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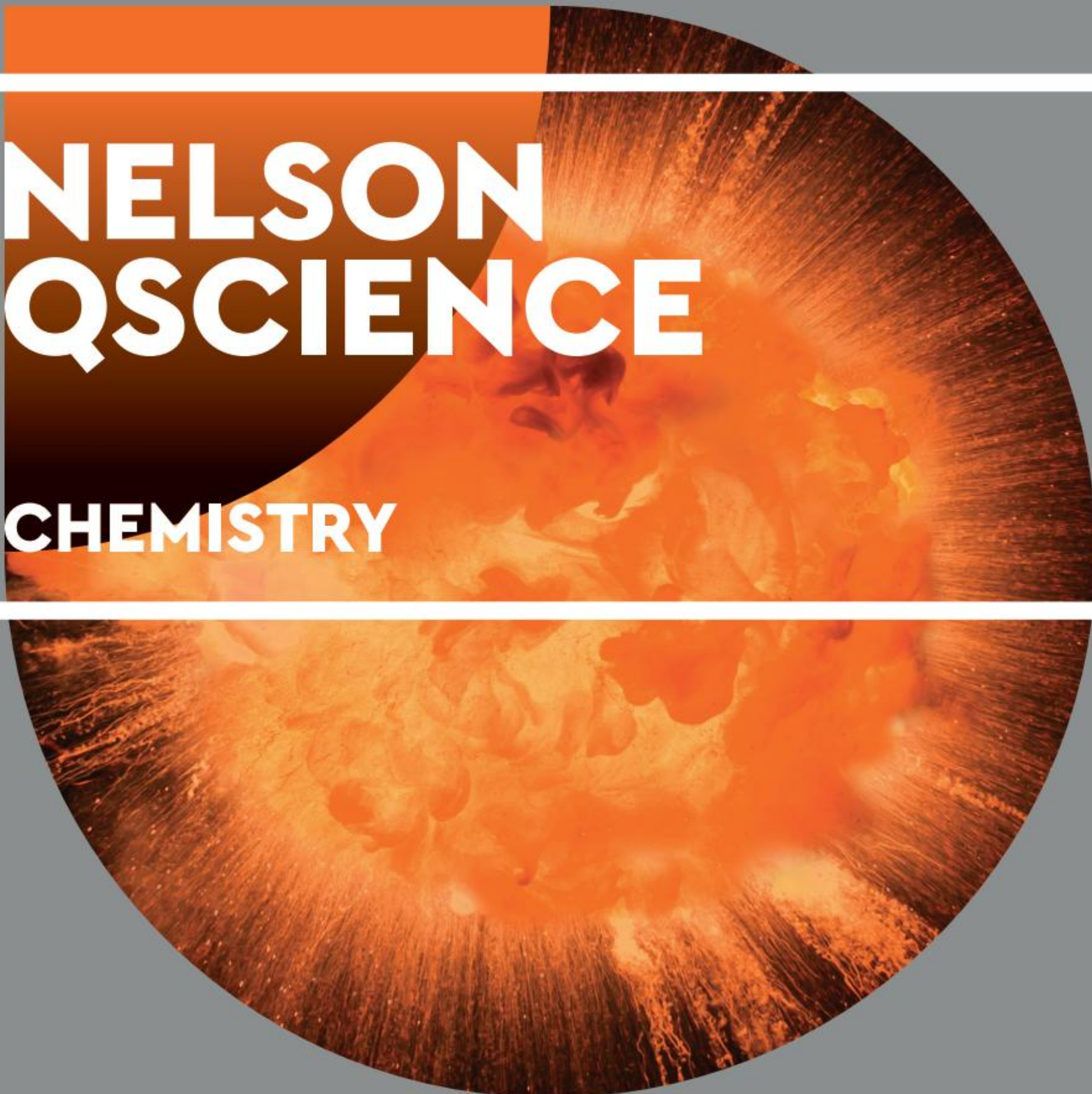
CHEMISTRY



UNITS 1 2

Nicholas Stansbie
Professor Alan E.W. Knight
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Dr Sarah Windsor

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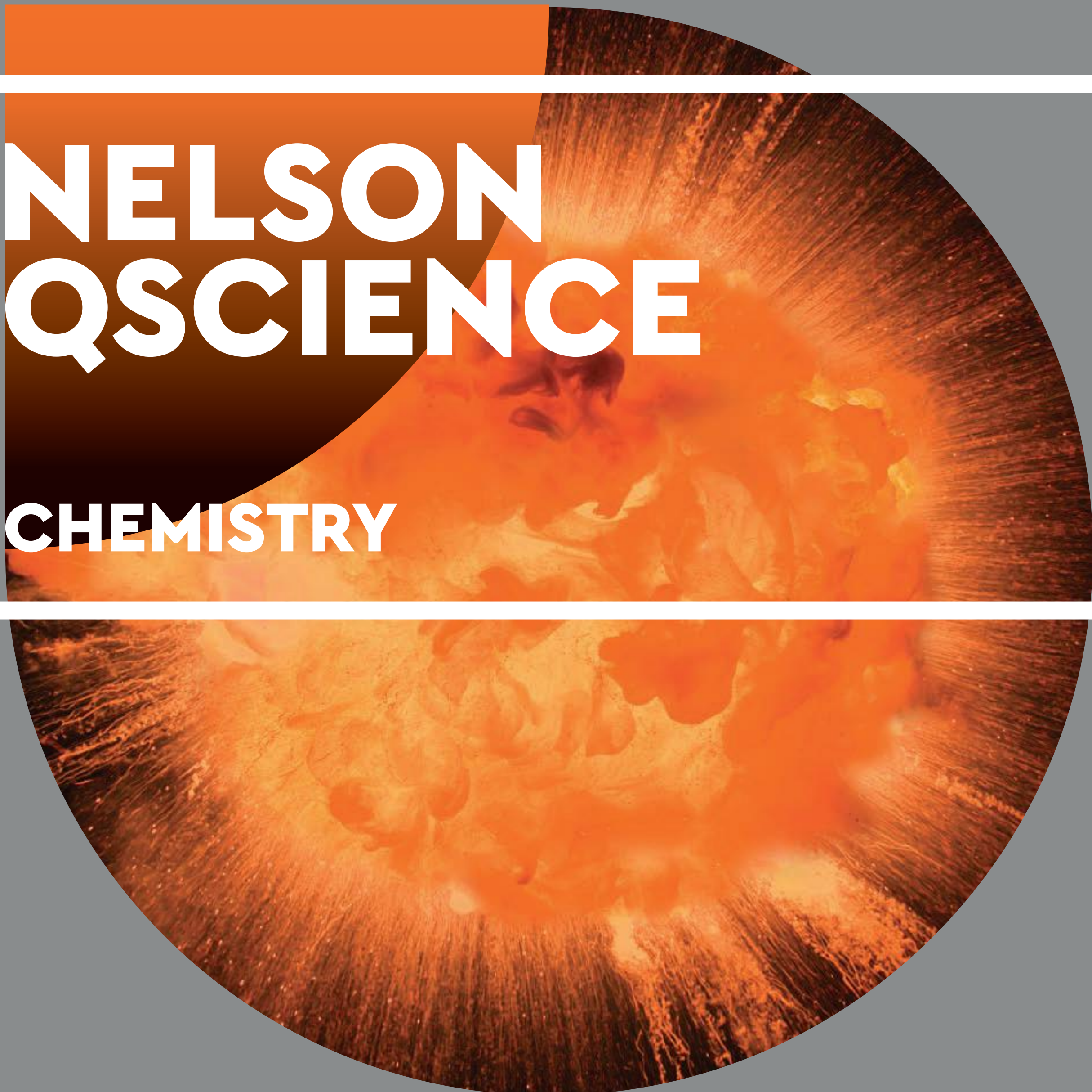

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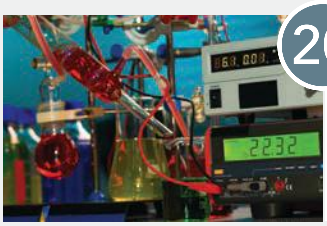

TOPIC 3: RATES OF CHEMICAL REACTIONS

PREFACE

Nelson QScience Chemistry Units 1 & 2 has been written to meet the requirements of the QCAA Senior Secondary Science Syllabus – Chemistry. Each page has been carefully considered to provide students with all of the information they need to meet the content and skills requirements of the new syllabus.

With the introduction of the QCE external examination, *Nelson QScience Chemistry* includes features such as practice exams at the end of each section, a Units 1 & 2 practice exam, chapter quizzes (available on NelsonNet) and ExamView (available on NelsonNet).

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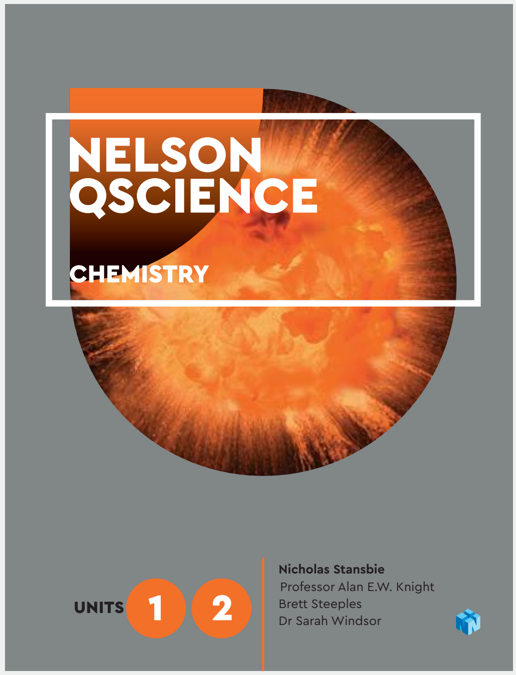

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Nicholas is an experienced science teacher and teacher of senior chemistry in the UK and in Australia. Nicholas' dedication to quality teaching and learning has enabled him to lead the *QScience Chemistry* team in the development of this text.

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SYLLABUS REFERENCE GRID

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ABOUT THIS BOOK

At the beginning of unit and topic

- Unit introductions are an overview of the key content in the unit.
- Topic introductions are an overview of the key content in the topic.

At the beginning of each chapter

- A short chapter summary introduces students to the key content and skills covered.
- Stimulus questions can be used to focus on the topic and highlight existing knowledge.

In each chapter

- **Worked examples** are explained clearly step-by-step.
- **Key formulas** are highlighted in the margin.
- **Key glossary terms** are highlighted in the margin.
- **Science as a Human Endeavour** provides opportunities for students to connect to the importance of SHE and develop scientific research skills.
- **Inquiring further** provides opportunities for students to further investigate scientific concepts and develop scientific research skills.
- **Section reviews** are written in the style of Bloom's revised taxonomy.
- **Practical experiments** contain guided instructions on the materials, procedure, collection and analysis of the results and discussion.

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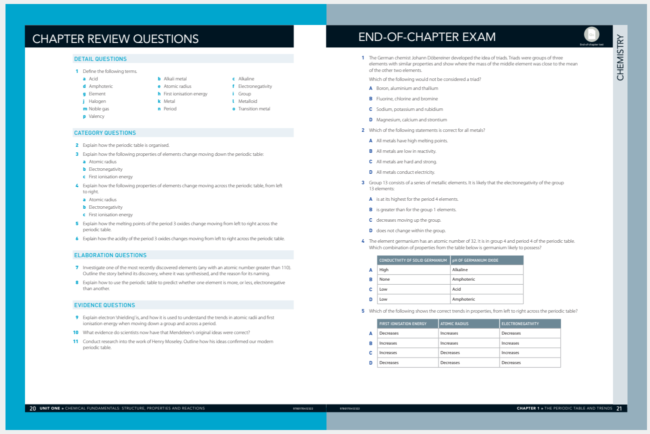
At the end of each chapter

- **Chapter review questions** written in the style of Marzano and Simms (2014) questioning sequences.
- **Practice examinations** occur at the end of each chapter to help students develop skills in decoding and answering exam-style questions.



At the end of the book

- **Practice exam** questions provide an extended practice of the content and skills learnt across the text.
- **Glossary** provides explanations of all of the new terms introduced in the text.
- **Answers** provide complete answers for student reference.

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» UNIT ONE

CHEMICAL FUNDAMENTALS: STRUCTURE, PROPERTIES AND REACTIONS

- Topic 1: Properties and structure of atoms
- Topic 2: Properties and structure of materials
- Topic 3: Chemical reactions: reactants, products and energy change

Chemical reactions occur when bonds are broken between the reactant atoms and molecules and new bonds are formed to produce the products. How this occurs depends on the structure of the reactants and products and the bonding that exists within them. Bonding depends on the arrangement of subatomic particles within atoms. The periodic table is a powerful tool that helps us to understand the patterns and trends observed in the properties of elements and to explain them in terms of atomic structure. Concepts such as 'the mole' enable chemists to measure the quantities of particles involved in a chemical reaction accurately, and to predict the amounts of products that will form.

UNIT OBJECTIVES

By the end of this unit, students should be able to:

1. describe and explain the properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change
2. apply understanding of the properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change
3. analyse evidence about the properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy
4. interpret evidence about the properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change
5. investigate phenomena associated with properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change
6. evaluate processes, claims and conclusions about the properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change
7. communicate understandings, findings, arguments and conclusions about the properties and structure of atoms and materials, and chemical reactions in terms of reactants, products and energy change.

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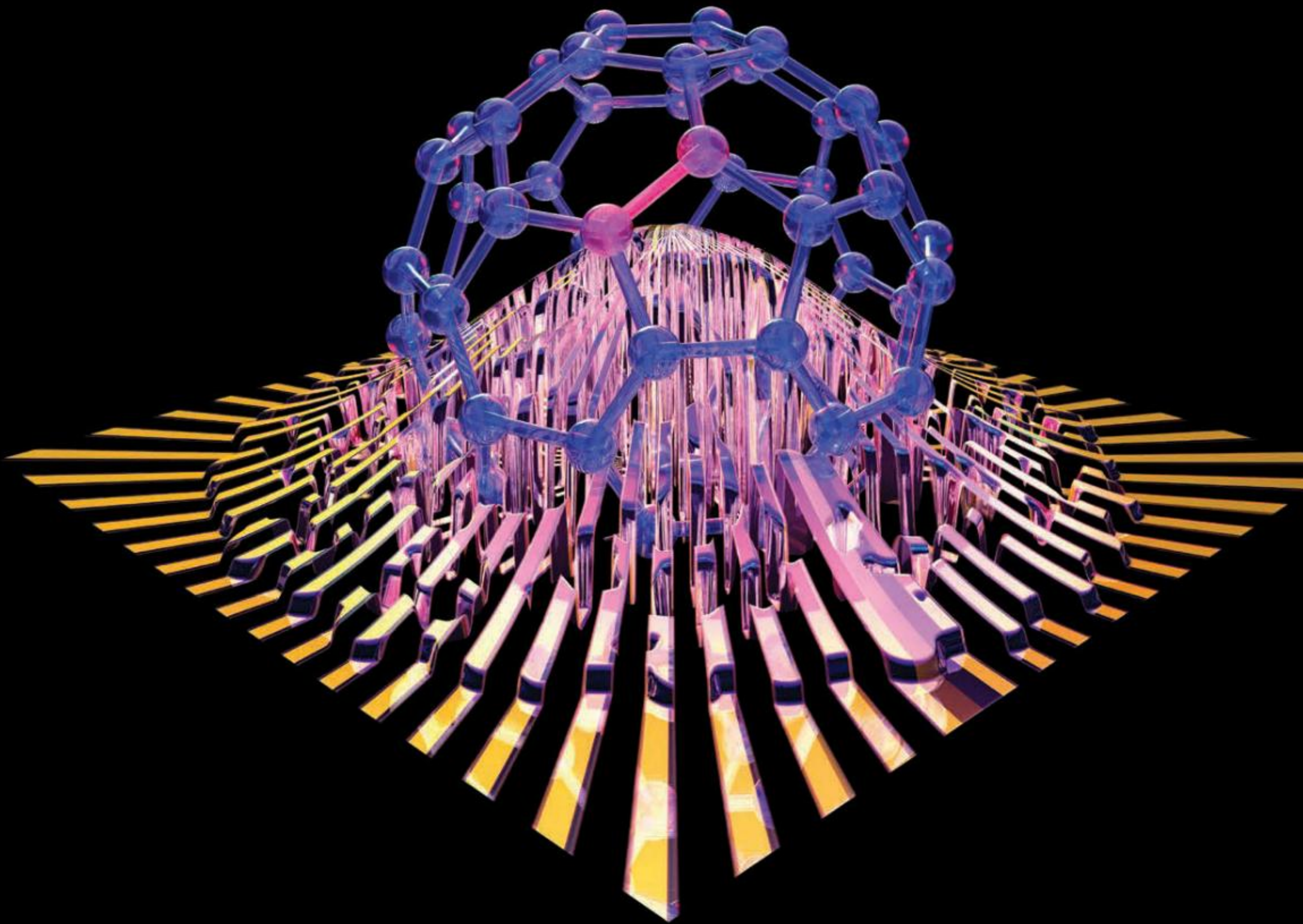

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» UNIT ONE

CHEMICAL FUNDAMENTALS: STRUCTURE, PROPERTIES AND REACTIONS

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Topic 1: Properties and structure of atoms

Understanding the structure of the atom is fundamental to understanding how materials behave and can be used. This topic outlines the relationship between the structure of an atom and the position of that element on the periodic table, and explains some of the trends and patterns that occur. It considers how different atoms bond together to form new substances, and how analytical techniques such as mass spectrometry, emission spectroscopy and atomic absorption spectroscopy can be used to determine the structure and bonding within a substance

SCIENCE AS A HUMAN ENDEAVOUR

Students should be given opportunities to investigate: models of the atom and the history of the development of scientific models, evaluate the use of radioistopes and investigate the distribution of elements.

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1 THE PERIODIC TABLE AND TRENDS

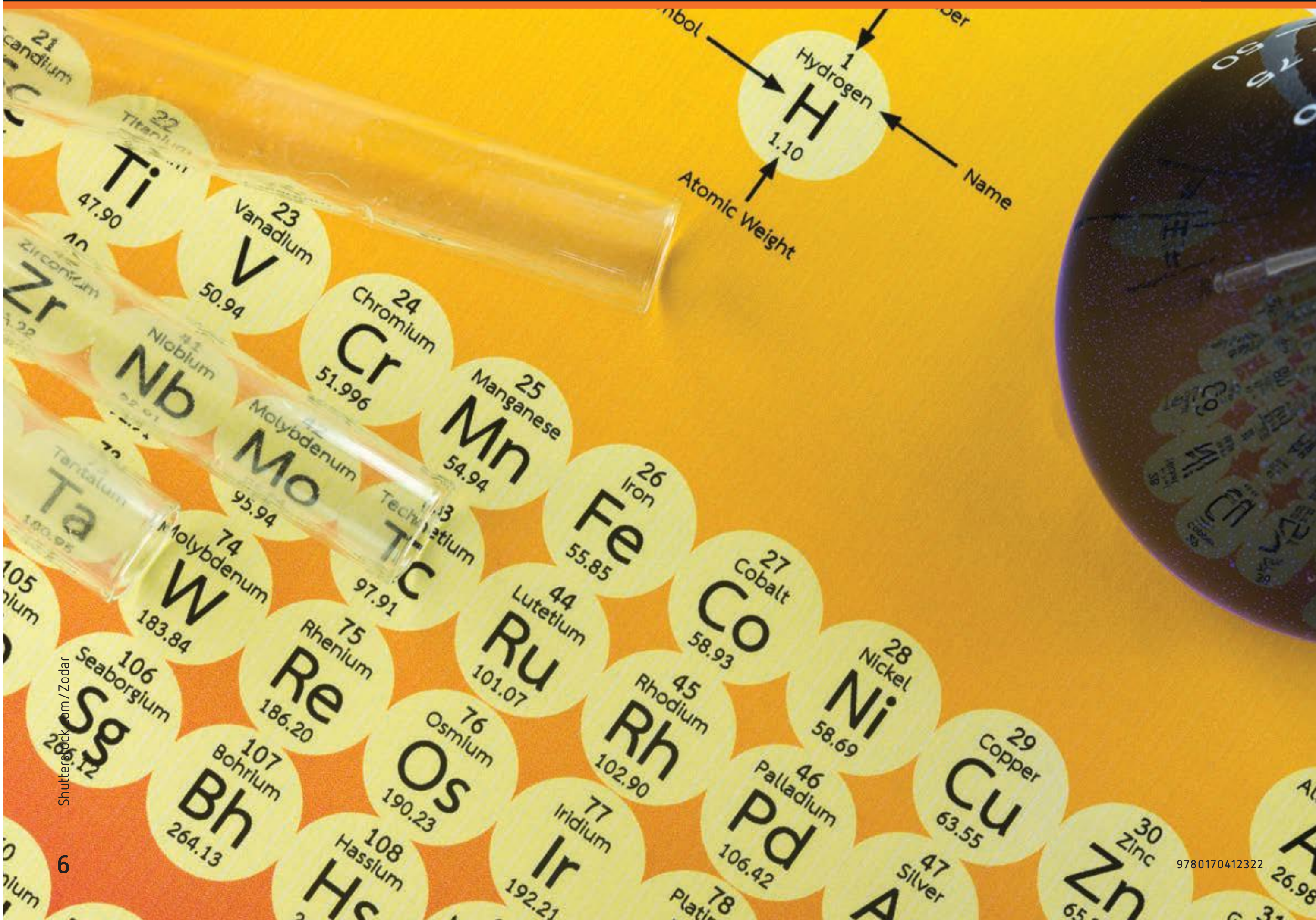
Introduction

The periodic table is the chemist's method of arranging all the elements in a systematic way so that regularly repeating or 'periodic' trends and their properties can be seen. This means that the properties of an individual element can be easily identified from its position in the table.

Stimulus questions

Why was Dmitri Mendeleev's work such an important breakthrough?

How does the periodic table help us to compare and explain the properties of elements?



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1.1 The periodic table of elements

Elements

Many elements, such as gold, silver, copper, iron, tin and lead, were discovered and used in ancient times. During the time of the alchemists, mercury, sulfur, carbon, arsenic, antimony, phosphorus and zinc were all added to the list of known elements. However, the modern scientific concept of an **element** was not devised until a German scientist, Hennig Brand, isolated phosphorus from urine, keeping the process secret, but selling it to other alchemists. Alchemists, as well as the new type of scientists following in the footsteps of Robert Boyle, subsequently identified many more elements, with almost all of the naturally occurring elements becoming known by 1900.

element
a substance consisting of only one type of atom

An element is a substance consisting of only a single type of atom. There are 118 known elements, from which all other substances and materials are derived.

Elements can be represented by symbols. Some elements are represented by the first letter of their names; for example, hydrogen is 'H' and carbon is 'C'. Some other elements have two letters; for example, chlorine is 'Cl' and aluminium is 'Al'. The first letter of an element's symbol is always a capital letter. The second letter is always lower case. Some elements are represented by letters from their names originating in other languages such as Latin (e.g. gold (Au) from *aurum*, meaning 'gold'; potassium (K) from *kalium*, meaning 'potash') and Greek (e.g. chlorine (Cl) from chlorós, meaning 'pale green').

Element names are determined by **the International Union of Pure and Applied Chemistry (IUPAC)**. Many of the elements have been known for a long time, but some elements have only recently been synthesised in laboratories. IUPAC approves names suggested by scientists, companies or universities that are instrumental in the discovery of the element. For example, element 116 was named livermorium (Lv) in 2012 after the Lawrence Livermore National Laboratory in California.

IUPAC
the International Union of Pure and Applied Chemistry

The periodic table of elements

The **periodic table of elements**, commonly known as the 'periodic table', is the chemist's method of arranging all the elements in a systematic way so that trends and patterns relating to their individual properties can be seen.

periodic table of elements
a table listing the known elements in order of atomic number, based on their chemical and structural properties, so that patterns and trends are easily recognisable

For many centuries, scientists had been trying to make sense of the elements and their properties. The first system, Johann Döbereiner's theory of 'triads' was published in 1829. He noticed that elements that had similar properties could be grouped into threes, which he called triads. Döbereiner noted that the groups of three elements had similar properties and their atomic masses showed a specific pattern. For example, calcium, strontium and barium were one triad. They all have similar chemical properties and the atomic mass of strontium is almost exactly halfway between that of calcium and barium. Döbereiner noticed the same relationship between lithium, sodium and potassium, so he called them another triad. Other triads were phosphorus, arsenic and antimony, and chlorine, bromine and iodine.

The periodic table was developed by the Russian chemist Dmitri Mendeleev and published in 1869. Mendeleev wrote out the names of the elements with their atomic masses and other properties. He laid them out in rows (periods) and columns (groups) by first ordering the elements by mass, and then grouping them so that elements with similar properties were in the same vertical column. Mendeleev left gaps for elements not discovered at the time such as gallium, germanium and scandium and could predict their mass and properties from the trends shown in his periodic table. These elements have been subsequently discovered and have the properties that Mendeleev predicted.

1.1.1 RSC: Periodic table
1.1.2 Solving the puzzle of the periodic table
1.1.3 The genius of Mendeleev's periodic table

SCIENCE AS A HUMAN ENDEAVOUR

I saw in a dream a table where all elements fell into place as required. Awakening, I immediately wrote it down on a piece of paper, only in one place did a correction later seem necessary. —Dmitri Mendeleev

Conduct research to find out about one of the elements that Mendeleev predicted was still to be discovered when he created the periodic table in 1869. Write a short discussion about how accurate his prediction was.

period
elements contained in a single horizontal row in the periodic table

group
elements contained in a single vertical column in the periodic table

atomic number
the number of protons contained within one atom of an element – each element has a unique atomic number

Arrangement of the periodic table

The periodic table is arranged in a series of horizontal rows called **periods** and vertical columns called **groups**. The elements are listed in the periodic table in order of **atomic number**, from element 1, hydrogen (H), to element 118, oganesson (Og). A simple periodic table showing groups and periods is seen in Figure 1.1.1.

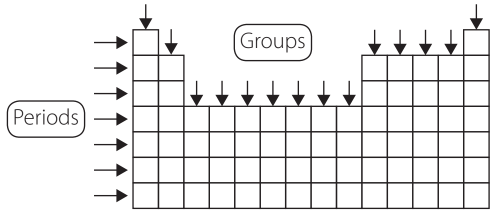

FIGURE 1.1.1 Groups and periods on the periodic table

Groups in the periodic table

Elements with similar properties are found in vertical groups numbered 1–18. For example, group 2 includes the elements beryllium, magnesium, calcium and strontium, while group 15 includes nitrogen, phosphorus and arsenic. Some of these groups have names (see Figure 1.1.2). Group 1 is often referred to as the alkali metals and includes lithium, sodium, potassium and rubidium. In the middle of the periodic table are the transition metals in groups 3–12.

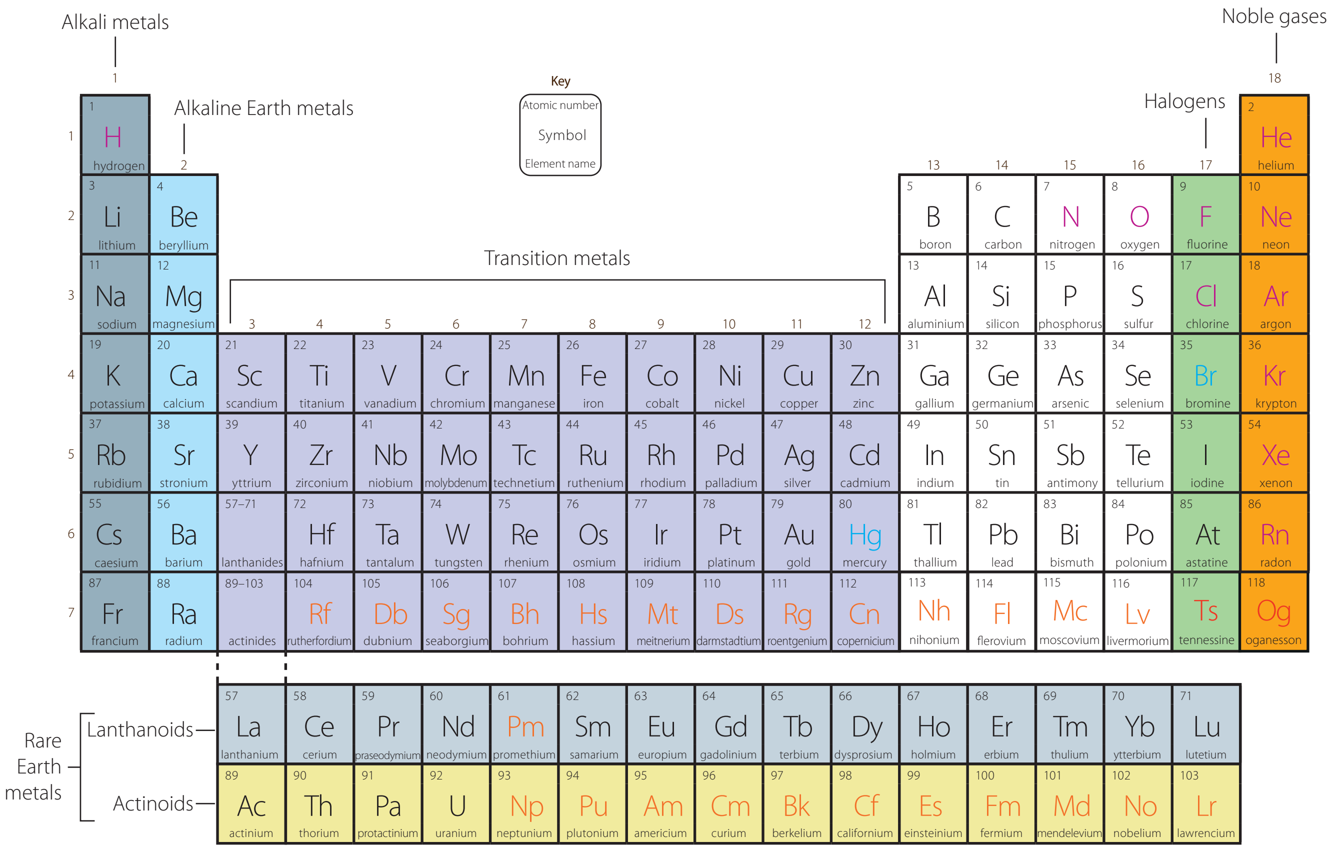

FIGURE 1.1.2 The modern periodic table

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SECTION REVIEW 1.1

REMEMBERING

1 Explain what the term 'element' means.

2 Describe how Mendeleev arranged the elements in the periodic table.

3 Distinguish between a group and a period on the periodic table.

4 Identify the groups that have been given common names, such as the noble gases. Make a table showing the group number and the common name of these groups.

UNDERSTANDING

5 Explain why some elements, such as carbon (C) and bromine (Br), have a symbol that matches their name, yet for other elements, such as mercury (Hg) and tungsten (W), the symbol does not appear to match their name.

APPLYING

6 A compound is a substance made from more than one element. A compound found in bleach has the formula NaClO. Interpret this formula to identify how many different elements it contains, and what their names are.

7 Choose two more groups of elements in the periodic table that Döbereiner could have identified as 'triads'.

ANALYSING

8 If beryllium reacts with fluorine gas to make beryllium fluoride, predict what would happen if you placed some magnesium metal in a gas jar containing chlorine gas. Explain your answer.

9 Argon does not react readily with any other element. (It is called an inert gas.) Which other elements could be called 'inert gases'?

10 Silicon is a well-known 'semiconductor' in electronics. What other element is also likely to display semiconductor properties?

1.2 Trends in the periodic table

The periodic table is a powerful tool because it reveals specific patterns in the properties of the elements. This means that scientists can gain a lot of information about an individual element and how it behaves simply from its position in the periodic table.

atomic radius
a measure of the size of an atom – measured as the distance from the nucleus to the boundary of the cloud of electrons surrounding it – the distances are very small and are measured in picometres (pm) (1 pm = 1×10^{-12} m)

Atomic radius

Atomic radius is a measure of the size of an atom, defined as the distance from the nucleus to the boundary of the cloud of electrons surrounding it (Figure 1.2.1). The boundary is not a well-defined physical entity. Rather, the electron cloud must be described as a probability distribution, which tapers off gradually as one moves away from the nucleus without a sharp cutoff. Even so, the boundary can be defined as the '90% probability sphere', which corresponds to the sphere within which there is a 90% probability of finding the electron.

An estimate of atomic radius can be obtained by taking half the experimentally measured distance between the nuclei of two neighbouring atoms of the same element, say metal atoms, in a metallic solid.

The concept of two 'neighbouring' atoms requires qualification. When the two atoms are chemically bonded together, scientists refer to the radius as the covalent radius. If the atoms are metals, existing together in a solid lattice, scientists term the 'half distance' as the metallic radius. Finally, for atoms that exist as individual particles that contact each other without bonding, the radius is called the 'van der Waals radius', a term that comes from the naming of the weak intermolecular forces between the atoms.

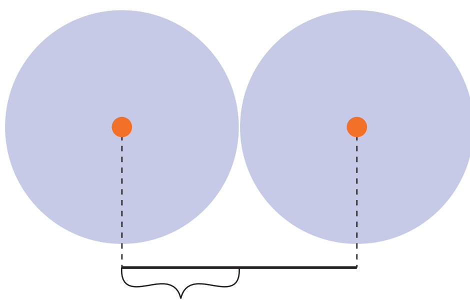

FIGURE 1.2.1 Atomic radius

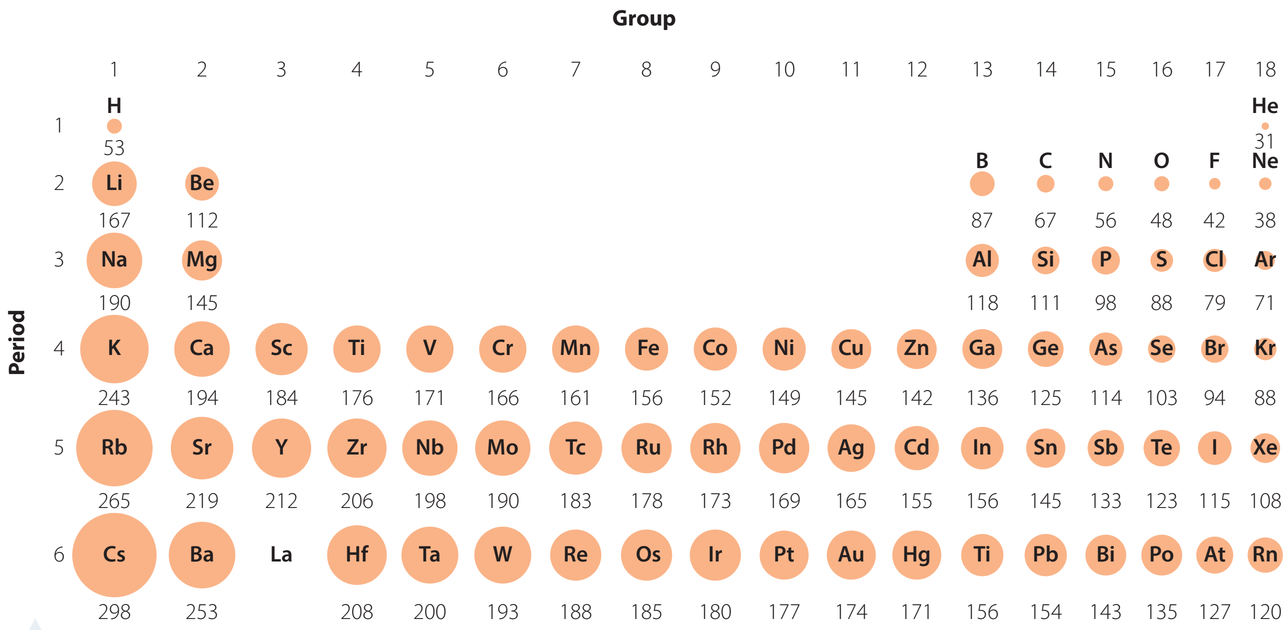

FIGURE 1.2.2 Trends in atomic radii (picometres) in the periodic table atomic radius decreases across a period and increases down a group. Note that no data is available for the calculated atomic radius of lanthanum.

The trends in atomic radii observed with respect to the periodic table remain consistent regardless of which measurement is used. The trends are illustrated in Figure 1.2.2.

- Atomic radius decreases from left to right across any period of the table. The positive charge of the **nucleus** increases across the table, due to the extra protons in the nucleus. As the nucleus becomes more positive, the electrons in the outer energy level are more strongly attracted to the positive nucleus and they move closer together, making the atom smaller.
- Atomic radius increases from top to bottom in any group. That is, the elements at the *bottom* of a group have more **electron shells** filled than those at the *top* of a group. As more energy levels are filled, the atomic radius increases.

nucleus
the positively charged centre of an atom, containing positively charged protons and electrically neutral neutrons, to which the orbiting electrons are attracted

electron shells
electrons exist in orbit around an atom in energy levels called shells; the greater the energy level, the further away the shell is from the nucleus

valence shell
the outermost electron energy level, or shell, in an atom

valency
the number of hydrogen atoms with which a single atom can bond when forming a compound (e.g. hydrogen has a valency of 1)

Valency

The number of electrons in the **valence shell** (valence electrons) of the elements in groups 1, 2 and 13–18 can be determined from the periodic table. The number of valence electrons also determines the **valency** of each element. Valency is an atom's ability to form chemical bonds with another atom. It is a number referring to the combining power of an element when compounds form. Table 1.2.1 shows the relationship between the group number, number of electrons and valency. Once the valency is known, the number of bonds an atom can form can be used to determine a ratio for combining different elements.

TABLE 1.2.1 How valency is related to groups in the periodic table

GROUP NUMBER	1	2	3–12	13	14	15	16	17	18
NUMBER OF VALENCE ELECTRONS	1	2	Variable	3	4	5	6	7	8
VALENCY	1	2	Variable	3	4	3	2	1	0

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Calculating the valency of atoms

Some simple rules for calculating the valency of atoms are as follows.

1. Refer to the periodic table and determine the number of valence electrons for the atom.
2. If the number of valence electrons is four or fewer, the valence is equal to the number of electrons. For example, magnesium has two valence electrons so magnesium has a valency of 2. Carbon has four valence electrons so it has a valency of 4.
3. If the number of valence electrons is greater than four, then the valency is calculated by subtracting the number of valence electrons from eight. For example, nitrogen has five valence electrons, so it has a valency of 3 (8 – 5). Chlorine has seven valence electrons so it has a valency of 1 (8 – 7).

Table 1.2.1 shows the relationship between the group number, number of electrons and valency. Once the valency is known, then the number of bonds an atom can form can be used to determine a ratio for combining different elements.

SYLLABUS LINK

Chapters 2 and 3 discuss the electron configuration of atoms in greater detail.

Ionic radius

To form ions, some atoms gain electrons, while other atoms lose electrons. When an atom gains or loses an electron, there is an imbalance in the amount of positively charged protons and negatively charged electrons. As such, the attractive forces between the nucleus and the electrons change, so the radius of the ion is different to that of the atom. Atoms in different groups form ions in different ways, so there is no clearly observable trend in the way that ionic radius changes across a period. However, within every group, the **ionic radius** increases down the group, as the number of electron shells increases.

ionic radius
the distance between the centre of the nucleus and the outermost electron shell in an ion

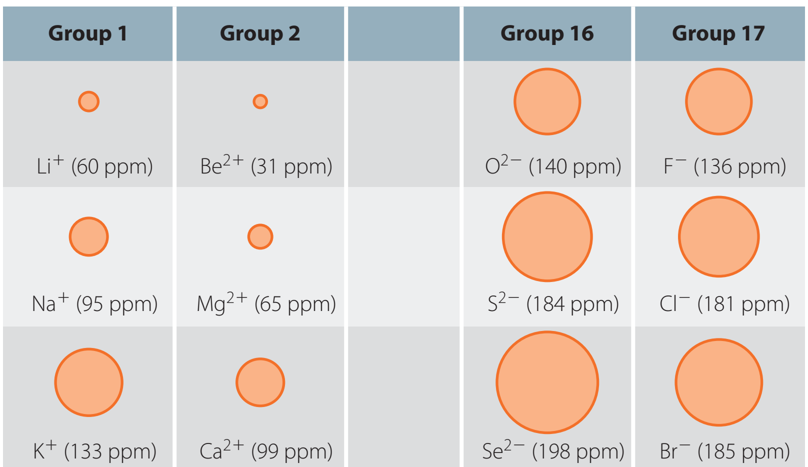


FIGURE 1.2.3 Ionic radii in groups 1, 2, 16 and 17, showing that ionic radius increases down each group. (ppm = parts per million)

First ionisation energy

The **first ionisation energy** is the amount of energy needed to completely remove an electron from a neutral atom when it is a gas. An atom that has a low first ionisation energy will become an **ion** (a charged atom) very easily.

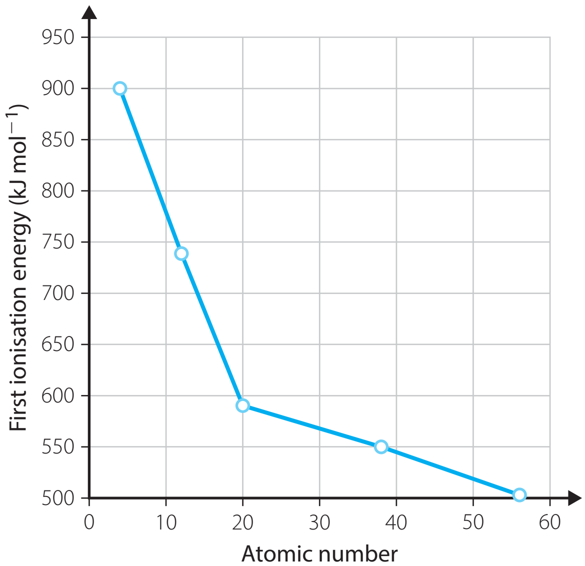
- First ionisation energy decreases down a group. Down a group, atoms in successive periods have an additional shell of electrons. The electrons in shells further away from the nucleus are there because they are less strongly bound to the positive nucleus. This means it takes less energy to remove an outer shell electron from the atom because there is a weaker attraction between the positive nucleus and the outermost negative electrons. The 'shielding effect', in which inner electrons 'shield' outer electrons from the attraction of the nucleus (lowering their ionisation energy in the process), also

first ionisation energy
the amount of energy required to remove a single electron from an atom in the gaseous form

ion
an atom that has become charged through the gain or loss of an electron

contributes to the weaker attraction of the outer electrons to the nucleus. This trend is observed among group 2 elements in Figure 1.2.4.

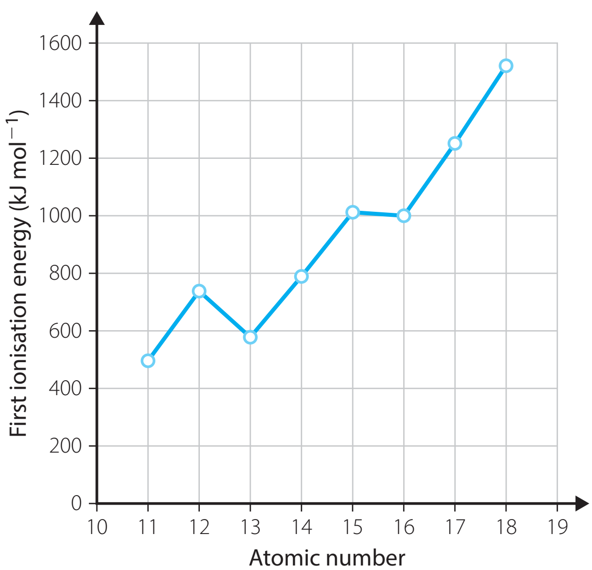
FIGURE 1.2.4 The first ionisation energies of the group 2 elements



ELEMENT	PROTON NUMBER	FIRST IONISATION ENERGY (kJ mol^{-1})
Beryllium	4	900
Magnesium	12	738
Calcium	20	590
Strontium	38	550
Barium	56	503

- First ionisation energy increases from left to right across a period. Within a period, the valence electrons of each atom are in the same shell so are all approximately the same distance from the positively charged nucleus. But as you go across a period, the number of protons increases, so the positive charge of the nucleus also increases. The attraction between the positive nucleus and the outermost negative electrons (in the same valence shell) becomes stronger. This makes it harder to remove an outer electron, so the first ionisation energy increases across a period. Figure 1.2.5 shows this trend for period 3. For example, the first ionisation energy is greater for magnesium than for sodium. Likewise, the first ionisation energy is greater for argon than for chlorine. The trend does display some minor exceptions. There are two relatively small decreases in ionisation energy from magnesium to aluminium and from phosphorus to sulfur. These apparent anomalies can be explained once the electron configuration of these elements is understood in terms of orbitals.

FIGURE 1.2.5 The first ionisation energies of the elements in period 3



ELEMENT	ATOMIC NUMBER	FIRST IONISATION ENERGY (kJ mol^{-1})
Sodium	11	496
Magnesium	12	738
Aluminium	13	578
Silicon	14	789
Phosphorus	15	1012
Sulfur	16	1000
Chlorine	17	1251
Argon	18	1521

Successive ionisation energies

The first ionisation energy is the energy required in order for the reaction in the equation shown below to take place, where the highest energy electron is removed from an atom (A), to form a positive ion (A^+):

$$A \rightarrow A^+ + \text{electron}$$

The energy required to remove a second electron and form A^{2+} is known as the second ionisation energy:

$$A^+ \rightarrow A^{2+} + \text{electron}$$

If this trend continued, then successive electrons could continue to be removed from the atom, giving rise to the term **successive ionisation energies**.

As previously discussed, the magnitude of the ionisation energy depends on the extent of the attraction between the nucleus and the electron being removed. With each additional electron that is removed, the ion becomes more positively charged; the attraction to the electrons that remain increases. This means that successive ionisation energies keep increasing.

Importantly, the size of the increase between successive ionisation energies provides the crucial information about an atom's electron configuration and the position of the element on the periodic table. Figure 1.2.6 graphs successive ionisation energies for an atom of sodium.

successive ionisation energies the energies required to remove the electrons from an atom, in sequence, starting with the outermost

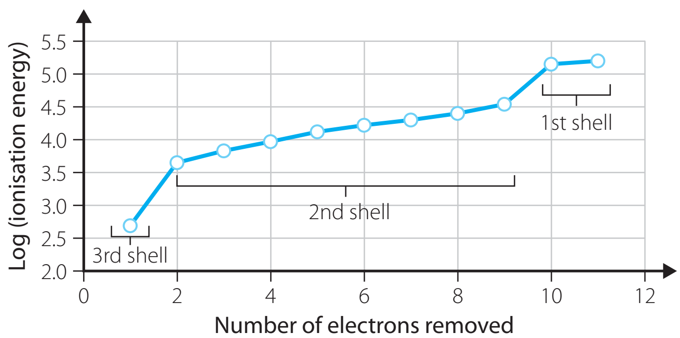


FIGURE 1.2.6 Successive ionisation energies of sodium

Sodium has eleven electrons, with the configuration 2:8:1. This means that there are two electrons in the lowest energy shell, 8 in the second and 1 electron in the highest energy shell. Figure 1.2.6 shows that there are particularly large increases in ionisation energy from the first to the second ionisation energy and from the ninth to the tenth ionisation energy. This corresponds to the change in electron shells. When an electron is removed from a new shell closer to the nucleus, the distance between the remaining electrons and the nucleus is reduced and the force of attraction between the electrons and nucleus is significantly greater. Accordingly, the increase in ionisation energy is larger than it would be if an additional electron were removed from the same shell.

When studying a graph of an atom's successive ionisation energies, you can quickly identify the electron configuration by counting the number of electrons removed between each large increase in ionisation energy.

SYLLABUS LINK

Chapter 2 discusses electron configurations in detail.

Electronegativity

Electronegativity is an atom's ability, when combined with another element, to attract electrons towards itself.

electronegativity the ability of an atom to attract electrons to itself, while being bonded to another atom

- Electronegativity increases from left to right across a period (Figure 1.2.7, page 14). As the atomic number (Z) increases, the number of protons in the nucleus increases. Thus, the nucleus is becoming increasingly positive across a period. Electrons in the same valence shell will each experience a stronger attraction to the nucleus moving across a period. The atomic radius decreases due to the increased nuclear charge across a period.

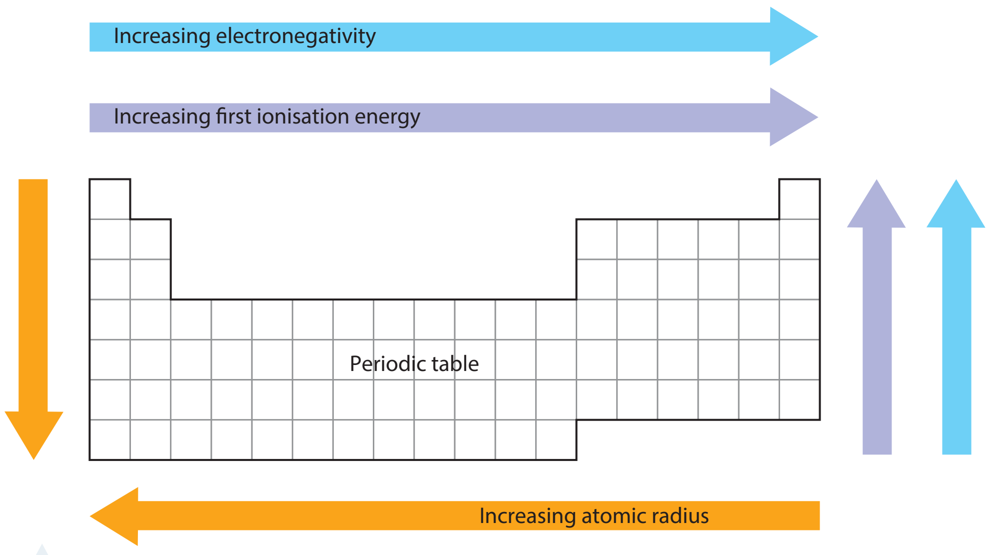


FIGURE 1.2.7 Trends in atomic radius, electronegativity and first ionisation energy in the periodic table

- Electronegativity decreases down a group. Going down a group, atoms have successively increasing numbers of electron shells. The outer shell electrons are further away from the nucleus consistent with a weaker attraction between the positive nucleus and the negative electrons for larger radius shells. The atomic radius also increases down a group. Increased shielding by the greater number of inner electrons also reduces the attraction of the outer electrons to the nucleus. Given that the nuclear attraction for electrons to the atomic nucleus is decreased, electrons associated with another bonded atom will also be less attracted and electronegativity also decreases.

  Electronegativity can be measured using a scale called the Pauling scale of electronegativity. The Pauling scale of electronegativity is a relative scale that includes every element, ranging from approximately 0.7 (francium) to 3.98 (fluorine).

The periodic table can be used to determine which of two elements is more electronegative. For example, oxygen is more electronegative than sulfur because it is above it in group 16. Chlorine is also more electronegative than sulfur as it is further to the right in period 3.

The most electronegative element is fluorine. The noble gases have zero electronegativity as they have a full valence shell, are stable and do not attract electrons towards themselves. This explains why the noble gases do not normally form compounds with other elements.

1.2.1 Atomic radius trends on periodic table
1.2.2 Valency
1.2.3 Atomic and ionic radius
1.2.4 Ionisation energy
1.2.5 First and second ionisation energies
1.2.6 Successive ionisation energies

SECTION REVIEW 1.2

REMEMBERING

1 State how many electrons are in the outer shells of:

- **a** beryllium
- **b** fluorine
- **c** phosphorus
- **d** sodium
- **e** argon.

2 Define 'first ionisation energy'.

UNDERSTANDING

3 Determine the valency of the following elements.

a Sodium

b Silicon

c Fluorine

4 Determine whether lithium or fluorine would have a bigger atomic radius.

APPLYING

5 In the following pairs of elements, state which element is more electronegative.

a Nitrogen and oxygen

b Magnesium and beryllium

c Phosphorus and fluorine

d Carbon and silicon

ANALYSING

6 Explain why atomic radius changes across a period and down a group.

7 Chlorine and oxygen are both strongly electronegative elements. Measured on the Pauling scale of electronegativity, chlorine has an electronegativity value of 3.16, and oxygen a value of 3.44. Suggest why the value for oxygen is higher.

1.3 Metallic and non-metallic character in the periodic table

Elements can be classified as metals, non-metals or metalloids.

Metals make up 75% of all elements. Metals are shiny in appearance, are good conductors of heat and electricity, and are **malleable** and **ductile**. Non-metals have the opposite properties of metals: they are dull in appearance, do not conduct electricity or heat, and are brittle. Metalloids are a small group of elements that have properties of both metals and non-metals.

Across the periodic table, there is a clearly observable trend, as the metallic character of elements decreases from left to right. Figure 1.3.1 shows this trend clearly. The 'staircase' of the metalloids, positioned between the metals and the non-metals, is also evident.

malleability
ability to be beaten into another shape or flattened into a thin sheet with a hammer without breaking

ductility
ability to be drawn out into a wire

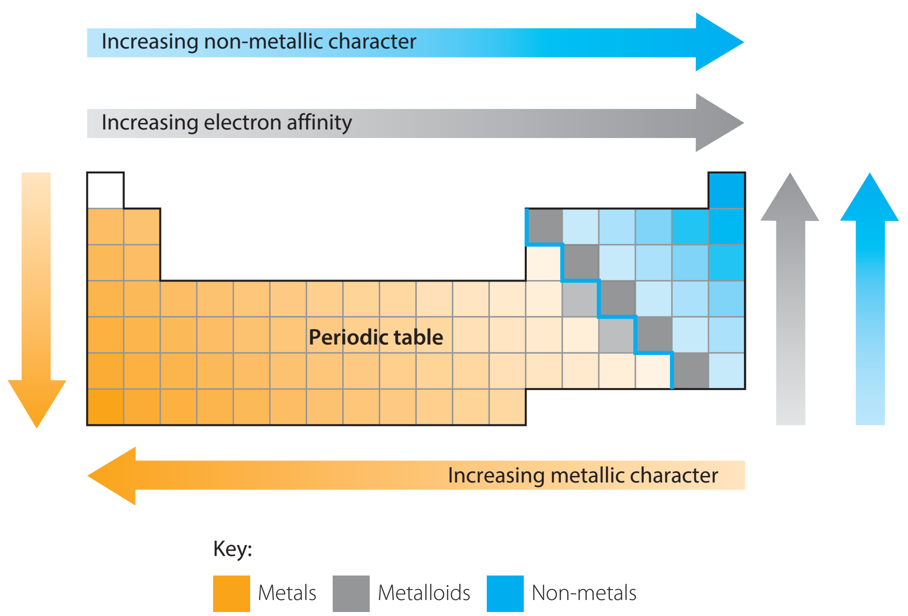


FIGURE 1.3.1 Metals, non-metals and metalloids in the periodic table

Group 1 – the alkali metals

The elements in group 1 are known as the alkali metals. Alkali metals occur in compounds in nature, rather than in their elemental form, because they are very reactive. They have a relatively low melting point compared with other metals and can be cut with a knife because they are so soft. These characteristics make them unsuitable for construction and fabrication, unlike metals such as iron and aluminium.

The alkali metals are so named because they all react with water to form a solution that is alkaline. This reaction can be quite violent, as in the case of potassium, which reacts spectacularly to form potassium hydroxide and hydrogen gas. The heat generated causes the hydrogen to burn vigorously, causing an explosion with a lilac flame. The reaction is shown in the following equation:

$$K(s) + H_2O(l) \rightarrow KOH(aq) + H_2(g)$$

Potassium + water → potassium hydroxide + hydrogen gas

1.3.1 The alkali metals
1.3.2 The halogens

Moving down group 1, the elements become more reactive. This can be explained by considering the first ionisation energy trend. The alkali metals all react to form a positive ion, by losing the one valence electron they all have. The amount of energy required to remove this electron decreases down the periodic table. This is because, with increasing period number, there is a greater number of electron shells for the atom and correspondingly increased shielding of the outer electrons from the nuclear attraction. This means it is easier it is to remove the valence electron. In summary, the further down the group the element occurs, the lower its first ionisation energy is, and the more its reactivity increases.

FIGURE 1.3.2 Properties of the halogens: all three physical states of the halogens: gas (chlorine), liquid (bromine) and solid (iodine)

Group 17 – the halogens

The trends and patterns of the elements in group 17, the halogens, provide a good contrast to those in group 1. Each halogen has a distinctive colour. All are non-metals, so they are poor conductors of heat and electricity. The halogens are toxic to life, which in turn makes them useful. For example, chlorine is used to treat swimming pool water and iodine is used as an antiseptic to help prevent wound infections. Importantly, the halogens have different characteristics when bonded with atoms from other groups. For example, halogens bonding with alkali atoms form alkali halides, such as sodium fluoride (NaF) and sodium chloride (NaCl), which are completely different because the halogen is now an anion. Sodium fluoride is effective in preventing dental decay and sodium chloride is common 'table' salt.

The halogens occur naturally in the form of diatomic molecules (e.g. F_2, Cl_2, Br_2) rather than the free atom. Along with other diatomic molecules such as hydrogen (H_2), oxygen (O_2) and nitrogen (N_2), scientists typically refer to the properties of the naturally occurring diatomic forms of these elements. The melting and boiling points of the halogens increase down the group: fluorine and chlorine are gases at room temperature, bromine is a liquid and iodine is a solid. Iodine has an unusual physical property in that, when it is heated, **sublimation** occurs, rapidly changing it from a grey solid into a purple gas.

sublimation
the process of a solid turning immediately into a gas when heated, without forming a liquid

Fluorine, the first element in group 17, is the lightest halogen and has the highest electronegativity value of all elements, making it extremely reactive. Almost all other elements form compounds with fluorine. Due to its reactivity, even the naturally occurring form of fluorine (i.e. F_2), is extremely difficult to manage in a laboratory and can be stored only in containers made from extremely stable, non-reactive materials such as platinum.

The halogens all have a high first ionisation energy, so they do not lose electrons when reacting. Instead, they try to gain a single electron to fill up their outer valence shell, to form a negatively charged ion. The ability to attract an electron is closely linked to the atomic radius.

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PRACTICAL ACTIVITY 1.3.1

Reactivity of the group 2 metals – the alkaline earth metals

Metals in group 2 of the periodic table are less reactive than those in group 1. When added to hydrochloric acid, the metals react to form the metal chloride, together with hydrogen gas.

AIM

To demonstrate the relative reactivity of elements within the group 2 metals

EQUIPMENT PER GROUP

- test-tube rack
- 2 test tubes
- splint
- 50 mL of 1 mol L^{-1} hydrochloric acid (HCl) (irritant)
- small piece of magnesium ribbon (highly flammable)
- small piece of calcium (flammable) (Do not use old calcium. Use fresh stock.)
- matches

WHAT ARE THE RISKS IN DOING THIS ACTIVITY?	HOW CAN THESE RISKS BE MANAGED?
	Wear eye protection

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, and ways to manage them.

PROCEDURE

1. Fill two test tubes one-quarter full with dilute hydrochloric acid.
2. Into one test tube, drop a small piece of magnesium.
3. Into the other test tube, drop a small piece of calcium.
4. Compare the reactivity of the two metals.
5. Drop another piece of magnesium into the first test tube and cover the end with your thumb.
6. When you can feel pressure, take your thumb off the test tube and test the gas with a lighted splint.
7. Record what happens.
8. Find a video of the reaction of sodium and potassium with water online. Note what you observe in comparison to the group 2 metals.

QUESTIONS

1. Write equations for the reactions you observe.
2. Moving down the periodic table, describe and explain the trends that are observed in the group 2 elements.
3. Compare the reactivity of the group 1 metals to the group 2 metals in the same period of the periodic table. Are group 1 metals less reactive or more reactive than group 2? Why is this so?

Fluorine is a small atom, which means that the distance between the positively charged nucleus at the centre of the atom and external electrons is relatively small, so the attraction to the nucleus is strong. Moving down the group, additional internal electron shells shield the valence shell from the nuclear attraction, so it is felt less strongly and attraction to the nucleus lessens, reducing electronegativity.

SECTION REVIEW 1.3

REMEMBERING

1. Define 'metalloid'.
2. Describe the metallic character trend in the periodic table.
3. State why alkali metals are so named.

UNDERSTANDING

4. List three reasons why chlorine is regarded as a non-metal.
5. Would you expect the reaction of sodium with water to be more, or less, vigorous than that of potassium? Explain why.

APPLYING

6. Show why the alkali metals all form an ion with a charge of +1.

ANALYSING

7. Explain why fluorine is the most electronegative element.
8. Explain the relationship between first ionisation energy and reactivity in groups 1 and 17.
9. Compare the reactivity of magnesium in group 2 to that of sodium in group 1.
10. Element number 117 was discovered in 2009 and named 'tennessine'. It is placed below astatine, at the bottom of group 17. So far, only a few atoms of tennessine have been obtained. If it were possible to collect a larger sample and investigate it, predict the properties it might have.

1.4 Properties of the oxides of the elements in period 3

oxide
a compound that forms when an element combines with oxygen

giant covalent network
a substance made from a series of covalently bonded atoms extending indefinitely throughout a whole crystal; very hard substances with high melting points (e.g. diamond is a giant covalent network formed from the element carbon)

covalent bond
a bond that is formed when the valence shells of two atoms overlap to share electrons

In examining the metallic character of the elements, you have seen examples of the trends that occur moving across the periodic table. Further trends can be observed by considering the properties of the oxides of the elements in period 3.

The formulas of the oxides of the elements vary according to the valency of the different elements in period 3. These can be seen in Table 1.4.1. Argon (not shown), together with the other group 18 elements, does not react with oxygen, so it does not form an oxide. Oxygen itself has a valency of 2.

TABLE 1.4.1 Formulas of the period 3 oxides

ELEMENT	Na	Mg	Al	Si	P	S	Cl
Valency	1	2	3	4	3	2	1
Formula of oxide	Na_2O	MgO	Al_2O_3	SiO_2	P_4O_6	SO_2	Cl_2O
Melting point (°C)	1132	2852	2072	1600	340	–72	–121
Bonding	Ionic	Ionic	Ionic	Macromolecular	Simple molecular	Simple molecular	Simple molecular

Table 1.4.1 demonstrates that the melting points vary for these compounds, which can be explained by the way in which the atoms bond together. Elements with strong metallic character tend to lose electrons when bonding, resulting in the formation of ions. Metallic character decreases from left to right across the periodic table. As such, the oxides of the metals sodium, magnesium and aluminium contain ions. Ionic substances such as these tend to have high melting points as the positively charged metal ions strongly attract the negatively charged oxide ions, requiring a lot of energy to disrupt the structure and cause the solid to melt.

Silicon dioxide is unique in period 3 in that it forms what is known as a **giant covalent network** solid. This means that silicon and oxygen atoms are bonded together in vast numbers, very strongly held by forces known as **covalent bonds**. As such, it too has a high melting point as these forces within the solid require a lot of heating to be overcome. Beach sand, for example, is mainly composed of silicon dioxide.

The oxides of phosphorus, sulfur and chlorine, however, are bonded differently. They exist as individual **molecules**, consisting of only a few atoms joined together. Due to the weak **intermolecular forces** that exist between them, the molecules can be separated easily from one another when heated, so their melting points are low.

The different bonding that exists within these oxides also affects the way in which they react in different circumstances.

Sodium oxide and magnesium oxides are known as **alkaline** oxides, because when added to water, they form an alkaline solution with a **pH** value of greater than 7. They also react when added to acids, in a **neutralisation** reaction.

Aluminium oxide is insoluble in water; however, it reacts with both acid and alkaline solutions and is classed accordingly as an **amphoteric** oxide.

Silicon dioxide does not dissolve in water; however, it does react when added to an alkaline solution, so it is classed as an **acidic** oxide.

Phosphorus, sulfur and chlorine oxides are all acidic oxides, because they form an acidic solution when added to water, and react to neutralise alkaline solutions.

Moving from left to right across the periodic table, the pattern is easy to observe: the oxides of the elements become less alkaline and more acidic.

molecule
a group of atoms covalently bonded together representing the smallest unit of an element or compound that can take part in a chemical reaction

intermolecular forces
attractive forces that cause molecules to aggregate together to form solids; they are significantly weaker than intramolecular forces, which are the ionic or covalent bonds that bond the atoms together inside a molecule

SYLLABUS LINK

Chapters 3 and 7 discuss bonding.

alkali
a substance that causes the concentration of hydrogen ions in a solution to decrease: reacts with acidic substances to form water, which is neutral, are often corrosive and have a pH value of greater than 7

pH
a scale used to measure the acidity of a solution: a solution of pH = 1 is strongly acidic; pH = 7 is neutral; and pH = 14 is strongly alkaline

neutralisation
a reaction that occurs when an acid is mixed with an alkali (or base), resulting in the formation of water

amphoteric substance
reacts with both alkaline and acidic solutions, so it is classified as neither an alkali nor an acid

acid
a substance that causes the concentration of hydrogen ions in a solution to increase: has a sharp, sour taste, are usually corrosive and have a pH value of less than 7

SECTION REVIEW 1.4

REMEMBERING

1 Define 'amphoteric oxide'.

2 What is the trend that is observed across period 3 in terms of the acidity of the oxides?

UNDERSTANDING

3 Discuss why magnesium oxide is regarded as alkaline.

APPLYING

4 Show the type of bonding that would exist in:
  a lithium oxide b nitrogen dioxide c boron oxide.

5 Classify the following oxides as acidic, alkaline or amphoteric.
  a Potassium oxide b Bromine oxide c Boron oxide

ANALYSING

6 Period 4 has 18 elements, compared with period 3, which has only eight. How does the trend in period 4 oxides compare to that of period 3? Is there a greater or smaller proportion of alkaline oxides?

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 Define the following terms.

- a Acid
- b Alkali metal
- c Alkaline
- d Amphoteric
- e Atomic radius
- f Electronegativity
- g Element
- h First ionisation energy
- i Group
- j Halogen
- k Metal
- l Metalloid
- m Noble gas
- n Period
- o Transition metal
- p Valency

CATEGORY QUESTIONS

2 Explain how the periodic table is organised.

3 Explain how the following properties of elements change moving down the periodic table:

- a Atomic radius
- b Electronegativity
- c First ionisation energy

4 Explain how the following properties of elements change moving across the periodic table, from left to right.

- a Atomic radius
- b Electronegativity
- c First ionisation energy

5 Explain how the melting points of the period 3 oxides change moving from left to right across the periodic table.

6 Explain how the acidity of the period 3 oxides changes moving from left to right across the periodic table.

ELABORATION QUESTIONS

7 Investigate one of the most recently discovered elements (any with an atomic number greater than 110). Outline the story behind its discovery, where it was synthesised, and the reason for its naming.

8 Explain how to use the periodic table to predict whether one element is more, or less, electronegative than another.

EVIDENCE QUESTIONS

9 Explain electron 'shielding', and how it is used to understand the trends in atomic radii and first ionisation energy when moving down a group and across a period.

10 What evidence do scientists now have that Mendeleev's original ideas were correct?

11 Conduct research into the work of Henry Moseley. Outline how his ideas confirmed our modern periodic table.

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END-OF-CHAPTER EXAM

End-of-chapter test

CHEMISTRY

1 The German chemist Johann Döbereiner developed the idea of triads. Triads were groups of three elements with similar properties and show where the mass of the middle element was close to the mean of the other two elements.

Which of the following would not be considered a triad?

A Boron, aluminium and thallium

B Fluorine, chlorine and bromine

C Sodium, potassium and rubidium

D Magnesium, calcium and strontium

2 Which of the following statements is correct for all metals?

A All metals have high melting points.

B All metals are low in reactivity.

C All metals are hard and strong.

D All metals conduct electricity.

3 Group 13 consists of a series of metallic elements. It is likely that the electronegativity of the group 13 elements:

A is at its highest for the period 4 elements.

B is greater than for the group 1 elements.

C decreases moving up the group.

D does not change within the group.

4 The element germanium has an atomic number of 32. It is in group 4 and period 4 of the periodic table. Which combination of properties from the table below is germanium likely to possess?

	CONDUCTIVITY OF SOLID GERMANIUM	pH OF GERMANIUM OXIDE
A	High	Alkaline
B	None	Amphoteric
C	Low	Acid
D	Low	Amphoteric

5 Which of the following shows the correct trends in properties, from left to right across the periodic table?

	FIRST IONISATION ENERGY	ATOMIC RADIUS	ELECTRONEGATIVITY
A	Decreases	Increases	Decreases
B	Increases	Increases	Increases
C	Increases	Decreases	Increases
D	Decreases	Decreases	Decreases

6 Explain why the first ionisation energy increases as you move across the periodic table.

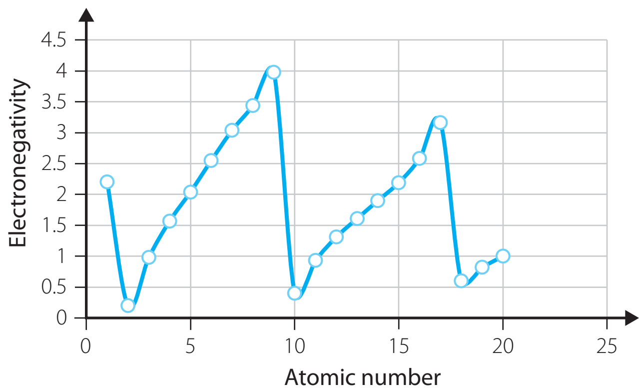
7 List the following atoms, in order of size, from smallest to largest: K, O, Li, Ba.

8 The following table shows the relationship between some properties.

Copy the table and complete it by stating whether the quantity in the left column is greater than (>), equal to (=) or less than (<) the quantity in the right column. Indicate your responses by placing a sign (>, = or <) in each row of the central column.

Radius of the K atom		Radius of the K^+ ion
Number of valence electrons in a Ca atom		Number of valence electrons in a Ba atom
Radius of an oxygen atom		Radius of a sulfur atom

9 The graph below shows how electronegativity changes with increasing atomic number for the first 20 elements.



a Define 'electronegativity'.

b Identify the element with the highest electronegativity.

c Explain why this element has the highest electronegativity value.

d Use data from the graph to identify the trend in electronegativity that occurs down a group of the periodic table and explain why this trend occurs.

e There are three elements on the graph that have very low electronegativity values. If the graph were extended to elements of higher atomic number, at what atomic number would the next low point occur? Explain your thinking.

10 Both aluminium and gallium are in group 13 in the modern periodic table. Before gallium was discovered, Mendeleev predicted its existence and listed its likely properties with great accuracy, as shown in the following table.

PROPERTY	MENDELEEV'S PREDICTION	LATEST DATA
Atomic mass	Approx. 68	70
Density ($g\,cm^{-3}$)	5.9	5.94
Melting point	Low	30.2
Formula of chloride	Ratio of 1:3	$GaCl_3$

Write a paragraph to explain:

a why Mendeleev left a 'gap' for gallium

b how he could predict the properties of gallium.

2 ATOMIC STRUCTURE

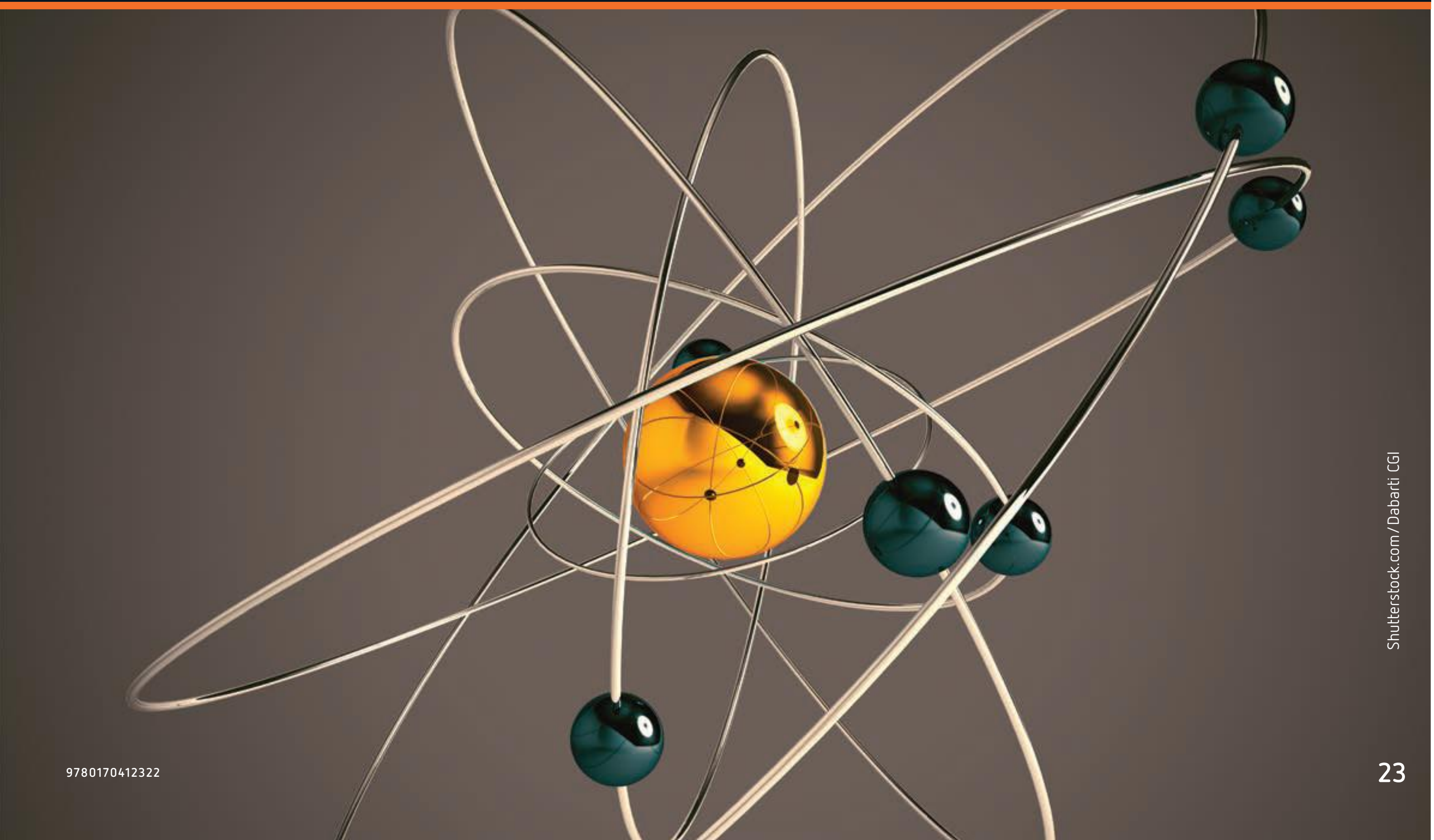
Introduction

Current understanding of the atom and its structure has been developed as a result of a number of important experiments over a long time. Our knowledge of atomic structure allows scientists to make predictions about the properties and reactions of different elements and compounds.

Stimulus questions

What is matter made from?

What evidence is there for the existence of particles so tiny we cannot see them?



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atom
the smallest stable unit of matter that can exist by itself

electron
a subatomic particle with a negative charge that is found in all atoms

alpha particle
a nucleus of a helium atom that is emitted by some radioactive substances as alpha radiation

nucleus
the positively charged centre of an atom, containing positively charged protons and electrically neutral neutrons, to which the orbiting electrons are attracted

2.1 The atom

In 1897, JJ Thomson discovered the electron, and in 1904 proposed what is known as the 'plum pudding' model of the **atom**. He realised that, within the atom, there were negatively charged subatomic particles, which are now recognised as **electrons**. He proposed that the atom was made of a positively charged material, throughout which electrons were distributed, like currants in a fruit cake or a plum pudding (Figure 2.1.1).

In 1911, almost 15 years after JJ Thomson's discovery, Ernest Rutherford led a team of scientists who carried out an experiment where they fired **alpha particles** at a thin sheet of gold foil. Following Thomson's plum pudding model, it was expected that the alpha particles would pass through the gold foil with little or no deflection.

Rutherford found that a small number of the alpha particles experienced a significant deflection as though they had struck something large. This did not support his hypothesis and led Rutherford to theorise that most of the atom is empty space, but with a small, incredibly dense, positively charged structure in its centre. All the mass of the atom was concentrated in this structure, which is now known as the **nucleus**.

PRACTICAL ACTIVITY 2.1.1

Gold foil experiment

AIM

To simulate Ernest Rutherford's gold foil experiment

EQUIPMENT PER GROUP

- 1 hula hoop
- 1 golf ball
- 1 ping-pong ball
- string

PROCEDURE

1. Work in pairs.
2. Tie the golf ball onto the hula hoop so it hangs approximately in the middle of the hula hoop when it is upright.
3. With their eyes closed, one student will lightly toss a ping-pong ball 30 times randomly through the hula hoop.
4. The other student is to record how many times the ping-pong ball hits the golf ball.

QUESTIONS

1. Explain what the hula hoop, the golf ball and the ping-pong ball each represent, in relation to Rutherford's experiment.
2. Compare the number of times the golf ball was hit by the ping-pong ball with the number of times it was missed. Explain whether this aligns with what Rutherford found in his experiment.
3. Explain how the results of Rutherford's experiment might lead you to the conclusion he made.

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Based on Rutherford's work and the contributions of many other scientists, such as Henry Moseley and James Chadwick, scientists now consider that the nucleus contains subatomic particles called **protons** and **neutrons**. The nucleus contains most of the mass of the atom, but occupies only a small volume of the atom. Electrons surround the nucleus. Electrons are very small, move extremely quickly, and are spread out over a relatively large distance. This creates an electron cloud around the nucleus (Figure 2.1.2). Despite the large amount of space they cover, electrons contribute almost no mass to the atom. If the atom were the size of The Brisbane Cricket Ground ('the Gabba'), including the grandstands, then each electron would be the size of a pea somewhere in the whole ground, while the nucleus would be no bigger than a peach in the centre of the playing field.

proton
a positively charged particle that exists in the nucleus of all atoms

neutron
an uncharged particle that exists in the nucleus of all atoms except hydrogen

2.1.1 The Rutherford experiment
2.1.2 Rutherford's gold foil experiment
2.1.3 Atoms: the space between
2.1.4 Build an atom

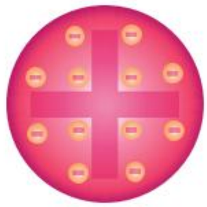

FIGURE 2.1.1 JJ Thomson's plum pudding model of the atom

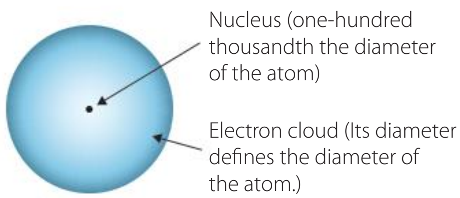

FIGURE 2.1.2 Structure of an atom, showing a nucleus and an electron cloud

Subatomic particles

An atom contains three types of subatomic particle – the proton, the neutron and the electron. The properties of each particle are shown in Table 2.1.1.

TABLE 2.1.1 Symbols, charges and relative masses of subatomic particles

PARTICLE	SYMBOL	CHARGE	RELATIVE MASS
Proton	p	+1	1
Neutron	n	0	1
Electron	e	–1	$\frac{1}{1800}$ (approx.)

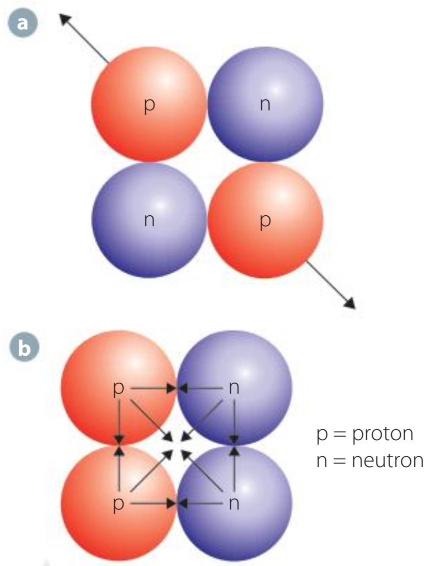

FIGURE 2.1.3 Forces in the nucleus. **a** Repulsive forces act between protons in a nucleus. **b** The attractive strong nuclear force acts between all particles in the nucleus.

What holds an atom together?

In an atom, the nucleus has an overall positive charge, because it consists of positively charged protons and neutrally charged neutrons. The electrons around the nucleus are negatively charged. The attraction between the positive nucleus and the negative electrons keeps the atom together. This is called **electrostatic attraction**.

An atom can be modelled as a nucleus surrounded by electrons, held together by electrostatic forces of attraction between the nucleus and the electrons.

electrostatic attraction
a force that occurs between particles with opposite charges

electrostatic repulsion
a force that occurs between particles with the same charge

strong nuclear force
an attractive force that occurs between all particles in the nucleus, regardless of charge

How does a nucleus stay together?

The nucleus of an atom is very stable despite being composed only of positive protons and neutral neutrons. There is a tendency for the protons to repel each other, due to **electrostatic repulsion** forces that act between them. However, presence of the neutrons in the nucleus increases the effect of the **strong nuclear force** to the extent that it balances the repulsion of the protons and creates a stable nucleus (Figure 2.1.3).

SECTION REVIEW 2.1

REMEMBERING

1 Name the three particles in an atom and state the charge of each particle.

2 Identify where each of the three particles is found in the atom.

UNDERSTANDING

3 Explain the balance of forces that exists in a nucleus that makes it stable.

4 Rutherford deduced that the nucleus of an atom is very small, extremely dense and positively charged. Explain how his experiment led him to draw these conclusions.

APPLYING

5 If Thomson's model of the atom had been correct, what results would Rutherford have observed from his gold foil experiment?

ANALYSING

6 Element 113 does not occur naturally. It was synthesised in 2012 by Japanese researchers. It was unstable so it decayed in less than a second. Propose the reason for the rapid decay of this element.

2.2 Protons, neutrons and electrons

The composition of elements can be represented to give information about the number of protons, neutrons and electrons in an atom (Figure 2.2.1).

2.2.1 Protons, neutrons and electrons
2.2.2 Atomic number, atomic mass and isotopes
2.2.3 Atomic number and mass number

FIGURE 2.2.1 Atoms can be represented by a symbol that includes the atomic number and mass number.

Mass number (total number of protons + neutrons)

$${}^{A}_{Z}\mathrm{X}$$

Symbol of element (depends on the value of Z)

Atomic number (number of protons)

atomic number
the number of protons contained within one atom of an element – each element has a unique atomic number

mass number
the number of protons and neutrons in the nucleus of an atom

relative atomic mass
the average mass of one atom of an element, relative to the mass of carbon-12

isotope
two atoms of the same element with the same number of protons but with different masses, due to their different number of neutrons

The **atomic number** (Z) represents the number of protons in the atom and it is the number that defines an atom. For example, an atom of carbon will always have six protons, even if the number of neutrons or electrons changes. In an atom that is uncharged, the number of electrons and protons are equal, so the atomic number will also give the number of electrons in the atom. On a periodic table, the atomic number is the smaller number inside the box, often located above the element symbol.

The **mass number** (A) is the total number of protons and neutrons in the atom. To find the number of neutrons in an atom, subtract the atomic number, Z, from the mass number, A.

$$\text{Number of neutrons} = \text{mass number } (A) - \text{atomic number } (Z)$$

Often, on a periodic table, the **relative atomic mass** (RAM) of an element is displayed, rather than the mass number. For example, carbon has an RAM of 12.01 in the periodic table. This number represents the average mass of all the different isotopes of the element that exist, so is not often a whole number. The mass number of the most common **isotope** of an element can usually be found by rounding the RAM to the nearest whole number.

The information about mass and atomic numbers found in the periodic table is shown in Figure 2.2.2.

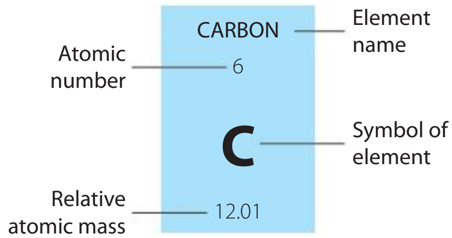

FIGURE 2.2.2 An example of how atomic information is displayed in the periodic table. Note the inclusion of the element name, atomic number, symbol of the element, and relative atomic mass.

WORKED EXAMPLE 2.2.1

Sodium is represented by $^{23}_{11}Na$. Determine the number of protons, neutrons and electrons in sodium.

ANSWER

- The atomic number of sodium is 11. This means that it has 11 protons.
- The mass number of sodium is 23, so it has $23 - 11 = 12$ neutrons.
- The atom is neutral, so there is an equal number of electrons and protons.

Therefore, sodium has 11 protons, 12 neutrons and 11 electrons.

SECTION REVIEW 2.2

REMEMBERING

1 Define:

a atomic number **b** mass number.

UNDERSTANDING

2 Outline how to determine the numbers of protons, neutrons and electrons in an atom.

3 Distinguish between mass number and relative atomic mass.

APPLYING

4 Identify each element below, and calculate the number of protons, neutrons and electrons in each.

a $^{19}_{9}F$ **b** $^{80}_{35}Br$

5 Represent each of the following as $^{A}_{Z}X$, using the periodic table as a reference.

a An atom of aluminium with a mass number of 27

b An atom with four protons and a mass number of 9

6 Copy and complete the following table. You will need to extract information from the periodic table for some elements.

ELEMENT	ATOMIC NUMBER	MASS NUMBER	NUMBER OF PROTONS	NUMBER OF NEUTRONS	NUMBER OF ELECTRONS
Hydrogen	1	1			
Magnesium			12	12	
Boron	5			6	
Chlorine		35			17
Nickel		59		31	

2.3 The electron configuration of atoms

Electrons are found orbiting the nucleus of every atom. Their behaviour can be described using a branch of advanced mathematics called 'quantum theory'. The 'shell model' of the atom, based on quantum theory, establishes the following principles.

electron energy levels, or shells
a group of electrons existing at the same distance from the nucleus of an atom; all the electrons in each shell have the same amount of energy

- Electrons exist in **electron energy levels, or shells**. Every electron in the same shell has the same amount of energy.
- Electrons in an atom cannot exist with energies that lie between the atom's defined energy levels, although they can move from one energy level to another when the atom absorbs or releases energy.
- The difference in the energy levels of the shells decreases the further away they are from the nucleus.
- The electron shell with the lowest energy level is closest to the nucleus. The further away from the nucleus the shell is, the greater the energy level it has (Figure 2.3.1).
- Each shell within an atom can hold only a specific maximum number of electrons. The general formula $2n^2$ can be used to determine the maximum number of electrons an energy level can hold, where n is the energy level number. For example, energy level $n = 3$ can hold $2 \times 3^2 = 18$ electrons. Table 2.3.1 shows the maximum number of electrons that can be found in each energy level.

2.3.1 Electron configuration

TABLE 2.3.1 Maximum number of electrons in each energy level

ENERGY LEVEL (n)	CALCULATION OF ELECTRON NUMBER ($2n^2$)	NUMBER OF ELECTRONS
1	2×1^2	2
2	2×2^2	8
3	2×3^2	18
4	2×4^2	32

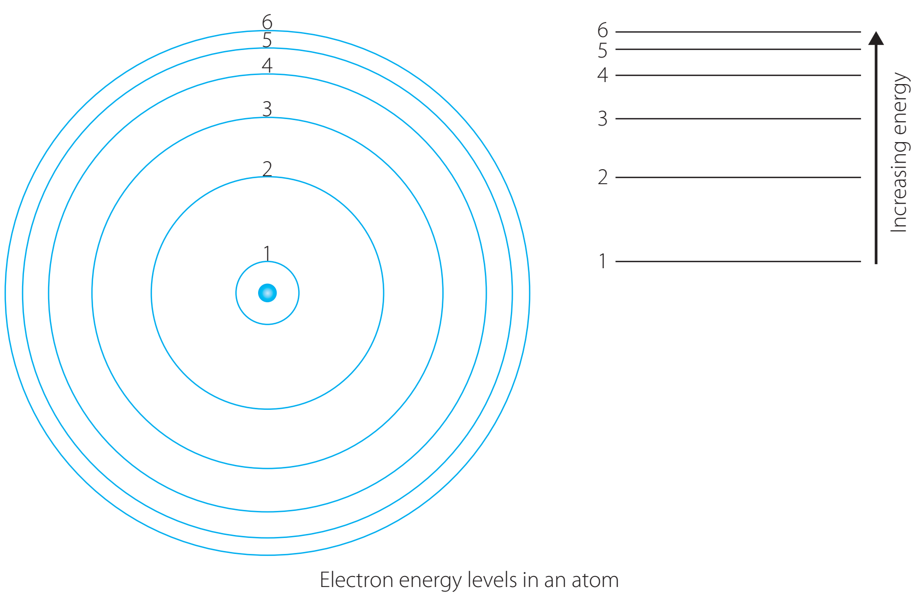

FIGURE 2.3.1 Electron shells in an atom

Writing an electron configuration

The electron configuration of each atom is unique. To determine the electron configuration of an atom a series of rules should be followed. The Aufbau principle states that electrons around a nucleus fill up in order and occupy the lowest energy levels first. This means that the single electron in hydrogen goes into the lowest energy level and the configuration is written as 1. The two electrons in helium also go into the lowest energy level. The electron configuration is written as 2.

electron configuration the specific arrangement of electrons within an individual atom

Aufbau principle electrons around an atom will fill the lowest available energy level first, before filling higher energy levels

Lithium has three electrons. The first two of these exist in the lowest energy level, which is now full. The remaining electron moves to the next highest energy level. The configuration of the lithium atom is written as 2,1.

Configurations are written to show how many electrons are in each level, and the order of the levels. The configurations for the first 10 elements are shown in Table 2.3.2.

TABLE 2.3.2 Electron configuration of the first 10 elements

ELEMENT	NUMBER OF ELECTRONS	ELECTRON CONFIGURATION
Hydrogen	1	1
Helium	2	2
Lithium	3	2,1
Beryllium	4	2,2
Boron	5	2,3
Carbon	6	2,4
Nitrogen	7	2,5
Oxygen	8	2,6
Fluorine	9	2,7
Neon	10	2,8

From neon, the first two energy levels are now full. The eight electrons of the next eight elements move to the third energy level. These configurations are shown in Table 2.3.3.

TABLE 2.3.3 Electron configuration of the elements of period 3

ELEMENT	NUMBER OF ELECTRONS	ELECTRON CONFIGURATION
Sodium	11	2,8,1
Magnesium	12	2,8,2
Aluminium	13	2,8,3
Silicon	14	2,8,4
Phosphorus	15	2,8,5
Sulfur	16	2,8,6
Chlorine	17	2,8,7
Argon	18	2,8,8

Atomic orbitals

subshell
a part of an electron shell that contains orbitals of the same energy

atomic orbital
the region of space around an atom that has a specific shape and may contain a maximum of two electrons

Each electron shell contains one or more **subshells**, designated 's', 'p', 'd' and 'f'. Each subshell contains groups of **atomic orbitals**, inside which a maximum of two electrons can exist. Each subshell type consists of a specific number of orbitals, as shown in Table 2.3.4.

TABLE 2.3.4 Subshells, orbitals and electrons

SUBSHELL DESIGNATION	NUMBER OF ORBITALS	NUMBER OF ELECTRONS IN THE SUBSHELL
s	1	2
p	3	6
d	5	10
f	7	14

The greater the shell number of the shell, the greater the number of subshells it holds. Table 2.3.5 shows the numbers of subshells, orbitals and electrons in each shell, confirming the maximum number of electrons that each shell can hold.

TABLE 2.3.5 The orbitals and electrons within the first three shells

SHELL NUMBER	SUBSHELL	NUMBER OF ORBITALS	NUMBER OF ELECTRONS	MAXIMUM NUMBER OF ELECTRONS IN SHELL ($2n^2$)
1	1s	1	2	2
2	2s	1	2	8
	2p	3	6	
3	3s	1	2	18
	3p	3	6	
	3d	5	10	

Pauli exclusion principle
no two electrons can occupy the same quantum state; hence, two electrons in the same orbital will have opposing spins

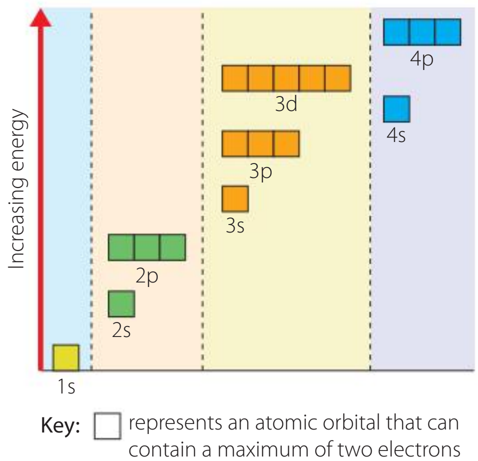

FIGURE 2.3.2 Energy order of the s, p and d subshells

These subshells have different energies. In a particular energy level, the s subshell always has the lowest energy, followed by the p subshell, then the d and f subshells. Figure 2.3.2 shows all the subshells in order of increasing energy.

According to the Aufbau principle this means that the 1s subshell is the lowest in energy and will fill first, followed by the 2s, the 2p, and so on.

The **Pauli exclusion principle** states that no two electrons can occupy the same quantum state. However, the maximum number of electrons that a single orbital can have is *two*, not one. This means that the two electrons in any orbital in a subshell must have different quantum states. The two states differ by their magnetic spin quantum numbers, which can be thought of as 'spin up' or 'spin down'. Scientists designate electrons in the same orbital by arrows, pointing up or down, to distinguish their spin states. This means that every electron has a unique location or address, so, a particular notation is used to represent this address. The combination of addresses for all the electrons is known as the electron configuration.

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TABLE 2.3.6 **Electronic configuration of first 18 elements using subshell notation**

ELEMENT	ATOMIC NUMBER	ELECTRON CONFIGURATION	ORBITAL DIAGRAM
Hydrogen	1	$1s^1$	[↑]
Helium	2	$1s^2$	[↑↓]
Lithium	3	$1s^2\,2s^1$	[↑↓] [↑]
Beryllium	4	$1s^2\,2s^2$	[↑↓] [↑↓]
Boron	5	$1s^2\,2s^2\,2p^1$	[↑↓] [↑↓] [↑][ ][ ]
Carbon	6	$1s^2\,2s^2\,2p^2$	[↑↓] [↑↓] [↑][↑][ ]
Nitrogen	7	$1s^2\,2s^2\,2p^3$	[↑↓] [↑↓] [↑][↑][↑]
Oxygen	8	$1s^2\,2s^2\,2p^4$	[↑↓] [↑↓] [↑↓][↑][↑]
Fluorine	9	$1s^2\,2s^2\,2p^5$	[↑↓] [↑↓] [↑↓][↑↓][↑]
Neon	10	$1s^2\,2s^2\,2p^6$	[↑↓] [↑↓] [↑↓][↑↓][↑↓]
Sodium	11	$1s^2\,2s^2\,2p^6\,3s^1$	[↑↓] [↑↓] [↑↓][↑↓][↑↓] [↑]
Magnesium	12	$1s^2\,2s^2\,2p^6\,3s^2$	[↑↓] [↑↓] [↑↓][↑↓][↑↓] [↑↓]
Aluminium	13	$1s^2\,2s^2\,2p^6\,3s^2\,3p^1$	[↑↓] [↑↓] [↑↓][↑↓][↑↓] [↑↓] [↑][ ][ ]
Silicon	14	$1s^2\,2s^2\,2p^6\,3s^2\,3p^2$	[↑↓] [↑↓] [↑↓][↑↓][↑↓] [↑↓] [↑][↑][ ]
Phosphorus	15	$1s^2\,2s^2\,2p^6\,3s^2\,3p^3$	[↑↓] [↑↓] [↑↓][↑↓][↑↓] [↑↓] [↑][↑][↑]
Sulfur	16	$1s^2\,2s^2\,2p^6\,3s^2\,3p^4$	[↑↓] [↑↓] [↑↓][↑↓][↑↓] [↑↓] [↑↓][↑][↑]
Chlorine	17	$1s^2\,2s^2\,2p^6\,3s^2\,3p^5$	[↑↓] [↑↓] [↑↓][↑↓][↑↓] [↑↓] [↑↓][↑↓][↑]
Argon	18	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6$	[↑↓] [↑↓] [↑↓][↑↓][↑↓] [↑↓] [↑↓][↑↓][↑↓]

Hydrogen has an electron configuration $1s^1$, with the superscript 1 representing the number of electrons in the subshell. Helium is $1s^2$. To write the configuration for lithium, the next electron needs to go into the next subshell, so lithium is $1s^2\,2s^1$. Table 2.3.6 shows the electron configuration for the first 18 elements, using this method. Each 'box' represents an orbital, and the electrons are represented by arrows with a pair of electrons in any orbital having different spins (i.e. 'spin up' or 'spin down').

Note that, when filling p, d, or f subshells, **Hund's rule of maximum multiplicity** applies. This states that for any given electron configuration, the lowest energy or 'ground' state is such that if two or more orbitals of equal energy are equally available, the electrons will occupy them all singly before filling them in pairs.

For example, nitrogen has the electron configuration of $1s^2\,2s^2\,2p^3$. The preferred **ground state** arrangement of the electrons in nitrogen is contrasted with an alternative '**excited state**', of higher energy as shown in Figure 2.3.3.

Hund's rule of maximum multiplicity
every orbital in a subshell is singly occupied with one electron before any one orbital is doubly occupied

ground state
the lowest energy electron configuration of an atom

excited state
an altered electron configuration from the ground state that occurs when an atom absorbs additional energy

FIGURE 2.3.3 Ground state and excited state electron configurations of nitrogen

Electron configurations for elements in period 4

As shown in Table 2.3.6, element number 18, argon, has the configuration $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$. This follows the Aufbau principle where electron orbitals fill from the lowest energy orbital first. In Figure 2.3.2, the subshells are filled in order of increasing subshell energy, as follows: $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p$.

Following on from the 3p subshell, the next highest orbitals are those from the 4s subshell, rather than the 3d. Consequently, potassium and calcium, elements 19 and 20, have electrons in the 4th shell, despite the 3rd shell not being completely filled. For scandium, element 21, the 4s subshell is filled, so the remaining electron is held in the 3d subshell. This pattern continues until element 31, gallium, where the 3d subshell is now complete, so the remaining electron is held in the next available orbital, in the 4p subshell. The electron configurations of elements 19 to 36 are shown in Table 2.3.7 on page 11. In this example, the convention of the [Ar] symbol has been used to save space, where [Ar] represents the configuration of an argon atom.

Electron configuration of ions

The subshell notation can be used to represent the electron configuration of ions in the same way as for atoms. An ion is an atom that has gained or lost an electron. A cation is an atom that has lost an electron and become positively charged. An anion is an atom that has gained an electron and become negatively charged.

According to the Aufbau principle, when a cation is formed, electrons are lost from the highest energy orbital first. When an anion is formed, electrons are added into the next available orbital.

SYLLABUS LINK

Chapter 3 discusses the formation of ions and ionic bonding in more detail.

Forming ions from period 4 elements

An anomaly is apparent relating to our previous explanations when considering how ions form in period 4 elements with outer electrons in the 3d subshell. When filling the subshells, electrons are added to the 4s subshell before the 3d subshell. However, once the 3d subshell is filled, its energy level falls to a level below that of the 4s. This means that when ions are formed, the 4s subshell is now the highest energy orbital, and electrons are lost from the 4s subshell before the 3d subshell.

WORKED EXAMPLE 2.3.1

Write the electron configuration for a sodium ion (Na^+).

ANSWER

1 Sodium has an atomic number of 11, so it has 11 electrons. Its configuration must be:

$$Na: 1s^2\ 2s^2\ 2p^6\ 3s^1.$$

2 To form a cation with a charge of +1, the sodium atom must lose its highest energy electron. Therefore, the electron configuration for a sodium ion is:

$$Na^+: 1s^2\ 2s^2\ 2p^6.$$

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TABLE 2.3.7 Electron configurations for elements 19 to 36

ELEMENT	ATOMIC NUMBER	ELECTRON CONFIGURATION	DIAGRAM
Potassium	19	[Ar] $4s^1$	[Ar] [↑]
Calcium	20	[Ar] $4s^2$	[Ar] [↑↓]
Scandium	21	[Ar] $4s^2\ 3d^1$	[Ar] [↑↓] [↑][ ][ ][ ][ ]
Titanium	22	[Ar] $4s^2\ 3d^2$	[Ar] [↑↓] [↑][↑][ ][ ][ ]
Vanadium	23	[Ar] $4s^2\ 3d^3$	[Ar] [↑↓] [↑][↑][↑][ ][ ]
Chromium	24	[Ar] $4s^1\ 3d^5$	[Ar] [↑] [↑][↑][↑][↑][↑]
Manganese	25	[Ar] $4s^2\ 3d^5$	[Ar] [↑↓] [↑][↑][↑][↑][↑]
Iron	26	[Ar] $4s^2\ 3d^6$	[Ar] [↑↓] [↑↓][↑][↑][↑][↑]
Cobalt	27	[Ar] $4s^2\ 3d^7$	[Ar] [↑↓] [↑↓][↑↓][↑][↑][↑]
Nickel	28	[Ar] $4s^2\ 3d^8$	[Ar] [↑↓] [↑↓][↑↓][↑][↑][↑]
Copper	29	[Ar] $4s^1\ 3d^{10}$	[Ar] [↑] [↑↓][↑↓][↑↓][↑↓][↑↓]
Zinc	30	[Ar] $4s^2\ 3d^{10}$	[Ar] [↑↓] [↑↓][↑↓][↑↓][↑↓][↑↓]
Gallium	31	[Ar] $4s^2\ 3d^{10}\ 4p^1$	[Ar] [↑↓] [↑↓][↑↓][↑↓][↑↓][↑↓] [↑][ ][ ]
Germanium	32	[Ar] $4s^2\ 3d^{10}\ 4p^2$	[Ar] [↑↓] [↑↓][↑↓][↑↓][↑↓][↑↓] [↑][↑][ ]
Arsenic	33	[Ar] $4s^2\ 3d^{10}\ 4p^3$	[Ar] [↑↓] [↑↓][↑↓][↑↓][↑↓][↑↓] [↑][↑][↑]
Selenium	34	[Ar] $4s^2\ 3d^{10}\ 4p^4$	[Ar] [↑↓] [↑↓][↑↓][↑↓][↑↓][↑↓] [↑↓][↑][↑]
Bromine	35	[Ar] $4s^2\ 3d^{10}\ 4p^5$	[Ar] [↑↓] [↑↓][↑↓][↑↓][↑↓][↑↓] [↑↓][↑↓][↑]
Krypton	36	[Ar] $4s^2\ 3d^{10}\ 4p^6$	[Ar] [↑↓] [↑↓][↑↓][↑↓][↑↓][↑↓] [↑↓][↑↓][↑↓]

WORKED EXAMPLE 2.3.2

Write the electron configuration for a sulfide ion (S^{2-}).

ANSWER

1 Sulfur has an atomic number of 16, so it has 16 electrons. Its configuration must be:

$$S: 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4.$$

2 To form an anion with a charge of −2, the sulfur atom must gain two electrons. Therefore, the electron configuration for a sulfide ion is:

$$S^{2-}: 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6.$$

WORKED EXAMPLE 2.3.3

Write the electron configuration for an iron(II) ion (Fe^{2+}).

ANSWER

1 Iron has an atomic number of 26, so its electron configuration is:

$$Fe: 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6.$$

2 To form an ion with a charge of +2, it must lose 2 electrons. When ions are formed, electrons are lost from the 4s subshell before the 3d subshell. The electron configuration for an Fe^{2+} ion is:

$$Fe^{2+}: 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6.$$

Chromium and copper – apparent exceptions in electron configuration

The electron configurations for chromium and copper do not follow the pattern of the other elements in period 4 (Tables 2.3.7 and 2.3.8). The reasons for these apparent exceptions stem again from the energy levels of the 3d orbitals changing when electrons are added.

In the case of chromium, when the 3d subshell is half filled, with one electron in each orbital, it becomes lower in energy than the 4s subshell, so electrons exist in the 3d subshell in preference to the 4s.

