better off just taking baking soda in water?

water quality

People need a reliable, clean water supply for drinking, cooking and cleaning. Indirectly, we need a good supply of water for the growing of food, care of livestock and other animals, growing timber for the construction of buildings, and for the production of consumer goods such as clothing. Less directly (though still essential), healthy water is needed to maintain the ecosystems that supply us with oxygen and cycle matter through the atmosphere, soil and waterways.

figure wq.1 We rely on water for many things in our everyday lives.

A sustainable water supply for each of these purposes can only be achieved through management of catchment areas, regular monitoring of quality and conservation principles.

Local water

Most water supplied to consumers in Australia is collected in large dams. This water is collected after flowing across land and then through creeks and rivers. Most problems with water quality are the result of human activity in these catchment areas. Erosion, animal wastes and pollutants such as fertilisers and pesticides entering the water system can affect water quality. Care needs to be taken to ensure that the catchment areas of such dams are clean and unpolluted so that the collected water is safe for humans.

figure wq.2 The Hinze Dam supplies water to the Gold Coast. It has a catchment area of 207 km^2, a storage capacity of 163 500 million litres and a surface area of 972 hectares. It was originally built in 1976 and the wall was raised a further 18 metres in 1989.

figure wq.3 Water monitoring requires regular and careful collection of water samples.

Local ecosystems also depend on a quality water supply. Poor water systems can result in the spread of disease, the depletion of fish stocks, problems with algae and the degradation of swimming and boating areas.

Local councils try to guarantee the quality of their water resources by taking protective measures that include:

- educating residents and businesses on water conservation and quality
- enforcing water restrictions to maintain our fresh water reserves
- fencing watercourses to keep livestock away
- ensuring that agricultural areas undertake practices to limit contamination by herbicides and pesticides
- limiting the influx of nutrient substances such as phosphates and nitrates that come from sources such as fertilisers and detergents
- limiting erosion on agricultural, development and building sites
- utilising effective bushfire management practices
- limiting the stormwater runoff from urban areas
- ensuring that chemicals are stored safely away from waterways
- making sure that wastewater is treated and disposed of by efficient and well maintained systems
- planning land use for residential, agricultural, industrial, recreational and tourism uses along ecologically sustainable principles.

* Section 4.1 **Dissolving** describes why some substances readily dissolve in water and others don’t (page 286).

Most local councils and some other organisations such as environmental groups, perform regular testing of waterways. They monitor the effects of land use on water quality. In this way, particular incidents (e.g. floods or power failures) or uses (e.g. factories or abattoirs) can be identified and their impact on the quality of the water assessed. Corrective and preventative measures can be put into place if necessary.

The following section describes the types of tests that can be used to monitor the quality of water.

Testing water quality

You can choose to perform all or some of the following tests depending on your concerns regarding the area you are monitoring. Some tests can be performed at the site and others require the collection of samples for laboratory testing.

* Section 2.2 **Ionic bonding** describes the use of chemical formulae. This will help identify many of the compounds used in these tests (page 229).

* Section 3.2 **Chemical reactions** describes how to write the chemical reactions. You should be able to write the equations for the reactions that take place in some of these tests (page 272).

* Section 4.2 **Concentration** describes the units scientists use for concentration and the formulae for the calculations performed in these tests (page 296).

Experiment wq.1

Qualitative analysis of water at the collection site

Method

1. Visually inspect the water sample for hydrocarbons (oil, petrol or organic matter). Look for a rainbow effect on the surface of the water and test for odour.
2. Record the temperature of the sample.
3. Half fill a small conical flask with a sample of water, stopper it and shake for 1 minute. Stop shaking it and immediately start the stopwatch to record the time taken for any bubbles to disappear. The slow disappearance of bubbles indicates the presence of a detergent.
4. Repeat step 3 for tap water and distilled water for comparison.
5. Lower a turbidity tube into water on location until the bottom of the tube (showing a cross) just disappears. Record the reading on the tube in NTU. Do the comparison with tap water.
6. If available, insert the probe of a portable pH meter into a water sample and record the reading.
7. If a pH meter is not available, test the water sample with indicator paper to establish the general pH range, then use the appropriate narrow-range pH paper and estimate pH as accurately as possible.
8. Tests for heavy metals can be performed as follows. Place five drops of 0.001 M lead nitrate solution in a clean test tube and add one drop of sodium sulfide solution. Add 5 mL distilled water to the test tube. Keep this test tube for comparison with other samples. Place a 5 mL sample of water into a test tube and add 1 mL sodium sulfide solution. The presence of a precipitate indicates heavy metals. Also perform this test on tap water.
9. Design a table to record your results and allow comparison at different times and/or sites.

materials

- thermometer
- 25 mL conical flask with stopper
- stopwatch
- turbidity tube
- pH meter (portable) or universal indicator paper and narrow range pH papers
- test tubes
- 5 mL 0.001 M lead nitrate solution ($Pb(NO_3)_2$)
- 5 mL 0.1 M sodium sulfide solution (Na_2S)

figure wq.4 Rainfall or flooding combined with soil that is unprotected by vegetation can result in turbid water.

Experiment wq.2

Qualitative analysis of dissolved oxygen and BOD in a water sample

Method

Use oxygen sensors from the StreamWatch kit or follow the laboratory procedure below.

1. At the collection site, fill four 300 mL reagent bottles with water and stopper them, ensuring that there are no air bubbles. On the label record all relevant information including the temperature. Label one bottle BOD (for biological oxygen demand) and put it aside. In the laboratory, fill a fifth bottle with distilled water. This is the control.
2. To each of the four other bottles add 2 mL manganese sulfate solution ($MnSO_4$) and 2 mL alkaline sodium iodide solution (NaI) using a graduated pipette or syringe.
3. Stopper well and invert several times to ensure thorough mixing. A brown sludge will form. If you need to remove some liquid before adding reagents, accurately record the volume removed.
4. Remove 4 mL liquid from the first reagent bottle, then add 4 mL of 9 M sulfuric acid solution (H_2SO_4). Stopper the bottle and mix. The sludge should disappear.
5. Transfer the entire contents of the reagent bottle into a 500 mL conical flask and add about ten drops of starch indicator solution. The solution should turn blue.
6. Fill a clean burette with 0.020 M sodium thiosulfate solution ($Na_2S_2O_3$). Record the initial burette reading.
7. Slowly titrate the blue solution with sodium thiosulfate, swirling constantly until the blue colour just disappears.
8. Repeat steps 4–7 with the other two water samples and the control.
9. Place the BOD sample of water in a dark space at 20°C for 5 days. At the end of this period, determine the dissolved oxygen as described above.

Calculations

Calculate the actual volume of the water used in the titrations. This is normally 300 mL minus 8 mL of reagents added before titration. If exactly 292 mL of sample was used, then 1 mL of 0.020 M sodium thiosulfate solution indicates the presence of 0.547 mg per litre of dissolved oxygen. Calculate the average volume of titre ($S_2O_3^{2-}$) used, and hence calculate the concentration of dissolved oxygen in mg L^{-1}.

Calculate the biological oxygen demand, BOD:

$$\text{BOD} = \text{concentration of dissolved oxygen (mg L}^{-1}\text{)} - \text{concentration of BOD sample (mg L}^{-1}\text{)}$$

You may also calculate the percentage saturation of oxygen in the sample by referring to table wq.1, which gives the amount of dissolved oxygen at various temperatures.

materials

- 5 × 300 mL airtight reagent jars
- thermometer
- 5 mL graduated pipette or graduated syringe
- 10 mL 2.15 M manganese sulfate solution ($MnSO_4$, dissolve 48.0g $MnSO_4.4H_2O$ in 100 mL water)
- 10 mL alkaline iodide solution (dissolve 50 g solid NaOH and 13.5 g NaI in 100 mL water)
- 50 mL 9 M sulfuric acid solution (H_2SO_4)
- 3 × 500 mL conical flasks
- 2 mL starch indicator solution (dissolve 0.5 g starch in 100 mL boiling water)
- 25 mL burette
- 100 mL 0.020 M sodium thiosulfate solution ($Na_2S_2O_3$)

safety

- Wear safety glasses and a laboratory coat for this experiment.
- Sulfuric acid is corrosive. Avoid contact with skin and eyes.

table wq. 1 The saturated concentration of oxygen in water at particular temperatures.

Temperature (°C)	Saturated concentration (mg L^{-1})
5	12.8
10	11.3
15	10.1
20	9.2
25	8.3

$$\% \text{ saturation} = \frac{\text{dissolved oxygen in sample}}{\text{maximum amount of oxygen}} \times 100$$

figure wq.5 Algal blooms cause a depletion of oxygen in water. When the large masses of algae die, bacteria break down the dead algae during the decay process. This uses up the oxygen in the water and other organisms die.