Similarly for copper, when the 3d subshell is completely filled, it too is lower in energy than the 4s subshell, and again there is only one electron in the 4s subshell.

TABLE 2.3.8 Electron configurations for copper and chromium

ELEMENT	ATOMIC NUMBER	EXPECTED ELECTRON CONFIGURATION	ACTUAL ELECTRON CONFIGURATION
Chromium	24	$[Ar]\ 4s^2\ 3d^4$	$[Ar]\ 4s^1\ 3d^5$
Copper	29	$[Ar]\ 4s^2\ 3d^9$	$[Ar]\ 4s^1\ 3d^{10}$

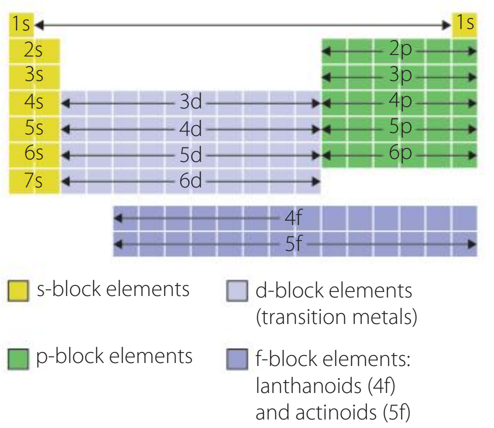


FIGURE 2.3.4 Subshells related to position in the periodic table

Electron configurations and the periodic table

The periodic table provides a very useful way of quickly determining the electron configuration of an element. Figure 2.3.4 shows how the periodic table can easily be divided into 'blocks'. For example, elements in the s-block have their highest energy electrons in an s orbital while the transition elements exist in the d-block, and the lanthanoids and actinoids, the f-block.

The periodic table can also be used to predict which subshells have higher or lower energy than others. The subshell with the lowest energy is the 1s subshell. The order of the subshells, in terms of their energy, is $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p$ and so on. This is found by simply reading across the periodic table from left to right, following the rows down. Each of the transition metal rows includes a subshell that belongs in the energy level above the row it is in. For example, the d subshell in the fifth row of the periodic table is the d subshell in the fourth energy level, subshell 4d.

SECTION REVIEW 2.3

REMEMBERING

1 State the maximum number of electrons that can be found in the first three energy levels.

2 Explain what the electron configuration of an atom is.

3 Explain how the Aufbau principle explains the electron configuration of an atom.

4 State the electron configuration of the:

a potassium atom and ion

b sulfur atom and ion

c nitrogen atom and ion

d calcium atom and ion.

UNDERSTANDING

5 Write electron configurations for the following elements using 2,8 notation.

a Beryllium

b Sulfur

c Calcium

d Helium

6 Write electron configurations for the following, using s, p, d, f notation.

a Nitrogen atom

b Phosphide P^{3-} ion

c Scandium atom

d Chromium

e Germanium Ge^{4+} ion

7 Identify the elements that have the following electron configurations.

a 2,8,2

b $1s^2$

c $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$

APPLYING

8 The following electron configurations are incorrect. Identify each element, explain why the configuration is incorrect and write the correct configuration for that element in the s, p, d, f form.

a 2,9

b 2,8,10

ANALYSING

9 The Na^+ and F^- ions are said to be 'iso-electronic'. Use the electron configurations for these ions to demonstrate what iso-electronic means.

10 Explain why the electron configuration of a copper atom is unusual compared with the other atoms in the first row of the transition metals.

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 Define the following terms.
- a Atomic number
- b Atomic orbital
- c Mass number
- d Shell
- e Subshell

2 Explain the models of the atom proposed by Thomson and Rutherford.

3 Explain how the number of subatomic particles can be determined from the atomic number and mass number.

4 Explain how the periodic table can be divided up into the s-block, p-block, d-block and f-block.

5 Explain how the following determine the electron configuration of an atom.
- a Aufbau principle
- b Pauli exclusion principle
- c Hund's rule of maximum multiplicity

CATEGORY QUESTIONS

6 Compare the sizes and charges of the three particles found in an atom.

7 What information about the structure of an atom can be determined by the following in the periodic table?
- a The group an element is in
- b The period an element is in

8 An element has seven protons and a mass number of 15. How many neutrons and electrons will it have?

9 Identify the number of valence electrons in each of the following elements.
- a Oxygen
- b Chlorine
- c Magnesium
- d Selenium

10 Write the electron configuration of the following, using the s, p, d, f notation.
- a Boron atom
- b Sulfide S^{2-} ion
- c Lithium
- d Vanadium
- e Manganese Mn^{2+} ion
- f Gallium Ga^{3+} ion
- g Krypton

ELABORATION QUESTIONS

11 Distinguish between the mass number and the relative atomic mass of an element.

12 An atom has 14 electrons. Show how they would be arranged around the nucleus, using the s, p, d, f notation.

EVIDENCE QUESTIONS

13 Explain why, despite scandium being in the fourth period of the periodic table, there is no element directly above it.

14 Explain why the electron configuration of chromium does not fit the pattern of the other elements in the first row of transition elements.

END-OF-CHAPTER EXAM

End-of-chapter test

CHEMISTRY

1 Which of the following is not the electron configuration of a transition metal atom?

A $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^2$

B $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$

C $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ 5s^1$

D $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2$

2 An ion that contains seven protons, eight neutrons and nine electrons will have a mass number and charge corresponding to:

	MASS NUMBER	CHARGE
A	15	+2
B	8	−2
C	15	−2
D	8	+2

3 What is the ground state electron configuration for $^{16}S^{2-}$?

A $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4\ 4s^2$

B $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$

C $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

D $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

4 The 3d subshell has:

A 3 orbitals and can hold up to six electrons.

B 5 orbitals and can hold up to 10 electrons.

C 3 orbitals and can hold up to nine electrons.

D 5 orbitals and can hold up to 15 electrons.

5 How many electrons are in the cation $^{59}Co^{3+}$?

A 27

B 30

C 59

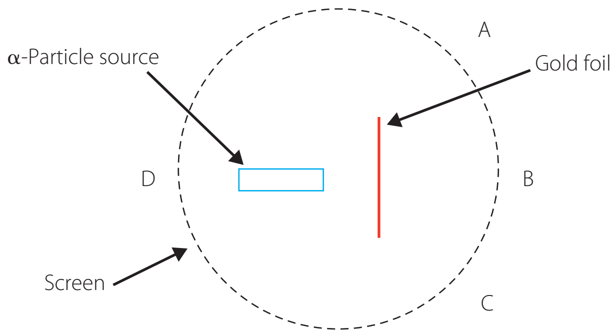
D 24

6 The atomic number of potassium is 19.

a Write the electron configuration, in terms of shells and subshells, for the potassium atom in its ground state.

b Write the electron configuration, in terms of shells and subshells, for the potassium ion.

7 In Rutherford's gold foil experiment, he bombarded a thin piece of gold foil with α-particles as shown in the simplified diagram.



a Referring to the diagram, rank the letters A to D from highest to lowest, to show where particles were detected after they had hit the gold foil.

________ > ________ > ________ > ________

b State three pieces of information about the nucleus that Rutherford deduced from the results from this experiment.

c The development of quantum theory added to Rutherford's atomic model. Briefly describe these additions.

9780170412322

3 INTRODUCTION TO BONDING

Introduction

Although there are 118 known elements, an infinite number of different materials are made from these elements. To form these new substances, atoms bond together to form compounds with their own unique properties. Understanding how this happens is fundamental to our ability to create new materials.

Stimulus questions

How do elements combine to form new substances?

How can scientists predict the types of substance elements will form?

3.1 Chemical bonding

element
a substance consisting of only one type of atom

compound
a substance consisting of atoms of two or more elements chemically bonded together in a fixed ratio

chemical bond
an interaction between the electron shells of one or more atoms causing them to combine to form a new substance called a compound

molecule
a group of atoms covalently bonded together representing the smallest unit of an element or compound that can take part in a chemical reaction

Although there are 92 naturally occurring elements on Earth, the vast majority do not exist in large quantities in their uncombined state. **Elements** exist more commonly in **compounds** because they form **chemical bonds** with atoms of other elements. In addition, there are a number of elements that exist as **molecules**, rather than single atoms, again due to the formation of chemical bonds between atoms. These bonds form spontaneously, due to the electrons in the valence shell of an atom interacting with the valence shells of other atoms to form a lower energy, more stable state.

The number of electrons in the valence shells of elements can be determined from the periodic table. Elements in the same group have the same number of electrons in the valence shell. The number of valence electrons increases from left to right across the groups of the periodic table. The noble gases (group 18) have full valence shells, so they are chemically inert. Elements in the same group have similar chemical properties because they have the same number of valence electrons.

Atoms tend to form chemical bonds to obtain a filled valence shell and become chemically stable. A chemical bond forms between atoms due to the electrostatic force of attraction between positive (the nucleus of an atom) and negative charges (electrons). A chemical bond forms when electrons from both atoms are attracted to both nuclei.

Forming chemical bonds can involve losing, gaining or sharing electrons. Which of these occurs depends on the position of the element in the periodic table and what element (or elements) it combines with. These conditions also determine the type of chemical bond formed. The properties of the resultant substance (element or compound) are directly related to the type of bonding within the substance.

The three types of chemical bonds that occur due to the transfer or sharing of electrons are:

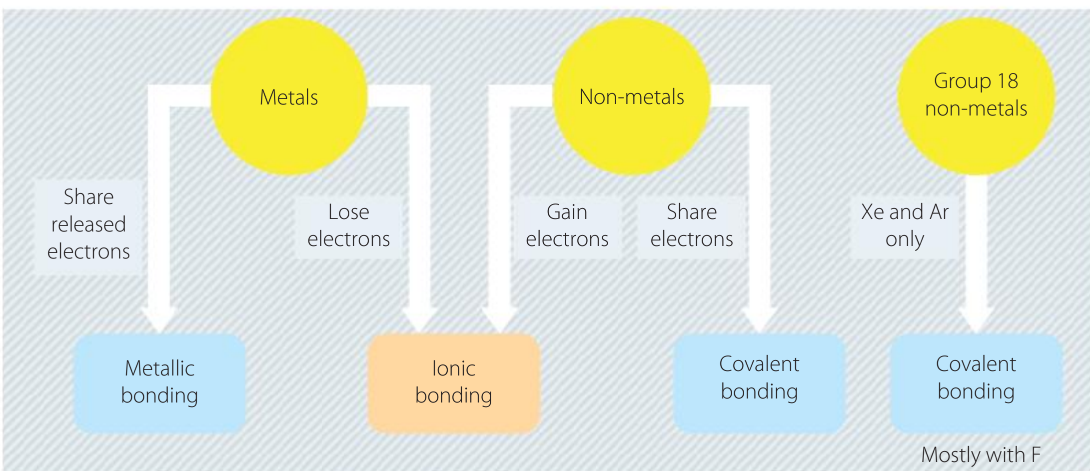
- metallic bonds, which involve the electrostatic attraction between metal ions that have released electrons

  (Note: Compounds do not form when bonds form between atoms of different metallic elements. Rather, an alloy, or amalgam, forms. This is a mixture of the two metals, which possesses properties of each of the metals that comprise it.)
- ionic bonds, which involve the electrostatic attraction between metal atoms that have lost electrons and non-metal atoms that have gained electrons
- covalent bonds, which involve electrostatic attraction between shared electrons and nuclei of non-metal atoms.

Figure 3.1.1 shows the relationship between metals and non-metals, and the type of bonding in which they participate.

3.1.1 Chemical bonding: the nature of the chemical bond
3.1.2 Chemical bonds
3.1.3 Chemical bonding

FIGURE 3.1.1 Types of bonding



SECTION REVIEW 3.1

REMEMBERING

1 Name the three types of chemical bonds that hold atoms together.

2 What determines the type of chemical bond that is formed between atoms?

UNDERSTANDING

3 Outline why chemical bonds form between atoms.

4 Explain why elements in the same group of the periodic table bond in the same way.

5 Identify the type of bonding that would occur between the following elements.

  a Magnesium and chlorine

  b Nitrogen and hydrogen

  c Sodium and mercury

3.2 Metallic bonding

The valence electrons in metal atoms are loosely held, as shown by their low ionisation energies. In order to attain a **noble gas electron configuration**, metal atoms tend to lose electrons, which are negatively charged, to form a positively charged cation.

In a metal, the atoms are surrounded on all sides by other metal atoms, which also tend to lose electrons. The valence electrons are attracted to the nuclei of the other atoms that are present and are drawn into the space between the atoms by the combined action of all attractive forces. The electrons that are detached from their atoms are called **delocalised electrons**. When the electrons break away from their atom, they leave behind a cation. However, the overall piece of metal is uncharged because the total positive charge of the cations is equal to the total negative charge of the delocalised electrons.

The delocalised electrons are free to move randomly between atoms within the three-dimensional arrangement and are often referred to as a 'sea of electrons' (Figure 3.2.1). The electrostatic force of attraction between the positively charged metal ions and the negatively charged delocalised electrons holds the structure together. **Metallic bonds** are non-directional because the electrons are free to move between cations. This movement of electrons explains why metals are able to conduct electricity when in the solid state.

noble gas electron configuration an arrangement of electrons that is identical to one of the elements in group 18, known as the noble gases

delocalised electrons electrons that are not held in a fixed orbit around one nucleus, but are shared by many atoms

metallic bonding the electrostatic forces of attraction that occur between cations and the electrons that are free to move within the regular lattice structure of a metal

3.2.1 Metallic bonding

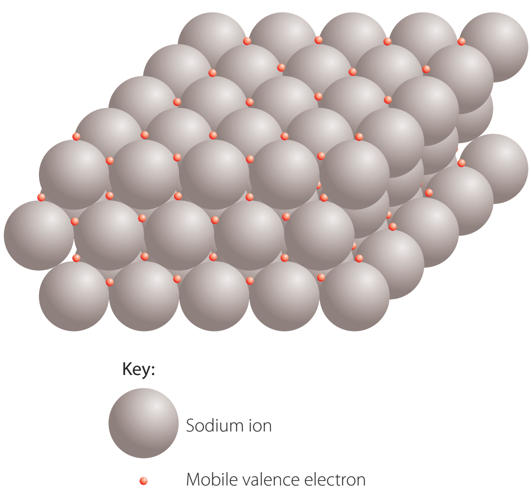


FIGURE 3.2.1 A model of a section of the metallic lattice of sodium showing the sodium cations in a 'sea of electrons'.

SECTION REVIEW

3.2

REMEMBERING

1 Identify the major characteristics of the valency electrons in a metal.

2 Define 'metallic bonding'.

UNDERSTANDING

3 Explain why metals have delocalised electrons.

4 Explain the following observations in terms of the metallic model.

a You can draw out metals into a wire.

b You should not stick a metal fork into a power socket.

c You pick up a metal spoon that has been sitting in a cup of hot water and burn your hand.

ANALYSING

5 Use the information in the table to determine if the position of an element in the periodic table relates to its melting point. Justify your reasoning.

ELEMENT	MELTING POINT (°C)
Sodium	98
Magnesium	650
Aluminium	660
Copper	1085

3.3 Ionic bonding

3.3.1 Ionic bonding

Although metals are common, they are not usually found naturally as pure metals. This is because metal atoms generally tend to lose electrons to obtain a filled outer shell and readily combine with elements that tend to gain electrons to form compounds.

For example, sodium metal is very reactive and is never found naturally in its metallic state. When sodium is brought together with chlorine, it forms the compound sodium chloride (Figure 3.3.1), which is more commonly known as table salt (Figure 3.3.2).

FIGURE 3.3.1 When sodium metal and chlorine gas meet, sodium reacts violently.

FIGURE 3.3.2 Crystals of sodium chloride are completely different in form from the constituent elements sodium and chlorine.

Ionic bonding involves the transfer of electrons. When atoms lose or gain electrons to attain a full outer shell, they form a charged particle called an ion. The **ions** that form on the loss of one or more electrons are positive (+) and are called **cations**. Ions that form on the gain of one or more electrons are negative (−) and are called **anions**.

ion
an atom that has become charged through the gain or loss of an electron

cation
a metallic atom which has become positively charged through losing one or more electrons

anion
a non-metallic atom that has become negatively charged through gaining one or more electrons

Sodium chloride is a combination of a metal and non-metal. The sodium atom with the electron configuration 2,8,1 has a tendency to lose one electron to become like neon (2,8), while the chlorine atom with the configuration 2,8,7 tends to gain one electron to become like argon (2,8,8). The sodium atom becomes a positively charged sodium ion and is represented as Na^+, while the chlorine atom becomes a negatively charged chloride ion and is represented as Cl^-. The resulting compound, sodium chloride (NaCl), is electrically neutral because the positive and negative charges balance each other (Figure 3.3.3).

When ionic compounds form, the total number of electrons lost and the total number of electrons gained must be equal. This means that the total positive charge on the cation(s) must equal the total negative charges on the anion(s), so the formula of an ionic compound will always represent an electrically neutral substance.

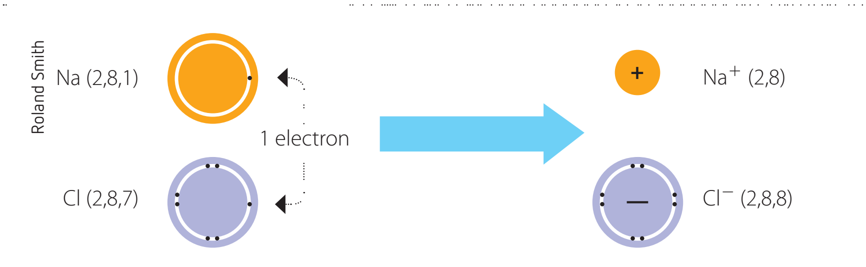


FIGURE 3.3.3 Formation of sodium chloride involving electron transfer

Lewis structure
a representation of the electron arrangement of atoms, using dots or crosses to represent electrons in their shells; Lewis structures are also referred to as Lewis dot structures, electron dot structures or electron dot formulas

In Figure 3.3.3 only the valence electrons in the outer shell of the atoms have been drawn. This reaction can also be represented using **Lewis structures**. These diagrams are often more convenient to use because the valence electrons can be represented as dots (or crosses), as shown in Figure 3.3.4.

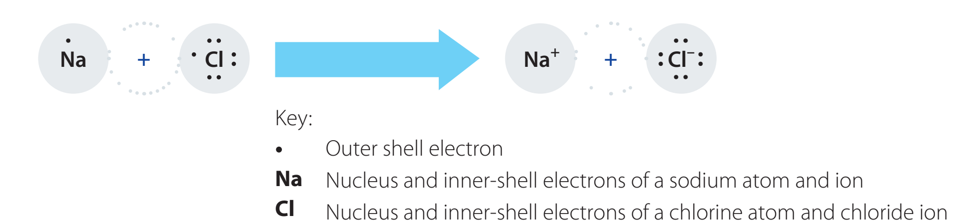


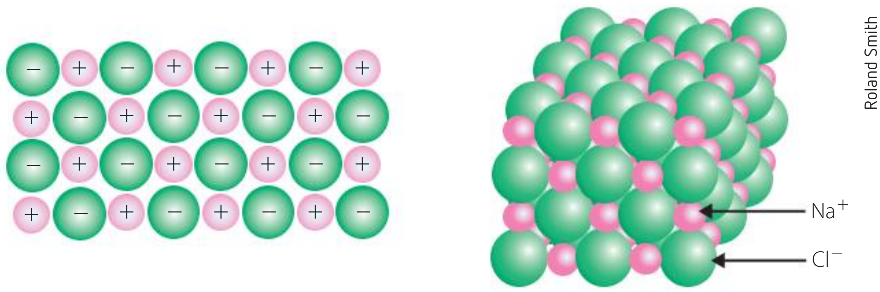
FIGURE 3.3.4 Lewis structures showing the ionic bonding in sodium chloride (NaCl)

There is a strong electrostatic attraction between the positive cations and the negative anions and this attraction results in a chemical bond being formed. This type of chemical bond is called an **ionic bond**. Ionic bonding is due to the electrostatic force of attraction between oppositely charged ions.

ionic bond
the mutual, non-directional, electrostatic attraction of positively and negatively charged ions, which holds them together in a lattice in a compound

Ionic compounds have a crystalline lattice structure in which the positive and negative ions are arranged in an orderly fashion with every positive ion being surrounded by negative ions and every negative ion being surrounded by positive ions (Figure 3.3.5). The electrostatic attraction between pairs of oppositely charged ions extends throughout the whole crystal. There are no separate units of NaCl, just a large array of cations and anions held together by ionic bonds.

FIGURE 3.3.5 An ionic substance consists of an orderly array of positive and negative ions.



The formula represents the ratio of ions, rather than the actual number of ions, because there are no discrete molecules in an ionic compound. A formula that gives the simplest ratio of atoms or ions is called an **empirical formula**. Formulas for ionic compounds are always empirical formulas.

empirical formula a chemical formula showing the simplest ratio of elements in a compound rather than the total number of atoms in the molecule

Ions and the periodic table

For the main group elements, the periodic table can be used to deduce the charge of an ion because the number of electrons an element loses or gains depends on the number of its valence electrons. Many transition metal elements (in groups 3–12) can form multiple ions because of the small energetic difference between the 4s and 3d subshells. This means they can lose both the 's' electrons and a variable number of 'd' electrons (Figure 3.3.6).

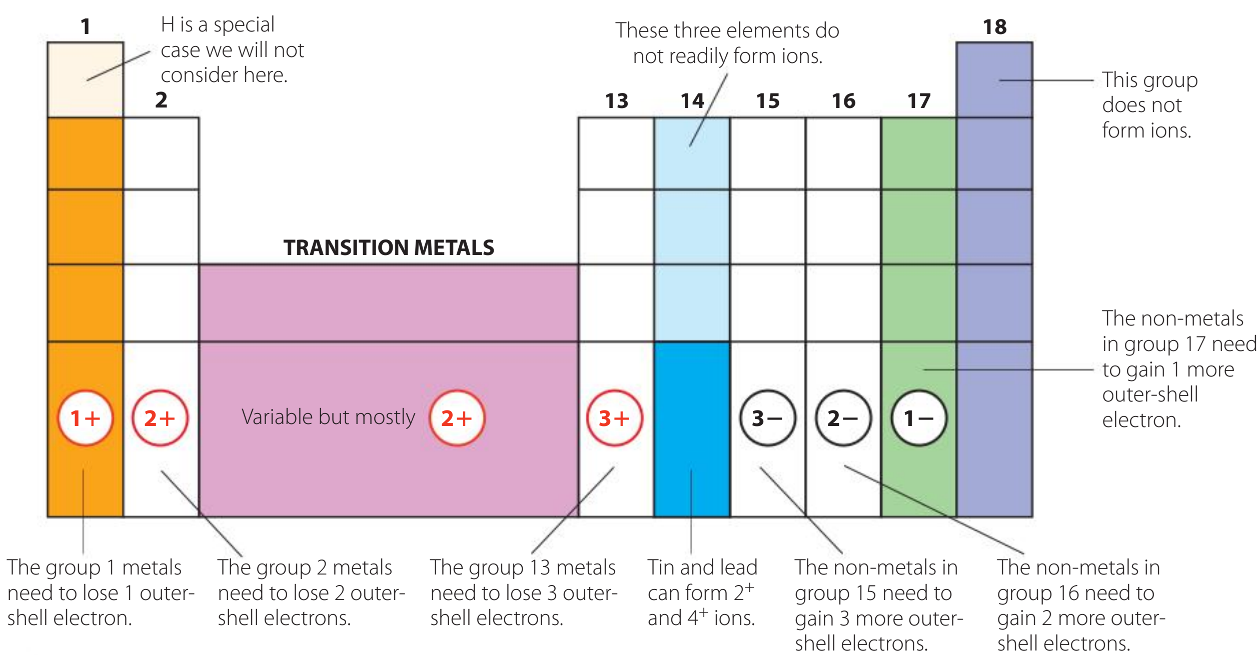


FIGURE 3.3.6 The charges of ions formed from elements according to their group of the periodic table.

Names of ionic compounds

In the names and formulas of ionic compounds composed of only two elements, the positively charged ion is always listed first and the negatively charged ion is listed second. The name of the positive ion is the same as the name of the element from which it was formed. However, in the name of the negative ion, the end of the non-metal parent element is replaced with '–ide'. For example, chlorine becomes chloride, oxygen becomes oxide, and bromine becomes bromide. Accordingly, the names 'sodium chloride', 'magnesium oxide' and 'potassium bromide'.

Transition metals have slightly different electron structures from other metals. As a result, many transition metals show more than one possible charge. A simple naming system is used to show which ion is present. For example, the metal iron (Fe) forms both Fe^{2+} and Fe^{3+} ions. Once, the ending of the name was used to indicate the charge of the ion. The one with a lower charge had '-ous' at the end of its name and the one with a higher charge had '-ic' at the end of its name. But now there is a much easier system in which the charge is shown as a Roman numeral, as shown in Table 3.3.1. Note that there is no gap between the parenthesis and the symbol of the element. The former names are still sometimes used in the chemical industry.

transition metal
the section of the periodic table in groups 3–12 containing metallic elements

3.3.2 Naming ions and ionic bonds

TABLE 3.3.1 Examples of naming transition metal ions

Ion	Fe^{2+}	Fe^{3+}	Cu^{+}	Cu^{2+}
Current naming convention	Iron(II)	Iron(III)	Copper(I)	Copper(II)
Former name	Ferrous	Ferric	Cuprous	Cupric

There are two different ionic compounds formed between iron and chlorine: iron(II) chloride ($FeCl_2$) and iron(III) chloride ($FeCl_3$). The particular compound present will be indicated by its name and the formula. Similarly, copper will also form two compounds with chlorine: copper(I) chloride (CuCl) and copper(II) chloride ($CuCl_2$).

In many ionic compounds, the positive ion or the negative ion (or both) consist of two or more atoms that are strongly bonded together and act as a single entity. These ions are called **polyatomic ions**. Polyatomic ions share their outer shell electrons where the charge on polyatomic ions is spread over the whole ion. As shown in Table 3.3.2, the names of the compounds involving polyatomic ions follow the convention discussed.

polyatomic ion
groups of non-metallic atoms that chemically bond together and have an overall charge, which can be positive or negative: they can take part in ionic bonding

TABLE 3.3.2 Some examples of common polyatomic ions and their charges

NAME OF ION	FORMULA	VALENCY	EXAMPLE OF A COMPOUND
Ammonium	NH_4^{+}	+1	Ammonium chloride
Hydroxide	OH^{-}	–1	Iron(III) hydroxide
Nitrate	NO_3^{-}	–1	Silver nitrate
Sulfate	SO_4^{2-}	–2	Copper(II) sulfate
Carbonate	CO_3^{2-}	–2	Calcium carbonate
Phosphate	PO_4^{3-}	–3	Sodium phosphate

Formulas of ionic compounds

When determining the formulas of ionic compounds, there is one simple rule: ionic compounds have no net charge. This means the total number of positive charges equals the total number of negative charges. When determining the formula, you need to determine the ratio of ions that will achieve this.

WORKED EXAMPLE 3.3.1

Write the formula for calcium sulfide.

ANSWER

1 Determine the ions present:

Calcium sulfide is composed of calcium ions and sulfide ions.

2 Determine the charge on the ions:

The ions present are Ca^{2+} and S^{2-}.

3 Compare the charges and determine the ratio that gives equal positive and negative charges:

These ions have the same-sized charge. So, one sulfide ion balances one calcium ion and the total positive charge equals the total negative charge.

4 Write the formula:

The charges on the ions are not shown in the overall chemical formula.

The formula is CaS.

SECTION REVIEW 3.3

REMEMBERING

1 Define 'ionic bonding'.

2 What is the overall charge on an ionic compound? How is this achieved?

UNDERSTANDING

3 Explain why metals generally form cations and non-metals generally form anions when these two types of substances react together.

4 Write the name and give the formula for the ions of the following elements.

a Bromide
b Sulfur
c Barium
d Potassium
e Nitrogen

5 Explain why the formula of ionic solids is an empirical formula and does not give the actual number of ions present in the compound.

APPLYING

6 Identify the names of the following ionic compounds.

a KI
b $BaCl_2$
c CaH_2
d PbO_2
e Na_3PO_4
f $Zn(NO_3)_2$

7 Write out the formulas for the following ionic compounds.

a Magnesium nitride
b Aluminium oxide
c Mercury(II) sulfide
d Ammonium hydroxide
e Copper(I) carbonate
f Zinc phosphate

8 State the charge that would be expected for an ion formed from an atom with an electron configuration of:

a $1s^2 2s^2 2p^6 3s^2$ b $1s^2 2s^2 2p^3$.

9 Explain why the formula of:

a sodium bromide is NaBr, not Na_2Br or $NaBr_2$

b aluminium fluoride is AlF_3, not AlF or Al_3F or Al_2F_6.

10 Identify the charged species that conduct electric current in:

a molten MgO b molten AgCl c a solution of LiI.

ANALYSING

11 Explain why hydrogen can form both a positive and a negative ion, and give the formula of each ion.

12 X represents different mystery elements. Deduce the charge X that must have in each of the following compounds.

a XOH b XI_2

c XNO_3 d X_2SO_4

e LiX f ZnX_2

g Al_2X_3

3.4 Covalent bonding

Like all elements in the periodic table (except the noble gases), non-metals form bonds to attain a filled valence electron shell. They can achieve this by accepting an electron from a metal. However, non-metal atoms can also share electrons to attain a filled valence shell. The electrostatic force of attraction between the shared electrons and the positively charged nuclei holds the atoms together and is called **covalent bonding**. Covalent bonding is also referred to as 'directional bonding' because there is a direct line between the bonded particles. An example of covalent bonding is shown for a hydrogen molecule in Figure 3.4.1.

Each hydrogen atom has one valence electron so it needs one more to fill its valence shell. When two hydrogen atoms come close enough together, the single electron of each atom will be attracted to both nuclei and drawn into the space between them. The electron shells of the individual atoms begin to merge and the two nuclei start to share the two electrons. Each hydrogen atom can now be considered to have two electrons and a filled valence shell. This shared pair is called a **bonding pair**.

covalent bonding
bonding that occurs between two or more atoms, where electron pairs are shared between the valence shells

bonding pair
a pair of electrons shared between two different atoms

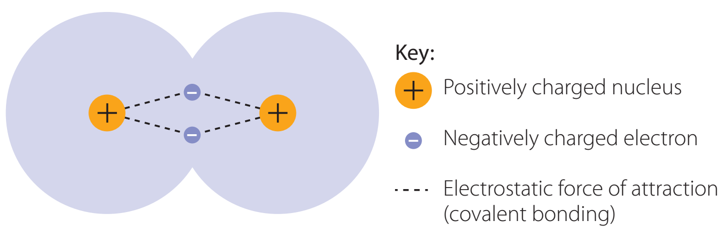


FIGURE 3.4.1 Covalent bonding in a molecule of hydrogen (H_2)

3.4.1 Covalent bonds

Representing covalent bonds

A covalent bond can be represented in different ways (Figure 3.4.2). Fluorine forms the diatomic molecule F_2 by sharing one pair of electrons. A Lewis structure can be used to show shared and unshared electrons. The pairs of unshared valence electrons are called **non-bonding pairs** or **lone pairs**. Another way to show one pair of shared electrons is by a dash. One dash means one shared (bonding pair) of electrons or one covalent bond.

non-bonding or lone pair electrons
a pair of electrons located on an individual atom and that are not shared

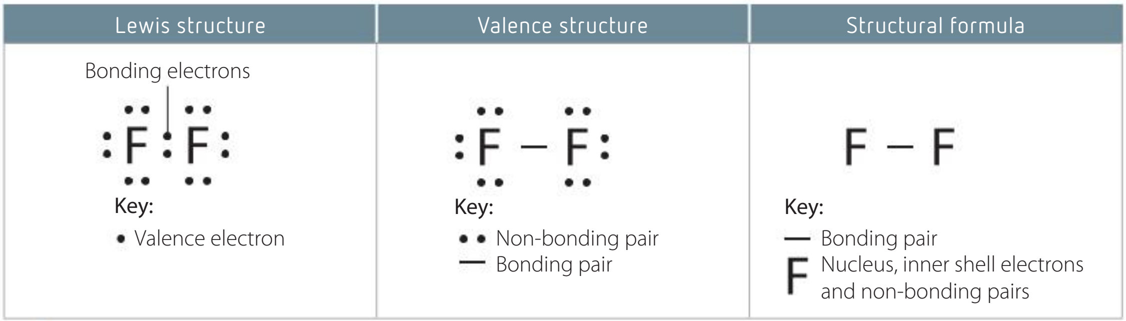


FIGURE 3.4.2 Representing covalent bonds

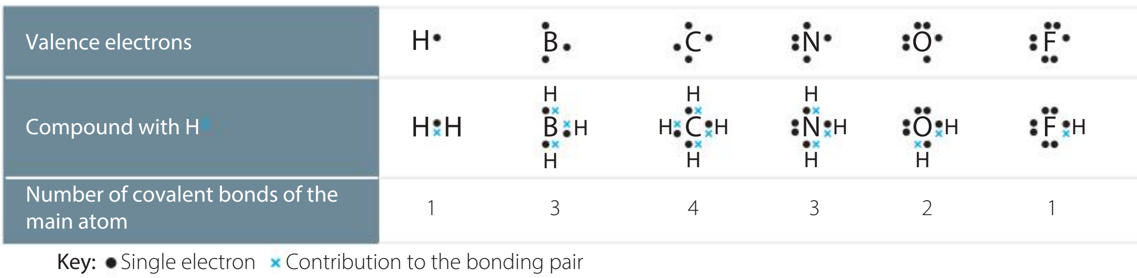
Covalent bonds and the periodic table

Covalent bonding occurs when both elements forming a given substance need to gain electrons to attain a noble gas configuration, such as H_2, H_2O and CO_2.

Hydrogen, metalloids and non-metals, except the noble gases, tend to form covalent bonds. The number of covalent bonds an atom forms is given by the valency of the element as shown in Table 1.2.1, which is directly related to its position in the periodic table. Figure 3.4.3 demonstrates how, in the Lewis structure formulas of some simple molecules, the number of covalent bonds of the main atom corresponds with the number of electrons needed to attain a noble gas configuration.

3.4.2 Single and multiple covalent bonds

FIGURE 3.4.3 Lewis structures of some simple molecules



Multiple covalent bonds

Sometimes, when atoms come together, they share more than one pair of electrons. When this occurs, a **multiple covalent bond** will be formed.

Consider what happens when an oxygen atom bonds to another oxygen to form O_2 or when a nitrogen atom bonds to another nitrogen atom to form N_2.

Each oxygen atom has a valency of 2, so it can form two covalent bonds. When two oxygen atoms come together, two pairs of electrons are drawn into the region between the two nuclei and their outer shells partly merge. This means there are four bonding electrons holding the nuclei closely together so this will be a stronger bond than that formed by one pair. Since two pairs of electrons are shared, the bond is called a **double covalent bond**.

Nitrogen has a valency of 3, which means it can form three covalent bonds. When two nitrogen atoms come together, they share three pairs of electrons and form a **triple covalent bond**. This bond is even stronger than a double bond and is the main reason why nitrogen gas is relatively unreactive because a lot of energy is needed to break the triple bond in a chemical reaction. Figure 3.4.4 shows how double bonds (e.g. in O_2) and triple bonds (e.g. in N_2) can be represented.

multiple covalent bond
two atoms share more than one pair of electrons: a double bond consists of two shared pairs of electrons and a triple bond of three shared pairs

double covalent bond
a double bond consists of two shared pairs of electrons

triple covalent bond
a triple bond consists of three shared pairs of electrons

Molecular formula	Lewis structure	Valence structure	Structural formula
O_2	$\ddot{\underset{\cdot\cdot}{O}}::\ddot{\underset{\cdot\cdot}{O}}$	$\ddot{\underset{\cdot\cdot}{O}}{:}=\ddot{\underset{\cdot\cdot}{O}}{:}$	$O=O$
N_2	$:N\vdots\vdots N:$	$:N\equiv N:$	$N\equiv N$

FIGURE 3.4.4 Ways of representing double and triple covalent bonds

Carbon, in group 14, has a valency of 4 so can form four covalent bonds. It cannot form a quadruple covalent bond because this type of bond cannot exist. A quadruple covalent bond would require four pairs of electrons to be located between the two nuclei. This cannot happen because the repulsive forces between the negative electrons becomes greater than the attractive forces between the electrons and the nuclei, preventing a fourth pair of electrons in the bond. Carbon can achieve its four bonds in many ways, as shown in Figure 3.4.5.

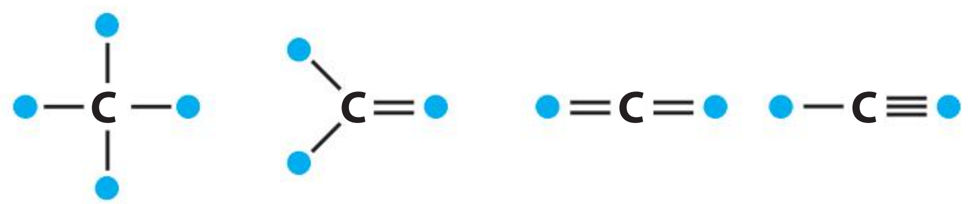


FIGURE 3.4.5 Different ways carbon can gain four electrons

PRACTICAL ACTIVITY 3.4.1

Molecular models

AIM

To use a molecular model kit to construct models of molecular substances and to represent these substances by Lewis structures and structural diagrams

EQUIPMENT PER GROUP

- molecular model kit

PROCEDURE

1. Copy the following table into your notebook. Make sure you leave enough space to draw the diagrams.
2. Use the molecular model kit to construct a model of each of the substances listed.
3. Use your models to draw structural diagrams of the molecules constructed.
4. Draw Lewis structure formula for each of the substances.

»

NAME	FORMULA	STRUCTURAL DIAGRAM	LEWIS STRUCTURE FORMULA
Water	H_2O		
Ammonia	NH_3		
Methane	CH_4		
Dihydrogen sulfide	H_2S		
Nitrogen	N_2		
Oxygen	O_2		
Carbon dioxide	CO_2		
Hydrogen cyanide	HCN		
Ethene	C_2H_4		

QUESTIONS

1 Which of the substances had multiple covalent bonds?

2 What were the different shapes?

3 Compare the shapes of H_2O and H_2S. Suggest a reason for similarities or differences. (Hint: Consider their positions in the periodic table.)

Names and formulas of covalent molecules

As you have already learnt, the valency of an element is an indication of the number of bonds it can form. Generally, this can be deduced from the periodic table as shown in Table 1.2.1. However, as with the transition metals, several non-metallic elements can display more than one valency in covalent compounds. For this reason, the names of covalent compounds need to provide an indication of the valency. The convention for naming covalent compounds differs from the naming of ionic compounds.

The following rules apply to covalent compounds made up of only two elements.

1 Use the element name for the first element and change the end of the name of the second element to 'ide'; for example, hydrogen fluoride.

2 The first named element is the one that is further to the left in the periodic table.

3 If both elements are in the same group, the element lower down the group is named first.

4 An exception to rules 2 and 3 is when oxygen is bonded to Cl, Br or I. In these compounds, oxygen is named last. The rationale is that oxygen is more electronegative than any of the halogens.

5 The number of atoms of each type is given by using the prefixes in front of each part of the name (although 'mono' may be omitted from the first named element). For example, CO is carbon monoxide, CO_2 is carbon dioxide and N_2O_5 is dinitrogen pentoxide.

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TABLE 3.4.1 Naming covalent molecules

PREFIX	MONO	DI	TRI	TETRA	PENTA	HEXA	HEPTA	OCTA	NONA	DECA
Number of atoms	1	2	3	4	5	6	7	8	9	10

Once the name of the compound is given, writing the formula is very straightforward. The order is given by the name and the prefix of the element becomes the subscript for that element in the formula. For example, diphosphorus pentoxide is P_2O_5 and sulfur hexafluoride is SF_6.

Sometimes compounds are referred to by their common names rather than their systematic names. For example, H_2O, dihydrogen monoxide, is called water; NH_3, nitrogen trihydride, is called ammonia; and CH_4, carbon tetrahydride, is called methane.

SECTION REVIEW 3.4

REMEMBERING

1 State the method used to determine which element is written first in a covalent compound.

2 a Define 'covalent bond'.

b Identify how many electrons share in the formation of the following.

i Single bond ii Double bond iii Triple bond

UNDERSTANDING

3 a How many electrons are there in the valence energy level of each carbon, hydrogen and chlorine atom in the molecule CH_3Cl?

b Which noble gas electron configuration does each atom resemble?

4 Determine the molecular formulas for the covalent molecules formed between:

a oxygen and fluorine

b boron and hydrogen.

5 Explain why the formula for methane is CH_4, not CH_3 or CH_5.

APPLYING

6 Write the formulas for:

a sulfur dioxide

b dinitrogen pentoxide

c carbon tetrachloride

d nitrogen trifluoride

e silicon tetrabromide.

7 Identify the following compounds.

a N_2O b NCl_3 c SO_3

d H_2S e N_2O_4

8 Draw Lewis structure formulas and structural diagrams for the following substances.

a Br_2 b HF c NF_3

d CS_2 e $HCHO$

9 Identify the bonding and non-bonding electron pairs in $SiCl_4$.

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 Define the following terms.
  a Covalent bonding
  b Ionic bonding
  c Metallic bonding

2 List each type of bond from Question 1 and represent it with a diagram.

CATEGORY QUESTIONS

3 Classify the bonding in each of the following solids as ionic, metallic or covalent.
  a Magnesium
  b Water
  c Barium chloride
  d Phosphorus triiodide
  e Lithium sulfide

4 Name the following compounds.
  a Ag_2O b $AlCl_3$
  c K_2SO_4 d NH_4Br
  e $Mg(OH)_2$ f $FeCO_3$

5 Write the correct formula for each of the following compounds.
  a Magnesium chloride
  b Calcium carbonate
  c Copper(II) nitrate
  d Ammonium chloride
  e Potassium sulfide
  f Lead(IV) oxide

6 Name the following compounds.
  a PCl_3 b SF_4
  c N_2O_3 d NO
  e Cl_2O_7

7 Write the formula for each of the following compounds.
  a Boron trichloride
  b Phosphorus pentaiodide
  c Hydrogen bromide
  d Dinitrogen tetroxide
  e Silicon dioxide

8 Draw Lewis structures for the following molecules.
  a CaF_2 b HCl
  c Na_2O d NH_3

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ELABORATION QUESTIONS

9 What is the valence of element X in each of the following compounds?

a X_2S_3

b $X(NO_3)_2$

10 What is the valence of Y in each of the following compounds?

a MgY_2

b $(NH_4)_3Y$

EVIDENCE QUESTIONS

11 Explain how our knowledge of atomic structure explains why different elements take part in different types of bonding.

12 When mixed together, metals do not form compounds, such as in covalent and ionic bonding, but form alloys. Carry out research to discover the key differences between a compound and an alloy.

END-OF-CHAPTER EXAM

End-of-chapter test

1 Element Q forms a compound with the formula $Q_2(SO_4)_3$. What is the formula for the compound that Q forms with chlorine gas?

A Q_3Cl_2

B Q_3Cl

C QCl_3

D Q_2Cl_3

2 Atom X has six valence electrons. Atom Y has seven valence electrons. What is the bonding type and most likely formula for the compound when the two atoms react?

	BONDING TYPE	FORMULA
A	Ionic	X_2Y
B	Ionic	XY_2
C	Covalent	X_2Y
D	Covalent	XY_2

3 When the ionic compound Ba_3P_2 is formed:

A six electrons are shared.

B three electrons move from the barium atoms to the phosphorus atoms.

C six electrons move from the phosphorus atoms to the barium atoms.

D three electrons move from the barium atoms to the phosphorus atoms.

4 The element X reacts with oxygen to form the compound X_2O_3, while element Y reacts with hydrogen to form the compound H_2Y.

What is the most likely formula of a stable compound formed by element X and element Y?

A XY

B X_2Y

C X_2Y_3

D X_3Y_2

5 The formulas (electrovalencies) of some common ions are shown in the following table.

NAME	FORMULA	NAME	FORMULA	NAME	FORMULA	NAME	FORMULA
Sulfate	SO_4^{2-}	Iron	Fe^{3+}	Sulfide	S^{2-}	Phosphate	PO_4^{3-}
Calcium	Ca^{2+}	Iodide	I^-	Lithium	Li^+	Barium	Ba^{2+}

a Select from the ions listed in the table to answer the following questions.

i Write a formula for an ionic compound that will have the formula XY_3.

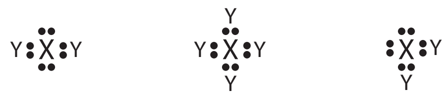
ii Write a formula for an ionic compound that will have the formula X_3Y_2.

iii Write the formulas of two ionic compounds that have the formula XY.

iv Name two ions that have the same electron configuration.

b The element calcium has an electron configuration of $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$. Its melting point is 842°C. Draw and label a diagram to explain the structure of solid calcium.

6 Three Lewis structures are shown in Figure 3.4.6, below.



Compound A Compound B Compound C

FIGURE 3.4.6 Lewis structures of three compounds

Copy the following table into your notebook and complete it by answering the questions that refer to the three structures shown in Figure 3.4.6.

	COMPOUND A	COMPOUND B	COMPOUND C
To which group in the periodic table does X belong?			
Suggest an identity for X.			
Suggest a possible molecule with this structure.			

4 ISOTOPES

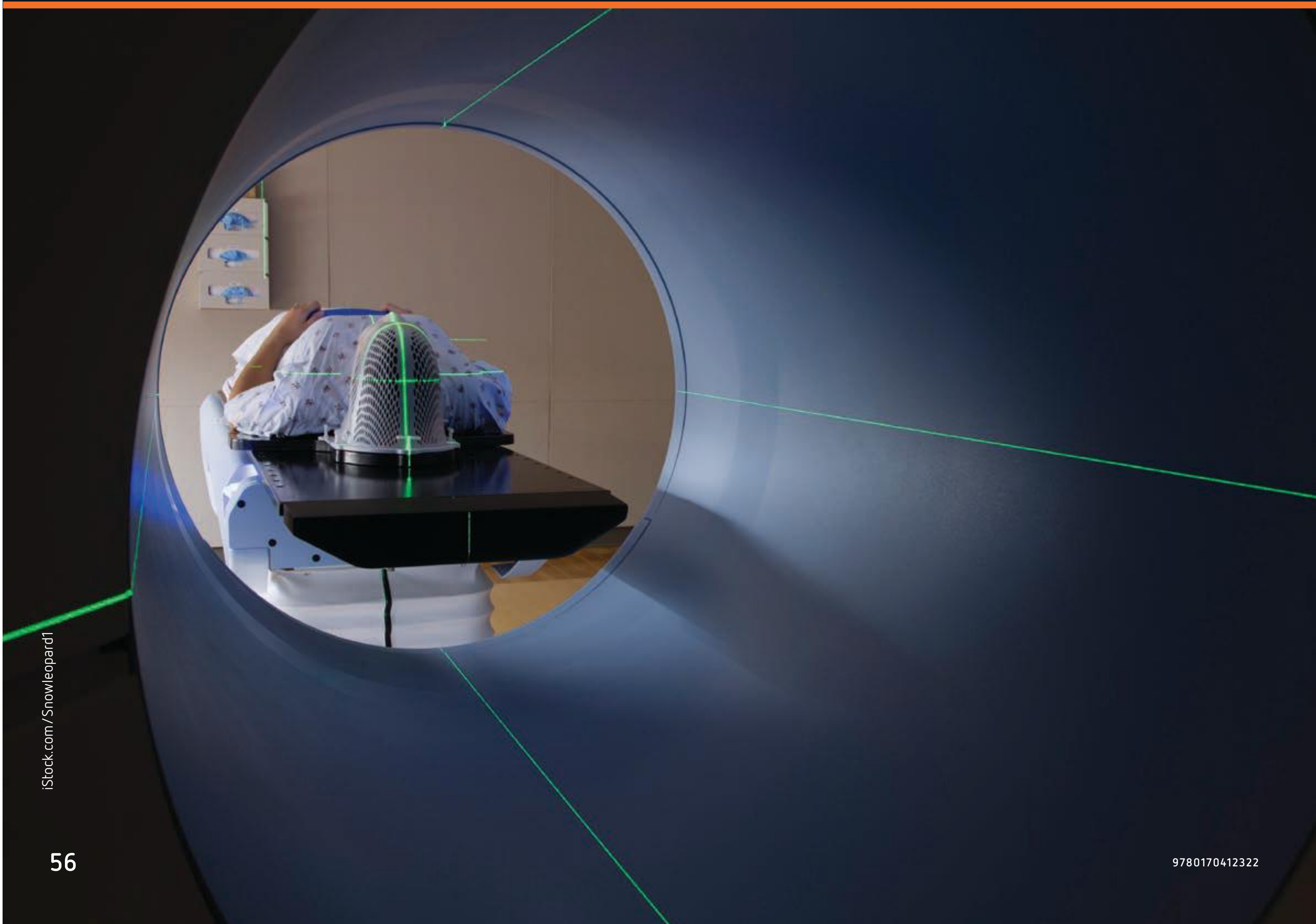
Introduction

Some remarkable techniques in a number of different fields, such as the detection of medical conditions and the dating of ancient objects, are the result of the existence of isotopes.

Stimulus questions

Are all atoms in an element identical?

Why is the chemistry of two isotopes the same, yet they have some different properties such as radioactivity?



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4.1 Different forms of the same element

Isotopes are atoms of the same element that have the same number of protons but different numbers of neutrons in their nuclei. This means that two isotopes of the same element will have the same atomic number, but different mass numbers. Isotopes of the same element show some similarities and some differences in their properties and their relative abundance on Earth is different. Isotopes of different elements can be naturally occurring or synthesised for a specific purpose.

Isotopes are named by writing the name of the element, followed by a hyphen and the mass number of the isotope. Carbon has a number of isotopes, all containing six protons but different numbers of neutrons. Carbon-12 ($^{12}_{6}C$) has six neutrons, while carbon-14 ($^{14}_{6}C$) has eight neutrons. Different elements have different numbers of naturally occurring isotopes and some, such as fluorine, have only one.

isotope
two atoms of the same element with the same number of protons but with different masses, resulting from their different number of neutrons

4.1.1 Isotopes

TABLE 4.1.1 Numbers of protons, neutrons and electrons in isotopes

ISOTOPE NAME	CARBON-12	CARBON-13	CARBON-14
Isotope symbol	$^{12}_{6}C$	$^{13}_{6}C$	$^{14}_{6}C$
Number of protons	6	6	6
Number of neutrons	6	7	8
Number of electrons	6	6	6
Isotope name	Uranium-234	Uranium-235	Uranium-238
Isotope symbol	$^{234}_{92}U$	$^{235}_{92}U$	$^{238}_{92}U$
Number of protons	92	92	92
Number of neutrons	142	143	146
Number of electrons	92	92	92

Chemical properties

Isotopes of the same element will have very similar **chemical properties**. Chemical properties relate to how an element participates in chemical reactions. When an atom reacts in a chemical reaction, its behaviour is caused by the arrangement and number of electrons. Isotopes of the same element have the same number and arrangement of electrons so their chemical properties are similar. For example, water (H_2O) is formed in exactly the same reaction, no matter whether the isotope of hydrogen used is ^{1}H or ^{2}H.

chemical properties
properties of a substance, such as acidity and flammability, that relate to the chemical reactions in which the substance takes part

Physical properties

Isotopes of the same element can have different **physical properties**. These properties are characteristics that you can observe or measure, such as colour, density and mass. Properties can vary because isotopes have slightly different masses due to the different numbers of neutrons. For example, the masses of identical amounts of helium-3 atoms and helium-4 atoms will be different. In addition, because of the differing numbers of neutrons, some isotopes of certain elements are radioactive while others are not.

Isotopes of an element have similar chemical properties but different physical properties, including variations in nuclear stability.

physical properties
properties of a substance, such as melting point and electrical conductivity, that do not involve a chemical reaction

SECTION REVIEW 4.1

REMEMBERING

1 Define 'isotope'.

UNDERSTANDING

2 Explain the difference in numbers of subatomic particles between the isotopes nitrogen-14 and nitrogen-13.

3 Explain why isotopes of the same element have the same chemical properties.

4.2 Radioactive isotopes

radioactive decay the process by which an unstable nucleus breaks down, releasing energy and matter

half-life the time taken for half of the number of radioactive nuclei in a sample to undergo decay

4.2.1 Radioisotopes

Some isotopes of elements are stable because the attractive and repulsive forces in the nucleus are balanced. Other isotopes have unstable nuclei in which the forces are not balanced. When an isotope has an unstable nucleus, the nucleus will undergo **radioactive decay** to become stable. During radioactive decay, high-energy particles or radiation is emitted, which can be used for purposes such as radioactive dating, medical diagnoses and medical treatments.

Most elements have at least two isotopes; some are stable and some are unstable. Both types of isotopes can be useful; however, unstable isotopes have particular properties that can be used and manipulated for a variety of purposes. For example, carbon-12 and carbon-13 are stable and do not undergo radioactive decay. Carbon-14 is unstable and undergoes decay when a neutron in the nucleus converts into a proton. This decay of carbon-14 nuclei is utilised in radiocarbon dating, a process of determining the age of a material that was once living. Another product of this conversion is an electron, which is emitted from the nucleus. This process is called beta decay. The electron emitted is known as a beta particle.

The **half-life** of a radioactive element is the time taken for half of the atoms in a sample to decay to a new element or isotope. The half-lives of these elements can vary from a few milliseconds to a year. The half-life of a particular element is constant and does not change. For example, an element may have a half-life of 1 minute. This means that, after 1 minute, half of the atoms in a sample of this element would have undergone radioactive decay. One minute later, half of the remaining atoms would have decayed. After another minute, half of the remaining atoms would have decayed. If there were 100 atoms at the start, then after 1 minute there would be 50 atoms, and after 2 minutes there would be 25 atoms, and so on. Figure 4.2.1 is a typical half-life graph.

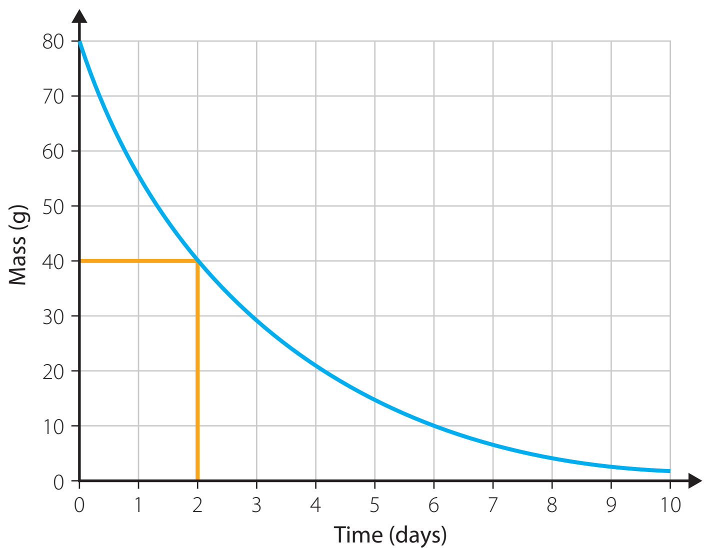

FIGURE 4.2.1 Graph of the half-life for the radioactive decay of an element. The half-life is when half of the original sample has decayed, in this case 2 days.

PRACTICAL ACTIVITY 4.2.1

Determining the half-life of an isotope

AIM

To calculate the half-life of an element

EQUIPMENT PER GROUP

- 100 M&M's® or 100 small discs where one side can be easily distinguished from another

  Note: If you use M&M's®, perform this activity outside the laboratory.
- plastic bag

PROCEDURE

1. Place the 100 M&M's® or discs into a bag and mix them thoroughly.
2. Pour out the M&M's® or discs over a clean piece of paper.
3. Remove the M&M's® that have the logo face down (or the discs with the blank face down). These have 'decayed'.
4. Count the M&M's® that have the logo face up (or the discs with the blank face up). These are still 'radioactive'. Copy the following results table and in the second column record the number of M&M's® with the face up.

TRIAL NUMBER	'ATOMS' REMAINING	COMBINED CLASS 'ATOMS' REMAINING
1		
2		
3		
4		

5. Return the remaining 'radioactive' M&M's® or discs to the bag and mix them thoroughly.
6. Repeat steps 3–6 until you have no more 'radioactive' M&M's® or discs.
7. Record the combined 'class' results in the third column of the results table.

QUESTIONS

1. On a graph, plot the trial number on the horizontal axis and the 'radioactive' atoms remaining on the vertical axis. Make sure you start your graph at zero. This is your half-life graph.
2. To find the half-life of your M&M's® or discs, draw a horizontal line from a point on the vertical axis that represents half of the original number of M&M's® or discs to where this line meets your graph line. Then draw a vertical line down to the horizontal axis. The place it cuts the horizontal axis is the half-life.
3. How many trials did it take for half the M&M's® or discs to 'decay' to half the original number?
4. Did half of the M&M's® or discs decay in one trial? Propose reasons why or why not.

Radiocarbon dating is a technique that measures the proportion of carbon-14, a radioactive isotope of carbon, in a piece of organic material that was once alive, to determine how old the material is. Conduct research into radiocarbon dating and explain how the age of an ancient object can be determined.

SCIENCE AS A HUMAN ENDEAVOUR

FIGURE 4.2.2 This image, produced by the radiation emitted by technetium-99, shows thyroid cancer that has spread (metastasised).

Uses of isotopes

In medicine, radioactive isotopes are used in two ways – to diagnose disease and to treat specific illnesses. Diagnosis of disease uses the radiation that is emitted from an unstable isotope. When an isotope is ingested by a patient, the radiation it emits forms a map of where it ends up in the body, revealing areas of unusual structure or activity, which can be an indicator of disease. The isotope can be used on its own or be attached to a molecule such as glucose that the body can use.

One of the most common radioactive isotopes used for diagnosis is technetium-99. This isotope has a short half-life of about 6 hours, which means technetium-99 breaks down very quickly and does not remain in the body. It is estimated that every year more than 20 million medical procedures use technetium-99. It has many uses, including imaging organs such as the kidneys, liver, lungs and brain. The isotope attaches to a specific pharmaceutical (a compound used as a medical drug) and is then transported to the required place in the body, where it decays, emitting radiation that is detected and measured. This is usually analysed by computer to produce images to help doctors diagnose their patients (Figure 4.2.2).

Radioactive isotopes are used in the treatment of diseases, especially cancer. Isotopes that emit radiation are ingested as a drink or in tablets. Cancer cells divide very rapidly so they absorb the radioisotope at high levels. The radiation emitted by the isotope as it decays damages the cells. The most common isotope used in this way is iodine-131. When iodine-131 is ingested, it concentrates almost exclusively in the thyroid. It is used to treat thyroid cancer because the radiation emitted will target cancer cells in that area.

4.2.2 Using radioisotopes

Mineral nutrition

The human body is made up of mostly organic materials, which contain carbon, hydrogen and often oxygen. The presence of other elements such as calcium, sodium, zinc and iron in different forms is important to keep your body functioning. Some people, especially children, have problems absorbing minerals from the food they eat. To diagnose this problem, isotopes are used to track how minerals move through the human body.

When children, pregnant women or older adults are the target group for treatment, radioactive isotopes are not recommended because the radiation that is emitted is particularly harmful for these people. Instead, enriched stable isotopes, which do not emit radiation, are used.

Calcium is an essential element for bone and teeth formation. The most common isotope of calcium is calcium-40. The heavier isotope calcium-42 is also usually present in very small amounts in any sample of calcium. Calcium-42 is not radioactive. Sometimes, children do not absorb calcium properly, so doctors measure the uptake of calcium by administering a sample of calcium enriched with calcium-42, which is not harmful but is detectable. Doctors measure the amount of the isotope that passes through the child's body and then calculate the amount of calcium absorbed. This allows doctors to identify whether that child has a problem absorbing minerals.

Tagging chemicals in reactions

Chemical reactions inside the human body can be examined by placing isotopic tracers into molecules. For example, a hydrogen-1 atom in a water molecule can be replaced with a hydrogen-2 atom. Hydrogen-2 is known as 'deuterium' or 'heavy hydrogen'. The heavy water molecule behaves in the same way as other water molecules so it is used by the human body in the usual manner. Hydrogen-2 can then be tracked through the human body and the chemical reactions that occur can be investigated.

Chemical reactions in industry or in a laboratory can be studied by the same methods. Many chemical reactions have been known about and even used in industry for years but not understood completely. When a specific atom in a reactant is tagged with an isotope, the path that atom takes during a chemical reaction can be followed and recorded. This helps chemists to understand the mechanisms of a reaction. When a reaction is understood properly, it can be manipulated. This allows chemists to make more of a particular product, or to alter a reaction to get a product that may not otherwise form as readily. For example, particular proteins can be tracked during protein synthesis to produce artificial proteins that can then meet specific requirements of the chemist.

Safety of isotopes

Most radioactive isotopes do not pose significant risk because the particles and radiation they emit either are present in very small amounts or they do not travel very far in air. Most people would not normally come into contact with isotopes that emit significant levels of dangerous radiation, apart from when undergoing certain medical procedures. People who work near radioactive isotopes, such as medical professionals, take appropriate precautions to limit their exposure through protective materials and secure storage. There are also restrictions on the number of procedures they can carry out each year, which further limits the amount of radiation to which they are exposed.

The risk of damage from radioactive isotopes increases with the length of exposure. Radioactive isotopes that are used on or around humans normally have very short half-lives. This means that the total amount of radiation to which humans are exposed is usually low because the isotopes decay to minimal amounts very quickly.

TABLE 4.2.1 Half-lives of some common radioactive isotopes used in medicine

ISOTOPE	USE	HALF-LIFE
Technetium-99	Medical imaging of internal structures	6.01 hours
Iodine-131	Treatment of thyroid problems	8.02 days
Thallium-201	Testing for stress in patients with heart problems	3.04 days
Fluorine-18	Detection of cancer, through attachment to glucose molecules	109.8 minutes

4.2.3 Seven things to know about radioisotopes

SECTION REVIEW 4.2

UNDERSTANDING

1. Explain why some isotopes are stable and some are unstable.
2. Use your understanding of isotopes to explain why they are useful in a range of contexts.

4.3 Relative atomic mass

An atom is extremely small, so it is difficult to measure the mass of one atom. Instead, scientists use the term **relative atomic mass** (A_r). The word 'relative' means that the mass of an atom is being compared with the mass of another atom that scientists use as a standard.

Scientists arbitrarily picked carbon-12 as the standard for comparison to calculate relative atomic mass. Hydrogen and oxygen were considered and even used for a time, but carbon was selected for ease of use in experiments and because it is a very common element on Earth. The relative atomic mass of an atom is the mass of the atom relative to the mass of a carbon-12 atom. The mass of a carbon-12 atom is considered to be exactly 12 and all others are compared to this.

relative atomic mass
the average mass of one atom of an element, relative to the mass of carbon-12

4.3.1 Relative atomic mass

Magnesium-24 has twice the mass of carbon-12 so has a mass of $2 \times 12 = 24$. Hydrogen-1 is one-twelfth the mass of carbon-12, so it has a mass of $\frac{1}{12} \times 12 = 1$. A visual representation of atomic mass comparisons for He-4, C-12 and Ti-48 are shown in Figure 4.3.1.

FIGURE 4.3.1 Relative atomic mass of the elements carbon, helium and titanium

Isotopes of the same element have different atomic masses. They also have different abundances on Earth, which means they are found in different amounts. For example, two of carbon's isotopes are carbon-12 and carbon-13. On Earth, 98.9% of all carbon is carbon-12 and only 1.1% is carbon-13. Other isotopes of carbon, such as carbon-14, are present but in negligible levels.

In the periodic table, the relative atomic mass for each element is displayed, rather than the mass number. As this quantity is an average of the masses of each isotope, it is not a whole number. The equation used to calculate relative atomic mass is as follows:

$$\text{Relative atomic mass} = \frac{(\text{abundance percentage} \times \text{atomic mass}) + (\text{abundance percentage} \times \text{atomic mass})}{100}$$

In the case of carbon, using the information in the previous equation, this becomes:

$$= \frac{(98.9 \times 12) + (1.1 \times 13)}{100}$$

$$= 12.01$$

Worked example 4.3.1 shows you how to calculate the relative atomic mass seen in the periodic table for any element. Worked example 4.3.2 shows you how to calculate the percentage abundance for an isotope if you are provided with the relative atomic mass.

WORKED EXAMPLE 4.3.1

Chlorine has two isotopes with relative atomic masses of 35 and 37, respectively. The percentages of each isotope present on Earth are 75% for chlorine-35 and 25% for chlorine-37.

QUESTION

What is chlorine's overall relative atomic mass?

ANSWER

1 Write out the correct formula to use:

$$\text{Relative atomic mass of chlorine} = \frac{(\text{abundance percentage} \times \text{atomic mass}) + (\text{abundance percentage} \times \text{atomic mass})}{100}$$

2 Insert amounts into the formula and calculate the answer:

$$\text{Relative atomic mass of chlorine} = \frac{(75 \times 35) + (25 \times 37)}{100}$$

$$= 35.5$$

Chlorine's overall relative atomic mass is 35.5%.

WORKED EXAMPLE 4.3.2

The element chlorine has two isotopic forms: chlorine-35 and chlorine-37.

QUESTION

If chlorine's relative atomic mass is 35.5, what is the percentage abundance of each isotope?

ANSWER

1 Write out the correct formula to use:

$$\text{Relative atomic mass} = \frac{(\text{abundance percentage} \times \text{atomic mass}) + (\text{abundance percentage} \times \text{atomic mass})}{100}$$

2 If the abundance of chlorine-35 is x, then the abundance of chlorine-37 is $(100 - x)$, because the total of the abundances must be 100. So:

$$35.5 = \frac{(x \times 35) + ((100 - x) \times 37)}{100}$$

3 Simplify the expression:

$$\begin{aligned} 3550 &= (35x + 3700 - 37x) \\ 2x &= (3700 - 3550) \\ x &= 75 \end{aligned}$$

Chlorine-35's percentage abundance is 75% and chlorine 37's percentage abundance is 25%.

SECTION REVIEW 4.3

REMEMBERING

1 Define 'relative atomic mass'.

UNDERSTANDING

2 Explain how the relative atomic mass shown in the periodic table is calculated.

APPLYING

3 A given element has a mass three times that of carbon. Calculate its relative atomic mass.

4 Copper has two isotopes: copper-63 with an abundance of 69.17% and copper-65 with an abundance of 30.83%. Calculate copper's relative atomic mass.

5 Vanadium (V) has two naturally occurring isotopes, vanadium-50 and vanadium-51. Given that the mass number for vanadium in the periodic table is 50.94, predict which isotope is more abundant and explain your reasoning.

6 Calculate the exact abundances of the vanadium isotopes in Question 5.

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 Define the following terms.
- **a** Half-life
- **b** Isotope
- **c** Radioactive
- **d** Relative atomic mass

CATEGORY QUESTIONS

2 The table below provides data about the number of subatomic particles in a series of different atoms. Copy the table below and complete the blanks, and from the information, identify which atoms are isotopes of one another.

ATOM	ATOMIC NUMBER	MASS NUMBER	NUMBER OF PROTONS	NUMBER OF NEUTRONS	NUMBER OF ELECTRONS
A	6	15		9	
B		15		8	7
C	12	25			
D		12	6		
E			11	14	
F			12	16	12
G		22		11	

ELABORATION QUESTIONS

3 Silicon has three isotopes, as shown in the following table.

MASS NUMBER	ABUNDANCE (%)
28	92.23
29	4.68
30	3.09

Calculate the relative atomic mass of silicon.

4 Bromine has two stable isotopes, bromine-79 and bromine-81. If the relative atomic mass of bromine is 79.90, calculate the percentage abundance of each isotope.

EVIDENCE QUESTIONS

5 Conduct some research to discover why radioisotope labelling is a valuable technique in understanding biological processes. Describe examples additional to those listed in this book of where radioisotope labelling has been particularly important.

END-OF-CHAPTER EXAM

End-of-chapter test

1 The element hydrogen consists of three isotopes: protium (^{1}H), deuterium (^{2}H) and tritium (^{3}H). If the relative atomic mass of hydrogen is 1.01, which isotope is the most abundant?

A Protium

B Deuterium

C Tritium

D It is not possible to determine which isotope is most abundant without examining the mass spectrum of hydrogen.

2 Data for an element is given in the table below.

RELATIVE ISOTOPIC MASS	63.9	65.9	66.9	67.9
PERCENTAGE ABUNDANCE	48.9	27.8	4.10	19.2

What is the relative atomic mass of the element closest to, based on the data provided?

A 65.2

B 65.3

C 65.4

D 65.5

3 The masses of three of the four isotopes of lead is shown in the following table.

ISOTOPE	MASS NUMBER	ABUNDANCE (%)
Lead-204	204	4
Lead-206	206	24
Lead-207	207	21

If the relative atomic mass of lead is 207.2, which of the following is the mass number and abundance of the remaining isotope of lead?

	MASS NUMBER	ABUNDANCE (%)
A	205	8
B	205	51
C	208	8
D	208	51

4 **a** Define 'isotope'.

b Explain how the properties of isotopes cause them to have useful applications in a number of fields.

c Explain how isotopes are used in one of these fields and briefly outline how the physical and chemical properties of the isotopes are critical to their application and use.

5 ANALYTICAL TECHNIQUES

Introduction

A variety of machines (analytical instrumentation) help scientists to determine: what kind of elements or molecular compounds are present in a solid, liquid or gas; how much of each element or compound is present; and the physical and chemical structure of the material. Analytical instruments need to have high sensitivity, low limits of detection, good selectivity, good time resolution, accuracy and reproducibility. The analytical information provided by such instrumentation, often from field-based devices, provides scientists with valuable information in areas such as: atmospheric and oceanographic testing; environmental and production analysis for agriculture; mining; metallurgy; chemical analysis for pharmaceutical and scientific industries; and forensic investigation.

Stimulus questions

- Why do atoms absorb and emit light?
- How can scientists separate and analyse a series of different atoms?
- What is the difference between quantitative and qualitative analysis?

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5.1 Mass spectrometry

Mass spectrometry is a technique that can be used to determine the specific identity and amount of atoms or molecules in a sample of material and, in particular, can determine the **isotopic composition** of an element. An isotopic composition tells you which isotopes are present in an element and the percentage of each isotope present.

Mass spectrometry is carried out using a mass spectrometer (Figure 5.1.1). A mass spectrometer is similar to a long tube, where a mixture of particles is separated according to their mass and then the number of particles of each different mass are counted. This involves five stages.

1 Vaporisation: the sample is vaporised into a gaseous form that can move easily through the mass spectrometer.

2 Ionisation: the sample is bombarded with high-energy electrons (or laser radiation). This knocks out or removes electrons from the atoms or molecules, leaving the atom with an overall positive charge. With electron bombardment, most atoms or molecules can be turned into positively charged ions by this method, even those that would not normally become positively charged, such as argon.

   For example, a sample of carbon that contains two isotopes, carbon-12 and carbon-13, will show two positive ions, each with a different mass. They will all have 6 protons because they are isotopes, but will have either 6 neutrons (carbon-12) or 7 neutrons (carbon-13). Ultimately, the mass : charge ratio will differ for each isotope.

3 Acceleration: the sample's positive ions are passed through an electric field. This causes the positive ions to be accelerated. Each ion's level of acceleration depends on the ion's mass. Smaller particles will be accelerated more than larger particles (assuming the charge on each ion is the same), so particles of different masses will be moving at different speeds.

4 Deflection: with a magnetic sector version of the mass spectrometer, particles are passed through a magnetic field. When a charged particle passes through a magnetic field, it experiences a force,

mass spectrometry
a technique used to determine the atomic or molecular composition of an element and its isotopic composition

isotopic composition
a list of the different isotopes present in a sample of an element and the relative amount of each isotope that is present

5.1.1 How does mass spectrometry work?

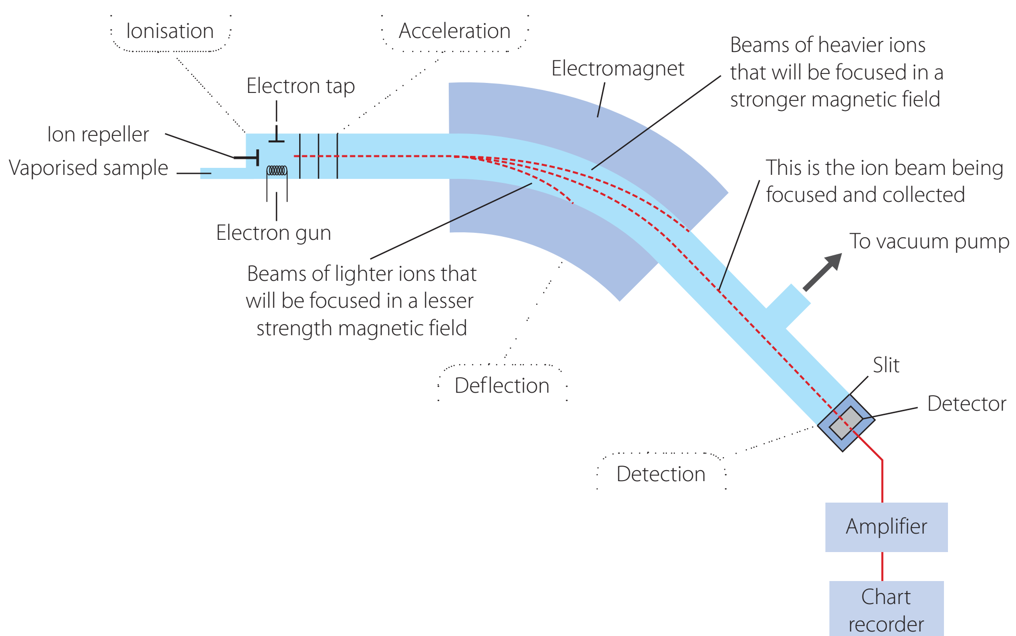

FIGURE 5.1.1 The operation of a mass spectrometer

mass spectrum
a graph of the information from a mass spectrometer, showing the number of particles of different mass (or more precisely, mass : charge ratio) that are present in a sample and the relative abundance of each

which changes its direction. How much an ion's path is deflected depends on the mass : charge ratio of the moving ion. The deflected ions move in an arc whose radius is inversely proportional to the mass : charge ratio of the ion. If all the ions were carrying a single positive charge, the lighter ions would be deflected more by the magnetic field and the heavier ions would be deflected less (Figure 5.1.1). The strength of the magnetic field varies during the process so that, at a particular field strength, an ion with just the right mass : charge ratio will travel on the right path to pass through the slit at the end of the curved tube. Gradually increasing the magnetic field strength ensures that all of the different mass : charge ratio particles will eventually pass through the slit to the detector.

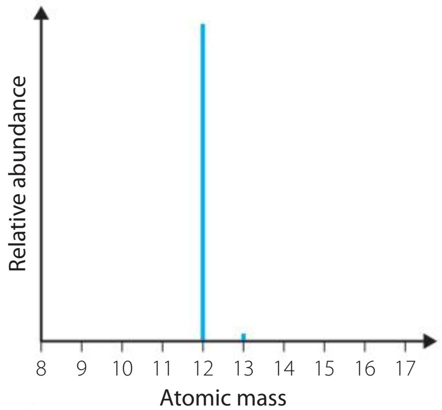

FIGURE 5.1.2 Mass spectrum of carbon

5 Detection: the detector is a negatively charged plate. When a positive ion hits the plate, there is a burst of electrical current, which is measured on a graph. The greater the number of particles, the greater the burst of current. The number of particles hitting the detector at any instant can be measured easily.

Information from the mass spectrometer is transformed into a graph called a **mass spectrum**. This graph shows the mass : charge ratio of the ions that are present and their relative abundance. If the sample consists of a mixture of isotopes – atoms of the same element with different masses – then each of the isotopes will have a different mass : charge ratio and will appear as a different peak on the mass spectrum. Assuming that the charge on each ion is 1, then the mass : charge ratio can be read as the mass number.

Figure 5.1.2 shows an ion with an atomic mass of 12 present in the greatest concentration, and a small amount of an ion with an atomic mass of 13. These are the two isotopes of carbon present in the sample.

WORKED EXAMPLE 5.1.1

QUESTIONS

1 Determine the number of isotopes of molybdenum and their relative atomic masses.

2 Which isotope is present in the highest concentration?

3 Which isotopes are present in equal amounts?

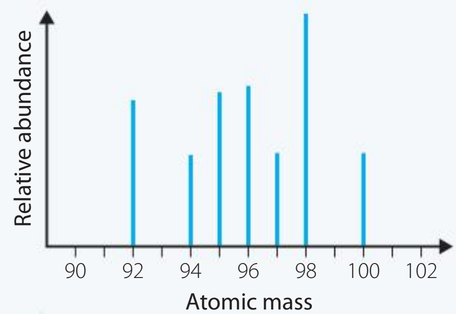

FIGURE 5.1.3 Mass spectrum of molybdenum

ANSWERS

1 There are seven peaks in the mass spectrum at 92, 94, 95, 96, 97, 98 and 100, so there are seven isotopes of molybdenum. They have relative atomic masses of 92, 94, 95, 96, 97, 98 and 100.

2 The isotope with a mass of 98 has the highest concentration in the sample. This isotope has the highest peak in the mass spectrum.

3 The isotopes at 94, 97 and 100 are present in equal amounts. Their peaks are the same height.

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SECTION REVIEW 5.1

REMEMBERING

1 Identify the name of the graph generated during mass spectroscopy.

UNDERSTANDING

2 Explain why isotopes of different mass can be separated by a mass spectrometer.

ANALYSING

3 Figure 5.1.4 shows the mass spectrum of an element.

a Identify the isotopes present in this element.

b Create a table showing the isotopes of the element and their relative abundances.

c Calculate the relative atomic mass of this element and use this information to identify the element.

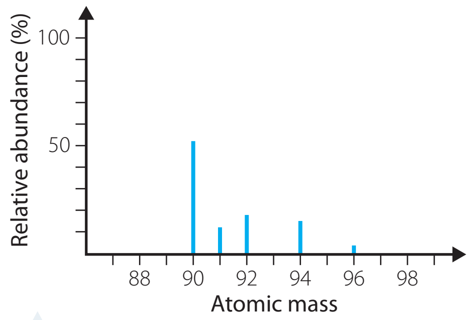


FIGURE 5.1.4 Mass spectrum of an element

5.2 Absorption spectroscopy

Spectroscopy is the process of analysing light. A spectroscope or spectrometer is a device used to separate light into its component wavelengths.

Within an atom, electrons are found in a specific arrangement around the nucleus. When all the electrons are at their lowest possible energy levels, the atom is in the **ground state**. An atom may gain energy as a result of:

- colliding with electrons – electric discharge
- electromagnetic radiation, such as light
- heat – collisions with particles of high kinetic energy.

When an atom absorbs energy, the electrons in the energy levels around the nucleus gain this extra energy and move to a higher energy level. An electron may rise by one or more levels, but cannot go part of the way between levels (Figure 5.2.1).

spectroscopy
a method of analysis performed by subjecting a sample to electromagnetic radiation and observing the frequencies absorbed or emitted

ground state
the lowest energy electron configuration of an atom

5.2.1 What is absorption spectroscopy?

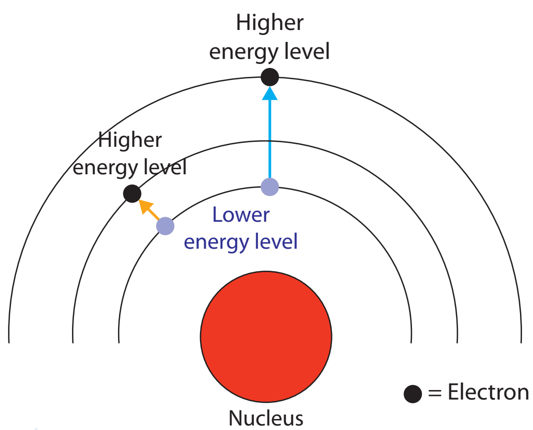


FIGURE 5.2.1 When electrons absorb energy, they can move between energy levels.

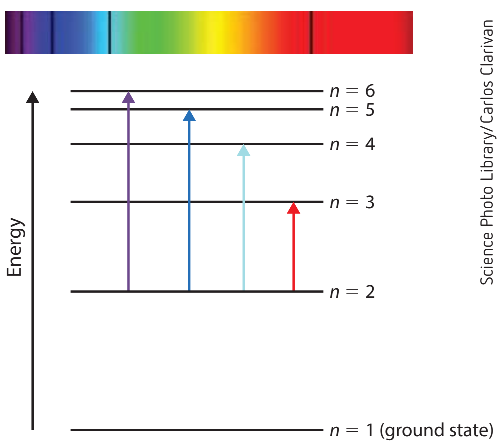


FIGURE 5.2.2 Hydrogen absorption spectrum

An electron moving from one energy level to another is undergoing a transition. A transition is associated with a specific amount of energy. When an amount of energy is applied to an atom, only the specific quantities of energy that correspond to the electron transitions between the shells can be absorbed (**absorption transition**). The transition energies are said to be 'quantised'. These specific quantities of energy can be supplied by specific wavelengths of electromagnetic radiation or other sources. This 'quantisation' was first explained by the Danish scientist Niels Bohr in his explanation of the observations of the absorptions of light by hydrogen atoms.

absorption transition
an electron moving to a higher energy level, when absorbing additional energy

absorption spectrum
a spectrum of electromagnetic radiation transmitted through a substance, showing dark lines or bands due to the absorption at specific wavelengths corresponding to transitions from excited energy levels to higher energy levels

Scientists had noticed the series of black lines in the **absorption spectrum** of hydrogen atoms (H) produced by the dissociation of molecular hydrogen (H_2) in an electric discharge. They had shone a full spectrum of visible wavelengths of light ('white light') at some hydrogen atoms and then dispersed the light through a prism (i.e. separated the light into its component wavelengths to get a 'rainbow' or the visible spectrum of light). They noticed that most of the light had passed through the hydrogen atoms unchanged, yet some wavelengths, shown in black in Figure 5.2.2, had been absorbed. Bohr made the link between these absorbed wavelengths and the electron absorption transitions between energy levels. Scientists now regard the existence of these absorptions as evidence for the existence of electron energy levels within atoms.

From Bohr's work, we have been able to establish that every element has a unique absorption spectrum. Each element's energy levels are unique because the electron configuration of each element is different. This means that every element's absorption spectrum shows a different pattern of absorbed wavelengths. This means that every element can be identified by its absorption spectrum.

SECTION REVIEW 5.2

REMEMBERING

1 Describe what happens to electrons when an atom absorbs energy.

UNDERSTANDING

2 Explain why every element has a unique absorption spectrum.

3 Explain why the lines in an absorption spectrum appear black.

5.3 Emission spectroscopy

excited state
an altered electron configuration from the ground state that occurs when an atom absorbs additional energy

photon
a particle representing a quantum of light

emission spectrum
the distribution of the frequencies of electromagnetic radiation emitted by an atom that has been heated or excited

An atom can have multiple energy levels, so it is possible for an electron to move up one, two or even more energy levels. An atom with electrons in upper energy levels is said to be in an **excited state**. Electrons in the excited, higher energy levels are unstable. After a very short time, less than one-millionth of a second, the electrons move back down to their original energy levels. As they do, they release the energy that they previously absorbed.

This energy is emitted as **photons** of light. Different transitions can occur as the electron falls to its original level because of the number of energy levels. Each transition between two shells represents a specific amount of energy, which equates to a photon of light with a specific wavelength. An atom will absorb and emit photons of different wavelengths because of the possible number of different transitions. Again, because the energy levels in each element are different, each element emits a unique series of wavelengths of light – its **emission spectrum**.

If the source of light is light emitted from an element or compound, a spectrometer can separate the emitted light into its component wavelengths to produce an emission spectrum. Figure 5.3.1 shows the operation of the spectrometer. The emitted light is composed of multiple wavelengths, which are dispersed through a prism and shone onto film (or digital array detector) to produce an emission spectrum.

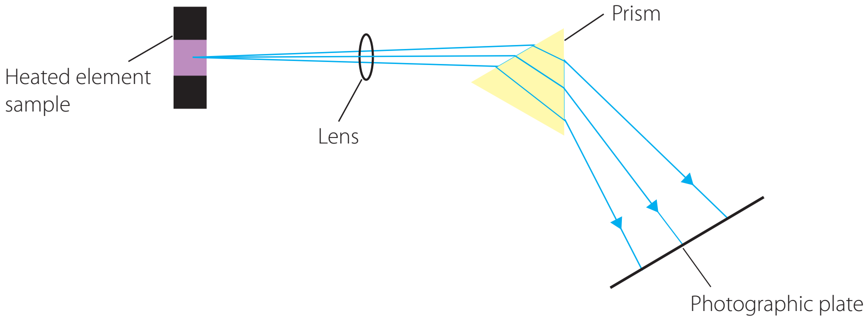


FIGURE 5.3.1 How an emission spectrum is formed

Spectra of some common elements are shown in Figure 5.3.2. Note that they are the *inverse* of an absorption spectrum; that is, series of coloured lines on a black background, rather than black lines on a coloured background. The coloured lines are the wavelengths of light emitted as the result of the electrons falling back into their original energy levels from an excited state. Scientists can determine the identity of elements in an unknown sample by comparing the lines in its emission spectrum to those known to be produced by specific elements, such as those in Figure 5.3.2.

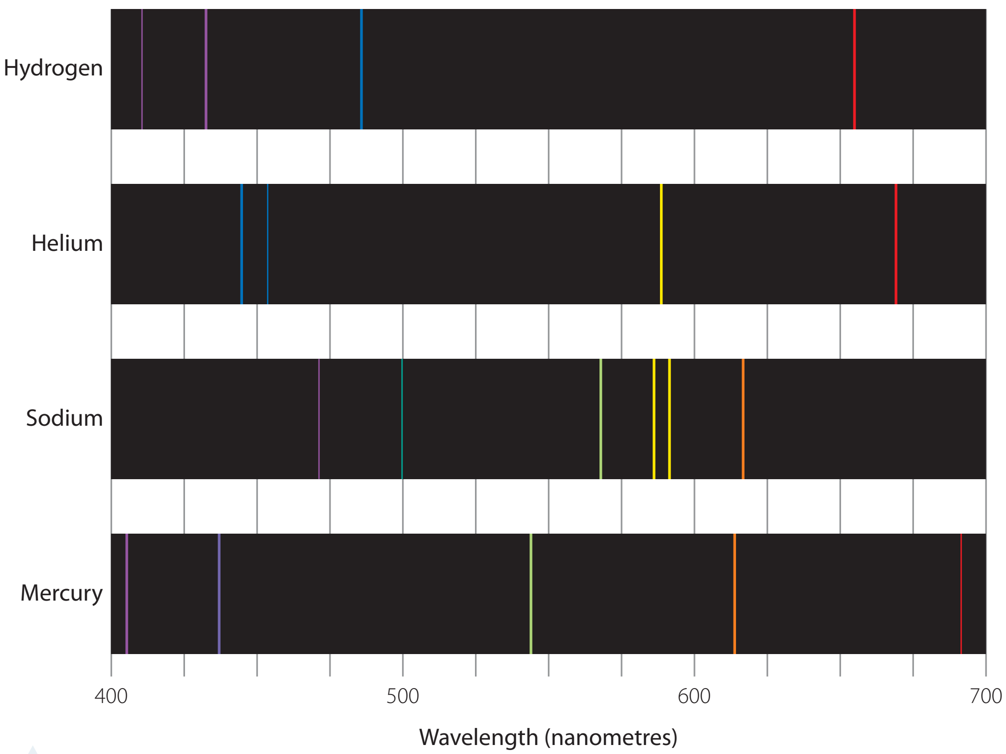


FIGURE 5.3.2 Emission spectra of some common elements

Hydrogen emission spectrum

The hydrogen emission spectrum is the simplest of all the elements, because hydrogen only has one electron. Nonetheless, a range of different lines are emitted in the infra-red, visible and ultraviolet regions of the electromagnetic spectrum because of the possible different transitions, as illustrated in Figure 5.3.3.

The Lyman series of lines are emissions in the ultraviolet region of the electromagnetic spectrum, which correspond to transitions of electrons to the lowest energy level, closest to the nucleus.

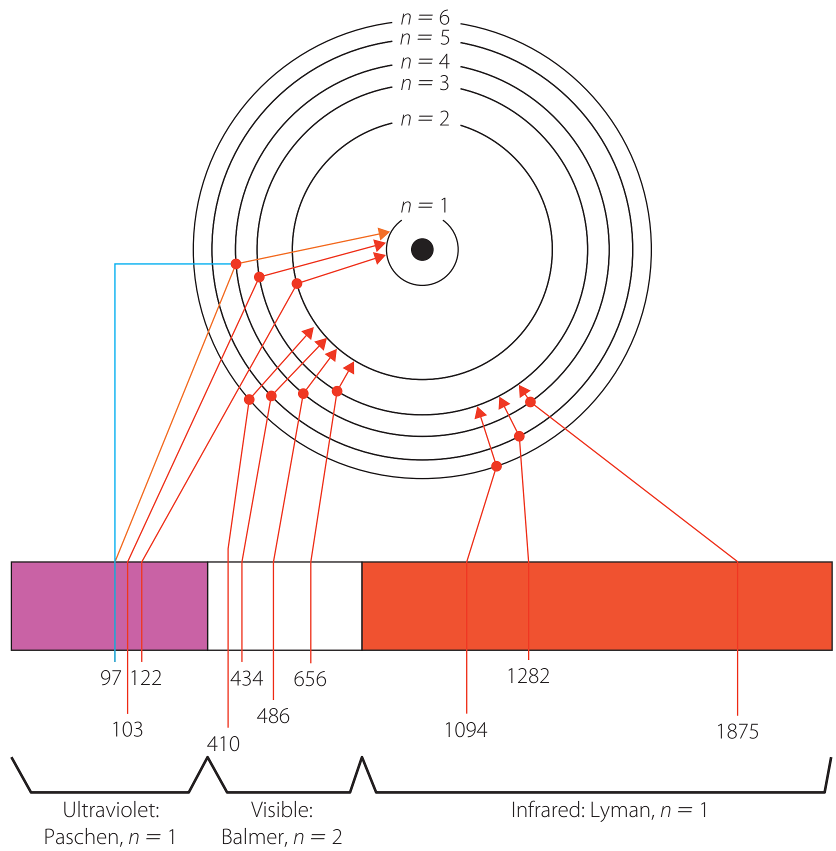
The Balmer series of lines are lower energy emissions in the visible region, involving transitions to the second energy level. These emissions are lower in energy than the Lyman series, because the differences in energy levels involved are smaller.

The Paschen series of lines are lower in energy again, because they are transitions to the third energy level, and occur in the infrared region.

A much simpler form of **emission spectroscopy** is that of flame tests, as described in Practical Activity 5.3.1 (page 73).

emission spectroscopy a technique where the discrete frequencies of radiation emitted are analysed to provide qualitative and quantitative information about the sample

FIGURE 5.3.3 A series of electron energy levels represented as $n = 1$, $n = 2$ and $n = 3$. It shows the different electron transitions that can occur in a hydrogen atom.



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PRACTICAL ACTIVITY 5.3.1

Flame tests

Some element ions give a characteristic colour when placed into the flame of a Bunsen burner. These different colours can be used to identify different cations (positive ions).

AIM

To observe and compare the colours of flames produced by different elements

EQUIPMENT PER GROUP

- solid samples (0.1 g of: calcium nitrate ($Ca(NO_3)_2$), calcium chloride ($CaCl_2$), sodium nitrate ($NaNO_3$), sodium chloride (NaCl), potassium nitrate (KNO_3), potassium chloride (KCl), strontium nitrate ($Sr(NO_3)_2$), strontium chloride ($SrCl_2$)
- 20 mL of 1.0 mol L^{-1} hydrochloric acid (HCl)
- 20 mL distilled water
- 2 × 100 mL beakers
- platinum or nichrome wire loop
- Bunsen burner
- matches

WHAT ARE THE RISKS IN DOING THIS EXPERIMENT?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?
The Bunsen burner and equipment will get hot and could cause burns.	Turn off the Bunsen burner or turn it to a yellow flame when not in use. Handle hot objects with care and do not place them directly onto bench tops; use a heatproof mat.
Use of chemicals	
Use of a spectroscope	

RISK ASSESSMENT

In your write-up, copy and complete the risk assessment table. Add any more risks you can think of, as well as ways to manage them. In particular, expand on the three risks listed to identify specific risks involved with each of them. Ask your teacher to check your risk assessment before you proceed.

PROCEDURE

1. Place approximately 20 mL of 1.0 mol L^{-1} HCl into a 100 mL beaker.
2. Place approximately 20 mL of distilled water into a 100 mL beaker.
3. Dip the wire loop into the distilled water, and then into the acid to clean it.
4. Hold the wire loop in the blue part of the Bunsen burner flame to remove any chemicals left on it. When the flame burns a normal blue colour, the wire loop is clean.
5. Dip the wire loop into the solid sample of calcium nitrate and hold it in the blue part of the Bunsen burner flame. Record the colour or colours observed.
6. Dip the loop into the distilled water and hydrochloric acid to clean it.
7. Repeat steps 5 and 6 for all the solids, ensuring you record the colour of each flame.

RESULTS

Create a table of your results, showing the chemical name and the colour of the flame. Compare your results with those of other groups because observation of colour can differ between individuals.

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QUESTIONS

1. Describe any patterns you can see in your results.
2. Describe any problems you had with determining the colours. Compare your results and observations with those of other groups.
3. Research the colours you should have seen for these elements. Your teacher may provide you with the expected colours.

DISCUSSION

1. Explain why the same elements will have the same colour flame when a flame test is conducted.
2. Explain why you may not have seen the expected colours in this experiment.

CONCLUSION

Write a conclusion linking elements and their flame colours.

SECTION REVIEW 5.3

REMEMBERING

1. Outline how an atom can emit energy.
2. Describe a spectroscope and how it works. Draw and label a simple diagram in your answer.

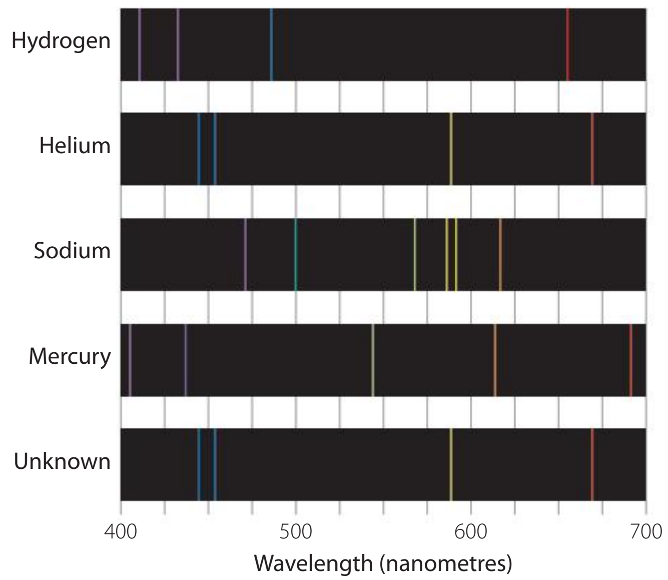
UNDERSTANDING

3. Explain 'ground state' and 'excited state' in reference to electrons in an atom. Draw and label a diagram in your answer.
4. Explain why all atoms of sodium will emit the same set of wavelengths of light when excited by intense heat or in a sodium vapour lamp (electric discharge).
5. Explain why sodium and magnesium have different emission spectra.

APPLYING

6. Figure 5.3.4 shows the emission spectra for several known elements and one unknown element. Identify the unknown element, giving reasons for your answer.

FIGURE 5.3.4 Emission spectra of some known elements and an unknown element



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5.4 Atomic absorption spectroscopy – a quantitative analysis

Analysing the wavelengths that make up an absorption or emission spectrum is a form of **qualitative analysis** – the spectrum can help identify the sample being tested.

Atomic absorption spectroscopy, or AAS, is a modified version of absorption spectroscopy that can be used to perform **quantitative analysis**; that is, to find the amount of an element present. It uses the principle of the Beer–Lambert law: the concentration of the sample element present is proportional to the amount of light that is absorbed. AAS's purpose is to measure the amount of light absorbed.

This analytical technique uses an atomic absorption spectrometer. Figure 5.4.1 outlines the technique.

qualitative analysis a technique that determines the identity of the sample, rather than the quantity

quantitative analysis a measurement of the amount of a substance present in a sample

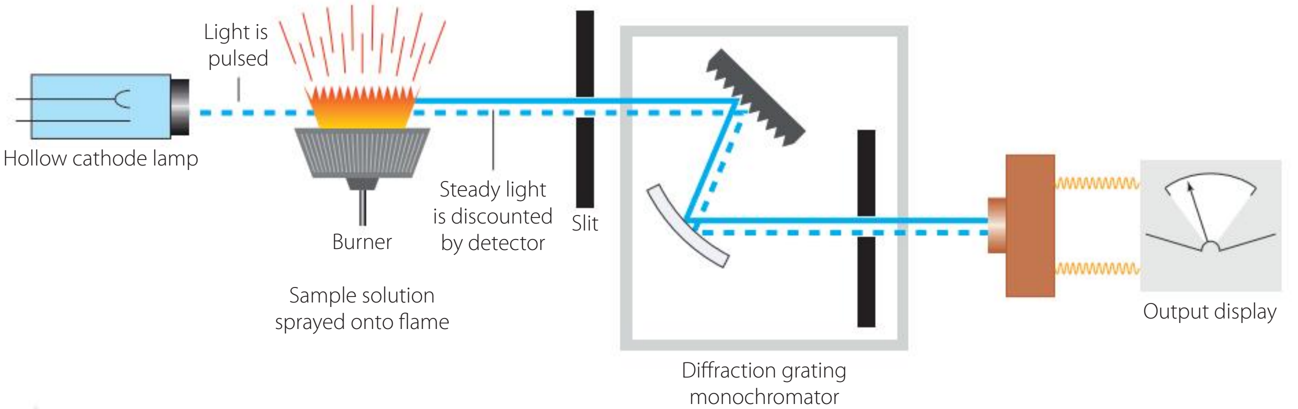

FIGURE 5.4.1 The process of atomic absorption spectroscopy

SCIENCE AS A HUMAN ENDEAVOUR

CSIRO scientist Alan Walsh (Figure 5.4.2) invented the concept of atomic absorption spectroscopy in 1952. Its development is considered one of the greatest achievements in chemical analysis.

AAS permits the measuring of tiny quantities of elements, which makes it useful to various fields worldwide: medicine, manufacturing, agriculture, mineral exploration, food analysis, metallurgy, biochemistry and environmental monitoring. Its commercialisation helped create Australia's scientific instrument industry, which is worth over $1 billion a year today.

When the original AAS patents expired circa 1969, more than 10 000 atomic absorption spectrophotometers were in use across hospitals, factories and laboratories all over the world. By 1977, this number jumped to around 40 000. Had sales continued to increase at the same rate as 1963–68, by the twenty-first century, the surface of Earth would have been covered by atomic absorption spectrophotometers!

FIGURE 5.4.2 CSIRO scientist Sir Alan Walsh, inventor of AAS

PRACTICAL ACTIVITY 5.4.1

Constructing a calibration curve to measure the concentration of copper(II) sulfate

A solution of copper(II) sulfate ($CuSO_4$) is blue and the intensity of the colour directly relates to the solution's concentration. Several methods can be used to determine the colour's intensity. For example, a simple light meter will measure the intensity of light passing through a solution. The more light that is absorbed, the more concentrated the solution is. However, a colorimeter shines light of a specific wavelength through a sample and measures an absorbance value similarly to an atomic absorption spectrometer.

In this experiment, you will measure the intensity of light passing through different concentrations of copper(II) sulfate solutions and construct a calibration curve. You will then be given a copper(II) sulfate sample of unknown concentration and will use your calibration curve to determine the solution's concentration.

AIM

To determine the concentration of a copper(II) sulfate solution by constructing and using a calibration curve

EQUIPMENT PER GROUP

- light source and light meter, colorimeter or colorimeter data probe with data logger or laptop
- 35 mL (approximately) of 1.0 mol L^{-1} of $CuSO_4$
- distilled water
- 50 mL beaker
- 10 mL measuring cylinder

RISK ASSESSMENT

WHAT ARE THE RISKS IN DOING THIS EXPERIMENT?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?
Use of chemicals	
Use of light meter or data logger	

Copy and complete the risk assessment table in your write up. Add any extra risks you can think of and how to manage them. Expand on the two listed risks to identify specific hazards involved with each. Ask your teacher to check your table before you proceed.

»

PROCEDURE

1. Collect approximately 35 mL of 1.0 mol L^{-1} $CuSO_4$ from your teacher.
2. You will use a piece of clear glassware to hold your copper(II) sulfate sample. Measure 10 mL into this piece of glassware, or fill a provided container to overflowing.
3. Take a reading of the light absorbance of the sample. The absorbance value will be provided as a read-out on your light meter or colorimeter.
4. Measure 8 mL of the 1.0 mol L^{-1} $CuSO_4$ solution into a 10 mL measuring cylinder and make up to 10 mL with distilled water. This will provide you with a 0.8 mol L^{-1} $CuSO_4$ solution.
5. Repeat steps 2 and 3 to determine an absorbance reading for this sample.
6. Measure 6 mL of the 1.0 mol L^{-1} $CuSO_4$ solution into a 10 mL measuring cylinder and make up to 10 mL with distilled water. This will provide you with a 0.6 mol L^{-1} $CuSO_4$ solution.
7. Repeat steps 2 and 3 to determine an absorbance reading for this sample.
8. Measure 4 mL of the 1.0 mol L^{-1} $CuSO_4$ solution into a 10 mL measuring cylinder and make up to 10 mL with distilled water. This will provide you with a 0.4 mol L^{-1} $CuSO_4$ solution.
9. Repeat steps 2 and 3 to determine an absorbance reading for this sample.
10. Measure 2 mL of the 1.0 mol L^{-1} $CuSO_4$ solution into a 10 mL measuring cylinder and make up to 10 mL with distilled water. This will provide you with a 0.2 mol L^{-1} $CuSO_4$ solution.
11. Repeat steps 2 and 3 to determine an absorbance reading for this sample.
12. Collect 10 mL of the sample of unknown concentration from your teacher.
13. Repeat steps 2 and 3 to determine an absorbance reading for this sample.

RESULTS

1. Construct a table showing your known concentrations and their absorbances. Include the result for the unknown concentration in this table.
2. Construct a fully labelled calibration curve for your known results.
3. Use the graph to determine the concentration of the unknown sample by interpolation.

QUESTIONS

1. Did your graph pass through the origin (0,0)? Why should you expect it to? Suggest reasons why your graph might not do this.
2. Compare your results with other groups. Did you all get the same answer? If possible, create a table showing a class set of results. Account for any differences between the results.

DISCUSSION

1. If your teacher can tell you the correct concentration of the unknown sample, discuss the accuracy of your results (that is, how close your results were to the true value).
2. Identify one error that may have affected your results. Discuss its effect on the results. Identify the error as random or systematic.

CONCLUSION

Write a conclusion discussing the results of this experiment.

random error
a variation that affects a measurement in a random way so that the measurement is as likely to change in any one direction as in any other

systematic error
an error that acts to give a consistent offset in data; for example, a zero error

AAS can be used only to measure the quantity of a specific element. This is because the lamp for this process is made from the same element that is being tested. If the amount of zinc is being tested, then the lamp filament is made of zinc. Similarly, if mercury is being tested, then the lamp is filled with mercury vapour.

An electric current is passed through the filament or via electrodes through a gaseous sample of the element. The electric current heats the filament to very high temperatures, or electron bombardment due to the current flowing in the gas excites the atomic vapour. The excited atoms will then emit light (see section 5.3). When the lamp contains a filament or vapour of a single element, then the light emitted has only the unique set of wavelengths particular to that element.

The sample being tested is **atomised**, changing the substances it contains into gaseous atoms. When the light from the lamp passes through the atomised sample, only the element being tested for will absorb the light from the lamp. This is because it has the same energy levels as the atoms that emitted the light from the lamp. Other elements in the vaporised sample will not absorb this light because the energy levels of all other atoms are different and their electrons cannot absorb the energies of the light present.

The light passes through the sample and is focused through a slit before entering a **monochromator**. The monochromator selects just one wavelength of the light for analysis by the detector. The detector measures the intensity of the light, which is then displayed as a number. This number, the **absorbance value**, is not a concentration but a measure of the amount of light that passed through the sample without being absorbed.

To measure the quantity of an element present, the absorbance of the sample is compared to that of a known concentration by constructing a **calibration curve**. First, a number of known concentrations of the element are prepared and their intensities are measured by atomic absorption spectroscopy. Then, a calibration curve of concentration against absorbance value is plotted. The concentration of an unknown sample can then be compared and determined by **interpolation** once its absorbance is measured.

atomised
a substance is heated until it exists as a vapour of gaseous atoms

monochromator
a filter used to ensure that only radiation of a specific frequency is shone at a sample

absorbance value
a measurement of the relative amount of radiation that is absorbed by a sample

calibration curve
a graph of absorbance values of a series of samples of known concentrations used to determine the concentration of an unknown sample

interpolation
to read or construct a new data point that has not been measured but is within the range of measured data

WORKED EXAMPLE 5.4.1

To determine the concentration of mercury in a sample of fish, the absorbances of some mercury samples of known concentration were measured by atomic absorption spectroscopy. Table 5.4.1 shows the results obtained. The fish sample was then analysed. Its absorbance value was 0.57. Note that Chapter 16 also discusses measures of concentration.

QUESTION

What is the concentration of mercury in the fish sample?

TABLE 5.4.1 Measurements of absorbance of known concentrations of mercury

MERCURY CONCENTRATION (PARTS PER MILLION, PPM)	ABSORBANCE
2.0	0.15
4.0	0.30
6.0	0.46
8.0	0.61
10.0	0.74

ANSWER

1 Construct a calibration curve using a computer spreadsheet or graphics calculator. The mercury concentration [Hg] in units of ppm is the independent variable, and the absorbance (A, arbitrary units) is the dependent variable. Figure 5.4.3 displays the result.

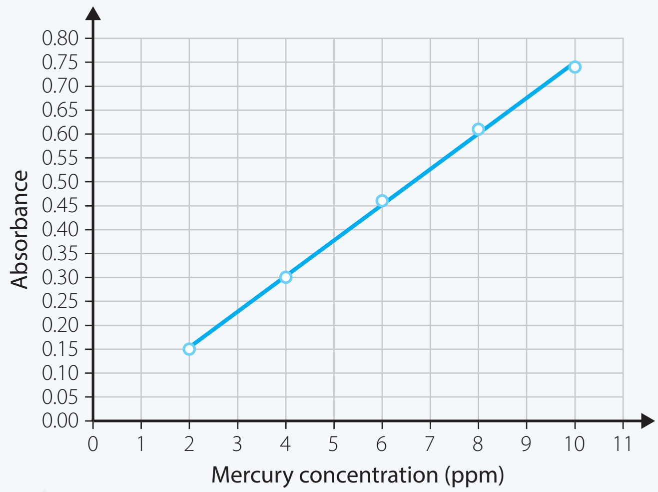

FIGURE 5.4.3 Mercury concentration

2 The linear trendline fit to the data shows the functional relationship:

$$A = 0.0745[Hg] + 0.005$$

On re-arrangement, this yields:

$$[Hg] = \frac{A - 0.005}{0.0745}$$

Hence:

$$A = 0.57$$

$$Hg = 7.58\,\text{ppm}$$

Rounded, $Hg = 7.6\,\text{ppm}$ to be consistent with precision of data

SECTION REVIEW 5.4

REMEMBERING

1 Outline the process of atomic absorption spectroscopy.

2 Describe how to construct a calibration curve for the process of atomic absorption spectroscopy.

UNDERSTANDING

3 Explain why the lamp in atomic absorption spectroscopy contains a filament or vapour of the same element as the element being tested.

ANALYSING

4 Cadmium is useful in small amounts but dangerous in large amounts. A sample of paint was tested for its cadmium content. Samples of known concentrations of cadmium were analysed by atomic absorption spectroscopy and the results in the following table were obtained.

CADMIUM CONCENTRATION ($mg\ L^{-1}$)	ABSORBANCE
0.0	0.000
1.0	0.038
2.0	0.082
3.0	0.120
4.0	0.160
5.0	0.200
6.0	0.240
Unknown	0.110

a Construct a calibration curve for the known concentrations of cadmium.

b From the curve, determine the concentration of cadmium in the sample of paint.

REFLECTING

5 Describe some ways that atomic absorption spectroscopy can help scientists and students learn more about the elements in materials, or how elements can be identified. Use examples in this section for ideas, or search the Internet for other uses.

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 Define the following terms.

a Atomic absorption spectroscopy

b Emission spectroscopy

c Flame test

d Mass spectrometry

e Qualitative analysis

f Quantitative analysis

CATEGORY QUESTIONS

2 Classify each of the following analytical techniques as quantitative, qualitative or both. Justify your choices.

a Mass spectrometry

b Flame test

c Emission spectroscopy

d Atomic absorption spectroscopy

ELABORATION QUESTIONS

3 The level of zinc in some food was labelled on the container. Students analysed samples of the food by atomic absorption spectroscopy to determine whether the level on the container was correct. Samples of known concentration of zinc were analysed and the following data were obtained, as shown in the table.

ZINC CONCENTRATION ($mg\,L^{-1}$)	ABSORBANCE
1	0.03
2	0.06
5	0.16
10	0.32
15	0.48

a Identify what the lamp was made of, from this analysis.

b Explain why the presence of a metal, such as calcium in the food, would not have interfered with this process.

c Construct a calibration curve with the concentration of zinc on the horizontal axis and the absorbance on the vertical axis.

d The sample of food gave an absorbance reading of 0.36. From the graph, determine the concentration of zinc in the food.

e The label on the container stated that the level of zinc in the food did not exceed 7.5 $mg\,L^{-1}$. Explain whether this statement was true.

4 Figure 5.4.4 represents the mass spectrum produced from the isotopic analysis of a particular element.

a Explain why the different isotopes can be separated by a mass spectrometer.

b How many isotopes were present in the element sample?

c Which isotope was present in the highest concentration?

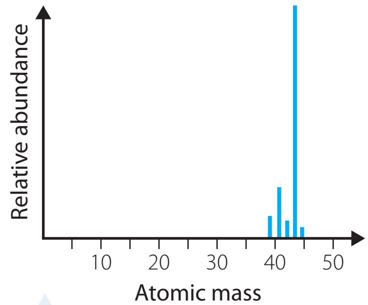

FIGURE 5.4.4 Mass spectrum of an element

EVIDENCE QUESTIONS

5 Figure 5.4.5 shows the line emission spectra of several common elements and one unknown element. Identify the unknown sample and describe it.

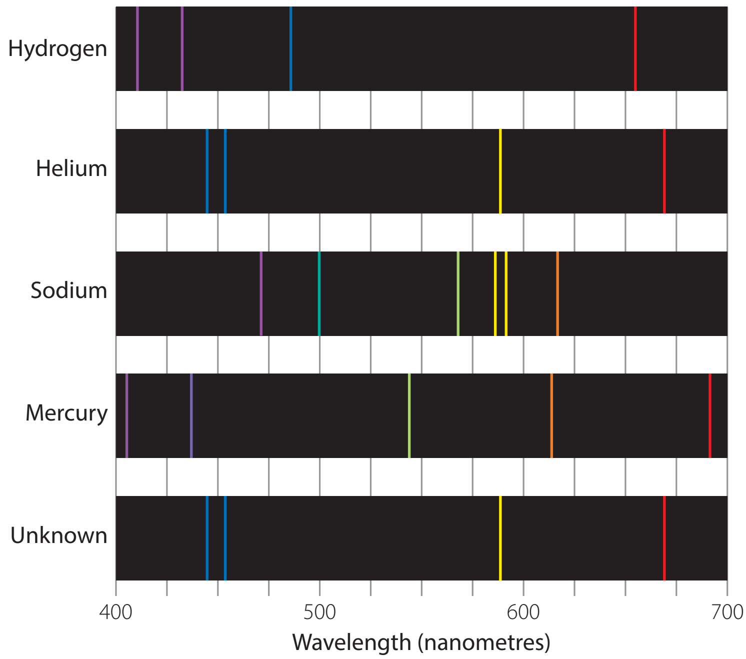

FIGURE 5.4.5 Emission spectra of known elements and an unknown element

6 Emission spectroscopy has been used to identify the presence of helium in the Sun. Conduct research and explain how the presence of helium was detected.

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END-OF-CHAPTER EXAM

End-of-chapter test

1 In terms of isotopic analysis, the purpose of a mass spectrometer is to:

A determine the isotopic molar masses of a chemical compound.

B identify the existence of isotopes of an element and determine the relative masses of the isotopes.

C identify the ratio of relative atomic masses of several chemical compounds.

D identify the existence of isotopes of an element and their abundances, to determine relative atomic masses of a chemical compound.

2 The lamp used in an atomic absorption spectrophotometer produces radiation:

A that will cause an electron in the element being analysed to be excited to a higher energy state.

B over a wide range of energies.

C with energies in the infra-red part of the spectrum that will excite electrons to higher states.

D that will cause an electron in the element being analysed to move to a lower energy state.

3 Boron has two isotopes: ^{10}B and ^{11}B.

RELATIVE ABUNDANCE (%)	MASS NUMBER
20	10
80	11

a Calculate the relative atomic mass of boron.

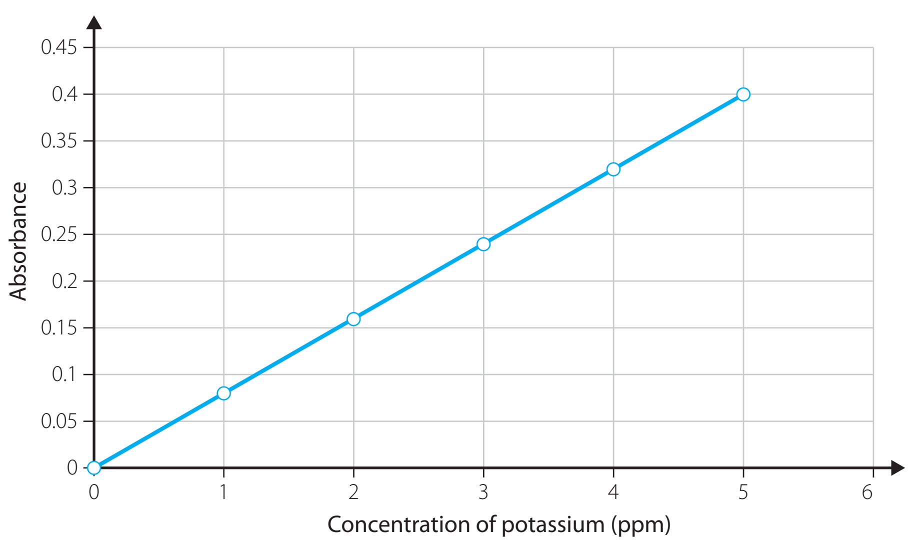
b Outline how a mass spectrometer can provide the relative abundance and mass number.

4 Bananas are an important source of the metal potassium, which is essential in trace quantities in virtually all animals. A 3.50 g sample of bananas was pulped and dissolved in 20.0 mL of acid. The mixture was filtered and diluted to 100.0 mL with distilled water. A small quantity of this solution was then passed through an atomic absorption spectrometer along with a number of solutions of known potassium concentration. The calibration curve generated by this experiment is shown in the following graph. The absorbance of the diluted banana solution was recorded as 0.19.

a Why was it necessary to generate a calibration curve for various potassium concentrations in this experiment?

b Use the calibration curve to determine the concentration of potassium in the diluted banana solution.

c Determine the concentration of potassium in the undiluted acid solution in parts per million (ppm).



» UNIT ONE

CHEMICAL FUNDAMENTALS: STRUCTURE, PROPERTIES AND REACTIONS

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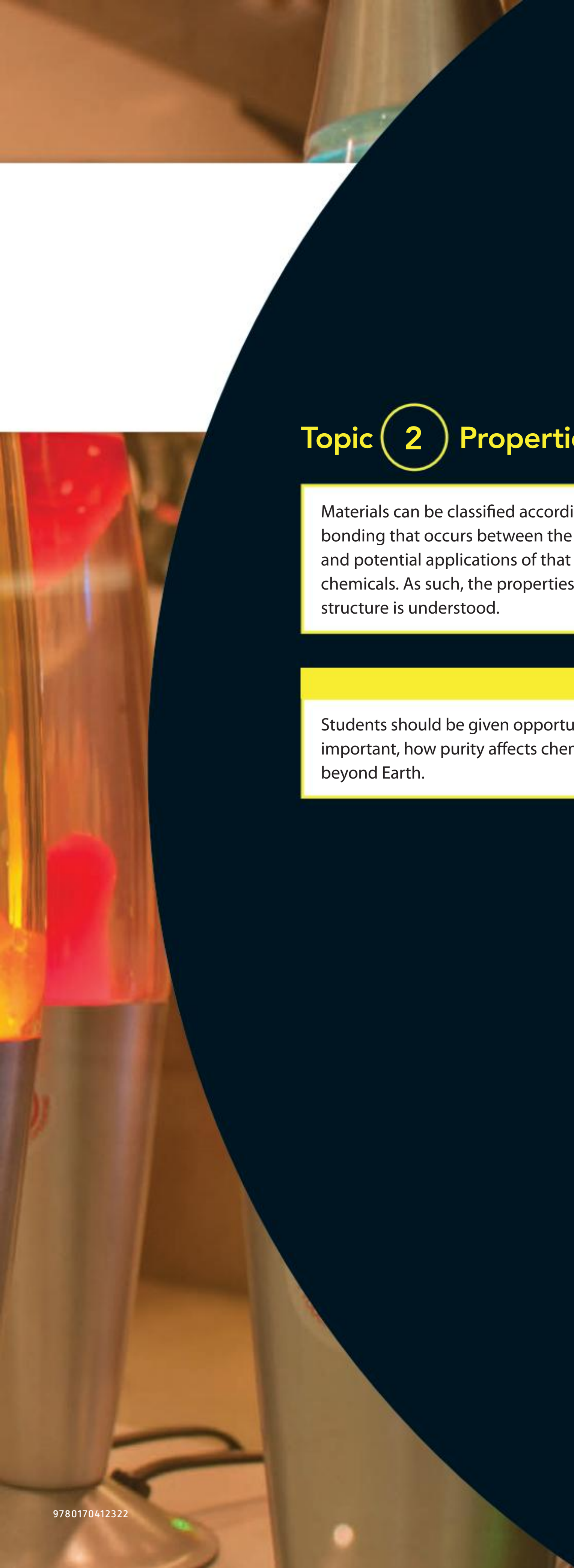

Topic 2 Properties and structure of materials

Materials can be classified according to their properties. These properties are the result of the bonding that occurs between the atoms of that material and are fundamental to the usefulness and potential applications of that material. This applies equally to organic and inorganic chemicals. As such, the properties of new materials can be predicted and modified if their structure is understood.

SCIENCE AS A HUMAN ENDEAVOUR

Students should be given opportunities to investigate: the uses of nanomaterials, why purity is important, how purity affects chemical reactions, and the possibility of carbon-based life existing beyond Earth.

6 COMPOUNDS AND MIXTURES

Introduction

Pure substances may be either elements or compounds and they have distinct physical and chemical properties. Elements cannot be separated into simpler substances. However, compounds can be separated into simpler substances using chemical processes. Mixtures can be separated into simpler substances using physical processes. The physical process used depends on the part or parts of the mixture that needs to be kept. While the inhomogeneity of some mixtures is visible to the naked eye, other heterogeneous mixtures must be viewed under an optical microscope for their inhomogeneity to be revealed. Electron microscopes are required to carry out nanotechnological investigations.

Stimulus questions

Are pure substances more valuable than impure substances?

Why is it important to understand the difference between pure substances and mixtures?

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6.1 Pure substances

Pure substances are composed of only one substance. Pure substances may be either **elements** or **compounds**. Some elements were discovered thousands of years ago, because they appear in nature in their elemental form. For example, copper, silver and gold (Figure 6.1.2, page 88) appear in nature in their elemental form because of their low chemical reactivity. As such, they have been used throughout the world to make coins. In the early 1800s, electrolysis technology enabled a British chemist, Humphry Davy, to isolate elemental potassium from the compound caustic potash (potassium hydroxide) and elemental calcium from the compound quicklime (calcium oxide). Some elements were discovered by laboratory synthesis, but were later found in trace amounts in nature. Technetium was the first element to be synthetically discovered in 1937 by Italian scientists Carlo Perrier and Emilio Segrè in a sample of molybdenum that had undergone particle acceleration. The element tennessine was synthesised in 2010, when Russian physicists led by Yuri Oganessian fused berkelium and calcium. Neither physical nor chemical processes can break down elements into simpler substances. Elements are the building blocks of all materials (Figure 6.1.1).

pure substance
composed of only one substance

element
a substance consisting of only one type of atom

compound
a substance consisting of atoms of two or more elements chemically bonded together in a fixed ratio

6.1.1 Pure chemical substances

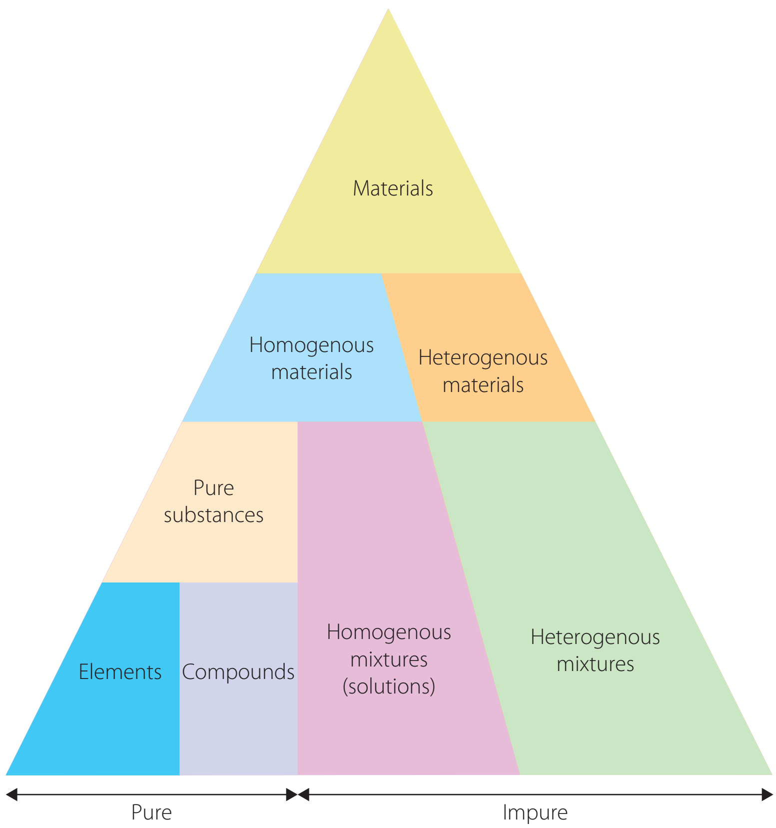

FIGURE 6.1.1 Classification of materials

Under the same conditions, pure substances exhibit the same physical and chemical properties. Physical properties include the following:

- melting and boiling points
- electrical and thermal conductivity
- density
- solubility.

FIGURE 6.1.2 Gold bars are still intact 150 years after seawater submersion, due to gold's low chemical reactivity.

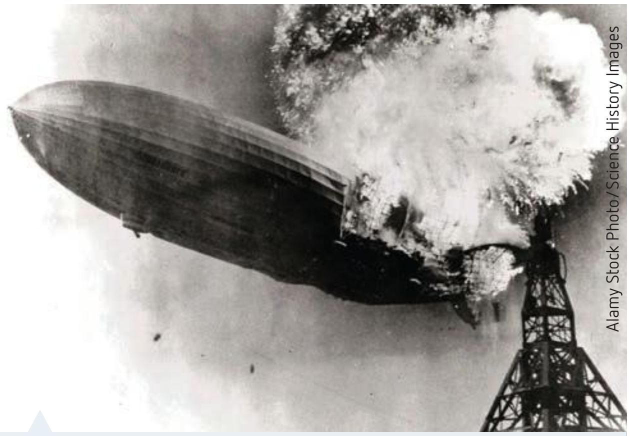

FIGURE 6.1.3 The *Hindenburg* disaster occurred because early airships were filled with light, but flammable, hydrogen.

Chemical properties include the following:

- reactivity
- stability
- toxicity
- flammability.

Throughout human history we have used substances initially because of their physical properties, but sometimes have been unaware of their chemical properties. For example, early airships were filled with hydrogen because of its low density. Unfortunately, the high flammability of hydrogen was not taken into consideration and this caused the *Hindenburg* disaster on 6 May 1937 at the Naval Air Station, Lakehurst, New Jersey, United States (Figure 6.1.3). The LZ 129 *Hindenburg* caught fire and was destroyed, resulting in 36 fatalities. In response to this incident, today's airships are filled with helium, which shares the same physical property of low density with hydrogen but has low flammability. Fire extinguishers are an example of humans exploiting both the physical and chemical properties of substances. Carbon dioxide has a higher density than air, which enables it to blanket the flame, and its low flammability does not support further combustion.

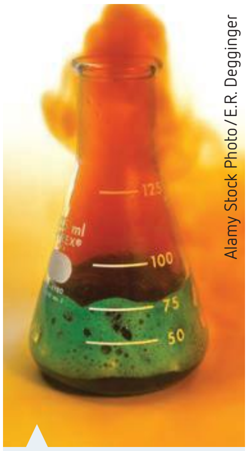

FIGURE 6.1.4 Observations indicating chemical reactions: solid disappearance, gas evolution, liquid colour change and temperature increase

Chemical processes

Chemical processes can break down compounds into simpler substances. For example, lead nitrate decomposes into lead oxide, nitrogen dioxide and oxygen on heating, according to the following equation:

$$2Pb(NO_3)_2(s) \xrightarrow{heat} 2PbO(s) + 4NO_2(g) + O_2(g)$$

$$\text{Lead(II) nitrate} \xrightarrow{heat} \text{lead(II) oxide} + \text{nitrogen dioxide} + \text{oxygen}$$

Observations can indicate that a chemical process (reaction) has occurred, such as a:

- solid forms (precipitation) or disappears
- gas is produced (evolved)
- colour change takes place
- temperature change takes place.

In Figure 6.1.4, all the following observations can be made in the reaction of copper with nitric acid:

- the copper solid disappears
- a brown–orange gas is evolved

- the liquid changes colour from colourless to blue
- the temperature increases.

The balanced chemical reaction for the reaction of copper with nitric acid forming nitrogen dioxide (brown–orange gas) and copper(II) ions (which give the liquid its blue colour) is given in the following equation:

$$Cu(s) + 4H^{+}(aq) + 2NO_3^{-}(aq) \rightarrow Cu^{2+}(aq) + 2NO_2(g) + 2H_2O(l)$$

Copper + nitric acid → copper(II) ions + nitrogen dioxide + water

INQUIRING FURTHER

Impure substances can be more valuable than pure substances. Pink diamonds are particularly valuable and are almost unique to Australia (Argyle mine), so of interest to Australian endeavour. Blue, pink and yellow diamonds are impure because they contain traces of elements in addition to carbon. Research the impurities contained in a blue, pink or yellow diamond and how these impurities affect the commercial value of the diamond.

PRACTICAL ACTIVITY 6.1.1

Observing chemical and physical properties of pure substances

Physical properties can be observed and measured without changing the composition of a substance. They can be used to identify substances. Chemical properties can only be observed when the substance is undergoing a change in composition.

AIM

To observe the physical and chemical properties of different substances

EQUIPMENT

- 5 g sulfur (S)
- 5 g iron filings (Fe)
- 5 g sodium hydrogen carbonate
- 5 g sodium chloride
- 5 g sucrose
- 2 × 1 cm strips magnesium ribbon (Mg)
- 30 mL of 2 mol L^{-1} hydrochloric acid (HCl)
- 5 g copper(II) chloride
- 3 × 5 cm squares of aluminium foil
- 50 mL beaker
- glass stirring rod
- heat mat
- magnifying glass
- 12 test tubes and test-tube holder
- magnet
- spatula
- Bunsen burner
- 10 mL measuring cylinder
- 8 pieces of 10 cm × 10 cm paper
- distilled water
- forceps
- test-tube tongs
- metal tongs
- matches

»

WHAT ARE THE RISKS IN DOING THIS EXPERIMENT?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?
Hydrochloric acid is corrosive.	Wear safety glasses and protective clothing. Take care when pouring and clean up spills immediately. If spilt on skin, wash the affected area with plenty of water and notify your teacher.
Copper(II) chloride is toxic.	Wear appropriate safety gear. Dispose of it in the chemical waste jar provided.
Substances may spit out of heated test tube.	Wear appropriate gear. Do not point the test tube at anyone; use tongs.
Powdered sulfur is an irritant to eyes, nose and throat.	Stir the mixture carefully and do not breathe in the fumes.

In your write-up, add any more risks you can think of, as well how to manage them. Ask your teacher to check your risk assessment before proceeding.

PROCEDURE PART A: OBSERVING PHYSICAL PROPERTIES

1 Obtain eight pieces of paper. Label each piece with: sulfur, iron filings, sodium hydrogen carbonate, sodium chloride, sucrose, magnesium, copper chloride or aluminium.
2 Using forceps, transfer one piece of magnesium and one piece of aluminium onto the appropriately labelled piece of paper. Using a clean spatula, transfer a small amount of each of the other substances onto the correspondingly labelled sheet.
3 Examine each substance carefully with the magnifying glass. Record your observations in the results table.
4 Test the effect of a magnet on each substance by moving the magnet under the sheet of paper. Record your observations.
5 Add a small amount of each substance to 5 mL of water in separate test tubes. Record your observations.
6 Save any solid samples remaining on the paper. Pour the copper chloride and water into the chemical waste jar. Dispose of solutions down the sink and add any solids to a waste container.

PROCEDURE PART B: OBSERVING CHEMICAL PROPERTIES AND CHANGES

7 Combine the remaining iron and sulfur and mix them with a spatula. Test with a magnet. Record your observations in the results table.
8 Place a strip of magnesium in a clean, dry test tube and add 5 mL of HCl. Record observations. Feel the test tube and record any temperature difference.
9 Repeat step 8 with small amounts of sodium hydrogen carbonate, sodium chloride, sucrose, copper chloride and aluminium. Remember that no reaction is also a result.
10 Place a watch glass close to the Bunsen burner. Light the Bunsen burner. Grasp a piece of aluminium with the metal tongs and hold it in the flame. Record your observations. Place any remaining aluminium on the watch glass.
11 Place 2 g of sucrose into a clean, dry test tube. Grasp the test tube with a test-tube holder and heat it gently over the Bunsen flame. Record your observations. Make sure to check for odour.
12 If no change has been observed, then heat the sample more vigorously for 1 minute. Remove it from the flame and place it in test-tube rack to cool. Scrape some of the residue from the test tube with a spatula and examine it.
13 Test the solubility of the residue and record your observations.
14 Discard the cooled test tube with the remainder of the residue into a waste glass container.
15 One-third fill a 50 mL beaker with distilled water. Add a spatula full of copper chloride to the water and stir it with a glass rod.
16 Loosely crumple a piece of aluminium foil and place it in the beaker. Record your observations. Note any changes in temperature.
17 Dispose of beaker contents in a chemical waste jar. Do not pour them down the sink.

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» RESULTS

1 Record your results for the physical changes in part **A** in a table like the one below.

SUBSTANCE	PHYSICAL STATE	COLOUR	ODOUR	SOLUBILITY IN WATER	EFFECT OF MAGNET
Sulfur					
Iron filings					
Sodium hydrogen carbonate					
Sodium chloride					
Sucrose					
Magnesium					
Copper chloride					
Aluminium					

2 Record your results for the chemical properties and changes in part B in a table like the one below.

SUBSTANCES	OBSERVATIONS
Fe and S – tested with magnet	
Mg and HCl	
Sucrose and HCl	
Sodium hydrogen carbonate and HCl	
Sodium chloride and HCl	
Copper(II) chloride and HCl	
Al and HCl	
Sucrose – heated	
Al – heated	
Copper(II) chloride and Al	

ANALYSIS

1 Classify each result in part B as a physical change, a chemical change or no change.

2 For those changes identified as chemical describe the evidence to support the decision.

3 Compare the physical properties of:

a heated aluminium with aluminium.

b heated sucrose with sucrose.

4 Summarise the chemical properties of aluminium and sucrose.

CONCLUSION

In your own words, state the difference between a physical and chemical change.

SECTION REVIEW 6.1

REMEMBERING

1 Define:

- **a** compound
- **b** element
- **c** pure substance.

UNDERSTANDING

2 Discuss the similarities and differences between hydrogen and helium.

3 Discuss the advantages and disadvantages of coins, metal bars and paper currency as investments.

4 Identify and discuss trends in the discovery of the elements.

APPLYING

5 Design an experiment (including safety considerations) to carry out the following reaction.

$$Cu(s) + 4H^+(aq) + 2NO_3^-(aq) \rightarrow Cu^{2+}(aq) + 2NO_2(g) + 2H_2O(l)$$

6.2 Separating mixtures

solute the substance that is dissolved in a solution

solvent the substance in which the solute of a solution is dissolved

Physical processes can separate mixtures into simpler substances. A mixture of two solids, such as sand and table salt, can be separated by dissolving the substance in water and using gravity filtration (Figure 6.2.1). In this example, the insoluble sand (residue) will be left in the filter paper and the soluble table salt will travel through the filter paper into the filtrate. A mixture of two **solutes** that are soluble in one **solvent**, such as table salt and copper sulfate in water, can be separated by recrystallising in a different solvent, such as ethanol. In this example, both table salt and copper sulfate are soluble in hot ethanol, but only table salt is soluble in cold ethanol, so copper sulfate crystals will precipitate out in cold ethanol. Subsequent gravity filtering will collect the copper sulfate (the residue in Figure 6.2.1) in the filter paper and the table salt will remain in the filtrate.

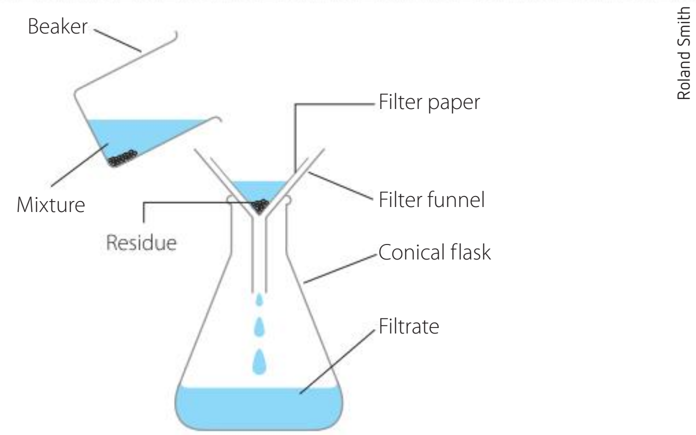


FIGURE 6.2.1 Gravity filtration separates insoluble residues in a mixture.

Distillation

A mixture of one solute in one solvent, such as table salt in water, can be separated by **distillation**. In the example in Figure 6.2.2, the salt water is boiled (in the round-bottomed flask) to give off pure water vapour, which is cooled (in the condenser) to form the pure water liquid **distillate**.

INQUIRING FURTHER

Distillation of salt water is the desalination technique used by several countries with ocean coastlines to alleviate chronic water deficits. Conduct research into the ethical, social, environmental and economic considerations of this practice.

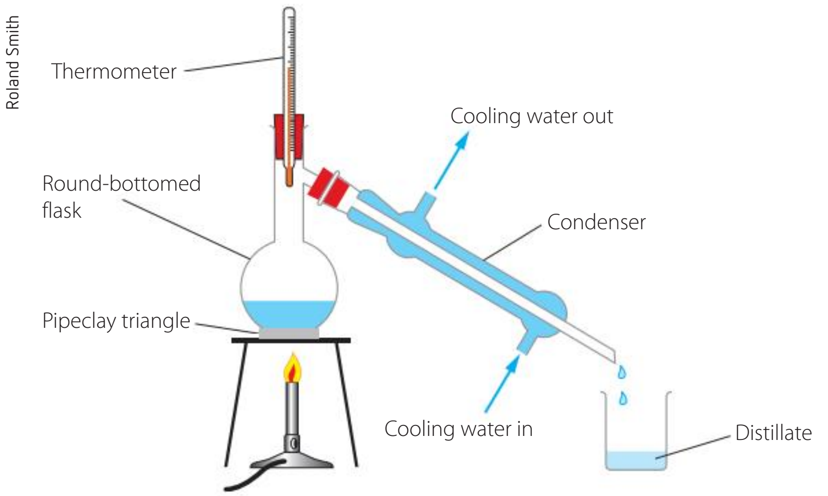


FIGURE 6.2.2 A simplistic set-up for carrying out distillation

distillation
a technique for separating the solvent from a solution, when the solvent is required to be kept

distillate
the solvent kept after distillation

A more complex set-up is required for distillation in oil refineries (Figure 6.2.3, page 94). In the fractional distillation of crude oil, the lowest boiling point hydrocarbons are extracted from the top of the column, and the further down the column the hydrocarbon is extracted, the higher is its boiling point. Natural gas, which is 95% methane (CH_4) with a boiling point of −161.4°C, is extracted from the top of the distillation column. Natural gas produces more heat and less carbon dioxide than coal or petroleum, but leakage and possible subsequent explosions from gas wells and pipelines has hampered its widespread use. Petrol ('gasoline') is actually a complex mixture of approximately 100 hydrocarbons. Petrol used in cars contains a mixture of numerous hydrocarbons, such as n-heptane ($CH_3(CH_2)_5CH_3$) and iso-octane ($(CH_3)_3CCH_2CH(CH_3)_2$), which have boiling points of approximately 99°C. Petrol is extracted from further down the fractional distillation column than natural gas. Petrol is currently the fuel of choice in mobile situations because it can be relatively easily and safely pumped from reservoirs. However, world crude oil supplies are running out and burning fossil fuels contributes to global warming through carbon dioxide production (see the following equation), so alternatives are being sought.

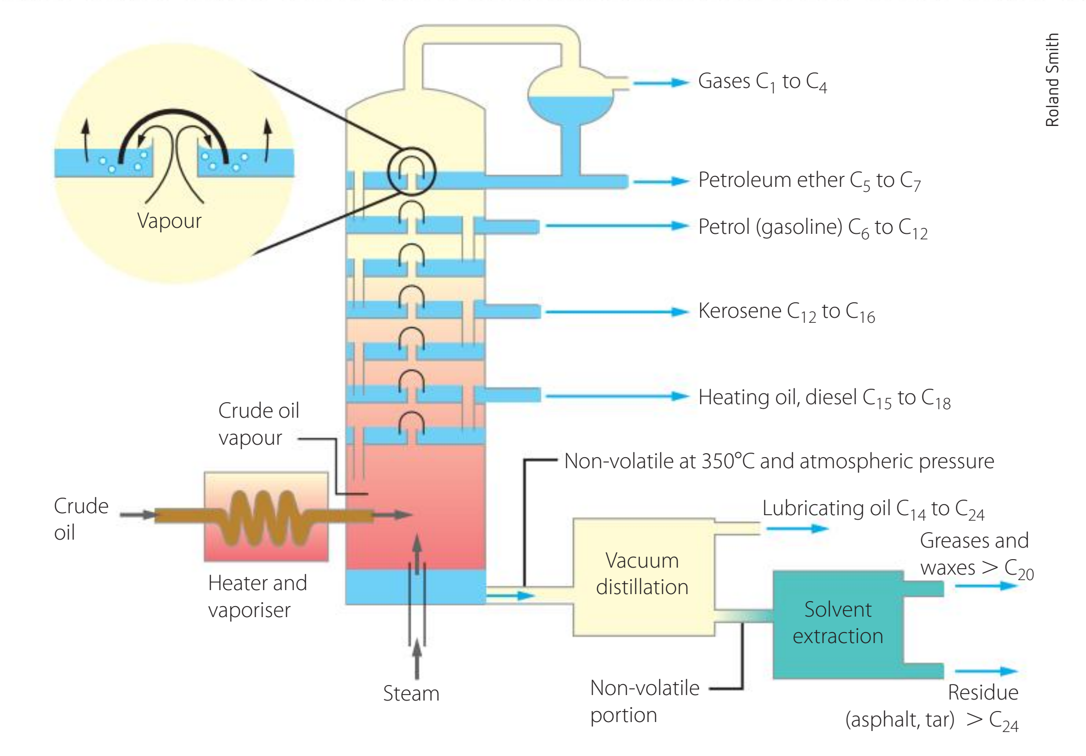
$$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(l)$$

$$\text{Octane} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water}$$

Diesel used in trucks has an average chemical formula of $C_{12}H_{23}$ and boiling points ranging from 200°C – 350°C and is extracted from near the bottom of the fractional distillation column. Although diesel undergoes more complete combustion than petrol, which results in less unburnt hydrocarbon emissions, it also results in more impurities being burnt and an increase in the emission of nitrogen and sulfur oxides.

6.2.1 Fractional distillation

FIGURE 6.2.3 Fractional distillation of crude oil used in oil refineries



Separatory funnels and vaporisation

immiscible liquid a liquid that does not mix with another liquid

vaporisation a technique for separating the solvent from a solution, when the solute is required to be kept

Separatory funnels are used in the extraction of compounds of interest from liquids. The initial liquid containing the compound of interest and an **immiscible liquid**, in which the compound of interest can dissolve, are added to the separatory funnel. For example, in Figure 6.2.4, in extracting caffeine from coffee, the coffee is dissolved in water (the denser liquid) and placed into a separatory funnel with hexane (the less dense liquid). Shaking the separatory funnel enables caffeine to dissolve preferentially in hexane. The water layer is removed from the tap at the bottom of the separatory funnel. The remaining solution of caffeine in hexane can undergo **vaporisation** (Figure 6.2.5) so that the caffeine will remain in the evaporating basin and the hexane will be vaporised. Note that this must be carried out in a fume hood because hexane fumes are toxic.

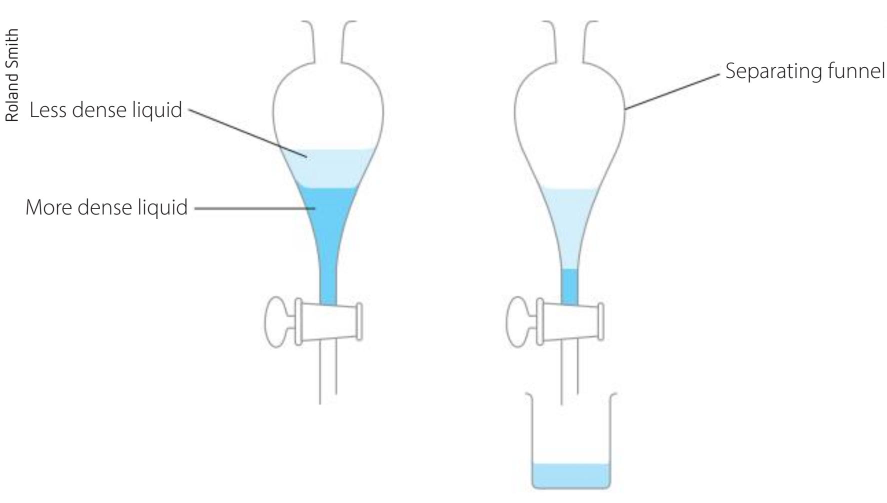


FIGURE 6.2.4 Separatory funnel

Roland Smith

FIGURE 6.2.5 Vaporisation

SCIENCE AS A HUMAN ENDEAVOUR

In 2014, Sarah Windsor and Amanda Neilen used separatory funnels to determine the concentration of an active ingredient (diethyl toluamide, DEET) in insect repellents. They needed eight separatory funnel extractions to remove DEET completely from the water solution into hexane to allow the concentration of DEET to be analysed by gas chromatography–mass spectrometry (GC-MS). They found that the DEET concentrations were 4% to 10% lower than the amount written on the bottle, meaning that the insect repellents protect us from insect bites for less time than labels suggest.

Liquid–liquid extraction

In the natural products industry, liquid–liquid extractions are used to extract bioactive compounds from natural materials. For example, in 2015 Daniel Meloncelli and co-workers used a combination of three extraction techniques to extract bioactives from the honey matrix:

1. ethyl ethanoate (acetate) was used to extract volatiles and semi-volatiles for GC-MS concentration analysis
2. ethanol was used to extract phenolic compounds for their identification using high-performance liquid chromatography (HPLC)
3. water doped with an ultraviolet light tag was used to extract antibacterial compounds for HPLC concentration analysis.

They found that Tasmanian honeys of different floral origins can be distinguished by their concentrations of antibacterial compounds such as methylglyoxal. The concentration of methylglyoxal is negligible in Leatherwood honeys but in Manuka honeys it ranges from 23 ppm to 307 ppm.

SECTION REVIEW 6.2

REMEMBERING

1 Define:
  - a distillate
  - b distillation
  - c immiscible liquid
  - d solute
  - e solvent
  - f vaporisation.

UNDERSTANDING

2 Discuss the similarities and differences between distillation and vaporisation.

6.3 Heterogeneous and homogeneous mixtures

homogeneous mixtures mixtures of uniform composition

heterogeneous mixtures mixtures of non-uniform composition

Homogeneous mixtures have uniform composition throughout. Non-uniform mixtures are called **heterogeneous mixtures**. Sometimes the non-uniformity of a heterogeneous mixtures is visible to the naked eye. For example, the minerals of granite are easily distinguishable: the pink grains are orthoclase feldspar, the white are quartz and the black grains are biotite/hornblende.

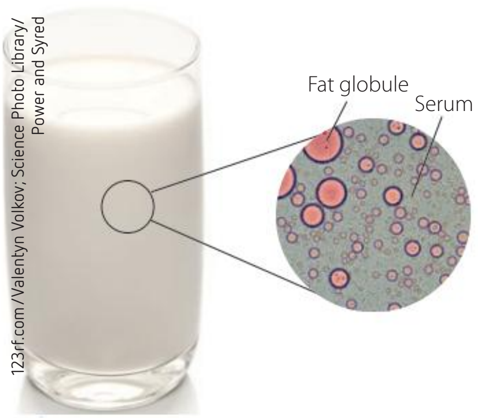


FIGURE 6.3.1 Milk appears to be homogeneous to the naked eye, but its inhomogeneity is revealed under an optical microscope.

Observing uniformity and non-uniformity

Sometimes you need to use an optical microscope to observe the non-uniformity of a heterogeneous mixture. You can see the non-uniformity of milk (fat globules suspended in serum) only under the microscopic view (Figure 6.3.1).

A comparison of water, salt water and soda water reveals that macroscopic observations can be both enlightening and misleading. You can see the carbon dioxide gas in the soda water easily, which leads to the correct classification of a heterogeneous mixture. Water and salt water are both colourless, transparent liquids; however, water is a pure substance composed only of the compound, water (H_2O), while salt water is a homogeneous mixture of solute salt dissolved in solvent water. In this instance, macroscopic observations were misleading. Measuring the boiling points can be used to distinguish between transparent, colourless liquids. At room temperature and pressure, the boiling point of water will always be 100°C, no matter how much water is left in the vessel. However, the boiling point of salt water will rise as the concentration of salt in the water increases. For example, a solution of seawater contains approximately 3.5 g of salt in every 100 g of seawater and an initial boiling point of 100.64°C, and this boiling point will continue to rise as the water is boiled off and the concentration of the remaining seawater increases.

6.3.1 How to identify heterogeneous and homogeneous mixtures

Carbon

amorphous shape a non-crystalline solid

Carbon constitutes 0.027% of Earth's crust. Some carbon occurs as an element. Some elemental forms of carbon have an **amorphous shape**. For example, carbon black is used in inks and tyres and charcoal is used in filters. Amorphous forms of elemental carbon convert to graphite at 2500°C and graphite converts to diamond at 100 000 atm and 3000°C (Figure 6.3.2). Diamonds are formed naturally at

6.3.2 Amorphous materials

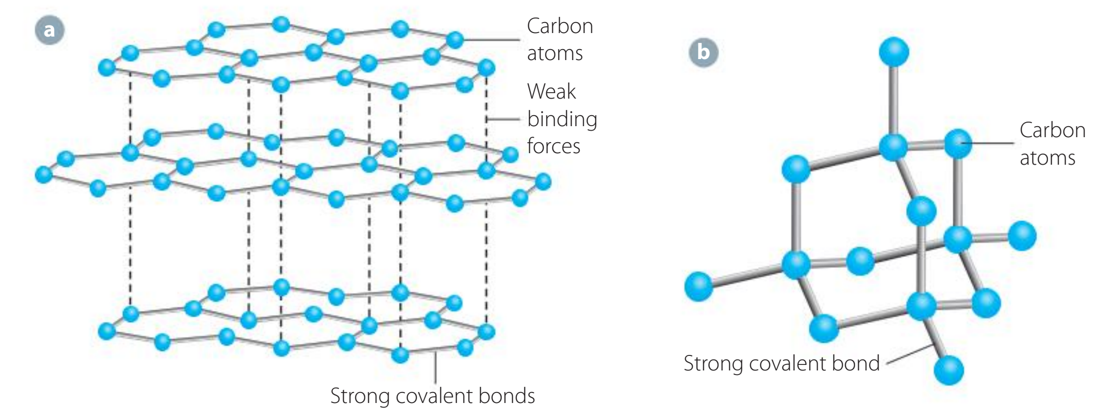


FIGURE 6.3.2 **a** Sheets of graphite and **b** covalent bonding in pure diamond

140–190 km below the surface of Earth. Diamonds can be synthetically produced in a laboratory by placing graphite at 50 000 atm and 1530°C for 4 days. Although only 0.01% of natural diamonds are coloured, it is exceptionally difficult to eliminate impurities from the synthetic process, particularly nitrogen, because it accounts for approximately 78% of atmospheric gases. Nitrogen impurities give diamonds a yellow colour and are an example of a homogeneous mixture. The structure of diamond doped with nitrogen impurities can be viewed only under an electron microscope.

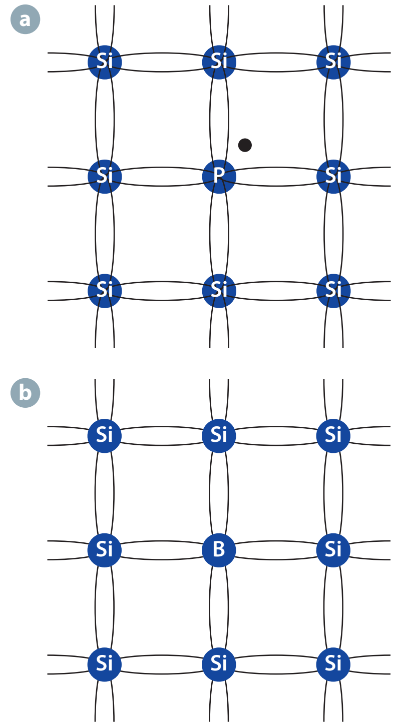


FIGURE 6.3.3 **a** The n-type region and **b** the p-type region of a photovoltaic solar panel

Doping impurities

The idea of doping impurities has been used to form photovoltaic solar panels. The upper n-type region and the lower p-type region of a photovoltaic solar panel are depicted in Figure 6.3.3. Like yellow diamonds, the structure of photovoltaic solar cells can be viewed only under an electron microscope, so they are considered homogeneous mixtures. Light strikes the upper n-type region of the solar panel and excites an electron in the phosphorus dopant, which donates this electron across the pn-junction to the acceptor dopant, boron, in the p-type region. This gives rise to a flow of electrons or, in other words, an electric current. The flow of current will give rise to a voltage difference between the n-type and p-type regions. The light energy is converted into electrical energy by connecting a load resistance across the two regions.

Homogeneous mixtures

One of the primary ways of modifying the properties of pure metallic elements is by alloying them with other elements in homogeneous mixtures. Pure gold is too soft to be used in jewellery; instead, it is alloyed with other coinage metals for jewellery. For example, an alloy that is 75% gold, 12.5% silver and 12.5% copper has a tensile strength > 10 times that of pure gold, > 7 times harder than pure gold and has the characteristic yellow colour associated with gold jewellery. **Substitution alloys** (Figure 6.3.4) are formed between metallic components that have atomic radii within 15% of each other (e.g. gold 4.078 Å, silver 4.085 Å, copper 3.615 Å). Steel is a homogeneous mixture of iron and carbon. However, the atomic radius of iron is 1.26 Å and the atomic radius of carbon is only 0.7 Å, so it forms an **interstitial alloy** (Figure 6.3.5). Changing the percentage composition of carbon in steel modifies its properties significantly. For example, < 0.2% carbon steel is malleable and ductile, so is used in cables and chains, but an increased carbon percentage increases its toughness; for example, 0.2–0.6 % carbon steel is used in girders and rails.

substitution alloy an alloy where the dopant has been substituted in the crystalline structure with the main atom

interstitial alloy an alloy where the dopant occupies the space between main atoms

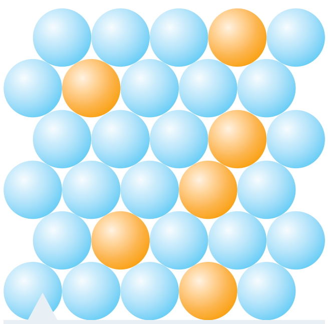


FIGURE 6.3.4 Substitution alloy

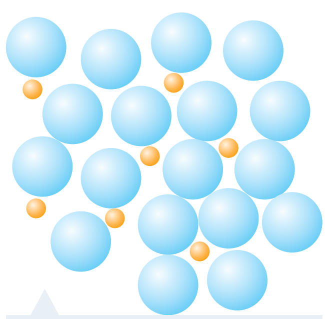


FIGURE 6.3.5 Interstitial alloy

SECTION REVIEW 6.3

REMEMBERING

1 Define:

a amorphous shape

b heterogeneous

c homogeneous

d interstitial alloy

e substitution alloy.

UNDERSTANDING

2 Discuss the similarities and differences between water, soda water and salt water.

APPLYING

3 Recommend the type of jewellery to wear for going out to dinner and swimming in a pool.

6.4 Nanomaterials

nanotechnology a branch of science dealing with particles in the range 1–100 nm

nanomaterial a substance that is made up of or incorporates particles in the range of 1–100 nm

Macroscopic observations made with the naked eye can discern objects at the millimetre scale (10^6 nanometres, Figure 6.4.1). Optical microscopes can distinguish an individual bacterium at the micrometre scale (1000 nanometres, Figure 6.4.1). Siemens produced the first commercial electron microscope in 1938. Electron microscopes have a resolution < 0.1 nm, so they are used in **nanotechnology** (1–100 nm) investigations of **nanomaterials**, such as viruses, antibodies and molecular structures (Figure 6.4.1).

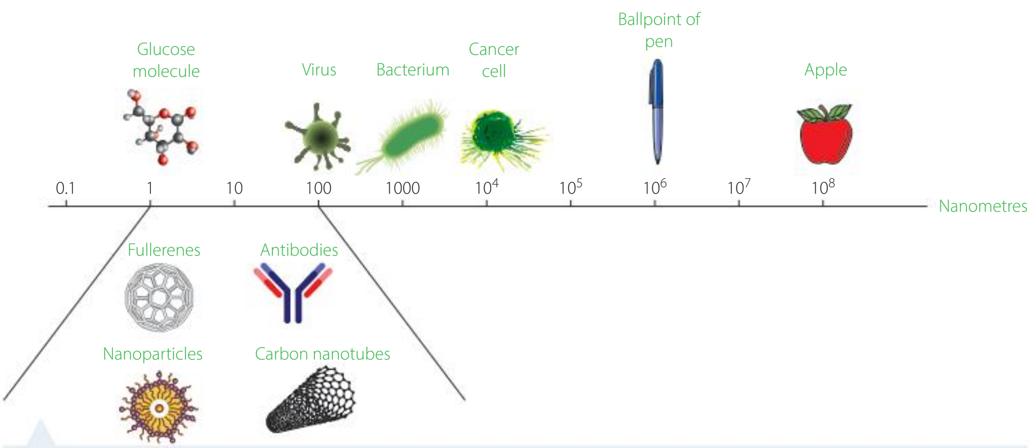


FIGURE 6.4.1 A log scale of objects in nanometres highlighting the nanomaterials: fullerenes, antibodies, nanoparticles and carbon nanotubes

fullerene a nearly spherical arrangement of 20–84 covalently bound carbon atoms

graphene a sheet of covalently bound carbon atoms

Fullerenes (Figure 6.4.1) were discovered by laboratory synthesis and mass spectrometry by Richard Smalley and co-workers in 1985. They were found to occur naturally in 1992 by Peter Buseck and co-workers in unusual carbonaceous rock in Russia. The first fullerene discovered contained 60 covalently bound carbon atoms in a nearly spherical arrangement of 32 faces, 12 pentagons and 20 hexagons, like a soccer ball. Since this initial discovery, fullerenes containing between 20 and 84 covalently bound carbon atoms have been identified. In medical applications, fullerenes have been used to encapsulate gadolinium(III) ions to reduce toxicity in magnetic resonance imaging and the performance and toxicity of iodinated fullerenes are being trialled for X-ray contrast of soft or vascular tissue. In 2004, 42 years after its first observation via electron microscope, **graphene** (Figure 6.4.2) was isolated and characterised by Andre Geim and Konstantin Novoselov. It is estimated that the graphene market will be worth around $280 million by 2020, predominantly in industries producing semiconductors and electronics. **Carbon nanotubes** (Figure 6.4.1) are graphene sheets rolled into a cylinder and capped at one or both ends with fullerenes. Single-walled carbon nanotubes are approximately 1000 nm long and 1 nm wide. Multi-walled carbon nanotubes consist of nested-together tubes within tubes. In 2013, Gaurav Lalwani and co-workers thermally crosslinked multi-walled carbon nanotubes into a macroscopic, three-dimensional, free-standing all-carbon scaffold that may be used for the next generation of energy storage, catalysis and biomedical devices and implants.

In the Middle Ages, stained glass window makers capitalised upon the deep red colour of finely divided nanoparticles of gold dispersed in molten glass. In 1857, Michael Faraday inadvertently made **colloidal** gold (Figure 6.4.3) when washing commercial gold leaf into thin enough sheets to become transparent. The washings yielded a fluid that was faintly ruby coloured as a result of the suspended gold particles. Although 20 nm gold particles melt at a far lower temperature than bulk gold and 2–3 nm gold particles become chemically reactive, the colloidal gold created by Faraday can still be observed in the Royal Institution of Great Britain's Faraday Museum in London. Colloidal gold is unusually stable – most colloids only last for a few months – but Faraday's samples are now more than 150 years old.

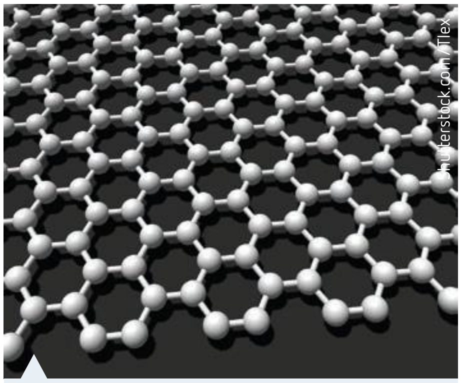


FIGURE 6.4.2 Sheet of graphene

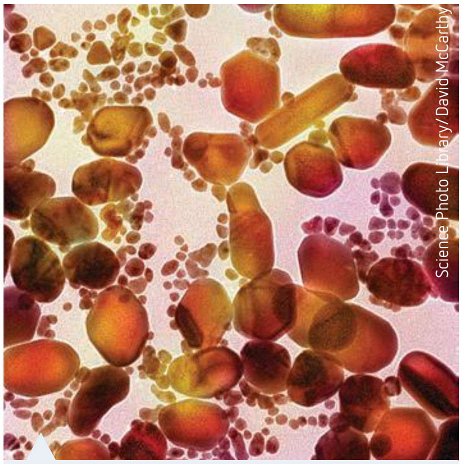


FIGURE 6.4.3 Colloidal gold containing nanoparticles of gold, which do not settle as a result of gravity, and are too small to be filtered

carbon nanotube
a rolled-up sheet of graphene capped with fullerene(s)

colloid
a mixture in which tiny clusters of particles are dispersed through another substance

6.4.1 What is nanotechnology?
6.4.2 The story of graphene

INQUIRING FURTHER

In 2007, Gibson and co-workers covalently functionalised 70 molecules of chemotherapeutic drug, paclitaxel, to individual 2 nm gold particles. Conduct research into optimising drug delivery via gold nanoparticle drug carriers in terms of the active drug being available at the target site for the correct duration and at the appropriate concentration.

SECTION REVIEW 6.4

REMEMBERING

1 Define:

- **a** carbon nanotube
- **b** colloid
- **c** fullerene
- **d** graphene
- **e** nanomaterials
- **f** nanotechnology.

UNDERSTANDING

2 Discuss the similarities and differences between different forms of carbon.

APPLYING

3 Apply your knowledge of intermolecular and intramolecular forces to determine why carbon nanoparticles are used in medicine and electronics.

4 Apply your knowledge of chemical reactivity to determine why gold is used in medical applications.

6.5 Inquiry skills: Analysing and interpreting data

PRACTICAL ACTIVITY 6.5.1

Separating sodium chloride and copper(II) chloride

You have been given a mixture of sodium chloride and copper chloride. Both substances have similar levels of solubility in water but one is more soluble in ethanol than the other. Use this fact along with what you have learnt about different separation techniques to design and conduct an investigation to separate these substances as effectively as possible.

AIM

Consider what you want to discover in this investigation

EQUIPMENT

Think about the equipment and chemicals will you need to complete the task. Be specific about quantities for all your materials; for example, chemicals and sizes of beakers. Remember to include basic materials such as matches, water and paper towel.

MANAGING RISKS

Construct a table similar to the one below. Identify the specific risks involved in the investigation and how to manage them to avoid injuries and damaging equipment. Ask your teacher to check your risk assessment before you proceed.

RISK ASSESSMENT

WHAT ARE THE RISKS IN DOING THIS EXPERIMENT?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?

HOW WILL YOU CONDUCT YOUR EXPERIMENT?

Hint: You will need to use the properties of solubility and boiling point. Write your method in a logical sequence of steps. Do not forget to be specific and include quantities.

WHAT RESULTS WILL YOU COLLECT?

Refer back to the aim you formulated. What data and observations will you need to record to determine whether you have achieved your aim?

WHAT HAVE YOU DISCOVERED?

Consider what your results show. Are they what you expected? Think about how to determine how effective your separation was.

WHAT DO YOU CONCLUDE?

Reflect on what you discovered to consider whether or not you achieved your aim. What do the results tell you about your investigation? How could you improve the effectiveness of the separation? Consider how successful your methodology was. What might you change or do differently to improve the separation?

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 Identify the major danger associated with hydrogen gas (H_2).

2 List and describe the equipment used during the determination of homogeneity of a liquid mixture.

3 Summarise what happens during a chemical process.

CATEGORY QUESTIONS

4 Explain the processes associated with separation of mixtures.

5 What places are associated with the desalination of water?

6 Identify the physical characteristics associated with pure and impure substances.

ELABORATION QUESTIONS

7 Investigate and outline the purposes associated with nanomaterials.

8 Investigate and outline the type and composition of the alloys associated with rose gold and white gold.

EVIDENCE QUESTIONS

9 Explain why helium is used in today's airships.

10 If you were to sell vitamins, would you focus your sales pitch on the partition coefficients of your vitamins or their potential illness preventative properties? Justify your decision.

END-OF-CHAPTER EXAM

End-of-chapter test

1 What observations would suggest that the following chemical reaction has occurred?

$$Cu(s) + 4H^+(aq) + 2NO_3^-(aq) \rightarrow Cu^{2+}(aq) + 2NO_2(g) + 2H_2O(l)$$

A A brown–orange gas evolved

B The liquid changed colour from colourless to blue.

C The solid disappeared.

D All of the above

2 What is Cu?

A An element

B A compound

C A homogeneous mixture

D A heterogeneous mixture

3 What is H_2O?

A An element

B A compound

C A homogeneous mixture

D A heterogeneous mixture

4 During the chemical reaction described in Question 1, the reaction vessel contains:

A An element.

B A compound.

C A homogeneous mixture.

D A heterogeneous mixture.

In Questions 5–7, fill in the gap.

5 ___________ cannot be broken down by either physical or chemical processes.

6 ___________ can be broken down by chemical processes.

7 ___________ can be separated by physical processes.

8 Distinguish between graphite and graphene.

9 In 2015, Sulway and George demonstrated an experiment at the Monash University ASELL workshop to determine the partition coefficient of ethanoic acid in two solvents: water and 1-octanol. Calculate this partition coefficient from the following data:

- Aqueous extraction volume in a separatory funnel = 25 mL
- Exact concentration of sodium hydroxide (NaOH) = 0.096 mol L^{-1}
- Average volume of NaOH added to the conical flask before extraction = 20.9 mL
- Average volume of NaOH added to the conical flask after extraction = 15.2 mL
- 1-octanol extraction volume in a separatory funnel = 25 mL
- Volume of aqueous ethanoic acid pipetted into the conical flask = 10 mL

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7 BONDING AND PROPERTIES

Introduction

All chemicals and substances can be classified and described according to their chemical and physical properties. These properties arise from how the atoms in a substance are bonded together. Chapter 3 outlined the different types of bonding that occur between different atoms. This chapter considers and explains the properties that substances containing each different type of bonding possess.

Stimulus questions

- Why do different substances have different physical and chemical properties?
- Why are most metals good conductors of electricity?
- What gives rise to the difference between graphite and diamond, both being pure carbon?
- Why do hydrocarbons usually burn readily in air?

Alamy Stock Photo/Elizabeth Nunn

7.1 Ionic compounds

ionic compounds compounds consisting of atoms that have lost and gained electrons to form charged ions that are electrostatically attracted to one another

7.1.1 Different substances and their properties

The physical properties of **ionic compounds**, such as sodium chloride, are different from the properties of the elements from which they are composed.

Most ionic compounds:

- are hard and brittle, so they do not scratch easily, but will shatter on impact
- do not conduct electricity in a solid state, but are good conductors in a molten or aqueous state
- have high melting and boiling points.

From these physical properties, scientists can infer that there are strong bonds between the particles in ionic compounds and they are made of charged particles, as shown by their ability to conduct electricity.

PRACTICAL ACTIVITY 7.1.1

Comparing the properties of a compound with those of its component elements

In this experiment, you will compare the properties of two elements with those of the resulting compound. The properties being compared experimentally are physical state, colour, odour, solubility in water, electrical conductivity, and reaction with hydrochloric acid. You may also consult data tables to add melting point and density to the comparison table.

AIM

To compare the properties of the elements magnesium and oxygen with its resulting compound, magnesium oxide

MATERIALS

- 2 strips of magnesium ribbon (20 cm and 1 cm)
- steel wool
- crucible tongs
- pipe clay triangle
- Bunsen burner
- tripod
- test tubes, stoppers and test-tube rack
- 250 mL beaker
- distilled water
- crucible and lid
- electrical conductivity apparatus
- matches
- 1 mol L^{-1} hydrochloric acid (HCl)

»

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WHAT ARE THE RISKS IN DOING THIS EXPERIMENT?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?
1 mol L^{-1} HCl is corrosive to skin and clothing.	Wear safety glasses and protective clothing. Take care when pouring and clean up spills immediately. If hydrochloric acid is spilt on the skin, wash the affected area with plenty of water and notify your teacher.
Magnesium burns with a very bright flame.	Do not look at the flame directly.
Magnesium oxide is an irritant.	Do not breathe in the powder; work in a well-ventilated area.
A hot crucible retains heat, especially if contents have significant mass.	Use tongs to handle the crucible; place hot crucible on the heatproof mat. Do not touch the hot crucible. If you burn yourself, place the affected area under cold running water for 10 minutes and inform your teacher.
The Bunsen burner will get hot.	Do not use the Bunsen burner if the gas tube is damaged. Ensure long hair is tied back and the flame is away from all flammable material. If you burn yourself, place the affected area under cold running water for 10 minutes and inform your teacher.

RISK ASSESSMENT

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, as well as ways to manage them.

PROCEDURE

1. Thoroughly clean the surface of both strips of magnesium ribbon with steel wool. Record the appearance of the cleaned magnesium.
2. Coil the longer piece of magnesium ribbon so that it fits inside the crucible.
3. Place the crucible on a pipe clay triangle or a tripod over the Bunsen burner and carefully heat the crucible without the lid until the magnesium begins to glow.

   Warning: Do not look directly at the burning magnesium.
4. Place the lid on the crucible with tongs and heat directly for about 10 minutes.
5. Remove the lid and heat for a further 5 minutes to ensure complete reaction.
6. Replace the lid and allow it to cool. This is the sample of magnesium oxide to be used for comparing properties.
7. Use the conductivity apparatus to test the electrical conductivity of magnesium metal, magnesium oxide and oxygen (air). Record the results.
8. Place 10 mL of distilled water in each of two test tubes. Add the 1 cm strip of magnesium to one and some of the magnesium oxide to the other. Stopper and shake. Record the results.
9. Consult data tables for the melting point and density of magnesium and magnesium oxide.
10. Add 10 mL of hydrochloric acid to each of three test tubes. To the first one, add the strip of magnesium ribbon. To the second, add some magnesium oxide. Record your observations. Stopper the third test tube and shake it to aerate the distilled water. Record your observations.

» **RESULTS**

Record your results in a table like the one below.

PROPERTIES	MAGNESIUM	OXYGEN	MAGNESIUM OXIDE
Physical state			
Colour			
Electrical conductivity			
Solubility in water		0.0359 g O_2 per litre of water at 25°C	
Reaction with HCl			
Melting point			
Density			

ANALYSIS OF RESULTS

What are the similarities and differences in the properties of the elements and compound?

DISCUSSION

What other properties could be compared?

CONCLUSION

Summarise your findings about the comparison of the properties of the elements and resultant compound.

Explaining properties of ionic compounds

In an ionic compound, each ion is strongly held in place by the attraction of adjacent oppositely charged ions. This bonding is described as non-directional because ions may be attracted to each other in any direction. The ions become arranged in a regularly repeating pattern, known as a **lattice**. Table 7.1.1 summarises the properties of ionic substances.

lattice
a regular repeating three-dimensional arrangement of ions, which results in the formation of crystals

TABLE 7.1.1 Properties of ionic compounds

PROPERTY	EXPLANATION
High melting point	There is strong attraction between ions, so a lot of energy is needed to move the ions out of their fixed position in the lattice.
Conducts electricity if in the molten state and in solution	Charged particles are free to move in the molten and aqueous state but are held in a fixed position in the solid states.
Hard but brittle	Ions are held rigidly in place in the lattice but the application of stress brings ions of the same charge closer together, so the crystal can shatter.
Mostly soluble in water	Ions are attracted to the water molecules and move out of position.
Giant lattice of repeating ions	Electrostatic attraction between positive and negative charges causes ions of opposite charge to surround each other.

malleable
a substance that can be repeatedly deformed, or hammered, without shattering

brittle
a substance that will shatter rather than bend, when a force is applied

The non-directional bonding holds the ions in place, preventing ionic crystals from being **malleable** like metals. Ionic crystals are **brittle**, and will shatter when hit with a strong force. Ionic bonds are strong, as shown by the hardness and high melting and boiling points of ionic substances.

Figure 7.1.1 (page 107) shows the information from Table 7.1.1 schematically, using sodium chloride as an example.

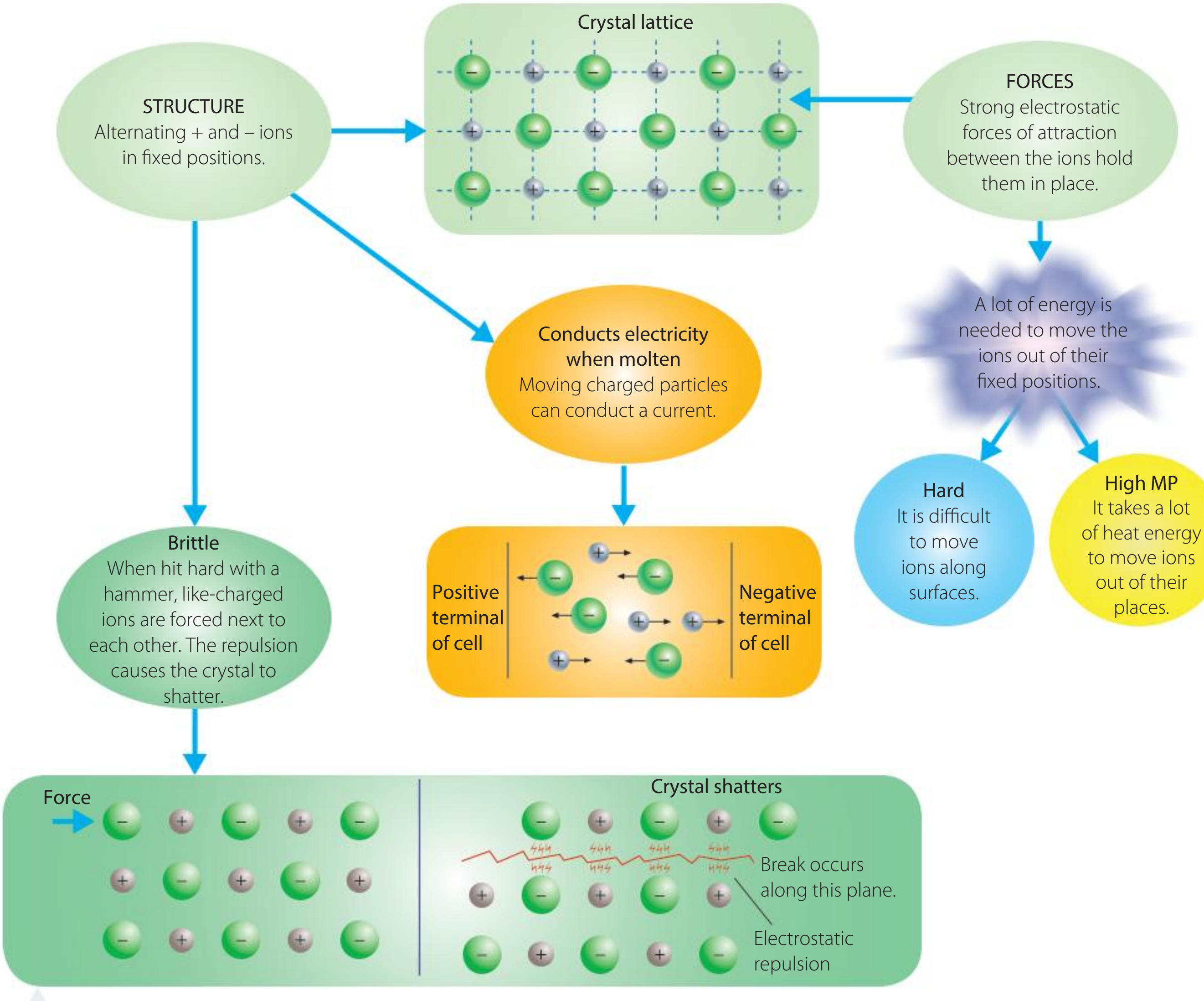


FIGURE 7.1.1 Explaining some of the properties of sodium chloride

SECTION REVIEW 7.1

REMEMBERING

1 Define 'ionic bonding'.

2 List three properties of ionic solids.

UNDERSTANDING

3 List four properties of ionic solids and explain these properties using the ionic-bonding model.

4 Explain why the formula of ionic solids is an empirical formula and does not give the actual number of ions present in the compound.

5 Explain why the bonding in an ionic solid is described as 'non-directional'.

ANALYSING

6 Two ionic solids have different melting points as shown below:

Sodium chloride: 801°C   lead iodide: 402°C

Use your knowledge of ionic bonding and the relative position of the elements involved on the periodic table to suggest why the melting points of these compounds are significantly different.

7.2 Metals

7.2.1 Metal structure and properties

SYLLABUS LINK

Chapter 3 discusses the model of bonding within a metal.

Elements are grouped on the periodic table on the basis of common properties. Metals make up 75% of the elements, while the remainder are non-metals or metalloids. Each of the three groupings has distinct physical properties by which it can be classified. The common properties of metals are listed in Table 7.2.1.

TABLE 7.2.1 Common properties of metals

PROPERTY	DESCRIPTION
Metallic lustre	Mirror-like shininess
Good conductor of heat	Allows heat to travel from one end to the other
Good conductor of electricity	Allows an electric current to easily pass through
Malleable	Able to be beaten into another shape or flattened into a thin sheet with a hammer without breaking
Ductile	Able to be drawn out into a wire

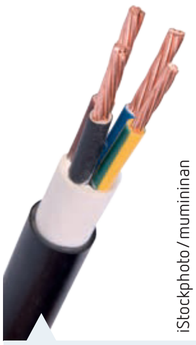


FIGURE 7.2.1 Copper wire is used in electrical wiring because it is a very good conductor of electricity.

SCIENCE AS A HUMAN ENDEAVOUR

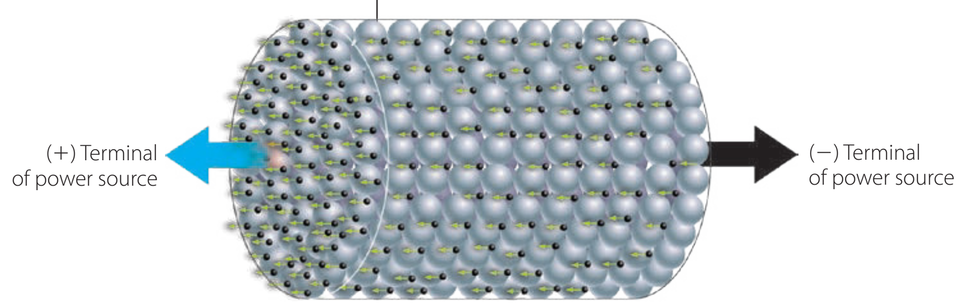
COPPER WIRE

Copper has been used in electrical wiring for communication since the telegraph services of the 1820s. Today, copper wire is still used for telephone, cable TV and ethernet connections (Figure 7.2.1). Advances in technology have led to changes in the way copper electrical wiring is structured and connected into circuits. Improvements have led to the development of copper wiring systems that can carry at least 1 gigabit (billion bits) of information per second, which equates to about 50 000 pages of text per second.

Conductivity

Metals are good conductors of electricity because of the highly mobile electrons within the lattice. When metals are connected into an electrical circuit, the mobile electrons have a net movement towards the positive terminal (Figure 7.2.2). Metals are also good conductors of heat because the mobile electrons acquire energy from the heat source and rapidly transfer it to cooler parts of the lattice (Figure 7.2.3, page 109).

FIGURE 7.2.2 Explaining the ability of a metal to conduct electricity



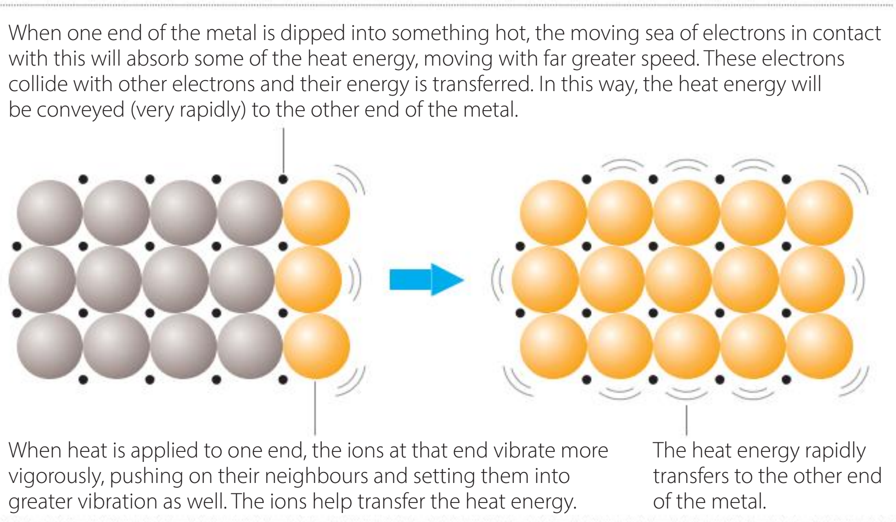


FIGURE 7.2.3 Explaining the ability of a metal to conduct heat

Lustre

A metal's lustre is explained by light rays being reflected off **delocalised electrons**. The close packing of the metal cations prevent light from passing through, which makes the metal opaque (Figure 7.2.4).

delocalised electrons
electrons that are not held in a fixed orbit around one nucleus, but are instead shared by many atoms

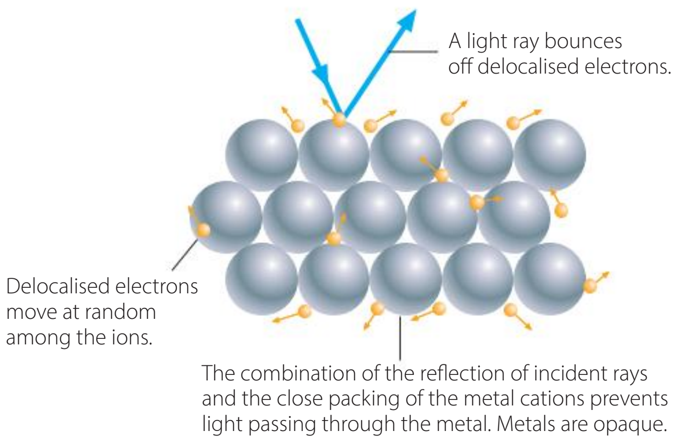


FIGURE 7.2.4 Explaining the lustre of a metal

Malleability and ductility

Metals are malleable and **ductile** because, when the orderly array of cations is sheared, the mobile electrons adjust to the new arrangement, allowing one layer of cations to slide over another without disrupting the metallic bonding (Figure 7.2.5).

ductile
a substance that can be drawn out to form a thin wire

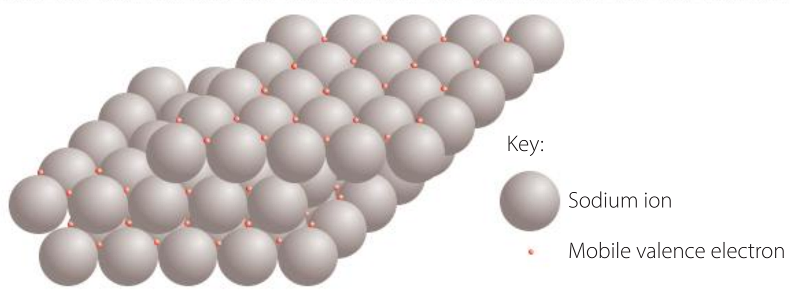


FIGURE 7.2.5 Explaining why metals are malleable and ductile

Properties that vary from metal to metal

Table 7.2.2 shows properties that vary from metal to metal.

TABLE 7.2.2 Some of the properties that vary from metal to metal

PROPERTY	MEANING	COMMENT
Density	The number of grams in each cubic centimetre of a material If a material is denser than water, then it will usually sink in water.	Density varies from metal to metal. For example, 1 cm^3 of gold weighs 19.3 g, but the same volume of aluminium weighs only 2.70 g. Gold is described as a very dense metal, while aluminium is considered a 'light' (not very dense) metal.
Hardness	Resistance to being scratched The hardest substance cannot be scratched by any other material.	Most metals are difficult to scratch. Gold, lead and calcium are examples of soft metals; that is, they are easily scratched.
Melting point	The temperature at which the material melts; that is, turns from a solid to a liquid	Melting point (MP) varies considerably from metal to metal. Mercury is already a liquid at room temperature but many metals require temperatures much higher than 1000°C to melt.
Tensile strength	A material's ability to withstand a stretching force; for example, the materials in cables that are suspended between poles must have high tensile strength	Tensile strength varies considerably from metal to metal. For example, steel (which is mostly iron) has greater tensile strength than aluminium.

The hardness of a solid is a measure of how difficult it is to scratch. The hardest known material in nature is diamond. When a substance is scratched, the particles in the surface layers are shifted out of position. For this to occur, the bonding forces between the particles in these layers must be disrupted. The more strongly a substance's particles are held in position, the harder it is to scratch. This is why scratch resistance is an indicator of the strength of bonds between particles. The fact that some metals are soft (e.g. gold) and easily scratched, while others are hard and resist scratching, supports the inference that bonding forces in metals vary in strength.

Metals also generally have high melting points and boiling points. High melting and boiling points mean a lot of energy is required to overcome forces holding atoms together in their position. Melting and boiling points are a good indicator of the strength of the forces holding particles together in the solid and liquid form. The high melting and boiling points of many metals reflects the strong metallic bond, while the low melting points of sodium and mercury indicate that they have weaker metallic bonds. Generally, the bonding forces in metals vary in strength.

Density is a measure of mass per unit volume. It is generally measured in grams per cubic centimetre ($g\,cm^{-3}$). Water has a density of 1 $g\,cm^{-3}$ so objects that are less dense will float on water while denser objects will sink.

The density of a metal depends on:

- how closely the ions are packed
- the volume of each ion
- the mass of each ion.

The volume and mass of the metal ions are related to their position in the periodic table. Metals with higher atomic numbers and more filled electron shells will have greater masses and volumes. Generally, it would be expected that elements of lower atomic numbers have lower density while those with higher atomic numbers have higher density. While the packing of ions in metals is highly organised, the way they are packed varies and this makes a difference to the amount of space between ions. Packing also affects the strength of the bonding forces – the more closely packed, the stronger the forces are within that metal.

Unlike metals, non-metals show a lot of variation in their properties. Non-metals tend to be classified by what they cannot do, such as conduct heat and electricity.

SECTION REVIEW

7.2

REMEMBERING

1 List three properties that are characteristic of metals.

2 Define, in reference to materials:

a brittle

b ductile

c malleable.

3 List three properties that vary between different metals.

UNDERSTANDING

4 Explain the following observations in terms of the metallic model.

a You can draw out metals into a wire.

b You should not stick a metal fork into a power socket.

c You pick up a metal spoon that has been sitting in a cup of hot water and burn your hand.

APPLYING

5 For each of the following purposes, suggest an important metallic property that makes it suitable for that purpose.

a Copper is used in electrical wiring.

b Iron is used in construction of bridges.

c Aluminium is used in saucepans.

d Jewellery is commonly silver and gold.

e Electric light filaments are made of tungsten.

ANALYSING

6 Use the information in the table to answer the following questions.

ELEMENT	MELTING POINT (°C)	THERMAL CONDUCTIVITY AT 25°C ($J\,s^{-1}m^{-1}K^{-1}$)	ELECTRICAL CONDUCTIVITY AT 25°C ($MS\,m^{-1}$)
Sodium	98	141	21
Magnesium	650	156	22
Aluminium	660	237	37
Copper	1085	401	58.4

The SI unit is Siemens per meter (Sm^{-1}). One Siemen per metre is one mho per metre where a mho is the unit for conductance (mho is the inverse of ohm, the unit for resistance).

a Compare the thermal and electrical conductivities of the metals. Is there a pattern in the data?

b Explain any pattern identified in part **a** in terms of delocalised electrons.

c Is a relationship apparent between position in the periodic table and melting point? Suggest why this may be so.

7.3 Covalent compounds

As discussed in section 3.4, non-metallic atoms are held together as elements and compounds by covalent bonding. Covalent substances can be divided into two distinct groups on the basis of their properties: **covalent molecular substances** and **covalent network substances**. In each of these substances, while the bonding remains covalent, the arrangement of the atoms and ultimately the properties of the resultant substances differs, as shown in Table 7.3.1.

covalent molecular substances composed of discrete molecules whose atoms are held together by intramolecular bonds

covalent network substances carbon and silicon compounds with structures that are three-dimensional networks of covalently bonded atoms

7.3.1 Covalent compounds: simple molecules

TABLE 7.3.1 Contrasting properties of covalent substances

COVALENT MOLECULAR SUBSTANCES	COVALENT NETWORK SUBSTANCES
Low melting and boiling points (many are liquids and gases at room temperature)	Very high melting points (usually solids at room temperature)
Non-conductors of electricity in solid and liquid states	Non-conductors of electricity in solid and liquid states
Form solids that are generally soft	Form extremely hard and brittle solids
Variable solubility	Insoluble in water and most other solvents
Variable reactivity	Chemically unreactive
Examples are oxygen, water and propane	Examples are diamond and quartz (silicon dioxide)

Covalent molecular substances

intramolecular bonding bonds that occur between atoms inside a molecule

intermolecular bonding interactions that exist between molecules – these are much weaker than intramolecular bonds but are important in determining the physical properties of the molecule such as melting and boiling points and solubility

Covalent molecular substances, whether elements or compounds, are composed of discrete molecules. Covalent bonding holds the atoms in the molecule together and is referred to as **intramolecular bonding** because it operates within the molecules.

The forces between molecules are quite weak compared to those within the molecule. These weak forces act to bond molecules together and are called **intermolecular bonding**. The nature of intermolecular bonds will be discussed in Chapter 13. Table 7.3.2 shows the properties of covalent molecular substances in terms of their bonding.

TABLE 7.3.2 Explaining properties of covalent molecular substances

COVALENT MOLECULAR SUBSTANCES	EXPLANATION
Low melting and boiling points (many are liquids and gases at room temperature)	The forces that hold the molecules together in the solid and liquid state are weak.
Non-conductors of electricity in solid and liquid states	There are no charged particles that can move through the substance.
Form solids that are generally soft	The forces that hold the molecules together in the solid state are weak.
Tend to be malleable rather than shatter	The forces between molecules are weak so molecules are easily moved relative to each other.
Variable solubility	Solubility depends on the intermolecular forces, which vary between substances.

Examples of covalent molecular substances include methane, water, and carbon dioxide. Although some variation occurs among the properties of different covalent molecular substances, the intramolecular forces inside all of them are significantly stronger than the intermolecular forces between them.

Covalent network substances

Carbon and silicon both have a valency of four. This means they can form four covalent bonds. As a result, they can form three-dimensional networks of covalently bonded atoms. In these networks, the covalent bonding extends indefinitely throughout the whole crystal. The three-dimensional nature of the network means the atoms are held rigidly in position by very strong bonds.

Elemental carbon exists as a range of different forms or **allotropes**, including graphite, diamond and fullerenes, with significantly different structures and physical properties.

Carbon in the form of diamond is an example of a covalent network. Each carbon atom is covalently bonded to four other carbon atoms in an orderly pattern that continues throughout the whole structure to make a giant crystalline lattice (Figure 7.3.1). The bonding electrons are tightly bound and highly localised.

allotrope
different forms of the same element, with different physical properties; for example, diamond and graphite are both different forms of carbon

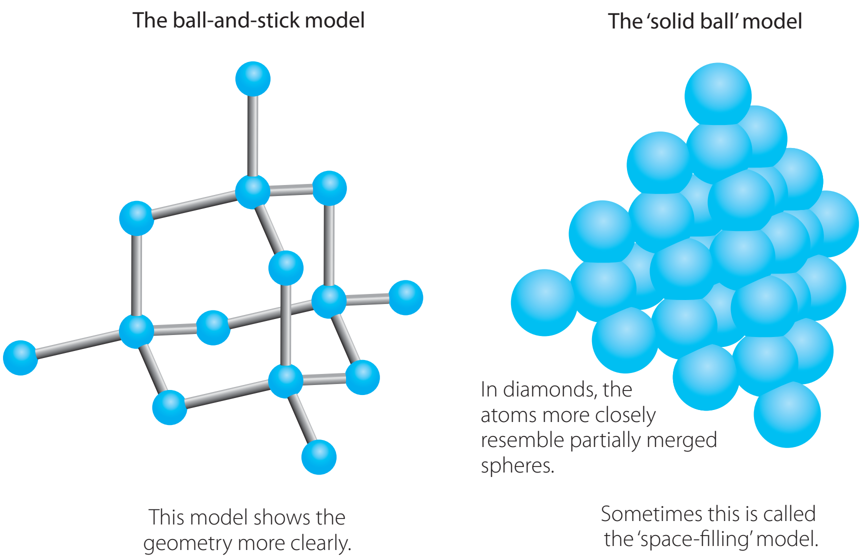


FIGURE 7.3.1 Models of the structure of diamond

SYLLABUS LINK
Chapter 6 also discusses forms of carbon.

Carbon's properties in diamond form are consistent with those listed in Table 7.3.1 (page 111). Silicon has the same valency as carbon and exists in a structure similar to diamond.

There are also compounds that exist as covalent network substances. Silicon and carbon covalently bond to form the crystalline compound silicon carbide (SiC), commonly known as carborundum. This compound has a rigid diamond-like structure with alternating silicon and carbon atoms.

Silicon dioxide (SiO_2), also known as quartz, is another, more common covalent network compound. Each silicon atom is bonded to four oxygen atoms, and each oxygen atom is bonded to two silicon atoms (Figure 7.3.2). This pattern extends throughout the entire crystalline lattice. Quartz is the main component of sand. The properties of silicon dioxide and silicon carbide are also consistent with Table 7.3.1.

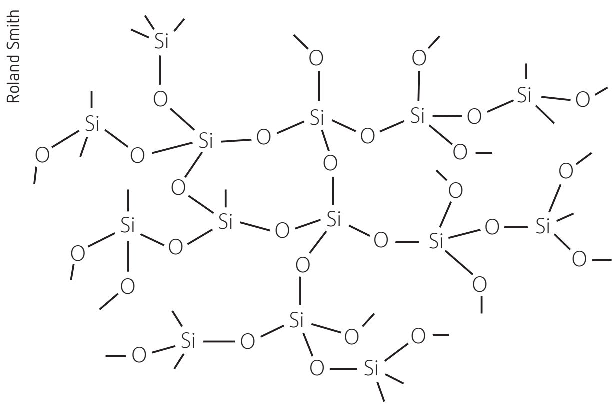


FIGURE 7.3.2 The covalent network structure of silicon dioxide

In a covalent network substance, the forces are strong and extend through the whole network, which explains why their properties differ from those of covalent molecular substances. Table 7.3.3 shows the properties of covalent network substances.

TABLE 7.3.3 Explaining properties of covalent network substances

COVALENT NETWORK SUBSTANCES	EXPLANATION
Very high melting points (usually solids at room temperature)	The covalent bonds that hold the atoms together are strong.
Non-conductors of electricity in solid and liquid states	There are no charged particles that are free to move throughout the structure.
Extremely hard and brittle solids	The bonds are very strong between atoms so it is difficult to scratch, but an impact force disrupts the positions of the atoms and causes the network to shatter.
Insoluble in water and most other solvents	There is no attraction between the atoms in the network and water molecules.

Other covalent substances

Graphite

Although graphite is a covalent compound and an allotrope of carbon, it has different properties from those of diamond. Like diamond, graphite has a high melting and boiling point, but graphite is soft and a good conductor of electricity.

7.3.2 Graphite

In graphite, the carbon atoms are arranged in flat parallel layers. Within each layer, each carbon atom is covalently bonded to three other carbon atoms, forming a hexagonal arrangement. This means that each layer is a two-dimensional network of carbon atoms. Weak bonding holds the layers together (Figure 7.3.3).

In graphite, only three of carbon's four electrons are involved in covalent bonding. The fourth electron is delocalised, which means it is free to move along the layer occupied by its 'parent' atom. The attraction between the delocalised electrons and the layers above and below it hold the layers together.

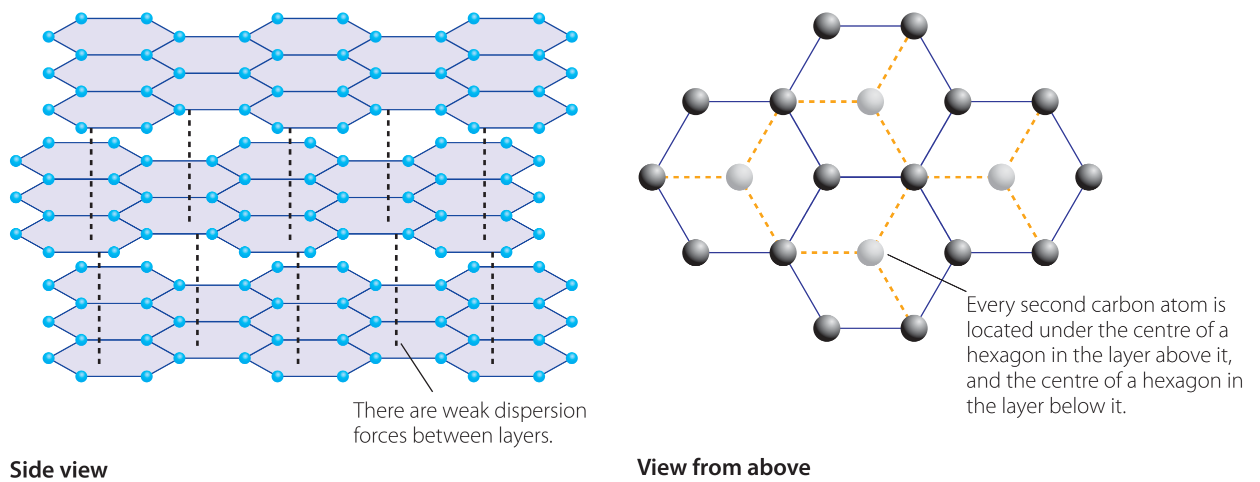


FIGURE 7.3.3 Modelling the structure of graphite

Graphite's high melting and boiling points can be explained by the strong covalent bonds within the layers, while its softness and slippery feel can be explained by the weak bonds between layers that allow them to easily slide over each other. Graphite is an electrical conductor because of the presence of the delocalised electrons that can move and conduct electricity along the layers.

Fullerenes

Although fullerenes are not covalent network substances, they are an allotrope of carbon and are held together by covalent bonds so it is appropriate to include them here.

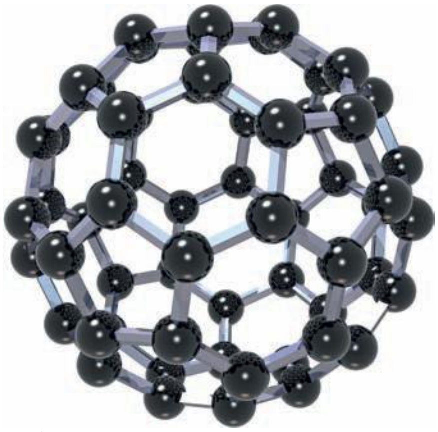
7.3.3 Fullerenes

The first fullerene molecule was discovered in 1985 by a group of researchers at Rice University, Texas, United States. It was found to have the formula C_{60} and was named after Buckminster Fuller because it resembles the geodesic domes he designed. The molecules are more commonly referred to as buckyballs.

Fullerenes have been found in nature and in space. The discovery of fullerenes expanded the number of known carbon allotropes. They have become the subject of intense research, especially in materials science, electronics and nanotechnology.

Fullerenes are molecules composed entirely of carbon. They are similar in structure and bonding to graphite. Each carbon is covalently bonded to three other carbons and the fourth electron is delocalised and free to move throughout the structure. They differ because they contain pentagonal and hexagonal rings that prevent them being planar. The most common fullerene is C_{60}. The structure resembles a soccer ball and is often referred to as an open cage structure (Figure 7.3.4).

Like graphite, fullerenes are soft and slippery because of the weak forces between molecules. They do not conduct electricity because there is no movement of electrons from one molecule to the next. Like diamond and graphite, they are insoluble in water because there is no attraction between the carbon atoms and water molecules. They typically have low melting points due to weak bonding between molecules.



Science Photo Library/Russell Kightley

FIGURE 7.3.4 A model of the C_{60} fullerene

SECTION REVIEW 7.3

REMEMBERING

1 List three differences in the properties of covalent molecular substances and covalent network solids.

2 Identify the elements most likely to form covalent network solids.

UNDERSTANDING

3 Explain why covalent molecular substances and covalent network substances have different properties when they both have covalent bonding.

4 Explain why graphite can conduct electricity while diamond cannot.

5 Explain why graphite and diamond both have high melting points, yet graphite is soft and diamond is extremely hard.

6 Identify the similarities and differences in the properties of graphite and fullerenes.

ANALYSING

7 Most substances that have an odour are covalent molecular substances. Suggest a reason for this.

7.4 Hydrocarbons

organic chemistry the study of carbon containing compounds

hydrocarbons compounds containing only hydrogen and carbon

The vast majority of compounds known to chemists are carbon compounds, because carbon is found in more places, in more forms, than any other element. A branch of chemistry, **organic chemistry**, is devoted to carbon-based compounds. A major reason for this is that carbon readily bonds to other carbon atoms (as seen in its allotropes) and each atom can form a total of four bonds from single to double to triple.

Chemists divide organic compounds into 'families' based on their structural features. This makes managing the huge number of carbon compounds easier. Families are organised according to which atoms or groups of atoms are bonded to carbon, as well as whether multiple bonds are present.

Hydrocarbons are compounds made up of only hydrogen and carbon. The simplest group of hydrocarbons are the alkanes.

Alkanes

7.4.1 Alkanes

The alkane family is very important because it includes the fuels that supply us with energy every day. Most Australians use alkanes to run our cars, supply us with hot water, heat our homes and cook our foods. Alkanes also supply us with the raw materials to make many other substances, including plastics.

The simplest alkane is methane (CH_4). Ethane (C_2H_6), propane (C_3H_8) and butane (C_4H_{10}) are next in the series. Figure 7.4.1 shows models and structural formulas for propane and butane.

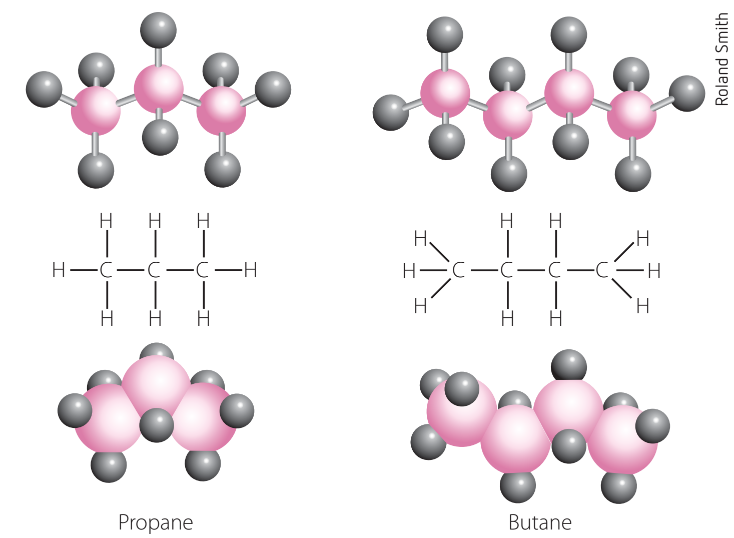


FIGURE 7.4.1 Models and structural formulas for propane and butane

Table 7.4.1 lists the first 10 alkanes. They are called *straight-chain alkanes* because the carbon atoms are joined in one continuous string. The chain is not straight but more like a zig-zag because of the way carbon bonds are arranged tetrahedrally.

TABLE 7.4.1 First 10 members of the alkane family

NAME	FORMULA
Methane	CH_4
Ethane	C_2H_6
Propane	C_3H_8
Butane	C_4H_{10}
Pentane	C_5H_{12}
Hexane	C_6H_{14}
Heptane	C_7H_{16}
Octane	C_8H_{18}
Nonane	C_9H_{20}
Decane	$C_{10}H_{22}$

Alkanes are also classified as **saturated compounds**, which means they contain only single bonds between carbon atoms.

Each successive member of the series is formed by replacing one H with a C and then adding three more hydrogens. This pattern follows for all straight-chain alkanes, so the general formula for alkanes can be expressed as C_nH_{2n+2}, where n is a whole number.

saturated compound
an organic compound where each carbon is bonded to four different atoms, without any double bonds present

Naming alkanes

As with ionic and covalent compounds, there is a systemic naming system for alkanes that allows us to work out the formula and structure of the compound.

The name of a carbon compound consists of a stem, which gives the length of the carbon chain, and an ending, which denotes the family of hydrocarbons the compound belongs to. All compounds that are in the alkane family have names ending in '–ane'. The stems of the first 10 members of the series vary according to the number of carbon atoms in the chain, as outlined in Table 7.4.2.

TABLE 7.4.2 The stems of organic compounds, according to the number of carbon atoms in the chain

NO. C ATOMS	1	2	3	4	5	6	7	8	9	10
Stem	Meth–	Eth–	Prop–	But–	Pent–	Hex–	Hept–	Oct–	Non–	Dec–

The alkanes considered so far have been straight-chain alkanes; however, many carbon compounds contain branched chains. In branched-chain compounds, at least one carbon atom is attached to more than two other carbon atoms.

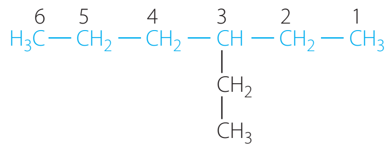
The rules for naming branched alkanes are as follows.

1. The end of the name indicates the hydrocarbon family to which the compound belongs. Alkanes always end in '–ane'.
2. Determine the longest continuous carbon chain and use the name of the corresponding alkane. This chain is the main chain and represents the parent structure. For example, if the longest chain had six carbon atoms, the parent structure would be hexane.

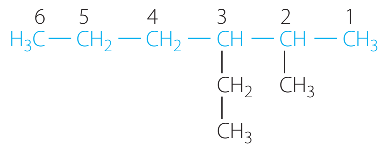
substituent
a carbon atom (or group of carbon atoms) that is not part of the longest chain of carbon atoms in an organic molecule

3 Other atoms or groups of atoms attached to the main (parent) chain are **substituents** and form 'branches'. If the substituent is a carbon group, it is called an alkyl group. The alkyl group is named according to the number of carbon atoms and given the ending '–yl'. For example, a substituent group CH_3CH_2– would be called ethyl. Other common substituents are F–, fluoro; Cl–, chloro; Br–, bromo; I–, iodo; and NO_2–, nitro.

4 Number the carbon atoms in the main (parent) chain so the branch or branches have the lowest possible number/s (numbering can be either from right to left or left to right). The position at which the group is attached to the main (parent) chain is specified by the number of the carbon to which it is attached. (The number is separated by a hyphen.) In the example below (numbered right to left), the branch is 3-ethyl and the compound is 3-ethylhexane. (Note: The alkane name is written as one word.)

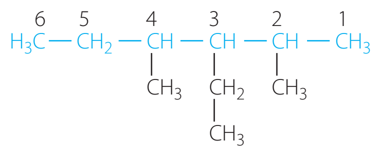


5 List the names of the substituents in alphabetical order (ignore di–, tri– etc.). In the example below, there are two branches: 3-ethyl and 2-methyl. The 3-ethyl group is listed before the 2-methyl. Words and numbers are separated by a hyphen.



The name of this compound is 3-ethyl-2-methylhexane.

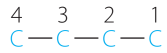
6 Where there is more than one substituent of the same type, use the prefixes di- (two), tri- (three), tetra- (four), penta- (five), and so on. Separate numbers by a comma.



The above compound contains two methyl groups so 'dimethyl' will be part of its name, which is 3-ethyl-2,4-dimethylhexane.

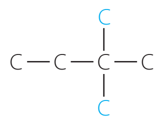
Drawing structural formulas of branched alkanes

1 Use the main name to draw the longest carbon chain and number the carbon chain, starting from 1 from any end of the chain. For example, in 2,2-dimethylbutane, the main chain is butane and the carbon atoms are numbered from 1 to 4.

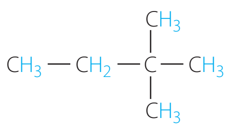


2 Identify the substituents that form the branches and to which carbon atom they are attached. For example, in 2,2-dimethylbutane, there are two methyl groups, both attached to carbon number 2.

3 Add the substituent group(s) to the parent chain.



4 Add hydrogen atoms to complete the structure. Check that each carbon atom has four bonds.



Properties of alkanes

Alkanes are covalent molecular substances, so their properties will be similar to those previously identified for covalent molecular substances. Table 7.4.3 summarises some physical properties of alkanes. These properties can be explained by the covalent molecular nature of the molecules.