Experiment wq.3

Quantitative determination of water hardness

materials

- 3 × 250 mL conical flasks
- 5 mL graduated pipette
- 3 mL pH 10 buffer (57 mL conc. NH_3 and 7.0 g NH_4Cl in 100 mL water)
- Eriochrome Black T indicator (dissolve 0.2 g powder in 15 mL commercial NH_3 and add 5 mL absolute ethanol)
- 15 mL 0.01 M EDTA (disodium salt solution, 1.861 g L^{-1})
- 10 or 25 mL burette
- 3 mL 0.01 M magnesium chloride solution ($MgCl_2$)

Method

Add 100 mL of your water sample to a conical flask. Add 1 mL buffer (pH 10) solution and two or three drops of Eriochrome Black T indicator. The solution becomes a red–violet colour. Fill a small burette with EDTA solution and slowly titrate, continually swirling the flask, until a permanent blue colour appears. If the colour change is not definite, magnesium ions are not present. Repeat these steps with another 100 mL sample, but add 1 mL magnesium chloride solution ($MgCl_2$) before titration. Record the reading on the burette. Repeat the titration with two more portions of the same sample of water.

Calculations

1. Calculate the average volume of EDTA added.
2. Since 1 mL EDTA indicates the presence of 1 mg $CaCO_3$, calculate the concentration of metal ions (Ca^{2+} and Mg^{2+}) in your 100 mL sample.
3. If you added 1 mL magnesium chloride solution, subtract this amount from the value calculated in step 2.

figure wq.6 Hard water is generally a result of the geology of the local area rather than pollutants. Magnesium and calcium ions in water react with soaps and detergents producing the scum that forms in a bath or shower. They also mean that greater amounts of detergents are needed to wash clothes. The build up of magnesium and calcium compounds as scale in water pipes can also cause problems for local councils.

Experiment wq.4

Determination of total solids in water

materials

- jar
- filter paper
- balance (accurate to 0.001 g)
- large evaporating basin
- filter funnel
- glass rod
- drying oven
- Bunsen burner
- heatproof mat
- 100 mL measuring cylinder

Method

In the field, test the water qualitatively for turbidity and for the presence of heavy metals and chlorides. If any of these are much higher than in drinking water, proceed with the test indicated below. If these tests do not indicate a high presence of dissolved or visible solids, test the water with the conductivity probe. (This may be part of the StreamWatch kit. A datalogger with conductivity probe is also suitable to measure the total dissolved solids, or TDS.)

1. Collect exactly 100 mL water in a jar.
2. In the laboratory, accurately weigh a dry filter paper. Weigh a large evaporating basin and position it as a receiver under the filter funnel as shown in figure wq.7.
3. Shake your water sample in the collecting jar and pour it through the filter paper over a glass rod.
4. Wash the jar several times with distilled water, adding each wash to the filter paper.
5. When filtration is complete, remove the filter paper and place it in a drying oven set to about 60°C. Leave the filter paper to dry overnight.
6. Place the evaporating basin on a tripod and gently heat over a Bunsen flame until there is only about 5 mL water left.
7. Transfer the evaporating basin into a drying oven and let it dry overnight.
8. Reweigh the filter paper and evaporating basin. Record all measurements.

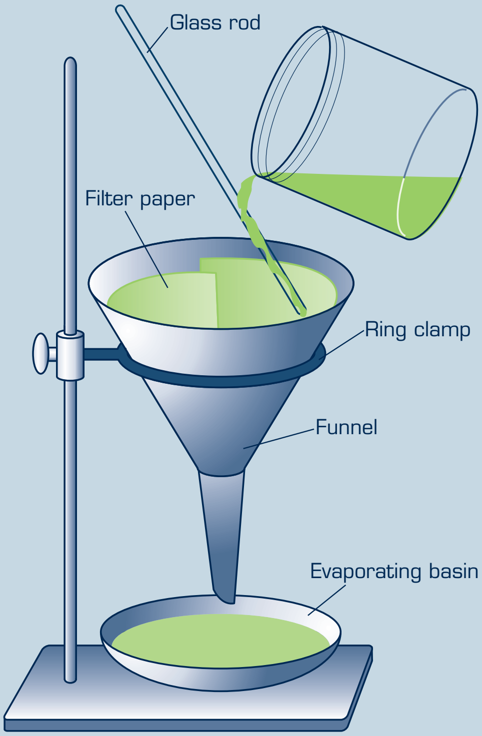


figure wq.7 Experimental set-up.