TABLE 7.4.3 Some physical properties of alkanes

PROPERTY	ALKANES
Appearance	Colourless and odourless
Melting point and boiling point	Low melting point and boiling point, which increases with molecular mass, the first four being gases at room temperature
Electrical conductivity	Non-conductors
Hardness	Solid alkanes are soft
Malleability	Solid alkanes are malleable
Solubility	Insoluble in water but soluble in some other solvents

Chemical reactivity of alkanes

Alkanes do not react with most substances. They only have two important reactions: combustion and substitution.

combustion
exothermic chemical reaction between a fuel and an oxidising agent (usually oxygen) that produces energy

1 **Combustion**: Combustion is the most significant reaction that alkanes undergo. When ignited in a plentiful supply of oxygen, alkanes react readily to produce water, carbon dioxide and a large quantity of heat. This makes alkanes important as fuel. The examples below show propane (C_3H_8) and octane (C_8H_{18}) reacting with oxygen:

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

Propane + oxygen → carbon dioxide + water

$$2C_8H_{18}(g) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$$

Octane + oxygen → carbon dioxide + water

substitution
a reaction where one atom of a compound is replaced by an atom of a different element or a group of atoms

2 **Substitution**: In a substitution reaction, a hydrogen atom is replaced with an atom of another element. This reaction usually occurs only with chlorine or bromine and requires sufficient energy to be supplied. If the mixture is subjected to ultraviolet (UV) light, the following reaction will occur.

$$CH_4(g) + Cl_2(g) \rightarrow CH_3Cl(g) + HCl$$

Methane + chlorine → chloromethane + hydrochloric acid

Substitution reactions can continue until all the hydrogen atoms in the compound have been replaced by a halogen. In methane's case, the final product would be CCl_4, which is tetrachloromethane. The name of the product gives the number of carbon atoms (methane) and the number of halogen atoms (tetrachloro). Substitution reactions can also occur when an atom is replaced by a group of atoms, such as a methyl group (CH_3).

Alkenes

The **alkene** family contains hydrocarbons in which one pair of carbon atoms is joined by a double bond and all the other carbon atoms are joined by single bonds. All compounds that are in the alkene family have names ending in '–ene'. The simplest alkene is ethene (C_2H_4) and the next member is propene (C_3H_6). The structures of these two alkenes are shown in Figure 7.4.2.

unsaturated compounds
organic compounds containing at least one double bond between carbon atoms

Compounds that contain multiple bonds between carbon atoms, such as alkenes, are called **unsaturated compounds**.

Table 7.4.4 shows the names and formulas of the first nine alkenes.

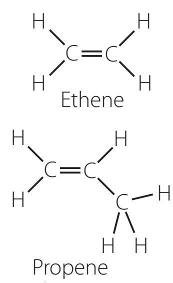


FIGURE 7.4.2 Ethene and propene

TABLE 7.4.4 The first nine members of the alkene family

NAME	FORMULA
Ethene	C_2H_4
Propene	C_3H_6
Butene	C_4H_8
Pentene	C_5H_{10}
Hexene	C_6H_{12}
Heptene	C_7H_{14}
Octene	C_8H_{16}
Nonene	C_9H_{18}
Decene	$C_{10}H_{20}$

7.4.2 Alkenes

Alkenes have a general formula of C_nH_{2n}. As with alkanes, each successive member of the series is formed by adding CH_2 to the formula of the previous member.

SCIENCE AS A HUMAN ENDEAVOUR

ETHENE

Ethene gas is a ripening hormone that stimulates the production of certain enzymes that cause fruit to ripen. The sweet smell produced by ripe fruit is ethene gas. To make fruit last longer, new plastics have been developed that absorb ethene, preventing the fruit from ripening.

Naming alkenes

Naming alkenes is not as straightforward as naming alkanes. This is because, in some alkenes, the double bond can occupy a number of different positions. Although this does not change the molecular formula of the compound, it does change the structural formula, which shows the position of the double bond. Figure 7.4.3's example demonstrates this.

Compounds that have the same molecular formula but different structural formula are called **structural isomers**.

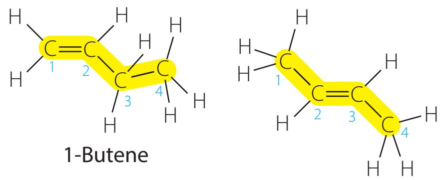


FIGURE 7.4.3 The structural isomers of butene

When naming alkenes, the location of the double bond needs to be included. Alkenes are named according to the following rules.

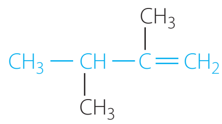
1. Take the usual stem name (such as 'eth-' and 'but-'), which indicates the number of carbon atoms in the chain, and add the ending '–ene'.
2. Determine the location of the double bond by numbering the carbon atoms from the end of the chain closest to the double bond (Figure 7.4.3).
3. Show the double bond's location by inserting, as a prefix, the number of the carbon at which the double bond starts, followed by a hyphen (-).

To name branched *alkenes*, add the following rules to those for naming branched *alkanes*.

4. When naming alkenes, the longest continuous carbon chain containing the carbon–carbon double bond becomes the main (parent) chain. The name corresponding to the alkane stem with the ending '-ene' forms the main name.

structural isomers
two compounds that contain the same number of each atom, but are arranged in a different way

5 Branched alkenes are named in a similar way to branched alkanes, except the lowest number is assigned to the double bond and *not* the substituent group.



2,3-Dimethyl-1-butene

Properties of alkenes

As might be expected, the physical properties of alkenes are similar to those of alkanes. As a family, alkenes show similar trends in melting point, boiling point and state under normal conditions. At room temperature, the first three alkenes (C_2–C_4) are gases, C_5–C_{15} are liquids, and the remainder are solids. Similar to alkanes, all alkenes are insoluble in water but will dissolve in some other solvents.

Chemical reactivity of alkenes

The presence of the double bond makes alkenes very reactive. They have two important reactions: combustion and addition.

Like alkanes, alkenes undergo combustion in a plentiful supply of oxygen to produce carbon dioxide and water. For example:

$$C_4H_8(g) + 6O_2(g) \rightarrow 4CO_2(g) + 4H_2O(g)$$

$$\text{Butene} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water}$$

addition reaction
a reaction where two molecules combine, resulting in a larger molecule being formed

Alkenes undergo a reaction that adds more atoms to the compound by breaking the double bond. This type of reaction is called an **addition reaction**. Alkenes undergo many addition reactions. Some of the more important reactions are as follows.

- Reaction with hydrogen: The alkene is converted to an alkane with the aid of a catalyst. For example, ethene is converted to ethane:

$$CH_2{=}CH_2(g) + H_2(g) \rightarrow CH_3CH_3(g)$$

$$\text{Ethene} + \text{hydrogen} \rightarrow \text{ethane}$$

- Reaction with a halogen: When a halogen such as chlorine or bromine reacts with an alkene, the halogen atoms add across the double bond. For example, 1-propene reacts with chlorine to produce 1,2-dichloropropane:

$$CH_3{-}CH{=}CH_2(g) + Cl_2(g) \rightarrow CH_3CHCl{-}CH_2Cl(g)$$

$$\text{1-Propene} + \text{chlorine} \rightarrow \text{1, 2-dichloropropane}$$

- Reaction with water: In the presence of a sulfuric acid catalyst, water adds an H and an OH across the double bond to produce a different group of organic compound – an alcohol. For example, ethene is converted to ethanol:

$$CH_2{=}CH_2(g) + H_2O(g) \rightarrow CH_3{-}CH_2OH(l)$$

$$\text{Ethene} + \text{water} \rightarrow \text{ethanol}$$

- Polymerisation: In the presence of a suitable catalyst and reaction conditions, ethene molecules will react with themselves to form long chains. This type of reaction is called addition polymerisation.

aromatic hydrocarbons
hydrocarbon containing one or more benzene rings

Benzene

Benzene is the simplest member of the family of **aromatic hydrocarbons**. This group originally got their name because many of them have a strong aroma (smell).

Benzene is a colourless liquid with the molecular formula C_6H_6, so it must be an unsaturated molecule. The structure is a six-sided flat ring with all bond angles of 120°. The carbon atoms are joined in a regular hexagon with one hydrogen atom attached to each carbon (Figure 7.4.4).

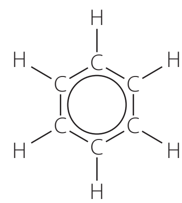
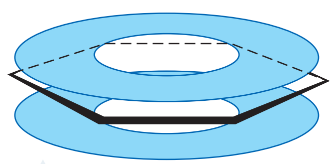


FIGURE 7.4.4 Structure of a benzene ring

On the basis of this representation, it was thought that benzene contained alternating single and double bonds, so they would undergo addition reactions similar to the alkenes and cycloalkenes. However, this rarely happens. The reactions of benzene are also completely different from those of other double-bonded hydrocarbons.

Further investigation of the benzene ring found that all the carbon–carbon bonds had the same bond length, somewhere between a single bond and a double bond. This can be explained if the electrons of the double bond are spread out around the ring of carbon atoms, making a cloud of delocalised electrons above and below the ring, shown schematically in Figure 7.4.5. This is similar to the delocalisation of electrons found in the structure of metals.

In chemical structure representations, the benzene ring is usually represented in either of the two forms shown in Figure 7.4.6.



Roland Smith

FIGURE 7.4.5 Bonding in benzene, shown schematically

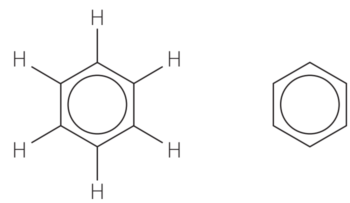


FIGURE 7.4.6 Different ways of drawing benzene

Properties of benzene

As with alkanes and alkenes, benzene has a low melting point, boiling point and density. It is also insoluble in water.

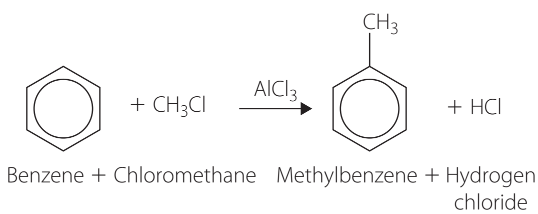
Chemical reactivity of benzene

The cyclic delocalisation of electrons in the benzene molecule makes it extremely stable. Benzene tends to undergo reactions in which the stable ring is preserved because reactions that disrupt the bonding due to the delocalised electrons are less favourable and will only occur at higher temperatures and under more vigorous conditions.

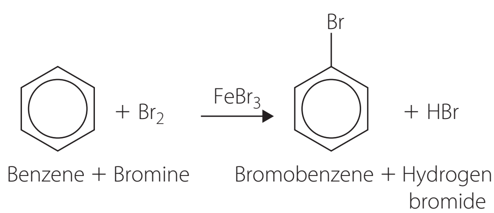
Most reactions of benzene are substitution reactions (like alkanes) where one or more of the hydrogens are replaced by a different atom or group of atoms.

Some examples of substitution reactions that benzene undergoes are as follows.

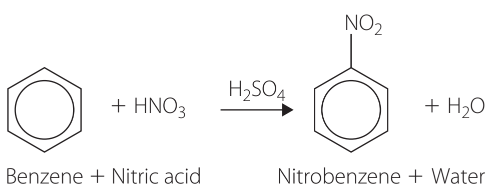
- Alkylation: substitution with an alkyl group



- Halogenation: substitution with a halogen



- Nitration: substitution with a nitro group



REMEMBERING

1 Give the general formula for:

a alkanes **b** alkenes **c** benzene.

2 Distinguish between saturated and unsaturated compounds.

3 Explain why benzene is extremely stable.

UNDERSTANDING

4 Identify each of the following as alkanes or alkenes. Justify your choice in each instance.

a C_6H_{10} **b** C_6H_{12} **c** C_6H_{14}

5 Draw a structural formula for each of the following.

a Butane **b** Octane **c** 1-Propene

d 3-Hexene **e** 2-Methylpentane **f** 2,3-Dimethylbutane

g 2-Methyl-2-pentene **h** 3,3-Dimethyl-1-butene

6 a Explain why the physical properties of alkanes and alkenes are similar but their chemical properties are different.

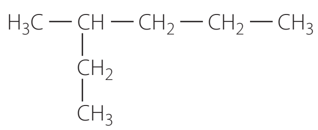
b Would you expect benzene to be soluble or insoluble in water? Justify your response.

7 Identify each of the following compounds.

a C_4H_{10} **b** C_9H_{20}

c $H_3C-CH_2-CH=CH-CH_3$ **d** $H_2C=CH-CH_2-CH_3$

e



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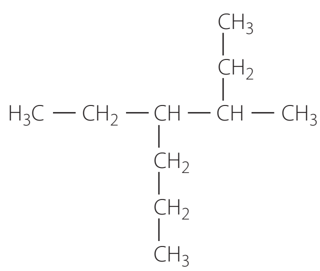

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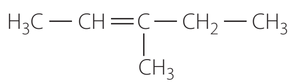

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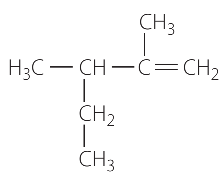


8 Distinguish between substitution and addition reactions.

APPLYING

9 Draw the structural formula and name the straight chain isomers of C_8H_{16}.

10 Write balanced equations for:

a pentane combustion

b butane reacting with chlorine gas (under UV light)

c propene reacting with water

d ethene reacting with bromine

e benzene reacting with chlorine

f the alkylation of benzene with $CH_3CH_2CH_2Cl$.

ANALYSING

11 A certain gaseous hydrocarbon is bubbled through bromine. The bromine is decolourised. What can you conclude about the unknown hydrocarbon?

7.5 Evaluating the properties, structure and bonding of different substances

The properties of a substance can be used to identify the type of bonding present. Table 7.5.1 summarises the main properties used to distinguish between metallic, ionic, covalent molecular and covalent network substances.

TABLE 7.5.1 Properties of substances

PROPERTY	METALLIC	IONIC	COVALENT MOLECULAR	COVALENT NETWORK
Melting and boiling points	Variable but most commonly high	High	Low	High
Electrical conductivity	Good conductor in solid and liquid states	Non-conductor in solid state but good conductor in molten and aqueous state	Does not conduct	Does not conduct (except graphite)
Hardness and malleability	Hard and malleable	Hard and brittle	Mostly soft	Hard and brittle
Solubility	Generally insoluble	Generally soluble in water	Variable – depends on solvent	Insoluble

»

PRACTICAL ACTIVITY 7.5.1

Comparing different types of substances

Solids can be divided into four categories on the basis of their properties. Solids can be metallic, ionic, covalent molecular and covalent network. The property of electrical conductivity effectively distinguishes between metallic, ionic and covalent solids. To distinguish between covalent molecular and covalent network types, melting point must be considered.

In this experiment, you will classify substances on the basis of their ability to conduct electricity in the solid, molten and/or aqueous states.

Note: Because of the amount of equipment and time involved, this experiment is best done by rotating through stations, each of which tests the properties of one substance.

AIM

To classify substances according to their physical properties

MATERIALS

ALL STATIONS

- Power pack, kit for measuring conductivity (Stations 1–5 need 2 kits)

STATIONS 1 AND 2

- 2 crucibles with lids, Bunsen burner, tripod, pipe clay triangle
- Station 1: sodium hydroxide pellets (enough to half-fill 2 crucibles)
- Station 2: silver nitrate crystals (enough to half-fill 2 crucibles). Platinum electrodes work better at this station because of contamination of the electrodes in silver nitrate.

STATIONS 3–5

- 2 crucibles with lids, hot plate
- Station 3: candle wax (enough to half-fill 2 crucibles)
- Station 4: sulfur (enough to half-fill 2 crucibles). Place this station in a fume cupboard.
- Station 5: camphor or naphthalene (enough to half-fill 2 crucibles). Place this station in a fume cupboard.

STATIONS 6–12

- 100 mL beaker
- Station 6: 50 mL distilled water
- Station 7: 50 mL kerosene
- Station 8: 50 mL ethanol
- Station 9: 50 mL of 0.1 mol L^{-1} sucrose solution
- Station 10: 50 mL of 0.1 mol L^{-1} NaCl solution
- Station 11: 50 mL of 0.1 mol L^{-1} NaOH solution
- Station 12: 50 mL of 0.1 mol L^{-1} HCl solution

STATION 13

- 1 piece each of copper sheet, tin foil, aluminium foil and quartz

»

»

WHAT ARE THE RISKS IN DOING THIS EXPERIMENT?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?
Hydrochloric acid and sodium hydroxide are corrosive.	Wear safety glasses, gloves and protective clothing. Take care when pouring. Clean up spills immediately; if spiled on skin, wash affected area with plenty of water and notify your teacher immediately.
Silver nitrate is toxic.	Wear appropriate safety gear. Dispose of the chemical waste in the jar provided.
Substances may splash out of heated crucible.	Wear appropriate safety gear; use tongs.
Sulfur and naphthalene are irritants to eyes, nose and throat.	Work in a fume cupboard. Do not breathe in the fumes.
Kerosene is flammable and can be an irritant.	Place kerosene well away from Bunsen burners and hotplates. Work in a well-ventilated area; do not breathe the fumes.

RISK ASSESSMENT

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, as well as ways to manage them.

PROCEDURE

1. At all stations, start the power pack at 2V. If you get a reading, record it. If not, turn the voltage up one step at a time until you get a reading, or get to a maximum of 6V. Record the result, either a current reading or no current.
2. At stations 1 and 2, place one of the crucibles on the tripod and heat it with a blue flame until the contents melt. Turn the Bunsen burner to a very low flame or off. Use one of the conductivity kits to test the molten substance. Use the other kit to test the solid in the other crucible.
3. At stations 3–5, there is a chance of the substance catching fire. If it does, place the lid on the crucible to put it out. Use one of the conductivity kits to test the molten substance. Use the other kit to test the solid in the other crucible.
4. At stations 6–12, dip the conductivity kit into the solution and record the result.
5. At station 13, place the electrodes against each sample.

RESULTS

Record your results in a table similar to the following.

CHEMICAL	STATE	CONDUCTIVITY	CLASSIFICATION

ANALYSIS OF RESULTS

What, pattern if any, is there to the conductivity results?

DISCUSSION

Use your knowledge of theory to identify the types of bonding in each of the substances tested. Compare your results with the theoretical classifications and suggest any reasons for differences.

CONCLUSION

What generalisations can you make about particles in substances and conductivity?

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 Construct a table comparing metallic, ionic, covalent molecular and covalent network substances, using the properties of electrical conductivity, melting and boiling point, and hardness.

2 Explain the differences in bonding and properties of diamond, graphite and fullerenes.

3 Describe the similarities and differences between alkanes, alkenes and benzene.

4 Define ´isomer´ and list one example.

CATEGORY QUESTIONS

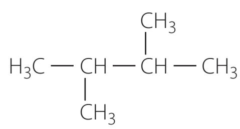
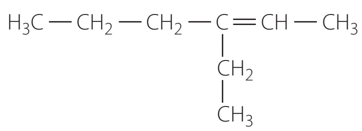
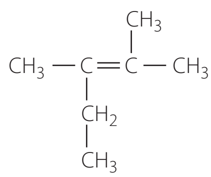
5 Phosphorus trichloride is a liquid with a boiling point of 74°C; it does not conduct electricity. Calcium chloride is a solid with a melting point of 772°C; when molten, it conducts electricity. Explain, in terms of bonding, why these compounds have such different properties.

6 Classify each of the following solids as ionic, covalent molecular, metallic or covalent network:

- **a** magnesium
- **b** tetrabromomethane
- **c** barium chloride
- **d** phosphorus triiodide
- **e** silicon dioxide
- **f** lithium sulfide
- **g** iodine
- **h** diamond.

ELABORATION QUESTIONS

7 Identify the following hydrocarbons.

- **a** $CH_3CH_2CH_2CH_3$
- **b** $CH_3—CH_2—CH_2—CH_2—CH_2—CH_3$
- **c** $CH_2 = CH—CH_2—CH_2—CH_3$
- **d** $CH_3—CH = CH—CH_3$
- **e** $CH_3—CH_2–CH = CH—CH_3$
- **f** 
- **g** 
- **h** 

8 Draw a structural formula for each of the following.

- **a** Ethane
- **b** Hexane
- **c** 3-Octene
- **d** 1-Pentene
- **e** 3,4-Dimethyl-2-pentene
- **f** 3-Methyl-1-pentene
- **g** 2-Ethyl-3-methyl-1-hexene

9 What is wrong with the following names? Give the correct names.

- **a** 5-Hexene
- **b** 4-Heptene

10 **a** Write equations for the combustion of:

- **i** Pentane
- **ii** 2-Heptene.

b Write equations for the reactions of the following alkanes in the presence of UV light.

- **i** Propane with $Br_2(aq)$
- **ii** Octane with $Cl_2(g)$

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11 Write balanced equations for the following reactions and draw the structural formulas of the products.

a 1-Pentene with Cl_2

b Propene with H_2O (with H_2SO_4 catalyst)

c 3-Hexene with bromine water

d Benzene with iodine

e Benzene with nitric acid

EVIDENCE QUESTIONS

12 Five solids A–E have the properties listed below. The relevant properties of sodium chloride and copper are also given.

SOLID	MELTING POINT (°C)	RELATIVE CONDUCTIVITY OF		SOLUBLE IN		'HAMMER' TEST
		SOLID	LIQUID	WATER?	HEXANE?	
A	327	5	2	No	No	Flattens
B	2030	0	0	No	No	Shatters
C	91	0	0	No	Yes	Forms powder
D	734	0	0.2	Yes	No	Forms powder
E	2870	0	0	No	No	Shatters
NaCl	801	0	0.2	Yes	No	Forms powder
Cu	1083	60	4	No	No	Flattens

Note: The hammer test describes what occurs when the material is continually hit with a hammer.

a Classify each of the solids A–E as ionic, covalent molecular, covalent network or metallic.

b Explain why sodium chloride and copper have the conductivity properties listed in the table.

c For either covalent molecular compounds or covalent network compounds, explain why they have the melting points, conductivities and solubilities shown in the table.

13 Four bottles containing colourless and transparent liquids were on the laboratory shelf but the labels had fallen off. The missing labels were found on the floor and had the names: hexene (C_6H_{12}), pentane (C_5H_{12}), water (H_2O) and methylated spirits.

A group of chemistry students was asked to devise a series of tests to identify the contents of each of the four bottles. The students labelled the bottles A–D and conducted their tests. Their results are shown in the following table.

BOTTLE	SOLUBLE IN WATER	FLAMMABLE	DECOLOURISES BROMINE IN THE ABSENCE OF UV LIGHT
A	Yes	No	No
B	No	Yes	Yes
C	No	Yes	No
D	Yes	Yes	No

Using the information from the tests, identify which bottle contained which chemical, giving reasons for your decision.

END-OF-CHAPTER EXAM

End-of-chapter test

1 Solid sodium fluoride (NaF) melts at 992°C, while the melting point of magnesium oxide (MgO) is 2800°C.

Which of the following is not a reasonable statement concerning the compounds MgO and NaF?

A There is ionic bonding between the particles in the lattice of each solid.

B There are stronger forces holding the particles together in MgO than in NaF.

C Electrons were transferred between atoms before lattices of the solids were formed.

D The electron configuration of the ions accounts for the different values of melting temperature of the two solids.

2 At room temperature, carbon disulfide is likely to:

A be a liquid that has high electrical conductivity.

B have ionic bonds and be hard but brittle.

C consist of triatomic molecules with weak bonds between the molecules.

D be a covalent network solid with a high melting temperature.

3 Which of the following substances conduct electricity as solids?

A Cu and Al

B Cu, Al and Hg

C Cu, $CuCl_2$, Al and Hg

D Cu, $CuCl_2$, H_2O, Al, Hg and CH_4

4 Which of the following substances conduct electricity as liquids?

A Cu and Al

B Cu, Al and Hg

C Cu, $CuCl_2$, Al and Hg

D Cu, $CuCl_2$, H_2O, Al, Hg and CH_4

5 How many hydrogen atoms are there in 2-pentene?

A 5

B 9

C 10

D 12

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6 The substances H_2O, $AlCl_3$, SO_2, NH_3, Cu, NaBr, Li_2O, Cl_2 that are covalent molecular substances are:

A $AlCl_3$, NaBr and Li_2O.

B H_2O, SO_2, NH_3 and $AlCl_3$.

C H_2O, SO_2, NH_3 and Cl_2.

D H_2O, $AlCl_3$, SO_2, NH_3, Cu, NaBr, Li_2O and Cl_2.

7 A sample of substance Z was analysed and found to have a low density and a low melting point. The substance was found to be insoluble in water and did not conduct electricity in either the liquid or solid form. The type of bonding in substance Z is most likely:

A metallic.

B covalent molecular.

C covalent network.

D ionic.

8 Provide examples to distinguish between the following terms.

a Allotrope and isotope

b Covalent molecular and covalent network structure

9 Compound A is a liquid at room temperature. It does not conduct electricity. When heated, it turns to a gas at 78°C. It can be easily condensed back to the original liquid.

a What evidence do you have that the liquid is not mercury metal?

b Give a reason why the liquid cannot be an ionic material.

c Which forces are stronger in this compound: the forces between molecules or the forces within the molecule? Explain your answer.

10 **a** Butane and diamond both contain carbon, and both contain covalent bonds between neighbouring carbon atoms. The melting point of butane is −140°C, while diamond is still a solid at over 3000°C. Explain why the melting points are so different.

b Alkanes and alkenes are both groups of hydrocarbons. Explain the main difference between these groups of compounds.

11 Give the semi-structural formulas and systematic names for all of the isomers of the fifth member of the alkane homologous series.

» UNIT ONE

CHEMICAL FUNDAMENTALS: STRUCTURE, PROPERTIES AND REACTIONS

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Topic 3: Chemical reactions: reactants, products and energy change

SCIENCE AS HUMAN ENDEAVOUR

Students should be given opportunities to investigate: the use of green energy to minimise the use of energy by industry, exothermic reactions in the body to provide energy, the use of and economic, social and environmental impact of biofuels.

KEY FORMULAS

$\Delta H = H_{(\text{products})} - H_{(\text{reactants})}$

ΔH = energy required to break bonds – energy released when bonds are formed.

$q = mc\Delta T$

$$\text{Percentage uncertainty (\%)} = \frac{\text{absolute uncertainty}}{\text{measurement}} \times 100$$

$$\text{Percentage error (\%)} = \left(\frac{\text{measured value} - \text{true value}}{\text{true value}}\right) \times 100$$

$$\text{percentage yield (\%)} = \frac{\text{experimental yield}}{\text{theoretical yield}} \times 100$$

8 CHEMICAL REACTIONS

Introduction

Chemical reactions involve the breaking of bonds in reactants as well as interactions between atoms and molecules to form new atomic and molecular substances as products. In this topic you will learn about how to describe chemical reactions using balanced chemical equations. A chemical reaction can involve either the production of energy (heat), or require heat for it to occur. You will learn how to measure and quantify the heat changes involved in chemical reactions and link those measurements and calculations with the internal (bond) energies of molecules.

Substances can also change from solid to liquid, liquid to gas, or vice versa. These are not chemical reactions as such, but are called phase changes. Phase changes can also involve temperature changes, such as when ice melts, it absorbs energy. The phase is whether the substance is in solid, liquid or gaseous form.

Stimulus questions

The human body is said to be a 'chemical factory'. Outline some of the most important chemical reactions that are associated with living matter.

Photosynthesis is responsible for the conversion of the largest mass of matter from one form to another on Earth. What chemical processes are involved?

Is the corrosion of iron a physical or chemical process? How do you know?

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8.1 Observing energy changes

chemical reaction the creation of new substances and associated energy transformations

Chemical reactions are important in our lives, in our homes and in industry. Chemical reactions occur every time we light a gas stove or a barbecue. Rusting of the steel in a car body is an example of a chemical reaction. Cooking food causes the ingredients to change chemically and form new substances. In industry, new materials are always being developed, and they require chemical reactions to change the starting materials into new materials. WikiCell is an edible form of packaging developed at Harvard University. Based on gelatin and sugarcane fibre, WikiCell may substantially decrease our need for plastic food packaging.

Some materials, such as liquefied natural gas (LNG) (which is mostly methane), require very little chemical processing to make them useful. Yet most materials are produced by chemical reactions, such as reactions, such as paints, fuels, textiles, hard finishes for cupboards and benchtops, paper and cardboard, electronic devices, components of household appliances, cleaning agents, and pharmaceuticals.

123RF / achiartistul

FIGURE 8.1.1 A car engine – the site of fuel combustion, a chemical reaction that releases energy by the conversion of a carbon–hydrogen compound, together with oxygen, to carbon dioxide and water vapour

Photosynthesis, respiration, rusting and combustion all involve chemical reactions. Photosynthesis occurs in plants in the presence of light. Respiration occurs in both plants and animals all the time. Iron rusts when it is exposed to oxygen and moisture. Petrol burns in the combustion chamber of cars. In each case, chemical changes have occurred.

All chemical reactions involve the creation of new substances and associated energy transformations, which scientists commonly observe as changes in the temperature of the surroundings, the emission of light, or both.

The main signs that can indicate that a chemical process has occurred are:

- solid precipitates from a solution or a substance is consumed
- a gas or vapour is produced (evolved)
- a colour change takes place
- there is a change in smell
- there is an input or release of energy, indicated by temperature changes or light being given off
- it is difficult to reverse what has happened.

123RF / firina

FIGURE 8.1.2 Green leaves and stems – the site of photosynthesis, a reaction that builds molecules that store energy. Photosynthesis is more important to life on Earth than any other chemical process. It is the largest chemical 'factory' on the planet!

8.1.1 Energy changes

Phase change

phase change transformation of a substance between solid, liquid or gaseous forms

Phase change is also referred to as 'change of state'. Melting, freezing, vaporisation (evaporation and boiling) condensation, sublimation and deposition are all changes of state, involving temperature changes (exchanges of heat energy) but not involving chemical reactions.

Familiar examples of a phase change include ice freezing and melting (liquid to solid and vice versa), water boiling (liquid to vapour), steam condensing and forming water droplets (vapour to liquid). On hot days, rainwater lying on a road surface evaporates as the rain water undergoes a phase change from liquid to vapour. If you leave moth balls (made from the organic chemical naphthalene) lying in a cupboard, they slowly disappear because they change state directly from solid to vapour, which is a process called 'sublimation'. Leaving a little salty water in a glass for a few days results in salt crystals depositing on the glass surface after the water has evaporated. The salt (NaCl) has undergone 'deposition' – a change from being in solution (dissolved in water) to being deposited as a solid.

Chemical equations

A chemical reaction can be represented by a chemical equation, which can use words or formulas. In a chemical equation, the reactants (the substances that exist before a chemical change (or reaction) takes place) are written to the left of the arrow. The products (the new substances formed) are written to the right of the arrow. The single arrow always points in the direction of the reaction; that is, from the reactants to the products.

For example, the word equation for photosynthesis is:

Carbon dioxide + water → glucose + oxygen

Carbon dioxide and water are the reactants. Glucose and oxygen are the products. The arrow points towards the products. The '+' is used when there is more than one reactant or more than one product. Energy changes associated with the chemical reaction can also be included in the writing of the chemical equation. This is discussed in Chapter 9.

SYLLABUS LINK

Chapter 9 discusses energy changes associated with the chemical reaction.

WORKED EXAMPLE 8.1.1

Respiration is the combustion reaction of glucose and oxygen in body cells to form carbon dioxide and water. The word equation is:

Glucose + oxygen → carbon dioxide + water

1 Identify the reactants.

2 Identify the products.

ANSWERS

1 Glucose and oxygen are the reactants, because they are to the left of the arrow.

2 Carbon dioxide and water are the products, because they are to the right the arrow.

SECTION REVIEW 8.1

REMEMBERING

1 Define:

a chemical reaction   b phase change.

UNDERSTANDING

2 Write word equations and identify the reactants and products for the following reactions.

a Rust (iron(III) oxide) forms when iron reacts with oxygen in the air.

b Petrol is composed of many hydrocarbon compounds, including pentane. Pentane reacts with oxygen in the combustion chamber of cars to form carbon dioxide and water.

8.2 Balanced chemical equations

Chemists normally represent chemical reactions using chemical symbols and formulas for the reactants and products. Numbers are written in front of chemical formulas to balance the equation so that the same number of atoms of each element is shown on each side. The number '1' is not written because it is assumed when no other number appears before the formula.

Generally, equations are written as the lowest whole number ratio between reactants and products. This ensures that the law of conservation of mass is maintained.

The states for each of the reactants and products are often included:

- (s) for solid
- (l) for liquid
- (g) for gas
- (aq) for aqueous solution.

In some sources, (aq) is not included; if a reactant or product does not have a state included, it is assumed that it is an aqueous solution (dissolved in water).

Chemists use balanced chemical equations to summarise information about chemical reactions. These equations communicate information concisely about reactants, products and numbers of atoms of each element present.

For example, as seen in Figure 8.2.1, magnesium (Mg) reacts with hydrochloric acid (HCl) to form magnesium chloride ($MgCl_2$), releasing bubbles of gas. The gas can be identified as hydrogen (H_2) using a flame to cause a 'pop'. The reaction can be described by the equation:

$$Mg + HCl \rightarrow MgCl_2 + H_2$$

The equation then needs to be balanced by altering the coefficients. Equations must be balanced to achieve the same number of atoms of each element on each side of the equation. In this example, the amount of hydrochloric acid is doubled:

$$Mg + 2HCl \rightarrow MgCl_2 + H_2$$

- There is one 'atom' of magnesium at the beginning of the reaction (left-hand side of the equation) and one 'atom' of magnesium at the end (right-hand side of the equation).
- There are two 'atoms' of hydrogen at the beginning of the reaction and two 'atoms' of hydrogen at the end.
- There are two 'atoms' of chlorine at the beginning of the reaction and two 'atoms' of chlorine at the end.

The states for each of the reactants and products are now included to complete the equation for the reaction:

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

Magnesium is a solid (s); hydrogen is a gas (g). Both hydrochloric acid and magnesium chloride are dissolved in water (aq).

The method for balancing (by adjusting the coefficients) is the same for complex reactions.

SYLLABUS LINK

Chapter 12 discusses the law of conservation of mass.

Science Photo Library/Martyn F. Chillmaid

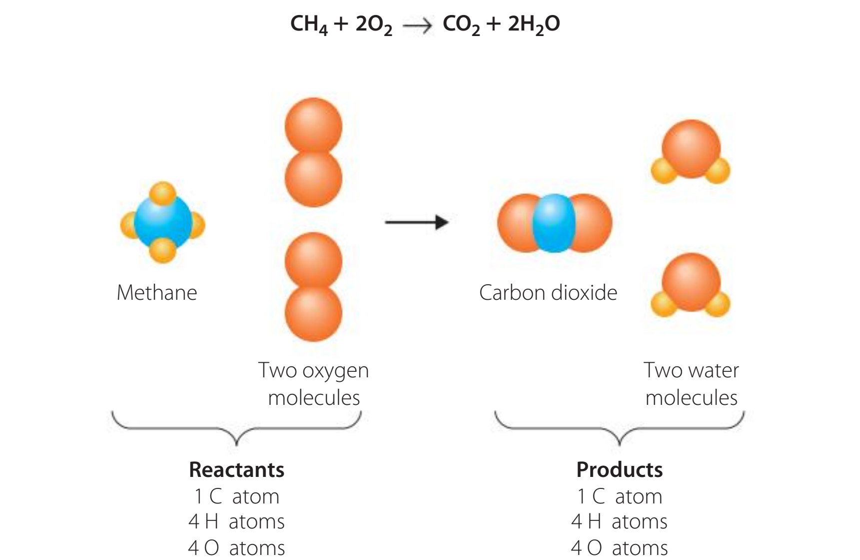
FIGURE 8.2.1 Magnesium reacting with hydrochloric acid. The bubbles seen are the release of hydrogen gas.

The combustion of methane

Methane, the main constituent of LNG, has the molecular formula CH_4. The combustion of methane in air involves the reaction with oxygen (O_2) to produce carbon dioxide (CO_2) and water vapour (H_2O). The balanced reaction is shown in Figure 8.2.2 (page 138).

In Figure 8.2.2, the reaction between methane and oxygen is shown as a chemical equation, but also as pictures showing the constituent atoms and shapes of the molecules. The reactants are on the left side of the equation, and the products are on the right. The equation is balanced – the numbers of C, H and O atoms are the same on both sides of the equation.

FIGURE 8.2.2
The reaction between methane and oxygen to form carbon dioxide and water



WORKED EXAMPLE 8.2.1

Write a balanced chemical equation, including states, for the reaction of octane (C_8H_{18}) with oxygen in a car's combustion chamber to produce carbon dioxide and water. Note that the combustion chamber is at a high temperature so, in this case, the water formed will be gaseous until it condenses later on after cooling through the exhaust pipe.

ANSWER

1 Write reactants and products using their chemical formulas.

$$C_8H_{18} + O_2 \rightarrow CO_2 + H_2O$$

2 Balance the equation.

$$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$$

a Start with the elements that appear in only one substance (choose C_8H_{18}) and balance atom numbers on each side of the equation for that substance and the atoms it contains. That is, balance only the C and H atoms to start because C_8H_{18} contains only C and H ($8 \times$ C and $18 \times$ H), so $8 \times CO_2$ and $9 \times H_2O$ are needed on the right to balance the atom numbers:

$$C_8H_{18} + O_2 \rightarrow 8CO_2 + 9H_2O$$

b Consider the O atom numbers. There are now $(8 \times 2) + (9 \times 1) = 25$ atoms of oxygen on the right and two atoms of oxygen on the left. O_2 is a molecule with 2 O atoms, so the fraction $\frac{25}{2}$ must go in front of the O_2, on the left, to balance with the 25 O atoms on the right. Note that, when balancing equations, whole numbers are used, so all quantities in the equation are doubled in this example, so 50 O atoms end up on both left and right-hand sides, as needed.

$$2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$$

Each side of the equation now has the same number of 'atoms' for each element.

3 Add the states for each of the reactants and products.

$$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$$

Octane is a liquid; oxygen and carbon dioxide are gases. This is a combustion reaction, so the water will be created at a high temperature and therefore it will be gaseous.

SECTION REVIEW 8.2

REMEMBERING

1 Identify the purpose of a chemical equation.

2 Identify the steps required to write balanced chemical equations.

UNDERSTANDING

3 Balance the following chemical equations.

a $Fe(s) + HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$

b $NaHCO_3(s) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + CO_2(g) + H_2O(l)$

c $CuFeS_2(s) + SiO_2(s) + O_2(g) \rightarrow Cu_2S\ (s) + FeSiO_3(s) + SO_2(g)$

d $PbS(s) + O_2(g) \rightarrow PbO(s) + SO_2(g)$

e $C_3H_8(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

APPLYING

4 Identify whether the following equations are balanced. Correctly balance the equations that are not balanced and justify your answer.

a $H_2(g) + O_2(g) \rightarrow H_2O(l)$

b $H_2SO_4(aq) + Ba(OH)_2(aq) \rightarrow BaSO_4(s) + 2H_2O(l)$

c $C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(l) + CO_2(g)$

d $2C_2H_5OH(l) + 6O_2(g) \rightarrow 4CO_2(g) + 6H_2O(g)$

e $TiO_2(s) + 2Cl_2(g) + 2C(s) \rightarrow TiCl_4(aq) + 2CO(g)$

8.3 Reactivity

Chemicals react to achieve stability; that is, the products are more stable than the reactants. During chemical reactions, bonds are broken and new bonds are formed. Energy is needed to break bonds, to break down chemically the forces of attraction that occur between:

- metal ions and the delocalised electrons in metallic lattices
- cations and anions in ionic lattices
- non-metallic elements in covalent molecules and covalent lattices.

Energy is released when bonds are formed.

spontaneous chemical reaction a reaction that occurs without the input of additional energy such as heat, light or electricity

Spontaneous chemical reactions

Spontaneous chemical reactions are those that occur without any addition of energy from the surroundings. Spontaneous reactions that take place at room temperature include the ripening of fruit, salt dissolving in water and the rusting of iron. They do not require any added energy for the reaction to occur. Substances that are more reactive are more likely to undergo reactions spontaneously at room temperature. Some spontaneous reactions occur quickly and others occur more slowly, but all occur without the input of additional energy, as shown in practical activities 8.3.1 and 8.3.2.

Metals can react spontaneously. All metals have some common properties; for example, except for mercury, they all have relatively high melting points. However, they do not have identical melting points. Similarly, they may have similar chemical properties, but they are not identical.

INQUIRING FURTHER

Alkaline batteries involve the reaction between zinc metal and manganese dioxide with potassium hydroxide (the alkaline compound) acting as a conductor. There are many other types of batteries that use spontaneous chemical reactions to generate electrical power. Research one of these (such as the lithium battery used in electrical cars) and explain the chemical reactions taking place.

Some metals react spontaneously with water, acid and oxygen. The generalised equations for these reactions are as follows.

$$\text{Metal} + \text{water} \rightarrow \text{a salt} + \text{hydrogen}$$

$$\text{Metal} + \text{acid} \rightarrow \text{a salt} + \text{hydrogen}$$

$$\text{Metal} + \text{oxygen} \rightarrow \text{metal oxide}$$

Science Photo Library

FIGURE 8.3.1 Batteries spontaneously change chemical energy to electrical energy

123RF/Jose Angel Astor

FIGURE 8.3.2 Metals in cars react with oxygen in the air to corrode

PRACTICAL ACTIVITY 8.3.1

Reaction of metal with oxygen

Corrosion is the spontaneous reaction of a metal with oxygen in a moist environment.

AIM

To compare the reactivity of metals by comparing their reactions in a moist environment

HYPOTHESIS

Write a hypothesis for this experiment.

EQUIPMENT PER GROUP

- 5 cm × 0.5 cm piece of copper
- 5 cm × 0.5 cm piece of zinc
- 5 cm × 0.5 cm piece of aluminium foil
- 5 cm piece of magnesium
- 5 cm iron nail (ungalvanised)
- 5 medium test tubes
- test-tube rack
- 5 cm × 5 cm piece of sand paper

»

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	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?
Metal residue may remain on your hands.	Wash your hands after handling metals.

RISK ASSESSMENT

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, as well as ways to manage them.

PROCEDURE

1 Clean each of the metals with the sand paper.
2 Place a piece of each metal into separate test tubes.
3 Place the test tubes into the test tube rack.
4 Add water to each of the test tubes until the piece of metal is covered.
5 Record observations daily for 1 week.

RESULTS

Record your observations in an appropriate format. For example, you could take a series of digital photos of each test tube.

ANALYSIS OF RESULTS

Write a balanced chemical equation for any reaction that has occurred.

QUESTIONS

1 Which metal was the most reactive? Justify your answer.
2 Which metal was the least reactive? Justify your answer.
3 Do your results support your hypothesis? Justify your answer.
4 Rank the metals from most reactive to least reactive. Justify your answer.

PRACTICAL ACTIVITY 8.3.2

Reactivity of metals

AIM

To compare the reactivity of metals by observing their reactions with water and acids

HYPOTHESIS

Write a hypothesis for this experiment.

EQUIPMENT PER GROUP

- 3 × 5 cm × 0.5 cm pieces of copper
- 3 approx. 5 cm long ungalvanised iron nails
- 3 × 5 cm pieces of magnesium
- 3 × 5 cm × 0.5 cm pieces of zinc
- bottle of calcium metal
- 50 mL × 1 mol L^{-1} hydrochloric acid (HCl)
- 50 mL × 4 mol L^{-1} HCl
- 5 × test tubes
- spatula

» RISK ASSESSMENT

WHAT ARE THE RISKS IN DOING THIS EXPERIMENT?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?
Acid could splash or spill onto your hands.	Wash your hands immediately if they come in contact with acid and after using acids.
Vigorous bubbling may splash into your eyes.	Wear safety glasses.

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, as well as ways to manage them.

PROCEDURE

PART A: REACTING METALS WITH WATER

1. Place one-quarter of a spatula of calcium into a test tube.
2. Place a piece of each of the other metals into separate test tubes.
3. Add approximately 3 cm depth of water to each test tube.
4. Record your observations.

PART B: REACTING METALS WITH 1 $mol\,L^{-1}$ HYDROCHLORIC ACID

1. For any metals that did not violently react with water, place a small quantity or piece of each into separate test tubes.
2. Add approximately 3 cm depth of 1 $mol\,L^{-1}$ HCl to each test tube.
3. Record your observations.

PART C: REACTING METALS WITH 4 $mol\,L^{-1}$ HYDROCHLORIC ACID

1. For any metals that did not violently react with 1 $mol\,L^{-1}$ hydrochloric acid, place a piece of each into separate test tubes.
2. Add about 3 cm depth of 4 $mol\,L^{-1}$ HCl to each test tube.
3. Record your observations.

RESULTS

Draw a table to show your results.

ANALYSIS OF RESULTS

Write a balanced equation for any chemical reaction that occurred in parts A, B and C.

QUESTIONS

1. Which metal was the most reactive? Justify your answer.
2. Which metal was the least reactive? Justify your answer.
3. Do your results support your hypothesis? Justify your answer.
4. Rank the metals from most reactive to least reactive. Justify your answer.
5. Why were the metals that reacted with water not tested with the HCl?

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INQUIRING FURTHER

CORROSION

Iron (as steel) is the main metal used in cars, and it corrodes readily unless it is protected. Review the chemical reaction that is responsible for the corrosion. How can you reduce the corrosion of steel?

Aluminium components do not corrode as readily as steel; explain why this is the case.

Plastics and rubber in cars also degrade slowly with time. Conduct research into and explain the spontaneous chemical reactions that cause the degradation of these materials.

SECTION REVIEW 8.3

REMEMBERING

1 Define:

a spontaneous chemical reaction

b corrosion.

UNDERSTANDING

2 Classify reactivity as either a chemical property or a physical property and justify your choice.

8.4 Representing chemical reactions

Chemical reactions can be represented in various ways, depending on the detail that is required. So far, balanced chemical equations have been used. These give general information about the reactants and the products. However, when the compounds involved in the reactions are bonded as ions, ionic equations are applied, to give information about each species present in the reaction. When ionic compounds are dissolved in water, they can be broken up (dissociated) into their constituent ions. The ionic species that undergo chemical change are the ones that take part in the reaction. The species that remain the same on both sides of the ionic equation are called **spectator ions**. They are present but do not take part in the chemical reaction. **Net ionic equations** are more specific again; they include only information about the species that take part in the reaction. Net ionic equations do not show the spectator ions but must be balanced for charge, as shown in Worked example 8.4.1 on page 144.

spectator ion
ions that remain in solution before and after a chemical reaction has taken place, which are not involved in the reaction

net ionic equation
shows only the species that are taking part in the chemical reaction; spectator ions are not included

WORKED EXAMPLE 8.4.1

Tin reacts with hydrochloric acid to form tin(II) chloride and hydrogen gas:

$$\text{Tin} + \text{hydrochloric acid} \rightarrow \text{tin(II) chloride} + \text{hydrogen}$$

Write the net ionic equation for the reaction.

ANSWER

1 The balanced equation is:

$$Sn(s) + 2HCl(aq) \rightarrow SnCl_2(aq) + H_2(g)$$

2 The ionic equation is:

$$Sn(s) + 2H^+(aq) + 2Cl^-(aq) \rightarrow Sn^{2+}(aq) + 2Cl^-(aq) + H_2(g)$$

Both the tin and hydrogen undergo chemical changes during the reaction. The chloride ion remains the same on both sides of the equation and does not take part in the chemical reaction. The chloride ion is a spectator ion, so it is not included in the net ionic equation.

This means that the net ionic equation is:

$$Sn(s) + 2H^+(aq) \rightarrow Sn^{2+}(aq) + H_2(g)$$

SECTION REVIEW 8.4

REMEMBERING

1 Identify two different ways that chemists represent chemical reactions.

APPLYING

2 Write net ionic equations for each of the following reactions:

- **a** $Fe(s) + CuSO_4(aq) \rightarrow Cu(s) + FeSO_4(aq)$
- **b** $Na_2CO_3(aq) + MgCl_2(aq) \rightarrow MgCO_3(s) + 2NaCl(aq)$

3 Write net ionic equations for the reaction of:

- **a** calcium metal and water
- **b** magnesium and hydrochloric acid
- **c** zinc and sulfuric acid
- **d** lead(II) nitrate solution and potassium iodide solution to form potassium nitrate solution and solid lead(II) iodide.

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CHAPTER REVIEW QUESTIONS

ELABORATION QUESTIONS

1 Explain how stalactites and stalagmites form in limestone caves and write the appropriate ionic equations.

EVIDENCE QUESTIONS

2 Write balanced chemical equations for:

a combustion of cyclohexane (C_6H_{12}) to yield carbon dioxide and water vapour

b rusting of iron when it reacts with oxygen to form iron(III) oxide

c respiration of glucose ($C_6H_{12}O_6$) with oxygen to produce carbon dioxide and water.

3 Write balanced chemical equations to represent each of these chemical reactions.

a The Haber process is an important chemical reaction where reacting hydrogen gas with nitrogen gas produces gaseous ammonia (NH_3).

b Magnesium hydroxide is an ingredient in common antacids. It reacts with the dilute hydrochloric acid in the stomach to produce a solution of magnesium chloride and water.

c Zinc metal reacts with dilute sulfuric acid to form dilute zinc sulfate and hydrogen gas.

4 Write net ionic equations for each of the following reactions.

a $2Al(s) + 3H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3H_2(g)$

b Copper metal reacts with nitric acid to form copper(II) nitrate, nitrogen dioxide gas and water.

END-OF-CHAPTER EXAM

End-of-chapter test

1 Which of the following is not an example of a chemical reaction?

A An iron nail rusting

B Lighting a match

C Respiration

D Sugar crystals dissolving in water

2 A student performed the 'pop' test by holding a lit taper over the top of a closed test tube containing hydrogen gas. When the gas was released, a loud 'pop' was heard and small clear, colourless droplets of liquid were formed at the mouth of the test tube. These were found to be water. What is the correct word equation to represent this reaction?

A Hydrogen + oxygen → water

B Hydrogen + heat → water + sound

C Hydrogen → water + sound

D Hydrogen → water

3 What does (aq) mean when written in a chemical equation?

A Water is a reactant in the chemical reaction.

B Water is a product in the chemical reaction.

C All of the substances are dissolved in water.

D The substance directly before it is in solution.

For Questions **4** and **5** refer to the unbalanced chemical equation below:

$$Mg + HCl \rightarrow MgCl_2 + H_2$$

4 Which of the following shows the correct coefficients for the balanced chemical equation?

A 1:1:1:1

B 1:2:1:2

C 1:2:1:1

D 2:4:2:2

5 Which of the following statements is false?

A The coefficients represent the particle ratios of each substance.

B The coefficients represent the mass ratios of each substance (e.g. g, kg).

C The coefficients represent the mole ratios of each substance.

D The coefficients represent the volume ratios of each substance.

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6 The reaction between sodium and water can be written as:

Sodium + water → sodium hydroxide + hydrogen

Which of the following equations is the correct balanced chemical equation for the reaction?

A $2Na(s) + 2H_3O \rightarrow 2NaOH(aq) + 2H_2(g)$

B $2Na(s) + H_2O \rightarrow 2NaOH(aq) + H_2(g)$

C $Na(s) + 2H_2O \rightarrow NaOH(aq) + 2H_2(g)$

D $2Na(s) + 2H_2O \rightarrow 2NaOH(aq) + H_2(g)$

7 Iron(III) oxide reacts with aluminium to produce aluminium oxide and metallic iron. Which of the following is the correct balanced chemical equation for the following reaction?

A $Fe_2O_3 + Al \rightarrow Fe + Al_2O_3$

B $Fe_2O_3 + 2Al \rightarrow 2Fe + Al_2O_3$

C $Fe_2O_3 + 3Al \rightarrow 2Fe + 3AlO$

D $Fe_2O_3 + Al_2 \rightarrow Fe_2 + Al_2O_3$

For Questions **8** and **9** refer to the table below.

METAL	REACTS WITH WATER	REACTS WITH DILUTE ACID	REACTS WITH OXYGEN
Ag			✓
Ca	✓	✓	✓
Al		✓	✓
Pb		✓	✓
Cu			✓
Mg		✓	✓
Au			
K	✓	✓	✓

8 Which is the least reactive metal?

A Ag

B Au

C Ca

D K

9 Which metal would most likely need to be stored in oil so it is not in contact with the air?

A K

B All except Au

C Ag

D None

10 If energy does not need to be added to a reaction to make it occur at room temperature, then it can be referred to as a:

A reversible reaction.

B combustion reaction.

C spontaneous reaction.

D non-spontaneous reaction.

11 Which of the following does not include spectator ions in the equation?

A Word equation

B Balanced chemical equation

C Ionic equation

D Net ionic equation

12 Copper(II) nitrate is heated in a crucible over a Bunsen burner, forming solid copper(II) oxide, nitrogen dioxide and oxygen gas.

Write a balanced chemical equation for this reaction.

13 A sodium hydroxide solution is added to a magnesium chloride solution. Solid magnesium hydroxide is formed with another clear solution.

a Write a balanced chemical equation for this reaction.

b Write an ionic equation for this reaction.

c Write a net ionic equation for this reaction.

14 Write balanced chemical equations for each of the following reactions.

a The reaction of nitrogen with hydrogen to form ammonia (gas)

b The decomposition of iron(II) sulfate ($FeSO_4$) to yield Fe_2O_3, SO_2 and SO_3

c The combustion of propane (LPG gas) to produce carbon dioxide and water vapour

d The combustion of ethanol to produce carbon dioxide and water vapour

e Photosynthesis: the reaction of carbon dioxide and water to produce $C_6H_{12}O_6$ (glucose) and oxygen

9 EXOTHERMIC AND ENDOTHERMIC REACTIONS

Introduction

In a chemical reaction, bonds are broken in the reactants, and bonds are formed to make the products. These changes involve energy.

This topic will explore how endothermic and exothermic reactions relate to the law of conservation of energy. When bonds break or are re-formed, heat energy is released or absorbed by the system to or from the surrounding environment. Heat is a form of energy and temperature is a measure of the average kinetic energy of the particles. The relationship between temperature and enthalpy changes will be applied to identify thermochemical reactions as exothermic or endothermic enthalpy level diagrams. Thermochemical equations will be used to identify the relative stabilities of reactants and products, and the sign of the enthalpy change (ΔH) for a reaction. The reason why reactions are exothermic or endothermic will be understood in terms of average bond enthalpies. Calorimetry measurements will be carried out to estimate the heat changes for a substance given the mass, specific heat capacity and measured temperature change. Calorimetry can be applied to obtain estimates for the enthalpy change for a chemical reaction.

Stimulus questions

- What is the source of energy in chemical bonds?
- How can the energy contained in a chemical bond be understood in terms of protons and electrons?
- When chemical reactions occur, what physical changes might be experienced by the surrounding solids, liquids or gases?
- How can the energy contained in a chemical bond be measured using calorimetry?

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9.1 The law of conservation of energy

Energy is stored in the chemical bonds of the reactants and the products. The law of conservation of energy states that energy cannot be created or destroyed; it can only be changed from one form to another. So if more energy is stored in the bonds of the reactants than the products, then when the products are formed, the extra energy is released as heat, sound or light. An example of this is the burning of magnesium. A small amount of energy, the activation energy, is needed to start this reaction. Burning magnesium produces a very bright white light.

SYLLABUS LINK
Chapter 21 discusses activation energy.

Once the reaction is kick-started by placing the magnesium ribbon in the flame, overcoming the activation energy, no further energy input is required. The burning of magnesium is an **exothermic reaction** because energy is released overall, as shown in Figure 9.1.1 and the following equation:

exothermic reaction
a reaction in which energy is released to the surroundings, causing a temperature increase

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s) + \text{light}$$

Another exothermic reaction is the reaction between magnesium and hydrochloric acid. This differs from the burning of magnesium in two ways.

1 Heat is produced instead of light.
2 The reaction is spontaneous at room temperature; no energy is needed to start the reaction:

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g) + \text{heat}$$

SYLLABUS LINK
Chapter 13 discusses intermolecular forces.

Scientists write the equations for these reactions to include both the chemicals involved and an energy term (energy gained or lost, as heat or light). They are called thermochemical equations.

Physical processes, such as changes of state, can also involve the release of energy. For example, water vapour (steam) cools down when it forms liquid water – it releases energy as heat. This phase change is an exothermic process (but not a reaction). Here, the hot, fast-moving gaseous water molecules (steam) lose kinetic energy in collisions with colder air molecules (mainly nitrogen and oxygen), or a colder surface, progressively slowing down. Once the kinetic energy of the water molecules is less than the forces of attraction between them (intermolecular forces), the water molecules combine, at first just in pairs (dimers), then trimers, tetramers, pentamers . . . and eventually they form droplets. The droplets grow in size and finally coalesce to a continuous liquid. The equation representing the phase change is:

FIGURE 9.1.1 Burning magnesium releases energy in the form of light

$$H_2O(g) \rightarrow H_2O(l) + \text{heat}$$

Other chemical reactions are **endothermic**, which means that the products contain more chemical energy than the reactants. This means that energy must be continuously added to allow the chemical reaction to proceed. The decomposition of copper(II) carbonate ($CuCO_3$) is an example of this type of reaction. This reaction only occurs while the copper(II) carbonate is being heated. The products contain more chemical energy than the reactants so extra energy is put into the reaction. Heat energy is converted to chemical energy.

endothermic reaction
a reaction in which energy is absorbed from the surroundings, causing a temperature decrease

$$CuCO_3(s) + \text{heat} \rightarrow CuO(s) + CO_2(g)$$

Physical processes, such as the boiling of water to produce steam, also involve the input of energy. This makes the boiling of water an endothermic process. The water molecules are gaining kinetic energy from the heat supplied, meaning that they are colliding with the hot surface of the heated container and rebounding with increased kinetic energy. The increased kinetic energy permits water molecules with sufficient kinetic energy to overcome the intermolecular forces holding the liquid together and the molecules leave as fast-moving gaseous molecules and are dispersed into the surrounding air. The energy required to change a liquid to a gas is called the **latent heat** of vaporisation. The latent heat of vaporisation for water is 2256 kJ kg^{-1}:

$$H_2O(l) + \text{heat} \rightarrow H_2O(g)$$

The energy required to change a solid to a liquid (the process of melting) is called the latent heat of fusion.

Enthalpy is the heat absorbed in a chemical reaction at constant pressure. The change in enthalpy, also referred to as the 'heat of reaction', uses the symbol ΔH. It is measured in kJ. Enthalpy is defined as the heat absorbed during a chemical reaction, so, by definition, ΔH is positive. A negative ΔH would indicate that energy has been released during a chemical reaction.

The difference between the relative energies of reactants and products, and which is the higher or lower, allows one to distinguish between exothermic and endothermic reactions. In Figure 9.1.2, the reaction is exothermic, energy is released, so ΔH is negative. The red arrow indicates that the surroundings gain heat. In Figure 9.1.3, the reaction is endothermic, energy is absorbed, so ΔH is positive. The blue arrow indicates that the reactants absorb energy from the surroundings. The surroundings lose heat and will become cooler.

latent heat
energy required to change a specific quantity of a substance from solid to a liquid

enthalpy
the total energy possessed by a chemical substance, which changes as bonds are broken or made during a chemical reaction

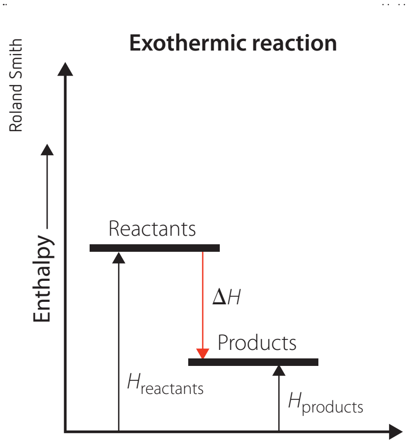

FIGURE 9.1.2 In an exothermic reaction, energy is released. The energy of the reactants is greater than the energy of the products. ΔH is negative.

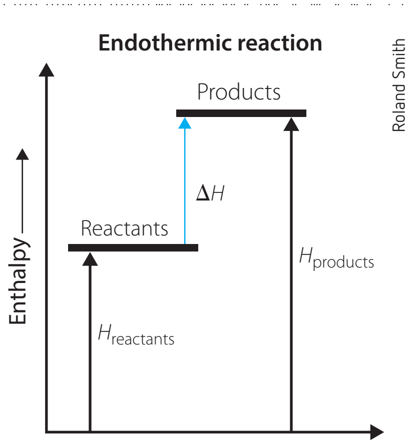

FIGURE 9.1.3 In an endothermic reaction, energy is absorbed. The energy of the reactants is less than the energy of the products. ΔH is positive.

KEY LAW

The law of conservation of energy
Energy cannot be created or destroyed; it can only be changed from one form to another. The law of conservation of energy dictates that in the breaking and re-forming of bonds, heat energy is released or absorbed by the system to, or from, the surroundings.

KEY LAW

Exothermic and endothermic reactions
Exothermic reactions have **negative** ΔH values.

Endothermic reactions have **positive** ΔH values.

KEY FORMULA

ΔH = (energy required to break bonds) − (energy released when bonds are formed)

$\Delta H = H_{\text{products}} - H_{\text{reactants}}$

Examples of how to calculate the enthalpy change for a chemical reaction, and how to determine whether a reaction is exothermic or endothermic, are provided in Worked example 9.1.1, on page 152.

9.1.1 Exothermic reactions

9.1.2 Energy changes and reversible reactions

9.1.3 Exothermic vs endothermic reactions

WORKED EXAMPLE 9.1.1

Write thermochemical equations for the following reactions.

1 Magnesium reacts with oxygen to form solid magnesium oxide. This is an exothermic reaction during which 1202 kJ are released.

2 Solid copper(II) carbonate decomposes to form solid copper(II) oxide and carbon dioxide. This is an endothermic reaction during which 44 kJ are absorbed.

ANSWERS

1 This is an exothermic reaction, so the ΔH is negative.

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$
$$\Delta H = -1202\,kJ$$

2 This is an exothermic reaction, so the ΔH value is positive.

$$CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$$
$$\Delta H = +44\,kJ$$

SECTION REVIEW 9.1

REMEMBERING

1 Define:
   a endothermic reaction
   b exothermic reaction.

2 Decide whether the following statements true or false.
   a Exothermic reactions have negative ΔH values.
   b Endothermic reactions have negative ΔH values.

APPLYING

3 Write thermochemical equations for the following reactions.
   a Liquid water is formed from hydrogen gas and oxygen gas in a reaction that releases 242 kJ heat.
   b Ammonia gas (NH_3) absorbs 92 kJ of heat to form nitrogen and hydrogen gases.
   c Calcium oxide solid is formed when 180 kJ of heat is absorbed by solid calcium carbonate. Carbon dioxide is the other product.
   d Ammonia gas reacts with oxygen to form water vapour and nitrogen oxide (NO) gas. The reaction releases 905 kJ of energy.
   e Liquid water freezes to form ice. The latent heat of fusion is 6.02 kJ mol^{-1}.

translational kinetic energy given by *KE* (translational) = $\frac{1}{2}mv^2$

rotational kinetic energy kinetic energy contained in rotational motion; determined by the mass and angular velocity

vibrational kinetic energy kinetic energy associated with the oscillation of atoms within a molecule or solid lattice as they vibrate

9.2 Heat, energy and temperature

Temperature is a common indicator of the change in energy of a substance. For example, when water is heated, it absorbs heat energy and its temperature increases. Conversely, when water is cooled, it releases heat energy and its temperature decreases.

'Heat' refers to a form of energy. Other forms of energy include light energy, electrical energy, sound energy, nuclear energy and dark energy. In terms of atoms and molecules, the forms of energy associated with temperature changes are **translational**, **rotational** and **vibrational kinetic energy** (Figure 9.2.1, page 153).

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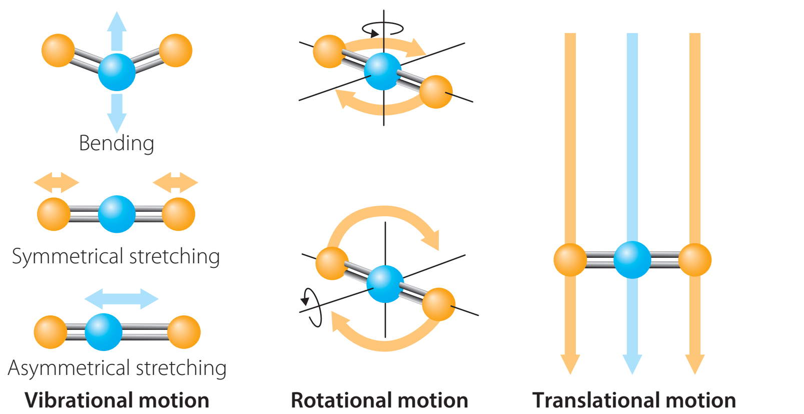

FIGURE 9.2.1 Vibrational, rotational and translational motion for the carbon dioxide molecule

Kinetic energy

Translational motion

Kinetic energy (*KE*) due to translational motion causes atoms and molecules to move through space at a variety of speeds. For gases, most of their kinetic energy is due to random translational motion. The translational kinetic energy (*KE*) is given by the equation $KE = \frac{1}{2}mv^2$, where m is the mass of the atom or molecule and v is its velocity.

In liquids, which are more 'ordered', a significant proportion of the *KE* can be due to translational motion, but *KE* due to both rotational and vibrational motion will also contribute.

kinetic energy (*KE*) the energy possessed by moving particles: a particle of mass m, moving with speed v has a translational kinetic energy equal to $\frac{1}{2}mv^2$; molecules also possess rotational and vibrational kinetic energy

Rotational energy

Rotational *KE*, molecules tumbling end over end, adds to the total *KE*. Atoms cannot have rotational *KE*, but molecules rotating have rotational energy due to the atomic masses rotating about their centre of mass (with an angular velocity). Molecular rotational energy is important in mechanisms for transfer of energy between liquids and gases.

Vibrational energy

Vibrational energy caused by atoms vibrating can be thought of as atoms joined together by springs, where the springs are the chemical bonds. In solids, which are tightly bound lattices of atoms or molecules, most of the *KE* is contained in vibrational *KE* as the atoms oscillate backwards and forwards. When a single gas molecule strikes a solid, the exchange of energy is between the solid's vibrational energy and the translational *KE* of the gas.

9.2.1 What is kinetic energy?

9.2.2 Kinetic energy

Chemical bond energy

The **chemical bond energy** or 'bond enthalpy' is the energy associated with the bond. It is the net sum of the attractive and repulsive electrostatic interactions between electrons and nuclei. Nuclear–nuclear repulsion increases as atoms come closer. In all stable molecules, this repulsion is balanced, or overcome, by electron–nuclear attraction at some interatomic distance. Chemical bond energy is measured as the difference between the energy of all the atoms totally separated in the gas phase, and the energy of the atoms when chemically combined, at the equilibrium bond distance for the stable molecule. Figure 9.2.2 (page 154) illustrates this for the ionically bonded sodium chloride molecule. The net *negative* potential energy of –589 kJ mol^{-1} at the interionic distance of 0.236 nm (236 pm) corresponds to the chemical bond energy or bond enthalpy. Bond enthalpy is always stated for the molecule and the separated atoms in the gas phase.

chemical bond energy also known as bond enthalpy; the amount of energy required to break one mole of the stated bond to give separated atoms, usually expressed in units of kJ mol^{-1}

9.2.3 Chemical bond energy

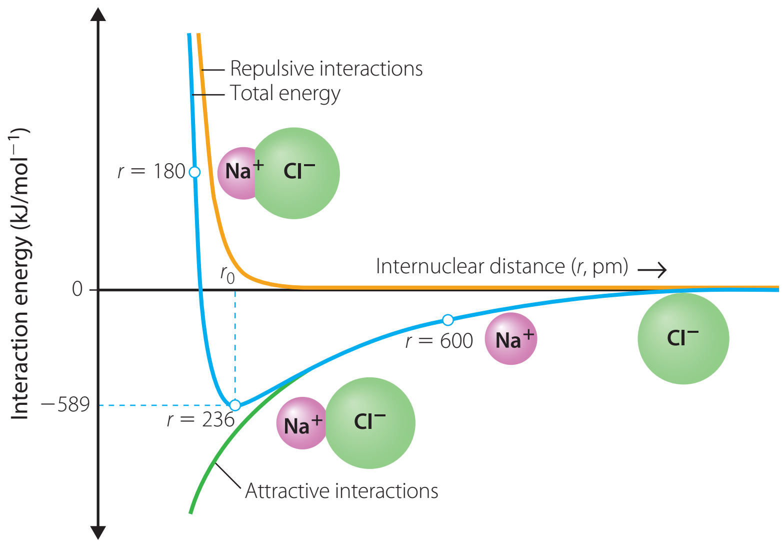

FIGURE 9.2.2 Schematic showing how the energy of interaction between a sodium cation and a chloride anion changes as a function of internuclear separation due to the changing balance between forces of attraction and repulsion

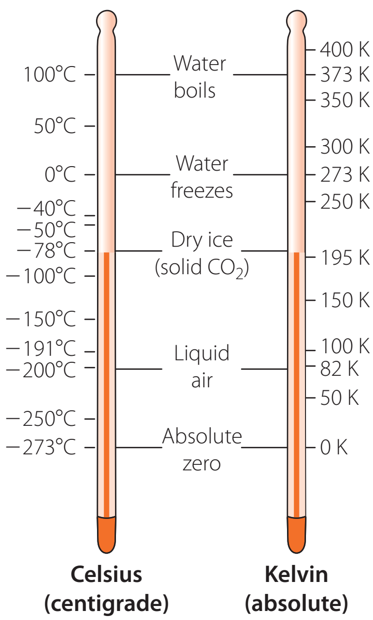

FIGURE 9.2.3 Celsius and Kelvin temperature scales

temperature
a measure of the average kinetic energy of the particles that constitute a solid, liquid or gas

Temperature

Temperature is a measure of the average *KE* of the particles in the system being measured. In gases and liquids, this is easy to conceptualise. In solids, the lattice of atoms and molecules is tightly bonded and the average *KE* refers mainly to the *KE* of the vibrational motion associated with the atoms and/or molecules in the solid. When a molecule in a solid sublimes, it converts its internal vibrational *KE* to translational *KE*. The molecule leaves the surface of the solid, travelling faster, but vibrating much less, to join the surrounding gaseous environment.

Temperature scales

Temperature can be measured on various scales. In Australia, temperature is commonly reported in degrees Celsius (°C). The Celsius temperature scale can have both positive and negative values. On a hot day, the temperatures are higher than 35°C. In some parts of Australia, temperatures can be as low as −10°C. Water freezes at 0°C and, at standard pressure, boils at 100°C (Figure 9.2.3). Standard pressure is 100 kPa or 1 bar.

Scientists often use the absolute temperature scale, which is measured in kelvin (K). This scale only has positive values. There is a linear relationship between the Celsius temperature scale and the absolute temperature scale, and the method to convert temperatures between the two scales is shown in Worked example 9.2.1 (page 155).

9.2.4 Temperature

KEY FORMULA

Temperature (°C) = temperature (K) − 273

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WORKED EXAMPLE 9.2.1

What is the:

1 freezing point of pure water in kelvin?

2 boiling point of pure water in kelvin?

ANSWER

1 The freezing point of water is 0°C.

$$\text{Temperature (°C)} = \text{temperature (K)} - 273$$

Rearrange the equation to make temperature (K) the subject:

$$\text{Temperature (K)} = \text{temperature (°C)} + 273$$

$$\text{Temperature (K)} = 0 + 273 = 273$$

The freezing point of water is 273 K.

2 The boiling point of water is 100°C.

$$\text{Temperature (°C)} = \text{temperature (K)} - 273$$

Rearrange the equation to make Temperature (K) the subject:

$$\text{Temperature (K)} = \text{temperature (°C)} + 273$$

$$\text{Temperature (K)} = 100 + 273 = 373\text{ K}$$

The boiling point of water is 373 K.

SECTION REVIEW 9.2

REMEMBERING

1 Define:

a chemical bond energy

b temperature

c absolute temperature scale.

UNDERSTANDING

2 Convert the following temperatures from °C to K.

a 25°C

b −77°C

3 Convert the following temperatures from K to °C.

a 635 K

b 138 K

4 Explain the forms of energy that contribute to total *KE* for a molecule.

5 Explain how the main sources of *KE* differ in solids, liquids and gases.

APPLYING

6 The mean temperature range for Bundaberg, Queensland, from 1959 to 2012 was 10.2°C. Calculate the mean temperature range over this period for Bundaberg in Kelvin.

9.3 Relationship between temperature and enthalpy

During a chemical reaction heat energy is either gained by, or absorbed from, the surroundings.

The amount of heat energy gained or lost (q) depends on the:

- capacity of the substance to absorb heat, the 'specific heat capacity' (C), as defined later in this section
- mass of the substances reacting (m), (the mass of the *limiting reagent* in the reaction)
- change in temperature (ΔT) (final temperature – initial temperature).

The amount of heat gained or lost (q) is directly proportional to each of these three factors, so it depends on the product of these three factors. This can be expressed mathematically as:

KEY FORMULA

$$q = mC\Delta T \text{ (units in kJ)}$$

where q is measured in joules (J); m is measured in grams (g); C is measured in $\text{J K}^{-1}\text{g}^{-1}$; and ΔT is the change in temperature (final temperature–initial temperature), measured in kelvin (K).

enthalpy change for a chemical reaction is the heat gained or lost per mole of reactant: enthalpy change = (enthalpy of products) – (enthalpy of reactants)

The heat gained or lost in a chemical reaction can be measured experimentally for a known mass of substance, with known specific heat, by measuring the temperature change from the beginning to the end of the reaction. This method is known as calorimetry. The **enthalpy change** associated with the chemical reaction is usually quoted as the heat gained or lost per mole of reactant. The enthalpy change is obtained by dividing q by the number of moles of reactant ($n = \frac{m}{M}$, where M is the molar mass of the reactant).

KEY FORMULA

$$\Delta H = \frac{q}{n} \text{ (units kJ mol}^{-1}\text{)}$$

It can also be expressed as the heat gained or lost per gram (or kilogram) of reactant, in which case:

specific heat capacity heat energy required to raise the temperature of 1 gram of a substance by 1 kelvin

KEY FORMULA

$$\Delta H = \frac{q}{m} \text{ (units J g}^{-1}\text{ or kg}^{-1}\text{)}$$

Specific heat capacity

FIGURE 9.3.1 Sand and grass have different capacities to retain heat.

Each substance has a different ability to hold heat. On a hot day, sand feels hot underfoot but grass feels cool (Figure 9.3.1). Even though the Sun is shining on both of them and they are receiving the same amount of heat energy, they feel very different. This is because they have different capacities to hold heat.

The capacity of different substances to hold heat is defined in terms of a standard unit of measure called the **specific heat capacity**, expressed as:

KEY FORMULA

The specific heat capacity (C) is the amount of heat (q, kJ) needed to increase the temperature of 1 g of a substance by 1 K.

Values for specific heat capacity provide a scale by which the ability of substances to retain (or lose) heat can be compared directly because the mass of the substances and the increase in temperature are taken to be the same. The only variable is the material's capacity to hold (or lose) heat. (Specific heat capacities are normally quoted in units of joules per degree kelvin, per gram. The actual value of the quantity quoted as $J K^{-1} g^{-1}$ is also the same in the alternate units, $kJ K^{-1} kg^{-1}$.)

The specific heat capacity of water is $4.18 J K^{-1} g^{-1}$. This means that 4.18 joules of energy is needed to increase the temperature of 1 gram of water by 1 kelvin (or 4.18 kJ per kg).

Analogous to the amount of heat energy associated with a chemical reaction, the total amount of energy needed to heat a substance (without it reacting) depends on the:

- mass of the substance (usually expressed in grams, g)
- specific heat capacity of the substance (usually in the units $J K^{-1} g^{-1}$)
- increase in temperature required (units Degrees; since this is a temperature change, the magnitude of the change is the same regardless of whether we calculate the difference of two temperatures in °C or in K).

As above, this may be written as:

$$q = mC\Delta T \text{ (units in kJ)}$$

If the temperature change is positive, then the temperature has increased and the substance has gained heat. If the temperature change is negative, then the temperature has decreased and the substance has released heat.

9.3.1 Specific heat capacity

WORKED EXAMPLE 9.3.1

Calculate how much heat was added if the temperature of 100 g water increased from 18°C to 60°C.

ANSWER

1 Extract the data from the equation and the textbook:

$$m = 100\,g,\ C = 4.18\,J\ K^{-1}\,g^{-1},\ \Delta T = 60 - 18 = 42°C = 42\,K$$

(This is a temperature *difference*, so the value for the difference is the same in °C as in K. Check for yourself from the temperature scales illustrated in Figure 9.2.3.)

2 Substitute the data into the equation:

$$q = mC\Delta T = 100 \times 4.18 \times 42 = 17\,556\,J = 18\,kJ$$

18 kJ of heat was added.

SECTION REVIEW 9.3

APPLYING

1 Calculate how much heat was added if the temperature of 100 g water increased from 10°C to 50°C.

2 Calculate how much heat was added if the temperature of 200 g water increased from 18°C to 60°C.

9.4 Inquiry skills: determining the heat change for a physical process

WORKED EXAMPLE 9.4.1

In an experiment, 26.8 kJ of heat was used to increase the temperature of 180 g of water. The initial temperature of the water was 22°C. What was the final temperature of the water?

ANSWER

1 Extract the data from the equation and the textbook:

$$m = 180\,\text{g};\ C = 4.18\,\text{J}\,\text{K}^{-1}\,\text{g}^{-1};\ q = 26.68\,\text{kJ} = 26\,800\,\text{J}$$

2 Rearrange the equation and substitute data:

$$\Delta T = \frac{q}{m \times c} = \frac{2\ 6800}{180 \times 4.18} = 35.6\,\text{K}$$

$$\Delta T = 35.6\,\text{K} = 35.6°\text{C}$$

$$\Delta T = T(\text{final}) - T(\text{initial})$$

Where: $T(\text{final}) = \Delta T + T(\text{initial}) = 35.6 + 22 = 57.6°\text{C}$

The final temperature of water was 57.6°C, rounded up to 58°C.

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PRACTICAL ACTIVITY 9.4.1

Comparing heat capacities

A common way to measure energy changes in the laboratory is to measure the change in temperature of water or a solution.

AIM

To compare the specific heat capacities of vegetable oil and water

EQUIPMENT PER GROUP

- 150 g of water
- 50 g of vegetable oil
- 3 × 150 mL beakers
- hot plate
- thermometer or temperature probe
- waste bottle for disposal of vegetable oil

RISK ASSESSMENT

WHAT ARE THE RISKS IN DOING THIS EXPERIMENT?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?
Vegetable oil is flammable.	Use a hot plate to heat the oil, not a Bunsen burner.
Splashing could occur as one liquid is added to another.	Wear safety glasses and pour carefully. Wash off any spills onto your skin immediately with water.

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, as well as ways to manage them.

PROCEDURE

1. Pour 50 g water into beaker A. Heat the water till its temperature reaches 50°C.
2. Pour 50 g water into beaker B and measure its temperature.
3. Pour the water from beaker A into beaker B, stir it and measure the final temperature of the water.
4. Empty beaker B.
5. Pour 50 g vegetable oil into beaker A. Heat the oil till its temperature reaches 50°C.
6. Pour 50 g water into beaker B and measure its temperature.
7. Pour the oil from beaker A into beaker B, stir it and measure the final temperature of the mixture.
8. Empty beaker B into a waste bottle.

RESULTS

1. Record the initial temperature of the water in beaker B for each experiment.
2. Record the final temperature of the water and water–oil mixture in beaker B for each experiment.

ANALYSIS OF RESULTS

In which experiment did beaker B have the greater temperature change?

DISCUSSION

Justify whether water or vegetable oil had the greater specific heat capacity.

SECTION REVIEW 9.4

APPLYING

1 Use the following data to answer the questions.

SUBSTANCE	SPECIFIC HEAT CAPACITY ($J K^{-1} g^{-1}$):
Ethanol	2.46
Vegetable oil	2.00

a How much heat is needed to increase the temperature of 15 g of ethanol by 40°C?

b What mass of vegetable oil was heated by 2346 J, if the temperature of the oil increased from 30°C to 60°C?

9.5 Investigating exothermic and endothermic reactions

When classifying a process as exothermic or endothermic, scientists measure changes in the temperature of the surroundings. An exothermic process releases heat, so the temperature of the immediate surroundings rises. In an endothermic process, heat is absorbed from the surroundings, and the surrounding temperature decreases.

Table 9.5.1 shows examples of endothermic and exothermic processes.

TABLE 9.5.1 Endothermic and exothermic processes

EXOTHERMIC PROCESSES	ENDOTHERMIC PROCESSES
Making ice cubes in a freezer	Melting of ice cubes or snow
Snow formation in clouds	Converting frost to water vapour
Water vapour condensing to liquid water (e.g. clouds turning to rain)	Evaporating water
Burning gas or petrol (combustion)	Producing sugar and O_2 by photosynthesis
Mixing sodium hydroxide pellets with water	Mixing water with ammonium nitrate
Mixing water with strong acids	Mixing water with cooking salt
Nuclear fission	Splitting apart a gas molecule (e.g. O_2 or N_2)
The rusting of iron in air	Baking bread
Burning sugar	Cooking an egg

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Hess's Law and enthalpy change calculations

In calculations involving enthalpy changes for reactions or physical processes, you can apply Hess's Law when there is more than one possible mechanism (reaction path) for the reactions that lead to a specific set of products:

> **Hess's Law**
> An enthalpy change accompanying a chemical change is independent of the route by which the chemical change occurs.
>
> KEY LAW

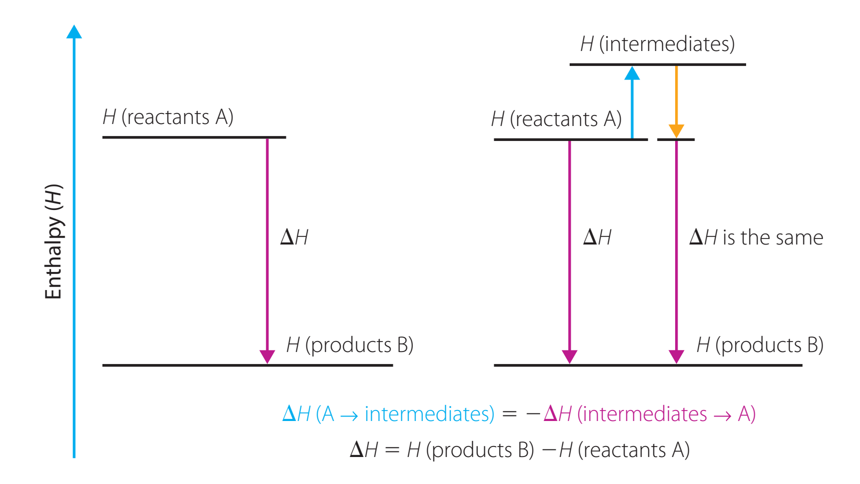

FIGURE 9.5.1 The enthalpy change for a reaction A → B is independent of the path taken for the reaction.

9.5.1 Hess's Law and reaction enthalpy change

Figure 9.5.1 shows the enthalpy change for an exothermic reaction where there are two different paths from reactants A to products B. The path at the left of Figure 9.5.1 is a direct reaction A → B. The other path involves some intermediates. For the reaction via the intermediates, write the reaction as A → (intermediates) → B. In both cases, the net enthalpy change from reactants to products is the same, because:

$$\Delta H = H_{\text{products}} - H_{\text{reactants}}$$

The starting reactants and final products are the same regardless of the intermediates that might participate in the reaction process. Even if some energy needs to be added to the reaction system to create the intermediates, that quantity of energy is released in the next step leading to the products, so you get that energy back again, as shown on the right of Figure 9.5.1.

In Practical activity 9.5.1 (page 163), you will examine and measure chemical and physical processes to classify them as endothermic or exothermic.

WORKED EXAMPLE 9.5.1

Given the following chemical equations, calculate the enthalpy change (ΔH) for the reaction that produces SO_3.

a $S + O_2 \rightarrow SO_2 \qquad \Delta H = -297\,kJ$

b $2SO_3 \rightarrow 2SO_2 + O_2 \qquad \Delta H = 198\,kJ$

ANSWER

1 The desired chemical equation is:

$$2S + 3O_2 \rightarrow 2SO_3$$

According to Hess's Law, an enthalpy change accompanying a chemical change is independent of the route by which the chemical change occurs. Hence, if a series of reactions are added together, the enthalpy change for the net reaction will be the sum of the enthalpy changes for the individual steps. This law enables you to calculate the enthalpy change for a chemical reaction using the strategy that any reaction can be represented by a series of steps for which the enthalpy changes are known. In these types of problems, the overall or net equation for which the enthalpy change is required guides you in how to construct the individual steps.

The desired equation shows two moles of sulfur reacting, so equation **a** must be doubled, multiplying all the coefficients by 2. When you double a reaction, ΔH must also be doubled because twice the energy will be released. Applying these changes, you have:

c $2S + 2O_2 \rightarrow 2SO_2 \qquad \Delta H = 2(-297\ kJ) = -594\ kJ$

2 For the desired reaction, SO_3 must be a product, so equation **b** should be reversed. When an equation is reversed, the sign of ΔH must be changed. The reverse of equation **b** is the following:

d $2SO_2 + O_2 \rightarrow 2SO_3 \qquad \Delta H = -198\ kJ$

3 Now, add equations **c** and **d** to obtain the overall equation, and add the ΔH values to determine ΔH for the overall or net equation. Any terms that are common to both sides of the combined equation can be cancelled.

c $2S + 2O_2 \rightarrow 2SO_2 \qquad \Delta H = 2(-297\ kJ) = -594\ kJ$

d $2SO_2 + O_2 \rightarrow 2SO_3 \qquad \Delta H = -198$

$2S + 3O_2 \rightarrow 2SO_3 \qquad \Delta H = -792\ kJ$

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PRACTICAL ACTIVITY 9.5.1

Endothermic and exothermic reactions

AIM

To observe chemical and physical processes and classify them as endothermic or exothermic

EQUIPMENT PER GROUP

- 4 g solid sodium hydroxide (NaOH)
- 10 g solid potassium nitrate (KNO_3)
- 5 g granulated zinc
- 50 mL 4 mol L^{-1} hydrochloric acid (HCl)
- 5 g sodium carbonate (Na_2CO_3)
- 50 mL 2 mol L^{-1} ethanoic (acetic) acid (CH_3COOH)
- 4 polystyrene cups
- thermometer or temperature probe
- stirring rod

WHAT ARE THE RISKS IN DOING THIS EXPERIMENT?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?
Acid could splash or spill onto your hands.	Wash your hands immediately if they come in contact with acid and after using acids.
Vigorous bubbling may cause splashes in your eyes.	Wear safety glasses.

RISK ASSESSMENT

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, as well as ways to manage them.

PROCEDURE

PART A

1. Pour 100 mL water into a polystyrene cup. Measure its temperature using the thermometer or temperature probe.
2. Add 4 g solid sodium hydroxide to the polystyrene cup, stir it and record the final temperature.

PART B

1. Pour 100 mL water into a polystyrene cup and measure its temperature.
2. Add 10 g solid potassium nitrate to the polystyrene cup, stir it and record the final temperature.

PART C

1. Pour 50 mL hydrochloric acid into a polystyrene cup and measure its temperature.
2. Add 4–5 pieces of granulated zinc to the polystyrene cup, stir it and record the final temperature.

PART D

1. Pour 50 mL ethanoic acid into a polystyrene cup and measure its temperature.
2. Add 5 g sodium carbonate to the polystyrene cup, stir it and record the final temperature.

»

»

RESULTS

1. Draw a table to clearly show your results.
2. Draw an appropriate graph to display your results visually.

DISCUSSION

1. Classify each of the processes as endothermic or exothermic, and justify your reasoning.
2. Write equations to communicate these processes.
3. Why were polystyrene cups used for these experiments?

Combustion reactions

combustion
exothermic chemical reaction between a fuel and an oxidising agent (usually oxygen) that produces energy

In **combustion** reactions, substances burn in the presence of oxygen and release large amounts of energy. Generally, combustion refers to burning fuels to release the energy needed for our daily activities. The general equation used for respiration (Worked example 8.1.1) is a combustion reaction that occurs in our body cells. Glucose ($C_6H_{12}O_6$) and oxygen gas (O_2) react to produce carbon dioxide gas CO_2(g) and water (H_2O):

$$C_6H_{12}O_6 + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l) \qquad \Delta H = -\sim 3000\,\text{kJ mol}^{-1}$$

During respiration, energy is released to enable us to breathe, talk, walk and carry out all of our everyday activities.

SYLLABUS LINK
Chapter 8 discusses respiration.

Since the Industrial Revolution, fossil fuels have been burned on a large scale to release energy. Power stations convert coal to electricity. Refined petroleum fuels our cars. Fossil fuels' sources have been produced over a period of millions of years from decayed animal and plant matter. Coal (predominantly elemental carbon), crude oil (hydrocarbons) and natural gas (predominantly methane, CH_4) are fossil fuels. They are the predominant sources of energy used in Australia.

FIGURE 9.5.2 Burning coal is an example of the combustion of a fossil fuel

Many fuels contain hydrocarbons, but can also be organic molecules composed of carbon, hydrogen and oxygen, such as ethanol (C_2H_5OH). Complete combustion occurs when these fuels react with oxygen (O_2) at standard atmospheric pressure to produce the oxides of carbon and hydrogen – carbon dioxide (CO_2) and water (H_2O). However, sometimes incomplete combustion occurs and carbon monoxide (CO) or carbon, as well as water, are the products. Incomplete combustion occurs when there is not enough oxygen for complete combustion.

For example, methane (CH_4), a major component of natural gas, undergoes complete or incomplete combustion depending on the amount of oxygen available.

Complete combustion produces carbon dioxide and water:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

Incomplete combustion produces carbon monoxide or carbon and water:

$$CH_4(g) + \frac{3}{2}O_2(g) \rightarrow CO(g) + 2H_2O(l)$$

$$CH_4(g) + O_2(g) \rightarrow C(s) + 2H_2O(l)$$

The formation of carbon or carbon monoxide requires less oxygen per mole of methane than the formation of carbon dioxide. In the incomplete burning of methane, the carbon produced ends up as small particles made up of clusters of many carbon atoms, called 'soot'.

Heat of combustion

The enthalpy change involved in combustion reactions is usually called the **heat of combustion**. In reality, this is an enthalpy of reaction, but it is given a special name. The 'heat of combustion' is further defined as the heat released when a fuel undergoes complete combustion at **standard atmospheric pressure**. Combustion reactions are always considered exothermic reactions. It is difficult to measure directly the heat released when a fuel undergoes combustion, so an indirect method is used by applying the law of conservation of energy. This is the method of calorimetry. The procedure used is to make use of the relationship described in section 9.3, connecting heat released in a reaction with mass, specific heat and temperature change:

heat of combustion heat released when a fuel undergoes complete combustion at standard atmospheric pressure

standard atmospheric pressure atmospheric pressure on Earth's surface, described as 1 atm, or 101.325 kPa

KEY FORMULA

$$q = mC\Delta T$$

Where:

q = heat content (J)

m = mass

C = specific heat capacity

T = temperature

ΔT = change in temperature

SYLLABUS LINK

Chapter 15 includes the definition of standard conditions.

Calorimetric experiments measure changes in heat. A known quantity of water is heated by the combustion of a fuel. The specific heat of water is known. The temperature change is measured and q may be calculated. Changes in heat can be measured using the equation:

Heat absorbed by water = heat released by combustion of fuel

The calorimetric method for determining an enthalpy change experimentally is reliable provided there is no heat loss to the surroundings from the water used to absorb the heat released by the combustion reaction.

SECTION REVIEW 9.5

REMEMBERING

1 Define:

- **a** Hess's Law
- **b** combustion
- **c** heat of combustion
- **d** standard atmospheric pressure.

APPLYING

2 Calculate ΔH for the following reaction:

$$Sn + 2Cl_2 \rightarrow SnCl_4$$

given these two equations:

$$Sn + Cl_2 \rightarrow SnCl_2 \quad \Delta H = -325\,kJ$$

$$SnCl_2 + Cl_2 \rightarrow SnCl_4 \quad \Delta H = -186\,kJ$$

3 Calculate ΔH for the following reaction:

$$2CO + 2NO \rightarrow 2CO_2 + N_2$$

given these two equations:

$$2CO + O_2 \rightarrow 2CO_2 \quad \Delta H = -566.0\,kJ$$

$$N_2 + O_2 \rightarrow 2NO \quad \Delta H = -180.6\,kJ$$

4 Calculate ΔH for the following reaction:

$$CS_2 + 2H_2O \rightarrow CO_2 + 2H_2S$$

given these two equations:

$$H_2S + \frac{3}{2}O_2 \rightarrow H_2O + SO_2 \qquad \Delta H = -563\,kJ$$
$$CS_2 + 3O_2 \rightarrow CO_2 + 2SO_2 \qquad \Delta H = -1075\,kJ$$

5 Calculate ΔH for the following reaction:

$$4Al + 3MnO_2 \rightarrow 2Al_2O_3 + 3Mn$$

given these two equations:

$$4Al + 3O_2 \rightarrow 2Al_2O_3 \qquad \Delta H = -3352\,kJ$$
$$Mn + O_2 \rightarrow MnO_2 \qquad \Delta H = -521\,kJ$$

9.6 Inquiry skills: determining enthalpy change for a reaction by calorimetry

WORKED EXAMPLE 9.6.1

A student clamped a conical flask containing 150 g water above a spirit burner that contained hexane (Figure 9.6.1). The temperature of the water was measured before and after the experiment and was found to increase by 20°C. The spirit burner containing hexane was weighed before and after the experiment. The mass of hexane that underwent combustion was 0.26 g. How much heat is released by the complete combustion of 1 g of hexane?

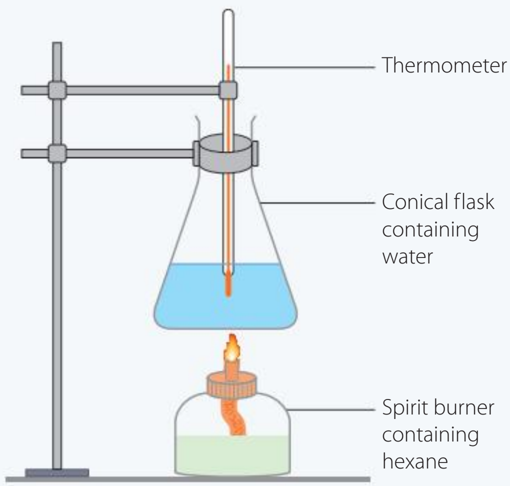

FIGURE 9.6.1 Equipment set-up to determine the heat of combustion for hexane.

ANSWER

1 Write down all of the data for water and determine the heat absorbed by the water.

- $m = 150\,g$

  $C = 4.18\,J\,K^{-1}\,g^{-1}$

  $\Delta T = 20°C$

  $q = mC\Delta T = 150\,g \times 4.18\,J\,K^{-1}\,g^{-1} \times 20°C = 12\,540\,J$

The law of conservation of energy states:

Heat absorbed by the water = heat released by combustion of hexane

2 Write down the data for hexane and use ratios to determine the heat released by 1 g of hexane.

Heat released by combustion of hexane = 12 540 J

Mass of hexane that was combusted = 0.26 g

0.26 g of hexane releases 12 540 J (i.e. 12.54 kJ)

$$1\text{ g of hexane releases } \frac{1}{0.26} \times 12540 = 48\,230\,J = 48\,kJ$$

1 g of hexane releases 48 kJ.

SECTION REVIEW 9.6

APPLYING

1 A student clamped a flask with 100 mL water above a spirit burner that contained 1-propanol (C_3H_7OH). The temperature of the water increased by 30°C. A mass of 0.372 g 1-propanol was used during this experiment. How much heat is released by 1 g of 1-propanol?

2 A student conducted an investigation to determine the heat of combustion ($kJ\,g^{-1}$) of 1-butanol. The specific heat capacity of water is $4.18\,J\,g^{-1}\,K^{-1}$. The student collected the following data:

- mass of water heated = 150 g
- initial temperature of water = 16.3°C
- final temperature of water = 39.6°C
- mass of 1-butanol used = 0.48 g.

a Use the student's data to calculate the heat of combustion ($kJ\,g^{-1}$) for 1-butanol.

b The accepted value for the heat of combustion for 1-butanol published in the scientific literature is $36.1\,kJ\,g^{-1}$. Explain why the student's value can be different from the literature value.

3 Calculate the enthalpy change for the reaction between ethene (C_2H_4) and hydrogen chloride (HCl) gases to make chloroethane gas (C_2H_5Cl) from the standard enthalpy of formation values given in the following table (this is an application of Hess's Law discussed in section 9.5).

COMPOUND	ΔH_F ($kJ\,mol^{-1}$)
$C_2H_4(g)$	+52.2
$HCl(g)$	−92.3
$C_2H_5Cl(g)$	−109

9.7 Mandatory practical

PRACTICAL ACTIVITY 9.7.1

Determining the heat of combustion of alkanols by calorimetry

Calorimetry was introduced in sections 9.3, 9.5 and 9.6. Here, a method similar to that outlined in Worked example 9.6.1 is applied to determine the heat of combustion of three alkanols (alcohols), namely propanol (C_3H_7OH), butanol (C_4H_9OH) and pentanol ($C_5H_{11}OH$), using the calorimetry method. The heat of combustion (ΔH_c) will be quoted as the heat (kJ) released by 1 g of the alkanol.

You must compare your measured average value with the values quoted in the scientific literature (NIST 1972): propanol: 33.61 (±0.05) $kJ\,g^{-1}$, butanol: 36.10 (±0.05) $kJ\,g^{-1}$ and pentanol 38.66 (±0.05) $kJ\,g^{-1}$. You must discuss the significance of any difference found between the literature and your experimental values. The values obtained for ΔH_c can be graphed against the number of C–H bonds in each alcohol in order to seek a correlation and rationalise the results in terms of energy stored in chemical bonds as well as the theoretical calculation of enthalpy change for these combustion reactions (see section 9.5).

AIM

To determine the heat of combustion for 1 g each of propanol (C_3H_7OH), butanol (C_4H_9OH) and pentanol ($C_5H_{11}OH$) using the method of calorimetry

EQUIPMENT PER GROUP

- approx. 0.1 g potassium permanganate ($KMnO_4$)
- copper calorimeter (beaker or pot) to contain water
- retort stand and clamps or tripod stand and heat mat
- spirit burner and cap
- 9 aluminium sheets – 15 cm × 15 cm (for insulating copper pot)
- thermometer or temperature probe
- digital mass balance (± 0.01 g)
- 100 mL measuring cylinder
- 50 mL measuring cylinder
- 100 mL each of propanol (C_3H_7OH), butanol (C_4H_9OH) and pentanol ($C_5H_{11}OH$)
- 1 L Tap water
- wood spacer blocks to raise height of spirit burner if a tripod stand is used for the calorimeter

RISK ASSESSMENT

Construct a table similar to the one below. Identify specific risks involved in the investigation and ways that you will manage the risks to avoid injuries or damage to equipment. Ask your teacher to check your risk assessment before you proceed.

»

WHAT ARE THE RISKS IN DOING THIS EXPERIMENT?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?
Alcohols are flammable liquids.	Keep away from flame; avoid spillage.
Equipment may get hot during this investigation.	Avoid touching hot equipment.
Spirit burner could leak and spilt alcohol could burn.	Wear safety glasses, gloves, aprons

RISK ASSESSMENT

PROCEDURE

1. Weigh the dry copper pot (calorimeter) and record the weight.
2. Use a 100 mL measuring cylinder to pour 50 mL tap water into the copper pot.
3. Reweigh the pot + water and record the mass of water.
4. Set up the equipment using the diagram in Worked example 9.6.1 (Figure 9.6.1) as a guide.
5. Use a 50 mL measuring cylinder to pour approximately 30 mL propanol into the spirit burner (or enough volume to cover the wick), weigh and record the mass of the spirit burner and propanol.
6. Place the spirit burner, allowing 2 cm space between the wick and the bottom of the calorimeter.
7. Use the aluminium foil to cover the water sample in the calorimeter, and wrap it around the sides of the calorimeter to help insulate the calorimeter from heat loss.
8. Place the thermometer carefully through a small hole pierced in the aluminium foil and clamp the thermometer so that its bulb is approximately in the middle of the 50 mL volume of water.
9. Record the initial temperature.
10. Use any strategy available to reduce the effect of air currents on the burner flame.
11. Light the spirit burner and wait until the water temperature has risen by approximately 35°C. Keep watch on the temperature. The thermometer can be used to gently stir the water to ensure uniformity.
12. Cap the spirit burner to extinguish the flame.
13. Record the highest temperature displayed by the thermometer.
14. Wait until the spirit burner has cooled. Weigh and record the mass of the spirit burner and remaining propanol to determine the mass of propanol that has undergone combustion.
15. Empty the calorimeter, rinse with cold water to restore its temperature to room temperature and refill with cold water as previously described; weigh it to determine the mass of water.
16. Refill the spirit burner to the same level as previously and reweigh.
17. Repeat the experiment to gather data (temperature change and mass propanol combusted) for three separate trials of propanol samples.
18. Repeat the experiment (three times each) for butanol and pentanol.

ANALYSING THE RESULTS

1. Write a balanced equation for the combustion of each alcohol in air.
2. Construct a table like the one below (a spreadsheet is recommended). Show the measured masses for the calorimeter, calorimeter and water, the temperature rise of the water in the calorimeter, the masses for spirit burner (and alcohol) before and after combustion.

EXPERIMENT NUMBER	MASS OF EMPTY CALORIMETER	MASS OF CALORIMETER + WATER	MASS OF WATER	INITIAL TEMPERATURE	FINAL TEMPERATURE	TEMPERATURE RISE	INITIAL MASS OF BURNER + ALCOHOL	FINAL MASS OF BURNER + ALCOHOL	MASS OF ALCOHOL COMBUSTED

3. Calculate the mass of water, temperature rise and mass of alcohol combusted for each trial.
4. Confirm the calculations by carrying out one calculation by hand (calculator).

5 Determine the heat absorbed by the water in the calorimeter for each alcohol sample using $q = mC\Delta T$ where the specific heat of water C is taken to be $4.18\,JK^{-1}g^{-1}$ (section 9.3 and Worked Example 9.3.1). This equals the heat released by the mass of alcohol combusted. Use the procedure shown in Worked example 9.6.1 to determine the heat released by 1 g of each of the alcohols. Once again, use your spreadsheet for your calculations but demonstrate the validity of one calculation by hand.

6 Using your results for the nine trials (three for each alcohol), determine the average experimental value for the heat of combustion of each alcohol (units: $kJ\,g^{-1}$). Include an estimate of the uncertainty of the experimental measurement (refer to Chapter 10 for advice on how to arrive at estimates for uncertainty in an experimental procedure).

DISCUSSION

Compare your experimental value for the heat of combustion of each alcohol with the accepted literature values: propanol: 33.61 (±0.05) $kJ\,g^{-1}$, butanol: 36.10 (±0.05) $kJ\,g^{-1}$ and pentanol 38.66 (±0.05) $kJ\,g^{-1}$. Include uncertainties (as percentages or absolute uncertainties) for both.

1 State the outcome of the experiment.

2 Evaluate the consistency of the results for the nine trials and decide whether it is justifiable to retain the three estimates for the heat of combustion for each alcohol. If not, revise your result for the average by eliminating any outlier data (refer to Chapter 10 to see how to identify outliers).

3 For each alcohol, indicate whether your value was greater than, equal to or less than the literature value.

4 Specify the magnitude of any difference and discuss whether the difference is significant when experimental uncertainties are taken into consideration.

5 Identify possible sources of errors and uncertainties and describe how each would affect the final result. It is likely that not all the heat from the spirit flame was absorbed by the calorimeter (why may this be so?). Explain how this may help to identify differences between your experimental values and the literature values.

6 For each source of error and uncertainty, suggest ways to modify your investigation to minimise the errors and uncertainties.

7 Plot both experimental and literature values for ΔH_c for each alcohol versus the number of CH bonds in the alcohol (all on the same graph). Discuss whether a correlation is observed. Comment on how any correlation observed can be understood in terms of the meaning and theoretical calculation of the enthalpy change for a chemical reaction (section 9.5).

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 Define:

a enthalpy
b temperature
c endothermic
d specific heat capacity
e kinetic energy (*KE*)
f chemical bond energy
g Hess's Law
h exothermic.

CATEGORY QUESTIONS

2 Explain how temperature relates to *KE*.

3 Show the relationship between temperature in degrees Celsius and temperature in kelvin.

4 Classify the following processes as endothermic or exothermic and justify your answers.

a Water freezing to become ice
b Sublimation of dry ice (solid CO_2) to gaseous carbon dioxide
c $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$ $\Delta H = -58\,kJ$
d Photosynthesis

ELABORATION QUESTIONS

5 Explain why biodiesel and diesel have different heats of combustion.

6 Evaluate this statement: 'The kelvin temperature scale is a useful scale for chemists.'

EVIDENCE QUESTIONS

7 Exothermic reactions are more important than endothermic reactions. Assess this statement with reference to specific examples.

8 A student wanted to heat some water in the science laboratory using as little heat as possible. Use the information in the following table of specific heat capacities to justify whether the student should put the water into a Pyrex glass beaker or a copper beaker during the heating process.

SUBSTANCE	SPECIFIC HEAT CAPACITY ($J\,K^{-1}\,g^{-1}$)
Aluminium	0.897
Copper	0.385
Pyrex glass	0.750
Tin	0.228
Vegetable oil	2.00

Refer to the following table of specific heat capacities to justify:

a why polystyrene cups were used instead of Pyrex glass beakers in Practical activity 9.5.1.
b whether copper, Pyrex glass or polystyrene should be used to hold water when it is placed over a lit spirit burner when conducting a heat of combustion investigation.

END-OF-CHAPTER EXAM

End-of-chapter test

SUBSTANCE	SPECIFIC HEAT CAPACITY ($J\,K^{-1}\,g^{-1}$)
Copper	0.385
Pyrex glass	0.750
polystyrene	1.3

1 Which of the following statements is false?

A Chemicals react to achieve stability.

B More energy is required to break bonds than to re-form bonds in a chemical reaction.

C Mass is conserved in chemical reactions.

D Bonds are broken and new bonds form in chemical reactions.

2 If energy does not need to be added to a reaction to make it occur at room temperature, then it can be referred to as a:

A reversible reaction.

B combustion reaction.

C spontaneous reaction.

D non-spontaneous reaction.

3 A student adds a chemical to water and stirs it with a thermometer until it has dissolved. They correctly deduce that it is endothermic. What observation could the student have made to deduce this?

A The chemical dissolved; it would not dissolve if it were exothermic.

B The temperature did not change, so energy needs to be added.

C The temperature of the solution was lower than the initial temperature of the water.

D The temperature of the solution was higher than the initial temperature of the water.

4 At what temperature does absolute zero occur?

A 0 K

B 0°C

C −273 K

D 273°C

5 Water has a specific heat capacity of $4.18\,J\,K^{-1}\,g^{-1}$ and ethanol has a specific heat capacity of $2.46\,J\,K^{-1}\,g^{-1}$. Which is true?

A Water requires more energy to raise the temperature by 1 K than the same mass of ethanol.

B Water requires less energy to raise the temperature by 1 K than the same mass of ethanol.

C Ethanol is always hotter than water.

D 2 grams of ethanol will have a higher specific heat capacity than water.

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6 A dietician used the phrase 'Food is fuel! It provides us with energy'. Which of the following reactions does this describe?

A Combustion

B Respiration

C Both A and B

D None of the above

7 Which of the following demonstrates that complete combustion has occurred/is occurring?

A The flame of a Bunsen burner is blue.

B The flame of a Bunsen burner is orange.

C The bottom of a beaker is black after being heated with a spirit burner containing a fuel.

D Limewater does not turn cloudy when the gas released upon burning a fuel is bubbled through it.

8 Copper(II) nitrate is heated in a crucible over a Bunsen burner, forming solid copper(II) oxide, nitrogen dioxide and oxygen gas.

a Write a balanced chemical equation for this reaction.

b Is this an exothermic or endothermic reaction? Explain your answer.

9 Draw an energy profile diagram to help you describe endothermic reactions. Include the following on your diagram: enthalpy, reactants, products, ΔH, $H_{\text{reactants}}$, H_{products}.

10 a Define specific heat capacity.

b 26.8 kJ of heat was used to increase the temperature of 180 g of a substance from 22°C to 57.6°C. Determine the specific heat capacity of this substance.

11 How much heat is absorbed during photosynthesis when 9.22 grams of glucose $C_6H_{12}O_6$ is produced?

$$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2 \qquad \Delta H = +2803\text{ kJ}$$

12 Octane (C_8H_{18}) is one of many components of gasoline. The heat produced by burning 1.00 g of octane is collected in 80.60 g of water in a calorimeter. If the temperature rise is 16.24 K, calculate the heat of combustion of octane in units $JK^{-1}g^{-1}$. (Use $C_{\text{water}} = 4.18\text{ J K}^{-1}\text{ g}^{-1}$)

13 Calculate the enthalpy change for the reaction between NO and O atoms that occurs in the atmosphere. (The O atoms are produced from the catalytic destruction of ozone). The NO reaction is:

$$NO + O \rightarrow NO_2$$

The following data is required to make use of Hess's Law:

$$O_2 \rightarrow 2O \qquad \Delta H = +495\text{ kJ}$$
$$2O_3 \rightarrow 3O_2 \qquad \Delta H = -427\text{ kJ}$$
$$NO + O_3 \rightarrow NO_2 + O_2 \qquad \Delta H = -199\text{ kJ}$$

10 SCIENTIFIC MEASUREMENTS – UNCERTAINTY AND ERROR

Introduction

Scientific inquiry is about finding things out through observation and experiment. This is why investigations based on unbiased observations and rigorous data analyses are central to science, and why they are an important ingredient in many other disciplines.

Large-scale scientific investigations can sometimes take years to complete and may involve collaboration among many scientists. They may require access to special equipment in Australia or other countries. They may cost a lot of money, sometimes millions of dollars, to complete. Scientists invest time in planning investigations before they begin. When scientists apply for grants to conduct investigations, they need to show that they have planned their research project carefully and justify how any money provided will be spent. Good planning is crucial to an investigation's success.

Measurements, observations and recording of results are central to scientific investigation. Scientists keep records of all their experiments. This is often a legal requirement. Typically, experimental results need to be kept for five to seven years. There are also requirements on how and where data are stored.

Scientific data, once collected, need to be analysed. There are various ways to do this; in the physical sciences (physics, chemistry, geology), it almost always involves constructing tables and/or graphs. Once a correlation is established graphically, a mathematical relationship can be derived and often, predictions can be made.

Finally, the results of the investigation must be communicated. Scientists working in industry will be required to produce a report, often with an executive summary, together with recommendations, for a company manager or the board of directors. Research scientists working in pure research organisations report their findings by publishing a scientific paper in either a journal or as conference proceedings. Reporting on an investigation often includes presenting the results as a spoken presentation or as a poster at a conference or company workshop. If the research is funded by a grant, then a research report must be submitted. If the results are really exciting, then the scientists may write a media release. Whichever way the results are communicated, a report is required for the investigation to be complete.

Stimulus questions

What is an uncertainty in a measured quantity? What are measurement standards?

How do scientists determine the precision of a measurement and the accuracy of measured quantities?

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10.1 Making measurements

When you perform scientific investigations, you need to:

- plan carefully
- find out about the topic by reading about what other people have done
- collect data from your own experiments and from secondary sources
- analyse the data to draw conclusions about how the data support or refute hypotheses or theories
- communicate your findings.

When planning an experimental procedure, consider the following questions.

- How many variations of your **independent variables** will you test? For example, how many different times or temperatures will you have?
- How are the **measurements** to be made?
- What apparatus is required?
- What **dependent variables** (e.g. concentrations or absorbance) are you planning to measure?
- How will you measure your dependent variable?
- How many times will you repeat the experiment to test the reproducibility of measurements?

independent variable a variable upon which another variable depends

measurements using a measurement procedure together with a measuring standard to determine a measurement

dependent variable the variable that changes as a result of changes to the independent variable

variable something that can change or be changed, as distinct from a constant, which does not

In a typical experiment, one independent variable is varied. A chosen domain (range of values) for that **variable** forms the basis of the testing. Everything else that could influence any of the dependent variables (those that depend on the independent variable) is kept constant. If a scientist measures the temperature of a sample of heated liquid (e.g. in a range from 25°C to 80°C) as a function of time as it cools (e.g. in a time domain from 0 to 30 minutes), the volume of that sample, its concentration, and any aspects of its environment that may affect temperature gain or loss are kept constant.

Figure 10.1.1 graphs the data obtained from measuring the concentrations (dependent variables) versus time (the independent variable) for dinitrogen pentoxide, nitrogen dioxide and oxygen following

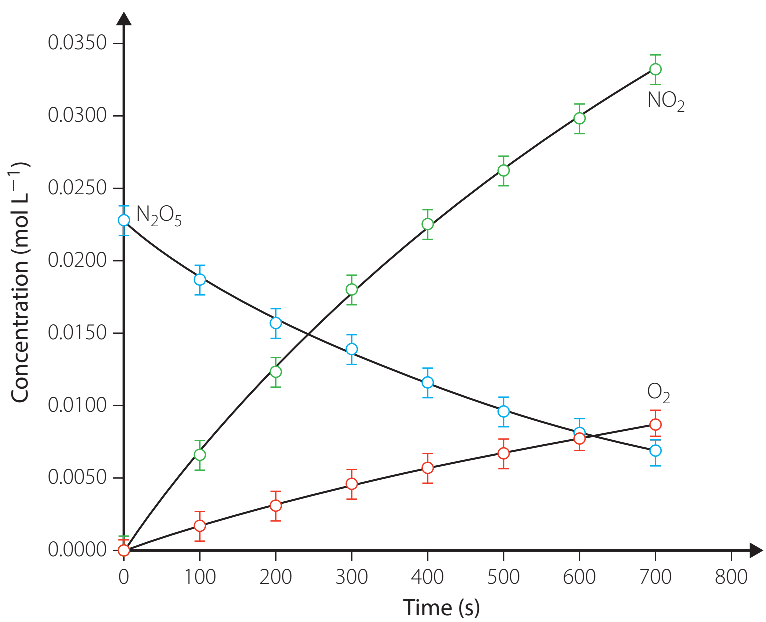

FIGURE 10.1.1 Graph showing concentration versus time for the dissociation of dinitrogen pentoxide (N_2O_5) to yield products nitrogen dioxide (NO_2) and oxygen (O_2). The curves fitted to the data points are based on a theoretical model for the chemical reaction.

the dissociation of dinitrogen pentoxide. This is an important process in atmospheric chemistry. The blue data points show that the concentration of dinitrogen pentoxide N_2O_5 decreases as it decomposes, while the concentrations of the products (nitrogen dioxide and oxygen) increase with time. The curves fitted to the data points are based on a theoretical model for the chemical reaction. They are not obtained merely by 'joining the dots'. Figure 10.1.1 displays a typical example of obtaining data, plotting the data on a graph, establishing an experimental hypothesis for the mechanism of the reaction (dissociation of N_2O_5) and confirming the hypothesis by fitting curves to the data using mathematical relationships derived from the theoretical model. Worked example 10.7.1 (page 189) explores a related experiment in more detail.

All experimental scientific reports present data based on measurements, observations and analyses of the data. The credibility of the conclusions presented in the report depends on the data's reliability. This requires the scientist to be honest and clear about the level of certainty of the measurements. For example, if a soil analysis concludes that toxic compound levels are above the legal limits for agricultural activities in the region, the confidence limits of the tests and analyses carried out must be included in the report.

For example, the Environment Protection Agency (EPA) set the legal maximum limit for polychlorinated biphenyl (PCB) concentrations in soil at $50\,mg\,kg^{-1}$. Suppose you conduct a soil analysis in five places and find that the measured levels are $5\,mg\,kg^{-1}$, $11\,mg\,kg^{-1}$, $53\,mg\,kg^{-1}$, $56\,mg\,kg^{-1}$ and $170\,mg\,kg^{-1}$. How do you report these findings? Do you say that the levels measured are sufficient to indicate that the area contains PCB concentrations above the legal limit? If so, how confident are you in your measurements? The EPA will ask you to state the uncertainties in each measurement as well as to provide a statistical analysis to demonstrate the probability that PCB levels are in excess of the legal limit over the whole soil region. Unless the data are sound in terms of errors, uncertainties and a statistical analysis, the EPA cannot make a recommendation. Nor can you as a scientist.

Whenever possible, measurements should be repeated. This allows you to check that the measurements are **reproducible**. If the results are similar each time, then your results are reproducible and are probably valid. If a result is not reproducible, it is probably not a valid result. A result is reproducible if you make exactly the same measurement more than once and, within the limits of experimental uncertainty, get the same result.

If a result is not reproducible, then a variable other than the one you are controlling may be affecting the measurement of the dependent variable. If this is the case, you need to determine what this other variable is, and control it if possible.

Estimating uncertainties

A measurement of some quantity (e.g. the volume of a liquid, the mass of a chemical sample, or the diameter of a pipe and length of a wire) is carried out by using a **measuring standard** (such as a measuring cylinder, a digital balance, vernier calipers or tape measure) and applying a **measurement procedure** to determine the measurement.

Once you have taken the measurement, the underlying question is, 'what is the reliability of the measurement?'

If scientists measure the weight of a glass of water using bathroom scales (with a digital display reading to 0.1 kg) and report it as '0.3 kg', what does that mean? If the same scientists measure the glass again on a laboratory scale, they might report it now as 303.5 g (or 0.3035 kg). The first measurement is an **estimate** of the weight of the glass of water. The second measurement is quoted to a higher **level of precision** (0.0305 kg compared with 0.3 kg – four digits rather than only one are now quoted after the decimal place). Both are legitimate measurements, but they provide different levels of information.

The meaning of reporting the first measurement (0.3 kg) is that the weight of the glass of water is known only to within approximately 0.1 kg. Scientists may believe it is not as heavy as 0.4 kg, nor is it 0.2 kg. However, it might be only a bit above or below 0.3 kg – but cannot say by how much. The **standard of measure** (the bathroom scale) does not have the capability to measure more precisely. In

reproducible
obtaining the same result, within the limits of experimental uncertainty, after carrying out exactly the same measurement more than once

measuring standard
calibrated device for making measurements

measurement procedure
a defined strategy for making a measurement using a measuring device

estimate
the outcome of a measurement

level of precision
the resolution of a measurement, typically defined by the number of decimal digits given for a measurement

standard of measure
an object, system, or experiment that bears a defined relationship to a unit of measurement of a physical quantity

addition, the bathroom scale may be covered in dust and is possibly worn. Does the scientist know if it has been calibrated to an accuracy of within half a kilogram? The measurement of 0.3 kg may be only coincidentally close to the 'real' weight of 303.5 g, so scientists establish the accuracy of a standard of measure before reporting a measurement using that standard.

Standards of measure are maintained by government agencies, including the National Institute of Standards Technology (NIST) in the United States. At NIST in Boulder, Colorado, the NIST-F2 atomic clock was launched in 2014. It was the latest in a series of caesium-based atomic clocks developed by NIST since the 1950s. Despite its stability of 1 part in 10^{16} (i.e. ±1 s in 300 million years), a new standard has been developed. In 2016, NIST reported a new dual atomic clock based on ytterbium atoms, with a stability of 1.5 parts in 10^{18} (Figure 10.1.2). Many modern devices taken for granted, including mobile phones, satellite TV transmission, global positioning systems, and even the operation of the electric power grid, rely on the ultra-high accuracy of atomic clocks.

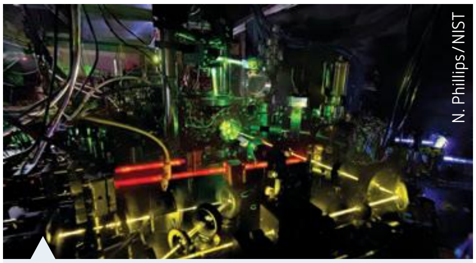

FIGURE 10.1.2 The National Institute of Standards Technology (NIST)-ZDT dual atomic clock, a new time standard

accuracy the degree to which the result of a measurement, calculation, or specification conforms to the correct value or a standard

NIST also conducts scientific research in fundamental physics and chemistry and, from this research, has developed a new method of defining the kilogram. The definition uses a 'scale', called the NIST-4 watt balance (Figure 10.1.3). This new scale has been used to measure a fundamental physical quantity called Planck's constant to within 34 parts per billion. However, to accept this new value of Planck's constant, and the kilogram, in the record book of standards, multiple measurements are required. Scientists in five countries have built watt balances to make independent measurements that can be compared. For the redefinition of the kilogram to be accepted, at least three experiments must produce values with a relative standard uncertainty of no more than 50 parts per billion, and one with no more than 20 parts per billion. These values must agree within a statistical confidence level of 95%.

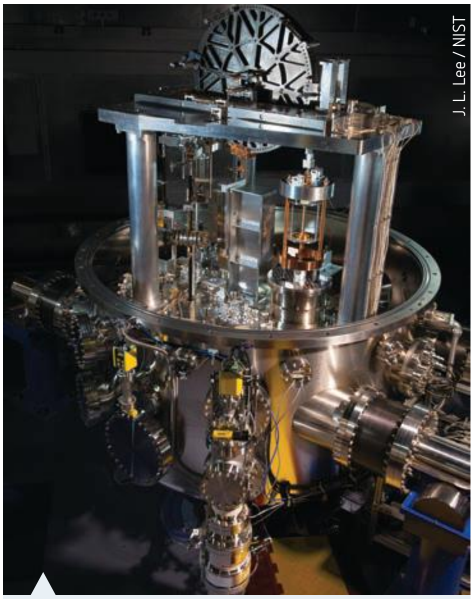

FIGURE 10.1.3 The National Institute of Standards Technology (NIST)-4 watt balance – providing a redefinition of the kilogram

Uncertainties in measurement procedures

A standard of measure requires a measurement device, such as a rule or tape measure to measure a length, a digital scale to measure a mass, or a burette to dispense a volume of liquid. A measurement procedure must be determined and evaluated regardless of the device. A rule or tape measure must be aligned with the object being measured. A digital scale can be zeroed before making the measurement. A burette is filled above the marking for 'zero' volume and a small volume of liquid is dispensed to waste before taking an initial reading. In the case of the burette, it does not matter if the initial reading is not exactly at zero. After dispensing a volume, the dispensed volume is simply the difference between the final and initial readings.

Parallax uncertainty

For all devices, uncertainty occurs because of the interface or interaction of the person taking the reading and the physical limitations of the measurement device. With a rule or tape measure, this may be parallax uncertainty, as well as the estimate of the reading if it falls between the smallest scale divisions. A digital scale displays a fixed number of digits – what about the digit that is not shown?

To read the level of a liquid in a burette, you need to decide from which part of the meniscus to take the reading; parallax uncertainty is also present (Figure 10.1.4, page 178). The level of the liquid might also be somewhere between the minimum markings. If the meniscus level lies between the burette markings (0.1 mL), you need to estimate the value of the second digit after the decimal point. There is always some uncertainty in making a visual estimate.

standards the fundamental reference for weights and measures, against which all other measuring devices are compared; administered by government agencies

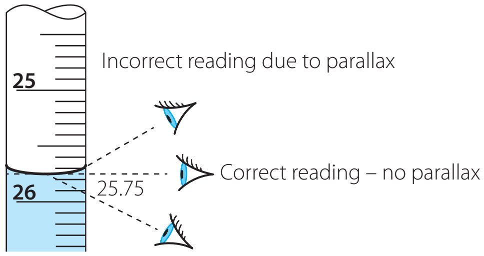

FIGURE 10.1.4 Reading a meniscus level on a burette can give rise to parallax error as well as uncertainty due to estimation of the liquid level if it is between minimum markings.

Reaction time

When you measure the time elapsed during an event using a stopwatch, you need to press a button to start the watch after observing the beginning, and to stop the watch at the end. There is a small time lapse between seeing the beginning of the event and pressing the start button. Typically this is 0.2 s, but it may vary; for example, from 0.1 s to 0.3 s. This is called the **reaction time**. A reaction time is associated with pressing both the start and the stop buttons. The measurement of the elapsed time will always have an uncertainty caused by a combination of the two reaction times. This uncertainty will normally always *add* to the actual start time for the observed event (unless you 'jump the gun'). If, during the stopwatch time reading, your reaction time was exactly +0.2 s at the beginning and at the end, the reaction times would cancel each other out, and the stopwatch reading would be correct. However, because reaction times vary, 'random uncertainty' occurs due to the reaction times being randomly different at the beginning and at the end.

How do scientists estimate the reaction time uncertainty? Generally, they establish a mean value for reaction time by pressing 'start', then 'stop', as quickly as possible a few times, averaging the readings and noting the range of times. For example, if a scientist obtained a mean reaction time via this operation of 0.20 s with a range of 0.09–0.31 s. They could conclude that there will be up to 0.2 s uncertainty at each end of the measurement of elapsed time, but it may be 0.2 ± 0.1 s. If a scientist has recorded that a reaction time is 0.2 ± 0.1 s, the uncertainty in the measured elapsed time will be 0.3–0.1 = 0.2 s. The convention often used with stopwatch readings is to record the time elapsed as:

$$\text{Time elapsed} = (\text{recorded time}) \pm 2 \times (\text{mean reaction time})$$

Responsible scientists conclude that the variation in the ability to synchronise the triggering of the stopwatch with the start and finish of the event introduces uncertainty in the measurement. This uncertainty will usually be larger than a digital stopwatch's instrument resolution, typically ± 0.01s.

There is uncertainty in every measurement. A measurement can be a best estimate with respect to a standard, but the measurement is meaningful only when it is accompanied by a quote of its level of uncertainty. This is often also expressed as the confidence factor for the measurement.

relative standard uncertainty
the standard uncertainty of a measurement result is the standard deviation of the measurement (as defined for a statistical normal distribution). The relative standard uncertainty is the ratio of the standard uncertainty and the measurement itself; that is, the fractional standard uncertainty

parallax uncertainty
the uncertainty that arises from the change in the apparent position of an object (e.g. scale reading for a liquid level) when viewed from different points

reaction time
the time taken for a person or system to respond following a given stimulus or event

Limitations of measurements

The limitations of a standard of measure, or limit of reading, are different for analogue and digital devices.

Analogue devices include swinging needle multimeters, liquid in glass thermometers and clocks with hands. Analogue devices have continuous scales.

For an analogue device, the limit of reading, sometimes called the 'resolution', is conventionally taken as ± half the smallest division on the scale. The **uncertainty of analogue scale** is taken as half the smallest division because you will generally be able to see which division mark that the indicator is closest, such as a needle or a fluid level. You may be able to estimate the measurement to one-fifth or even one-tenth of the smallest division if the spacing between divisions is large. However, the limit of reading uncertainty is normally defined as half of the smallest division. So, for a liquid in a glass thermometer with a scale marked in units of 1°C, the uncertainty is ± 0.5°C.

Digital devices such as digital multimeters, clocks and thermometers have a scale that gives you a number. The measurement is limited to a specific number of digits, typically three or four, so it is a discrete scale. A digital device has an uncertainty of a whole division since the next digit is totally unknown, but it could be anywhere from zero to 0.999 of one unit. So a digital thermometer that reads to whole degrees has an uncertainty of ± 1°C, which is an indicator of the **uncertainty of a digital scale**.

The resolution or limit of reading is the minimum uncertainty in any measurement. Usually the uncertainty is greater than this minimum.

uncertainty of analogue scale
± (half the smallest division)

uncertainty of a digital scale
± (the smallest division)

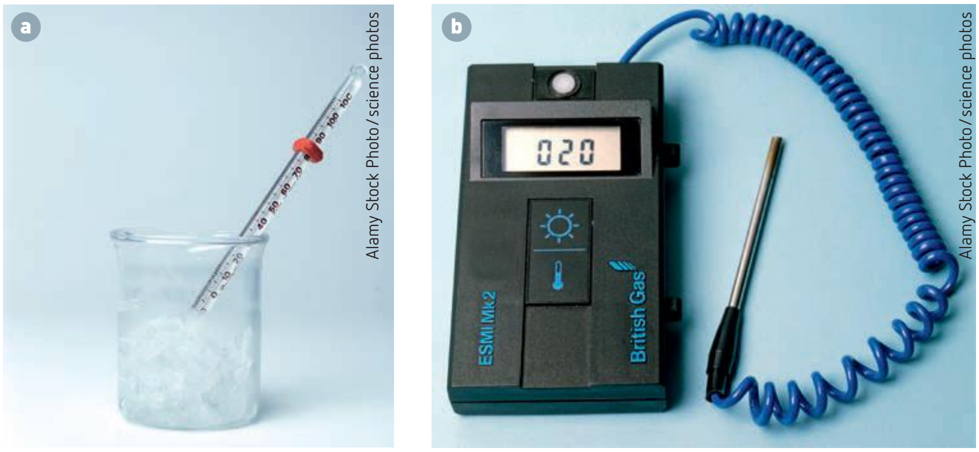

FIGURE 10.1.5
a This analogue thermometer has an uncertainty of half the smallest division on the scale.
b This digital thermometer has an uncertainty of a degree.

Precision of measuring devices

The measuring device used will have a **precision**, usually given in the user manual. The precision defines the closeness of several measurements of the same quantity to each other; that is, the repeatability of the measurement (see section 10.2). For example, a multimeter may have a precision of 0.5% on a voltage scale. If you measure a potential difference of 12.55 V on this scale, the uncertainty due to the precision of the meter is:

precision
the closeness of several independent measurements of the same quantity to each other

$$\frac{0.5}{100} \times 12.55\,\text{V} = 0.005 \times 12.55\,\text{V}$$
$$= 0.06\,\text{V}$$

This is greater than the limit of reading uncertainty, which is 0.01 V in this case (equal to a whole division of the last digit displayed).

Digital devices may seem more precise than analogue devices, but this is often not the case. A digital device may be easier for you to read, but this does not mean it is more precise. The uncertainty due to the limited precision of the device is generally greater than the limit of reading.

Accuracy of measuring devices

Any measuring device has an accuracy that can be determined only by comparison with a precisely known primary standard of measure. For example, the National Measurement Institute is responsible for maintaining Australia's official standard of length. In 1983, the metre was defined as the distance travelled by light in $1/c$ seconds, where c is the velocity of light in a vacuum and has the value 299 792 458 m s^{-1}. The Australian standard of length is realised by three He–Ne lasers that operate in the red (633 nm), orange (612 nm) and green (543 nm) and are frequency stabilised to iodine absorption lines. The length of a gauge block (e.g. a length of rigid metal bar) is determined using interferometry, a highly precise measurement technique based on interference of light waves.

Calibration is the comparison of the measuring device against a known primary standard. A metre rule is calibrated with respect to the standard metre, but usually the calibration is carried out by making the marks on the ruler with a machine that provides a secondary standard of measure based on a gauge block. The accuracy of the ruler is determined by the accuracy of the secondary standard as well as the calibration procedure (the precision of the device making the marks on the steel rule).

calibration
a comparison, and if required, adjustment of a measuring device against a known primary standard

SECTION REVIEW 10.1

REMEMBERING

1 Define:

a measurement **b** measurement procedure **c** standard of measurement.

UNDERSTANDING

2 How do you calibrate an instrument?

3 What is a secondary standard?

10.2 Precision versus accuracy

A ruler can expand and contract with temperature, so it will give a different reading of length depending on the ambient temperature. A 25 mL volumetric pipette may be both precise and accurate to within ± 0.1 mL at 20°C, but if the liquid being measured is at 80°C, the expansion of the pipette will introduce an inaccuracy to the delivered volume, even though the delivery of the volume of liquid may be precise to within ± 0.1 mL.

For instance, if a scientist delivers 10 samples of what she believes is 25 mL of liquid at 80°C from the pipette, she estimates that they will all be precise to within ± 0.1 mL but the delivered volume may be 25.5 mL; that is, inaccurate by 0.5 mL.

A swimmer might break the 100-metre freestyle world record three times in a row in their 50-metre training pool with times of 46.30 s, 46.31 s and 46.36 s. That is a precise result that will possibly get the coach quite excited. However, if the pool is short by 1 metre, the times recorded are all invalid because the length of the pool does not comply with the standard length. The swim times recorded were precise to within 0.1%. This is calculated from the mean recording and spread of times (46.32 ± 0.03 s), which is equivalent to 46.32 ± 0.1%. However, because the pool was not 50 m long, the recorded swim times were inaccurate; they were 2% less than the accurate value that would have been obtained in an exactly 50 m pool. The 2% inaccuracy arises because the pool is short by 1 m in 50 m; i.e. 2% too short.

Good precision means your measurements will all lie within a small range (Figure 10.2.1b). However, the measurements may be a long way from the truth if the device you are using to make the measurements, although precise, is inaccurate because of calibration or non-compliance with the accepted standard of measure.

a

Courtesy of John Morris Scientific

b

Function	Precision
pH range	0.00 to 14.00 pH
Resolution	0.01 pH
Accuracy	+/− 0.01 pH
pH slope range	80 to 120%
No. of calibration points	1 to 3 points (push-button)
Buffer options	pH 4.01, 7.00, 10.01 (USA) pH 4.01, 6.86, 9.18 (NIST) pH 4.10, 6.97 (Pb)
Temperature range	0.0 to 100.0°C
Resolution	0.1°C
Accuracy	± 0.5°C
Temperature comp.	Automatic/Manual (0 to 100°C)

Courtesy of John Morris Scientific

FIGURE 10.2.1 **a** A typical pH meter. **b** A page from the user manual giving the precision of various scales.

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In summary, the precision of the measurement is still determined by the precision of the measuring device, together with the precision associated with the measurement procedure. For a metre rule, this is given by half of the value of the smallest division; that is, ±0.5 mm. However, if the metre rule is 1005 mm long instead of 1000 mm, the accuracy of our measurement is compromised even if precise measurements are taken. Table 10.2.1 shows the relationship between precision and accuracy for measuring length.

10.2.1 Accuracy vs precision

TABLE 10.2.1 The relationship between precision and accuracy for measuring length

INSTRUMENT	LENGTH TO BE MEASURED	SMALLEST SCALE DIVISION (mm)	UNCERTAINTY (mm)	TYPICAL ACCURACY (mm)
Steel metre rule	To measure length to within 1 m	1	± 0.5	0.25; depends on manufacturer
Vernier calipers	To measure small lengths, internal and external diameters of objects	0.1	± 0.05	0.03–0.1; depends on manufacturer
Micrometer screw gauge	To measure very small lengths, diameters or thicknesses	0.01	± 0.005	Often about 0.01; depends on manufacturer 0.001 for some ultra-high accuracy instruments

Variation of the measured quantity

The object, or the quantity being measured, may itself vary. For example, reaction rate depends strongly on the temperature, concentration and other factors. Even keeping the conditions as close to identical as possible, it is unlikely that repeat experiments will give you exactly the same results. Making repeat measurements allows you to estimate the size of the variation.

Sometimes you can see how the measured quantity varies during a measurement by watching a needle move or the readings change on a digital device. Watch and record the maximum and minimum values. The difference between these is the range:

$$\text{Range} = \text{maximum value} - \text{minimum value}$$

The quoted measurement is the average value (mean), or the centre of the range:

$$\text{Quoted measurement} = \frac{\text{maximum value} + \text{minimum value}}{2}$$

The uncertainty in the measurement is half the range:

$$\text{Uncertainty} = \frac{1}{2}\text{range} = \frac{1}{2}(\text{maximum value} - \text{minimum value})$$

For example, if you are using an analogue multimeter and you observe that the needle fluctuates between 12.2 V and 12.6 V, then your measurement should be recorded as:

$$(12.4 \pm 0.2)\,\text{V}$$

Note that the measurement and the uncertainty are provided together in the brackets, indicating that the unit applies to both the measurement and its uncertainty.

Repeatability

When you take repeat measurements, the best estimate of the measured quantity is the average value. If you have taken fewer than 10 measurements, then the best estimate of the uncertainty is half the range. If you have more than 10 measurements, then the best estimate of the uncertainty is the standard deviation, usually denoted by the symbol sigma (σ). The calculation of standard deviation is given by:

$$\sigma = \sqrt{\frac{\sum_{i=1}^{n}(x_i - x)^2}{n-1}}$$

where x_i is an individual value of the measured value, x is the average value of the measured value and n is the total number of measurements. The sum is over all values of x_i (from $i = 1$ to n). Most calculators have built-in statistical functions and usually include standard deviation.

Spreadsheet software (e.g. Microsoft Excel) also provides the means to calculate functions such as standard deviation. Remember that 'repeat measurements' means repeating under the *same* conditions, which is not the same as collecting data points under different conditions.

10.2.2 What is the difference between repeatability and reproducibility?

SECTION REVIEW 10.2

REMEMBERING

1 Define:

a accuracy **b** precision **c** uncertainty **d** repeatability.

UNDERSTANDING

2 How do scientists estimate the uncertainty spread in a set of measurements?

3 Distinguish between uncertainty and error.

APPLYING

4 Calculate the standard deviation in the following set of measurements:

Experiment number	1	2	3	4	5	6	7	8	9	10
Mass of sample (g)	1.09	1.01	1.10	1.14	1.16	1.11	1.04	1.13	1.17	1.08

10.3 Qualitative and quantitative data

Many measurements refer to quantities that cannot be specified using a numeric scale. The colour of a mineral sample can often aid in its identification, but it cannot be quoted as a number that is useful for any form of mathematical processing. A solution of potassium permanganate ($KMnO_4$) is purple; one of copper(II) sulfate ($CuSO_4$) is blue; sodium chloride (NaCl) is colourless. These are qualitative observations of colour. Colour is usually considered as an example of **qualitative data**.

qualitative data/measurement quantities that are described by descriptive values and cannot be specified using a numeric scale

In contrast, if scientists observe the colour of light emitted by a hydrogen discharge lamp (Figure 10.3.1), while they can qualitatively observe that it is pinkish-purple, they can also disperse the light

10.3.1 Qualitative and quantitative data

FIGURE 10.3.1 Hydrogen discharge lamp emitting a pinkish-purple glow

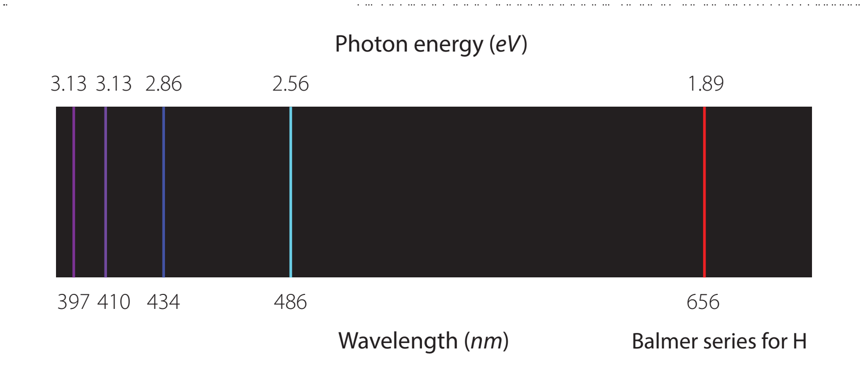


FIGURE 10.3.2 The spectrum obtained from a hydrogen discharge lamp by dispersing the light through a spectrograph

through a prism or with a diffraction grating to reveal that there are discrete wavelengths of light being emitted in the red, blue-green, and blue-to-violet regions of the visible spectrum. Scientists can measure the wavelengths of the emitted radiation using a spectrometer (Figure 10.3.2). The wavelengths are measured at 656 nm (red), 486 nm (blue-green), 434 nm (deep blue), 410 nm (indigo) and 397 nm (violet).

The wavelengths of the hydrogen atom emission spectrum are **quantitative data**. They can be used in a quantitative way; that is, by manipulating relationships between the numbers to determine the energies for the electronic orbitals in the hydrogen atom.

quantitative data/measurement numerical data acquired through counting or measuring

Qualitative data or measurements are descriptions such as colour, formation of a precipitate or the presence of bubbles. For example, a chemical reaction may lead to a colour change. You would usually describe the colour in words, such as 'pink' or 'green', rather than using a number. Shapes (spherical, cubic, tetragonal and octahedral) are also examples of qualitative data. However, the measured volumes for objects conforming to those shapes are quantitative data. The shapes can be used to classify (or sort) the quantitative data into separate classifications. Qualitative data is valuable, and can provide important measures, such as in surveys of gender, nationality and buying preferences. Qualitative data are often combined with quantitative measurements to permit statistical analysis of the qualitative information.

When variables have a numerical value, you make quantitative measurements. You measure that numerical value in the appropriate units. For example, you may measure temperature in degrees Celsius, mass in grams or volume in litres.

continuous variables quantitative data for which there are an infinite number of possible values within a selected range

Continuous variables may take any possible value, usually within some range. Length, time, mass, temperature and volume are examples of continuous variables used in chemistry. A variable that may take only fixed values is called a **discrete variable**. Often these are whole numbers of things that cannot be broken into smaller parts, such as electrons, protons or atomic number.

discrete variables quantitative or qualitative information that can be categorised into a classification, or is based on counts, with a finite number of values possible; the values cannot be subdivided meaningfully

Quantitative data provide numerical measures of quantities which can be manipulated using mathematical operations. They lend themselves to being graphed readily. Graphs and functional analyses can provide the means to generate mathematical relationships and make quantitative predictions from trends that might be observed.

SECTION REVIEW 10.3

REMEMBERING

1 Define:

a quantitative data

b qualitative data

c continuous variable

d discrete variable.

UNDERSTANDING

2 What advantages are provided by quantitative data compared to qualitative data?

10.4 Random and systematic errors

random error
a variation that affects a measurement in a random way so that the measurement is as likely to change in any one direction as in any other

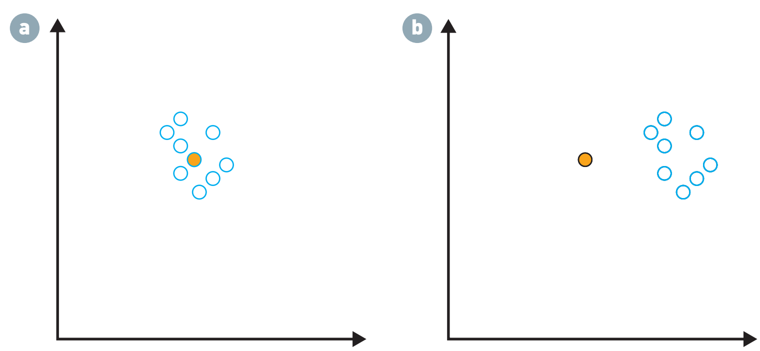
systematic error
an error that acts to give a consistent offset in data; for example, a zero error

10.4.1 The difference between systematic and random knowledge

Sources of uncertainty all give rise to **random errors**. This means that repeated measurements will be randomly spread around the 'true value', and centred on that value.

You may also have **systematic errors** in your data. These typically occur when there is a calibration error, such as a zero error, in a measuring device. Equipment should always be checked to ensure that a zero reading is obtained when it is expected. For example, in distilled water, a pH meter should read 7.0 (at 25°C). An electronic balance should be tared to zero before weighing the mass of a sample. If the measuring device is not calibrated correctly, then readings will be faulty.

FIGURE 10.4.1
a Results are clustered about the true value when the errors are random. **b** Results are clustered about some other value as a result of systematic errors.



Uncertainty and human error

The phrase 'human error' is not used by professional scientists to describe uncertainty in measurement. Human error is a term that describes a mistake, such as if a doctor gives a patient the wrong injection or amputates the wrong limb. If you make an error (mistake) in a scientific measurement, then you need to repeat the measurement.

A scientist taking measurements contributes to uncertainty through the measurement procedure, but the error must be reported as an uncertainty caused by an identified source. Always define the source of the uncertainty (e.g. reaction time, parallax, estimation between minimum markings or limit of reading) and make sure that the magnitude of any such error is quantified. The use of the term 'human error' in reference to measurement uncertainty has no place in a scientific report.

SECTION REVIEW 10.4

REMEMBERING

1 Define:

a random error **b** systematic error.

UNDERSTANDING

2 Why is it inappropriate to classify experimental uncertainties as occurring due to 'human error'?

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10.5 Significant figures

After all the sources of uncertainties have been identified and their magnitudes estimated, it is necessary to quote results to the correct number of **significant figures**.

Consider the numbers 0.00123, 1230 and 2345.678. How many significant figures should be reported in each case? The answer to this question is based on the following approach.

significant figures
all the nonzero digits of a number and the zeros that are included between them or that are final zeros and signify accuracy

1. In a decimal number, do not count the 'place holders' (the zeros after a decimal point and before the first non-zero integer), so for 0.00123, the two zeros after the decimal point are not significant. There are three significant figures (123).
2. The number 1230 can also be expressed as 1.230×10^3. In either case the trailing zero is significant. Its existence implies a meaning of 1230; that is, somewhere between 1229.5 and 1230.4 (rounding up or down would both give 1230). In the form 1.230×10^3, the third digit after the decimal is zero, not 1, 2, 3 or 4. By writing 1.23×10^3, you would not know whether the third digit after the decimal was 0, 1, 2, 3, or 4. If only two digits were included after the decimal point, you would need to round down to 1.23×10^3. This means that 1.2×10^3 has two significant figures, 1.23×10^3 has three significant figures, and 1.23456×10^n (for any power n) has six significant figures.
3. Quoting the correct number of significant figures is important. The number of significant figures must be consistent with the uncertainties in the measurement. In other words, can the number of significant figures quoted be believed? For example, if scientists reported that the distance between Brisbane and Cairns was 1691.4635 km, would this be believable? The implied precision of the quoted distance is given by the last digit (5). This means quoting the distance as somewhere between 1691.46354 km and 1691.46345 km since rounding the last digit in either of those numbers would give us 1691.4635. This is a quoted precision of ±0.00005 km, or 5 mm! The distance from Brisbane to Cairns depends on from where in Brisbane you leave and where you arrive in Cairns, an uncertainty of ±1 km or more. It is sufficient to state, for example, that the distance from Brisbane to Cairns is 1690 ± 10 km. For most practical purposes, quoting 1700 km (just two significant figures) would probably be enough.

Scientific notation is typically a good way to provide measurements because the number of significant figures determined by measurement uncertainties can be seen easily. For example, 2.45×10^{-4} is usually a preferable way to express the measurement 0.000245.

The answer as to how many significant figures should be displayed when quoting a measurement comes from estimating the uncertainty and estimating the consequences of any error propagation through calculations (section 10.6).

10.5.1 Introduction to significant figures

10.5.2 Significant figures

SECTION REVIEW 10.5

REMEMBERING

1. Define:
   a. significant figures
   b. place holder in a decimal number.

UNDERSTANDING

2. How many significant figures should be displayed when quoting a measurement?

10.6 Absolute uncertainty, absolute error, percentage uncertainty and percentage error

absolute error
the magnitude of the difference between the observed/measured value and the true value

absolute uncertainty
the spread or interval of measured values within which the true value is expected to lie

percentage uncertainty
percentage uncertainty

$$= \frac{\text{absolute uncertainty of measurement}}{\text{measurement value}} \times \frac{100}{1}$$

The **absolute error** is the estimate that you have made for the error, and the **absolute uncertainty** is your estimate of uncertainty, expressed as numbers (with units).

For example, if the uncertainty in measuring 60 mL of a liquid based on an estimate of measurement precision is ±1 mL and the accuracy of the measuring cylinder is 0.5 mL, then the absolute uncertainty is ±1.5 mL.

The **percentage uncertainty** is a percentage calculated relative to the measured quantity:

$$\text{Percentage uncertainty} = \frac{\text{absolute uncertainty}}{\text{measured quantity}} \times \frac{100}{1}\%$$

So, for the example above, the percentage uncertainty is given by:

$$\text{Percentage uncertainty} = \frac{\pm 1.5\ \text{mL}}{60\ \text{mL}} \times \frac{100}{1}\% = \pm 2.5\%$$

Note that the measured quantity, the uncertainty, and the percentage uncertainty are all quoted to the same number of significant figures (2).

If the measured quantity were 10 mL instead of 60 mL, then the absolute uncertainty would still be ±1.5 mL, but the percentage uncertainty would now be 15% – a much larger uncertainty relative to the measured quantity:

$$\text{Percentage uncertainty} = \frac{\pm 1.5\ \text{mL}}{10\ \text{mL}} \times \frac{100}{1}\% = \pm 15\%$$

If the measured quantity is being compared with an accepted value (a measurement made by a very high-precision instrument, or a value that has been accepted as a standard of measurement) then it is appropriate to state a percentage error relative to the accepted value.

Suppose you conduct an experiment to measure the enthalpy of combustion for ethanol. In your experiment, you obtain a measurement of $\Delta H = -1360 \pm 12\ \text{kJ mol}^{-1}$. The value for ΔH for ethanol listed by the Material Measurement Laboratory of NIST is $-1366.3 \pm 0.4\ \text{kJ mol}^{-1}$. You would state that your measurement is *consistent with the accepted value* because your estimate of uncertainties indicates that the accepted value lies within your stated uncertainty range of -1348 to $-1372\ \text{kJ mol}^{-1}$.

The percentage uncertainty in your measurement is given by:

$$\text{Percentage uncertainty} = \frac{\pm 12\ \text{kJ mol}^{-1}}{1360\ \text{kJ mol}^{-1}} \times \frac{100}{1}\% = \pm 0.88\%$$

percentage error
percentage error

$$= \frac{\text{accepted value} - \text{measured value}}{\text{accepted value}} \times \frac{100}{1}$$

Your uncertainty and percentage uncertainty both contain 2 significant figures, so the statement that your uncertainty is ±0.88% is acceptable. However, the *mean* of your measurement displays a **percentage error** relative to the accepted value, given by:

$$\begin{aligned}\text{Percentage error} &= \frac{\text{accepted value} - \text{measured value}}{\text{accepted value}} \times \frac{100}{1}\% \\ &= \frac{-1366.3\ \text{kJ mol}^{-1} - 1360\ \text{kJ mol}^{-1}}{1366.3\ \text{kJ mol}^{-1}} \times \frac{100}{1}\% \\ &= \pm 0.46\%\end{aligned}$$

Again, uncertainty and percentage error both contain 2 significant figures, so the statement that your error is ±0.46% is acceptable if the uncertainty in your measurements was less than ±0.46%.

However, given that your measurement uncertainty is ±0.88%, which is greater than the estimate for the percentage error for the mean of your measurement (1360 kJ mol^{-1}), it is necessary to state that your percentage error is *likely to be in a range*:

$$\text{Percentage error}(l) = \frac{-1366.3\ \text{kJ mol}^{-1} - 1348\ \text{kJ mol}^{-1}}{1366.3\ \text{kJ mol}^{-1}} \times \frac{100}{1}\% = -1.34\%$$

$$\text{Percentage error}(h) = \frac{-1366.3\ \text{kJ mol}^{-1} - 1372\ \text{kJ mol}^{-1}}{1366.3\ \text{kJ mol}^{-1}} \times \frac{100}{1}\% = +0.42\%$$

Here (l) and (h) indicate that the error calculation is for the lower and higher values in your range for the uncertainty in your measured value for ΔH for ethanol. Notice that the percentage error can be asymmetric about the mean.

The final statement would be that your value of $\Delta H = -1360 \pm 12$ kJ mol^{-1} for ethanol displays a percentage error relative to the accepted value of -1366.3 ± 0.4 kJ mol^{-1} in the range: (−1.34% to 0.42%).

Error propagation

Whenever two or more measurements are combined in a calculation, the uncertainties in each of those measurements will combine to produce uncertainty in the calculated result.

A simple treatment of **error propagation** is sufficient for most purposes. When adding or subtracting, give the final answer to the least number of decimal places. When multiplying or dividing or taking powers, give the final answer to the least number of significant figures.

error propagation uncertainty produced as a result of combining measurements, each with uncertainties, in a calculation

SECTION REVIEW 10.6

REMEMBERING

1 Define:

a absolute error

b absolute uncertainty

c percentage error

d percentage uncertainty.

UNDERSTANDING

2 How many significant figures should be displayed when quoting the results of a calculation in the following?

a When adding or subtracting measurements

b When multiplying or dividing or taking powers of measurements

10.7 Presenting and analysing data

The outcome of a quantitative or qualitative scientific investigation is usually the collection of a set of experimental measurements or observations known as 'raw data'. The presentation and analysis of data is the task that yields usable and useful information. The analysis, regardless of whether the data is qualitative or quantitative, will:

- describe and summarise the data
- identify relationships between variables
- compare trends for variables
- identify differences between behaviour of variables
- interpret outcomes
- forecast implications arising out of the analysis.

Organising data

The first step is to organise your data, usually by tabulating it. If you have more than a few data points, then it is advisable to display them in a table.

Record your data in a logical manner. Clearly label each set of data so that they are identified unambiguously when you need to analyse them. If you have multiple experiments, a separate table for each may be advisable. For example, if you are investigating the effect of concentration on the rate of a reaction for a set of different temperatures, then a separate table for each temperature that you tested makes good sense. Each table (assigned to a particular temperature) would contain the measured rates of reaction for each of the different concentrations.

After the data are tabulated, they can be displayed as a graph. Graphs are a powerful technique for identifying trends and relationships because they visually represent how dependent variables change as a function of independent variables. There are many different types of graphs that can be used to organise and display data.

Calculations based on the measured data can assist in answering a question posed in the experimental design or can address any hypothesis proposed. Always assign units to all quantities, so that any derived values have the correct units. Uncertainties must be quoted for any derived quantities.

Identifying trends, patterns and relationships

You may be able to see a pattern simply by looking at a list of numbers in a table. However, the most reliable way to identify a pattern in data or a relationship between variables is to examine a graph. If you have derived a mathematical equation describing the relationship between the dependent and independent variables based on a theoretical hypothesis, then use it to generate a fit on a graph of your data. Do not substitute your data into your hypothesised equation and try to show that it fits.

A graph should be large and clear. The axes should be labelled with the names of the variables and their units. Choose a scale so that your data take up most of the plot area. This will often mean that the origin is not shown in your graph. Usually, there is no reason why it should be. Furthermore, if the point (0,0) is not a measured data point, it must not be included in any theoretical fitting of the data.

When you are seeking a relationship between variables, plot a **scatter graph** (Figure 10.7.1b). This shows your data as points. *Do not join them up as in a dot-to-dot picture.* Generally, the independent

scatter graph
a graph that uses horizontal (x) and vertical (y) axes to plot a set of data points represented by (x, y), without a line joining the points; used to demonstrate or determine a mathematical relationship between variables

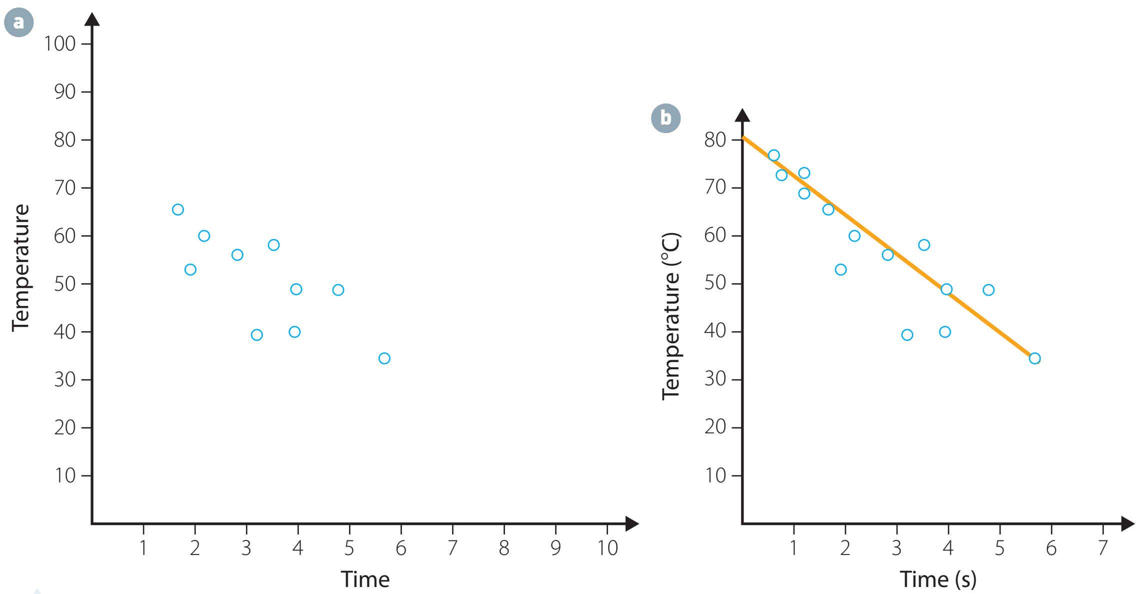

FIGURE 10.7.1 Examples of **a** a poor graph and **b** a good graph. How many differences can you see?

variable is plotted on the x-axis and the dependent variable is plotted on the y-axis, unless there is a good reason to do otherwise. If reaction rates are being measured by taking readings of concentration (C) as a function of time (t), the time t would be the independent variable, plotted on the x-axis and concentration, C, the dependent variable, would be plotted on the y-axis.

Number of data points

To determine a relationship, you need to have enough data points and the range of your data points should be as large as possible. A minimum of six data points is generally considered adequate if the relationship is expected to be linear, but always collect as many as you reasonably can, given the available time.

For non-linear relationships, you need more data points than this. Try to collect more data in regions where you expect rapid variation. For example, if you are measuring the pressure of a fixed mass of gas at different volumes (Figure 10.7.2), then you should expect an inverse relationship $\left(P \propto \frac{1}{V}\right)$ and a hyperbola graph. If the number of data points are limited, it is easy to mistakenly deduce the relationship as linear.

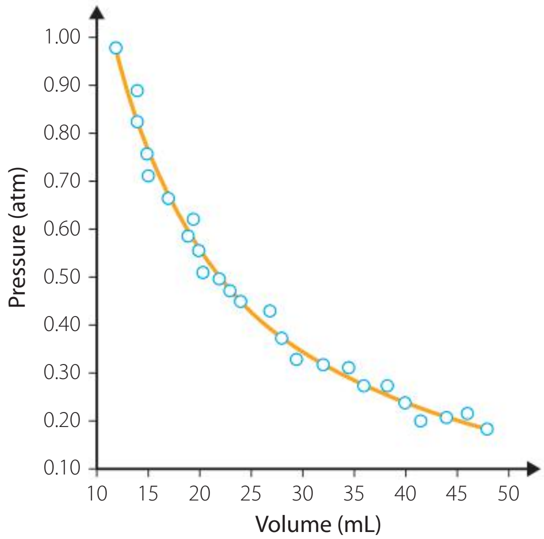

FIGURE 10.7.2 A pressure–volume graph for a gas needs many data points to show the hyperbola shape.

WORKED EXAMPLE 10.7.1

Nitrogen dioxide (NO_2) is a combustion product present in automobile exhausts. It is one of the nitrogen oxides known as NO_x and is partly responsible for the brown colour of city smog, as shown in Figure 10.7.3.

FIGURE 10.7.3 Photochemical smog over the city of Shanghai, China. The brown colour is contributed to by nitrogen dioxide (NO_2), which absorbs light broadly over the blue-green region of the visible spectrum.

During daylight hours, the NO_2 molecules absorb solar radiation in the 400–550 nm (blue-green) region of the spectrum and dissociate to nitrogen oxide (NO) and oxygen (O) atoms. The O atoms can react readily with O_2 molecules to produce ozone (O_3). Ozone is an important component of the stratosphere (atmosphere above approximately 12–50 km from Earth's surface) because it protects us from solar ultraviolet radiation that is harmful to biological systems. However, in the troposphere (the lower layer of the atmosphere up to approximately 10–12 km), ozone is a harmful component of what is called 'photochemical smog' because ozone is a powerful oxidising agent and can irritate the respiratory tract.

Laboratory studies of NO_2 dissociation have helped us to understand the complex chemistry in polluted atmospheres.

QUESTION

How can you use laboratory data on the dissociation of nitrogen dioxide to determine a mathematical relationship for how its concentration varies with time and to present the data scientifically?

ANSWER

1 Tabulate your data. Table 10.7.1 presents data for the dissociation of NO_2 at 300 K. The concentration of NO_2 is measured as a function of elapsed time after a short pulse of 420 nm light has initiated the reaction.

TABLE 10.7.1 Concentration of nitrogen dioxide (NO_2) measured as a function of time following the dissociation of NO_2 at 300 K

TIME (t)	CONCENTRATION OF NO_2 (mol L^{-1}) (± 3%)
0	0.0096
100	0.0064
200	0.0048
300	0.0038
400	0.0031
500	0.0027
600	0.0024
700	0.0021

2 Estimate uncertainties. Note the number of significant figures contained in the measurements for each of the measured variables. The time is quoted to 3 significant figures. Concentrations are quoted to 2 significant figures and represent the mean of a number of experimental repetitions. Uncertainties can be quoted as standard deviations or percentages; uncertainties in concentrations are estimated as ± 3%.

3 Identify the relationship between variables. The independent variable is time (t) measured in units of seconds. The dependent variable is the concentration of NO_2 in the usual units (mol L^{-1}). Inspecting the data table indicates that, as the time increases from $t = 0$ to $t = 700$, the NO_2 concentration decreases. However, the table does not reveal too much about the shape or functional form of the dependence of NO_2 on time.

You can graph the raw data readily using software such as Microsoft Excel (scatter plot) as shown in Figure 10.7.4. The error bars on the graph (which estimate the range of uncertainty for each data point) have been set at ± 3% for the vertical axis (concentration) in accordance with the uncertainties stated in the table. The uncertainty in time measurements based on the number of significant figures would not be visible on the graph, so it is omitted.

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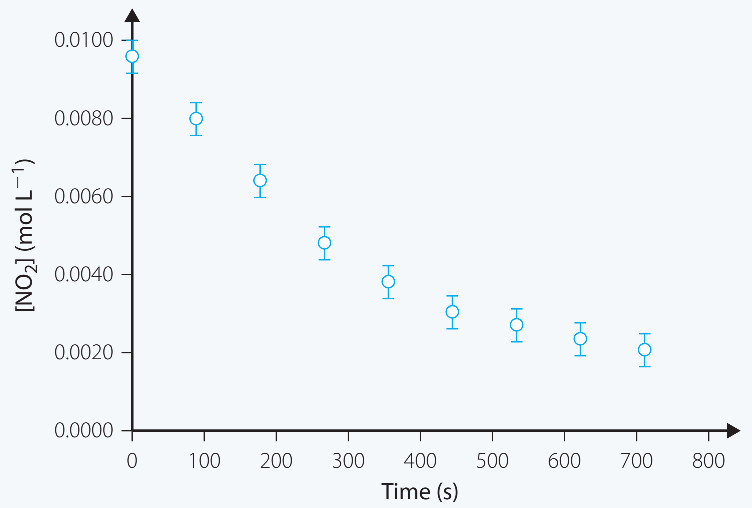

FIGURE 10.7.4 Concentration versus time graph for the dissociation of nitrogen dioxide (NO_2)

The graph shows clearly that the dependence of nitrogen dioxide concentration on time is not linear (i.e. not a straight line).

4 Construct a model. To assist in analysing the data, construct a mathematical model. Any model usually starts with a theoretical framework. Here, we begin with a balanced chemical reaction for the dissociation of nitrogen dioxide (NO_2) where the products are the nitrogen monoxide (NO) molecule and O atoms:

$$NO_2 \rightarrow NO + O$$

We will explore two possible mathematical relationships that might help to establish the relationship between nitrogen dioxide concentration and time.

5 Test the models. We will use a log relationship, using the natural logarithm of nitrogen dioxide concentration ($\ln[NO_2]$) and an inverse relationship ($1/[NO_2]$). The calculations are shown in Table 10.7.2 and plotted in Figure 10.7.5.

TABLE 10.7.2 Concentration of nitrogen dioxide ($[NO_2]$) measured as a function of time following its dissociation at 300 K as well as values for $\ln[NO_2]$ and $1/[NO_2]$ derived from the measured values for $[NO_2]$

TIME (s)	$[NO_2]$ (mol L^{-1})	$\ln[NO_2]$	$1/[NO_2]$
0	0.00962	–4.64	104
50	0.00800	–4.83	125
100	0.00641	–5.05	156
200	0.00481	–5.34	207
300	0.00382	–5.57	262
400	0.00305	–5.79	328
500	0.00271	–5.91	369
600	0.00236	–6.05	424
700	0.00207	–6.18	483

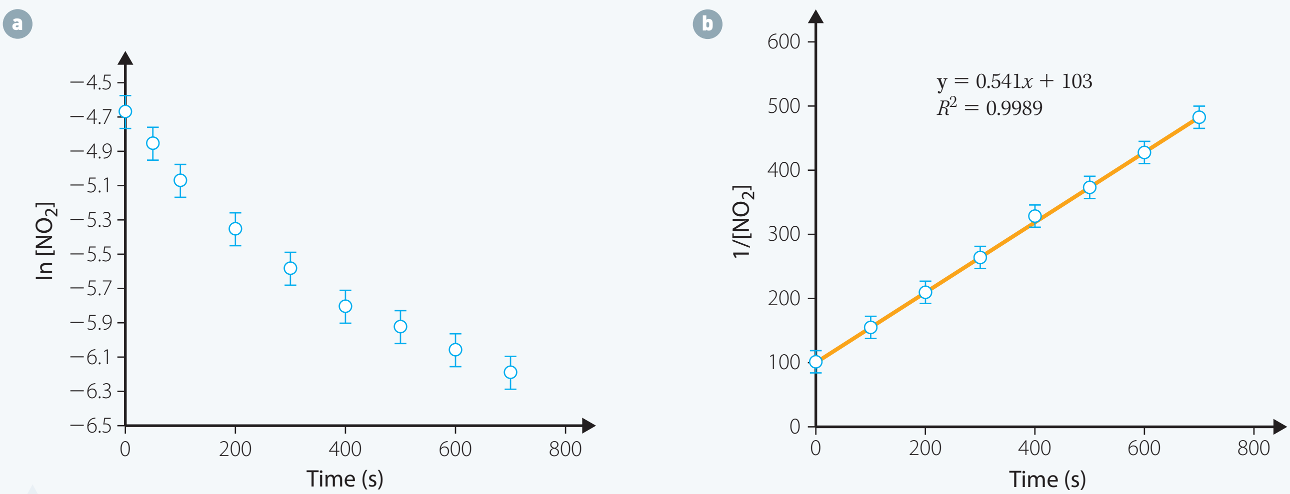

FIGURE 10.7.5 **a** Graph of ln[NO_2] versus time for the dissociation of nitrogen dioxide (NO_2). The dependence is non-linear. **b** Graph of 1/[NO_2] versus time for the dissociation of NO_2. A linear relationship is observed.

When we graph ln[NO_2] versus time (Figure 10.7.5a), the values for ln[NO_2] fall off with time, as did the raw data for nitrogen dioxide concentration. However, the fall-off remains non-linear, so it is not possible to conclude that there is any justifiable linear relationship between ln[NO_2] and the time elapsed. In other words, the rate of dissociation cannot be directly proportional to ln[NO_2].

In contrast, the graph of 1/[NO_2] versus time shown in Figure 10.7.5b displays a linear relationship. Notice that the calculated values for 1/[NO_2] increase with time. This is expected because we have derived an inverse relationship based on the raw data. This linear relationship means that we can conclude that the rate of dissociation is directly proportional to 1/[NO_2]; that is, an inverse or hyperbolic relationship:

$$\left(\frac{\Delta[NO_2]}{\Delta t}\right)$$

Mathematically, our conclusion can be stated as follows:

$$\text{Rate of dissociation of } NO_2 = \frac{\Delta[NO_2]}{\Delta t} \propto \frac{1}{[NO_2]}$$

Hence:

$$\text{Rate of dissociation of } NO_2 = \frac{\Delta[NO_2]}{\Delta t} = k\left(\frac{1}{[NO_2]}\right)$$

where k is the proportionality constant.

6 Determine the proportionality constant. The value of k is obtained from the slope of the graph as shown in Figure 10.7.5b:

$$\left(\text{Gradient} = \frac{\Delta y}{\Delta x}\right)$$

Find the gradient using the 'add trendline' command in Microsoft Excel to add a line of best fit to the data. Choosing the linear trendline option gives a line of best fit of the form:

$$y = mx + c$$

This is the value for k, the constant of proportionality. The y-intercept is also obtained from the linear fit.

CONCLUSION AND DATA PRESENTATION

1 Draw a conclusion. Having obtained the parameters m and c for the straight line fit to the graph of $1/[NO_2]$ versus time, a mathematical expression for the rate of dissociation of $[NO_2]$ has been derived:

$$\text{Rate of dissociation of } NO_2 = 0.54\left(\frac{1}{[NO_2]}\right) + 102$$

2 Present your data. Note that the value for k and for the intercept has been rounded to 2 significant figures because the raw data for $[NO_2]$ were also presented with 3 significant figures. The y-intercept of 102 represents the value for $1/[NO_2]$ at time $t = 0$ estimated from our line of best fit. This can be compared with the first (calculated) data point for $1/[NO_2]$, namely 104, given in Table 10.7.2. The difference between them is 2, which means they are in agreement to within 2% of each other. This agreement is within the 3% uncertainty in the raw data points.

The overall analysis is consistent with the level of uncertainty in the original measurements.

This worked example provides an investigation of the time dependence of NO_2 concentration in a chemical reaction where NO_2 dissociates due to light absorption. It is an example of how to establish a possible mechanism for a chemical reaction using mathematical relationships. This kind of analysis is commonly employed in the study of the rates of chemical reactions, a topic introduced in Chapter 21.

Outliers

When you plot your raw data, you may find that one or two points are data **outliers**. These are points that do not fit the pattern of the rest of the data. These points may be mistakes; for example, they may have been incorrectly recorded or a mistake was made during measurement. They may also be telling you something important. For example, if they occur at extreme values of the independent variable, then it might be that the behaviour of the system is linear in a certain range and then changes. You may choose to ignore outliers when fitting a line to your data, but you should be able to justify why. In the graphs for nitrogen dioxide dissociation in Table 10.7.3, all the data points seem to lie on a smooth curve, so they may all be retained as legitimate data.

outlier
a data point that does not fit the pattern shown by other measured data points; sometimes they are defined quantitatively as lying more than 1.5 times the interquartile range above the third quartile or below the first quartile

interpolation
to read or construct a new data point that has not been measured but is within the range of measured data

extrapolation
to estimate a value for a data point that lies outside the range of measured values using a mathematical relationship based on the measured data, or extending a perceived sequence of values, and assuming that the trend identified will continue

Interpolation and extrapolation

Reading points, other than data points, from a line of best fit within the region in which you have data is called **interpolation**. You cannot be sure that this is exactly what you would find if you measured that point. However, if your line of best fit truly represents the behaviour of the system (and the derived parameters are within the limit of uncertainties in the raw data, as in the case of the y-intercept for the $1/[NO_2]$ graph in Worked example 10.7.1), then using interpolated points in the analysis is justifiable.

Extrapolation is used to predict a value of a dependent variable for a chosen value of the independent variable that lies outside the range of measured values for the data. For example, we could use the linear relationship established in Figure 10.7.5b to predict a value for the NO_2 concentration at times longer than 800 seconds.

Stating conclusions based on data analyses

If the hypothesis is not supported, it is not enough to simply say, 'our hypothesis is wrong'. Rather, it is important to identify what is wrong with it.

Go through your method, results and analysis. Check that your equipment was correctly calibrated, and that you were using it correctly. Check that data are recorded in the correct units, and that units are carried correctly through all calculations during analysis. Check your analysis carefully. If you are working in a group, ask another person to repeat the calculations.

Be rigorous with calculating uncertainties. If the uncertainties are ±0.1%, then any parameters derived from mathematical correlations need to be consistent with this level of uncertainty in the data. If the uncertainties are much greater, such as ±15%, then any derived parameters, for say a linear fit to data, that lie within this level of uncertainty, can still support the hypothesis of a linear fit even if the data points are much more scattered on the graph.

Consider what other factors may have affected your results. Were there variables that you were not able to control? Were there variables that you forgot to control?

It is not sufficient to conclude that 'the experiment did not work'. Either a mistake was made or the theoretical or mathematical model you have proposed is inappropriate.

WORKED EXAMPLE 10.7.2

The enthalpy change (ΔH) associated with the combustion of different linear hydrocarbons (alkanes with the general formula C_nH_{2n+2}) has been measured to a very high degree of precision and accuracy by researchers at the National Bureau of Standards (NBS) in the United States. (Enthalpy change (ΔH) for combustion reactions is covered in Chapters 9 and 11).

QUESTION

Is the heat of combustion of alkanes linearly proportional to the number of carbon atoms in an alkane?

ANSWER

1 Examine the data. Table 10.7.3 displays some of the published data. The units for ΔH given here are kJ mol^{-1}. In Chapter 9, enthalpy changes were quoted in units of kJ per gram (kJ g^{-1}). Chapter 12 introduces the common chemical unit of a mole.

The values for ΔH are all negative, so the reactions (yielding carbon dioxide (CO_2) and H_2O as products) are exothermic.

TABLE 10.7.3 Enthalpy change (ΔH) associated with the combustion of linear hydrocarbons (alkanes)

ALKANE NAME	FORMULA (C_nH_{2n+2})	n	ΔH (kJ mol^{-1})
Methane	CH_4	1	–212.80
Ethane	C_2H_6	2	–372.82
Propane	C_3H_8	3	–530.61
Butane	C_4H_{10}	4	–687.98
Pentane	C_5H_{12}	5	–845.16
Hexane	C_6H_{14}	6	–1002.57
Heptane	C_7H_{16}	7	–1160.01
Octane	C_8H_{18}	8	–1317.45
Decane	$C_{10}H_{22}$	10	–1632.34
Dodecane	$C_{12}H_{26}$	12	–1947.23
Hexadecane	$C_{16}H_{34}$	16	–2577.00

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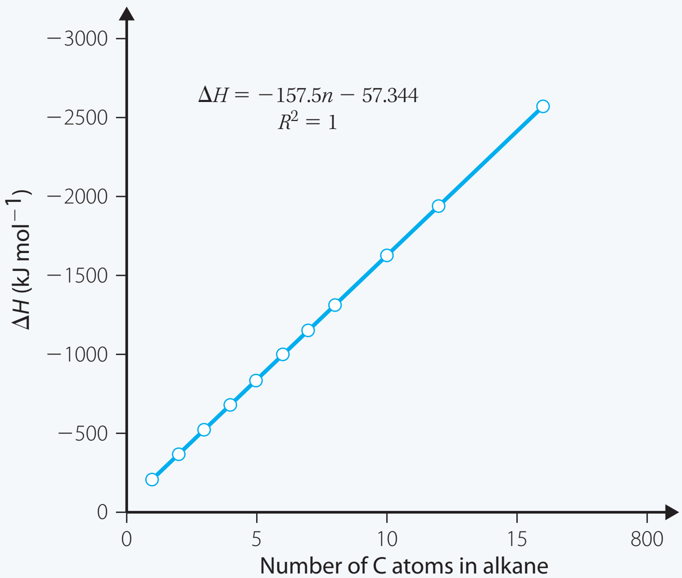
2 What is your hypothesis? Scientists look for a correlation between the measured heats of combustion and one or more molecular properties of the hydrocarbons.

A first approach might be to plot the measured value of ΔH as a function of the number carbon atoms in the alkane (n). A chemical justification for this is that, as the number of carbons in an alkane increases, there is a linear increase in the number of C—C bonds that need to be broken in the combustion reaction. Accordingly, the enthalpy of combustion should increase. This hypothesis turns out to be quite fortuitous.

Figure 10.7.6 graphs the measured values for ΔH versus the alkane carbon number (n) together with a linear trendline fit to the data. The correlation is essentially perfect.

The hypothesis that the heat of combustion of alkanes is linearly proportional to the number of carbon atoms in the alkane is upheld with a high degree of certainty.



Data sourced from: U. S. DEPARTMENT OF COMMERCE NATIONAL BUREAU OF STANDARDS RESEARCH PAPER RP1642, Part of Journal of Research of the National Bureau of Standards, Volume 34, March 1945. HEATS OF COMBUSTION AND FORMATION OF THE PARAFFIN HYDROCARBONS AT 25 C. By Edward J. Prosen and Frederick D. Rossini

FIGURE 10.7.6 Measured enthalpies of combustion (ΔH) for alkanes plotted versus the number (n) of carbon atoms in the alkane

CONCLUSIONS AND PREDICTIONS

1 Formulate a conclusion. Derive derived a useful relationship between ΔH and the number of carbon atoms (n) in the alkane chain from a database of precise measurements of ΔH for alkanes, together with a mathematical correlation.

2 Make a prediction. This correlation can be used to predict what the value of ΔH might be for *any* value of n. The relationship derived from the linear fit to the experimental data is:

$$\Delta H = -157.5n - 57.344\,\text{kJ mol}^{-1}$$

For example, an interpolated data point can be obtained for the alkane $C_{15}H_{32}$, known as pentadecane, by substituting $n = 15$ in this relationship. A prediction is that:

$$\Delta H = -157.5(15) - 57.344 = -2419.84\,\text{kJ mol}^{-1}$$

The value published in the NBS tables of data is $-2419.55\,\text{kJ mol}^{-1}$. The difference is a mere 0.01%. The conclusion is that the derived correlation may be used with confidence for interpolation.

Similarly, extrapolate from the correlation to predict values for ΔH for higher n alkanes. For example, for the alkane heptadecane ($C_{17}H_{36}$), $n = 17$, a prediction is that:

$$\Delta H = -157.5(17) - 57.344 = -2734.84\,\text{kJ mol}^{-1}$$

Once again, the extrapolated value is consistent to within 0.01% of the published value, $-2734.44\,\text{kJ mol}^{-1}$.

Worked example 10.7.3 demonstrates an application of statistical analysis and calculation of the standard deviation (σ) for a set of measurements to reach a rigorous conclusion.

WORKED EXAMPLE 10.7.3

Worked example 10.7.1 analysed the dissociation of nitrogen dioxide (NO_2). It demonstrated a mathematical relationship – an appropriate representation of the data was that the rate of the reaction depended on the inverse of the concentration ($1/[NO_2]$).

This Worked example examines measurements of nitrogen dioxide concentrations obtained in an air quality monitoring experiment to determine whether nitrogen dioxide levels in the Sydney central business district (CBD) comply with environmental standards. Table 10.7.4 provides measurements for the concentration of nitrogen dioxide recorded for the first 60 days of 2016. Uncertainties are shown in the table footnote and the units are explained. The air quality target set by the EPA is that average nitrogen dioxide levels should not exceed 80 $\mu g\,m^{-3}$ over a two month period.

QUESTION

Based on the measurements provided, is the EPA target, that average NO_2 levels should not exceed 80 $\mu g\,m^{-3}$ over a two-month period, being met?

ANSWER

1 Examine and plot the data. Graph the data to display how nitrogen dioxide concentrations vary over the 60-day period. Figure 10.7.7 (page 197) graphs measured nitrogen dioxide concentrations (dependent variable) versus the day measured (independent variable). The data points include error bars of ±0.5.

TABLE 10.7.4 Concentrations ($\mu g\,m^{-3}$) of nitrogen dioxide (NO_2) measured in the Sydney CBD for the first 60 days of 2016

DAY	$[NO_2]$	DAY	$[NO_2]$	DAY	$[NO_2]$
1	38	21	30	41	30
2	43	22	62	42	72
3	40	23	49	43	60
4	38	24	32	44	41
5	66	25	38	45	69
6	62	26	57	46	36
7	53	27	32	47	66
8	66	28	45	48	79
9	57	29	80	49	51
10	36	30	78	50	41
11	64	31	49	51	56
12	47	32	45	52	78
13	40	33	39	53	30
14	64	34	59	54	74
15	79	35	71	55	40
16	36	36	35	56	77
17	40	37	32	57	40
18	58	38	67	58	60
19	31	39	69	59	76
20	56	40	71	60	30

Units for nitrogen dioxide concentration ($[NO_2]$) are micrograms (µg) per m^3 (1 microgram (µ) = 10^{-6} g)

Two significant figures are quoted for NO_2 concentrations

This is consistent with estimated uncertainties for $[NO_2]$ of ±0.5 µg.

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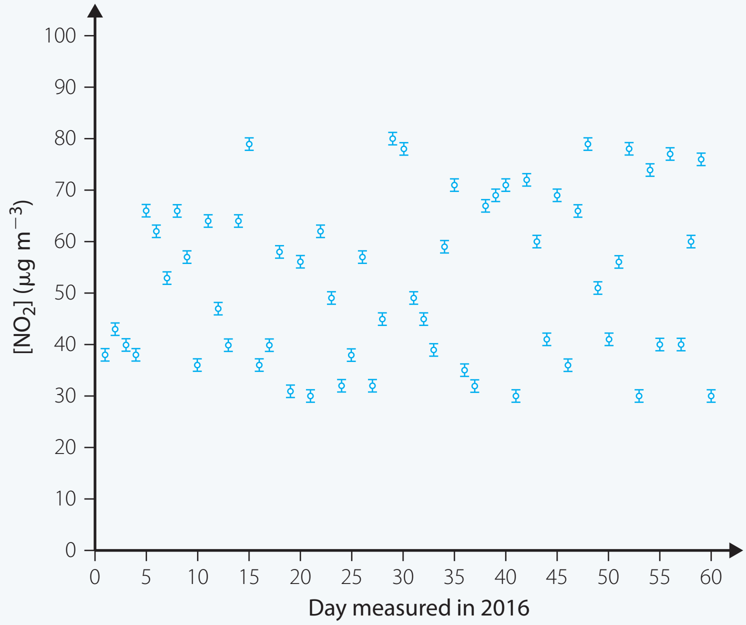

FIGURE 10.7.7 Measured nitrogen dioxide (NO_2) concentrations in Sydney CBD versus day of measurement

2 The measured concentrations are scattered widely. Generally, the bulk of measured concentrations lie below the EPA target of 80 µg m^{-3}. However, to reach a definitive scientific conclusion as to the meaning of the measurements, a statistical analysis is required. Using the formula for mean and standard deviation (section 10.2), the mean nitrogen dioxide concentration for the 60-day period can be calculated by entering Table 10.7.4's data into a spreadsheet and then carrying out the calculation. The result obtained is expressed as:

$$\text{Mean value for } NO_2 \text{ concentration} = 53\,\mu g\,m^{-3}$$

$$\sigma = \text{standard deviation} = 16\,\mu g\,m^{-3}$$

3 The standard deviation (σ) provides a measure of the likelihood that a certain percentage of measured data points will lie within 1 standard deviation; that is, within σ of the mean. This can be seen visually using a 'normal distribution' (Figure 10.7.8).

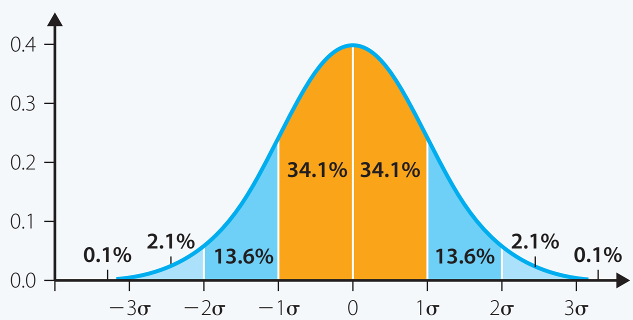

FIGURE 10.7.8 Plot of the normal distribution. Each band has a width of one standard deviation (σ).

CONCLUSION

Figure 10.7.8's probability distribution curve allows us to conclude that $2 \times (34.1)\% = 68.2\%$ of measurements will lie within 1 standard deviation (σ) of the mean; that is, in the range $53 \pm 16\,\mu g\,m^{-3}$.

It can also be stated that $2 \times (34.1 + 13.6)\% = 95.4\%$ of measurements will lie within 2σ of the mean; that is, in the range $53 \pm 32\,\mu g\,m^{-3}$. This means that the range has an upper limit of $85\,\mu g\,m^{-3}$, which is a little above the EPA target of $80\,\mu g\,m^{-3}$.

A slightly more detailed analysis of Figure 10.7.8, together with the use of calculated values for the normal distribution (readily available in most graphics calculators or in mathematics and statistics textbooks), allows us to deduce that there is a 94% probability that nitrogen dioxide concentrations will *not* exceed the EPA target of $85\,\mu g\,m^{-3}$ in a 60-day period. Stated in reverse, this means that there is a 6% probability that nitrogen concentrations *will* exceed the EPA target of $85\,\mu g\,m^{-3}$ in a 60-day period.

In summary, our report can state that there is a 6% probability that nitrogen concentrations will exceed the EPA target of $85\,\mu g\,m^{-3}$ in a 60-day period. This is a quantitative estimate based on a set of measurements. This conclusion is more useful than 'a few measurements of nitrogen concentrations were found to exceed the EPA target of $85\,\mu g\,m^{-3}$ in the 60-day period'.

Policy decisions on reducing nitrogen dioxide concentrations in the Sydney CBD are up to the appropriate administrative authorities but, as a scientist, you have provided clear, unambiguous and unbiased conclusions based on measurements that include an honest and rigorous determination of uncertainties.

SECTION REVIEW 10.7

REMEMBERING

1 Define:

- **a** scatter graph
- **b** outlier
- **c** interpolation
- **d** extrapolation.

UNDERSTANDING

2 To identify a *linear* relationship in a set of data points, what is the minimum number of data points required?

3 To identify a *non-linear* relationship in a set of data points (for example, an inverse relationship such as $\propto \frac{1}{V}$), what is considered an adequate minimum number of data points?

4 Identify the methods used to establish the uncertainties in estimating the mean of data points or the uncertainty in fitted trendlines, which might establish relationships between independent and dependent variables.

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 For each of the following, give the answer to the correct number of significant figures.

a 2.55 m × 3.266 m
b 5.23 g ÷ 35.45 g
c $(6.022 \times 10^{23}) \times 0.100$ g
d 4.3456 g + 3.3357 g – 2.2 g
e 8.0001 cm + 2.51 mm
f 17.648 s – 1.251 s

2 After heating a 5.00 g sample of potassium chlorate ($KClO_3$), a volume of oxygen is collected and the mass is calculated to be 1.85 g. Theoretically, 5.00 g of $KClO_3$ should yield 1.96 g oxygen. Determine the absolute error and the relative percentage error in this experiment.

3 An experimental determination of the boiling point of ethyl alcohol yields a result that is quoted by the experimenter as 76.5 ± 2°C. The accepted literature value for the boiling point of ethyl alcohol is 78.37°C.

a What is the percentage uncertainty for the experiment?
b Based on the quoted uncertainty, is the experimental result consistent or inconsistent with the accepted literature value?
c What is the percent error of the experimental value relative to the accepted literature value?

4 Consider the following results for experimental measurements of the same quantity (enthalpy change ΔH for a reaction, time interval t for an event, mass of a sample, m).

a Determine if they agree with the accepted result listed in the second column.
b Calculate the % uncertainty as stated for each measurement.
c If the stated uncertainty is such that the measurement does not agree with the accepted result, how would you classify the source of error?
d Calculate the percentage error for each measurement.

EXPERIMENTAL RESULT	ACCEPTED VALUE
$\Delta H = 1054 \pm 5\,J\,K^{-1}\,g^{-1}$	$\Delta H = 1049.6\,J\,K^{-1}\,g^{-1}$
$t = 22.5 \pm 0.1$ sec	$t = 22.0$ sec
$m = 136 \pm 4$ g	$m = 130 \pm 10$ g

5 Classify each of the following as quantitative (discrete or continuous) or qualitative data.

a Electronegativity of elements in the periodic table
b Covalent radii of atoms
c Positive charge on cations
d Colours of flames used to identify elements in emission spectroscopy (e.g. red, orange, yellow, green, violet)
e Number of protons in elements in the periodic table

6 A set of data are provided below. The data are:

i a calibration curve of Ca^{2+} concentrations ($[Ca^{2+}]$) versus the reading on an atomic emission spectrometer (arbitrary scale) for two sets of calibration trials
ii the reading on the atomic emission spectrometer (AES) for two samples of river water (last two measurements in the second and third columns). (The AES technique is discussed in Chapter 5.)

CONCENTRATION OF CA^{2+}		
UNITS: PARTS PER MILLION	AES READINGS (ARBITRARY SCALE)	
PPM (CA^{2+})	SAMPLE 1	SAMPLE 2
630	1.214	1.208
460	0.956	0.949
275	0.531	0.529
100	0.191	0.188
0	0.000	0.005
River water	0.825	0.839

a Graph the data for the calibration curve and use a line of best fit to determine the linear relationship for AES reading versus ([Ca^{2+}]. You can use Excel (LINEST function) to determine the uncertainties associated with the line of best fit, or use appropriate manual techniques to obtain an estimate for the uncertainties in the slope and intercept for the linear relationship.

b Using interpolation, determine the concentration of ([Ca^{2+}] in the river water sample.

c Determine the uncertainty in this estimate for the concentration of ([Ca^{2+}] in the river water sample.

7 In a laboratory experiment using atomic absorption spectroscopy, six values were obtained for the concentration of Pb^{2+} ions in a sample of water taken from a stagnant pond near a highway (six separate trials using small vials of water extracted from the same sample):

SAMPLE	[Pb^{2+}] (mol L^{-1})
1	0.330
2	0.325
3	0.308
4	0.309
5	0.333
6	0.329

(Note that Chapter 5 also discusses atomic absorption spectroscopy.)

a What is the mean value for the estimate for [Pb^{2+}] from this experiment?

b Are there any outliers in the data?

c Estimate the appropriate uncertainty in the measurements (e.g. using standard deviation or inspecting the range of values and establishing an appropriate range for the uncertainty).

d Are the number of significant figures provided in the data consistent with the uncertainty?

e Are the sources of error likely to be random, systematic, or both?

END-OF-CHAPTER EXAM

End-of-chapter test

1 A measure of how close your measurements are to a true or accepted quantity is called:

A precision.

B accuracy.

C Neither A nor B

D Both A and B

2 A measure of how closely repeated measurements agree with each other is called:

A accuracy.

B precision.

C percentage error.

D None of the above

3 You measure a mass of a sample of KNO_3 six times using a school laboratory scale and get the following measurements: 0.505 g, 0.503 g, 0.498 g, 0.499 g, 0.501 g, 0.502 g

The mass measured on a calibrated scale at the National Measurement Laboratory yields a result of 0.5270 ± 0.00005 g. How would you characterise your measurements?

A Neither accurate nor precise

B Both accurate and precise

C Accurate but not precise

D Precise but not accurate

4 Which of each of the following pairs of measurements has the greater uncertainty?

a 1 cup ± 1 tablespoon or 1 litre ± 1 tablespoon

b 1 m^2 ± 10^2 cm or 1 km ± 1 cm

c 5 hours ± 15 minutes or 1 day ± 3 hours

d (778.9 ± 0.3) kg or (395.2 ± 0.5) kg

5 How many significant figures are in each of the following?

a 1381 kg

b 0.0232 m

c 9 293 000 cm

d 264.0 mL

e 8.00×10^4 km

f 80 020 g

6 Round each number to three significant figures.

a 0.0003546

b 0.00007529

c 9.090012×10^6

7 Classify each of the following data examples as either qualitative or quantitative, and discrete or continuous.

a The volumes of 10 samples of river water

b Risks associated with conducting an experiment

c The number of seconds in a 1-minute time interval

d Diameters of 10 mL, 50 mL and 100 mL measuring cylinders

e Atomic masses of all the naturally occurring elements

f Number of electrons in each of the second-row elements (Li to F)

g The enthalpy of combustion of the alkanes

h The number of carbon atoms in the linear alkanes

8 An object with a pre-weighed mass of exactly (and correctly) 0.7450 g is given to two students, who must to measure its weight with two digital balances, both with a resolution of 0.01 g. The first student measures the object and states the weight as 0.77 ± 0.005 g. The second student estimates the weight as 0.74 ± 0.01 g.

a Which of the students has quoted the absolute uncertainty correctly? Correct the other student's statement of uncertainty.

b Of the two student measurements, which has the greater percentage error?

c Given that the balances are different instruments, what is the most likely source of error in either of the measurements?

d Are there any random errors in this experiment?

9 A student determined the specific heat capacity of a metal using calorimetry. The results from six separate measurements, using different samples of metal, were as follows.

Trial number	1	2	3	4	5	6
Estimate of specific heat (J K^{-1} g^{-1})	0.393	0.376	0.410	0.365	0.402	0.387

a Determine the mean value for the specific heat capacity of the metal.

b Determine the uncertainty for each measurement based on the number of significant figures given.

c Determine the uncertainty for the estimated specific heat capacity of the metal based on the spread of the data about the mean. Quote the measurement for the specific heat capacity including this estimate of uncertainty.

d Determine the standard deviation for the six measurements and quote the measurement for the mean specific heat capacity including the standard deviation.

e Comment on the relationship between these two estimates of uncertainty.

The accepted value for the specific heat capacities of several metals are as follows (J K^{-1} g^{-1}):

Iron	0.444
Nickel	0.461
Zinc	0.390
Copper	0.385
Brass (an alloy)	0.375

f From the experimental result, can the identity of the metal be established?

g Are any of these metals definitely not the metal for which the specific heat has been measured? Justify your answer in terms of uncertainties.

h Discuss the likely sources of contributions to the uncertainties from random error and systematic error.

11 FUELS

Introduction

Fuels are used for domestic, industrial and transport purposes. Every time you cook your food in your oven or microwave you are using energy generated from fuel burned at a power station. You can see the smog emitted from industrial stacks as a result of burning fuels. Apart from steam trains such as the Valley Rattler in Gympie, Queensland, which burned coal to drive them, trains, cars, trucks and buses are driven by burning other fossil fuels. At current usage rates fossil fuels are predicted to run out in 50 to 100 years. Renewable forms of fuels are being investigated.

Stimulus questions

A Saudi oil minister said, 'The Stone Age didn't end because they ran out of rocks'. How does this relate to the 'Coal Age'?

What Age do you think will emerge from the current Diversified Era of fuel usage?

Discuss the projected beneficial, harmful and unintended consequences to the economy, society and environment of a new Age of fuel usage, particularly for the Asia–Pacific region.

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11.1 Fossil fuels

fossil fuels
decomposed organic matter under high temperature and pressure for a long time

non-renewable fuels
unable to be regenerated at a rate similar to which it is consumed

Fossil fuels are organic matter that has decomposed over time under high temperature and pressure, such as coal (Figure 11.1.1). Fossil fuels are **non-renewable** fuels because they take approximately 250 million years to form and are consumed by most of the world's population of 7 billion people.

There are several types of fossil fuels used for different purposes. Coal is the most reliable form of energy. It has been in use since the 1700s, so scientists have had a long time to perfect the technology to harness coal power for human use. Coal is also the least expensive source of energy. Oil is the fossil fuel of choice for transport, because it is the only commercial non-renewable fuel found naturally in liquid form. In 2015, nearly 75% of total oil consumption in the United States was in the transport sector. While natural gas produces more heat and less carbon dioxide, and contains fewer undesirable impurities than coal or oil; leakage at gas wells and along pipelines have prevented its widespread use, due to fear of explosions. Steam methane (95% of natural gas) re-forming is used to extract hydrogen as a source of power. Although a car driven by hydrogen would be twice as efficient as one driven by petrol, hydrogen's density at normal temperature and pressure is so low that it needs to be compressed into a higher pressure gas or chilled into a liquid to allow for transportation and storage.

11.1.1 300 years of fossil fuels in 300 seconds

11.1.2 Fossil fuels

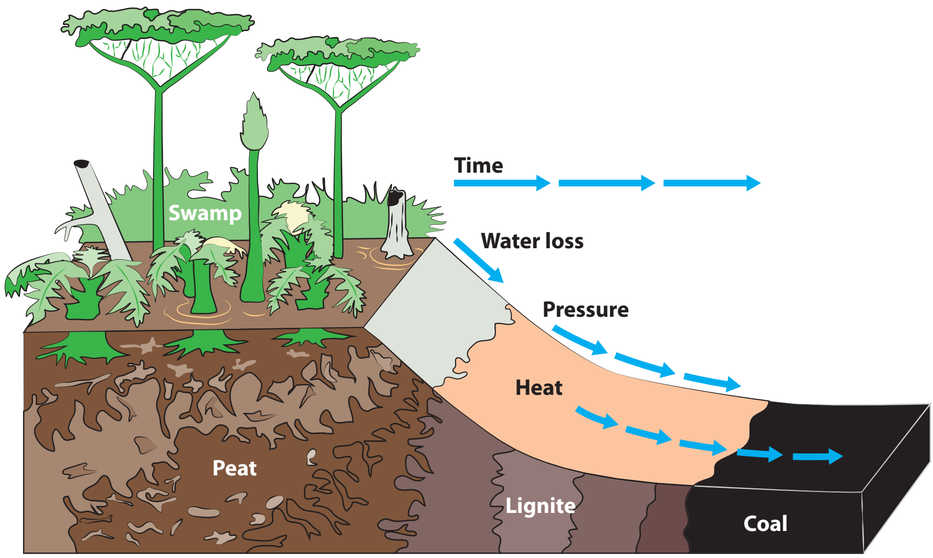

FIGURE 11.1.1 Coal, a fossil fuel, is produced from the decomposition of organic matter over time under high temperature and pressure.

combustion
an exothermic chemical reaction, often rapid, between a fuel and an oxidising agent (usually oxygen) that produces energy

Combustion is a chemical reaction that produces energy involving the rapid combination of a substance with oxygen. The heat from coal combustion was first used to boil water and the steam produced turned a piston to generate kinetic energy to drive a locomotive. In 1882, Thomas Edison built the first coal power plant, which capitalised on Michael Faraday's 1831 discovery of electromagnetic induction. In a coal power plant, coal combustion is used to boil water and the steam turns a turbine. The turbine turns a coil in the presence of a magnetic field, which generates electricity (Figure 11.1.2, page 205).

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Coal

The complete combustion of the major element in coal (carbon) produces carbon dioxide, as shown in the following equation. Ice cores from Antarctica have revealed that prior to the Industrial Revolution, atmospheric carbon dioxide concentrations were approximately 280 ppm and have now risen to about 400 ppm. Carbon dioxide is one of the major greenhouse gases. A **greenhouse gas** provides the **greenhouse effect**, which results in heating of Earth's near-surface atmosphere by the trapping of out-going infrared radiation. Its pre-industrial concentration of 280 ppm raised Earth's average surface temperature from −19°C to +17°C. However, there are several long-term, worldwide problems associated with the large, post-industrial increase in atmospheric carbon dioxide concentration. The increase in atmospheric carbon dioxide concentration increases Earth's surface temperature, melting the ice caps and causing rising sea levels with the potential to inundate many of the world's coastal cities, an economic and social disaster. The increase in atmospheric carbon dioxide concentration also increases the amount of carbon dioxide dissolved in the oceans. Carbon dioxide reacts with water to form carbonic acid, which acidifies oceans. Ocean acidification has the potential to destroy shelled organisms, an environmental and economic disaster.

greenhouse gas gas that traps out-going infrared radiation

greenhouse effect trapping of out-going infrared radiation heating Earth's near-surface atmosphere

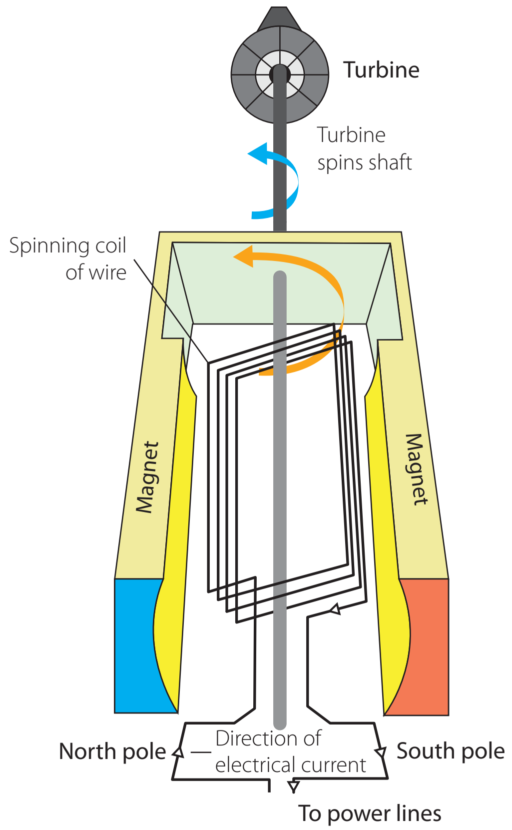

FIGURE 11.1.2 In a power plant, the turbine turns a coil in the presence of a magnetic field to generate electricity

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

Carbon + oxygen → carbon dioxide

The complete combustion of one of the minor elements in coal (sulfur) produces sulfur dioxide:

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

Sulfur + oxygen → sulfur dioxide

An example of the devastating effects of sulfur dioxide was observed during The Great Smog of 1952 in London, UK. At that time, Londoners used coal fires to keep warm. A temperature inversion choked with black smog held the toxic fumes over London for four days, killing more than 4000 people. The disaster ended when high winds dissipated the noxious cloud. Lessons were learnt from the event and domestic coal fires were replaced by internal heating at great cost to households. Industrial sources of sulfur dioxide were dealt with by raising the factory stack heights to increase sulfur dioxide dispersion in the higher altitude winds instead of decreasing the absolute amount of sulfur dioxide emitted into the atmosphere. Although this helped prevent a similar disaster, it also exposed communities and environments to industrial emissions that were previously outside affected areas.

Oil

Oil is composed of many different fossil fuels. The complete combustion of most of the compounds in oil (mainly hydrocarbons) produces carbon dioxide and water. The unleaded petrol used to fuel for our cars contains a large fraction of octane (C_8H_{18}) while the diesel fuel used in trucks and buses contains a large proportion of hexadecane ($C_{16}H_{34}$). The complete combustion of octane is represented by:

$$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$$

Octane + oxygen → carbon dioxide + water

Although petrol combustion leaves more unburnt hydrocarbons than diesel, diesel produces more nitrogen and sulfur oxides and dioxides than petrol. Unlike the Great Smog of 1952, emissions from internal combustion engines are associated with photochemical smog. Chemical reactions that lead to photochemical smog begin with formation of the hydroxyl radical (•OH). Photolysis of nitrogen dioxide produces atomic oxygen, which rapidly reacts with molecular oxygen to form ozone. Photolysis of ozone produces an excited state atomic oxygen, which reacts with water to form a hydroxyl radical. Hydroxyl radicals abstract hydrogen atoms from saturated hydrocarbons (like octane and hexadecane), which initiates the hydrocarbon oxidation into aldehydes, which concurrently converts nitrogen monoxide into nitrogen dioxide.

Natural gas

Methane makes up approximately 95% of natural gas. The complete combustion of methane produces carbon dioxide and water:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

Methane + oxygen → carbon dioxide + water

Water and carbon dioxide absorb much of the radiation below 7700 nm and above 13 000 nm, which leaves a 'spectral window' through which thermal energy escapes. However, methane is one of the gases that closes this spectral window because methane absorbs between 7100 nm and 8300 nm, partially closing the window and providing a greenhouse effect of $0.48\,W\,m^{-2}$.

Hydrogen

The technology of steam methane re-forming produces 95% of the hydrogen produced in the United States per year. Steam methane re-forming is used to obtain hydrogen in its pure form, which can then be used as a fuel. Hydrogen is extracted from methane by reaction with water:

$$CH_4(g) + 2H_2O(g) \rightarrow 4H_2(g) + CO_2(g)$$

Methane + water → hydrogen + carbon dioxide

The major hindrance to steam methane re-forming is its production of carbon dioxide waste. Several carbon sequestration methods are in prototype or industrial trial stage of development. For example, in 2008, Joshuah Stolaroff and co-workers demonstrated the technical feasibility of capturing carbon dioxide directly from ambient air by using a sodium hydroxide spray-based contactor.

Hydrogen, in its pure form, makes an excellent fuel. Combining hydrogen with oxygen provides a large amount of energy with the only waste product being water:

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

Hydrogen + oxygen → water

Although water vapour has the largest greenhouse effect ($110\,W\,m^{-2}$) due to a combination of opposing processes, no human activities directly increase atmospheric water vapour concentrations. Global warming increases water evaporation from oceans, lakes and rivers, which increases atmospheric water vapour concentration, again increasing the greenhouse effect. Conversely, global warming increases evaporation from water bodies, which increases cloud formation, increasing the reflection of the Sun's flux and decreasing global warming. The interplay between these two opposing processes has yet to be elucidated completely.

SECTION REVIEW 11.1

REMEMBERING

1 Define:

- **a** fossil fuels
- **b** non-renewable
- **c** combustion
- **d** greenhouse effect
- **e** greenhouse gas.

UNDERSTANDING

2 Discuss the similarities and differences of fossil fuel waste products.

3 Discuss the trends in energy types generated by burning of fossil fuels.

APPLYING

4 Recommend the best fossil fuel usage for your class, taking into account economic, social and environmental considerations.

11.2 Nuclear fuels

Nuclear fission has been a power source since the 1950s. Although a nuclear reaction does provide the heat to boil water and is comparable to a coal power plant, it is still steam that turns the turbine. The turbine rotates a coil in the presence of a magnetic field, which generates electricity (Figure 11.1.2). Nuclear fission involves the splitting of heavy nuclei, usually uranium. Although nuclear fission is one of the cleanest forms of power while running, producing only water vapour, its potential for environmental disasters in fuel extraction, waste storage and reactor meltdowns have held back an efficient energy source from worldwide long-term proliferation. **Nuclear fusion** is the joining of small nuclei. It occurs in our Sun and all other stars. This type of power generation is free from the environmental disasters of nuclear fission, but technological hurdles, such as containment of plasmas and generation of high temperatures and pressures, need to be overcome to increase the ratio of energy outputs to inputs.

nuclear fission splitting of heavy nuclei

nuclear fusion joining of small nuclei

11.2.1 Nuclear fuels
11.2.2 Nuclear fuel

Nuclear fission

On 27 June 1954, the USSR's Obninsk Nuclear Power Plant was the first to generate electricity for a power grid. In all nuclear fission reactors, slow-moving neutrons strike heavy nuclei to induce splitting. The splitting of a common fuel uranium-235 is shown by:

$$^{1}_{0}\mathrm{n} + {}^{235}_{92}\mathrm{U} \rightarrow {}^{137}_{52}\mathrm{Te} + {}^{97}_{40}\mathrm{Zr} + 2\,{}^{1}_{0}\mathrm{n}$$

Neutron + uranium-235 → tellurium-137 + zirconium-97 + neutrons

When uranium is extracted from the ground it spontaneously emits alpha radiation. Although alpha radiation cannot penetrate the skin from outside the body, it is accompanied by gamma radiation that can go straight through anyone near the uranium mine. Alpha radiation can also contaminate waterways and agricultural land; people who drink this water or eat the resulting crops may develop cancer from alpha radiation ingestion.

Once the uranium has been used up, the fission products left in the reactor need to be stored for a theoretical estimate of 30 half-lives prior to reaching safe levels for human exposure. Although fission products such as zirconium-97 have a half-life of 16.744 hours and need to be stored for only 21 days, other fission products, such as strontium-90, have a half-life of about 20 years and need to be stored for

600 years. The time estimate of 30 half-lives for fission products to reach safe levels for human exposure will be tested practically for the first time in about the year 2554. The logistics of storing toxic waste in 10 centimetres of lead or 1 metre of concrete for at least six generations has only begun to be understood.

Although the well-publicised examples of reactor meltdowns – Chernobyl in 1986 and Fukushima in 2011 – occurred under dissimilar circumstances, both resulted in the explosion of the reactor core contents. Chernobyl was a combination of inherent reactor design flaws and reactor operator mistakes. However, in Fukushima a 9.1 magnitude earthquake and resultant 40.5 m tsunami caused power loss and subsequent cooling system failure.