Calculations

1. By appropriate subtraction, find the mass of:
   - **a** solids left on the filter paper (g per 100 mL)
   - **b** solids left in the evaporating basin (g per 100 mL).
2. Pollutants are normally reported in mg L^{-1} or parts per trillion.

   Calculate the amount of solids in your water sample in mg L^{-1} and record as:
   - **a** suspended solids (from the filter paper)
   - **b** dissolved solids (from evaporating basin)
   - **c** total solids (dissolved + suspended).

Experiment wq.5

Quantitative analysis of chloride ions

This test is suitable for water that shows a strong presence of chloride ions when tested with 0.1 M silver nitrate solution ($AgNO_3$).

materials

- 3 × 500 mL conical flasks
- 1 mL phenolphthalein indicator
- 1 mL 0.1 M sodium carbonate solution (Na_2CO_3)
- 1 mL 0.1 M ethanoic acid solution (CH_3COOH)
- 25 mL burette
- 50 mL standard 0.1 M silver nitrate solution ($AgNO_3$)
- 15 mL 0.1 M potassium chromate solution (K_2CrO_4)

figure wq.8 Salt-affected land is most often the result of land clearing and irrigation. These cause the groundwater level to rise, bringing dissolved salt to the surface, which kills salt-sensitive vegetation.

Method

1. Transfer 250 mL of your water sample into each of three conical flasks. If the pH of the original water sample is less than 8, use the following pH adjustment procedure.
2. To one water sample add two drops of phenolphthalein indicator. The solution should be colourless. Add 0.1 M sodium carbonate solution (Na_2CO_3), dropwise while stirring, until the water turns a faint pink colour.
3. Repeat this pH adjustment for the other two samples.
4. If the pH of the original sample is above 8.2, add 2 drops of phenolphthalein indicator to one water sample. The solution turns pink. Add 0.1 M ethanoic acid solution, dropwise while stirring, until the pink colour disappears. Repeat for the other two samples.
5. Fill a clean burette with silver nitrate solution ($AgNO_3$). One fill is enough for all titrations.
6. Add 5 mL potassium chromate solution (K_2CrO_4) to the water sample in one conical flask.
7. Slowly add silver nitrate solution from your burette, while gently swirling the flask. The end point is reached when the first permanent reddish-brown colour appears.
8. Repeat the titration with the other two water samples.

Calculations

1. Calculate the average volume of silver nitrate solution added.
2. Calculate the moles of silver nitrate needed to react with the chloride ions using the relationship:

   amount of solute = concentration × volume of solution

   $$n = c \times V$$

   (This is equivalent to the number of moles of Cl^- ions.)
3. Calculate the concentration of Cl^- ions in your 250 mL sample in mol L^{-1}. The concentration of chloride ions is usually reported in parts per million or mg L^{-1}. Calculate the concentration of Cl^- ions in mg L^{-1} (the molar mass of chloride ions is 35.45 g mol^{-1}).

Experimental investigations

You may be concerned about the health of your local waterways, especially if some uses of the land adjoining these waterways have the potential to pollute or degrade them. Select your site or sites of testing carefully. You can compare water quality upstream and downstream of an area used for activities such as residential, agricultural, industrial, a sewage treatment plant or a factory. You may look at the effect of land clearing or a bushfire on a waterway.

Keep in mind the uses of the water course. Is it part of the catchment of your town water supply? Do people use it for swimming or to water crops? Is it a breeding habitat for fish and other aquatic creatures eaten by humans?

Water from the selected sites might be tested regularly over a period of weeks or months. In this way you may see if the waterway is improving or getting worse. You may be able to identify a particular incident that has brought about a change in water quality.

Reporting your results

Analyse the results of your tests and comment on whether or not the results are acceptable in terms of odour, pH, turbidity, levels of heavy metals, BOD, hardness, total solids and chloride ion concentration.

Which of the poor results could be caused by pollutants? Which pollutants? Where did they come from?

Which of the poor results are caused by inappropriate land use, or degradation or oversuse of the land?

What possible harmful effects could result from pollutants or poor water qualities indicated by your testing?

What tests indicate acceptable or good quality of water at your selected sites?

What policies or practices regarding the waterway have brought about good water testing results?

Devise a system and rate the quality of the water at your chosen sites from very poor to very good. Make recommendations for continued good practice, and for changes or rehabilitation practices that should be applied to ensure that the waterway is healthy.