In 2006, Cardis and co-workers estimated that 4000 people have died from Chernobyl's radiation fallout. In addition, 31 first responders lost their lives within days to months of the incident. It is difficult to determine how many of the 15 900 deaths after the Fukushima incident were directly related to the nuclear reactor meltdown or the earthquake itself. Although Sydney's Lucas Heights facility was designed as Australia's first nuclear power plant, it has never been used for this purpose and today generates radiation for research purposes. Nevertheless, approximately 16% of the world's energy production is derived from nuclear power plants. Many European countries produce more than half their power from nuclear energy, with France producing over 70% of its energy needs from nuclear fission reactors. The United States produces 20% of its energy from nuclear power and is responsible for the largest amount of nuclear-based power production in the world (805 billion kWh).

11.2.3 Nuclear fission and fusion

Nuclear fusion

Helium is the reaction product from the fusion of hydrogen isotopes. Helium is a non-toxic, unreactive gas that already constitutes 5 ppm of Earth's atmosphere. Unlike nuclear fission, nuclear fusion does not generate high-level nuclear waste. Although nuclear fusion's fuels are theoretically non-renewable, they are abundant. Deuterium is readily extracted from ordinary water and tritium is produced from lithium available from sea water. Nuclear fusion is the same reaction that occurs within stars:

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

$$\text{Deuterium} + \text{tritium} \rightarrow \text{helium} + \text{neutron}$$

At the high temperatures and pressures required for this reaction to occur, hydrogen becomes a plasma. Plasmas are essentially a sea of charged particles because electrons are free to move independently of the nucleus. The Russian tokamak design (Figure 11.2.1) for a nuclear fusion reactor that uses toroidal and poloidal magnetic fields to contain the plasmas is still in use today. The Liley tokamak at the Australian National University in Canberra was the only tokamak operating outside the USSR between 1964 and 1969. Like its fission counterpart at Lucas Heights, this fusion facility is used for research rather than power generation. Although the optimal ratio of energy outputs to energy inputs for a nuclear fusion power plant is 50, current facilities are operating at a performance of 3.5 in San Diego, United States, and 3.8–4.7 in Hefei, China. The tokamak under construction in Cadarache, France, is proposed to have a performance of greater than 10.

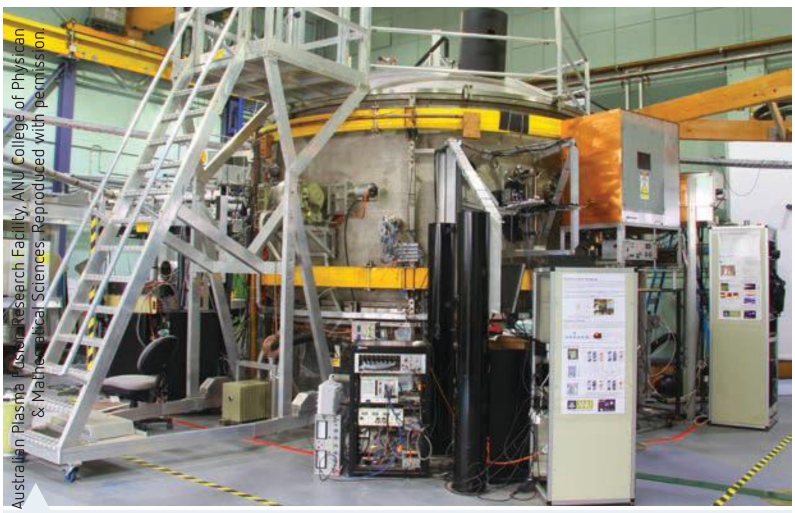

FIGURE 11.2.1 Containment of plasmas by toroidal and poloidal magnetic fields in a nuclear fusion tokamak

SECTION REVIEW 11.2

REMEMBERING

1 Define:

 a nuclear fission

 b nuclear fusion.

UNDERSTANDING

2 Discuss the advantages and disadvantages of nuclear fission as a power source.

3 Discuss the advantages and disadvantages of nuclear fusion as a power source.

APPLYING

4 Recommend the best nuclear fuel usage for Australia taking into account economic, social and environmental considerations, as well as the practicality of generating useful energy in the near future.

11.3 Biofuels

Biofuels are fuels that are made primarily of biomass. Biofuels are **renewable** fuels because they have the potential to be replenished and are available over an essentially limitless time period. In 2015, researchers Lim and Shaw demonstrated an experiment at the Monash University ASELL workshop, which trans-esterified used vegetable oil into biodiesel using an acid-catalysed reaction. Although this prototype process decreases viscosity and increases volatility so that biodiesel can replace fossil-fuel diesel, it has yet to be industrialised due to the high cost of the vegetable oil reactant. An energy investment of the gasification of wood, straw or husks is required to produce methanol as a biofuel. Sugar from sugar cane can be fermented and distilled to produce ethanol as a biofuel. The complete combustion of ethanol produces carbon dioxide and water:

$$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$$

$$\text{Ethanol} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water}$$

Biofuels represent a short-term solution to our reliance on fossil fuels because biofuel combustion still produces carbon dioxide, which contributes to global warming and the subsequent melting of ice caps and ocean acidification (section 11.1). Many Australian petrol stations sell E10 petrol, which contains 10% ethanol. However, Brazilian flex-fuel cars run on blends from 20% ethanol in petrol to 100% ethanol. Brazil's large amount of arable land and efficient sugarcane cultivation has made its 40-year ethanol program possible. This raises questions about the transferability of this program to other countries.

Wind power (such as in South Australia), hydroelectric power (such as the Snowy Mountains scheme) and solar power (such as the Sunshine Coast Regional Council in Queensland) are renewable energy sources that do not produce carbon dioxide. However, they supplement rather than replace reliable coal power plants. Geothermal energy is led by Iceland, a country that straddles the edge of a tectonic plate, so transferability of this power source is of limited use in a country such as Australia, which is positioned in the middle of a tectonic plate.

biofuels
fuels that are primarily made of biomass

renewable fuel
have the potential to be replenished and are essentially limitless

11.3.1 Biofuels
11.3.2 Biofuels and bioprospecting for beginners
11.3.3 Explainer: the evolution of biofuels
11.3.4 Queensland biofuel mandate
11.3.5 Why aren't we only using solar power?
11.3.6 Can 100% renewable energy power the world?

INQUIRING FURTHER

Biofuels have raised a number of questions. Is it fair to take up valuable agricultural land for transport purposes? Is it a good idea to deforest virgin land for cultivating biomass for fuels, when reforestation is viewed as a potential carbon sink for the carbon dioxide produced from burning fossil and biomass fuels? Research one of these questions and present a research report on your findings.

SCIENCE AS A HUMAN ENDEAVOUR

Research a renewable power scheme currently in use in Australia. Identify its economic, social and environmental advantages and disadvantages.

SECTION REVIEW 11.3

REMEMBERING

1 Define:

a biofuel

b renewable energy.

2 Distinguish between supplementing and replacing coal power plants.

APPLYING

3 Recommend the best supplemental fuel usage for different countries across the Asia-Pacific region, taking geographical and prevailing weather conditions into consideration.

11.4 Inquiry skills: analysing and interpreting data

PRACTICAL ACTIVITY 11.4.1

The heat content of ethanol

In 2009, Steehler and colleagues published a comparison of the heat content of several biofuels. A simplification of their methods is described here to determine the heat content of ethanol. (This is equivalent to ΔH_c, the heat of combustion of a chemical compound, as described in Chapter 9, sections 9.5, 9.6 and 9.7.)

Ethanol's energy content (E_c, kJ g^{-1}) will be determined by the energy absorbed by a can of water by burning ethanol:

$$E_c = \frac{m_w \times C_w \times \Delta T}{\Delta m}$$

$$\text{Energy content} = \frac{\text{(mass of water} \times \text{specific heat of water} \times \text{temperature difference)}}{\text{mass of fuel lost}}$$

The energy absorbed by the water will be the mass of water (m_w)(g), multiplied by the specific heat of water (C_w) (= 4.184×10^{-3} kJ g^{-1} °C^{-1}), multiplied by the temperature difference [ΔT (°C) = final temperature (T_f) – initial temperature (T_i)]. The amount of ethanol burned [Δm (g)], will be determined by the mass lost from the fuel can [initial mass (m_i) – final mass (m_f)].

AIM

To determine the energy content of ethanol

EQUIPMENT PER GROUP

- 200 mL ethanol
- 500 mL water
- electronic balance
- empty soft-drink can
- glass stirring rod
- 100 mL measuring cylinder
- thermometer
- fuel can
- heat-proof gloves
- lighter or matches
- ruler
- 1 L beaker

»

WHAT ARE THE RISKS IN DOING THIS ACTIVITY?	HOW CAN THESE RISKS BE MANAGED?
Equipment can get extremely hot.	Wear heat-proof gloves.

RISK ASSESSMENT

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, as well as ways to manage them.

PROCEDURE

1. Set up the apparatus as shown in Figure 11.4.1.
2. Zero a balance with a soft-drink can with a glass stirring rod through it.
3. Pour 100 mL water from a measuring cylinder into the soft-drink can.
4. Record the mass of the water (m_w).
5. Measure and record the initial temperature of the water (T_i).
6. Measure and record the initial mass of the fuel can (m_i).
7. Set up the apparatus with a 2.5 cm wick and 2 cm between the top of the wick and bottom of the soft-drink can.
8. Don heat gloves.
9. Ignite the fuel can with a lighter.
10. Stir the water in the soft-drink can with the thermometer as the water heats.
11. When the water temperature has been raised by about 30°C, remove the fuel can from beneath the soft-drink can and extinguish it by placing the 1 L beaker over the fuel can.
12. Continue to stir the water in the soft-drink can with the thermometer until it reaches a maximum value.
13. Measure and record the final temperature of the water (T_f).
14. Measure and record the final mass of the fuel can (m_f).
15. Empty the heated water from the soft-drink can.
16. Repeat steps 1–15 twice more.

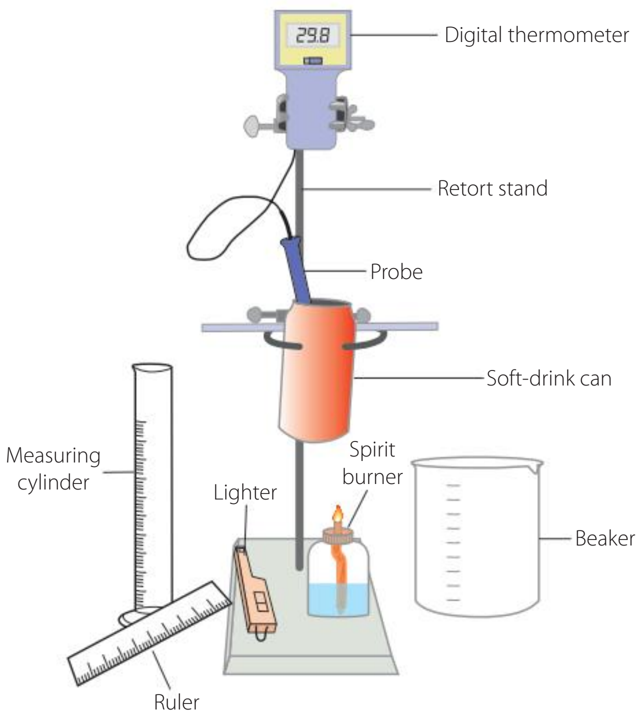

FIGURE 11.4.1 Apparatus set-up to determine the heat content of ethanol

CALCULATIONS

- m_w (g) was recorded in step 4
- $C_w = 4.184 \times 10^{-3}$ kJ g^{-1} °C
- ΔT (°C) = T_f (step 13) – T_i (step 5)
- Δm (g) = m_i (step 6) – m_f (step 14)
- Calculate the energy content of ethanol for each trial using the equation from the beginning of this practical activity (page 210).
- Calculate the average energy content of ethanol for all trials using the equation:

$$\text{Average } E_c = \frac{\text{sum of } E_c \text{ of each trial}}{\text{number of trials}}$$

QUESTIONS

1 Is the average energy content of ethanol calculated in this experiment the same as the literature value published by NIST (29.67 ± 0.05 kJ g^{-1})? If not, why not?

2 The heat of combustion (heat content) of octane (one of the more abundant hydrocarbons in petrol for cars) is by comparison, 47.9 kJ g^{-1}. Discuss the relevant efficiencies of each fuel in terms of energy per litre. (Convert mass to volume using data researched for the density of ethanol and octane.)

SECTION REVIEW 11.4

APPLYING

1 Apply your knowledge of the terms 'precision' and 'accuracy' to assess the precision and accuracy of this technique for determining the heat content of ethanol.

2 Modify the experiment to obtain a value of ethanol heat content closer to 29.9 kJ g^{-1}, which was determined by Hammerschlag in 2006.

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CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 Identify the dangers associated with the fuel extraction, waste storage and meltdown of nuclear fission power plants.

CATEGORY QUESTIONS

2 List the equipment used during the determination of heat content of fuels in this chapter compared to industry.

3 State the chemical reactions that take place in:

a a coal power plant

b an internal combustion engine.

4 Identify the chemical reactions associated with the burning of biofuels.

5 Identify the locations associated with the use of:

a wind energy

b solar energy

c geothermal energy.

6 Construct a list of fuels associated with the production of carbon dioxide.

ELABORATION QUESTIONS

7 Identify the purposes with which deuterium and tritium are associated.

8 State the physical traits of hydrogen that are associated with the hindrance of its widespread use as a fuel for the transport sector.

EVIDENCE QUESTIONS

9 Explain why electric trains are used today, rather than steam trains.

10 If you were an energy company representative, would you focus your future power supply pitch on supplements or replacements to coal power plants? Justify your reasoning.

END-OF-CHAPTER EXAM

End-of-chapter test

Questions 1 to 4 refer to the following equation:

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

1 What observation(s) suggest(s) that the chemical reaction has occurred?

A Solid disappears

B Liquid colour changes

C Colourless gas evolves

D Both A and C

2 The element carbon (C) is the principal component in one of the following.

A Fossil fuels

B Nuclear fuels

C Biofuels

D Renewable fuels

3 CO_2 is:

A a greenhouse gas.

B contributing to melting of the ice caps.

C contributing to ocean acidification.

D All of the above

4 The above chemical reaction:

A is a precipitation reaction.

B always goes to completion.

C is a combustion reaction.

D is an endothermic process.

5 ___________ fuels cannot be replenished on the timescale of consumption.

6 ___________ fuels can be replenished on the timescale of consumption.

7 ___________ nuclei can be split into smaller nuclei by nuclear ___________.

8 ___________ nuclei can be joined into larger nuclei by nuclear ___________.

9 What is the difference between the Great Smog of 1952 and photochemical smog?

10 What is the difference between the natural and anthropogenic (man-made) greenhouse effect?

11 In 2009, Steehler and colleagues determined the heat content of ethanol as part of their investigation of several biofuels. Calculate the heat content of ethanol using the following data.

- $m_w = 98.3\,g$
- $C_w = 4.184 \times 10^{-3}\,kJ\,g^{-1}\,°C$
- $T_f = 45.2°C$
- $T_i = 24.9°C$
- $m_i = 153.4\,g$
- $m_f = 151.9\,g$

12 MOLE CONCEPT AND THE LAW OF CONSERVATION OF MASS

Introduction

Calculating chemical quantities requires use of measurements that are related to atoms and molecules, their masses, how they are distributed in space (or in a volume), and how that distribution can be quantified. Chemical changes that occur as a result of reactions also need to be measured and identified. The relative atomic mass of an element is defined as the ratio of the weighted average mass per atom of the naturally occurring form of the element to $\frac{1}{12}$ the mass of an atom of carbon-12. Relative atomic masses reflect the isotopic composition of the element. The measurement of concentrations of atoms or molecules requires scientists to know how many atoms or molecules occupy a given volume. Scientists need to define relationships that permit calculations using quantities that represent atomic and molecular masses. However, they also need to measure these quantities using instruments that are readily available, such as a digital balance for measuring mass, or a pipette for dispensing a volume of liquid.

Stimulus questions

Can chemical methods in the laboratory be used to work out the actual masses of the tiny particles, atoms and molecules that make up matter?

Can laboratory instruments be used to weigh out a specific number of molecules of a substance?

How can equations be written for chemical reactions so that it is easy to identify how much of one compound might react with another, and how much of one or more products will be formed?

What are the connections between atomic mass, molecular mass and molar mass?

iStockphoto.com/Claudia Prommegger

12.1 Mole and Avogadro's number of particles

Individual atoms and molecules are too small to see even with the most specialised equipment. For this reason, chemists devised a unit that is practical for counting and handling a standard number of particles.

INQUIRING FURTHER

The International Avogadro Coordination (IAC), often referred to as the 'Avogadro project', begun in the early 1990s as collaboration between various national metrology institutes. The project aimed to measure the Avogadro constant by the X-ray crystal density method to a relative uncertainty of 2×10^{-8} or less. Their paper, published in January 2011, summarised the result of the International Avogadro Coordination and presented a measurement of the Avogadro constant to be $6.02214078 \times 10^{23}$.

atomic mass unit one-twelfth of the mass of on atom of carbon-12

Science Photo Library/Andrew Lambert Photography

FIGURE 12.1.1 A mass of 12 g of powdered carbon – a couple of teaspoons in volume

Avogadro constant $6.02214078 \times 10^{23}$ (usually approximated as 6.022×10^{23})

12.1.1 How big is the mole?
12.1.2 The mole and Avogadro's number
12.1.3 How was Avogadro's number determined?

The Avogadro constant

Chemists deal with amounts of substances in terms of their masses, so they use a relationship between mass and number of particles.

The mass of an atom is its relative atomic mass (represented by the symbol A_r). The relative atomic mass of an atom is its mass relative to a carbon-12 atom. The relative masses of the atoms of an element are devised by comparing them to the mass of carbon-12. The mass of a carbon atom is considered to be exactly 12 and all others are compared to this. One **atomic mass unit** (amu) is equal to one-twelfth the mass of one atom of carbon-12. Chemists were able to work out that one carbon-12 atom, with a relative atomic mass of 12 amu, has an actual mass of 1.99265×10^{-23} grams. However, no laboratory scale can measure the mass of an individual atom. For everyday use, chemists need to measure quantities of substances in a larger unit, such as a gram, but they still need to know how many particles there are.

Chemists chose a number of particles that would have a mass, in grams, equivalent to the mass of one atom in atomic mass units. The same number fits all elements because equal numbers of different atoms always have the same mass ratio. The number was determined to be about 6.02×10^{23} and is called the Avogadro constant (represented by the symbol N_A) in honour of Amedeo Avogadro, a 19th-century Italian scientist. The **Avogadro constant** (N_A) is the number of atoms (6.022×10^{23}) in exactly 12 grams of the carbon-12 isotope. Figure 12.1.1 shows a sample of carbon powder weighing exactly 12.00 g. Therefore, we are looking at a sample that must contain 6.022×10^{23} atoms of carbon! The Avogadro constant is a scaling factor between macroscopic and microscopic (atomic scale) observations of nature.

Naming the Avogadro constant

The Avogadro constant was initially determined by French physicist Jean Perrin, who proposed it be named after Amedeo Avogadro to honour his contributions to chemistry. One of Avogadro's most important contributions was his resolution of the confusion surrounding atoms and molecules (although he did not use the term 'atom'). Avogadro believed that particles could be composed of molecules and that molecules could be composed of still simpler units, atoms.

The Avogadro constant is a number so large it is difficult to comprehend. If you had the same number of oranges as the Avogadro constant, they would form a sphere the size of Earth. Alternatively, this number of grains of rice would cover Australia to a depth of approximately 1 kilometre.

The relationship of Avogadro's constant to the mole

Just as it is convenient to group sheets of paper into reams (500 sheets) and eggs into cartons of a dozen, chemists measure the amount of substances in a batch size called the mole (mol) and represent the number of moles by the symbol n.

Chemists have chosen 1 mole as a standard unit for large numbers of atoms, ions or molecules. A mole is the amount of substance containing N_A particles of that substance. For most calculations, it is sufficiently precise to take $N_A = 6.02 \times 10^{23}$. The **mole** is the SI base unit representing the chemical quantity of a substance.

mole
the amount of a substance (g) containing $N_A = 6.022 \times 10^{23}$ particles (atoms, ions or molecules) of that substance

Converting between moles and number of particles

The relationship between the number of moles (n) of a substance and the number of particles (atoms, ions or molecules) is given by:

$$\text{Number of moles } (n) = \frac{\text{number of particles}}{\text{number of particles per mole}}$$

where the number of particles per mole $= 6.02 \times 10^{23}$.

So:

$$n = \frac{\text{number of particles}}{6.02 \times 10^{23}}$$

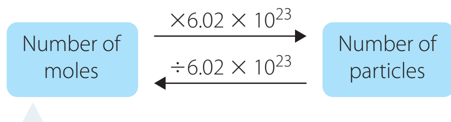

FIGURE 12.1.2 Converting between moles and number of particles

Conversion between number of moles and number of particles, or vice versa, can be carried out using this relationship, as shown in Figure 12.1.2.

Moles and chemical formulas

The molecular formulas of elements and compounds contain subscripts that represent the numbers of atoms of each element present in a molecule of the substance. For example, in one molecule of water (H_2O), there are two atoms of hydrogen and one atom of oxygen. In two molecules of water, there are double the number of hydrogen (4) and oxygen (2) atoms, and so on.

The subscripts also represent the number of moles of each atom present in one mole of molecules of the substance. So, if there were one mole of water molecules, there would be two moles of hydrogen atoms and one mole of oxygen atoms. Similarly, if there were 2 moles of water molecules, then there would be 4 moles of hydrogen atoms and 2 moles of oxygen atoms.

This relationship also applies to ionic compounds. One mole of sodium sulfate (Na_2SO_4) formula units contains 2 moles of sodium ions (Na^+) and 1 mole of sulfate ions (SO_4^{2-}).

The number of individual atoms in a known mass (g) can be calculated using the relationship between moles and the Avogadro constant.

SECTION REVIEW 12.1

REMEMBERING

1 Define:

a Avogadro constant **b** the mole.

UNDERSTANDING

2 Explain the relationship between the Avogadro constant and the mole.

APPLYING

3 Calculate the number of moles of each of the following substances, given the number of particles.

- **a** 4×10^{24} atoms of helium
- **b** 1.03×10^{22} molecules of chlorine gas
- **c** 1.204×10^{23} ions of copper
- **d** 8.22×10^{21} formula units of aluminium oxide

4 Calculate the number of particles of each of the following substances, given the number of moles.

- **a** 2.5 mol of oxygen molecules
- **b** 0.33 mol of gold atoms
- **c** 1.6 mol of iron(II) ions
- **d** 0.035 mol of silver nitrate

5 Copper(II) carbonate has the formula $CuCO_3$. In 3 mol of $CuCO_3$ how many of each of the following are there?

- **a** Moles of Cu^{2+} ions
- **b** Moles of CO_3^{2-} ions
- **c** Moles of oxygen atoms
- **d** Individual atoms of oxygen

6 Octane (C_8H_{18}) is an important constituent of petrol. In 3×10^{26} molecules of octane, how many moles of each of the following are present?

- **a** C_8H_{18} molecules
- **b** C atoms
- **c** H atoms

ANALYSING

7 How many moles of oxygen atoms are in each of the following?

- **a** 2.5 mol of H_2SO_4 molecules
- **b** 0.2 mol of CH_3COOH molecules
- **c** 3.02×10^{24} formula units of $Al_2(CO_3)_3$

12.2.1 The law of conservation of mass
12.2.2 Conservation of mass

12.2 The law of conservation of mass in chemical equations

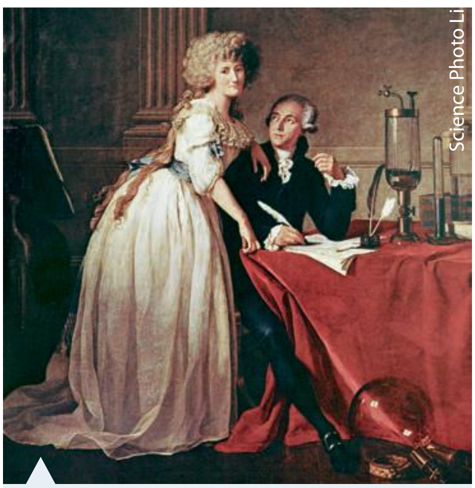

FIGURE 12.2.1 Marie and Antoine Lavoisier

When a chemical reaction occurs, new substances are formed. Chemical reactions can be represented using chemical equations. Balanced chemical equations represent the relationship between the number of particles that react and the number of particles that are produced.

Chemical equations must be balanced because atoms are conserved. This is a consequence of the law of conservation of mass: mass is neither created nor destroyed by chemical or nuclear reactions, radioactive decay or physical transformations. The law of conservation of mass was developed by Antoine Lavoisier.

Moles and chemical reactions

The reaction between magnesium (Mg) and oxygen (O_2) can be represented by the balanced equation:

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

The coefficient of each species in the equation indicates the ratio in which the species react. In this case, two atoms of magnesium react with one molecule of oxygen gas to produce two formula units of magnesium oxide (MgO).

KEY LAW

The law of conservation of mass
Mass is neither created nor destroyed by chemical or nuclear reactions, radioactive decay or by physical transformations

The chemical equation can also be read in terms of moles, because the number of particles is directly proportional to the number of moles through the Avogadro constant. This means the coefficients of the species in the equation also indicate the molar ratios for the species taking part in the reaction; that is, 2 moles of magnesium react with 1 mole of oxygen gas to produce 2 moles of magnesium oxide.

The study of the amounts of reactants and products in a chemical reaction is called **stoichiometry**. This word comes from the Greek words *stoicheion*, meaning 'element', and *metron*, meaning 'measurement'.

stoichiometry the relationship between the relative quantities of substances taking part in a reaction or forming a compound, typically a ratio of whole integers

Consider the reaction for the complete combustion of methane gas (CH_4) to carbon dioxide (CO_2) and water:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

The relationships between the number of moles of reactants and products are represented in Table 12.2.1. Although the mole quantities (numbers of moles) are different, the ratios remain the same. All are in the ratio of 1:2:1:2.

TABLE 12.2.1 Ratio of particles involved when methane reacts with oxygen gas

CH_4	$+ 2O_2(g) \rightarrow$	$CO_2(g)$	$+ 2H_2O(g)$
1 mol	2 mol	1 mol	2 mol
2 mol	4 mol	2 mol	4 mol
0.5 mol	1 mol	0.5 mol	1 mol
5 mol	10 mol	5 mol	10 mol

If the number of moles of any species in the reaction are known, apply the balanced equation to work out the number of moles of all the other species in that reaction.

Mass and chemical equations

Chemical equations indicate a relationship between numbers of moles of reactants and products, and masses are related to moles. This means that chemical equations can also be used to determine relationships between masses of reactants and products.

Remember that chemical reactions obey the law of conservation of mass. The direct consequence of this law is that, in chemical reactions, the combined mass of the products must equal the combined mass of the reactants.

For example, in the reaction between magnesium and oxygen, the mass–mass relationships between reactants and products are:

$$2Mg(s) + 1O_2(g) \rightarrow 2MgO(s)$$

$$2\,mol + 1\,mol \rightarrow 2\,mol$$

$$48.6\,g + 32\,g \rightarrow 80.6\,g$$

Notice that, although atoms and mass are conserved, moles are not. This is because moles count the number of molecules and formula units, not necessarily individual atoms. This means that 1 mole of a substance may contain many moles of individual atoms.

SECTION REVIEW 12.2

REMEMBERING

1. Define 'stoichiometry'.
2. State the law of conservation of mass.
3. Why do reactants and products always need to be related using a balanced chemical equation?

UNDERSTANDING

4. Copy and complete the following table to determine the missing numbers of moles using the equation:

$$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$$

Fe_2O_3	CO	Fe	CO_2
1 mol			
		3 mol	
	0.6 mol		
			0.09 mol

12.3 Empirical and molecular formulas

Relative mass of substances

Chemical symbols and formulas, such as He, Na, Cl_2 and H_2O, are short-hand representations for elements and compounds. Symbols and formulas may also represent a group of atoms or formula units. Just as relative atomic mass is used to describe the mass of an atom, relative molecular mass (symbol M_r) is used to describe the mass of one molecule of a molecular substance on a scale in which the mass of an atom of the carbon-12 isotope is exactly 12 (units are atomic mass units). The relative molecular mass (M_r) of a substance is the mass of one molecule of the substance on a scale in which the mass of an atom of the carbon-12 isotope is exactly 12.

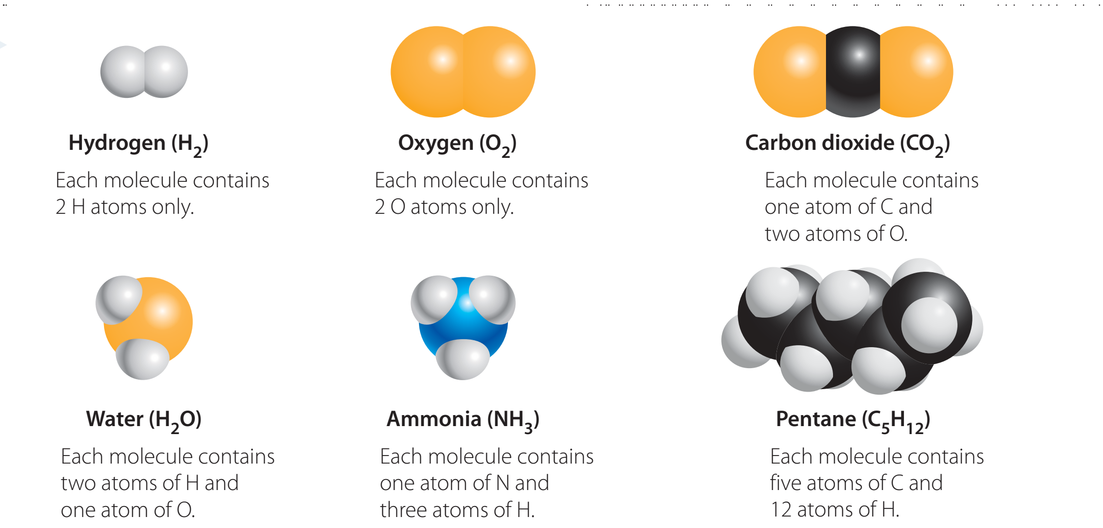

FIGURE 12.3.1 Space-filling models of molecules. The sizes of atoms are shown as their covalent radii

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The relative molecular mass of a molecule, expressed in atomic mass units, is calculated by adding the relative atomic masses of all the component atoms of the molecule.

Not all pure substances exist as molecules. Ionic compounds consist of a crystalline lattice of cations and anions combined in a definite fixed ratio (Figure 12.3.2), as indicated by the formula of the compound, such as Al_2O_3. The term 'relative molecular mass' is incorrect applied to ionic substances and '**relative formula mass**' should be used instead. This is also given the symbol M_r. The relative formula mass is the mass of one formula unit of an ionic compound on a scale in which the mass of an atom of the carbon-12 isotope is exactly 12. The relative formula mass of a compound (M_r) is the sum of the relative atomic masses of the atomic species as given by the formula of the compound.

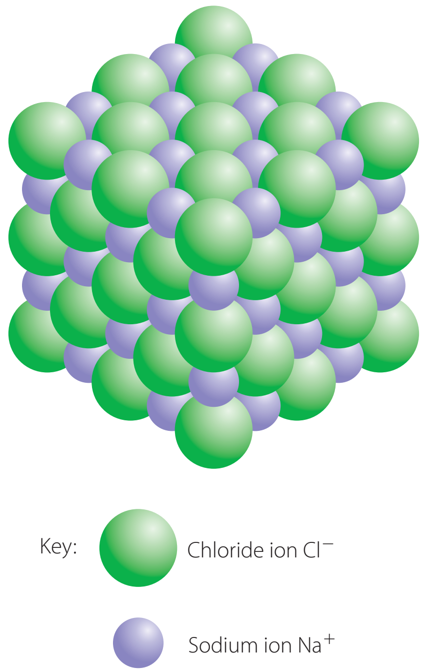

FIGURE 12.3.2 Sodium chloride Na^+Cl^- – an example of an ionic compound

Percentage composition

The chemical formula of a compound gives information about the elements present in a compound and the ratio in which the atoms of those elements are present. For example, in phosphoric acid (H_3PO_4), H, P and O are present in the ratio 3:1:4.

It is important to remember that the ratio of the atoms differs from the ratio of the masses of the atoms because atoms have different relative atomic masses. Sometimes, scientists need to know the ratio by mass or the percentage composition.

The percentage composition of a compound is the percentage by mass of each of the different elements in the compound. For example, Figure 12.3.3 shows aluminium foil, which has approximately 98.5% of the element Al compared to bauxite, which is the mineral from which aluminium is derived and aluminium oxide (alumina). Aluminium oxide (Al_2O_3) contains approximately 53 % Al (as shown in Worked example 12.6.2).

Knowledge of percentage composition provides useful information in many different instances. For example, a chemist can determine which mineral ores contain the higher percentage of a particular metal to advise mining and metal extraction industries. A chemist often compares the percentage composition of an unknown compound with the percentage composition calculated from a known formula. If the percentages match, the identity of the unknown can be confirmed.

relative formula mass (for ionic compounds) the mass of one formula unit of an ionic compound on a scale in which the mass of an atom of the carbon-12 isotope is exactly 12

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FIGURE 12.3.3 Bauxite, aluminium oxide (alumina) and aluminium have different percentage compositions of aluminium

Calculating percentage composition

percentage composition the percentage by mass of each of the different elements in the compound

The **percentage composition** can be calculated from the chemical formula and the relative atomic masses of the elements in the compound. This composition is independent of how much of the compound there is.

German chemist Martin Klaproth achieved prominence in the late 18th to early 19th centuries through his development of new procedures for analysing compounds. Although early chemists, including Antoine Lavoisier, recalculated analysis results so they totalled 100%, Klaproth believed that, when using pure reagents and attention to detail, a less than 100% result suggested the presence of another element. He was the first to discover uranium and zirconium and characterised them as distinct elements. He also elucidated the composition of numerous compounds containing tellurium, strontium, cerium and chromium.

Moles and mass

Chemists often need to calculate the number of moles present in a given mass of a substance. The mass (in grams) of 1 mole of a particular element or compound is numerically equal to its relative atomic, molecular or formula mass. The relative formula mass, relating to a single molecular, ionic or formula unit, is expressed in atomic mass units. The mass of 1 mole of a substance is called the molar mass (M). The units of **molar mass** are grams per mole (g/mol or $g\,mol^{-1}$).

molar mass mass of 1 mole of a substance (equals mass of $N_A = 6.022 \times 10^{23}$ atoms, molecules or ions of a substance)

For example, the molar mass of magnesium (Mg) is $24.31\,g\,mol^{-1}$, the molar mass of oxygen gas (O_2) is $(2 \times 16.0) = 32\,g\,mol^{-1}$ and the molar mass of CO_2 is $(1 \times 12 + 2 \times 16) = 44\,g\,mol^{-1}$.

KEY FORMULA

One mole of any substance has a mass equal to the relative atomic, molecular or formula mass of that substance, but expressed in grams. The molar mass (M) is an *actual* mass that can be measured in grams.

FIGURE 12.3.4 One mole each of five different substances. The mass of 1 mole varies depending on the substance.

Converting between moles and mass

If you know the amount of a given substance in grams, then you can calculate the number of moles and vice versa. For example, if you have 18 g of water, then you must have 1 mol of water because the molar mass is $18.0\,g\,mol^{-1}$. If you want to measure out 2 mol of water, then you would measure $(2 \times 18) = 36$ g of water.

The relationship between the number of moles (n), mass (m) and the molar mass (M) of a substance is given by:

$$\text{Number of moles } (n) = \frac{\text{mass (g)}}{\text{molar mass (g mol}^{-1})}$$

$$n = \frac{m}{M}$$

This formula can be rearranged if you want to calculate the mass of a given number of moles:

$$m = n \times M$$

Alternatively, the conversion between mass and moles can be represented schematically as shown in Figure 12.3.5.

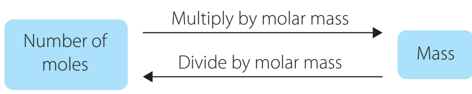

FIGURE 12.3.5 Converting between the number of moles and mass

Empirical and molecular formulas

The **empirical formula** of any compound is the simplest whole number ratio in which the atoms of the elements are present in the compound. For an ionic compound, this is the formula that is given because ionic compounds have an infinite lattice structure. In contrast, a molecular substance consists of individual molecules, so its molecular formula states the actual number of each type of atom present in the molecule.

empirical formula a chemical formula showing the simplest ratio of elements in a compound rather than the total number of atoms in the molecule

For example, the compound hydrogen peroxide has the molecular formula H_2O_2. It is composed of two atoms of hydrogen and two atoms of oxygen bonded together. The empirical formula of hydrogen peroxide is HO because this is the simplest whole number ratio. However, for water (H_2O), the empirical and molecular formulas are the same because the ratio of atoms cannot be simplified any further.

Other examples of empirical and molecular formulas are shown in Table 12.2.1.

TABLE 12.3.1 Empirical and molecular formula of different compounds

SUBSTANCE	MOLECULAR FORMULA	EMPIRICAL FORMULA
Methane	CH_4	CH_4
Octane	C_8H_{18}	C_4H_9
Glucose	$C_6H_{12}O_6$	CH_2O
Calcium nitride	Not applicable (ionic)	Ca_3N_2
Silicon dioxide	Not applicable (ionic)	SiO_2

It is possible for different compounds to have the same empirical formula.

Determining empirical and molecular formulas

The empirical formula of a substance can be determined experimentally. The term 'empirical' means obtained from experiment. This can be done by experimentally determining the percentage composition and then applying the mole concept to determine the ratio of particles. (Remember, the ratio of particles and moles of particles are the same.)

Sometimes, dividing by the smallest number does not give a ratio close to a whole number. For example, if the ratio is 2.33 to 1, multiplying by 3 gives a ratio of 7 to 3. Other common ratios are 1.33 to 1, or 4 to 3; and 1.67 to 1, or 5 to 3. When these non-integer ratios occur, another step is needed in the calculation; that is, multiplication by an integer that converts the ratio to a whole number (integer) ratio.

SECTION REVIEW 12.3

REMEMBERING

1 Distinguish between relative molecular mass and molar mass.

2 Distinguish between an empirical formula and a molecular formula.

APPLYING

3 Calculate the molar mass of:

a hydrogen peroxide (H_2O_2)

b magnesium bromide ($MgBr_2$)

c calcium carbonate ($CaCO_3$)

d pentane (C_5H_{12}).

4 Identify the mass of each of the following.

a 7.1×10^{23} atoms of neon gas

b 5.43 mol of lead

c 3.14×10^{24} molecules of phosphorus tetramer (P_4)

d 1.236 mol of magnesium sulfate ($MgSO_4$)

e 0.452 mol of potassium phosphate (K_3PO_4)

5 Write out the empirical formula of each of the following compounds.

a $COCl_2$

b C_4H_{10}

c N_2O_4

d $Na_4P_2O_6$

6 How many moles are there in 10.00 g of each of the following?

a Water

b Ethanoic (acetic) acid (CH_3COOH)

c Ethane (C_2H_6)

d Sodium bromide

e Barium chloride.

7 An oxide of chlorine contains 18.4% oxygen by mass. State the empirical formula of the compound.

8 Nicotine was analysed to determine its composition. It was found to contain the following percentages by mass: 74.0% C, 8.7% H and 17.3% N. Its relative molar mass was found to be 162. State the:

a empirical formula

b molecular formula.

ANALYSING

9 The artificial sweetener NutraSweet® is the compound aspartamine, which has the molecular formula $C_{14}H_{18}N_2O_5$.

a What is the mass of 1.00 mol of aspartamine?

b How many moles of aspartamine are in 6.22 g of the substance?

c What is the mass of 0.245 mol of aspartamine?

d How many molecules are there in 4.28 mg of aspartamine?

10 a How many molecules of oxygen gas are there in 4.00 g of O_2?

b How many atoms of oxygen would be obtained if the molecules in Question **12a** were split in half (i.e. into O atoms)?

11 a If the body of a person is 80% water, how many moles of water would be in the body of a person who weighs 80 kg?

b How many water molecules is this?

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12.4 Experimental and theoretical yield

Application of the relationships between masses of products and reactants is very important in chemistry. It is a vital factor in calculating how much of a desired product is made and how much by-product (waste) is also made. This is important in all chemical manufacturing, including the plastic, metal, pharmaceutical, fertiliser and petroleum industries.

When one or more chemicals react to give one or more product molecules, all but one of the reagents may be in excess. This means that this reagent is a **limiting reagent**. It will have reacted completely at the end of the reaction, meaning it will have all been used up as a result of reacting with the excess reagent(s). However, if there is a **stoichiometric ratio** of reactants, then each of the reactants will be present in just sufficient quantity to be completely used up at the end of the reaction. Whichever situation prevails in a chemical reaction process, the maximum amount of product that can be produced is called the **theoretical yield**.

limiting reagent the reactant that it is totally consumed when the chemical reaction is complete

stoichiometric ratio where the mole ratios for reactant atoms or molecules present in a chemical equation are in the exact proportion for the reaction to occur with no reactant in excess

theoretical yield the amount (mass) of a product that is produced from the complete reaction of the limiting reagent

Mass–mass calculations

Mass–mass calculations use the mass of a known substance and a balanced chemical equation to calculate the mass of all other species (reactants and products) involved in the reaction.

When solving problems involving mass–mass stoichiometry, it is important to remember that the mass relationships between reactants and products are derived from the mole ratios in a balanced chemical equation.

Figure 12.4.1 shows the steps involved in mass–mass calculations.

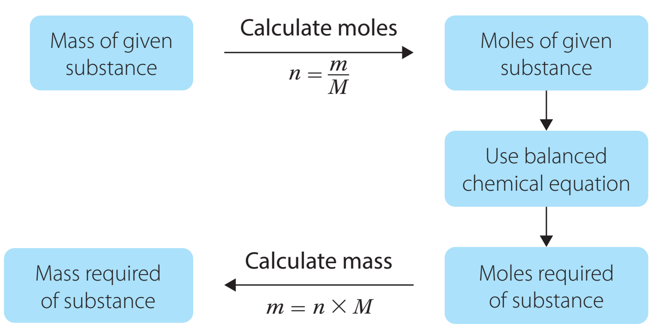

FIGURE 12.4.1 Mass–mass calculations

experimental yield the actual yield (mass) of a product obtained for a reaction conducted in the laboratory

A balanced chemical equation provides the information required to determine what the theoretical or ideal yield of the reaction should be. However, the theoretical yield assumes perfect completion of the reaction. In an experimental situation, the actual **experimental yield** may be less than the theoretical yield. There are a number of possible reasons for this.

1. There may be competing reactions that have not been considered in the written, balanced equation.
2. The mechanism (stoichiometry) of the reaction may be different due to experimental conditions differing from what is required for the assumed reaction.
3. The conditions required for the reaction to take place, such as temperature and pressure, may be imperfect.
4. The reaction mixture may contain impurities that influence the reaction.

Generally, experimentally obtained yields fall short of calculated theoretical yields.

The ratio of the actual yield to the theoretical yield is a quantity known as the 'fractional' yield. This is usually quoted as a percentage of the theoretical yield and is defined as the percentage yield:

KEY FORMULA

$$\text{Percentage (\%) yield} = \frac{\text{experimental yield}}{\text{theoretical yield}} \times \frac{100}{1}$$

Calculating yields

Consider the combustion reaction of methane (CH_4). Methane burns in air by combining with oxygen molecules to yield carbon dioxide (CO_2) and water. The balanced chemical reaction is:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

One mole of methane requires 2 moles of oxygen for the stoichiometry to be satisfied. If methane is burning on a gas stove, there will be a plentiful supply of oxygen from the surrounding air. For every mole of methane that undergoes combustion, scientists predict obtaining 1 mole of carbon dioxide.

Suppose a scientist measures that 32 g of methane have been consumed in a certain amount of time during cooking rice on a stove. What is the theoretical yield of carbon dioxide produced? The number of moles of methane are calculated to provide the number of moles of carbon dioxide:

$$CH_4\,(n) = \frac{m}{M} = \frac{32}{16} = 2\text{ mol of }CH_4 = \text{number of moles of }CO_2\text{ produced}$$

$$\begin{aligned}\text{The mass of }CO_2\text{ produced} &= \text{number of moles of }CO_2 \times \text{molar mass of }CO_2\\ &= 2 \times 44\\ &= 88\text{ g}\end{aligned}$$

The theoretical yield of carbon dioxide produced from burning 32 g of methane in an excess of oxygen is 88 g.

Scientists can calculate the theoretical yield of water produced from burning 32 g of methane similarly. For each mole of methane, 2 moles of water are produced, so 4 moles of water will be produced in total. This gives a theoretical water yield of $4 \times 18\text{ g} = 72\text{ g}$.

WORKED EXAMPLE 12.4.1

In a laboratory, scientists collected the water formed from the combustion of 32 g of methane using a condensation apparatus and then weighed the water.

QUESTION

What is the experimental yield of water resulting from the combustion of 32 g of methane?

ANSWER

1 When the reaction was run in the lab, only 60 g of water were collected. The fractional (experimental) yield of water for this experiment is given by:

$$\text{Fractional yield} = \frac{\text{experimental yield}}{\text{theoretical yield}}$$

$$\text{Percentage (\%) yield} = \frac{\text{experimental yield}}{\text{theoretical yield}} \times \frac{100}{1}$$

2 Where 60 g of water were collected, the % yield is:

$$\text{Percentage yield} = \frac{60}{71} \times \frac{100}{1} = 83\%$$

The percentage yield is 83%.

In some reactions, two steps may take place overall. For example, pure zinc metal can be produced in a two-step process that starts with zinc sulfide (ZnS). ZnS is the main constituent of the mineral sphalerite. The mineral will contain impurities. The two-step reaction sequence, starting with ZnS, is:

$$2ZnS + 3O_2 \rightarrow 2ZnO + 2SO_2$$

$$ZnO + CO \rightarrow Zn + CO_2$$

To calculate the theoretical yield, you will need to calculate the number of moles of zinc produced in the two-step reaction. The strategy is to obtain an overall reaction, which 'sums' the individual steps and excludes any 'intermediates'. Although zinc oxide (ZnO) is produced in the first step, it is then consumed in the second step. Zinc oxide is an intermediate and does not appear among the final products for the overall reaction. To exclude an intermediate from the overall reaction, equations for the individual steps must be added in a specific way so that the intermediates cancel. The coefficients of intermediates need to be equalised, just as you would solve simultaneous equations in algebra. The coefficient of ZnO is '2' in the first step, but just '1' in the second step. If the second step is multiplied by 2 and added to the first step, this equation is obtained:

$$2ZnS + 3O_2 + 2ZnO + 2CO \rightarrow 2ZnO + 2SO_2 + 2Zn + 2CO_2$$

As in an algebraic equation, the 2ZnO appearing on both sides of the equation can be cancelled.

The overall reaction is shown in the following equation, which indicates that, in theory, one mole of zinc can be obtained for every mole of ZnS:

$$2ZnS + 3O_2 + 2CO \rightarrow 2SO_2 + 2Zn + 2CO_2$$

For example, beginning with 1 mole of ZnS (97.5 g), the theoretical yield would be 1 mole of Zn; that is, 65.4 g. This is the basis for a calculation of experimental yield, as shown in Worked example 12.4.2.

WORKED EXAMPLE 12.4.2

QUESTION

If 50 kg of ZnS yielded 29.5 kg of pure zinc (Zn), what is the experimental yield for the process?

ANSWER

1 Refer to the overall reaction for the process to calculate the theoretical yield.
2 mol of ZnS yields 2 mol of Zn in the overall reaction.
Therefore, 1 mole of ZnS would produce 1 mole of Zn metal.

2 Calculate the molar mass of ZnS.
The atomic weight of Zn is 65.4. The atomic weight of S is 32.1.
The molar mass of ZnS = 65.4 + 32.1 = 97.5 g mol^{-1}

3 Convert mass of ZnS to moles of ZnS:

$$\text{50 kg of ZnS contains } \frac{50 \times 10^3}{97.5} \text{ moles} = 513 \text{ moles}$$
$$= \text{number of moles of Zn metal produced}$$
$$\text{The mass of Zn produced} = 513 \times 65.4 \text{ g}$$
$$= 33\,538 \text{ g}$$

The theoretical yield for the two-step reaction of 50 kg of ZnS is 33.5 kg of Zn.

4 Calculate actual (experimental) % yield:

$$\text{Experimental \% yield} = \frac{29.5}{33.5} \times \frac{100}{1} = 88\%$$

The experimental yield is 88%.

In industrial chemistry, including pharmaceutical chemistry, multistep reactions are often used in producing end products. It helps greatly to have as high a percentage yield as possible in each step. This is because the overall percentage yield will be given by the product of the percentage yields for each step. For example, consider a five-step process, each with only 50% yield. The fractional yield for each step is 0.5. Multiplying 0.5 by itself five times is given by $(0.5)^5 = 0.031$ = overall fractional yield. This corresponds to an overall 3.1% percentage yield. Compare this with the case if each step had a 95% yield. An overall fractional yield of $(0.95)^5 = 0.77$ would be obtained, which corresponds to an overall 77% yield. The latter is obviously a preferable outcome!

SECTION REVIEW 12.4

REMEMBERING

1 Define 'limiting reagent'.

2 What is the theoretical or ideal yield of a reaction?

3 Explain why the experimental yield for a reaction is usually less than the theoretical yield.

UNDERSTANDING

4 The combustion of ethanol yields carbon dioxide and water.

$$C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$$

If the reaction is carried out in air:

a what reactant will be in excess?

b what is the limiting reagent?

c what is the molar ratio between the reactant ethanol and the product carbon dioxide?

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APPLYING

5 For the combustion of butane in air to yield CO_2 and water:

- **a** write a balanced chemical reaction
- **b** identify the limiting reagent
- **c** calculate the number of moles of O_2 needed for each mole of butane reacted
- **d** calculate the number of grams of CO_2 produced when 6 g of butane undergoes combustion in air
- **e** calculate the number of grams of H_2O produced when 6 g of butane undergoes combustion in air
- **f** determine the experimental % yield if 5.5 g water is collected in a laboratory experiment from the combustion of 6 g of butane.

ANALYSING

6 Dimethylhydrazine ($(CH_3)_2NNH_2$) was used as a fuel for the Apollo Lunar Descent Module, with N_2O_4 being used as the oxidant. The products of the reaction are H_2O, N_2 and CO_2.

- **a** Write a balanced chemical equation for the combustion reaction.
- **b** If 150 kg of $(CH_3)_2NNH_2$ react with 460 kg of N_2O_4, what is the theoretical yield of N_2?
- **c** If a 30 kg yield of N_2 gas represents a 68% yield, what mass of N_2O_4 would have been used up in the reaction?

7 Adipic acid ($C_6H_{10}O_4$,) is a raw material for making nylon and it can be prepared in the laboratory by the following reaction between cyclohexene (C_6H_{10}) and sodium dichromate ($Na_2Cr_2O_7$) in sulfuric acid.

$$3C_6H_{10}(l) + 4Na_2Cr_2O_7(aq) + 16H_2SO_4(aq) \rightarrow 3C_6H_{10}O_4(aq) + 4Cr_2(SO_4)_3(aq) + 4Na_2SO_4(aq) + 16H_2O$$

Side reactions occur. These, plus losses of product during its purification, reduce the overall yield. A typical yield of purified adipic acid is 68.6%.

- **a** Identify how many grams of cyclohexene are required to prepare 12.5 g of adipic acid in 68.6% yield.
- **b** The only available supply of sodium dichromate is its dihydrate ($Na_2Cr_2O_7.2H_2O$). (The reaction occurs in an aqueous medium, so the water in the dihydrate does not cause a problem, but does mean that you will need a larger mass of the dihydrate form of potassium dichromate ($K_2Cr_2O_7 \cdot 2H_2O$) than if you were able to obtain just $K_2Cr_2O_7$). Identify how many grams of this dihydrate are also required in the preparation of 12.5 g of adipic acid, in a yield of 68.6%.

12.5 Inquiry skills: mathematical representations, Part I

Converting between moles and number of particles

Knowing the number of particles (atoms, ions or molecules) in a sample enables calculation of the number of moles (n) in that sample.

WORKED EXAMPLE 12.5.1

How many moles of magnesium are there in 1.45×10^{23} atoms of magnesium?

ANSWER

To convert from particles to moles, divide the number of particles given by N_A:

$$n(\text{Mg}) = \frac{1.45 \times 10^{23}}{6.02 \times 10^{23}} = 0.24 \text{ mol}$$

The number of moles of magnesium is 0.24 mol.

WORKED EXAMPLE 12.5.2

Knowing the number of moles (n) in a sample of a substance allows you to calculate the number of particles (atoms, ions or molecules) in that sample.

QUESTION

How many atoms are there in 0.25 mol of neon atoms?

ANSWER

To convert from moles to particles, multiply the number of moles given by N_A:

$$\begin{aligned}\text{Number of atoms of Ne} &= 0.25 \times 6.02 \times 10^{23} \\ &= 1.505 \times 10^{23}\ \text{atoms}\end{aligned}$$

There are 1.5×10^{23} neon atoms.

Moles and chemical formulas

To calculate the number of moles of individual atoms in a molecule, use the relationship between moles and the chemical formula of a substance.

WORKED EXAMPLE 12.5.3

1 Calculate the number of moles of carbon atoms and oxygen atoms in 5 mol of carbon dioxide (CO_2).

2 How many moles of atoms are present in total?

3 How many individual atoms are present?

ANSWERS

1 Using the formula, you can observe that 1 mol of CO_2 contains 1 mol C and 2 mol O.
For 5 mol CO_2:

$$n(\text{C}) = 5 \times 1\ \text{mol C} = 5\ \text{mol of C}$$
$$n(\text{O}) = 5 \times 2\ \text{mol O} = 10\ \text{mol of O}$$

Therefore, there are 5 mol of C and 10 mol of O.

2 The total number of moles of atoms is the sum of number of moles of individual atoms:

$$n(\text{CO}_2) = 5 + 10 = 15\ \text{mol of atoms}$$

Therefore, there are 15 mol of atoms.

3

$$\text{Using } n = \frac{\text{number of particles}}{6.02 \times 10^{23}}$$

$$\begin{aligned}\text{Total number of individual atoms} &= 15 \times 6.02 \times 10^{23} \\ &= 9.03 \times 10^{24}\end{aligned}$$

Therefore, there are 9.03×10^{24} individual atoms present.

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Moles and chemical reactions

If the number of moles of any species in the reaction are known, then use the balanced equation to work out the number of moles of all the other species in that reaction.

WORKED EXAMPLE 12.5.4

Propane gas (C_3H_8) burns in oxygen to produce carbon dioxide and water. If 2.5 mol propane is burnt, calculate the number of moles of:

- oxygen gas needed for complete combustion
- carbon dioxide produced
- water produced.

ANSWERS

1 Write a balanced equation and determine the reaction ratio from the coefficients:

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$
$$1\text{ mol} + 5\text{ mol} \rightarrow 3\text{ mol} + 4\text{ mol}$$

2 Use the ratio to calculate the number of moles from the reacted amount given.

a 1 mol of C_3H_8 reacts with 5 mol of O_2.

$$n(O_2) = \frac{5}{1} \times n(C_3H_8)$$
$$= \frac{5}{1} \times 2.5 = 12.5 \text{ mol of } O_2$$

12.5 mol of O_2 are needed.

b 1 mol of C_3H_8 produces 3 mol of CO_2.

$$n(CO_2) = \frac{3}{1} \times n(C_3H_8)$$
$$= \frac{3}{1} \times 2.5 = 7.5 \text{ mol of } CO_2$$

7.5 mol of CO_2 are produced.

c 1 mol of C_3H_8 produces 4 mol of H_2O.

$$n(H_2O) = \frac{4}{1} \times n(C_3H_8)$$
$$= \frac{4}{1} \times 2.5 = 10 \text{ mol of } H_2O$$

10 mol of H_2O are produced.

SECTION REVIEW 12.5

APPLYING

1 Calculate the number of moles of each of the following substances, given the number of particles.

  a 3.3×10^{25} carbon atoms
  b 1.2×10^{22} molecules of water
  c 6.6×10^{21} sodium ions
  d 3.01×10^{24} formula units of calcium carbonate
  e 9.05×10^{23} molecules of hydrogen gas

2 Calculate the number of particles in each of the following substances, given the number of moles.

  a 0.1 mol of methane molecules
  b 6.0 mol of hydrogen atoms
  c 0.5 mol of aluminium ions
  d 0.082 mol of argon gas
  e 2.1 mol of ammonia molecules

3 In each of the following, specify how many:

  a moles of K^+ ions and Cl^- ions there are in 3 mol of KCl
  b moles of H atoms there are in 6 mol of $C_6H_{12}O_6$
  c moles of atoms there are in 0.5 mol of ammonia (NH_3)
  d individual oxygen atoms there are in 2.5 mol of calcium carbonate ($CaCO_3$)
  e individual atoms there are in 1.25 mol of ethanol (C_2H_5OH).

4 a If 4.5 mol of magnesium were burned in oxygen gas, calculate the number of moles of:
    i oxygen required for complete combustion of the magnesium
    ii magnesium oxide produced.

  b Iron reacts with hydrochloric acid (HCl) to produce iron(II) chloride and hydrogen gas:

$$Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$$

    i Identify how many moles of Fe are needed to completely react with 1.5 mol of HCl.
    ii Identify how many moles of H_2 are produced when 3 mol of HCl react.
    iii If 4 mol of $FeCl_2$ are produced, how many moles of Fe react?

  c Butane (C_4H_{10}) undergoes combustion in the presence of oxygen gas to produce carbon dioxide and water. If 5.5 mol of butane react, calculate the number of moles of:
    i oxygen consumed
    ii carbon dioxide produced
    iii water produced.

12.6 Inquiry skills: mathematical representations, Part II

Relative mass of substances

The relative molecular mass of an element or compound is calculated by adding the relative atomic masses of all the component atoms of the molecule.

WORKED EXAMPLE 12.6.1

Calculate the relative molecular mass of ethane (C_2H_6).

ANSWER

1 Identify the number of each type of atom:

Number of C atoms = 2
Number of H atoms = 6

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2 Multiply the number of each type of atom by their relative atomic mass (from the periodic table):

$$\text{Mass of C} = 2 \times A_r(\text{C}) = 2 \times 12.0 = 24.0$$
$$\text{Mass of H} = 6 \times A_r(\text{H}) = 6 \times 1.0 = 6.0$$

3 Add the masses of each element to obtain the relative mass of the compound:

$$M_r(C_2H_6) = 24.0 + 6.0 = 30.0$$

This can be simplified to:

$$M_r(C_2H_6) = 2 \times A_r(\text{C}) + 6 \times A_r(\text{H}) = 2 \times 12.0 + 6 \times 1.0 = 30.0$$

Therefore, the relative molecular mass is 30.0.

The relative formula mass of a compound (M_r) is the sum of the relative atomic masses of the atomic species as given by the formula of the compound.

WORKED EXAMPLE 12.6.2

Calculate the relative formula mass of calcium hydroxide ($Ca(OH)_2$).

ANSWER

1 Identify the number of each type of atom:

Number of Ca atoms = 1
Number of H atoms = 2
Number of O atoms = 2

2 Multiply the number of each type of atom by its relative atomic mass (from the periodic table):

$$\text{Mass of Ca} = 1 \times A_r(\text{Ca}) = 1 \times 40.1 = 40.1$$
$$\text{Mass of H} = 2 \times A_r(\text{H}) = 2 \times 1.0 = 2.0$$
$$\text{Mass of O} = 2 \times A_r(\text{O}) = 2 \times 16.0 = 32.0$$

3 Add the masses of each element to obtain the mass of the compound:

$$M_r(Ca(OH)_2) = 40.1 + 2.0 + 32.0 = 74.1$$

This can be simplified to:

$$\begin{aligned} M_r(Ca(OH)_2) &= 1 \times A_r(\text{Ca}) + 2 \times A_r(\text{H}) + 2 \times A_r(\text{O}) \\ &= 1 \times 40.1 + 2 \times 1.0 + 2 \times 16.0 \\ &= 74.1 \end{aligned}$$

The relative formula mass is 74.1.

Percentage composition, empirical formula

The percentage composition of a compound is the percentage by mass of each of the different elements in the compound.

WORKED EXAMPLE 12.6.3

Calculate the percentage composition of aluminium oxide (Al_2O_3).

ANSWER

1 Calculate the relative mass of each element in the compound:

$$\text{Mass of Al} = 2 \times 27.0 = 54.0$$
$$\text{Mass of O} = 3 \times 16.0 = 48.0$$

2 Calculate the relative formula (or molecular) mass:

$$\begin{aligned} \text{Formula mass of } Al_2O_3 &= \text{mass of Al} + \text{mass of O} \\ &= 54.0 + 48.0 \\ &= 102.0 \end{aligned}$$

3 Calculate the % of each element:

$$\begin{aligned} \%\ \text{Al} &= \frac{\text{mass of Al in compound}}{\text{formula mass of } Al_2O_3} \times 100 \\ &= \frac{54.0}{102.0} \times 100 \\ &= 52.9\% \end{aligned}$$

$$\begin{aligned} \%\ \text{O} &= \frac{\text{mass of Al in compound}}{\text{formula mass of } Al_2O_3} \times 100 \\ &= \frac{48.0}{102.0} \times 100 \\ &= 47.1\% \end{aligned}$$

The percentage composition is 52.9% Al and 47.1% O.

The empirical formula of a substance can be determined experimentally. You can do this by experimentally determining the percentage composition and then applying the mole concept to determine the ratio of particles, as shown in Worked example 12.6.4.

WORKED EXAMPLE 12.6.4

An unidentified compound was analysed in the laboratory and found to have the following percentage composition by mass: 26.09% C, 4.35% H and 69.56% O.

1 What is its empirical formula?

2 If the relative molecular mass of the compound is 92, what is its molecular formula?

ANSWERS

1 a Assuming 100 g of the compound, determine the mass of each element:

ELEMENT	C	H	O
% (GIVEN)	26.09	4.35	69.56
MASS (IN 100 g)	26.09	4.35	69.56

b Find the number of moles of each element by dividing mass by the molar mass:

ELEMENT	C	H	O
MASS (IN 100 g)	26.09	4.35	69.56
MOLES (n)	$\frac{26.09}{12.0} = 2.17$	$\frac{4.35}{1.01} = 4.31$	$\frac{69.56}{16.00} = 4.35$

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c Find the simplest whole number ratio by dividing the number of moles by the smallest number then rounding to a whole number:

ELEMENT	C	H	O
MASS (IN 100 g)	26.09	4.35	69.56
MOLES (n)	$\frac{26.09}{12.0} = 2.17$	$\frac{4.35}{1.01} = 4.31$	$\frac{69.56}{16.00} = 4.35$
SIMPLEST RATIO	$\frac{2.17}{2.17} = 1$ $= 1$	$\frac{4.31}{2.17} = 1.99$ $= 2$	$\frac{4.35}{2.17} = 2$ $= 2$

The empirical formula is CH_2O_2.

2 **a** Relative empirical formula mass $= 1 \times 12.01 + 2 \times 1.01 + 2 \times 16.00$

$$= 12.01 + 2.02 + 32.00$$
$$= 46.03$$

b To calculate the molecular formula, determine how many empirical units there are in the molecule:

$$\frac{\text{Relative molecular mass}}{\text{Relative empirical mass}} = \frac{92}{46.03}$$
$$= 2$$

This means that molecular formula contains two empirical formula units.

c Molecular formula $= 2 \times CH_2O_2$

$$= C_2H_4O_4$$

The molecular formula is $C_2H_4O_4$.

SECTION REVIEW 12.6

APPLYING

1 Calculate the relative molecular mass of:

- **a** phosphorus (P_4)
- **b** carbon dioxide (CO_2)
- **c** nitric acid (HNO_3)
- **d** glucose ($C_6H_{12}O_6$).

2 Calculate the relative formula mass of:

- **a** aluminium nitride (AlN)
- **b** barium nitrate ($Ba(NO_3)_2$)
- **c** sodium carbonate (Na_2CO_3)
- **d** copper(II) sulfate ($CuSO_4$).

3 Calculate the percentage composition of each element in:

- **a** methane (CH_4)
- **b** sodium chloride (NaCl)
- **c** calcium cyanide ($Ca(CN)_2$)
- **d** ammonium oxalate ($(NH_4)_2C_2O_4$).

4 **a** Find the empirical formula of the following compounds, given their percentage composition by mass:

- **i** 7.8% C, 92.2% Cl
- **ii** 64.8% C, 13.6% H, 21.6% O
- **iii** 32.0% C, 6.7% H, 42.6% O, 18.7% N.

b Identify the molecular formula of the compound in part **a ii** if its relative molecular mass is 74.12.

12.7 Inquiry skills: experimental and theoretical yield

Moles and mass

To determine the theoretical yield of a product formed in a chemical reaction starting with a given mass of reactants, make use of a specific relationship: 1 mole of any substance has a mass equal to the relative atomic, molecular or formula mass of that substance expressed in grams. This is shown in Worked example 12.7.1.

WORKED EXAMPLE 12.7.1

Calculate the mass of 1 mole (molar mass) of ethane (C_2H_6).

ANSWER

The relative molecular mass of ethane is calculated from:

$$M = 2 \times m(\text{C}) + 6 \times m(\text{H}) = 2 \times 12.0 + 6 \times 1.0 = 30.0$$

$$M(\text{C}_2\text{H}_6) = 30.0\,\text{g}\,\text{mol}^{-1}$$

The molar mass of ethane is $30.0\,\text{g}\,\text{mol}^{-1}$.

Converting between the number of moles and mass

If you know the amount of a given substance in grams, then you can calculate the number of moles and vice versa, as shown in Worked examples 12.7.2 and 12.7.3.

WORKED EXAMPLE 12.7.2

How many moles of copper(II) sulfate ($CuSO_4$) are there in 12.2 g of copper(II) sulfate?

ANSWER

1 Write the relationship of copper(II) sulfate:

$$n = \frac{m}{M}$$

n =?, m = 12.2 g, M can be calculated

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2 Calculate the molar mass of copper(II) sulfate:

$$M(CuSO_4) = M(Cu) + M(S) + 4 \times M(O) = 63.6 + 32 + 4 \times 16$$
$$= 159.6 \text{ g mol}^{-1}$$

3 Substitute:

$$n(CuSO_4) = \frac{12.2}{159.6}$$
$$= 0.0764 \text{ mol}$$

There are 0.0764 mol in 12.2 g of copper(II) sulfate.

WORKED EXAMPLE 12.7.3

What is the mass of 0.75 mol of sodium hydroxide (NaOH)?

ANSWER

1 Write the relationship:

$$n = \frac{m}{M}$$

$n(NaOH) = 0.75$ mol, $m = ?$, M can be calculated

2 Calculate the molar mass of sodium hydroxide:

$$M(NaOH) = 23 + 16 + 1 = 40 \text{ g mol}^{-1}$$

3 Rearrange to find m:

$$m = n \times \mathrm{M}$$

4 Substitute and solve:

$$m(NaOH) = 0.75 \times 40$$
$$= 30 \text{ g}$$

There are 30 g of sodium hydroxide in 0.75 mol.

Calculating yields in chemical reactions

The mass relationships between reactants and products are derived from the mole ratios in a balanced chemical equation. All but one of the reagents may be in excess. Always identify the limiting reagent if other reagents are in excess. Alternatively, if the quantities of reactants are present in exactly the right stoichiometric ratio, then each of the reactants will be available in precisely sufficient quantity for them to all be used up completely at the end of the reaction.

WORKED EXAMPLE 12.7.4

Sulfuric acid is formed when sulfur dioxide (SO_2) and water react with oxygen in the presence of a catalyst.

1 If 3.20 g of SO_2 react, calculate the mass of:

a O_2 that reacts

b sulfuric acid (H_2SO_4) that is produced.

2 If 4.0 g of H_2SO_4 is obtained in an experiment from 3.2 g SO_2, determine the experimental % yield of sulfuric acid produced.

ANSWERS

1 Write a balanced equation:

$$2SO_2(g) + O_2(g) + 2H_2O(l) \rightarrow 2H_2SO_4(l)$$

2 Calculate the number of moles of SO_2 using the relationship: $n = \frac{m}{M}$:

$m = 3.20$ g, $n =$?, M can be calculated.
$M = 32.1 + 2 \times 16 = 64.1\ \text{g mol}^{-1}$

$$n = \frac{3.20}{64.1} = 0.0499\ \text{mol} = 0.05\ \text{mol}$$

3 Use the balanced chemical equation to calculate the number of moles of oxygen gas and sulfuric acid.

a From the equation, 2 mol of SO_2 reacts with 1 mol of O_2:

$$n(O_2) = \frac{1}{2} \times n(SO_2)$$

$$n(O_2) = \frac{1}{2} \times 0.05 = 0.025\ \text{mol of } O_2$$

b 2 mol of SO_2 produces 2 mol of H_2SO_4:

$$n(H_2SO_4) = \frac{2}{2} \times n(SO_2)$$

$$n(H_2SO_4) = \frac{2}{2} \times 0.05 = 0.05\ \text{mol of } H_2SO_4$$

4 Calculate the mass of the required substance: $m = n \times M$.

a Mass of O_2:

$n = 0.025$ mol, $m =$?, M can be calculated
$M = 2 \times 16 = 32\ \text{g mol}^{-1}$
$m = 0.025 \times 32 = 0.8$ g

The mass of oxygen reacted is 0.8 g.

b Mass of H_2SO_4:

$n = 0.05$ mol, $m =$?, M can be calculated
$M = 2 \times 1.01 + 32.1 + 4 \times 16 = 98.13\ \text{g mol}^{-1}$
$m = 0.05 \times 98.13 = 4.9$ g

The mass of sulfuric acid produced is 4.9 g.

c $$\%\ \text{yield} = \frac{4.0}{4.9} \times \frac{100}{1} = 82\%$$

The experimental percentage yield is 82%.

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SECTION REVIEW

12.7

APPLYING

1 a Calcium carbonate decomposes when heated strongly. This equation represents the decomposition:

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

If 25 g of calcium carbonate decomposes, calculate the mass of:

i CaO produced

ii CO_2 produced.

b If an experiment produces 10.5 g of CaO, calculate the experimental % yield.

2 a Zinc reacts with hydrochloric acid to form zinc chloride and hydrogen gas.

If 2.6 g of zinc reacts with excess HCl, calculate the theoretical values for:

i the mass of zinc chloride produced

ii the mass of hydrogen gas produced.

b If this experiment produces 5.2 g of zinc chloride, calculate the experimental % yield.

3 a Octane (C_8H_{18}) is a one of the many hydrocarbons that are components of petrol. It burns in air to produce carbon dioxide and water. Write the balanced equation for the reaction.

If 45.6 kg of octane (this is about 65 L of fuel) is burned, calculate the mass of:

i oxygen that reacts.

ii carbon dioxide produced.

b If an experiment where 45.6 kg of octane burns in air produces 105.6 kg of carbon dioxide, calculate the experimental % yield.

Explain why the yield is less than 100%.

Comment on the amount of carbon dioxide produced by a tank of petrol (assume 100% octane) and the consequence for greenhouse gas emissions. Do this by estimating the average number of kilometres driven per year by a motorist, a fuel economy of 10 L/100 km, and work out how much carbon dioxide they produce annually (in tonnes).

4 In the wine industry, yeast ferments glucose ($C_6H_{12}O_6$) from crushed grapes to produce ethanol (C_2H_5OH) and carbon dioxide.

a Write a balanced overall equation for the production of ethanol and CO_2 from glucose to determine the number of moles of ethanol produced per mole of glucose.

b In a small scale laboratory experiment, if 45 g of glucose are fermented, calculate the theoretical mass of ethanol produced.

c If only 13.8 g of ethanol are produced, calculate the experimental % yield and offer reasons why a 100% yield is not likely in the fermentation of crushed grapes in winemaking.

12.8 Mandatory practicals

PRACTICAL ACTIVITY 12.8.1

Measuring molar quantities

In the chemistry laboratory, the amount of substance is usually measured in mass and sometimes in volume. If there is the same number of moles of different substances, they will have different masses because the substances are made up of different types of molecules (or atoms). Different molecules will have different atomic compositions and different molar masses.

AIM

To determine the number of moles in a measured quantity of substance

EQUIPMENT PER GROUP

- sulfur powder
- water
- aluminium foil
- copper wire
- sodium chloride
- copper(II) sulfate ($CuSO_4$)
- sucrose
- carbon powder
- electronic scales
- measuring cylinder
- eyedropper
- spatula
- 2×50 mL beakers

RISK ASSESSMENT

WHAT ARE THE RISKS IN DOING THIS EXPERIMENT?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?
Copper(II) sulfate is toxic.	Wear appropriate safety gear, clean up spills and dispose of chemicals as directed.
Powdered sulfur is an irritant to eyes, nose and throat.	Do not agitate the powder when measuring it.

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, as well as ways to manage them.

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PROCEDURE

PART A

1 Calculate the molar mass of each of the substances provided.
2 Measure 1 mole of carbon powder, water and sodium chloride.
3 Compare the apparent amounts of each of these substances.

PART B

Use your understanding of the mole to solve each of the following problems. List the steps you followed to solve each problem and show the relevant calculations.

1 How many moles of carbon are in 1 level teaspoon of carbon?
2 How many atoms of copper are in a 1 cm piece of copper wire?
3 How many atoms of aluminium are in a 2 cm^2 piece of aluminium foil?
4 What contains the greater number of moles – a level teaspoon of sugar or a level teaspoon of copper(II) sulfate?
5 How many molecules are there in 1 drop of water?

PART C

Collate the class results for each problem in **part B**.

DISCUSSION

1 Compare your results with those of other groups in the class and suggest possible reasons for any variation.
2 How accurate do you consider the results to be? Support your answer with experimental data.
3 What assumptions were made about the samples provided?
4 Suggest how the accuracy of the experiment could be improved.

PRACTICAL ACTIVITY 12.8.2

Determining the empirical formula of magnesium oxide

The empirical formula of a substance can be calculated from experimental data. In this experiment, you will carry out the reaction:

$$\text{Magnesium} + \text{oxygen} \rightarrow \text{magnesium oxide}$$

By determining the masses of each element and using the relationship:

$$\text{Number of moles} = \frac{\text{mass}}{\text{molar mass}}$$

the number of moles of each element can be determined. The empirical formula is derived from this.

AIM

Write an aim for this experiment.

EQUIPMENT PER GROUP

- 1 strip of magnesium ribbon (about 20 cm)
- steel wool
- crucible tongs
- pipe clay triangle
- Bunsen burner
- tripod
- balance
- crucible and lid
- matches

RISK ASSESSMENT

WHAT ARE THE RISKS IN DOING THIS EXPERIMENT?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?
Burning magnesium emits a bright light.	Do not look directly at burning magnesium.
Magnesium oxide is an irritant to eyes, nose and throat.	Lift only the edge of the crucible lid. Work in a well-ventilated area.

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, as well as ways to manage them.

PROCEDURE

1. Heat the crucible and lid strongly for 5 minutes over a hot Bunsen flame. Allow it to cool.
2. Weigh the crucible and lid.
3. Thoroughly clean the surface of the magnesium ribbon with steel wool.
4. Record the appearance of the cleaned magnesium.
5. Coil the piece of magnesium ribbon so that it fits inside the crucible. Weigh the crucible, lid and magnesium ribbon.
6. Place the crucible on a pipe clay triangle on a tripod over the Bunsen burner and carefully heat the crucible without the lid until the magnesium begins to glow. Use tongs to then put the lid on before the magnesium bursts into flames. (Warning: Do not look directly at the burning magnesium.)
7. Once the lid is in place, heat the crucible strongly for about 10 minutes, lifting the lid occasionally to admit oxygen. Try to prevent any magnesium oxide smoke escaping.
8. Remove the lid and heat the crucible for a further 5 minutes to ensure complete reaction.
9. Replace the lid and allow the reaction mixture to cool. Reweigh the crucible, lid and magnesium oxide.

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RESULTS

Record your results in a results table similar to the one below.

ITEMS WEIGHED	MASS (g)
Mass of crucible and lid	
Mass of crucible, lid and magnesium	
Mass of magnesium	
Mass of crucible, lid and magnesium oxide	
Mass of magnesium oxide	

ANALYSING THE RESULTS

1 Calculate the number of moles of magnesium and oxygen.
2 State the ratio of moles of magnesium to moles of oxygen.
3 Using the mole ratio, determine the empirical formula of magnesium oxide.

DISCUSSION

1 Compare your results with those of other groups in the class and suggest possible reasons for any variations.
2 Consider the sources of error in this experiment and suggest how the accuracy of the results might be improved.

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 Define the following terms.

- **a** Atomic mass unit
- **b** Avogadro constant
- **c** Law of conservation of mass
- **d** Mole
- **e** Molar mass
- **f** Percentage composition
- **g** Relative formula mass
- **h** Relative molecular mass
- **i** Stoichiometry

CATEGORY QUESTIONS

2 Explain the relationship between relative atomic mass and relative molecular mass.

3 Distinguish between relative molecular mass and relative formula mass.

4 Outline how percentage composition provides useful information to chemists.

5 Explain the relationship between the Avogadro constant and moles.

6 Distinguish between relative molecular mass and molar mass.

7 **a** Distinguish between empirical formula and molecular formula.

b Is it correct to refer to the formula of an ionic compound as the empirical formula? Explain why or why not.

8 Identify this law: In a chemical reaction, mass is neither created nor destroyed.

ELABORATION QUESTIONS

9 Calculate the relative molecular or formula mass of:

- **a** zinc hydroxide ($Zn(OH)_2$)
- **b** ethanoic (acetic) acid (CH_3COOH).

10 Calculate the percentage composition of:

- **a** calcium in calcium hydroxide ($Ca(OH)_2$)
- **b** oxygen in aluminium carbonate ($Al_2(CO_3)_3$).

11 Calculate the number of moles in:

- **a** 3.01×10^{24} carbon atoms
- **b** 1.2044×10^{28} ethanoic (acetic) acid molecules (CH_3COOH).

12 Calculate the mass of:

- **a** 2.05 mol of rubidium atoms
- **b** 0.5 mol of glucose ($C_6H_{12}O_6$).

EVIDENCE QUESTIONS

13 Urea has the formula CH_4ON_2. In 2 mol of urea, calculate the number of:

- **a** moles of H atoms
- **b** moles of O atoms
- **c** individual atoms of N.

14 In an experiment to find the formula of a compound, copper(II) oxide was heated until only pure metallic copper remained. The following results were obtained:

- mass of crucible = 35.030 g
- mass of crucible + copper = 38.205 g
- mass of crucible + copper(II) oxide = 39.005 g

Find the empirical formula of the compound.

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15 The mass of the bones in an average adult is about 11 kg. Calcium phosphate ($Ca_3(PO_4)_2$) makes up 50% of this mass. What is the mass of phosphorus in the bones of an average adult?

16 Identify the mass of zinc that has the same number of atoms as 1.00 g of hydrogen gas.

17 Elemental analysis of a compound of iron gave the following percentages by mass:

Fe 20.09%, S 11.55%, H 5.04%, O 63.32%.

a Calculate the empirical formula for the compound.

b If all the hydrogen comes from the water of crystallisation, determine how many molecules of water there are in a formula unit of this compound. Deduce the formula of the compound in the usual hydrate form (for example, $Na_2CO_3{\cdot}10H_2O$).

18 A silver coin was analysed and found to contain only silver and copper. A mass of 1.580 g of this coin was dissolved in concentrated nitric acid and the resultant solution was diluted. Reaction of the solution with excess hydrochloric acid yielded 1.050 g of silver chloride. Calculate the percentage of silver in the sample taken from the coin.

END-OF-CHAPTER EXAM

End-of-chapter test

1 Which of the following is closest to the relative molecular mass of carbon dioxide?

A $28\,g\,mol^{-1}$

B $44\,g\,mol^{-1}$

C $14\,g\,mol^{-1}$

D $22\,g\,mol^{-1}$

2 The relative mass of silver chloride is more correctly called:

A relative atomic mass.

B relative molecular mass.

C relative formula mass.

D relative compound mass.

3 How many feathers are there in 2 moles of feathers (correct to 2 significant figures)?

A 6.0×10^{23}

B 1.2×10^{24}

C Need to know average mass of feathers

D 3.0×10^{23}

4 The mass of 1 mole of nitrogen dioxide is closest to:

A 6.02×10^{23}.

B 30 g.

C 46 g.

D 44 g.

5 What is required to calculate the percentage composition of a substance?

A Chemical formula and relative atomic masses of the elements in the compound

B Chemical formula only

C Chemical formula and relative mass of the compound

D Relative mass of the compound only

6 The Avogadro constant is equal to:

A 6.02×10^{23} atoms.

B 6.02×10^{23} g.

C 12.01 atoms.

D 12.01 g.

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7 Which of the following describes the mass of Avogadro's number of particles?

A 35 g of methane

B 40 g of lithium fluoride

C 16 g of oxygen gas

D 36.5 g of hydrogen chloride

8 How many moles of water are produced when 0.03 moles of aluminium hydroxide decompose according to the following equation?

$$2Al(OH)_3(s) \longrightarrow Al_2O_3(s) + 3H_2O(l)$$

A 0.045 mol

B 0.03 mol

C 0.06 mol

D 0.015 mol

9 Which of the following is both an empirical and a molecular formula?

A H_2O_2

B $C_6H_{12}O_6$

C C_2H_4

D $C_{12}H_{22}O_{11}$

10 The total number of particles in 4.5 g of KCl is:

A 6.02×10^{23} particles.

B 3.6×10^{22} particles.

C 0.06 particles.

D 2.7×10^{24} particles.

11 How many moles of aluminium are in 0.15 mol $Al_2(SO_4)_3$?

A 0.3 mol

B 0.15 mol

C 0.45 mol

D 1.0 mol

12 The coefficients of a chemical equation can represent:

A mole ratios.

B particle ratios.

C mass ratios.

D both mole ratios and particle ratios.

13 The ratio of Mg:P:O atoms in $Mg_3(PO_4)_2$ is:

A 1:2:2.

B 1:3:3.

C 3:1:4.

D 3:2:8.

14 1 mole of chlorine gas and 1 mole of lithium chloride are weighed. The total mass is:

A 77.8 g.

B 113.3 g.

C 2 moles.

D 1.2×10^{24} particles.

15 Determine the percentage composition of iron(II) sulfate.

16 Respiration can be represented by the following equation:

$$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O + \text{energy}$$

a Calculate the molecular mass of glucose.

b How many moles of carbon dioxide are produced per mole of glucose?

c How many particles of water are produced per mole of oxygen?

17 Hydrogen sulfide gas reacts with oxygen gas to form sulfur dioxide gas and water according to the equation:

$$\_\_\_\_ H_2S + \_\_\_\_ O_2 \rightarrow \_\_\_\_ SO_2 + \_\_\_\_ H_2O$$

a Balance the equation by writing the correct coefficients in the spaces above.

b Calculate the mass of sulfur dioxide produced from 1.0 kg of hydrogen sulfide.

18 An unknown substance has the following percentage composition by mass:

49.4% C, 5.2% H, 16.6% O, 28.8% N.

a Determine the empirical formula of this compound.

b If the molar mass of the compound is 198.18 g mol^{-1}, determine the molecular formula.

19 A piece of zinc metal of mass 1.308 g is added to 2.0 g of hydrochloric acid. The following reaction takes place.

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

a Calculate the moles of each reactant present before the reaction takes place.

b One reactant is in excess (there is too much of it). Identify which reactant is in excess and by how many moles.

c Calculate the mass of hydrogen gas produced in this reaction.

20 Silver nitrate ($AgNO_3$) reacts with iron(III) chloride ($FeCl_3$) to give silver chloride (AgCl) and iron(III) nitrate ($Fe(NO_3)_3$). In a particular experiment, it was planned to mix a solution containing 25.0 g of $AgNO_3$ with another solution containing 45.0 g of $FeCl_3$.

- **a** Write the chemical equation for the reaction.
- **b** Which reactant is the limiting reactant?
- **c** What is the maximum number of moles of AgCl that could be obtained from this mixture?
- **d** What is the maximum number of grams of AgCl that could be obtained?
- **e** How many grams of the reactant in excess will remain after the reaction is over?

21 A chemist was required to make 100 g of chlorobenzene (C_6H_5Cl) from the reaction of benzene (C_6H_6) with chlorine and to expect a yield no higher than 65%. What is the minimum quantity of benzene that can give 100 g of chlorobenzene if the yield is 65%?

The equation for the reaction is:

$$\underset{\text{Benzene}}{C_6H_6} + Cl_2 \rightarrow \underset{\text{chlorobenzene}}{C_6H_5Cl} + HCl$$

» UNIT TWO

MOLECULAR INTERACTIONS AND REACTIONS

- Topic 1: Intermolecular forces and gases
- Topic 2: Aqueous solutions and acidity
- Topic 3: Rates of chemical reactions

Water is essential for life on Earth and, despite being a common substance, it is unusual compared with other compounds. It is an excellent solvent because of the hydrogen bonds it can form with solutes. The acidity or alkalinity of solutions can be measured using the pH scale, which also determines the reactions in which a solution will take part. It is important to understand how quickly these reactions will occur by measuring the rate of a chemical reaction, particularly when external conditions are varied, and to be able to represent and explain these changes.

UNIT OBJECTIVES

By the end of this unit, students should be able to:

1. describe and explain intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions
2. apply understanding of intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions
3. analyse evidence about intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions
4. interpret evidence about intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions
5. investigate phenomena associated with intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions
6. evaluate processes, claims and conclusions about intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions
7. communicate understandings, findings, arguments and conclusions about intermolecular forces and gases, aqueous solutions and acidity, and rates of chemical reactions.

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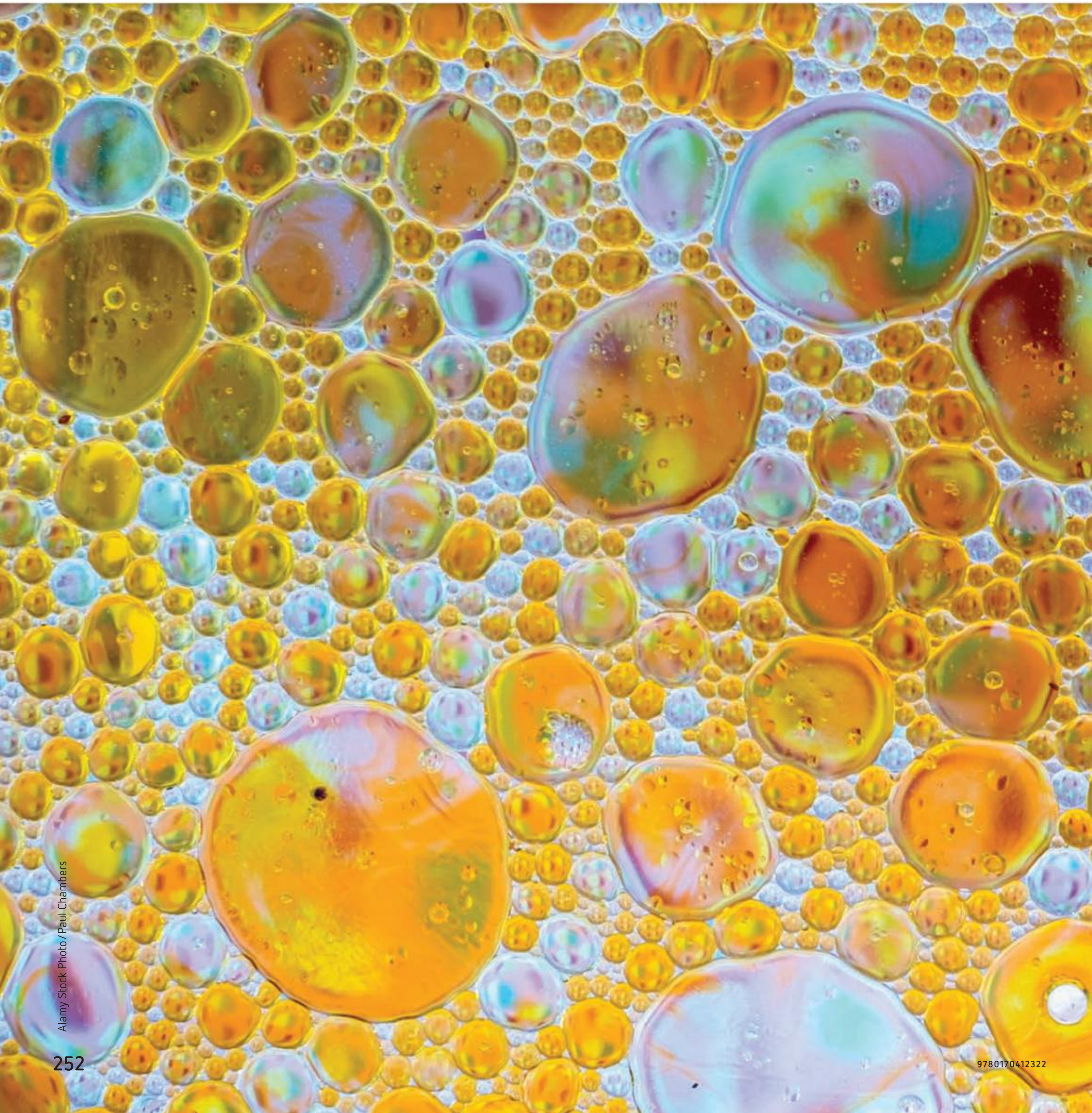
» UNIT TWO

MOLECULAR INTERACTIONS AND REACTIONS



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Topic 1: Intermolecular forces and gases

Molecular structure depends on the valence electrons in the atoms that are bonded together to form a stable molecule. Structure can be predicted successfully using the valence shell electron pair repulsion (VSEPR) theory) Furthermore, an understanding of the forces between molecules, (intermolecular forces) is necessary to predict how materials will behave and their properties. The properties of materials vary according to the type of atoms or molecules involved. Analytical techniques such as gas and High-performance liquid chromatography (HPLC) can be used to separate the atoms in a sample and determine their composition and purity. The **kinetic theory of gases** explains their behaviour and the quantitative relationships between temperature, volume and pressure.

SCIENCE AS A HUMAN ENDEAVOUR

Students should be given opportunities to investigate how forensic science employs the processes of chemistry, how a knowledge of chemistry is essential for safe scuba-diving practices and investigate the use of two- and three-dimensional models in the development and communication of VSEPR.

13 INTERMOLECULAR FORCES

Introduction

Chemical bonding and molecular structure are crucial to our understanding of the natural world. However, bonding between atoms is only part of the story. Scientists also need to understand the concept of intermolecular forces – the forces that act *between* molecules, binding them together and influencing their physical properties. Without this, the idea that lumps of solid water could float in liquid water would be unthinkable! Scientists would not be able to understand why oil and water do not mix. We would be shocked to see a balloon filled with radon gas falling to the ground faster than a lump of lead.

Stimulus questions

Why can you smell perfume from a bottle of scented oil faster if the bottle is warm rather than cold?

Why do some household powders, such as table salt and sugar, dissolve readily in water while others, such as plain flour, do not?

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13.1 Molecular structure: valence shell electron pair repulsion (VSEPR) theory

The shape of simple molecules can be predicted by inspection of the outer or valence electron pairs around the central atom of the molecule. These pairs may be bonding or non-bonding. In bonding pairs, electrons form a covalent bond with another atom. In non-bonding pairs, the electrons from the central atom have not formed covalent bonds.

Bonding electrons and non-bonding electrons (lone pairs)

Bonding electrons are electrons that are shared in a covalent bond between atoms. **Non-bonding electrons**, also known as lone pairs, are electrons belonging to the central atom that are not used in bonding.

bonding electrons electrons that are shared in a covalent bond between atoms

non-bonding electrons electrons belonging to the central atom that are not used in bonding; also known as lone pairs

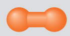
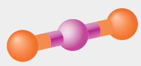
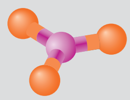
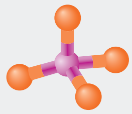
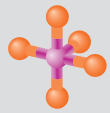
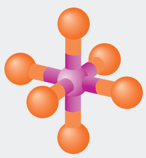
Predicting the shape of molecules

The pairs of electrons around the central atom, being negatively charged, repel each other, and try to get as far away as possible from each other. This repulsion of the pairs of electrons contributes to determining the shape of a molecule. This theory is known as the valence shell electron pair repulsion (VSEPR) theory. In this theory, multiple bonds are counted as one bond. For example, the carbon=oxygen double bond in carbon dioxide counts as one bonding pair, even though four electrons are present in the bond.

The bonding and non-bonding pairs of electrons arrange themselves around the central atom to minimise the electrostatic repulsion between them, forming the shapes shown in Table 13.1.1. These shapes are modified by the presence of non-bonding pairs of electrons.

13.1.1 The shape of molecules

TABLE 13.1.1 Shapes associated with the number of electron pairs

NUMBER OF ELECTRON PAIRS	SHAPE	VISUAL REPRESENTATION
1	Linear	
2	Linear	
3	Trigonal planar	
4	Tetrahedral	
5	Trigonal bipyramidal	
6	Octahedral	

13.1.2 Covalent bonding in ammonia

Ammonia

The shape of the ammonia molecule (NH_3) is modified by a non-bonding pair of electrons.

A formula of NH_3 suggests at first glance that the molecule has a trigonal planar shape because there are three N—H bonds, so three bonding pairs of electrons. However, the Lewis structure shows an extra pair of electrons, a non-bonding pair, around the central nitrogen atom (Figure 13.1.1a). Having four pairs of electrons means that each pair of electrons appears at the corners of a tetrahedron with the nitrogen atom in the centre.

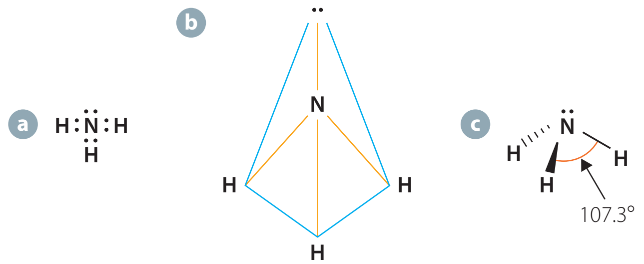


FIGURE 13.1.1 The Lewis structure for ammonia indicates a tetrahedral shape but the non-bonding pair of electrons at the top corner of the tetrahedron does not take part in the shape, leaving the pyramidal shape.

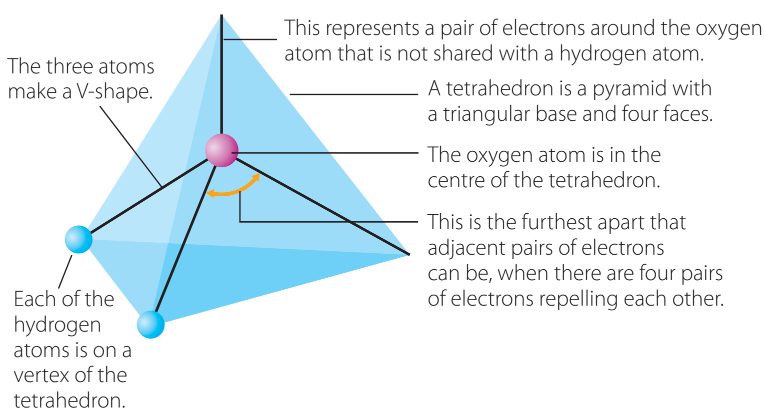
13.1.3 Hydrogen bonds in water

Water

Another example of how non-bonding pairs of electrons can affect the shape of a molecule is seen in water (H_2O). H_2O has two bonding pairs of electrons, so it may be anticipated that it would form a linear shape: H—O—H.

However, the Lewis structure shows four pairs of electrons around the central oxygen atom: two bonding pairs with the hydrogen atoms, and two non-bonding pairs (Figure 13.1.2).

FIGURE 13.1.2 The Lewis structure for water indicates a tetrahedral shape but the two non-bonding pairs of electrons at the top and bottom corners of the tetrahedron do not take part in the shape, leaving the bent or 'V'- shape.



Bond angles

Non-bonding pairs of electrons affect the bond angles in a molecule. A bond angle is the angle between the atoms bonded to the central atom. Methane and ammonia are two examples. Methane has no non-bonding pairs of electrons, so the H—C—H bond angle is 109.5°. In ammonia, the presence of the non-bonding electron pair gives an H—N—H a bond angle of 107.3°. Water has two non-bonding electron pairs, so the H—O—H bond angle is 104.5°.

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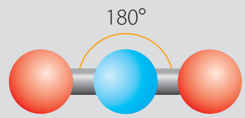
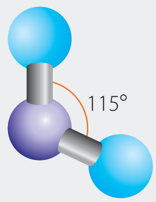
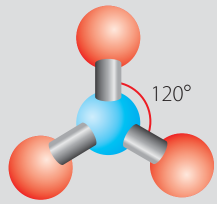
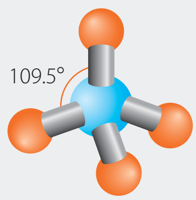
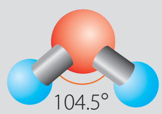
Non-bonding pairs of electrons affect the angles. This is because they exert stronger repulsive forces between each other, and between a non-bonding pair and a bonding pair, than the repulsive forces *between* bonding pairs.

13.1.4 VESPR and bond angles

TABLE 13.1.2 **Shapes of molecules associated with bonding and non-bonding pairs of electrons**

GENERAL FORMULA	NUMBER OF BONDING ELECTRON PAIRS	NUMBER OF NON-BONDING ELECTRON PAIRS	MOLECULAR SHAPE
MX_2	2	0	Linear – 180° (e.g. CO_2) 
MX_2E	2	1	V-shaped (bent) – 115° (e.g. NO_2) 
MX_3	3	0	Trigonal planar – 120° (e.g. BF_3) 
MX_4	4	0	Tetrahedral – 109.5° (e.g. CH_4) 
MX_2E_2	2	2	V-shaped (bent) –104.5° (e.g. H_2O) 

M = central atom, X = bonding pairs, E = non-bonding pairs

SECTION REVIEW

13.1

REMEMBERING

1 Define:

a bonding electron pair
b non-bonding electron pair
c valence electrons.

UNDERSTANDING

2 Carbon dioxide (CO_2) and sulfur dioxide (SO_2) both have their central atoms double-bonded to 2 oxygen atoms. Draw their shapes and explain the differences, if any, between them.

13.2 The polarity of molecules

SYLLABUS LINK

Chapter 3 revises electron sharing in covalent molecules.

In elemental molecules such as nitrogen (N_2), hydrogen (H_2) and chlorine (Cl_2), the electrons are evenly shared between the atoms that make up the covalent bond. However, with heteroatomic molecules such as hydrogen chloride (HCl), ammonia (NH_3) and water (H_2O), the electrons are not evenly shared. This is caused by a property of elements called electronegativity (electron-attracting power).

Electronegativity

SYLLABUS LINK

Chapters 1 and 7 revise electronegativity.

Electronegativity is the relative ability of an atom in a molecule to attract electrons to itself. This relative ability is measured using the Pauling scale of electronegativity. Table 13.2.1 shows some electronegativity values of important elements. Helium, neon and argon do not have values because, as noble gases, they have full outer electron shells and no electron-attracting power.

The polarity of a covalent bond depends on the difference in electronegativity values of the elements that make up the bond. The greater the difference, the greater the polarity of the bond.

TABLE 13.2.1 Electronegativity values of some important elements

ELEMENT	ELECTRONEGATIVITY
H	2.2
He	–
Li	1.0
C	2.5
N	3.0
O	3.5
F	4.0
Ne	–
Cl	3.0
Br	2.8
Ar	–

Hydrogen

The hydrogen molecule (H_2) consists of two hydrogen atoms bonded together by a single covalent bond. The two atoms in the bond are identical, so they have the same electronegativity. This means that the H—H bond is non-polar and hydrogen is a non-polar molecule.

Hydrogen chloride

The hydrogen chloride (HCl) molecule consists of a hydrogen atom bonded to a chlorine atom by a single covalent bond. The two atoms making up this bond are different; they have different electronegativity values, which means that there is a difference in electronegativity (Figure 13.2.1).

The H—Cl bond is polar. It has a slightly positive pole ($\delta+$) and a slightly negative pole ($\delta-$). It can help to think of a covalent bond as a tug of war between the two elements. In Figure 13.2.2, the chlorine atom is more electronegative than the hydrogen atom and can 'pull' the electrons towards itself, thereby becoming slightly more negative. The hydrogen atom will become slightly more positive.

$Cl \longrightarrow H$

3.0 − 2.2 = 0.8

FIGURE 13.2.1 The polarity of the H—Cl bond, showing the difference in electronegativity between chlorine and hydrogen

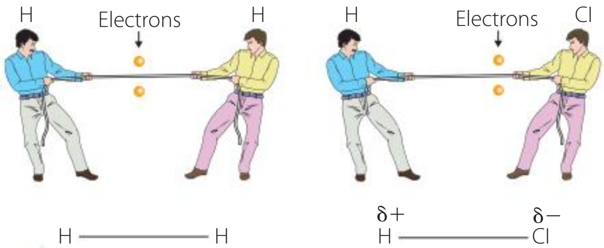


FIGURE 13.2.2 The 'tug of war' between two elements, showing the polarity of the H—Cl bond compared to the H—H bond

Polar and non-polar polyatomic molecules

For molecules with more than two different atoms, the presence of a polar covalent bond does not necessarily mean that the whole molecule is polar. Molecules can be non-polar despite the presence of polar bonds. For the molecule to be polar, a polar bond must be present and there must be asymmetry in the molecular shape. Carbon dioxide (CO_2) is a linear molecule with a central carbon atom and an oxygen on each side. The molecule is symmetrical (Figure 13.2.3). The C═O bond is polar, with the dipole pointing towards the oxygen because it is more electronegative, but there is no net separation of charge in the overall molecule. Each oxygen will attract the electrons but, as they are on either end of the molecule, the charge is distributed over the whole molecule. It is like having an equally strong person on each end of a rope in tug of war; the rope does not move because the pull from each end is equal. The poles cancel each other out. Although the molecule has polar covalent bonds, there is no net separation of the charges, so no net dipole. This molecule is non-polar.

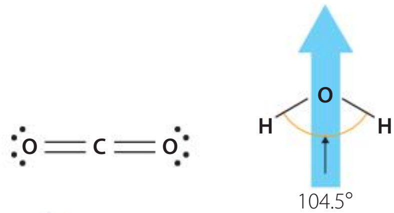


FIGURE 13.2.3 The non-polar carbon dioxide molecule (CO_2, left) and the polar water molecule. The arrows point from the positive end to the negative end of the dipole.

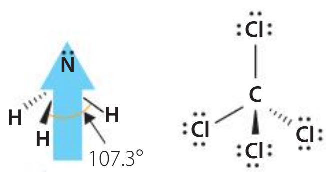


FIGURE 13.2.4 The molecular shape of ammonia (NH_3) and tetrachloromethane (carbon tetrachloride, CCl_4)

Contrast the polarity of carbon dioxide with water as shown in Figure 13.2.3. In water, the O—H bond is polar. Oxygen is central and the two hydrogens are at the bottom of the 'V'-shape. This means that the bottom end of the molecule is slightly positive. Water is asymmetrical; there is a net dipole present so water is polar. The presence of polar covalent O—H bonds and a V-shape means that the water molecule is polar. The dipoles cancel out in carbon dioxide but add together in water.

The ammonia (NH_3) and tetrachloromethane (carbon tetrachloride, CCl_4) molecules have different shapes and polar bonds. Ammonia, a pyramidal-shaped molecule, is polar. The nitrogen atom at the top of the pyramid is slightly negative due to its greater electronegativity. The hydrogen end of the molecule is slightly positive. Due to the molecule's asymmetry, it is a polar molecule (Figure 13.2.4). Tetrachloromethane has polar bonds due to differences in electronegativity between the carbon and chlorine atoms – it is tetrahedral in shape with a central carbon. The dipole in each of the C—Cl bonds points to the chlorine atom. However, the negative regions on chlorine in CCl_4 cancel themselves out due to the symmetry of the molecule, which is why CCl_4 molecules are non-polar (Figure 13.2.4).

13.2.1 Polar and non-polar molecules

SECTION REVIEW 13.2

REMEMBERING

1 Define:

a dipole

b polar covalent bond

c polar.

UNDERSTANDING

2 The most electronegative elements, in order of decreasing strength are: F > O > N = Cl > Br > C = S = I > P = H > Si. Consider the following pairs of elements: O—H, Cl—H, C—H, P—H, N—O, N—Cl, N—H, S—H, C—S, F—O, Cl—O, C—O, P—O, C—Cl

a With the covalent bonds between the above elements, classify each as slightly more negative or slightly more positive by labelling with δ^- and δ^+ as appropriate.

b Determine the electronegativity difference for each pair of elements, and use the values to classify each pair as polar, slightly polar or non-polar.

APPLYING

3 Consider the following molecules.

a SiH_4

b CH_3Cl

c CS_2

d PH_3

i Draw Lewis structures and valence structures.

ii Determine the shape of the molecule.

iii Predict if the molecule has polar bonds.

iv Show the direction of the dipole, if present, by drawing an arrow towards the negative end of each polar bond.

v State whether the molecule has a net dipole, making it polar.

13.3 Observable properties and intermolecular forces

An understanding of why some molecules are polar while others are not is of crucial importance when explaining the observable properties of substances.

These properties include:

- melting point
- boiling point
- solubility in polar and non-polar solvents
- vapour pressure.

The degree of polarity of molecules determines the intermolecular forces present within substances.

Intermolecular forces

intermolecular forces
attractive forces that cause molecules to aggregate together to form solids

Intermolecular forces occur between molecules. They can occur between molecules of the same substance and between molecules of different substances. Intermolecular forces should not be confused with intramolecular forces. Intramolecular forces are covalent bonds and are much stronger than intermolecular forces. VSEPR theory provides a basis for predicting the shapes of molecules based on the repulsive forces between electron pairs, both bonding pairs and non-bonding pairs. These are intramolecular forces. However, VSEPR theory also permits the understanding of the distribution of electrons within a molecule, which allows estimation of the molecular polarity. Polarity in turn affects the strength of intermolecular forces.

Three types of intermolecular force, in order of increasing strength, are:

- dispersion forces
- dipole–dipole forces
- hydrogen bonding.

13.3.1 Intramolecular and intermolecular forces
13.3.2 Intermolecular forces
13.3.3 Intermolecular forces of attraction

Dispersion forces

Polar molecules can form dipole–dipole or hydrogen bonds between the molecules. Non-polar molecules cannot, so there must be some other force that enables non-polar molecules, such as dry ice, liquid nitrogen and oils and fats, to form liquids and solids. This force is a **dispersion force**.

dispersion force the weak attractive force between atoms and molecules caused by an instantaneous temporary change in dipole moment, arising from the movement of orbiting electrons

The simplest molecule, hydrogen (H_2), can be used to explain a dispersion force (Figure 13.3.1). The force is very weak because hydrogen has an extremely low melting point (about −259.2°C). Hydrogen molecules have a cloud of electron density surrounding them. When the electrons, which can be thought of as constantly 'orbiting' the molecule, are on one side of the molecule there is, for an instant, a temporary dipole. One end of the molecule will be very slightly more positive and the other end will be very slightly more negative. The temporary positive end attracts a neighbouring molecule's electrons, inducing them to come towards the positive area. This forms the weak electrostatic attractions of dispersion forces. The fluctuating changes in charge result in a very weak force between molecules.

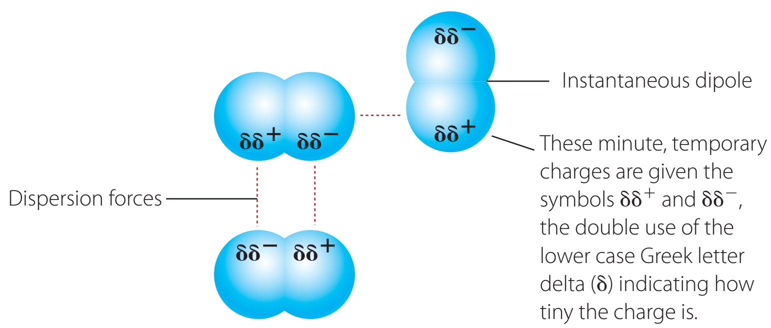


FIGURE 13.3.1 The dispersion force between hydrogen molecules in solid hydrogen, showing the instantaneous dipole–dipole interactions.

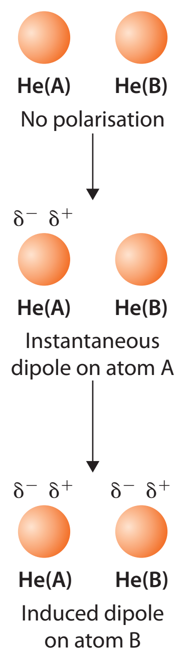


FIGURE 13.3.2 Noble gases can explain how dispersion forces occur through instantaneous dipoles induced due to the movement of electrons.

Dispersion forces exist in every molecule because every molecule has rapidly fluctuating electrons orbiting the molecule. The force is due to the rapid movement of electrons in the cloud of electron density surrounding the molecule. Dispersion forces are due to instantaneous temporary dipoles that occur with the movement of electrons.

Dispersion forces are much weaker than other intermolecular forces. When dipole–dipole or hydrogen bonding is present, the effect of dispersion forces is discounted.

Dispersion forces are also present in the noble gases (Figure 13.3.2). Electrons in two adjacent atoms occupy positions so that the atoms form temporary dipoles. With more electrons, there are more chances for dispersion forces to occur. Molecules are electrically neutral, so when there are more electrons, there are more protons. Increasing molecular weight of molecules means an increased chance of dispersion forces and a higher melting point (Table 13.3.1).

The electrons in larger atoms are also further from the nucleus and more readily form temporary dipoles. This explains why the larger noble gas argon has a boiling point of −186°C, much higher than that of helium, −269°C. A similar trend of increasing dispersion forces with increasing molecular weight is seen with the halogens (Table 13.1.1). The more electrons present (the larger the molecular weight), the bigger the dispersion force will be between the molecules.

TABLE 13.3.1 Boiling points of noble gases and halogens

NOBLE GAS	BOILING POINT (°C)	HALOGEN	BOILING POINT (°C)
Helium	–269	Fluorine	–188
Neon	–246	Chlorine	–34.6
Argon	–186	Bromine	58.8
Krypton	–152	Iodine	183
Xenon	–108	Astatine	337
Radon	–62		

Dipole–dipole forces

dipole–dipole force the attraction between molecules with permanent dipoles

In polar molecules, there is a net dipole. The dipole–dipole force occurs when there is a permanent dipole present. One end of the molecule is slightly negative while the other end is slightly positive. Hydrochloric acid (HCl) is a polar molecule. The presence of dipoles creates an attraction between neighbouring molecules, a **dipole–dipole force**, which is an attraction between two dipoles in two individual molecules. Molecules line up so that the slightly positive (δ^+) end of one molecule attracts the slightly negative end (δ^-) of another molecule (Figure 13.3.3). The dipole–dipole force (dipole–dipole attraction) is relatively strong because the dipole is permanent.

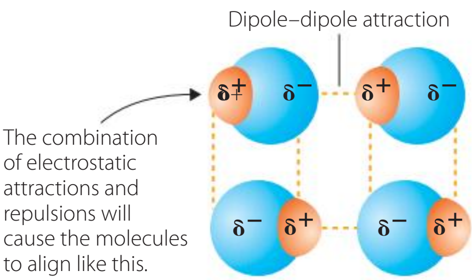
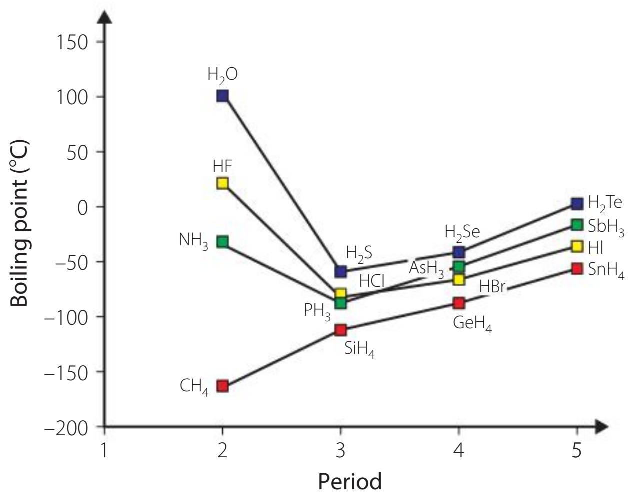


FIGURE 13.3.3 Dipole–dipole attraction between polar molecules of hydrogen chloride (HCl)

Hydrogen bonds

Boiling and melting points give an estimation of the strength of the intermolecular forces between molecules. Figure 13.3.4 shows a general increase in boiling points as you go down the periodic table groups, if you could exclude period 2 hydrides.

FIGURE 13.3.4 Boiling point of hydrides of groups 14–17



When nitrogen, oxygen or fluorine is attached to a hydrogen atom, the resulting molecule has an unusually high boiling point. When these very electronegative elements bond to hydrogen, a special type of dipole–dipole force occurs called a **hydrogen bond**. This is the strongest of the weak intermolecular forces, being about one-tenth the strength of a covalent bond. In diagrams, it is shown as a dotted line between the molecules.

hydrogen bonds when an H atom that is bonded to an O, N or F atom in one molecule, becomes attracted to the lone pair of electrons of N, O or F of an adjacent molecule

Hydrogen bonds are due to the hydrogen nucleus (a single proton) being extremely small. When hydrogen is bonded to the most electronegative elements (fluorine, oxygen and nitrogen), the charge difference over the polar covalent bond is at a maximum. Water displays hydrogen bonding. Oxygen's strong ability to attract electrons means that the electrons in the O—H covalent bond are more attracted to oxygen. Oxygen becomes slightly negatively charged and hydrogen is left with a slightly positive charge. The hydrogen atom is so small that it also means that two adjacent molecules can get very close to each other. Similarly, in ammonia, the nitrogen attracts the electrons of the covalent bonds, leaving the hydrogen atoms more positive. These hydrogens can attract the nitrogens of adjacent molecules. Hydrogen bonding then occurs between the hydrogen in –OH, –NH or –FH and the lone pairs on nitrogen, oxygen or fluorine on the adjacent molecule (Figure 13.3.5).

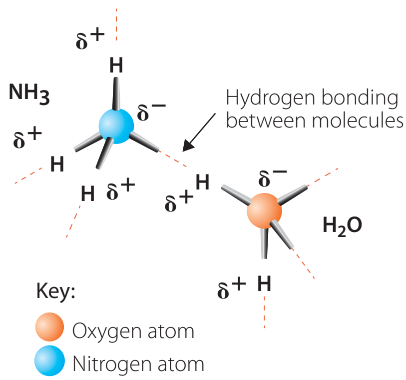


FIGURE 13.3.5 Hydrogen bonding occurring between a hydrogen (H) on one molecule and the lone pair of electrons on nitrogen (N) in another molecule

The importance of bonding

Bonding is responsible for many of the properties of materials and biomolecules. The nucleic acids in the DNA double helix are held together by hydrogen bonds. This allows them to unzip easily, enabling replication and transcription. Hydrogen bonding helps maintain the structures of proteins, enzymes, ribosomes and cell membranes.

Analytical techniques such as chromatography work because of surface interactions and the intermolecular bonds of different substances to different surfaces or solvents. Utilising this knowledge allows chemists to determine the organic substances in mixtures.

Water and hydrogen bonding

There are three types of bonding in water. Strong covalent bonding exists between the hydrogen and oxygen within the molecule. Hydrogen bonding and dispersion forces occur between the water molecules. Hydrogen bonding is responsible for the unique properties of water.

The unique properties of water include unusual melting and boiling points, density in solid and liquid phases, surface tension and an ability to act as a very good solvent. Hydrogen bonding explains all the properties that relate to temperature. A relatively large amount of energy is required to break the bonds in water, which is why water has a relatively high boiling point, 100°C. For a certain amount of energy, fewer molecules escape into the air. The fewer vaporised water molecules means lower vapour pressure, the downwards pressure exerted by the gas molecules on the surface of the liquid. To change state from liquid to gas (high latent heat) or to increase the average temperature of the water (high specific heat) energy must be absorbed. The change of phase from liquid to gas is endothermic.

SYLLABUS LINK

Chapter 16 discusses surface tension and capillary action.

The surface tension of water is high as hydrogen bonding occurs in all directions and is relatively strong. It holds the molecules together (Figure 13.3.6). In the bulk of liquid water, each molecule has forces operating in all directions. However, on the surface, there is an imbalance – the liquid molecules are attracted to each other and exert a net force that pulls them together.

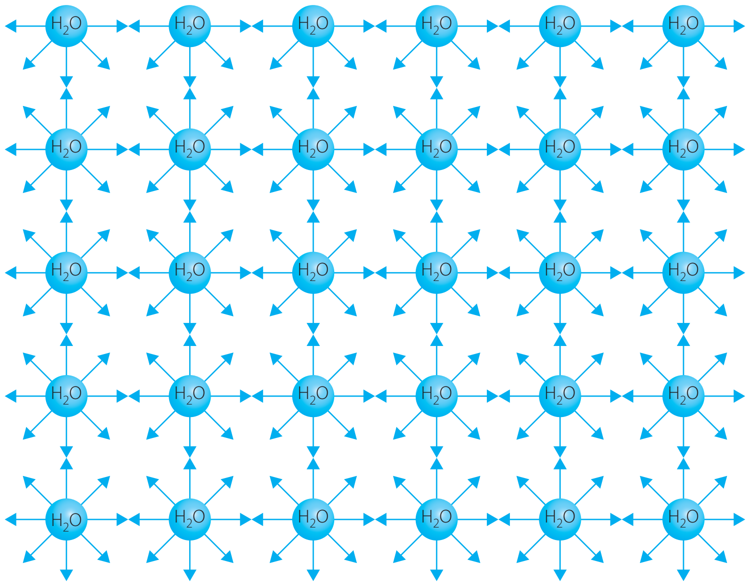


FIGURE 13.3.6 Molecular representation of surface tension in water

SYLLABUS LINK
Chapter 17 discusses solubility in greater depth.

Intermolecular forces and solubility

When considering the solubility of substances (solutes) in other substances a general rule applies:

Like dissolves like

The 'like dissolves like' rule means that polar substances (including ionic) dissolve in polar solvents, and non-polar substances dissolve in non-polar solvents.

Water, being a polar substance that contains hydrogen bonds, can dissolve ammonia, another polar substance that contains hydrogen bonds. Methane has dispersion forces only and is insoluble in water.

SECTION REVIEW 13.3

REMEMBERING

1. List three strong bonding forces.
2. Identify the weak intermolecular forces.

UNDERSTANDING

3. Explain why you should not put sealed bottles full of water in the freezer.
4. Rank the three weak intermolecular forces from weakest to strongest. Explain how polarity influences this order.
5. Explain the rule 'like dissolves like'.

APPLYING

6. Explain why dihydrogen sulfide (H_2S), a heavier molecule than water with more electrons, has a lower boiling point.

13.4 Mandatory practical

PRACTICAL ACTIVITY 13.4.1

Molecular model building

AIM

To use VSEPR theory to predict the shapes of simple molecules and apply this to build models of the molecules to demonstrate their three-dimensional shape

PROCEDURE

1 Copy the following table and use your understanding of VSEPR theory to fill it in. (Symbolising)
2 Conduct research to find out the bond angles in each molecule. (Investigating)
3 Build each model using a molecular model kit or straws and beads. (Symbolising)

MOLECULAR FORMULA	LEWIS STRUCTURE	SHAPE INDICATED BY ELECTRON PAIRS	PREDICTED SHAPE USING VSEPR (INCLUDING BOND ANGLES)
HF			
BF_3			
PH_3			
CO_2			
H_2O			
SO_2			
CH_4			
$CHCl_3$			

QUESTIONS

1 How did your models compare with your predictions? (Matching)
2 Compare the bond angles for H_2O and SO_2. Can you explain any differences? (Matching)
3 Compare the bond angles for CH_4 and CH_3Cl. Can you explain any differences? (Matching)
4 Using the information you have gathered in this activity, write a conclusion addressing the relative repulsive power of single bonds, double bonds and lone pairs of electrons. (Investigating)

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 Define the following terms.
- **a** Bonding electrons
- **b** Non-bonding electrons
- **c** Melting point
- **d** Boiling point
- **e** Solubility in polar and non-polar solvents
- **f** Vapour pressure
- **g** Dispersion forces
- **h** Dipole–dipole forces
- **i** Hydrogen bonding

CATEGORY QUESTIONS

2 Determine whether the following are polar.
- **a** Ammonia
- **b** Iodine chloride
- **c** Dichloromethane (CH_2Cl_2)
- **d** Nitrogen trichloride
- **e** Sulfur difluoride

3 What are the common groups that exhibit hydrogen bonds?

ELABORATION QUESTIONS

4 Geckos have utilised dispersion forces to allow them to support their full weight with just one toe. If scientists could replicate the ability of a gecko to attach to and detach from a smooth surface, how might that knowledge benefit society?

EVIDENCE QUESTIONS

5 Hydrogen peroxide is polar. Predict the shapes of H_2O_2. Explain why at first this appears unusual.

6 What effect, if any, do non-bonding pairs of electrons have on the bond angles in a molecule. Illustrate your answer by comparing the bond angles in water and nitrogen dioxide.

7 Ethyne (acetylene) has the formula C_2H_2. It contains the $C{\equiv}C$ triple bond. Use VSEPR theory to predict its shape.

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END-OF-CHAPTER EXAM

End-of-chapter test

CHEMISTRY

1 Choose the type of geometry found in the water molecule.

A Linear

B Trigonal planar

C Tetrahedral

D Pyramidal

2 The central atom in NF_3 is surrounded by:

A 3 single bonds and two non-bonding pairs of electrons.

B 3 single bonds only.

C a double bond, a single bond and a non-bonding pair of electrons.

D 3 single bonds and a non-bonding pair of electrons.

3 The boiling point of water is much higher than would be predicted from the boiling points of the other group 16 hydrides such as H_2S, H_2Se and H_2Te. Which of the following statements explains this?

A Water is less polar than H_2S, H_2Se and H_2Te.

B Water contains hydrogen bonding.

C The covalent bonds in water are stronger than H_2S, H_2Se and H_2Te.

D Water has a lower molecular mass.

4 In which molecule is there at least one polar bond but no overall dipole?

A CF_4

B SO_2

C $CHCl_3$

D NH_3

5 What is the name of the geometrical shape formed when a central atom is surrounded by two double bonds, a single bond and a non-bonding pair of electrons?

6 From the following list of elements, select the pair that would form the strongest dipole:

Cl, S, C, H, F, N, Ar, Br and P.

7 Name the property of substances that involves the downwards pressure exerted by gas molecules on the surface of the liquid of the same substance.

8 Explain why the boiling point of CH_4 (–161.5°C) is lower than that of CCl_4 (76.2°C).

9 Explain why ammonia (NH_3) is highly soluble in water but phosphine (PH_3) is virtually insoluble.

10 How would the dipoles present in a Cl_2 molecule differ from the dipoles present in a HCl molecule?

11 Which liquid at room temperature would evaporate more quickly: CCl_4 or H_2O? Justify your decision.

14 CHROMATOGRAPHY

Introduction

Chromatography is one of the most powerful and, in many ways, simplest analytical tools at our disposal. It is used to separate mixtures, identify unknown substances and give information on amounts of substances present in a mixture.

The word 'chromatography' is derived from the Greek words for 'colour', *khroma*, and 'to write', *graphein*. The term was first used by Russian scientist Mikhail Tsvet in 1900 in his work on separating plant pigments such as chlorophyll, carotenes and xanthophylls, which are coloured.

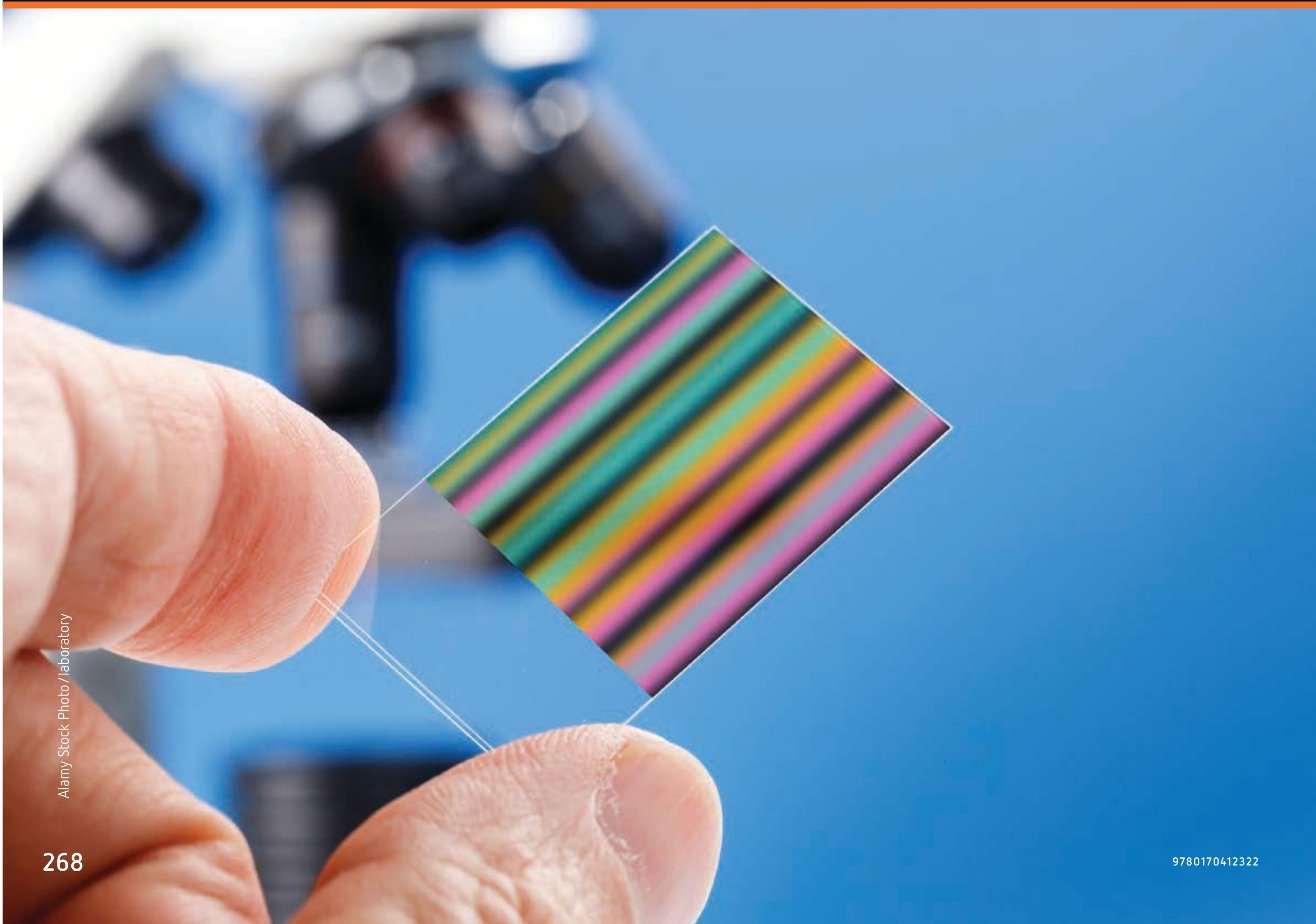
Many new forms of chromatography have been developed, and with technological improvements in recent decades, these new techniques are becoming indispensable to the analytical chemist.

Stimulus questions

What happens to ink dropped onto a piece of paper?

What happens when water is dropped onto an ink spot on a piece of paper?

Is ink a pure substance?



Alamy Stock Photo/laboratory

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14.1 Chromatography techniques

Knowledge of how solutes dissolve in solvents and interact with the surface of materials is the basis of the analytical techniques of **chromatography**. There are three main techniques in popular use today: thin-layer chromatography, gas chromatography and high-performance liquid chromatography. The chemical basis for these techniques is that different substances will adsorb onto a surface (called the stationary phase) and desorb into a solvent (the mobile phase) at different rates.

chromatography
a group of techniques that separate substances based upon differential distribution between a stationary phase and a mobile phase.

Paper chromatography

Paper chromatography is the simplest form of chromatography but it can still be useful in the qualitative analysis of components in a mixture, particularly for coloured substances such as inks and food colourings.

In this method, the stationary phase is a uniform piece of absorbent paper while the mobile phase is a solvent or mixture of solvents in which the components are dissolved.

Separation occurs as a result of differences in the relative solubility of each component in the solvent and the water in the paper fibres.

While paper chromatography is useful as an introduction to chromatography, its analytical use has been superseded by other techniques such as thin-layer chromatography (TLC).

14.1.1 Chromatography
14.1.2 What is paper chromatography?

PRACTICAL ACTIVITY 14.1.1

Chromatography of plant pigments

The plant pigment chlorophyll absorbs the light energy used in photosynthesis. When investigating how chlorophyll works, chemists first needed to isolate chlorophyll to find out how many different types there are, then determine whether all plants had the same pigment (Figure 14.1.1).

AIM

To separate plant pigments using chromatography

PROCEDURE

1 Design a method to separate the plant pigments in plants by chromatography. Hint: Plant pigments will dissolve best in an ethanol and water mixture. Filter paper or chalk can be used to separate the pigments. Sand can help to break open the cells by grinding the plants. Calcium carbonate can neutralise acid released from the cells as they are crushed.

2 Copy and complete the risk assessment table. Add any more risks you can think of, as well as ways to manage them. Ask your teacher to check your risk assessment before you proceed.

WHAT ARE THE RISKS IN DOING THIS INVESTIGATION?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?
Heating a mixture of ethanol and water will cause the beaker to get hot	If you burn yourself, place the affected area under cold running water for 10 minutes and inform your teacher.
Ethanol is a toxic and flammable liquid	Only directly heat solutions that have less than 50 % v/v of ethanol. Keep bottles away from flames and heat sources. Do not come into contact with the liquid or breathe the fumes.
Plants may be poisonous	
Glassware	
Filter paper	

RISK ASSESSMENT

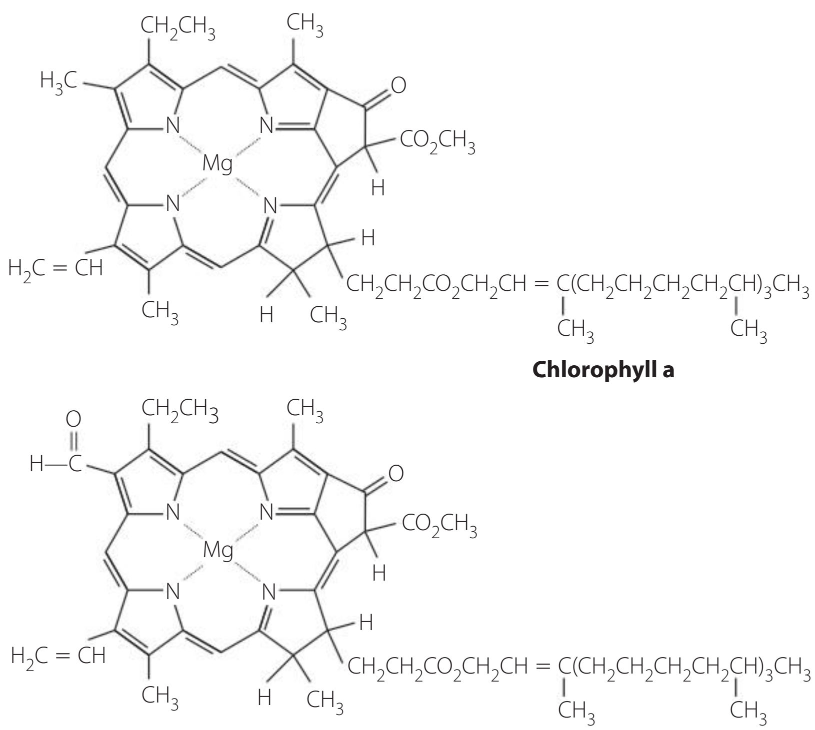

FIGURE 14.1.1 Chlorophyll pigments in plant cells absorb light energy. There are many different pigment molecules, each a different colour. Two are shown here.

QUESTIONS

1. How will you carry out your investigation? Which method will you try first?
2. Will large pieces of plant interfere with your technique? If so, what should you do?
3. For how long will you boil the plants? How will you judge whether you have achieved the maximum extraction?
4. How will you use the filter paper or chalk to separate the pigments from the solvent?
5. How fine will the plant pieces have to be? What factors would increase the amount of pigment released?
6. How will you analyse your results?

DISCUSSION

1. How well did you manage to extract plant pigments?
2. Did other students succeed in extracting plant pigments? Compare the effectiveness of their method with yours.

TAKING IT FURTHER

1. What were some of the problems encountered? What errors were encountered?
2. How would you refine this method? How will the refined method provide more consistent results?
3. If you suspected that one of the components was chlorophyll, how would you confirm this?

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Thin-layer chromatography

Thin-layer chromatography (TLC) is used to separate and analyse a wide variety of molecular mixtures. It is used to analyse the presence of particular drugs and amino acids; TLC is faster and provides more separation than paper chromatography.

The TLC plate typically consists of a 0.1-mm-thick layer of absorbent material bonded to a glass or plastic support. The absorbent material consists of a specially prepared, finely ground matrix of microscopic particles of silica gel, alumina or similar materials. These particles provide a large surface area for chromatographic separation. This is the stationary phase. The origin line is drawn in pencil about 1 cm from the base. This is where the samples are placed.

Separation is achieved by the solvent (mobile phase) moving up the stationary phase. This movement is due to capillary action. The solvent moves through the samples towards the top of the plate. The faint wavy line seen is called the solvent front. The samples will desorb into the solvent and travel up the plate with the solvent. If all the components spent the same amount of time in the solvent, then these components would all travel at the same rate. However, components have differences in chemical structure, which alters the adsorption and desorption rates. The particles most attracted to the stationary phase move the slowest. The particles that are most strongly attracted to the solvent move the fastest.

In TLC, you measure the distance the spot has moved from the origin compared to the distance the solvent has moved from the origin. This ratio is called the **retardation factor (R_f)**, as shown in Figure 14.1.2. The R_f value must be between 0 and 1 and is defined as the distance travelled by the sample from its origin divided by the distance of the solvent front from the sample origin. A substance that does not migrate from the sample origin has an $R_f = 0$, while one that is not adsorbed (i.e. has migrated with the front) has an $R_f = 1$. The R_f value is characteristic for a particular substance with a given absorbent and solvent system. The R_f value cannot be greater than 1.

retardation factor (R_f) the distance travelled by the sample from its origin divided by the distance of the solvent front from the sample origin

The separation of the components depends upon the length of time the plate is in contact with the solvent. If this is too short, then there is only a small separation. If this is too long, then all the spots would eventually make it to the top. A longer plate and a reasonable time will provide the best separation. It is not always possible to identify all the components of a sample with one TLC plate. Some samples may have very similar R_f values. Different solvent mixtures and stationary phases can be tried to achieve better separation.

14.1.3 Thin-layer chromatography

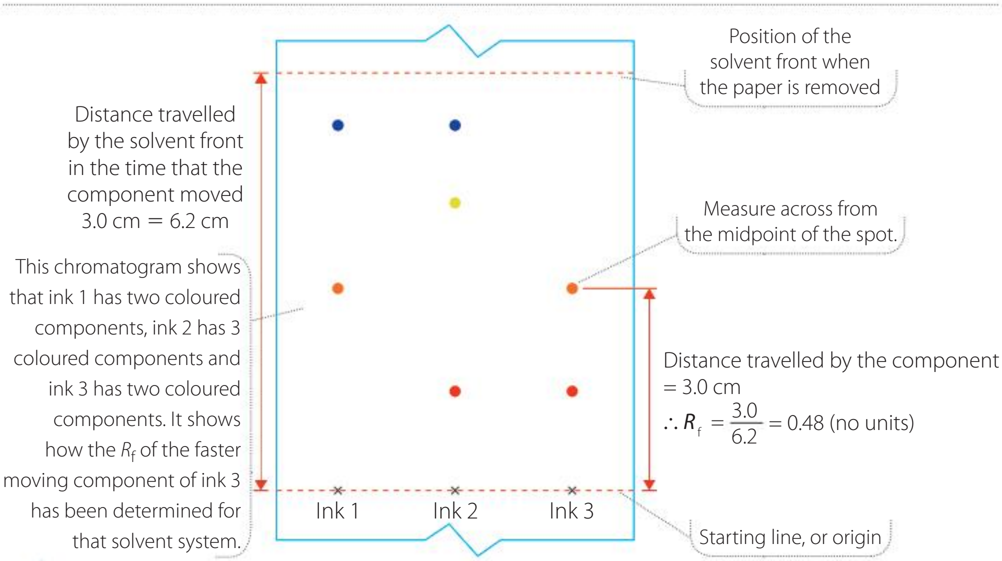

FIGURE 14.1.2 Determining the retardation factor (R_f) from a chromatogram

Gas chromatography

gas chromatography (GC)
a separation technique for small organic molecules that can withstand relatively high temperatures

14.1.4 Gas chromatography

Gas chromatography (GC) is a separation technique for small organic molecules that can withstand relatively high temperatures. Blood alcohol levels are measured by GC.

The gas chromatograph consists of a gas bottle containing a carrier gas, an oven, a column, a detector and a recorder. The sample is injected into the oven where the carrier gas pushes the sample into the long, thin column. In the column, smaller particles and those that adsorb onto the stationary phase the least leave the column first. Larger particles and those that adsorb more readily take longer to leave the column.

The time the sample takes to elute (leave the column) is characteristic for the substance. This time is called the retention time (R_t). The R_t value is used to identify the component, as the R_f is in TLC and paper chromatography.

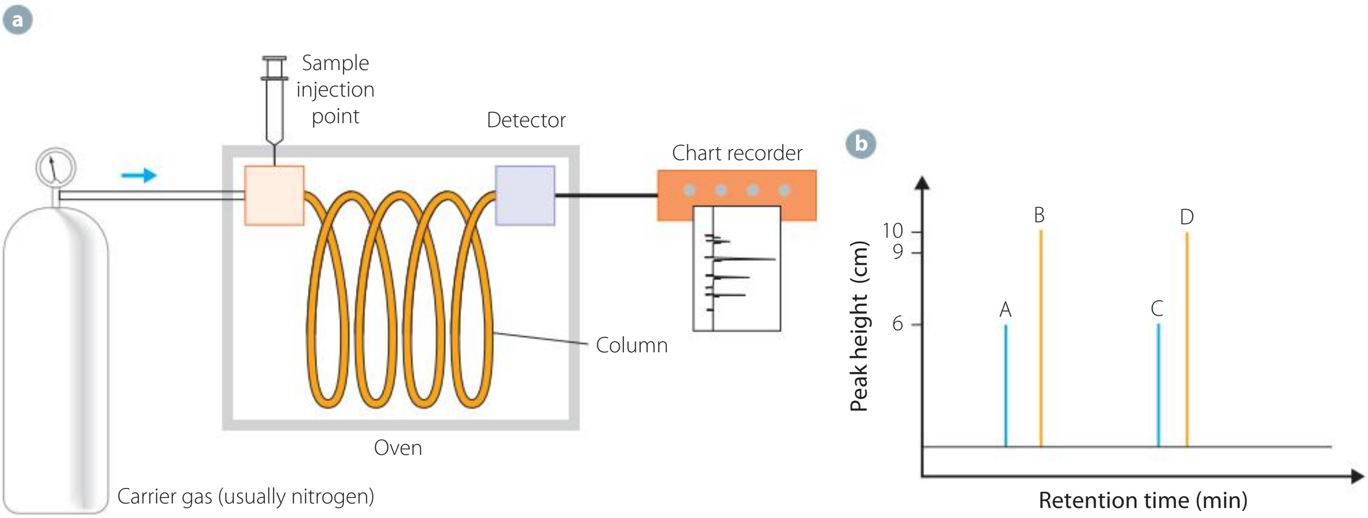

FIGURE 14.1.3 A schematic diagram of **a** a gas chromatograph and **b** a hypothetical gas chromatogram

GC can also measure the amount of the analyte. If a wine is thought to contain 12% alcohol, then a series of standards with alcohol ranging from 5% to 20% is run through the gas chromatograph and the peak areas are measured (Figure 14.1.3). A calibration curve is drawn and the unknown value is read off. It is important that the unknown value lies within the range of the standard values.

The R_t value will be the same for each standard. What is different is the height of the peak. The peak area will alter in proportion to the amount of the substance present. A calibration curve is plotted and the peak area of the unknown wine is used to determine the amount of alcohol present.

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High-performance liquid chromatography

High-performance liquid chromatography (HPLC) is used for larger organic molecules. It can be used for substances that are unstable to heat because there is no oven. The stationary column is shorter than in the gas chromatograph. The mobile phase is a liquid, not a gas. A pump is used because pressure is required to move the liquid through the column. As with GC, the R_t values are used to identify compounds, standards are run to confirm the identity (have the same R_t) and a calibration curve is used to determine the amount of the substance present.

high-performance liquid chromatography
a separation technique for large organic molecules that are unstable at high temperatures

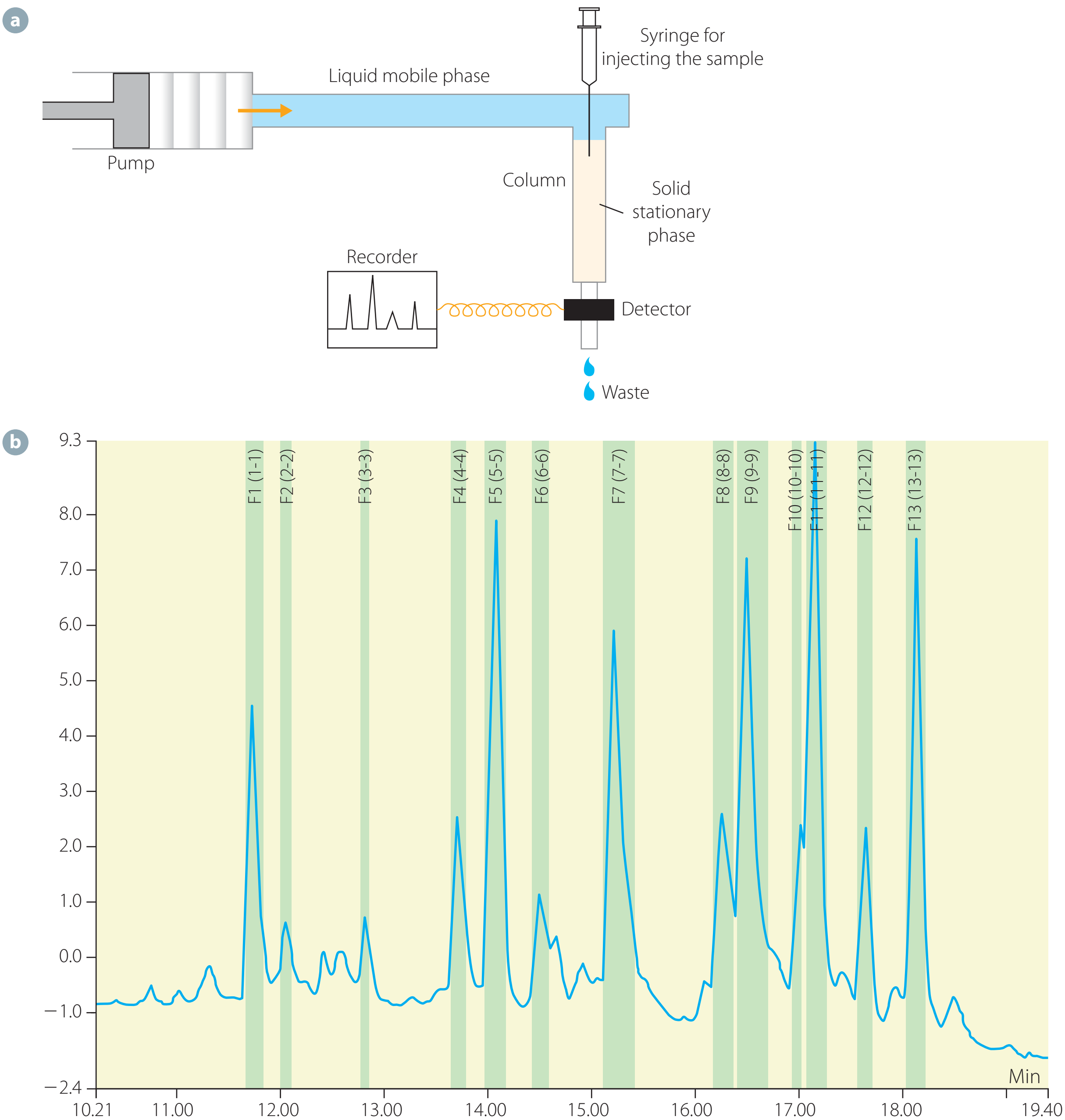

FIGURE 14.1.4 A schematic diagram of **a** a high-performance liquid chromatograph and **b** a high-performance liquid chromatogram of pigments extracted from a *Dianella* plant

TABLE 14.1.1 A summary of chromatography techniques

TECHNIQUE	BASIS	STATIONARY PHASE	MOBILE PHASE	MEASURES	USES
Thin-layer chromatography (TLC)	Solvent moves over the stationary phase by capillary action. Components adsorb and desorb at different rates due to chemical structure and ability to form intermolecular bonds to the stationary and mobile phases	Fine powder on glass or plastic	Solvent may be organic or water mixture	R_f is characteristic for each sample under the same conditions Standards are used to confirm identity (qualitative)	Organic mixtures such as plant pigments and drugs
Gas chromatography (GC)	Carrier gas is the mobile phase. Sample is injected into the oven. The gas moves the vaporised particles through a fine long column. The smallest solutes that adsorb the least, elute first. R_t depends on temperature, length of column, flow rate and chemical structure	Long, thin column	Inert gas such as N_2, CO_2 or He	R_t indicates what the sample is Area under peak indicates amount of substance present Standards are used to confirm identity (qualitative) and to prepare a calibration curve for calculations of amounts (quantitative analysis)	Small, heat-stable organic compounds such as ethanol
High-performance liquid chromatography (HPLC)	Similar to GC but mobile phase is a liquid. Unlike the gas in GC, the liquid plays an important role in adsorption and desorption. Needs a pump to push the liquid through the denser packed column	Short column (particle size of stationary phase is extremely small, allowing better separation)	Water–methanol or water–acetonitrile mixes	R_t indicates what the sample is Area under peak indicates amount of the substance present Standards are used to confirm identity (qualitative) and to prepare a calibration curve for calculations of amounts (quantitative analysis)	Larger organic compounds

R_f = retardation factor, R_t = retention time

SECTION REVIEW 14.1

REMEMBERING

1 List four chromatography techniques.

2 List the substances most suitable for analysis by:

- **a** thin-layer chromatography
- **b** gas chromatography
- **c** high-performance liquid chromatography.

APPLYING

3 Explain how a gas chromatogram is produced.

4 Explain why TLC would not be suitable for separating a large quantity of mixture.

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14.2 Separating components

Separating the components of mixtures depends largely on differences in the strength of the intermolecular forces of the components.

All forms of chromatography techniques work on the same principle. They all have a stationary phase and a mobile phase. The stationary phase can be a solid or a liquid supported on a solid. The mobile phase can be a liquid or a gas. During chromatography the mobile phase, carrying the components to be separated, moves through the stationary phase. The components spend differing amounts of time in the stationary phase and travel at different rates. To understand the role of intermolecular forces in chromatography, it is necessary to look at the different types of chromatography.

Thin-layer chromatography

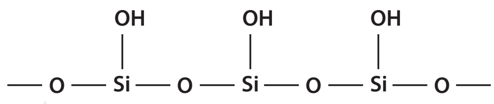


FIGURE 14.2.1 The surface of silica deposited on a TLC plate showing the —OH groups attached to the silicon atoms

In thin-layer chromatography (TLC), the most common material used for the stationary phase is silica (a form of silicon dioxide). A very thin layer of this is deposited on a glass or plastic plate. The important thing about the silica is that, at its surface, silicon atoms are bonded to —OH groups.

The presence of the —OH groups makes the stationary phase very polar and enables it to form hydrogen bonds with any suitable compounds in the mobile phase.

An important part of TLC is the choice of solvent (the mobile phase). The solvent must be able to dissolve the components of the mixture to be separated and must be attracted, to some degree, to the stationary phase silica.

During TLC, as the chromatogram develops, the solvent begins to soak up the plate. As it travels, the solvent dissolves the mixture of compounds placed at the base line of the plate. As the solvent travels further up the plate, the compounds travel with it. How far they travel depends on how long they stay dissolved in the solvent. This depends on the forces of attraction between the compounds and:

1. the solvent
2. the stationary phase.

For example, a mixture of 1-propanol and propanone was to be separated using TLC. The structures of these compounds are shown in Figure 14.2.2.

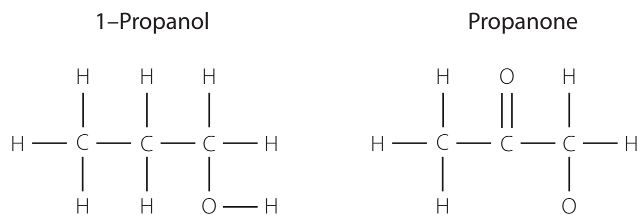
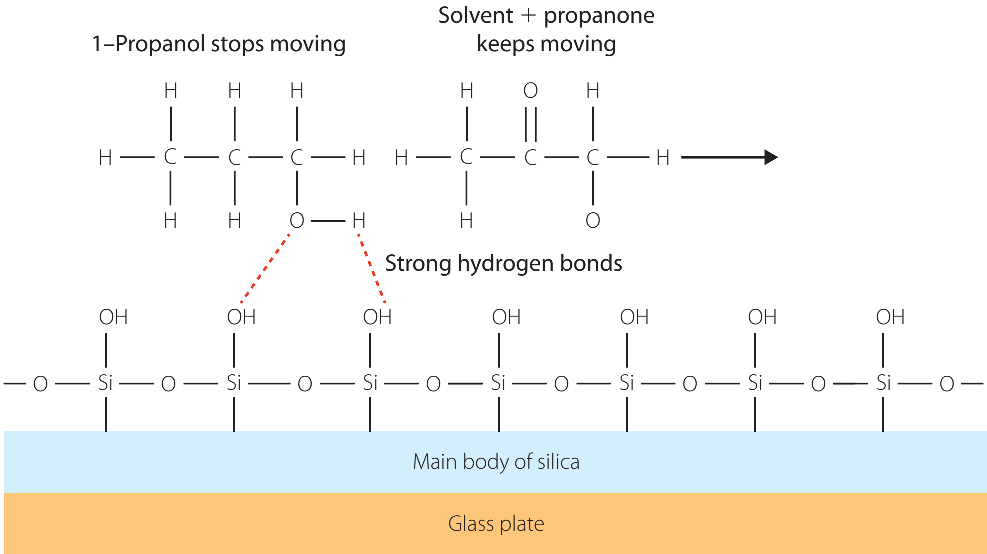


FIGURE 14.2.2 The structures of 1-propanol and propanone

Propanol contains the —OH group, which can form hydrogen bonds. Propanone is a polar molecule and can undergo dipole–dipole attractions.

These molecules will both dissolve in a solvent such as ethyl ethanoate (ethyl acetate) and when travelling up a TLC plate coated with silica will travel some distance before one of the compounds becomes stuck to the silica. In this case, the compound that will stick first is the 1-propanol due to its ability to form strong hydrogen bonds to the silica. The propanone, being simply polar, will not be as strongly attracted to the silica and will travel further up the plate dissolved in the solvent.

FIGURE 14.2.3 The movement of 1-propanol and propanone along a TLC plate. The 1-propanol forms strong hydrogen bonds with the silica stationary phase, so it will appear on the plate closer to the base line.



Gas chromatography

The principle of gas chromatography (GC) is essentially the same as TLC.

The column in a gas chromatograph consists of a glass or steel tube between 1 and 4 metres in length that is coiled so it can fit into an oven. The column is packed with a solid such as small aluminium silicate spheres coated with a high boiling point liquid.

The samples are injected into the column and are carried through the column in the mobile phase (typically an inert gas such as nitrogen).

As with TLC, the compounds being separated will spend differing times in the thin layer of liquid that coats the aluminium silicate spheres.

High-performance liquid chromatography

The principle of HPLC is essentially the same as for TLC. The main difference is that, instead of the liquid mobile phase being drawn across the stationary phase by capillary action, in HPLC a pump forces the liquid phase through the column at high pressure (up to 400 atm). This makes it much faster. The liquid phase travels through the column very quickly, so the packing material can have a much smaller particle size, giving it a larger surface area. This enables superior separation of component compounds.

SECTION REVIEW 14.2

REMEMBERING

1 Define:

a solvent **b** solute **c** hydrogen bond

d polar bond **e** non-polar bond.

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UNDERSTANDING

2 Explain how compounds with hydrogen bonds, dipole–dipole attraction and dispersion forces only would be separated on a TLC plate with:

- **a** a non-polar stationary phase
- **b** a polar stationary phase.

APPLYING

3 Figure 14.2.4 shows the gas chromatogram for a sample of common pesticides. Predict which of compounds 1–5 is the most non-polar. Explain your reasoning, given that the stationary phase is a non-polar waxy polymer.

4 It is possible to increase the effectiveness of chromatography by changing the polarity of the solvent. Would it be possible to do the same by changing the polarity of the stationary phase?

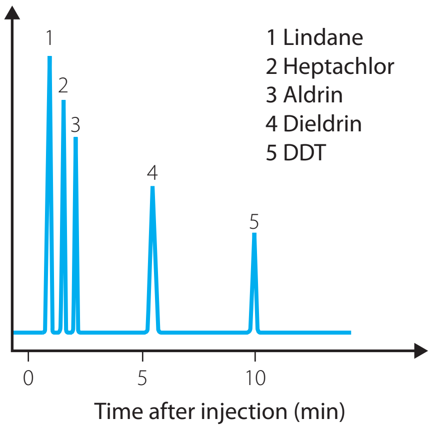


FIGURE 14.2.4 Gas chromatograms for some common pesticides

14.3 Inquiry skills: analysing, interpreting and evaluating data

Chromatography techniques are very important in the identification of unknown compounds; that is, in qualitative analysis. They are also excellent techniques for providing information about the amount of substances present; that is, quantitative analysis.

Thin-layer chromatography

TLC is used mainly for qualitative analysis when all that is required is a quick observation of the number of compounds present in a sample, or the qualitative detection of an unknown compound in a sample.

The most useful information taken from a TLC plate is the retardation factor (R_f) This is determined by:

$$R_f = \frac{\text{distance travelled by component}}{\text{distance travelled by solvent}}$$

For example, in an experiment, if the distance travelled by the solvent is 3.4 cm and the distance travelled by the sample is 1.9 cm, then the R_f value would be:

$$R_f = \frac{1.9}{3.4} = 0.5588$$

An R_f value has no units and is usually quoted to two decimal places, so the R_f value calculated above should be:

$$R_f = 0.56$$

As long as the conditions under which this experiment was carried out remain constant, such as temperature, solvent and type of plate, the R_f value will be the same.

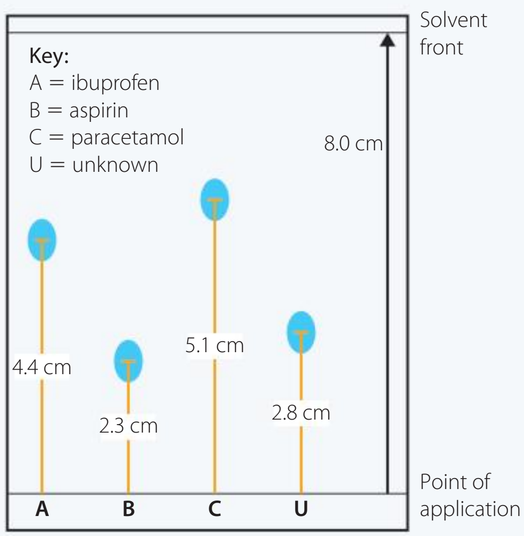
WORKED EXAMPLE 14.3.1

A painkiller medication is tested to see whether it contains aspirin, paracetamol or ibuprofen. The painkiller is dissolved in ethyl ethanoate (acetate) and placed on a TLC plate, marked 'U' (unknown). Pure samples of aspirin, paracetamol and ibuprofen are placed on the same plate. Once the solvent front reaches the top of the plate, the plate is removed and dried. An ultraviolet light is shone onto the plate and the spots are marked with iodine solution stain. Figure 14.3.1 shows the results.

QUESTION

Does the unknown painkiller contain aspirin, paracetamol or ibuprofen?

FIGURE 14.3.1 Thin-layer chromatography plate of an unknown painkiller and three known samples



ANSWER

1 Calculate the retardation factor (R_f) values for each component, as follows.

$$R_f\,\mathrm{A} = \frac{4.4}{8} = 0.55$$

$$R_f\,\mathrm{B} = \frac{2.3}{8} = 0.29$$

$$R_f\,\mathrm{C} = \frac{5.1}{8} = 0.64$$

$$R_f\,\mathrm{U} = \frac{2.8}{8} = 0.35$$

2 Compare the R_f values for the unknown painkiller with the known samples to determine the most likely constituent of the unknown painkiller:

Component B (aspirin) is the most likely constituent with an R_f value of 0.29 compared to a value of 0.35 for the unknown.

3 Suggest a reason why the R_f values are not identical:

The unknown sample may contain 'filler' materials that have affected the aspirin's R_f value.

4 Suggest a reason why an ultraviolet (UV) light was shone onto the plate:

Many compounds are not coloured. UV light can be used to show where they appear on the plate.

Gas chromatography

In GC it is possible to use the data to not only separate and identify substances but also to determine their relative amounts.

Figure 14.3.2 shows a simplified gas chromatogram of a sample containing methane, carbon dioxide and carbon monoxide.

This simplified gas chromatogram demonstrates that the peaks are represented by triangles. The quantity of each peak can be determined by measuring the *area* of each 'triangle'. The values for area are then multiplied by a proportionality constant to determine the actual amount of substance present. The proportionality constant depends on the specific gas chromatograph and the measurements taken (Worked example 14.3.2).

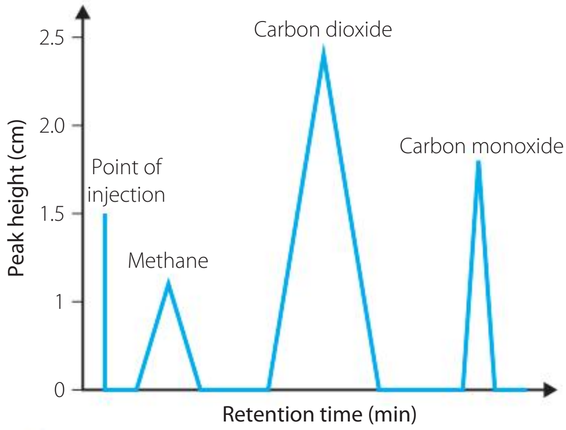


FIGURE 14.3.2 Simplified gas chromatogram of a gas mixture

WORKED EXAMPLE 14.3.2

A gas sample was analysed by GC for methane, carbon dioxide and carbon monoxide. Using the data provided, and with reference to Figure 14.3.2, calculate how much of each gas is present in the gas sample. The area: concentration proportionality constant is given to be 3×10^{-5} mol mm^{-2} for this case.

ANSWER

1 Methane:

$$\text{Peak height} = 11\text{ mm}$$
$$\text{Peak base} = 4\text{ mm}$$
$$\text{Area (triangle)} = \tfrac{1}{2} \times \text{base} \times \text{height}$$
$$= \tfrac{1}{2} \times 4 \times 11 = 22\text{ mm}^2$$

Proportionality constant $= 3 \times 10^{-5}$ mol mm^{-2} (in this case)

The amount of methane is:

$$22 \times 3 \times 10^{-5}$$
$$= 6.6 \times 10^{-4}\text{ moles}$$

2 Carbon dioxide:

$$\text{Peak height} = 24\text{ mm}$$
$$\text{Peak base} = 7\text{ mm}$$
$$\text{Area (triangle)} = \tfrac{1}{2} \times \text{base} \times \text{height} = \tfrac{1}{2} \times 7 \times 24 = 84\text{ mm}^2$$

Proportionality constant $= 3 \times 10^{-5}$ moles mm^{-2}

The amount of methane is:

$$84 \times 3 \times 10^{-5} = 2.5 \times 10^{-3}\text{ moles}$$

3 Carbon monoxide:

$$\text{Peak height} = 18\text{ mm}$$
$$\text{Peak base} = 2\text{ mm}$$
$$\text{Area (triangle)} = \tfrac{1}{2} \times \text{base} \times \text{height} = \tfrac{1}{2} \times 2 \times 18 = 18\text{ mm}^2$$

Proportionality constant $= 3 \times 10^{-5}$ moles mm^{-2}

The amount of methane is:

$$18 \times 3 \times 10^{-5} = 5.4 \times 10^{-4}\text{ moles}$$

SECTION REVIEW

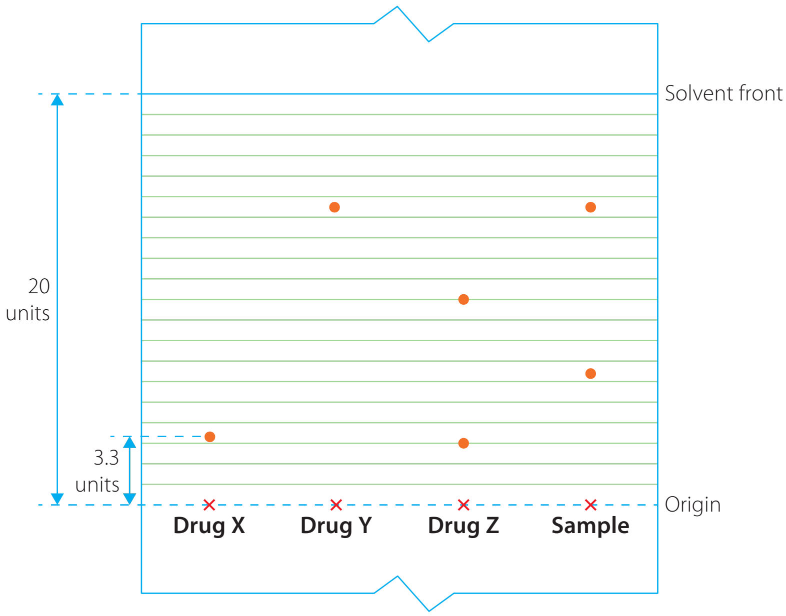
14.3

UNDERSTANDING

1 Explain why the spots on a TLC plate can sometimes appear as an elongated oval rather than a simple circle.

APPLYING

2 The following figure represents a thin-layer chromatogram of drug samples confiscated by the police. Included on the plate are commercial drugs, labelled X, Y and Z.



a Why do chemists include commercial drugs on the same TLC plate as the samples from the police?

b Does the confiscated sample contain any of the commercial drugs? Justify your answer using the data provided.

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CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 Define the following terms.

a Adsorption

b Desorption

c Analyte

d Mobile phase

e Stationary phase

f R_f

g R_t

CATEGORY QUESTIONS

2 Which factors determine the length of time an analyte spends in the mobile phase?

ELABORATION QUESTIONS

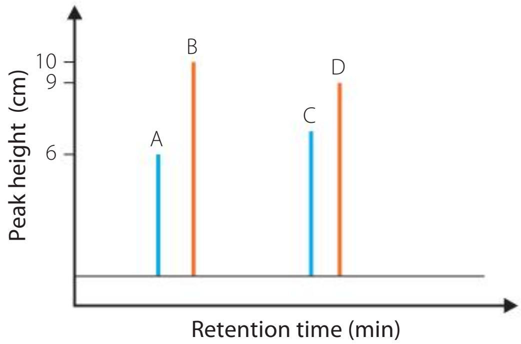
3 Explain why HPLC is a better technique to use than TLC for separating larger organic compounds.

4 When analysing a gas chromatogram, why does the area under a peak need to be multiplied by a proportionality constant?

EVIDENCE QUESTIONS

5 Honeys vary in their composition. Honey consists generally of the disaccharides maltose and sucrose, and the monosaccharides glucose and fructose.

HPLC is commonly used to analyse honey and to find out which sugars are present and in what percentage. The chromatogram in the following figure was obtained when a sample of Australian honey was tested.



a Which molecule spent most time in the stationary phase? Explain how you determined this.

b What is the percentage of 'C' in the mixture?

c Why was HPLC used in preference to GC in this analysis?

d It is suspected that peak C is fructose. How can this be verified?

END-OF-CHAPTER EXAM

End-of-chapter test

1 In chromatography, the mobile phase can be:

A solid or liquid.

B gas or liquid.

C gas only.

D liquid only.

2 In a thin-layer chromatogram of a mixture of butanoic acid, tetrachloromethane and butanol, using ethyl ethanoate (ethyl acetate) as a solvent and silica as the stationary phase, which of the following shows R_f values in order from highest to lowest?

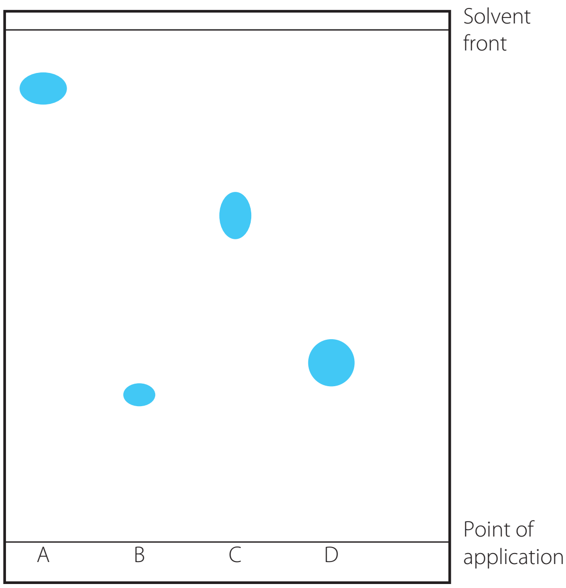
A Butanoic acid, butanol, tetrachloromethane

B Butanol, tetrachloromethane, butanoic acid

C Tetrachloromethane, butanol, butanoic acid

D None of the above

3 A TLC plate is shown in the following figure. The solvent used was pentane (a non-polar liquid).



Which sample contains the most-polar molecules?

A A

B B

C C

D D

4 What is the term given to the substance that coats a TLC plate?

5 In HPLC, what is the term used to identify an unknown substance, determined using the time taken for substances to leave the column?

6 In your opinion, what property of the carrier gas in GC is most important?

7 Which two aspects of the design of a high-performance liquid chromatograph make it so useful?

8 The following table provides the retention times and relative peak areas for a series of ethanol standards and a sample of champagne when they were run through a gas chromatograph.

ETHANOL % (w/v)	RELATIVE PEAK AREA	R_T
8.4	4.27	0.90
9.8	5.72	0.89
10.6	6.75	0.90
12.2	7.98	0.89
Champagne sample	5.24	0.90

A Construct a calibration curve for the ethanol and use it to determine the percentage of ethanol in the champagne.

B If the density of ethanol is 0.785 g mL^{-1}, express your answer as % v/v.

C Explain why you need a calibration graph and not just peak areas.

D Explain why this calibration curve would not be suitable to find the alcohol content of a low alcohol wine that typically has a value of around 7% v/v.

15 GASES

Introduction

Our atmosphere consists almost entirely of gas. Gases are all around us and inside us. An understanding of the behaviour of gases is of crucial importance to our industry, our economy and our environment.

An appreciation of the particulate nature of gases enables an understanding of the properties of gases, such as pressure, temperature and volume, which, in turn, leads to an understanding of how chemical reactions involving gases occur – knowledge that is vital in industry.

Stimulus questions

How are gases compressed?

What is the composition of the atmosphere?

How can pure gases such as oxygen and argon be obtained from the atmosphere?

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15.1 Gases

The relationship between volume, number of moles and molar volume

Gas volume

A gas is a substance – an element, compound or a mixture – that exists in a state with no defined volume or shape.

gas
a substance (element, compound or mixture) that exists in a state with no defined volume or shape

When considering the volume of gases, it is important to note the difference between the volume of the gas particles and the volume of the container in which the gas is held. The volume of the gas particles is fixed but, because the gas particles are so far apart, they can be brought together (compressed) into a smaller volume. When scientists refer to the volume of a gas, they are referring to the volume of the container. Under normal conditions, because gases expand freely to fill any available space, the total volume of space taken up by the gas particles themselves is a tiny fraction of the total volume of the container (Figure 15.1.1).

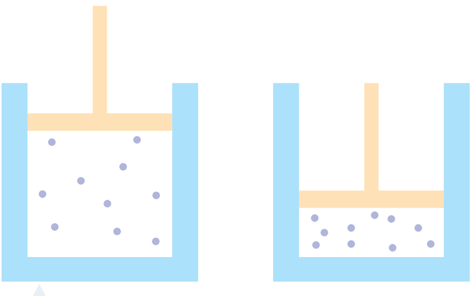


FIGURE 15.1.1 Gases take up the volume of their container. The containers have different volumes but there is the same amount of gas in each one.

Number of moles of gas and volume

For chemists, it is much more convenient to think in terms of the number of gas particles present in a container, rather than just its volume, because the number more accurately represents the amount of matter or mass of that gas.

In 1811, Amedeo Avogadro attempted to show how the volume of a gas was related to the amount of gas present. He proposed that equal volumes of gas contain equal numbers of particles at the same temperature. This became known as Avogadro's hypothesis. Avogadro's hypothesis states that equal volumes of any gas, measured at the same temperature and pressure, contain the same number of particles. This means that, regardless of the composition of the gas, the total number of particles would be the same under the same conditions. So, provided temperature and pressure are the same, 1 L of hydrogen gas has the same number of particles as 1 L of helium gas, or 1 L of oxygen and so on. Also, if the number of particles were doubled, then the volume would double.

Avogadro's hypothesis
equal volumes of any gas, measured at the same temperature and pressure, contain the same number of particles

Standard conditions

To compare changes in conditions that involve gases, scientists use a set of standard conditions for temperature and pressure. In 1982, the International Union of Pure and Applied Chemistry (IUPAC) established standard temperature and pressure (STP) as a temperature of 273.15 K (0°C) and an absolute pressure of 100.00 kPa (0.987 atm or 1.000 bar). This contrasts with the old standard of 0°C and 101.325 kPa (1 atm). The currently accepted standard conditions (STP and SLC) are defined in Table 15.1.1.

standard temperature and pressure (STP)
0°C and 100 kPa

TABLE 15.1.1 Standard conditions for temperature and pressure

STANDARD CONDITIONS	TEMPERATURE (°C)	PRESSURE (kPa)
Standard temperature and pressure (STP)	0	100
Standard laboratory conditions (SLC)	25	100

Molar volume

A mole is defined as a fixed number of particles, so it follows that equal volumes of different gases, measured at the same temperature and pressure, contain the same number of moles. This is called the **molar volume of a gas**.

The units of Avogadro's constant may be electrons, atoms, ions or molecules, depending on the nature of the substance.

- At standard temperature and pressure (STP) (0°C and 100 kPa) the molar volume of a gas is 22.71 L.
- At **standard laboratory conditions (SLC)** (25°C and 100 kPa) the molar volume of a gas is 24.79 L.

molar volume of a gas the volume occupied by 1 mole of gas at a known temperature and pressure

standard laboratory conditions (SLC) 25°C and 100 kPa

SYLLABUS LINK

Chapter 12 discusses Avogadro's constant, about 6.02×10^{23}.

KEY FORMULA

The relationship between number of moles and volume of gas at STP can be expressed as:

$$\text{Number of moles of gas} = \frac{\text{volume of gas at STP}}{\text{volume of one mole of gas at STP}}$$

$$n = \frac{V}{22.71}$$

Where: n = the number of moles

V = the volume of gas (L) at STP.

WORKED EXAMPLE 15.1.1

QUESTIONS

A How many moles of carbon dioxide gas are present in 350 mL of gas at STP?

B What is the mass of carbon dioxide in part A?

C What volume is occupied by 160 g of helium at STP?

ANSWER A

1. Write the relationship between the number of moles of a gas (n), gas volume (V, in L) and volume of 1 mole of gas at STP (22.71 L) using the formula:

$$n = \frac{V}{22.71}\ \text{mol}$$

2. Determine the quantities and check the units: $n = ?$ $V = \text{mL} = 0.350\ \text{L}$
3. Substitute in the formula:

$$n = \frac{0.350}{22.71}$$

So, $n = 0.154$ mol.

ANSWER B

1 Write the relationship between number of moles (n), mass (m) and molar mass (M) using the formula:

$$n = \frac{m}{M}$$

2 Determine the quantities of n and M. From part A, $n = 0.0154\,\text{mol}$:

$$\begin{aligned} M(CO_2) &= 12\text{ g (C)} + (2 \times 16)\text{ g } (O_2) \\ &= 44\text{ g mol}^{-1} \end{aligned}$$

3 Substitute in the equation:

$$0.0154\text{ mol} = \frac{m}{44}$$

$$m = n \times M$$

So, $m = 0.0154\text{ mol} \times 44\text{ g mol}^{-1} = 0.678\text{ g}$.

ANSWER C

1 Write the relationship between the number of moles of a gas (n), gas volume (V, in L) and volume of 1 mole of gas at STP (22.71 L) using the formula:

$$n = \frac{V}{22.71}\text{ mol}$$

2 Determine the quantities and check the units: $n = ?$ $V = ?$ $m = 160\,\text{g}$

3 Calculate the number of moles:

$$n = \frac{m}{M}$$

$$m = 160\,\text{g},\ M(\text{He}) = 4.00\,\text{g mol}^{-1}$$

$$n = \frac{160}{4.00} = 40\,\text{mol}$$

4 Calculate the volume:

$$n = \frac{V}{22.71}$$

$$40 = \frac{V}{22.71}$$

So, $V = 40 \times 22.71\,\text{L} = 908\text{ L}$.

SECTION REVIEW 15.1

REMEMBERING

1 Define:

a gas b Avogadro's hypothesis c molar volume.

APPLYING

2 If the volume of 0.13 mol of a gas at room temperature is 385.9 mL and the volume of the same amount of gas at the same temperature is 35.9 mL, what condition must be different? Explain your answer.

3 a What volume container would you need to store 0.095 mol of nitrogen at STP?

b i How many moles are present in 0.450 L of oxygen gas at STP?

ii What is the mass of oxygen in part **i**?

c i What volume will 10 g of hydrogen gas occupy at STP?

ii What volume will 10 g carbon dioxide occupy at STP?

iii Explain why these volumes differ when the mass is the same.

15.2 Kinetic theory

The pressure exerted by gases is one of the many similar physical properties they exhibit. Experimental observations by chemists that different gases behave in a similar manner led to the development of mathematical relationships linking volume, mass, pressure and temperature. These relationships are based on the notion of an **ideal gas**; that is, one that obeys the relationships perfectly. While no real gas behaves in a completely ideal manner, most gases are considered as being sufficiently close to ideal under normal conditions that these relationships can be used.

ideal gas
a gas that perfectly obeys all the proposals of the kinetic theory of gases and the gas laws

The kinetic theory of gases explains much of the physical behaviour of gases. The term 'kinetic' comes from the Greek *kinetikos*, which means 'moving', and indeed many of the properties of gases relate to their motion.

15.2.1 Kinetic theory
15.2.2 Kinetic theory of gases

The kinetic theory of gases proposes that:

1. gases consist of molecules (except the noble gases, which consist of atoms) that move in continual random straight-line motion.
2. the average distance between gas molecules is very large compared to the size of the molecule.
3. intermolecular forces between gas molecules are negligible because of the relatively large distances between molecules
4. all collisions of gas molecules are perfectly **elastic collisions**, which means no net energy is lost during these collisions.
5. pressure is due to collisions of the molecules with the walls of the container
6. temperature is a measure of the average kinetic energy of the molecules.

elastic collision
a collision with no net loss of energy

Note that this theory uses the model of an ideal gas. Gases, such as hydrogen and helium, which have small, light particles come close to 'ideal' under conditions of low pressure and high temperature. Chemists apply kinetic theory to real gases under normal conditions.

Considering temperature

The kinetic theory of gases states that temperature is a measure of average kinetic energy. Although the molecules of a sample of gas have the same average kinetic energy, the individual molecules move at various speeds. Some move faster and some move slower than the average. Collisions between particles change individual speeds.

However, at a given temperature, molecules of all gases, no matter what size, shape or mass, have the same average kinetic energy. Kinetic energy is given by the formula:

KEY FORMULA

$$\text{Kinetic energy} = \frac{1}{2}mv^2$$

Where:

m = mass

v = speed

Where m is the mass (kilograms) and v is the speed of the molecules (metres per second). The unit of kinetic energy is the Joule (J).

As the temperature increases, the kinetic energy of the particles increases, and vice versa. Although the gas particles have a range of speeds at any given temperature, the average speed of the particles of a gas increases as temperature increases (Figure 15.2.1).

This means that decreasing the temperature will result in the particles slowing down. Theoretically, at low enough temperatures, the particles will eventually stop moving. The theoretical lowest temperature possible is known as **absolute zero** (0K), which is equivalent to −273.15°C.

absolute zero
precisely 0 K on the Kelvin scale, which is a thermodynamic (absolute) temperature scale and −273.15 degrees Celsius on the Celsius scale

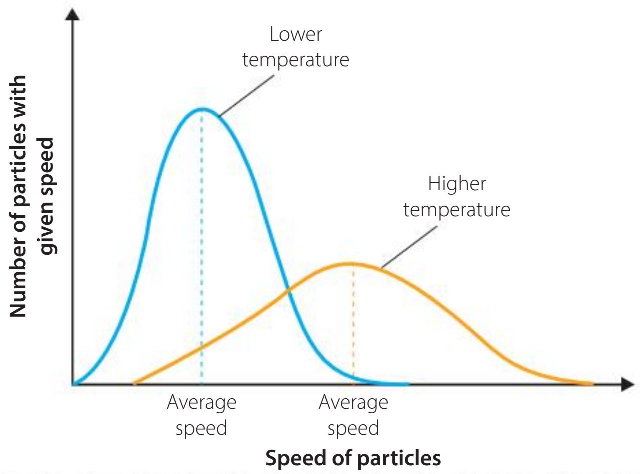


FIGURE 15.2.1 Average speed of chemicals at two different temperatures

Pressure

According to the kinetic theory of gases: *pressure is due to collisions of the molecules with the walls of the container*. Gas pressure is a measure of the force per unit area exerted by gas particles on the walls of their container.

The International System of Units (SI) unit of gas pressure is the **Pascal (Pa)**; 1Pa is equivalent to a force of 1 newton per square metre. This is a very small quantity, so hectopascals and kilopascals are more commonly used for measuring pressures.

Pascal (Pa)
the SI unit of gas pressure, equivalent to a force of 1 newton per square metre

The gases in the atmosphere exert pressure on every exposed surface. The pressure due to the atmosphere is called atmospheric pressure. This pressure varies from place to place and time to time. At sea level, atmospheric pressure is like having a 1-kilogram tub of margarine pushing on every square centimetre of every part of the surface of your body all the time. Usually, it goes unnoticed because the external air pressure exerted by the atmosphere is balanced by the internal pressure exerted outwards from the body's tissues and blood. To make communication easier, scientists agreed on a standard of pressure to represent the average air pressure at sea level. This unit of standard atmospheric pressure is 1.10325×10^5 Pa or 101.325kPa (kilopascals) at sea level. Another common unit for pressure is the atmosphere (atm): 1atm = 101.3kPa.

Atmospheric pressure varies with height above sea level, temperature and weather conditions, so it is not always 1.00atm. Meteorologists generally mention high and low pressure systems in weather reports and use the hectopascal (1 ha = 100Pa) or millibar (100Pa) units when discussing pressure. Chemists and physicists generally use the units Pascal (or kiloPascal) and atmosphere (Table 15.2.1).

TABLE 15.2.1 **Commonly used units of gas pressure**

NAME OF UNIT	SYMBOL	COMPARISON
Atmosphere	atm	101.3 kPa
Hectopascal	hPa	100 Pa
Kilopascal	kPa	1000 Pa
Millimetres of mercury	mmHg	760 mmHg = 1atm
Pascal	Pa	Standard unit, 1 newton per square metre ($N\,m^2$)

Considering volume

The space occupied by the molecules of a gas is its volume. Volume can be measured using different units (Table 15.2.2). The SI unit of volume is the cubic metre (m^3). Chemists generally use the units of litre (L).

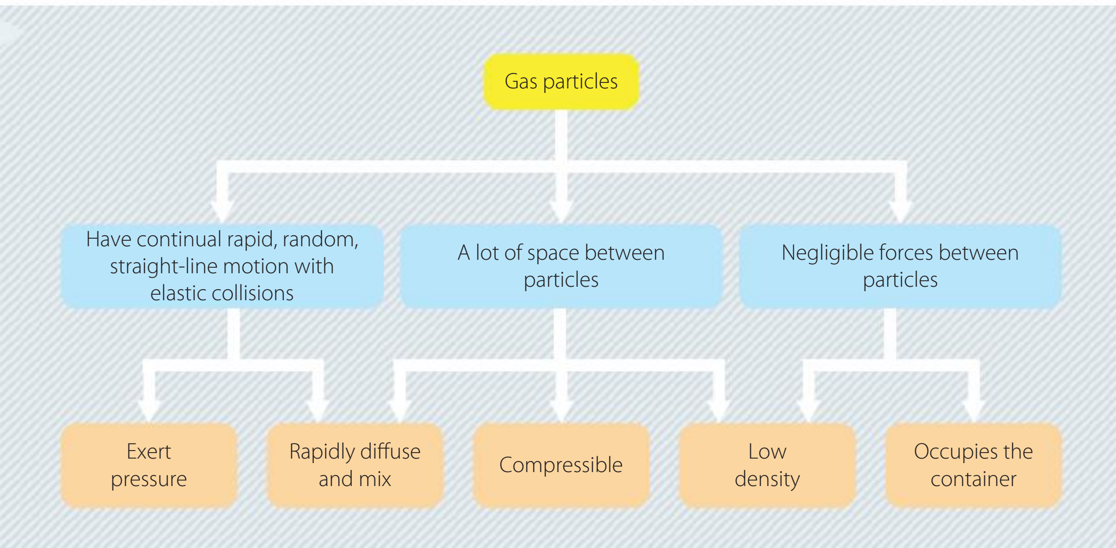
TABLE 15.2.2 Commonly used units of volume

NAME OF UNIT	SYMBOL	COMPARISON
Cubic metre	m^3	Standard unit
Cubic decimetre	dm^3	$1000\ dm^3 = 1\ m^3$
Cubic centimetre	cm^3	$1000\ cm^3 = 1\ dm^3$
Litre	L	$1\ L = 1\ dm^3$
Millilitre	mL	$1\ mL = 1\ cm^3$

Applying the kinetic theory

The kinetic molecular theory was developed using the findings of many different scientists. It can be used to explain many properties of gases (Figure 15.2.2).

FIGURE 15.2.2 A schematic diagram of the kinetic molecular theory of gases



Pressure, temperature and volume relationships

15.2.3 Pressure and temperature relationship of a gas

Consider a gas placed in a piston (Figure 15.1.1, page 285). If the piston is made smaller, the gas particles occupy a smaller space. If the temperature is kept constant then, according to the kinetic molecular theory, the gas particles will move randomly at the same speeds. More collisions will occur per unit area of the walls of the piston. In other words, the pressure will increase.

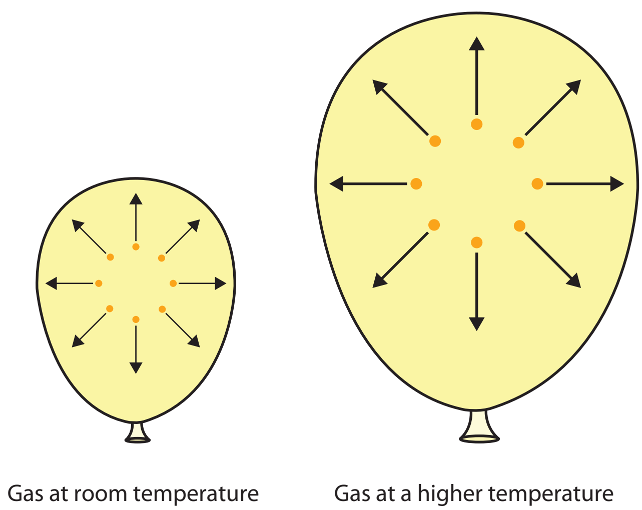


FIGURE 15.2.3 The relationship between temperature and volume for an enclosed gas where the pressure remains constant

Consider a balloon filled with gas at room temperature (Figure 15.2.3). If the temperature of the balloon is increased then, according to the kinetic molecular theory, the gas particles will move faster and the force per unit area exerted by the gas particles will increase. The volume of the balloon will increase.

The relations between pressure, temperature and volume are encapsulated quantitatively in what have become known as the gas laws.

Gas laws

Boyle's law

More than 300 years ago, Robert Boyle was the first chemist to perform quantitative experiments on gases. By measuring the pressures and volumes of many different gases, he observed that, at a constant temperature, the volume of a given mass of gas is inversely proportional to the pressure. This is known as Boyle's law.

15.2.4 Boyle's law

In mathematical terms, Boyle's law can be expressed as V ~ 1/P, which means the greater the applied pressure, the smaller the volume.

Another way to state Boyle's law is PV 5k, where k is a constant. For a particular sample of gas at a constant temperature, the product of the pressure and volume is a constant value.

A useful alternative form of Boyle's law is:

$P_1V_1 = P_2V_2$

Where:

P_1 is the initial volume

V_1 is the initial volume

P_2 is the final volume

V_2 is the final volume.

Note: Any units for pressure and volume can be used, as long as they are consistent.

KEY LAW

Boyle's law states that, at a constant temperature, the volume of a given mass of gas is inversely proportional to the applied pressure.

WORKED EXAMPLE 15.2.1

A sample of gas collected in a 242 mL container has a pressure of 87.6 kPa. What would the volume of this gas be at 101.3 kPa? (Assume temperature remains constant.)

ANSWER

The volume would be 209 mL.

1 Write the relationship:

$$P_1V_1 = P_2V_2$$

2 Determine the quantities:

$$P_1 = 87.6\,\text{kPa},\ P_2 = 101.3\,\text{kPa}$$
$$V_1 = 242\,\text{mL},\ V_2 = ?$$

3 Check the units are consistent.

4 Rearrange the formula:

$$V_2 = \frac{P_1V_1}{P_2}$$

Substitute quantities:

$$V_2 = \frac{87.6 \times 242}{101.3}$$

So, $V_2 = 209\,\text{mL}$.

15.2.5 Charles' law

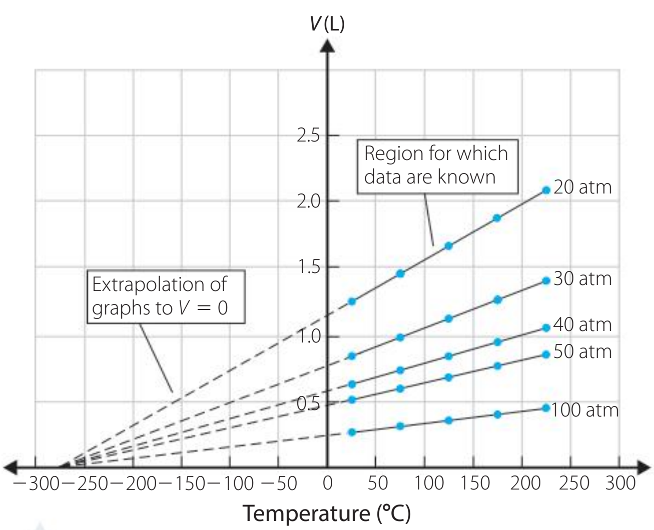


FIGURE 15.2.4
Graph showing how the volume of 1 mole of hydrogen gas varies with temperature for a number of different pressures.

Charles' law

About 100 years after Boyle investigated the relationship between pressure and volume, French physicist Jacques Charles studied the relationship between temperature and volume using gases. The results of his many experiments led him to determine that for a fixed quantity of gas at a constant pressure, the volume increases linearly with temperature. This relationship is called Charles' law. When experimental temperature and volume data are plotted, a linear relationship emerges (Figure 15.2.4).

Extrapolation of the data ($V = 0$) leads to a temperature of −273°C, now known as absolute zero. Charles's experiments led to the development of a new temperature scale. If the data is plotted using the absolute temperature scale, all lines pass through the origin of the graph, so temperature (K) is directly proportional to volume. Furthermore, in calculations involving gases, temperature must be given in kelvin.

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In mathematical terms, Charles' law can be expressed as $V \alpha T$ or:

$$\frac{v}{t} = \text{constant}$$

An alternative form of this equation useful for calculations is:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

where V_1 is the initial volume
T_1 is the initial temperature (K)
V_2 is the final volume
T_2 is the final temperature (K).

Note that any unit for volume can be used, as long as it is used consistently, but temperature must be in kelvin.

KEY FORMULA

Charles' law states that at constant pressure, the volume of a fixed quantity of gas is proportional to its absolute (kelvin) temperature.

WORKED EXAMPLE 15.2.2

A 225 mL volume of gas is collected at 58°C. What volume would this sample occupy at 25°C?

ANSWER

The volume would be 202 mL.

1 Write the relationship:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

2 Determine the quantities and convert temperature to kelvin by adding 273:

$$V_1 = 225\,\text{mL},\ V_2 = ?$$
$$T_1 = 58°\text{C} = 331\,\text{K},\ T_2 = 25°\text{C} = 298\,\text{K}$$

3 Check the units are consistent.

4 Rearrange the formula:

$$V_2 = \frac{V_1 V_2}{T_1}$$

5 Substitute quantities:

$$V_2 = \frac{225 \times 298}{331}$$

So, $V_2 = 202\,\text{mL}$.

Combined gas law

When dealing with gases, temperature and pressure typically change at the same time. By combining Boyle's law and Charles' law, a new relationship called the combined gas law is produced. This gives a relationship between volume, pressure and absolute temperature for a given mass of gas. It is most useful in the following form:

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

where P_1, V_1 and T_1 represent initial pressure, volume and temperature (K), and P_2, V_2 and T_2 represent final pressure, volume and temperature (K).

The temperature must be measured in kelvin and the units for volume and pressure must be consistent.

WORKED EXAMPLE 15.2.3

A sample of gas at a pressure of 105 kPa has a volume of 2.50 L at a temperature of 25°C. What pressure is needed to compress it to 1.0 L at 20°C?

ANSWER

The pressure would be 258 kPa.

1 Write the relationship:

2 Determine the quantities and convert temperature to kelvin by adding 273:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$P_1 = 105\,\text{kPa},\ P_2 = ?$
$V_1 = 2.50\,\text{L},\ V_2 = 1.0\,\text{L}$
$T_1 = 25°\text{C} = 298\,\text{K},\ T_2 = 20°\text{C} = 293\,\text{K}$

3 Check the units are consistent.

4 Rearrange the formula:

$$P_2 = \frac{P_1V_1T_2}{T_1V_2}$$

5 Substitute quantities:

$$P_2 = \frac{105 \times 2.5 \times 293}{298 \times 1.0}$$

So, $P_2 = 258\,\text{kPa}$.

15.2.6 Ideal gas law

The general gas equation ('ideal' gas law)

All the gas laws studied so far have required that at least one of the conditions of pressure, temperature, volume and quantity of gas remain constant. In reality, when dealing with gases in situations such as scuba diving and using anaesthetics, all the conditions can vary. However, if Boyle's and Charles's laws are combined with Avogadro's hypothesis, the resulting relationship, called the general gas equation, includes all conditions as variables.

The general gas equation is also known as the ideal gas law and the universal gas equation. It provides a relationship between pressure, volume, temperature and the amount of gas:

$$PV = nRT$$

KEY FORMULA

The general gas equation, $PV = nRT$, gives the relationship between pressure, temperature, volume and number of moles of a gas.

where P is pressure in kPa
V is volume in L
n is total number of moles of gas present
R is the universal gas constant = 8.31 J K^{-1} mol^{-1}
T is temperature in K.

Regardless of the chemical formula and properties of different gases, the general gas equation describes the physical properties of all gases. The equation is called the ideal gas equation because real gases deviate slightly from it, particularly at high pressures and low temperatures, because the intermolecular forces between real gases come into play as the particles slow down and move closer together.

It is important to use the correct value of R in the general gas equation. Always include the unit and the number when recording the value of a physical quantity to reduce mistakes.

The value of R, the universal gas constant, is dependent on units. The value above ($8.31\,J\,K^{-1}\,mol^{-1}$) is based on an SI unit for pressure. Sometimes, it is more convenient to use pressure in atmospheres. When the unit for pressure is atmospheres, the value of R is $0.0820\,L\,atm\,K^{-1}\,mol^{-1}$. Table 15.2.3 shows the units and values for R.

TABLE 15.2.3 Units and values for the universal gas constant, R

CONDITION	$R = 8.314\,J\,K^{-1}\,mol^{-1}$	$R = 0.082\,L\,atm\,K^{-1}\,mol^{-1}$
Pressure	kPa	atm
Volume	L	L
Temperature	K	K
Amount of gas	mole	mole

WORKED EXAMPLE 15.2.4

What is the mass of carbon dioxide in a sample that occupies 1.50 L at 20°C and 125 kPa?

ANSWER

The mass is 3.39 g.

1 Write the relationship:

$$PV = nRT$$

2 Determine the quantities and check the units.

$$m = ?,\ P = 125\,kPa,\ V = 1.50\,L,\ n = ?,\ T = 20°C = 293\,K$$

3 Decide on the appropriate value for R based on the units for P.

$$R = 8.314\,J\,K^{-1}\,mol^{-1}$$

4 Rearrange the formula:

$$n = \frac{PV}{RT}$$

5 Substitute quantities:

$$n = \frac{125 \times 1.5}{8.314 \times 293}$$

So, $n = 0.077\,mol$

6 Calculate mass from moles

$$n = \frac{m}{M} \therefore m = n \times M$$

Where: $M(CO_2) = 44$

So, $m = 0.077 \times 44 = 3.39\,g$.

SECTION REVIEW 15.2

REMEMBERING

1 List the proposals of the kinetic theory.

UNDERSTANDING

2 Explain why an increase in temperature of a gas-filled balloon would cause it to burst.

3 Explain why the particles of hydrogen gas move faster than those of oxygen gas at the same temperature.

APPLYING

4 Hydrogen chloride gas and ammonia gas react according to the equation:

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

The ammonium chloride produced is present as thick white smoke.

In the experiment depicted in Figure 15.2.5, both gases were released at the same time. Predict where you think the thick white smoke will form. Explain your answer.

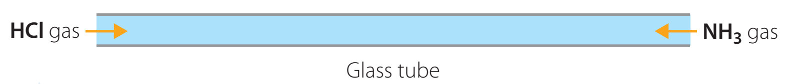


FIGURE 15.2.5 Apparatus for demonstrating the ammonia/hydrogen chloride reaction

5 Assuming the temperature is constant, carry out the following calculations.

- **a** If 4.42 L of gas was collected at a pressure of 94.2 kPa, what volume would it occupy at 1 atm pressure?
- **b** A 210 mL sample of nitrogen gas at 1.0 atm is compressed until the final volume is 150 mL. Determine its final pressure.
- **c** When the pressure is 522 kPa, a gas has a volume of 55 mL. Calculate its volume at atmospheric pressure.
- **d** A sample of hydrogen gas at 73.3 kPa in a 2.5 L flask is expanded into a total volume of 7.2 L. Calculate the final pressure.

6 Assuming pressure is constant, carry out the following calculations.

- **a** A sample of oxygen has a volume of 3.50 mL at 20°C. What volume would it occupy at 80°C?
- **b** If a sample of gas had a volume of 1.7 L at 25°C and 1 atm pressure, what:
  - **i** would its volume be at 100°C?
  - **ii** temperature would it need to be cooled to have a volume of 0.80 L?
- **c** A gas occupies a volume of 60.0 mL at 36°C. What volume would the gas occupy at 0°C?

7 Assuming pressure is constant, carry out the following calculations.

- **a** What is the volume of 0.34 mol of H_2S gas at 100°C and 75.0 kPa pressure?
- **b** A sample of oxygen gas is collected and found to occupy a volume of 527 mL at 1.05 atm and 20°C. What is the mass of the sample?
- **c** What is the pressure of 4.00 g of He in a container of volume 7.5 L at 30°C?

15.3 Inquiry skills: using mathematical representations

Chemical reactions involving gases

In Chapter 12, mass calculations based on amounts of reactants and products were considered. The calculations needed to use the mole relationships given by coefficients in a balanced chemical equation are used to determine the amounts of reactants and products. This relationship can now be extended to include volumes of gases (Figure 15.3.1 and Worked example 15.3.1).

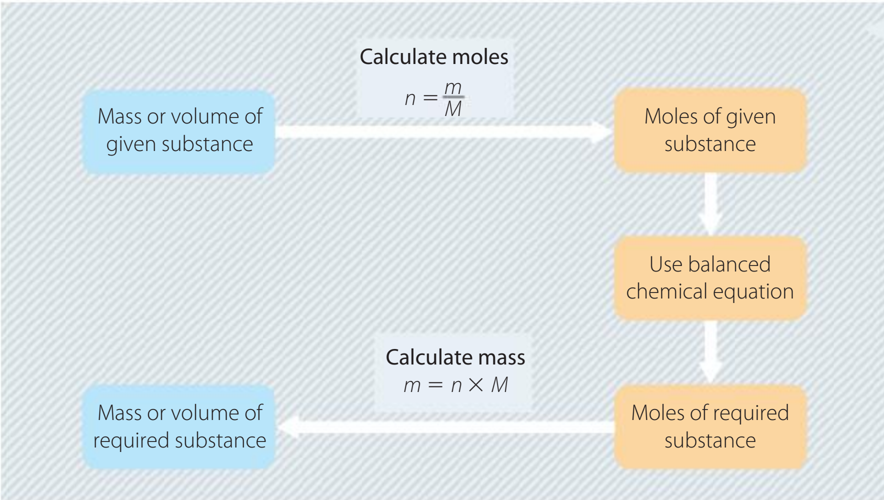


FIGURE 15.3.1 Mass–volume relationships, where m is mass in grams, n is number of moles and M is molar mass

WORKED EXAMPLE 15.3.1

Consider the reaction between an excess of dilute hydrochloric acid [HCl(aq)] and 0.325 g of zinc at STP. You already know how to calculate the mass of hydrogen (H_2) produced. By adopting the procedure outlined in Figure 15.3.1, it is now possible to determine the *volume* of the gas produced.

QUESTION

What volume of hydrogen is produced when 0.325 g of zinc metal reacts with an excess of hydrochloric acid at STP?

ANSWER

1 Write the balanced equation:

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

2 Convert mass of Zn to moles:

$$\text{Moles of zinc } (n) = \frac{0.325}{65.4} = 4.97 \times 10^{-3} \text{ moles}$$

3 Determine the mole ratio from the balanced equation, and the number of moles of $H_2(g)$ produced:

$$1 \text{ mole Zn} : 1 \text{ mole } H_2$$

4.97×10^{-3} moles of Zn produce 4.97×10^{-3} moles of $H_2(g)$.

4 Use Avogadro's hypothesis to determine the volume of $H_2(g)$ produced:

$$\text{Number of moles of gas} = \frac{\text{volume of gas at STP}}{\text{volume of one mole of gas at STP}}$$

$$n = \frac{V}{22.71}$$

If one mole of $H_2(g)$ at STP has a volume of 22.71 L and there are 0.076 moles of $H_2(g)$, then:

$$V = 4.97 \times 10^{-3} \text{ moles} \times 22.71 \text{ L} = 113 \text{ mL}.$$

WORKED EXAMPLE 15.3.2

Ammonia (NH_3) is produced by mixing nitrogen (N_2) and hydrogen (H_2). What volume of hydrogen is required to react with an excess of nitrogen to produce 3.95 L of ammonia at STP?

ANSWER

1 Write the balanced equation for the reaction:

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

2 Convert volume of $NH_3(g)$ to moles.

$$n = \frac{V}{22.71} = \frac{3.95}{22.71} = 0.174 \text{ mol}$$

3 Determine the mole ratio from the balanced equation and the number of moles of H_2:

$$3 \text{ moles } H_2(g) : 2 \text{ moles } NH_3(g)$$

If there are 0.174 moles of NH_3 then:

$$\frac{3 \text{ moles } H_2}{2 \text{ moles } NH_3} = \frac{x \text{ moles } H_2}{0.174 \text{ moles } NH_3}$$

$$= \frac{3}{2} \times 0.174$$

$$= 0.261 \text{ moles } H_2(g) \text{ required}$$

$n = 0.261$ moles of H_2 required.

4 Use Avogadro's hypothesis to determine the volume of $H_2(g)$ required.

$$0.261 = \frac{V}{22.71}$$

$$0.261 \times 22.71 = 5.93 \text{ L of } H_2(g) \text{ required}$$

The volume of $H_2(g)$ required is 5.93 L.

SECTION REVIEW 15.3

APPLYING

1 A mass of 4.86 g of magnesium is ignited in nitrogen dioxide. The products of the reaction are magnesium oxide and nitrogen gas. Calculate the volume of nitrogen gas produced at 273 K and 101.3 kPa.

2 The complete combustion of carbon in oxygen produces carbon dioxide. Calculate the volume of oxygen that would react at SLC with 10.0 g of carbon and the volume of carbon dioxide produced under the same conditions.

3 Hydrogen peroxide (H_2O_2) decomposes to give water and oxygen gas. How many grams of hydrogen peroxide are needed to produce 100 L of oxygen at SLC?

15.4 Mandatory practical

PRACTICAL ACTIVITY 15.4.1

Estimating the molar volume of a gas

In section 15.1 you learned about Avogadro's hypothesis, which states that equal volumes of any gas, measured at the same temperature and pressure, contain the same number of particles. In other words, at STP 1 mole of any gas occupies 22.4 L and, at SLC, 1 mole of any gas occupies 24 L. If a gas-producing reaction is carried out in the laboratory, it is possible to collect the gas and measure its volume. Using Avogadro's hypothesis, the molar volume of the collected gas can be estimated.

AIM

To estimate the molar volume of a gas

EQUIPMENT PER GROUP

- boiling tube
- rubber bung fitted with glass tube
- mass balance
- retort stand and clamp
- 100 mL measuring cylinder
- 0.50 g powdered calcium carbonate
- 150 mL of 1 mol L^{-1} ethanoic (acetic) acid
- 50 mL measuring cylinder
- test tube

WHAT ARE THE RISKS IN DOING THIS ACTIVITY?	HOW CAN THESE RISKS BE MANAGED?

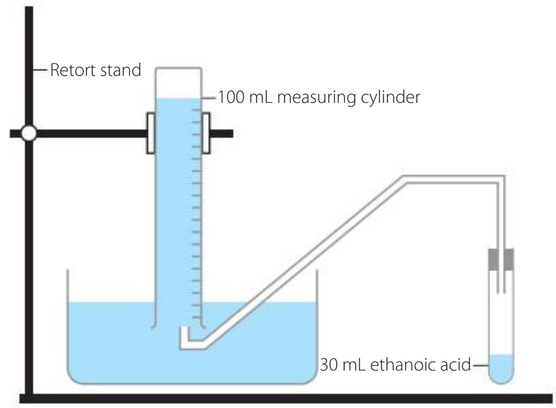
RISK ASSESSMENT

Copy and complete the risk assessment table. Write down any risks you can think of, as well as how to manage them. Show your table to your teacher before you proceed.

PROCEDURE

1. Set up the apparatus as shown in Figure 15.4.1.
2. Place 30.00 mL ethanoic acid in the boiling tube.

FIGURE 15.4.1 Experimental set-up for estimating the molar volume of a gas



3. Place a weighing boat (or filter paper) on a mass balance and zero the balance (press tare).
4. Place about 0.10 g calcium carbonate onto weighing boat and accurately record the mass.
5. Remove the bung from the boiling tube and carefully pour the calcium carbonate into the tube.
6. Quickly replace the bung.
7. When the reaction is complete, record the volume of gas in the measuring cylinder.
8. Repeat the experiment four times, increasing the mass of calcium carbonate by 0.10 g each time.

RESULTS

Copy the following table, inserting a new row for each experiment, and use it to record your results.

EXPERIMENT NUMBER	MASS OF CALCIUM CARBONATE (g)	VOLUME OF CARBON DIOXIDE (mL)
1		

ANALYSING THE RESULTS

1. Write a balanced equation for the reaction between calcium carbonate and ethanoic acid.
2. Graph volume of carbon dioxide (y axis) against the mass of calcium chloride.
3. What is the mathematical relationship of the regression line? How can the information provided by the regression line help you determine the volume of one mole of carbon dioxide?
4. Draw a line of best fit passing through the established data points.
5. Does your graph go through the origin? If it does, what does this suggest?
6. For each experiment that you carried out, determine the volume of one mole of carbon dioxide (the experimental *molar volume*).
7. Using the equation you recorded in step 1, determine the number of moles of carbon dioxide produced in each experiment to calculate the volume of one mole of carbon dioxide (the theoretical *molar volume*).

DISCUSSION

1. Compare the experimental molar volume of carbon dioxide produced in each experiment with the theoretical molar volume of carbon dioxide and calculate the percentage error in each case.
2. Within the experimental percentage error of your investigation, does your value for molar volume agree with the theoretical value?
3. Identify possible sources of error and describe how these would affect the result.
4. For each source of error, suggest how to improve the experiment to minimise these errors.

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 Define the following terms.

a STP

b SLC

CATEGORY QUESTIONS

2 Use the kinetic theory to explain why gases:

a can be compressed

b can mix rapidly

c have low density.

3 Explain why the general gas equation cannot be applied to gases under conditions of high pressure and low temperature.

EVIDENCE QUESTIONS

4 You have been advised that windows should be left slightly ajar as a cyclone approaches. Provide evidence to support or refute this claim.

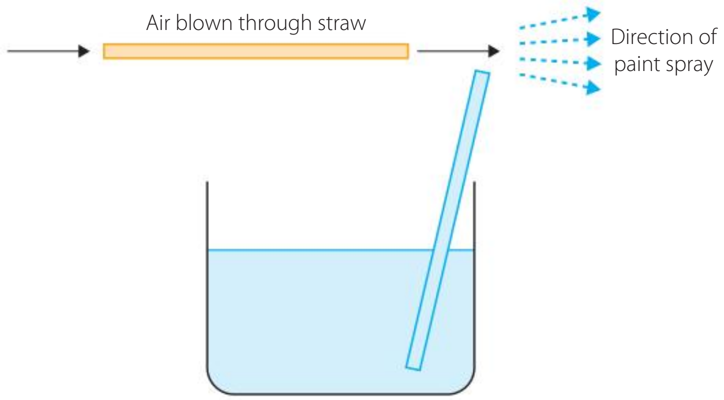
5 The user manual of a car suggests adding more air to the tyres in winter. Provide evidence to support or refute this claim.

6 In recent years it has been fashionable for people with high-performance cars to have the car tyres filled with nitrogen rather than air. Given that the composition of air is approximately 78% nitrogen and 21% oxygen with small amounts of gases, such as argon, carbon dioxide and nitrogen oxides, and an average of 1% water vapour at sea level, suggest why:

a this practice might seem sensible

b motor-racing teams fill their car tyres with air.

7 It is possible to make a very crude air brush using the apparatus shown below:



This is an example of Bernoulli's principle.

Using the diagram and your understanding of the behaviour of gases, suggest what Bernoulli's principle might be.

ELABORATION QUESTIONS

8 A car with typical petrol consumption of about 14 km L^{-1} will produce about 0.16 kg of carbon dioxide per kilometre.

a If the typical annual distance travelled is 20 000 km, how much carbon dioxide will it produce in a year?

b There are about 1 billion (10^9) motor vehicles in the world. How many kilograms of carbon dioxide do they produce per year?

c This is a lot of carbon dioxide but how does it compare with the amount of CO_2 in the atmosphere? The volume of CO_2 in the atmosphere is approximately 1.8 × 1018 L.

d Assuming the temperature is 0°C, calculate the mass of CO_2 in the atmosphere.

e Using your answer from part b, calculate the fraction of CO_2 that cars contribute to the atmosphere.

9 What is the difference between a real gas and an ideal gas?

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END-OF-CHAPTER EXAM

End-of-chapter test

CHEMISTRY

1 The pressure of a gas will __________ when the volume is decreased and will __________ when the absolute temperature is decreased.

A decrease ... decrease

B increase ... decrease

C decrease ... increase

D increase ... increase

2 Which one of the following statements is *not consistent* with the kinetic molecular theory of gases?

A The actual volume of the gas molecules themselves is very small compared to the volume occupied by the gas at ordinary temperatures and pressures.

B The average kinetic energies of different gases are different at the same temperature.

C There is no net gain or loss of the total kinetic (translational) energy in collisions between gas molecules.

D Individual gas molecules are relatively far apart.

3 A gas is held under conditions of STP. It is found that 71.0 g of the gas occupies a volume of 22.71 L under these conditions. What is the gas?

A CO_2

B N_2

C Cl_2

D KI

4 Name two standard units of pressure.

5 How many litres does one mole of a gas at STP occupy, according to Avogadro's hypothesis?

6 What property of an enclosed gas is represented by the force of the gas particles hitting the walls of its container?

7 What is wrong with saying the gas bottle is 'half-full'?

8 On a hot day, a metal can is half-filled with petrol and sealed. The next morning, after a cold night, it is observed that the can is partially crushed. Explain why this could happen.

9 A sample of an unknown gas has a mass of 3.19 g and occupies a volume of 3.97 L at SLC. Identify the gas.

10 How do you convert °C to K?

A Subtract 273

B Add 373

C Add 273

D It depends on the pressure.

11 Which of the following best expresses Boyle's law?

A $V \alpha \frac{1}{P}$

B $T \alpha V$

C $P \alpha V$

D $V \alpha \frac{1}{T}$

12 Gases are more easily compressed than liquids and solids. The reason for this behaviour is that:

A gas molecules move with much greater velocities than the molecules of liquids and solids, permitting gases to adjust more rapidly to a change in volume.

B gas molecules undergo elastic collisions with the walls of a container and elastic substances are easily compressed.

C the average distance between gas molecules is much greater than that between particles in liquids or solids, so the volume may be reduced more easily.

D attractive forces between gas molecules are much smaller than those between particles in solids and liquids, and these small forces can be readily overcome during compression.

13 Which of the following is not one of the assumptions of the kinetic theory of gases?

A All collisions of gas molecules are non-elastic. This means energy is lost during collisions.

B Gases consist of tiny particles called molecules, except for noble gases, which consist of atoms.

C The molecules of a gas exert negligible attractive or repulsive forces on each other.

D The average distance between molecules in a gas is very large compared with the size of each gas molecule.

14 A gas fills a 100 mL cylinder fitted with a piston at a particular temperature and pressure. If the volume of the gas is halved by pushing in the piston, and at the same time the absolute temperature is doubled, the pressure of the gas will:

A increase four times.

B be doubled.

C be unchanged.

D be halved.

15 Which of the following statements about the properties of gases is correct?

A The solubility of a gas in a liquid increases as the liquid is heated.

B One mole of any gas at SLC occupies 24.5 L.

C Larger gas molecules diffuse more quickly than smaller molecules.

D The thermal energy of a gas is given by its temperature.

16 As a man dives into the sea, the pressure of gas in his lungs changes from 100 000 Pa to 150 000 Pa. If his lungs initially held 6 L of gas, the volume of his lungs would become (assume temperature remains constant):

A 3 L.

B 4 L.

C 6 L.

D 9 L.

17 Ammonia was prepared by the following reaction:

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

If 1 L of nitrogen is reacted with excess hydrogen, how many litres of ammonia will be produced?

A 2 L

B 3 L

C 1 L

D 0.5 L

18 Decide whether each of the following statements is true or false. For those that are incorrect, write the correct statement.

a Halving the number of particles in a given volume of gas decreases the pressure by one-half if the temperature is kept constant.

b At absolute zero, the volume occupied by a gas would be zero.

c At constant temperature and pressure, the volume of a gas decreases as the number of moles of the gas increase.

d According to the combined gas law: $T_2 = P_1 \times \frac{V_2 T_1}{P_2 V_1}$

19 A sample of methane has a volume of 4.60 L at 27°C. What volume would it occupy at 120°C if the pressure remained constant?

20 What is the final pressure of gas (at a constant temperature) if 50 mL at 100 kPa expands to 60 mL?

21 An industrial plant manufacturing polythene (polyethylene) stores ethene (ethylene) gas (C_2H_4) in a vessel with a capacity of 2 L and capable of withstanding a pressure of 1500 kPa. What is the maximum mass of ethene that could be stored in the vessel at a temperature of 42°C?

22 350.0 mL of a gas at 26°C and 105.6 kPa has a mass of 0.35 g. Calculate the molar mass of the gas.

23 Carbon monoxide burns in oxygen according to the equation:

$$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$$

Calculate the volume of oxygen required at 25°C and 1.00 atm for the combustion of 28 g of carbon monoxide.

24 Nitroglycerine has a density of 1.59 g mL^{-1}. It explodes to form several gases according to the reaction:

$$4C_3H_5O_9N_3(s) \rightarrow 12CO_2(g) + O_2(g) + 6N_2(g) + 10H_2O(g)$$

A sealed 1.00 mL container is filled with nitroglycerine and detonated. Assuming standard temperature and that the container would not break upon detonation, calculate the pressure inside the container in atmosphere.

» UNIT TWO

MOLECULAR INTERACTIONS AND REACTIONS

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Topic 2: Aqueous solutions and acidity

SCIENCE AS A HUMAN ENDEAVOUR

Students should research sources of acid rain and its effects; the ionic and molecular substances found in blood plasma; water quality across different sources, and how this affects the way water is treated and used.

FORMULAS

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{volume of solvent}}$$

16 AQUEOUS SOLUTIONS AND MOLARITY

Introduction

Water has unique physical and chemical properties. For a small molecule, water has a big influence on life on Earth. Water's unique properties arise partly from the strength of the intermolecular forces between the water molecules. The shape of water molecules affects these forces, and they are important for some practical applications.

Stimulus questions

Water boils at 100°C yet, based on its molecular weight and comparisons with compounds such as hydrogen sulfide (H_2S) and hydrogen selenide (H_2Se), it should boil at about −50°C. Why is this so?

Why is water considered the 'universal solvent'?

How do scientists characterise a 'supersaturated' solution?

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16.1 The structure and properties of water

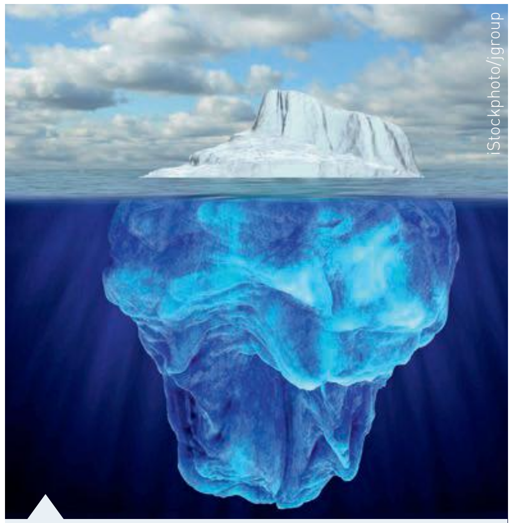

FIGURE 16.1.1 The tip of the iceberg: why do icebergs float in water?

Water is a tasteless, odourless liquid that is essential to our survival. All organisms are at least 60% water. Although water is very common, its particular properties make it an unusual material. These properties have enabled life to evolve on Earth and explain why water is vital for life.

The importance of water was recognised by Empedocles (490–430 BCE). He claimed that water was one of the four elements that formed everything that we know. However, he never realised the reasons why this substance is so vital for life.

Although water has been found beyond Earth, on Mars, Saturn's Moon Titan, and as water vapour surrounding black holes, life is only expected to occur where water exists as a liquid.

Water freezes at 0°C and boils at 100°C. It can act as a solvent over this entire temperature range, dissolving and transporting materials across a cell or over the oceans. It is the only substance on Earth that exists naturally in its three states of matter: solid, liquid and gas. Unlike other substances, solid water is less dense than the liquid. This explains why ice floats, while most other solids sink when mixed with their liquid form (Figure 16.1.1).

The boiling point, **density** in solid and liquid phases, **surface tension** and its ability to act as a solvent are the unique properties of water. These properties are due to the water molecule's shape, bonding within the molecule and the bonds it forms with other substances. Ultimately it is water's bonding that makes it vital in biological, chemical and physical processes on Earth.

density
the mass per unit volume of a substance

surface tension
the tension of the surface film of a liquid, due to the attraction between particles in the surface layer

Water and hydrogen bonding

Three types of bonding occur in water: covalent bonding, hydrogen bonding and dispersion forces. Strong covalent bonding exists between the hydrogen and oxygen within the molecule. Hydrogen bonding and dispersion forces occur between the water molecules. Hydrogen bonding is responsible for the unique properties of water.

There is a negative correlation between intermolecular force strength and **vapour pressure**. Compounds with strong intermolecular forces will have low vapour pressure.

vapour pressure
the pressure of a substance in the gaseous state above its liquid

The unique properties of water include unusual melting and boiling points, density in solid and liquid phases, surface tension and an ability to act as a very good solvent. Hydrogen bonding explains all the properties that relate to temperature. You need a relatively large amount of energy to break the intermolecular bonds between water molecules, which is why water has a relatively high boiling point, 100°C. In comparison, pentane (C_5H_{12}), with its much higher molar mass, has a lower boiling point of 36°C. At room temperature, fewer molecules of water will escape into the air relative to pentane molecules. Fewer vaporised water molecules means lower vapour pressure (the downwards pressure exerted by the gas molecules on the surface of the liquid in a closed container). Energy must be absorbed to change state from liquid to gas. Water has a high latent heat of vaporisation because of the strong intermolecular forces due to hydrogen bonding. Likewise, to increase the average temperature of water requires more heat than for almost any other liquid. This is because other than ammonia, water has the highest specific heat of any liquid. The change of state for water from solid to liquid or liquid to gas is endothermic.

16.1.1 Hydrogen bonds in water

The surface tension of water is high because hydrogen bonding occurs in all directions and is relatively strong. It holds the molecules together (Figure 16.1.2 on page 310).

The cohesive nature of water helps plants take up nutrients from the soil. Root hairs draw up water containing dissolved nutrients, such as nitrogen, phosphorus and potassium, from the soil. **Capillary action** helps the water to move up the xylem of the plant to the leaf surface. Capillary action relies on

capillary action
the ability of a liquid to flow through narrow spaces in opposition to gravity

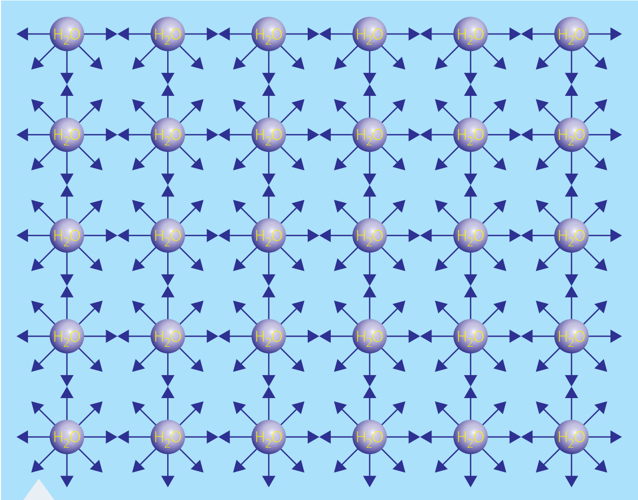

FIGURE 16.1.2 In the bulk of liquid water, each molecule has forces operating in all directions. However, on the surface, there is an imbalance – the liquid molecules are attracted to each other and exert a net force that pulls them together.

the adhesion of water molecules to the sides of the xylem vessel and the cohesive properties of water. The cohesion of water molecules means dissolved nutrients are transported to the top of tall trees, defying gravity. Capillary action is due to three main forces.

1. cohesive force: the intermolecular force between molecules in a substance that helps to maintain a certain shape of the liquid
2. surface tension: this is due to the cohesive forces at the surface of a material and results in the surface of the fluid being under tension
3. adhesive force: these are the forces of attraction between unlike molecules.

Capillary action can be seen when part of a towel is dipped in water – the whole towel eventually gets soaked. Capillary action can also be seen in narrow glass vessels, where adhesive forces between water molecules and the surface of the glass and cohesive forces in water cause the surface of the water to curve to form a meniscus. The water continues to be drawn up, until gravity overcomes the adhesive force.

As with all substances, as water cools, its particles slow down and pack closer together so the density increases. This explains why cooler water has a higher density than warmer water. When the temperature is less than 4°C, the molecules do not have as much kinetic energy and no longer move between the spaces, but form more hydrogen bonds, as shown in Figure 16.1.3. The shape of the water molecule means a regular hexagonal shape is formed. Water has the unique property of expanding as it freezes. The density decreases as it cools from 3°C to 0°C. Ice is about 8% less dense than water (Figure 16.1.3).

Another property of water is its ability to act as a solvent. Pure water is liquid from 0°C to 100°C. This means that it can act as a solvent over this temperature range. Water dissolves and transports a range of materials across the planet and through each cell of a living organism. Water does this by forming hydrogen bonds or ion–dipole bonds with a wide range of substances from salts to organic molecules, such as proteins.

16.1.2 Capillary action and why we see a meniscus

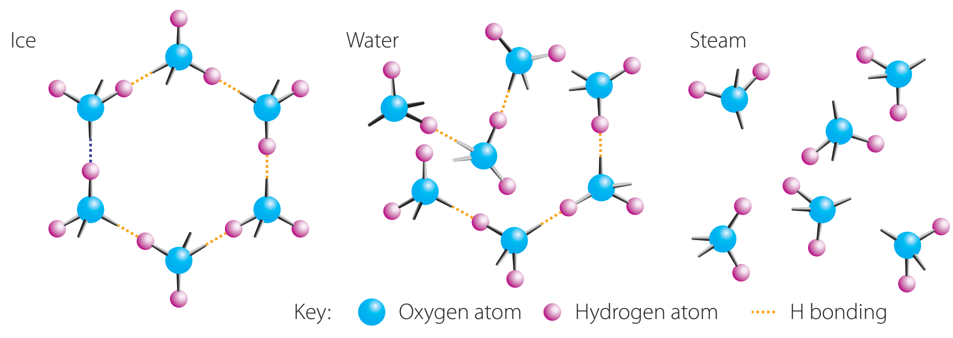

FIGURE 16.1.3 The structure of ice compared to that of liquid water and steam showing hydrogen bonding

water of crystallisation water molecules present in a fixed ratio that form an essential part of the crystal structure of some compounds

hydrated ionic crystals with a specific number of water molecules in their chemical formulas

Water of crystallisation

Water of crystallisation occurs when water molecules are attracted to the ions of a salt. Many ionic compounds that crystallise out of an aqueous solution include water molecules in a regular array through the crystal lattice. The water molecules are in set ratio to the ionic substance. These compounds are called '**hydrated**'. For example, hydrated copper(II) sulfate ($CuSO_4 \cdot 5H_2O$) has five water molecules of

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crystallisation per formula unit. The bonds between the ions and water are called ion–dipole bonds; the ion and the dipole in the water molecules form a weak bond. Water, or other molecules that form dipole bonds with metal atoms, are referred to as ligands. The water ligands take the positions around the central ion and will follow the VSEPR theory (Chapter 13). Hydrated copper(II) sulfate has an octahedral shape. The central ion is a cation, so the negative oxygen of water points towards the copper(II) ion, as shown in Figure 16.1.4.

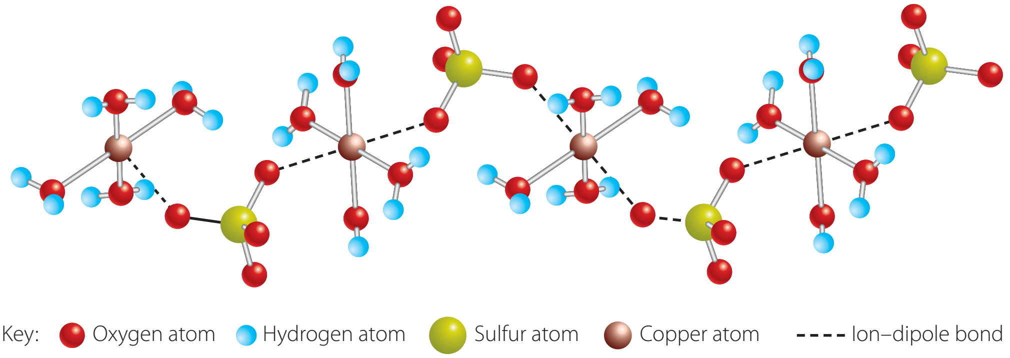

FIGURE 16.1.4 The crystals of copper(II) sulfate consist of an octahedral arrangement of water molecules and sulfate ions around the central copper ion.

The water of crystallisation can be evaporated by heating the hydrated compound. When all the molecules have been removed, the compound is said to be '**anhydrous**'. To do this, a known mass of the hydrated compound is heated so that the mass of water removed can be determined. From this, the empirical formula of the hydrate can be calculated.

anhydrous
ionic solids with no water present

SECTION REVIEW 16.1

REMEMBERING

1 Define:

a density

b surface tension

c vapour pressure

d capillary action

e water of crystallisation.

UNDERSTANDING

2 Give three reasons why water can be considered an unusual substance, despite it being a very common substance.

3 Explain how hydrogen bonding between water molecules can cause the phenomenon of capillary action.

4 Explain why the complex ion in $CuSO_4 \cdot 5H_2O$ has a shape that is described as octahedral.

APPLYING

5 Explain why you should not put full water bottles in the freezer.

solute
a substance that is dissolved in a solvent to form a solution

solvent
a substance (usually a liquid) in which another substance is dissolved

solution
a homogenous mixture with a solute dissolved in a solvent, where the particles of the solute cannot be observed and the solution may be distinct in colour

hydrophilic
a substance that mixes readily with water

hydrophobic
a substance that does not easily dissolve in water

16.2 Solutes, solvents, solutions and concentration

Water is considered the universal solvent. This is due to its polarity. **Solutes** are substances that dissolve in a **solvent**. The **solution** that is formed from a solute dissolving in a solvent is homogenous, as the particles are too small to be seen.

Substances that readily dissolve in water are called **hydrophilic**. These include ionic compounds and polar substances, such as some polar organic compounds. Substances that do not dissolve in water are **hydrophobic**. These tend to be gases or non-polar materials.

Substances may dissolve in water if they are polar or charged, by forming ion-dipole, dipole-dipole, or hydrogen bonds with water molecules. However, a substance will only dissolve if the energy of the bonds that it forms with the water is lower than the energy of the bonds between the water molecules or the bonds within the solid undissolved substance. Consequently, not all polar molecules and not all ionic substances are soluble in water. Chapter 18 discusses the process of forming solutions and how **solubility** varies with temperature further.

solubility
the amount of solute that can dissolve in 100 g water (solvent) at a particular temperature

SECTION REVIEW 16.2

REMEMBERING

1. List four properties of water that make it an unusual substance, despite it being very common on Earth.
2. Explain why water is so important to the existence of life.
3. Define 'water of crystallisation'.

UNDERSTANDING

4. Explain why ionic substances such as salt dissolve in water.

APPLYING

5. Despite being an ionic substance, barium sulfate does not actually dissolve in water. Explain why this might be.
6. A sample of copper(II) sulfate was tested in order to determine whether or not it was anhydrous. This was done by repeatedly heating, cooling and weighing the sample. Explain how the results of this experiment would be used to determine an outcome.

ANALYSING

7. The tallest trees in the world are redwoods, found on the western coast of the United States. Water is transported up the trunk and stems in tubes called xylem. A particularly notable feature of the redwood trees is that the further up the tree they get, the thinner the xylem become. Explain how this phenomenon might enable water to be transported up to the topmost leaves

16.3 Units of concentration

concentration
the number of particles in a given volume

Aqueous solutions can contain many different dissolved substances. The amount of dissolved solute is regularly monitored in many areas of chemistry, including wine production, water quality and swimming pools. In many of these areas, it is important to know the **concentration** of the solute in the solution. There are a variety of units used to measure concentration, all of which can be summarised as:

$$\text{Concentration} = \frac{\text{amount of solute}}{\text{amount of solvent}}$$

The specific unit used depends on the context in which the measuring is carried out. Table 16.3.1 lists some common units of concentration with some examples of their context.

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TABLE 16.3.1 Some commonly used units of concentration

UNIT OF ONCENTRATION	DEFINITION	CONTEXT
g/100g	Grams of solute per 100 g of solvent	The maximum solubility of sodium chloride at 20°C is 25 g NaCl per 100 g water
%v/v	Percentage by volume	The alcohol content of a bottle of wine might be 12.3%v/v
%w/w	Percentage by mass	The concentration of sugar in a solution is 25% w/w
%w/v	Percentage of mass by volume of solution	The concentration of potassium chloride is 5%w/v
ppm	Grams of solute in one million grams of solution (1 ppm is equivalent to 1 mg L^{-1})	The amount of fluoride in a city's water supply was 80 ppm
ppb	Grams of solute in one billion grams of solution (1 ppb is equivalent to 1 µg L^{-1})	The amount of lead in drinking water is might be 15 ppb
molarity	The number of moles of solute per litre of solution	The concentration of a solution of hydrochloric acid was 0.1 M

Remember that pure water has a density of 1 g mL^{-1}, so the mass of water in grams is always numerically equal to the volume in mL.

WORKED EXAMPLE 16.3.1

1 Red wine has an alcohol content of approximately 13 v/v %. In a 750 mL bottle, what amount of alcohol is present?

2 A hair sample was found to have a copper level of 41 ppm. It should be less than 2.5 mg % w/w. Is the amount in the hair an acceptable level?

ANSWERS

1 Calculate the alcohol content:

= total volume × percentage alcohol
= 750 mL × 13 mL/100 mL = 97.5 mL

The 750 mL of red wine contains 97.5 mL of ethanol.

2 No, level is too high.

- Convert the copper level in hair from parts per million to an amount in grams.

$$41 \text{ ppm} = 41 \text{ g per } 10^6 \text{ g} = \frac{41 \times 100}{10^6} = \frac{41}{10^4}$$

- The legal limit has the % w/w as a value in mg so convert from grams to milligrams.

$$41 \times 10^{-6} \times 100 \text{ \% w/w} = 41 \times 10^{-6} \times 100 \times 1000 = 4.1 \text{ mg \% w/w}$$

SECTION REVIEW 16.3

APPLYING

Calculate the following.

1 200 mL of a sport drink has 6% w/v of sugar. What amount of sugar is present?

2 A 300 mL glass of champagne contains 36 mL of alcohol. What is the concentration (v/v) of alcohol?

3 A solution has a cadmium ion level of 500 ppm. How much cadmium would be present in 100 mL of this solution?

unsaturated solution
a solution with less than the maximum amount of solute dissolved in a given quantity of solvent

saturated solution
a solution in which no more of a particular solute can dissolve in a given quantity of solvent

supersaturated solution
an unstable solution that has more than the maximum amount of solute dissolved in a given quantity of solvent; it has a higher amount of solute than the saturated solution

16.4 Saturated solutions

When adding sugar to a cup of water, there reaches a point at which no more sugar can dissolve at a particular temperature. The extra sugar stays at the bottom of the cup. The solution is said to be saturated with sugar. Before this point was reached, it was an unsaturated solution.

If the water is heated, then more sugar will dissolve. If this heated solution is allowed to cool, crystals will start to precipitate again. The temperature at which this happens is the temperature at which the solution is saturated. A solution can become supersaturated. This means that there is more solute dissolved than in a saturated solution at the same temperature. If this supersaturated solution is bumped, a sugar crystal is added or the side of the glass is scratched, then the extra sugar will precipitate out again. This is because the solution is unstable. Toffee is a supersaturated solution of sugar. To stop the toffee crystallising, a mixture of sugars or the addition of some fats can be used to stabilise the solution.

SECTION REVIEW 16.4

REMEMBERING

1 Fill the gaps: When a solvent holds as much of a solute as it normally can at a given temperature, the solution is ______________. When a solvent holds less dissolved solute than it normally can at a given temperature, the solution is ______________.

UNDERSTANDING

2 Samples of brackish waters were found to have $30\,g\,L^{-1}$ of salt. How much salt can be obtained from the following?

- **a** 50 mL
- **b** 300 mL

3 Bottles containing 40 mL of brand X, a liquid plant food, contains 3.2 g of seaweed extract. What is the percentage concentrate of seaweed extract in the plant food?

4 The fluoride concentration of drinking water should be less than 1 ppm on a mass to volume basis. A 500 mL sample of water was found to contain 0.3 mg of fluoride ions.

- **a** Calculate the concentration of fluoride ions in the water in % w/v and parts per million.
- **b** Is this below the recommended level?

5 How many grams of potassium nitrate would be needed to saturate 50 g of water at 20°C?

6 Sodium thiosulfate will dissolve in water readily. At 25°C, a saturated solution will have 120 g of $Na_2S_2O_3 \cdot 5H_2O$ for each 100 g of water.

- **a** You take a test tube and add a crystal of sodium thiosulfate to $2\,cm^3$ of water. Is it likely to be saturated?
- **b** You take another test tube and add a similar amount of water and then add crystals until no more dissolves. What term describes this?
- **c** In a third test tube, you add $2\,cm^3$ of crystals and a few drops of water. Why would you not expect all the crystals to dissolve?
- **d** You heat it until the solution is clear and then cool it back to room temperature. What would happen if a crystal were now added to the test tube?

7 A 20 g solution of a pesticide contains 0.5 g of active ingredient. What is the concentration in % w/w and ppm.

APPLYING

8 On the bottle of every sport drink is a label detailing the amount of salts and sugars present, usually in 1 litre of product. Compare sport drinks compositions. If you had to make 750 mL of a sport drink, what and how much would you add to 750 mL of water?

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9 A 200 mL sample of tap water was found to contain 0.2 mg of fluoride ions.

a Calculate the concentration of fluoride ions in the water as ppm.

b As the limit should be between 0.7 and 1.2 ppm, does this concentration exceed the recommended limit of 1 ppm?

c The recommended maximum daily intake of fluoride is about 10 mg/day. How many cups would a person need to drink to reach this limit? Assume that water is the only source of fluoride.

16.5 Molarity and reacting volumes of solutions

When calculating quantities in chemical reactions, the basic unit is the mole. It is convenient to have a measure of concentration in terms of moles. This measure of concentration is called **molarity**. The molarity of a solution is defined as the number of moles of solute per litre of solution.

molarity
the number of moles of solute per litre of solution

The concentration of a particular solution is given by:

$$\text{Concentration} = \frac{\text{amount of solute}}{\text{amount of solvent}}$$

$$c = \frac{n}{V}$$

This formula can also be rearranged appropriately to find n (given c and V) or V (given n and c):

$$n = cV$$

$$V = \frac{n}{c}$$

Molarity is also referred to as molar concentration and is given the symbol M. In your chemistry experiments, you may use a solution that is labelled as 0.1 M sodium chloride, for example. This means 0.1 moles of sodium chloride per litre of solution.

Sometimes, chemists may need to know the mass of solute in a solution so units of $g \cdot L^{-1}$ may also be used. The following describes the relationship between moles and mass:

$$n = \frac{m}{M}$$

Worked example 16.5.1 shows how moles, concentration and mass are calculated.

SYLLABUS LINK

Chapter 12 explores the relationship between moles and mass.

WORKED EXAMPLE 16.5.1

QUESTIONS

1 Calculate the concentration of copper(II) sulfate in a solution if 250 mL of the solution contains 0.01 mol of copper(II) sulfate.

2 What is this concentration in $g L^{-1}$?

3 Calculate the concentration of a solution prepared by dissolving 13.5 g of sodium hydroxide in 100 mL of solution.

4 What mass of silver nitrate is needed to prepare 250 mL of a 0.120 $mol L^{-1}$ solution of silver nitrate?

ANSWERS

1 Concentration is 0.04 mol L^{-1}:

a Write the relationship:

$$c=\frac{n}{V}$$

b Identify the quantities and check units:

$$c = ?, n = 0.01\,\text{mol}, V = 250\text{ mL} = 0.25\,\text{L}$$

c Substitute to find c:

$$c=\frac{0.01}{0.25}=0.040\text{ mol L}^{-1}$$

2 Concentration is 6.4 g L^{-1}.
Convert moles to grams using $n=\frac{m}{M}$

$$n=0.040\text{ mol}, m=?, M(\text{CuSO}_4)=63.55+32.06+4\times16.0=159.6\text{ g mol}^{-1}=0.25\,\text{L}$$

$$0.040=\frac{m}{159.6}$$

$$m=0.040\times159.6=6.4\,\text{g}$$

3 Concentration is 3.38 mol L^{-1}:

a Write the relationships:

$$c=\frac{n}{V}\quad n=\frac{m}{M}$$

b Identify the quantities and check units:

$$c=?\text{ mol L}^{-1}, n=?, V=100\text{ mL}=0.100\,\text{L}, m=13.5\text{ g}$$

c Calculate M and substitute to find n:

$$M(\text{NaOH})=40\text{ mol}^{-1}$$

$$n=\frac{13.5}{0.100}=0.3375\text{ mol}$$

d Substitute to find c:

$$c=\frac{0.3375}{0.100}\text{ mol L}^{-1}$$

$$c=3.38\text{ mol L}^{-1}$$

4 Mass is 5.09 g:

a Write the relationships:

$$c=\frac{0.3375}{0.100}\text{ mol L}^{-1}$$

b Identify the quantities and check units:

$$c=0.120\text{ mol L}^{-1}, n=?, V=250\text{ mL}=0.25\,\text{L}, m=?$$

c Substitute to find n:

$$0.120=\frac{n}{0.25}\text{ mol L}^{-1}$$

$$n=0.120\times0.25=0.03\text{ mol}$$

d Calculate M and substitute to find m:

$$M(\text{AgNO}_3)=169.9\text{ g mol}^{-1}$$

$$0.03=\frac{m}{169.0}$$

$$m=0.03\times169.9=5.09\,\text{g}$$

Calculating reacting quantities

Now that a relationship between moles and concentration has been developed, calculations can be extended to include chemical reactions that involve solutions.

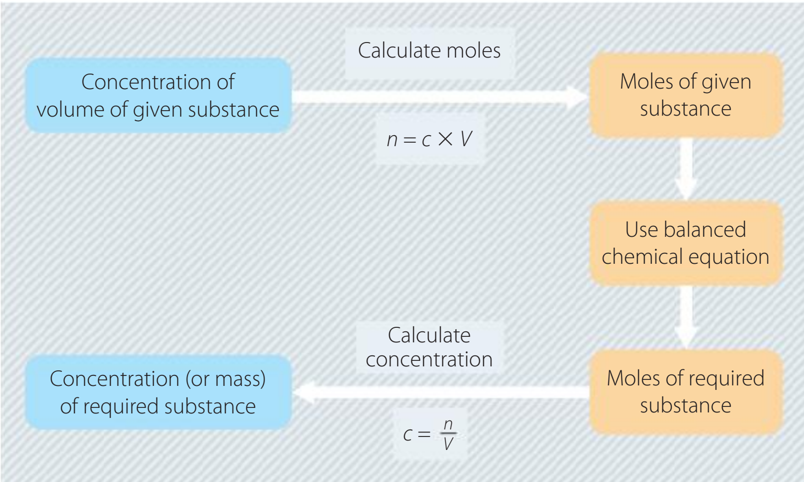

FIGURE 16.5.1 Calculating reaction quantities

WORKED EXAMPLE 16.5.2

QUESTION

What mass of barium sulfate will form when excess sodium sulfate is added to 50 mL of a 0.30 mol L^{-1} barium chloride solution?

ANSWER

Mass is 3.5 g.

1 Write a balanced equation:

$$Na_2SO_4(aq) + BaCl_2(aq)\ BaSO_4(s) + 2NaCl(aq)$$

2 Calculate the number of moles of $BaCl_2$:

$$c = \frac{n}{V}$$

$$\therefore n = cV$$

$$c = 0.30 \text{ mol } L^{-1}, V = 50 \text{ mL} = 0.05 \text{ L}$$

$$\therefore n = 0.30 \times 0.05 = 0.015 \text{ moles}$$

3 Use the balanced equation to calculate the number of moles of $BaSO_4$:

$$n(BaSO_4) = n(BaCl_2) = 0.015 \text{ mol}$$

4 Calculate the mass of $BaSO_4$ $\left(M(BaSO_4) = 233\right)$:

$$n = \frac{m}{M}$$

$$\therefore m = nM$$

$$\therefore m = 0.015 \times 233 = 3.5 \text{ g}$$

Dilutions

To reduce transport and packaging costs, many products such as cordials, detergents, laboratory chemicals and pesticides are sold in a concentrated form. Water is added when they are ready for use. The amount of water added to cordials depends on personal preference, but for pesticides there are instructions on the packaging for making solutions of specific concentrations. This ensures the final mixture is suitable for its intended purpose. The process of adding water to a solution to make it less concentrated is called **dilution**.

dilution
adding water to a solution to reduce the concentration

Dilution is an important laboratory technique used to produce solutions of required concentration for testing and analysis procedures. The dilution process must be as accurate as possible, so the concentration of the required diluted solution is correct.

The usual equipment to ensure accurate measures of volume are pipettes and volumetric flasks similar to those shown in Figure 16.5.2.

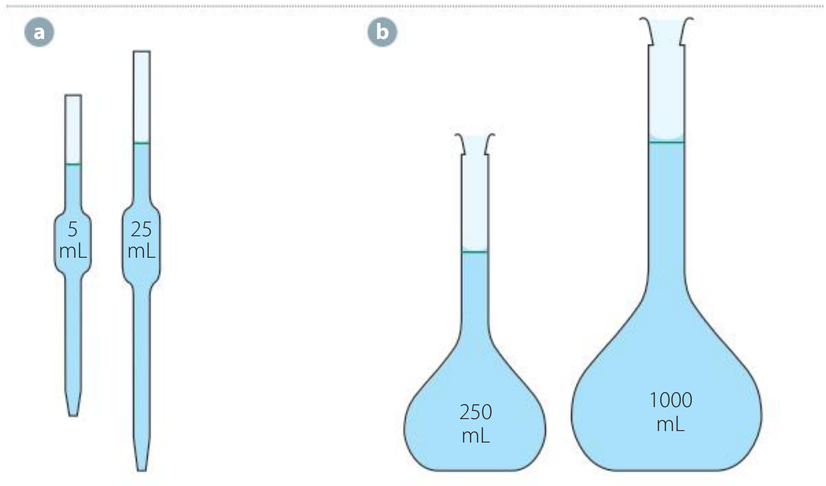

FIGURE 16.5.2 **a** Pipettes and **b** volumetric flasks are used in the laboratory when accurate volumes are required.

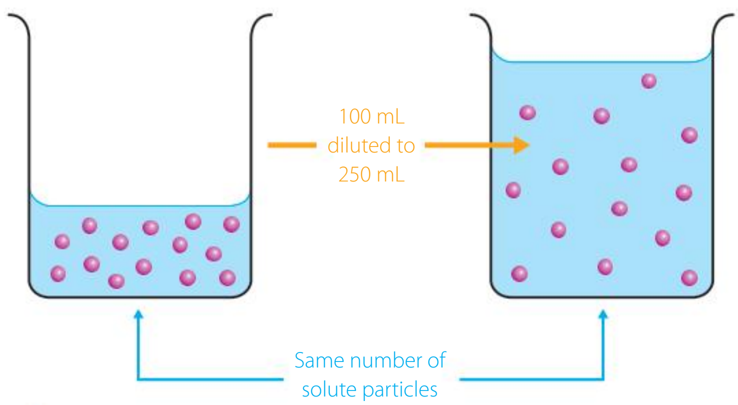

FIGURE 16.5.3 Dilution: the same number of moles in a larger volume of solvent

The dilution procedure involves:

- using a pipette to measure a definite volume of a solution of known concentration
- transferring this volume to a volumetric flask
- adding sufficient water to the volumetric flask to make up the solution to the calibration mark.

When calculating the concentration of the diluted solution, scientists make use of the fact that the number of moles (particles) in the diluted solution is the same as in the original volume of solution (Figure 16.5.3). The number of moles of solute remains constant, so scientists can apply the following formula to relate concentration and volume of the original solution to that of the diluted solution. This relationship is called the dilution formula:

$$c_1V_1 = c_2V_2$$

where c_1 and c_2 are the initial and final concentrations of the solution in the same units, and V_1 and V_2 are the initial and final volumes of the solution, in the same units.

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WORKED EXAMPLE 16.5.3

QUESTION

1 What volume of 12 mol L^{-1} hydrochloric acid is needed to prepare 500 mL of a 0.25 mol L^{-1} solution?

2 What is the concentration of a solution made by diluting 20 mL of a 5.0 mol L^{-1} solution to 1.00 L?

ANSWERS

1 The volume is 10.4 mL.

a Write the relationship:

$$c_1V_1 = c_2V_2$$

b Identify the quantities and check units for consistency:

$$c_1 = 12 \text{ mol L}^{-1}, V_1 = ?, c_2 = 0.25 \text{ mol L}^{-1}, V_2 = 500 \text{ mL}$$

c Substitute:

$$12 = V_1 \times 0.25 = 500$$

$$V_1 = \frac{0.25 \times 500}{12} = 10.4 \text{ mL}$$

2 The concentration is 0.10 mol L^{-1}.

a Write the relationship:

$$c_1V_1 = c_2V_2$$

b Identify the quantities and check units for consistency:

$$c_1 = 5.0 \text{ mol L}^{-1}, V_1 = 20 \text{ mL}, c_2 = ?, V_2 = 1.00 \text{ L} = 1000 \text{ mL}$$

c Substitute:

$$5.0 \times 20 = ? \times 100$$

$$c_2 = \frac{5.0 \times 20}{1000} = 0.10 \text{ mol L}^{-1}$$

PRACTICAL ACTIVITY 16.5.1

Preparing and diluting solutions

AIM

To prepare solutions of known concentration and use these to determine the unknown concentration of a solution

MATERIALS

- approx. 0.1 g potassium permanganate ($KMnO_4$)
- solution of unknown concentration
- 250 mL beaker
- 25 mL pipette
- 4 × 250 mL volumetric flasks
- stirring rod
- filter funnel
- electronic balance
- water
- paper towel

» RISK ASSESSMENT

WHAT ARE THE RISKS IN DOING THIS EXPERIMENT?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?
Potassium permanganate can stain skin and clothing.	Wear protective clothing and avoid contact.
Potassium permanganate can be irritating to the skin and eyes.	Wear safety glasses and avoid contact with the chemical.

In your write-up, add any more risks you can think of, as well as ways to manage them. In particular, expand on the two risks listed to identify specific risks involved with each of them.

PROCEDURE

1 Weigh out approximately 0.1 g of potassium permanganate in one of the beakers. (Make sure you measure the mass you are using accurately.)
2 Add just enough water to dissolve the crystals of potassium permanganate. Transfer the solution to a 250 mL volumetric flask.
3 Fill the flask to the 250 mL mark with water and mix thoroughly.
4 Draw up a table like the one in the results section and complete it for this first solution.
5 Pipette 25 mL of the first solution into a clean volumetric flask and fill to 250 mL.
6 Complete the table for this second solution.
7 Repeat steps **5** and **6** using the second and third solutions so you have successive dilutions.
8 Compare the colour of your solution of unknown concentration to the colours of your diluted solutions.

RESULTS

1 Complete the results table.

SOLUTION	RELATIVE COLOUR INTENSITY	INITIAL VOLUME	FINAL VOLUME	MASS $KMNO_4$	MOLES $KMNO_4$	CONCENTRATION ($MOL L^{-1}$)
1						
2						
3						
4						

2 What effect does dilution have on the following?
  a Colour intensity
  b Concentration of the solution
  c Mass of the solute in the diluted solution compared to the original mass

DISCUSSION

1 a What is the concentration of the unknown solution?
  b How easy was it to determine this?
  c What could be done to more accurately determine the unknown concentration?
2 How would you prepare a solution twice as concentrated as solution 2?

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SECTION REVIEW 16.5

REMEMBERING

1 Outline the relationship between molarity and concentration.

UNDERSTANDING

2 Calculate the concentration of the following solutions in the following.

i $mol\,L^{-1}$ **ii** $g\,L^{-1}$

a 2.00 L of solution containing 1.50 mol of potassium bromide

b 250 mL of solution containing 0.0025 mol of sodium iodide

3 Calculate the number of moles of solute that is needed to make the volume of the solution in the following.

a 100 mL of a 0.40 M solution of barium nitrate

b 500 mL of a 0.15 M solution of sodium sulfate

c 2.00 L of a 0.032 M solution of copper(II) chloride

APPLYING

4 What mass of solute is required for each of Questions 3a–c?

5 Calculate the molar concentration of the following solutions.

a 199 g $NiBr_2$ dissolved in 5.00 L of solution

b 0.059 g KF dissolved in 227 mL of solution

c 23 g ethanol (C_2H_5OH) in 750 mL of solution

6 What mass of solute is needed to make the following solutions?

a 25.0 mL of a 0.02 M solution of sodium carbonate

b 32.0 mL of a 0.10 M solution of copper(II) sulfate

7 Calculate the volume of a 0.025 M solution that contains 1.0×10^{-3} mol of magnesium nitrate.

8 Calculate the molarity of:

a 0.20 mol zinc iodide dissolved in 1.5 L solution

b 0.04 mol magnesium chloride dissolved 450 mL solution.

9 Calculate the concentration of the following solutions made by dissolving a known mass in a given volume.

a 8 g potassium chloride dissolved in 250 mL solution

b 1.46 g iron(II) sulfate dissolved in 100 mL solution

c 7.5 g silver nitrate dissolved in 500 mL solution

10 What mass of copper(II) hydroxide forms when excess copper(II) sulfate is added to 100 mL of a 0.450 $mol\,L^{-1}$ sodium hydroxide solution?

11 If excess NaOH is added to 10 mL of a 0.500 $mol\,L^{-1}$ solution of zinc chloride, calculate the mass of $Zn(OH)_2$ produced.

ANALYSING

12 What mass of silver iodide will be formed when excess sodium iodide is added to 25 mL of a 0.15 M solution of silver nitrate?

13 To measure the concentration of calcium ions in a water sample, an analyst took 500 mL of the sample and added sodium carbonate until no further precipitation occurred. The precipitate was filtered off, dried and weighed. It was found to have a mass of 1.72 g.

a What was the molarity of calcium ions in the original solution?

b How many grams of calcium ions were there per litre in the original solution?

c Water is considered to be 'hard' if it contains more than 201–300 $mg\,L^{-1}$ calcium carbonate. Would this water sample be classified as 'hard'?

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 Define the following terms.

a Surface tension
b Capillary action
c Solubility
d Solute
e Hydrophobic
f Hydrophilic
g Concentration
h Molarity

CATEGORY QUESTIONS

2 Explain the presence of surface tension as a property of a liquid, in terms of the interactions between the molecules.

3 In 16.4 mL of 0.117 M magnesium chloride solution, how many moles are there of each of the following.

a Magnesium chloride
b Magnesium ions
c Chloride ions

4 How many moles of solute do you need to make the following solutions?

a 2.00 L of 1.50 M sodium chloride
b 3.5 L of 0.20 M potassium hydroxide

5 Calculate the molarity of the solutions made by dissolving the following amounts of solute in water and making the volume up to the stated value.

a 5.0 mol nitric acid in 2.0 L
b 2.5 mol sodium hydroxide in 0.50 L

6 The volume of solution in column A was diluted to the volume in column B. Calculate the molarity of the diluted solution for each of the three examples below.

	COLUMN A	COLUMN B
a	50 mL of 0.242 M hydrochloric acid	500 mL
b	25 mL of 0.152 M sulfuric acid	2.0 L
c	10 mL of 0.114 M sodium hydroxide	250 mL

7 For each of the three examples below, what volume of solution in Column A is needed to prepare the solution in Column B?

	COLUMN A	COLUMN B
a	0.282 M hydrochloric acid	250 mL of 0.0113 M
b	2.42 M sulfuric acid	2.0 L of 0.121 M
c	0.318 M sodium hydroxide	1.0 L of 0.300 M

ELABORATION QUESTIONS

8 A commercial liquid medication is known to contain 400 mg of magnesium hydroxide per 10.0 mL of the medication. Express this as % w/v and ppm. (Assume that 1 mL of medication has a mass of 1 gram.)

9 Formamide is a liquid with a very high surface tension, which wets glass. If the end of a thin glass capillary is dipped into a beaker of formamide, what is observed?

10 Calculate the total volume of alcohol in a 375 mL of wine with an alcohol content of 8.5% v/v.

11 Serotonin ($C_{10}H_{12}N_2O$) is a compound that conducts nerve impulses in the brain and muscles. A sample of spinal fluid was found to contain a serotonin concentration of 1.5 mg L^{-1}. How much serotonin is there in 1 mL of spinal fluid?

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END-OF-CHAPTER EXAM

End-of-chapter test

1 You prepare an aqueous solution of ammonium sulfate, $(NH_4)_2SO_4$, by dissolving 15.864 g of solid solute in 500.0 mL deionised water. In this solution, the concentration of the ammonium ion would be closest to:

A 0.24 M.

B 0.12 M.

C 0.48 M.

D 0.06 M.

2 What volume of water (mL) must be added to 7.0 mL of 10 M nitric acid to produce a solution of 3.5 M HNO_3?

A 7

B 13

C 20

D 27

3 Which property of water is most relevant to its use for cooling when humans perspire and dogs pant?

A Relatively high boiling point

B Polar nature enabling a wide range of solutes to dissolve

C High latent heat of fusion

D High latent heat of vaporisation

4 Which of the following is not a property of water?

A As a solid, it is less dense than in its liquid form.

B It has a relatively low melting temperature compared to similarly sized covalent molecules.

C It has a high heat of vaporisation.

D It can dissolve a large range of substances.

5 A 7.0 L sample of 2.0 M nitric acid has an extra 16 L of water added to it. What is the new concentration of HNO_3?

A 0.61 M

B 0.875 M

C 4 M

D 14 M

6 A solution contains 2.0 g of KCl dissolved in 250 mL. Give the concentration of the solution in the following units.

a $mg\,L^{-1}$

b % (w/v)

c M

d ppm

7 Referring to water's physical properties, explain the following observations when making a cup of coffee.

a If sugar is added to hot coffee it readily dissolves . If flour is mistakenly added, it does not dissolve.

b A metallic spoon is used to stir the hot coffee. While stirring, the coffee and the spoon are at the same temperature, but when the spoon is removed from the coffee, its temperature drops far below the coffee's.

c The drinker prefers iced coffee, and adds ice-cubes to help cool the coffee. Explain why the ice-cubes float before melting.

d The drinker uses a straw to drink the iced coffee, but notices that the coffee's height in the straw is greater than that of the coffee in the cup.

8 What mass of solute is needed to make the following solutions?

a 2.00 L of a 0.200 mol L^{-1} solution of sodium iodide

b 250 mL of a 0.01 mol L^{-1} solution of potassium phosphate

c 500 mL of a 0.0075 mol L^{-1} solution of barium nitrate

9 Calculate the mass of magnesium carbonate formed when excess sodium carbonate is added to 75 mL of a 0.08 mol L^{-1} magnesium chloride solution.

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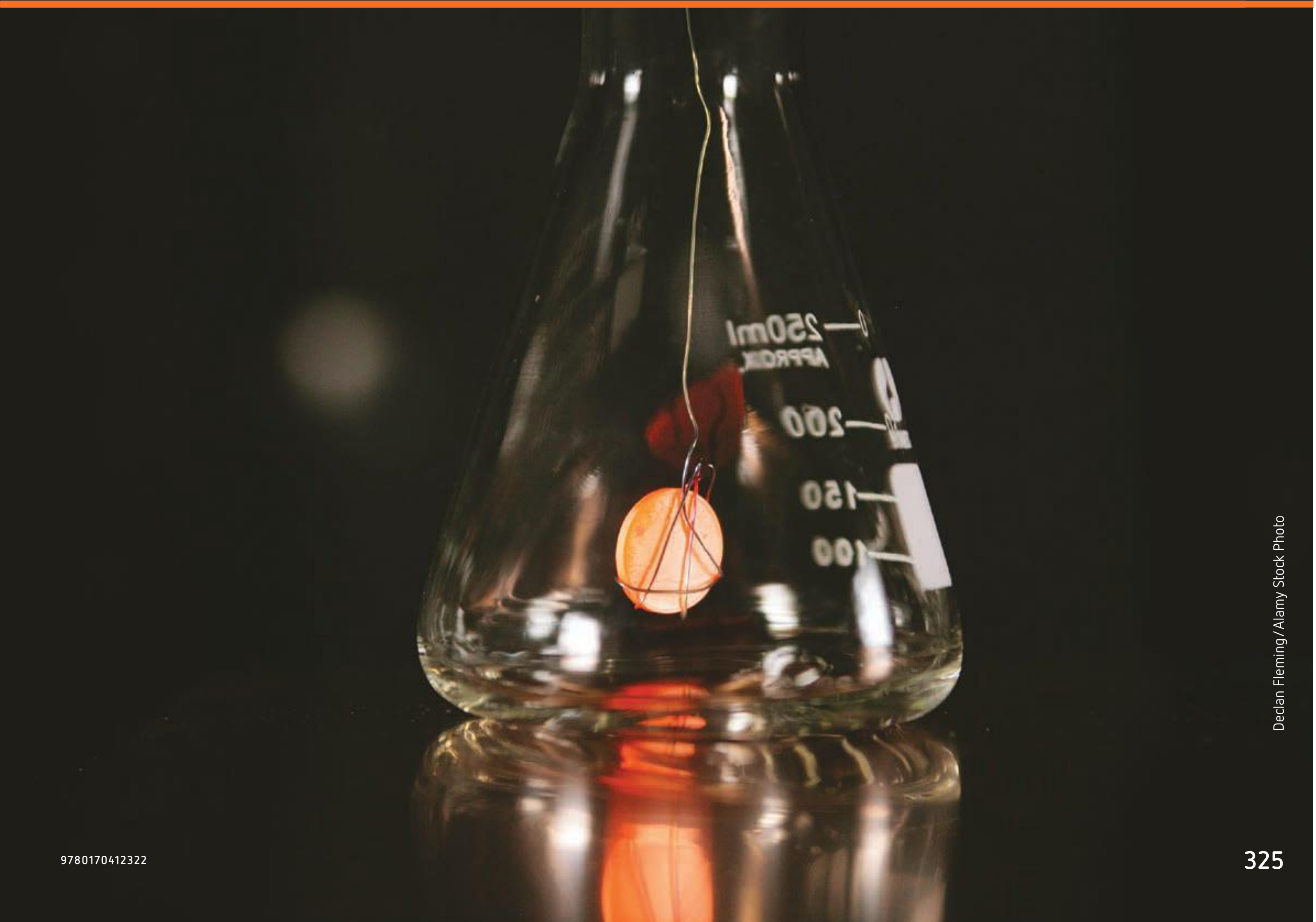
17 IDENTIFYING IONS IN SOLUTION

Introduction

Chapter 17 examines precipitation reactions and shows how to use solubility rules to predict the products of these reactions. Precipitation reactions can be applied practically in a qualitative analysis to determine the identity of unknown cations and anions in solution.

Stimulus questions

Solutions of sodium carbonate and sodium sulfate look the same. How do scientists tell them apart?



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17.1 Solubility rules

precipitate
the solid that forms when two different liquids react as a result of the insolubility of at least one product

precipitation reaction
a reaction between two solutions that results in the formation of a solid

suspension
a mixture of a liquid and an insoluble solid, where the solid distributes evenly throughout the liquid

17.1.1 Solubility

FIGURE 17.1.1 A silver chloride precipitate

When solutions of certain ionic compounds are mixed, they sometimes react, producing a solid known as a **precipitate**. This type of reaction is a **precipitation reaction**. Often, when a precipitate forms, it does not immediately sink to the bottom of the container; rather, it distributes evenly throughout the liquid, forming a cloudy-looking **suspension**.

For example, when sodium chloride solution is added to silver nitrate solution, a white solid of silver chloride forms. Sodium chloride solution is composed of sodium ions (Na^+) and chloride ions (Cl^-) dissolved in water. Silver nitrate solution is composed of silver ions (Ag^+) and nitrate ions (NO_3^-) dissolved in water. When the solution of Na^+ and Cl^- ions is mixed with the solution of Ag^+ and NO_3^- ions, the result is a solution containing all the ions moving randomly throughout the solvent. The ions in the original solutions were surrounded (solvated) by water molecules. When the solutions are mixed, the Ag^+ and Cl^- ions are more strongly attracted to each other than to the water, so they cluster together in groups and form a solid, which precipitates out of solution; that is, a precipitate of solid AgCl forms, as shown in Figures 17.1.1 and 17.1.2.

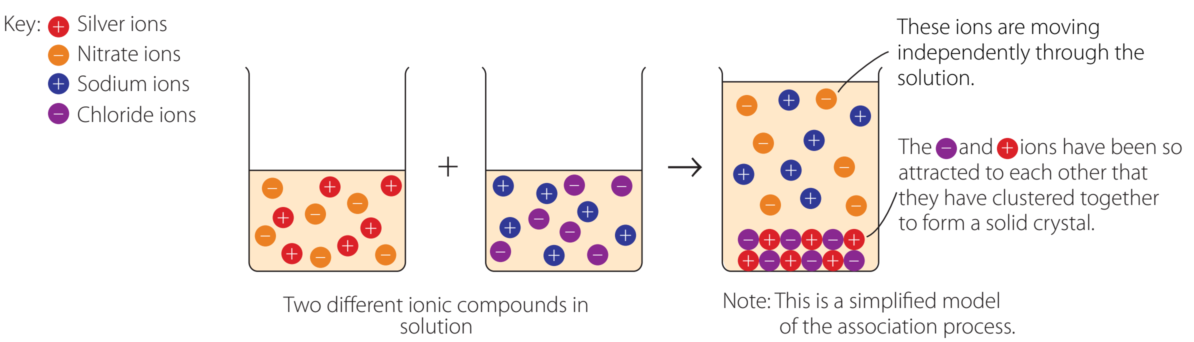


FIGURE 17.1.2 Explaining the formation of a silver chloride precipitate

The sodium and nitrate ions continue to move randomly through the solution because their attraction to water molecules is stronger than their attraction to each other. They are not involved in the reaction and they remain in the solution as ions. Ions that are not involved in a reaction are called **spectator ions**.

spectator ions
ions that remain in a solution before and after a chemical reaction has taken place, which are not involved in the reaction

By systematically mixing solutions that contain known cations and anions, chemists have experimentally identified which compounds are soluble and which are insoluble, forming a precipitate. These results have been summarised into a table of solubility data (Table 17.1.1, page 329), which you can use to predict whether precipitation will occur when two known solutions are mixed.

PRACTICAL ACTIVITY 17.1.1

Investigating precipitation reactions

Solubility rules can be formulated from experimental results. In principle solubility is predictable but the calculations are complex and require highly accurate estimates of intermolecular interactions, a subject that remains an active area of chemical research. In this experiment, you will determine which mixtures of solutions form precipitates and use these results to construct solubility rules. These can then be compared with existing experimental data from the chemical literature.

AIM

To compare the solubility of a range of ionic substances through precipitation reactions

EQUIPMENT PER GROUP

- cations (all 0.1 mol L^{-1}) in dropper bottles: sodium nitrate ($NaNO_3$), silver nitrate ($AgNO_3$), lead(II) nitrate ($Pb(NO_3)_2$), ammonium nitrate (NH_4NO_3), magnesium sulfate ($MgSO_4$), iron(II) sulfate ($FeSO_4$), copper(II) sulfate ($CuSO_4$), zinc sulfate ($ZnSO_4$), aluminium sulfate ($Al_2(SO_4)_3$), barium chloride ($BaCl_2$), calcium chloride ($CaCl_2$), iron(III) chloride ($FeCl_3$)
- anions (all 0.1 mol L^{-1}) in dropper bottles: potassium nitrate (KNO_3), potassium chloride (KCl), sodium sulfate (Na_2SO_4), sodium carbonate (Na_2CO_3), sodium acetate ($NaCH_3COO$), sodium hydroxide (NaOH)
- distilled water
- 6 test tubes and test-tube rack
- paper towel
- test-tube cleaner

PROCEDURE

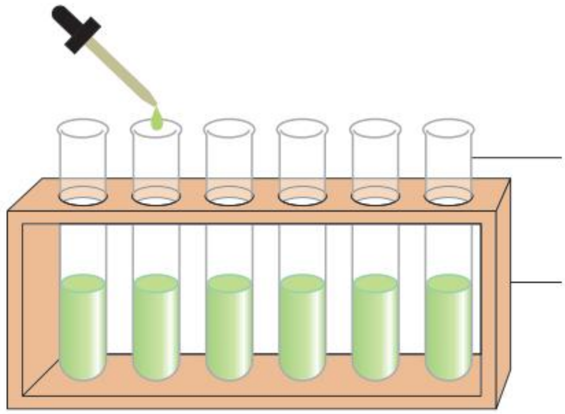
1 Copy and complete the risk assessment table. Add any more risks you can think of, as well how to manage them. Expand on the two risks listed to identify specific risks involved with each of them.

WHAT ARE THE RISKS IN DOING THIS EXPERIMENT?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?
Lead and barium salts are poisonous.	Wear safety glasses and avoid contact with skin. If contact occurs, wash thoroughly with soap and water.
Lead, barium, silver and copper salts are harmful to the environment.	Do not pour solutions containing these salts down the drain. Dispose of as directed by your teacher.

RISK ASSESSMENT

2 Put a few drops of each anion into each test tube, as shown in Figure 17.1.3. Take care not to mix droppers.

FIGURE 17.1.3 Experimental set-up



3 Add a few drops of the first cation to each. Record your observations.

4 Clean all test tubes thoroughly because incorrect results can arise from contamination of solutions.

5 Repeat step **3** for the next cation. Continue until all cations have been tested.

RESULTS

Draw up a table similar to the one below to record your observations. For no reaction, record 'NR'; for a precipitate, record 'P' and the colour of the precipitate.

CATION	ANION					
	NO_3^- NITRATE	CL^- CHLORIDE	SO_4^{2-} SULFATE	CO_3^{2-} CARBONATE	CH_3COO^- ETHANOATE	OH^- HYDROXIDE
Na^+						
Ag^+						
Pb^{2+1}						
NH_4^+						
Mg^{2+}						
Fe^{2+}						
Cu^{2+}						
Zn^{2+}						
Al^{3+}						
Ba^{2+}						
Ca^{2+}						
Fe^{3+}						

ANALYSING THE RESULTS

Summarise your results using statements such as 'All nitrates are …'

DISCUSSION

1 a Compare the results of the experiment with the solubility rules in Table 17.1.1 (page 329).

b What, if any, were the discrepancies between your results and Table 17.1.1? Suggest a reason for these discrepancies and a way of resolving them.

2 Write a balanced net ionic equation for each of the reactions in which a precipitate was obtained.

CONCLUSION

Summarise the solubility rules according to your results.

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Table 17.1.1 summarises the solubility of common ionic compounds. The table applies to some common cations and is organised by the solubility of various anions. An additional useful generalisation to note is that all salts containing Group 1, ammonium or nitrate ions are soluble.

TABLE 17.1.1 Solubility rules for some common ions

SOLUBLE ANIONS	EXCEPTIONS
NO_3^-	None
CH_3COO^-	Ag^+ slightly soluble
Cl^-	Ag^+ insoluble; Pb^{2+} slightly soluble
Br^-	Ag^+ insoluble; Pb^{2+} slightly soluble
I^-	Ag^+, Pb^{2+} insoluble
SO_4^{2-}	Ba^{2+}, Pb^{2+}, Sr^{2+} insoluble; Ag^+, Ca^{2+} slightly soluble
INSOLUBLE ANIONS	**EXCEPTIONS**
OH^-	Group 1, NH_4^+, Ba^{2+}, Sr^{2+} soluble; Ca^{2+} slightly soluble
O^{2-}	Group 1, NH_4^+, Ba^{2+}, Sr^{2+}, Ca^{2+} soluble
S^{2-}	Groups 1 and 2, NH_4^+ soluble
CO_3^{2-}	Group 1, NH_4^+ soluble
SO_3^{2-}	Group 1, NH_4^+ soluble
PO_4^{3-}	Group 1, NH_4^+ soluble

WORKED EXAMPLE 17.1.1

QUESTIONS

1 What precipitate, if any, will form when you mix aqueous solutions of magnesium iodide and silver nitrate? If a precipitate forms, write the neutral species and net ionic equations for the reaction.

2 Select a reagent that could be used to precipitate the cation in $BaBr_2$. Explain your selection.

ANSWERS

1 **a** Write the formula of the compounds present:

Magnesium iodide (MgI_2)
Silver nitrate ($AgNO_3$)

b Predict the products by swapping the partners of the reacting compounds and check that the formulas are correct.

$$MgI_2 + AgNO_3 \rightarrow AgI + Mg(NO_3)_2$$

c Use Table 17.1.1 to confirm if either product is insoluble:
All nitrates are soluble and all iodides are soluble except Ag^+ and Pb^{2+}, so AgI will precipitate.

d Write the required equations, checking they are balanced.
Neutral species equation is:

$$MgI_2(aq) + 2AgNO_3(aq) \rightarrow 2AgI(s) + Mg(NO_3)_2(aq)$$

Net ionic equation is:

$$Ag^+(aq) + I^-(aq) \rightarrow AgI(s)$$

2 a Identify the ions present and one to be precipitated:
Ba^{2+} and Br^- are the ions present in solution. Ba^{2+} is the ion to be precipitated.

b Use Table 17.1.1 to identify an anion that could be used to precipitate the identified cation:
Anions that will produce an insoluble compound with Ba^{2+}: S^{2-}, CO_3^{2-}, SO_3^{2-} and PO_4^{3-}.

c Choose one anion and identify a cation to form a compound. Check that the cation does not react with the anion from the original compound:
Generally, it is best to choose a cation from group 1 or NH_4^+ because compounds of these are soluble. It would be incorrect to choose Ag^+ and Pb^{2+} because they both precipitate with Br^-.

d Identify the reagent and write the formula:
Na_2CO_3 can be used to precipitate Ba^{2+} from a solution of $BaBr_2$.

PRACTICAL ACTIVITY 17.1.2

Identifying ionic compounds

The labels on the bottles of five different chemicals have fallen off. The bottles have been labelled A, B, C, D and E. The chemicals are known to be ammonium hydroxide, potassium iodide, sodium hydroxide, silver nitrate and zinc sulfate. You must identify the contents of each bottle.

AIM

Think about what you want to find out in this investigation.
Hint: Consider all the possible pairings of the chemicals listed and use solubility rules to predict what precipitates will actually form.

EQUIPMENT

Think about the equipment you will need. Will you need any additional chemicals to confirm your results? Be specific about quantities of chemicals and do not forget to include basic materials like distilled water and paper towel.

RISKS

Construct a table similar to the one below. Identify specific risks involved in the investigation and ways that you will manage the risks to avoid injuries or damage to equipment. Ask your teacher to check your risk assessment before you proceed.

RISK ASSESSMENT

WHAT ARE THE RISKS IN DOING THIS INVESTIGATION?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?

PROCEDURE

Hint: You need to use solubility rules. Write your method in a logical sequence of steps. Remember to be specific and include quantities.

»

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RESULTS

Refer back to your aim. Refer to the results of Practical activity 17.1.1 (page 328) for possible table structure and use those results to help identify products.

DISCUSSION

Consider what your results show. Look at theoretical predictions you made about possible precipitates and compare your results with these. Are there any differences or inconclusive results? Do you need to repeat any tests? Do you need to use an alternative solution for confirmation?

CONCLUSION

Reflect on what you found. Could you identify each of the solutions?
How could you improve the effectiveness of the identification process?

SECTION REVIEW 17.1

REMEMBERING

1 Define:

- **a** precipitate
- **b** spectator ion.

UNDERSTANDING

2 Which of the following substances are soluble in water?

- **a** Sodium hydroxide
- **b** Aluminium oxide
- **c** Copper(II) sulfate
- **d** Silver carbonate
- **e** Lithium chloride

3 Name a different pair of substances (using suitable amounts) that you could dissolve in the one sample of water to produce a solution identical with that containing:

- **a** potassium chloride and zinc nitrate
- **b** ammonium sulfate and sodium carbonate
- **c** magnesium bromide and copper(II) sulfate.

APPLYING

4 Solutions of what substances would you mix to prepare the following compounds by precipitation? Write equations for your reactions.

- **a** Magnesium carbonate
- **b** Lead sulfate
- **c** Silver bromide

5 When solutions of potassium sulfide and cobalt acetate are mixed, which salt, if any, will precipitate? If a precipitate does form, write the neutral species and net ionic equations.

6 Name and write the formula of the precipitate (if any) that forms when each of the following pairs of substances is mixed.

- **a** Copper(II) chloride and potassium hydroxide
- **b** Sodium carbonate and ammonium sulfate
- **c** Lead nitrate and copper(II) sulfate

ANALYSING

5 Separate quantities of an unknown solution, Solution A, form precipitates when mixed with potassium carbonate and sodium hydroxide solutions, but did not form precipitates when mixed with solutions of sodium chloride or ammonium sulfate. Solution A could have contained:

A Pb^{2+}, Zn^{2+}, Cu^{2+}

B Mg^{2+}, Al^{3+}, Zn^{2+}

C Mg^{2+}, Ba^{2+}, $A1^{3+}$

D Cu^{2+}, Zn^{2+}, Ag^{+}

6 Describe tests to determine whether a solution contained:

a lead nitrate or barium nitrate.

b copper(II) sulfate or iron(II) sulfate.

17.2.1 Common anions table and formulas list

17.2.2 Testing salts for anions and cations

17.2 Determining the presence of ions in solutions from experimental data

Several common tests can be used to identify particular ions in a solution (Table 17.2.1). The experimental results can be compared with known test results, allowing ions to be identified.

TABLE 17.2.1 Tests used to identify common anions

ANION	TEST
Carbonate (CO_3^{2-})	1 Solution has a pH between 8 and 11 (pH paper suffices). 2 Addition of dilute HNO_3 evolves a colourless gas (CO_2).A
Sulfate (SO_4^{2-})	1 Addition of $Ba(NO_3)_2$ to an acidified sample of the solution produces a thick white precipitate. 2 Acidification and addition of $Pb(NO_3)_2$ produces a white precipitate.
Phosphate (PO_4^{3-})	1 Addition of ammonia followed by $Ba(NO_3)_2$ produces a white precipitate. 2 Addition of Mg^{2+} in an ammonia/ammonium nitrate buffer produces a white precipitate, $Mg(NH_4)PO_4$. 3 Acidification with HNO_3 followed by addition of ammonium molybdate solution $[(NH_4)_2MoO_4]$ produces a yellow precipitate; warming the mixture for a few minutes may be necessary.
Chloride (Cl^-)	1 Addition of $AgNO_3$ to an acidified sample produces a white precipitate,B which dissolves in ammonia solution and darkens in sunlight.

A Any strong acid would do, but for analysing mixtures scientists do not want to introduce any Cl^- or SO_4^{2-}.
B In non-acidic solutions, silver nitrate also produces precipitates with carbonate and phosphate (and with sulfate at all pH values if sulfate concentration is moderately high), so this test alone does not prove the presence of chloride: it is also necessary to prove the absence of sulfate.

TABLE 17.2.2 Precipitates formed between cations and anions

CATIONA	ANION					
	OH^-	CL^-	SO_4^{2-}	CO_3^{2-} IN ALKALINE SOLUTIONB	PO_4^{3-}	
					SOLUTION pH 2	SOLUTION pH 6
Ba^{2+}	No	No	Yes	Yes	No	Yes
Pb^{2+}	Yes	YesC	Yes	Yes	No	Yes
Ag^{+}	YesD	Yes	YesC	Yes	No	Yes
Cu^{2+}	Yes	No	No	Yes	No	Yes

A In acidic or alkaline solution
B Cannot have carbonate in acid solution: it decomposes to CO_2(g)
C Provided concentration of Cl^- or SO_4^{2-} is not too low, say > 0.05 mol L^{-1}
D Precipitate is dark brown or black Ag_2O

In Practical activity 17.2.1, you will use the information from Tables 17.2.1 and 17.2.2 (page 332), along with what you have learnt, to identify some unknown anions. A similar process could be used to identify cations. In addition, you can identify cations with flame tests, by noting the distinctive colour they emit when placed in a flame.

TABLE 17.2.3 Tests used to identify common cations

CATION	TESTS
Ag^+	▪ With Cl^- forms a white precipitate that dissolves in ammonia solution
Pb^{2+}	▪ With Cl^- forms a white precipitate (if solution not too dilute, e.g. > 0.05 mol L^{-1}) and this precipitate does not dissolve in ammonia solution ▪ With I^- forms a yellow precipitate
Ba^{2+}	▪ With SO_4^{2-} forms a white precipitate ▪ Gives a pale green flame colour ▪ No precipitate with OH^- or F^- (compare Ca^{2+})
Ca^{2+}	▪ With SO_4^{2-} forms a white precipitate (if solution not too dilute, e.g. > 0.05 mol L^{-1}) ▪ With F^- forms a white precipitate ▪ Gives a brick-red flame colour
Zn^{2+}	▪ With hydroxide forms a white precipitate ▪ This precipitate dissolves in both excess OH^- solution and ammonia solution
Cu^{2+}	▪ With OH^- forms a blue precipitate ▪ This precipitate dissolves in NH_3 to form a deep blue solution ▪ Gives a blue-green flame colour
Fe^{2+}	▪ With OH^- forms a green or white precipitate, which may turn brown ▪ Decolourises acidified dilute potassium permanganate solution
Fe^{3+}	▪ With OH^- forms a brown precipitate ▪ With thiocyanate (SCN^-) forms a deep red solution

17.3 Writing equations for precipitation reactions

Scientists can determine a great deal about what is happening in a reaction from the states in a chemical equation. When salts dissolve, chemists generally do not write the ions separately. For example, dissolved sodium chloride is written as NaCl(aq). The state aqueous (aq) indicates that the ions have **dissociated**. The Na^+ and Cl^- ions are separated and surrounded by water molecules. They are charged ionic particles, so the solution can conduct electricity; it is an **electrolytic solution**. Two examples for the dissociation of salts are:

$$NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$$

Sodium chloride → sodium ions + chloride ions

$$Cu(NO_3)_2(s) \rightarrow Cu^{2+}(aq) + 2NO_3^-(aq)$$

Copper nitrate → copper ions + nitrate ions

When molecular substances such as carbon dioxide, hydrochloric acid (HCl) or glucose ($C_6H_{12}O_6$) dissolve, the state is aqueous. When acids such as HCl ionise in water, they are written as HCl(aq). This means that the solution contains separate H^+ and Cl^- ions.

$$C_{12}H_{22}O_{11}(s) \rightarrow C_{12}H_{22}O_{11}(aq)$$

Solid sucrose → aqueous sucrose

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

Hydrochloric acids → hydrogen ions + chloride ions

dissociated
when the ions from an ionic compound separate from one another and form a solution in water

electrolytic solution
a solution that contains ions, so it will conduct electricity

After a substance becomes aqueous, further reactions may take place. For example, carbon dioxide has to dissolve in water before it can react with the water. The reaction with water is written separately:

$$\text{Dissolution: } CO_2(g) \rightarrow CO_2(aq)$$

$$\text{Reaction: } CO_2(aq) + H_2O(l) \leftrightarrow H_2CO_3(aq)$$

Carbon dioxide + water ↔ carbonic acid

Complete and net ionic equations

The reaction between sodium chloride and silver nitrate solutions was discussed in section 17.1. When the two solutions are mixed, a precipitate of silver chloride is formed. The equation for this reaction can be written in a number of different ways, which may include or omit the spectator ions.

complete ionic equation
an equation showing all ions present in a chemical reaction – both those reacting and spectator ions

overall equation
a simple equation showing the formulae and ratio of all compounds taking part in a chemical reaction, showing no ions

net ionic equation
shows only the species that are taking part in the chemical reaction; spectator ions are not included

1 $Na^+(aq) + Cl^-(aq) + Ag^+(aq) + NO_3^- \rightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq)$

Sodium ions + chloride ions + silver ions + nitrate ions → silver chloride + sodium ions + nitrate ions

2 $NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

Sodium chloride + silver nitrate → silver chloride + Sodium nitrate

3 $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

Silver ions + chloride ions → silver chloride

Equation 1 is a **complete ionic equation**, showing all the ions present in the solution. Equation 2 is an **overall equation**, showing all the reactants and products as actual neutral compounds. Equation 3 is the **net ionic equation**, showing only the reacting ions. Removing the spectator ions from the complete ionic equation produces the net ionic equation. This equation represents the reaction taking place most accurately, because it focuses the attention on the newly formed substance, so it is most commonly used to represent precipitation reactions.

SECTION REVIEW 17.3

REMEMBERING

1 Define:
- **a** dissociation
- **b** electrolytic solution
- **c** complete ionic equation
- **d** overall equation
- **e** net ionic equation
- **f** spectator ion.

UNDERSTANDING

2 Write overall equations and net ionic equations for the reactions (if any) that occur when solutions of the following pairs of substances are mixed. If there is no reaction, write 'NR'.
- **a** Sodium chloride and silver nitrate
- **b** Copper(II) sulfate and potassium hydroxide
- **c** Nickel chloride and potassium sulfate
- **d** Sodium carbonate and iron(II) sulfate
- **e** Zinc nitrate and ammonium sulfide
- **f** Potassium carbonate and calcium chloride

17.4 Mandatory practical

PRACTICAL ACTIVITY 17.4.1

Identifying unknown ions in solution

In this experiment, you will conduct a series of tests to identify specified anions in a solution that contains only one anion. Many of the reactions that occur in this experiment are complex, so you are not expected to write equations for them. The main purpose of this experiment is to understand the processes used.

AIM

To conduct a series of chemical reactions to devise chemical tests for identifying the phosphate, sulfate, carbonate and chloride anions (PO_4^{3-}, SO_4^{2-}, CO_3^{2-} and Cl^-) in solution when they are the only ions that could be present

MATERIALS

- dropper bottles containing 0.1 mol L^{-1} of test solutions: $Pb(NO_3)_2$, $Ba(NO_3)_2$, HNO_3, $AgNO_3$ and NaOH
- dropper bottles containing 0.1 mol L^{-1} of anion solutions: Na_2SO_4, Na_2CO_3, Na_3PO_4 and NaCl
- 4 test tubes and test-tube rack
- distilled water
- water bath

PROCEDURE

1 Copy and complete the risk assessment table. Add any more risks you can think of, and how to manage them. In particular, expand on the two risks listed to identify specific risks involved with each.

WHAT ARE THE RISKS IN DOING THIS INVESTIGATION?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?
Concentrated NaOH and HNO_3 are corrosive.	Wear safety glasses and protective clothing. Do not allow contact with skin or clothes. If contact occurs, wash the area with large amounts of water for 10 to 15 minutes.
Some metal salts (silver, barium and lead) are poisonous. Silver nitrate will leave a brown stain on clothing and skin.	Avoid contact with skin and clothes. Wash hands thoroughly after use. Dispose of the chemicals as directed by your teacher.

RISK ASSESSMENT

2 Add 10 drops of each anion solution to each test tube.

3 Add 10 drops of HNO_3 to each test tube and warm them gently.

4 Record the results.

5 Use these results to describe a test for one particular anion.

6 Thoroughly clean all test tubes with distilled water between tests.

7 For the three remaining anions, add 10 drops each in three test tubes, add 5 drops of HNO_3 solution, and add 5 drops of $AgNO_3$ solution (containing Ag^+).

8 Record the results in a table like the one in the results section. Write 'NP' if no precipitate forms and 'ppt' if precipitate forms and give its colour.

9 Repeat steps **5–7** with $Pb(NO_3)_2$ (containing Pb^{2+}) and then $Ba(NO_3)_2$ (containing Ba^{2+}).

10 After testing the Ba^{2+} solution, add 10 drops of NaOH to each of the test tubes and record any changes.

EXTENSION

1 Use the test results to identify the anion (one per solution) in each of three unknown solutions.
2 Try to identify an unknown solution containing two anions.

RESULTS

TEST SOLUTION/ANION	H^+	Ag^+	Pb^{2+}	Ba^{2+}	Ba^{2+} AND OH^-
CO_3^{2-}					
Cl^{2-}					
SO_4^{2-}					
PO_4^{3-}					

ANALYSIS OF RESULTS

For each test and result below, decide which anion would be the most appropriate fit to these results.

ANION	TEST AND RESULTS
	Gives a ppt with acidified Ag^+ but not with Ba^{2+}
	Gives a ppt with acidified Ba^{2+}
	Gives a ppt with Ba^{2+} in alkaline solution but not acid solution
	Evolves gas with addition of dilute HNO_3

DISCUSSION

1 Is the lead ion useful in identifying anions? Explain why or why not.
2 When identifying anions in a solution in which two anions may be present, it is necessary to destroy any CO_3^{2-} before further testing for other anions can be carried out. How might this be done? (Consider the tests in this experiment.) Why is this necessary?

CONCLUSION

How effective was your identification?

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CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 Define:
  a spectator ion
  b precipitate
  c suspension
  d net ionic equation
  e electrolytic solution.

ELABORATION QUESTIONS

2 Use a solubility table to determine if a precipitation reaction will occur when the following reactants are mixed. Where a reaction occurs, write the balanced overall equation and net ionic equation for the reaction.
  a Potassium chloride and zinc nitrate
  b Ammonium bromide and lead nitrate
  c Ammonium sulfide and magnesium acetate
  d Strontium chloride and zinc sulfate

3 Identify solutions that would be mixed to produce each of the following precipitates.
  a Lead sulfate
  b Iron(II) sulfide
  c Magnesium hydroxide

4 A solution was thought to contain either lead nitrate or calcium nitrate. Describe the tests that could be carried out to identify the substance correctly.

EVIDENCE QUESTIONS

5 A solution is known to contain either sodium carbonate or sodium chloride. Samples of the solution evolved a gas with HNO_3 but no precipitate with $AgNO_3$. Which anion is present? Explain.

6 A solution is believed to contain either barium chloride or calcium chloride. Samples of the solution give a white precipitate with Na_2SO_4 and a pale green flame colour. Which of the ions is present? Explain.

7 A student is given four solutions. Each contains one of the following cations: Pb^{2+}, Cu^{2+}, Ba^{2+}, Ca^{2+}, Fe^{2+} or Fe^{3+}. The results of a series of tests that were conducted are given in the following table, where NP stands for no precipitate and ppt stands for precipitate.

REAGENT ADDED	A	B	C	D
KI	Yellow ppt	NP	NP	NP
H_2SO_4	ppt	NP	ppt	ppt
NaOH	ppt	Brown ppt	NP	ppt

Use the results to identify possible cations for A–D. If there is not enough information, suggest other tests and their expected results that could be used to identify any still unknown solutions.

8 A chain of coffee shops started giving away brightly coloured drinking mugs. It was thought that lead compounds had been used in the paint and that this dissolved in acidic drinks such as soft drinks and fruit juices.

Design an experiment to determine:
  a whether lead is present in the paint.
  b the quantity of lead dissolved in 100 mL of an acidic drink.

END-OF-CHAPTER EXAM

End-of-chapter test

1 Which of the following solutions, when added to an aqueous solution of iron(II) sulfate, $FeSO_4(aq)$, will not form a precipitate?

A Barium chloride, $BaCl_2(aq)$

B Sodium hydroxide, $NaOH(aq)$

C Lithium carbonate, $Li_2CO_3(aq)$

D Ammonium nitrate, $NH_4NO_3(aq)$

2 Which of the following reactants, added to a solution of sodium sulfate, would form a precipitate?

A Potassium sulfate solution

B Copper(II) chloride solution

D Ammonium carbonate solution

C Barium nitrate solution

3 What is the precipitate that forms when lead chloride solution is mixed with sodium iodide solution?

A Lead solid

B Lead iodide

C Sodium chloride

D Sodium metal

4 If you added an aqueous solution to an aqueous solution of barium chloride and a precipitate formed, what might the added solution have contained?

A Ammonium sulfate

B Copper(II) nitrate

C Lithium hydroxide

D Potassium iodide

5 When an aqueous lithium chloride solution is added, which of the following would form a precipitate?

A $AgNO_3(aq)$

B $CuSO_4(aq)$

C $Fe_2(SO_4)_3(aq)$

D $Ca(NO_3)_2(aq)$

6 When aqueous solutions of barium nitrate and chromium(III) sulfate are mixed, a white precipitate forms. Write a net ionic equation for this reaction.

7 What solution would you mix with calcium chloride to obtain a precipitate of the cation?

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8 Solid sodium chloride is found in deposits well below ground level. Cold water is pumped underground at high pressure to bring the sodium chloride to the surface as brine, a concentrated aqueous solution.

a Suggest why it might be better to use hot water rather than cold water for this process.

b Suggest one reason why cold water is usually used rather than hot water. Often, the brine produced can contain magnesium ions as contaminants. These magnesium ions can be removed by precipitation.

c Suggest a suitable chemical to add to the brine to precipitate out the magnesium ions.

d Write the net ionic equation for this precipitation reaction.

18 SOLUBILITY

Introduction

This chapter examines the relationship between solubility and temperature for different substances, including ionic and organic compounds, as well as gases. Solubility curves are used to determine whether or not a solution is unsaturated, saturated or supersaturated. For solutions of solid compounds we use solubility curves to predict the amount of crystallisation that can occur when a solution cools.

Stimulus questions

- What is the relationship between solubility and temperature?
- How can a scientist predict the amount of sodium chloride that will dissolve in water at 60°C?
- How does the relationship between solubility and temperature differ for gases relative to solids?

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18.1 Solubility and intermolecular bonding

Solutions form when a solute dissolves in a solvent, creating a **homogenous** mixture where pieces of solid solute can no longer be observed. When this occurs, the forces that have been holding together the solute particles are overcome, and new interactions form between the water and the solute particles, causing them to mix. Not all substances are soluble in water. This varying solubility of different substances with water or indeed any other solvent can be explained by comparing the relative strength of these bonds.

The process of breaking chemical bonds requires energy, so it is considered **endothermic**. However, energy is released when new bonds are made: this is an **exothermic** process. For a substance to be soluble in water, the energy released when the new bonds are formed with water must be greater than the energy needed to separate the solute particles.

A variety of different substances dissolve in water. The specific types of bonds that exist when a solution forms depend on the structures of the solute. As a polar solvent, water can form different intermolecular bonds, including ion–dipole, dipole–dipole and hydrogen bonds, which is why it can dissolve various types of compounds, both ionic and covalent in structure.

homogeneous outwardly appearing to be made up of only one substance

endothermic reaction energy is absorbed from the surroundings, so the temperature falls

exothermic reaction energy is released to the surroundings, so the temperature increases

Solutions of ionic compounds

When an ionic salt such as sodium chloride dissolves in water, ion–dipole bonds are formed. The anion of the salt attracts the positive part of the water molecule. More water molecules surround the anion until it is perfectly hydrated. The cation attracts the negative part of the water molecule, and a similar process occurs until the ionic bonds within the salt crystal are broken and the salt has dissolved.

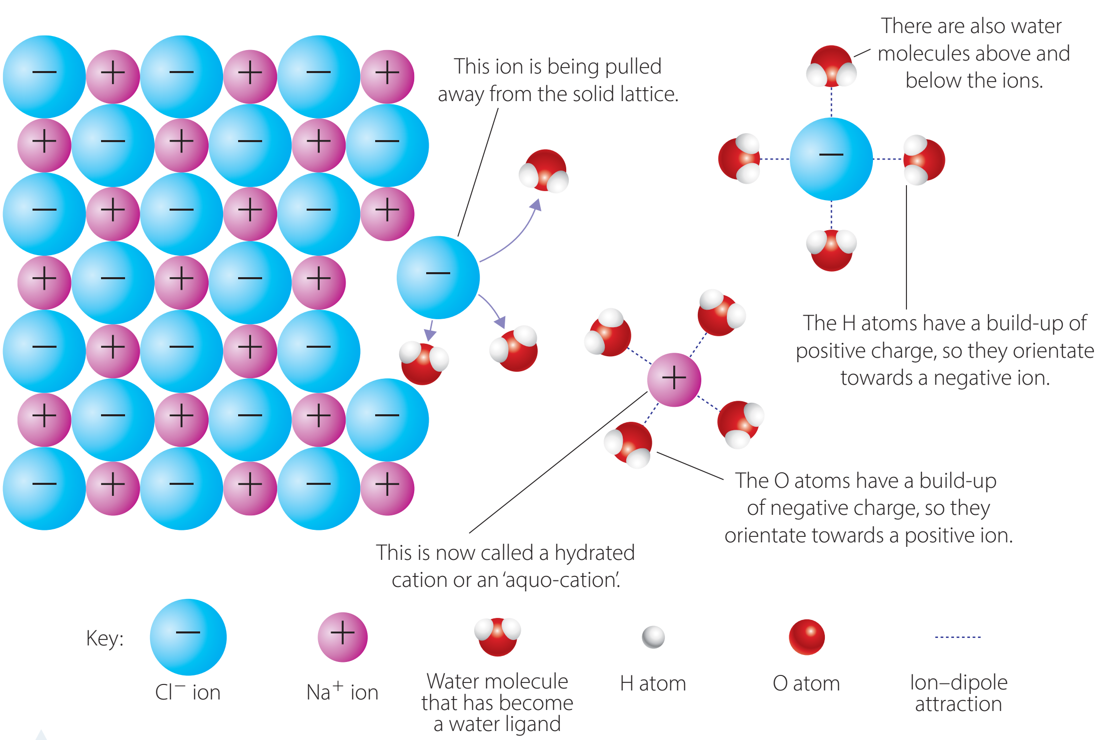


FIGURE 18.1.1 A simplified model of the solution process for a sodium chloride crystal in water

18.1.1 Analysis of ionic compounds: solubility

However, not all ionic compounds are soluble in water. Figure 18.1.2 illustrates this, clearly outlining the role of energy factors in controlling the solubility of ionic compounds.

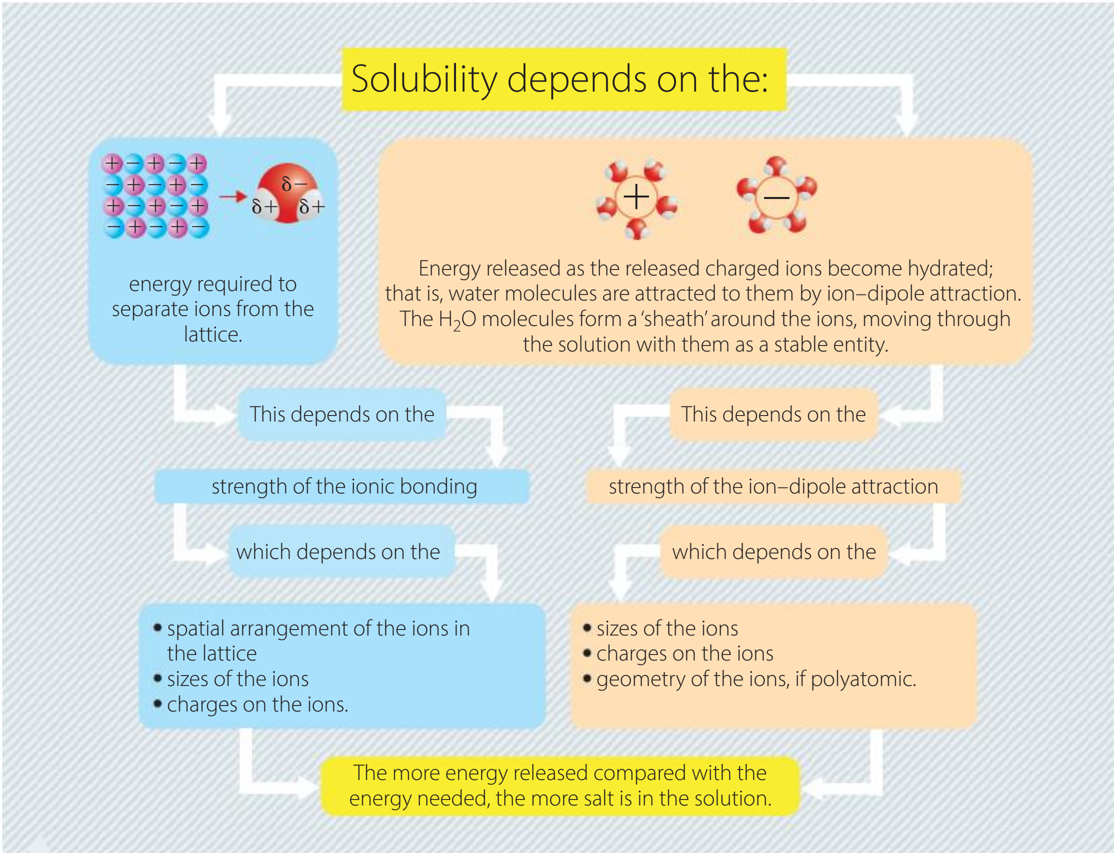


FIGURE 18.1.2 Factors affecting the solubility of ionic compounds

Organic compounds as solutes

Most organic molecules are non-polar, so dispersion forces are the only type of intermolecular bonding that exists between them. This means that, when they are mixed with water, a strong reaction with the highly polar water molecules does not occur, so these molecules are insoluble.

However, when polar groups are attached to the carbon chain, the hydrophobic nature of organic molecules decreases. Examples of groups that give polarity include –OH, C=O and –NH. These functional groups can form hydrogen bonds and dipole–dipole bonds with water molecules, increasing the solubility of the molecule as a whole.

Table 18.1.1 (page 343) shows the solubilities of a series of different alkanol molecules. These are molecules with a non-polar hydrocarbon chain and a polar –OH bond.

TABLE 18.1.1 The solubility of alkanols in water

NAME	MOLECULAR FORMULA	SOLUBILITY IN WATER (g/100 mL)
Methanol	CH_3OH	Infinite
Ethanol	C_2H_5OH	Infinite
Propanol	C_3H_7OH	Infinite
Butanol	C_4H_9OH	8.0
Pentanol	$C_5H_{11}OH$	2.2
Hexanol	$C_6H_{13}OH$	0.7

It is clearly evident that, the greater the number of carbon and hydrogen atoms (and the number of non-polar C–C and C–H bonds), the lower the solubility of the overall molecule.

If a large proportion of polar groups is present, then larger organic molecules such as glucose, vitamins, enzymes and hormones may be sufficiently polar that they can dissolve in water.

It should be noted that there are also some organic molecules, such as carboxylic acids, which dissociate to form ions when added to water. These molecules are capable of forming additional ion–dipole bonds with water molecules, so they have significantly greater solubility than other organic molecules that do not ionise when added to water. This can be represented by the equation below:

$$\underbrace{CH_3COOH + H_2O}_{\text{Ethanoic acid}} \rightarrow CH_3COO^- + H_3O^+$$

Ethanoic acid

the positive hydronium and negative acetate ion formed can now form ion–dipole bonds with water molecules and are highly soluble

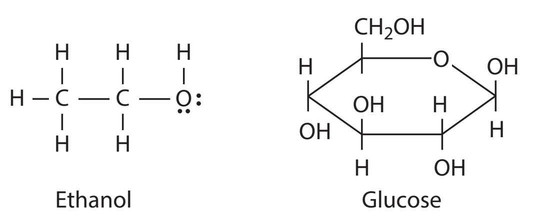


FIGURE 18.1.3 Ethanol and glucose molecules are polar due to the presence of polar regions.

PRACTICAL ACTIVITY 18.1.1

Disappearing act

AIM

To discover what happens when two liquids mix

EQUIPMENT PER GROUP

- 2 accurately measured identical volumes of ethanol and water
- large measuring cylinder

PROCEDURE

1. Any one of three things could happen when mixing the two liquids. The total volume could stay the same as the combined individual volumes, it could be greater, or it could be less. For example, if 50 mL ethanol and 50 mL water are combined, is the final volume 100 mL, more, or less? Predict what you think will happen.
2. Carefully pour the water and ethanol into a measuring cylinder.

QUESTIONS

1. How many of the class accurately predicted the result?
2. Would you expect the same result if ethanol was poured first, or poured faster or slower? What causes any gas evolution you have observed? Can you repeat the experiment without bubbles?
3. Was this a chemical reaction? What must occur for it to be a chemical reaction? Did you have any evidence for this?
4. What interaction could occur between the molecules?
5. Summarise what must have happened as the two liquids mixed.

PRACTICAL ACTIVITY 18.1.2

Vitamins: good or bad?

Vitamins are a group of essential nutrients that organisms only need in small amounts. They are a diverse group of organic compounds. For example, unlike other animals, humans must consume vitamin C. A lack of vitamin C causes scurvy – a deadly disease that used to kill sailors on long voyages. In 1770, Captain James Cook accepted the scientific view that a diet including citrus fruit was essential. He was the first to circumnavigate the world without losing a man to scurvy. An excess of vitamin C is never fatal.

Vitamin A is believed to have caused the death of Antarctic explorer Xavier Mertz in 1913. Short of food Mertz and fellow explorer Douglas Mawson resorted to eating their sled dogs. Canine liver is high in vitamin A and an excess of this vitamin could have caused his death. Vitamin A is a non-polar fat-soluble vitamin. It accumulates in the liver and is harder to remove from the body than water-soluble vitamins. This causes toxicity in large doses.

»

»

1 Copy the vitamin molecules listed below. For each vitamin, identify and circle the polar groups.

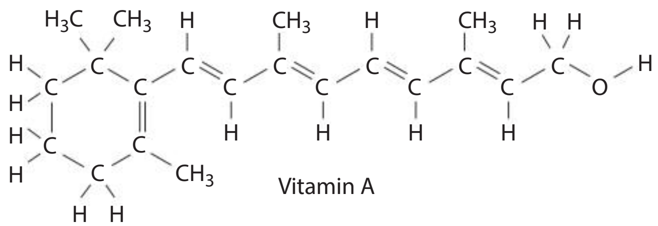
2 a Why is vitamin C considered polar and water-soluble?

 b Why is vitamin A considered non-polar and fat-soluble?

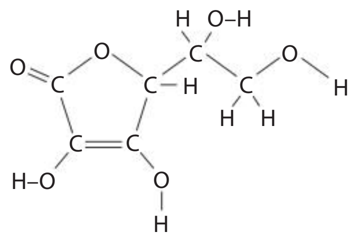
 c How does this explain why excessive vitamin C is never fatal but excessive vitamin A can be fatal?

3 Decide whether the other vitamin molecules overall are polar or non-polar. Justify your decisions.

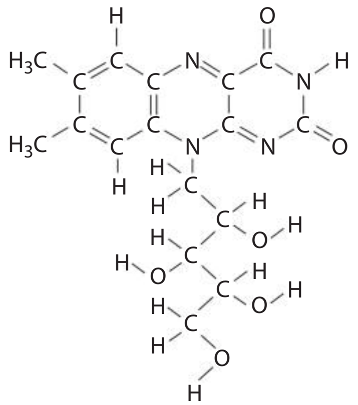
4 On the basis of polarity, classify vitamins B, D and E as likely to be water- or fat-soluble. Research to see if your predictions are correct.



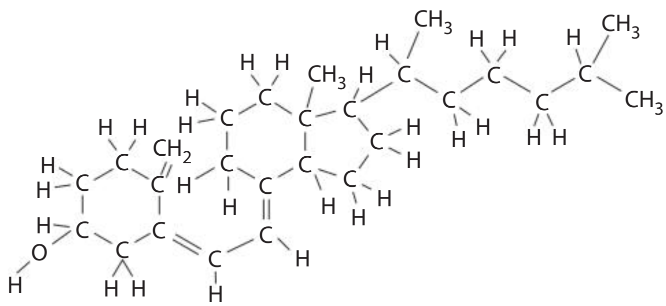
Vitamin A



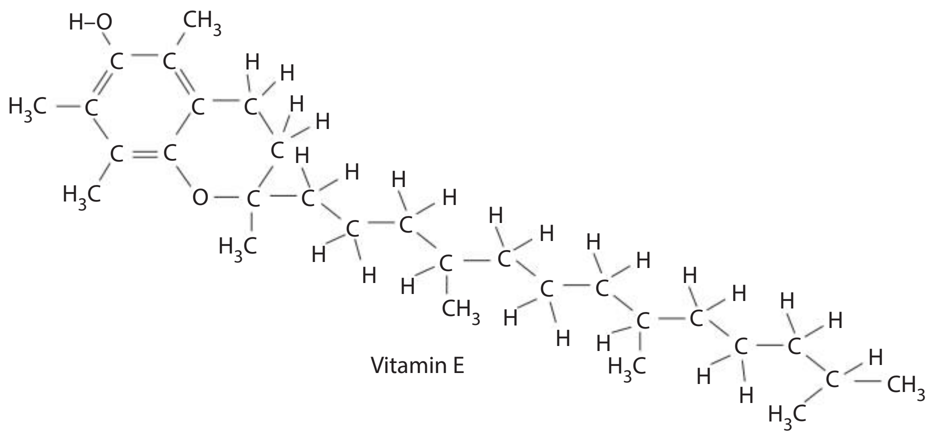
Vitamin C



Riboflavin
(Vitamin B_2)



Vitamin D_3



Vitamin E

CONCLUSION

Research how excess chemicals are removed from the body.

SECTION REVIEW 18.1

REMEMBERING

1 Draw a diagram to show sodium chloride (NaCl) dissolving in water.

2 Write the formula for:

- **a** lithium bromide (LiBr) crystallising as a dihydrate
- **b** calcium nitrate ($Ca(NO_3)_2$) crystallising as a tetrahydrate
- **c** magnesium nitrate ($Mg(NO_3)_2$) crystallising as a hexahydrate.

UNDERSTANDING

3 Hydrogen bromide and methane are small covalent molecules. Explain why methane does not dissolve in water but hydrogen bromide does.

4 Describe what happens when sugar is added to a cup of coffee. Identify the solute and solvent.

5 KBr dissolves in water by dissociation.

- **a** What does this statement mean?
- **b** Write an ionic equation to show what happens.
- **c** Is this solution an electrolyte?

6 Describe what happens on a molecular level when water and ethanol mix. Explain, using the terms 'solute' and 'solvent'. Which bonds are broken and which bonds are formed? Write an equation to show this.

APPLYING

7 Copy and complete the following table, filling in the gaps.

BONDING TYPE	SOLUBILITY IN WATER	EXAMPLES
Ionic	Most are soluble	a ______________
b ______________ ______________ ______________	Soluble if hydrogen bonds possible Soluble by reacting with water Otherwise insoluble	Ethanol, glucose HCl, HNO_3, NH_3 Dichloromethane
Non-polar molecular	c ______________ Most are insoluble	O_2, I_2 d ______________
Very large molecules	Insoluble if highly structured f ______________	e ______________ Starch, glycogen, enzymes
g ______________	h ______________	Diamond, SiO_2
Metals	i ______________ Unless they react with water	j ______________ Li, Na, K, Ca, Ba

8 Classify the following as either soluble in water or not soluble. Explain your reasoning.

a	HBr	**b**	$MgCl_2$	**c**	CS_2
d	$ZnCl_2$	**e**	H_2SO_4	**f**	$SiCl_4$
g	Glucose (a sugar, $C_6H_{12}O_6$)	**h**	C_8H_{18}	**i**	NH_2CONH_2

ANALYSING

9 A sample weighing 6.25 g of blue hydrated copper(II) sulphate, $CuSO_4 \cdot xH_2O$, (x unknown) is gently heated in a crucible until the mass remaining is a constant 4.00 g. Determine the formula of the hydrated salt (the value of x in $CuSO_4 \cdot xH_2O$).

10 A student determines the water of crystallisation of magnesium sulfate. The procedure heats a known mass of the hydrated salt in a crucible until no further water is driven off and the weight remains constant despite further heating. Unfortunately, the student becomes confused with how to carry out the calculations, and makes several errors.

The results and calculations are shown below. What are the errors? Identify the initial weight for the hydrated sulfate and determine the correct value for x.

Mass of hydrated magnesium sulfate = ? Dried weight = 1.65 g

Empirical formula: $MgSO_4 \cdot xH_2O$	$\frac{1.65}{120}$	:	$\frac{3.228}{138}$	This is to convert mass to moles.
	0.01375	:	0.02339	Next calculate simplest ratio
	1		1.7	The calculated whole number ratio
	2		3	The formula for hydrated sulfate: $2MgSO_4 \cdot 3H_2O$

18.2 The effect of temperature on solubility

Solubility can be measured as the amount of solute dissolved in a fixed amount of water. A common unit used is g/100g – the mass of solute dissolved in 100 g of water. The solubility of a substance depends upon temperature. Exactly how solubility varies with temperature depends on other factors, including the state of the solute.

Temperature dependence of the solubility of solids

For most solids dissolving in water, the solubility of the substance increases with temperature.

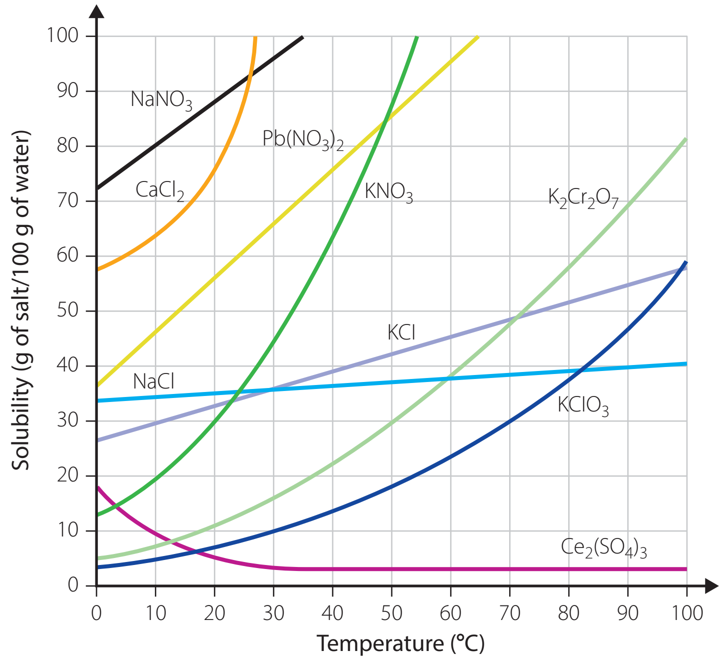
Everyday experience teaches us that sugar or salt can be dissolved in water more quickly by stirring the water with a spoon, causing the particles of solute and solvent to move more quickly. As a result, increasing the temperature has a very similar effect to stirring on the solubility, although it can be a more marked effect because the consequent increase in the movement of the particles is greater.

Applying heat to a liquid increases the temperature, which is a measure of the kinetic energy of the molecules, causing the molecules to move more quickly. The solvent molecules collide with the solute molecules with a greater impact, and effectively break apart the pieces of solute and overcome the attractive forces that hold them together. In addition, when they are moving at a higher rate, the pieces of solute are less likely to hold together, so they dissolve more readily.

A **solubility curve** is a graph that shows the variation of solubility with temperature. Figure 18.2.1 (page 348) shows solubility curves for several different ionic compounds. As shown, the solubility of different compounds increases at markedly different rates with temperature, and the solubility of cerium sulfate ($Ce_2(SO_4)_3$) actually decreases with temperature. This is due to the varying strength of the ionic bonds inside the ionic solids, which need to be overcome, and the variable nature of the ion–dipole interactions between the ions and water molecules that form during the process of dissolving.

solubility curve a graph showing the variation of the solubility of a substance in water with temperature

FIGURE 18.2.1 Solubility curves of various ionic compounds



The solubility of gases in water with temperature

18.2.1 Solubility of gases in water

The solubility of gases in water is much lower than that of solids. In addition, as the temperature is increased, the solubility of gases decreases rather than increases. Figure 18.2.2 demonstrates this trend for the gases carbon dioxide and oxygen.

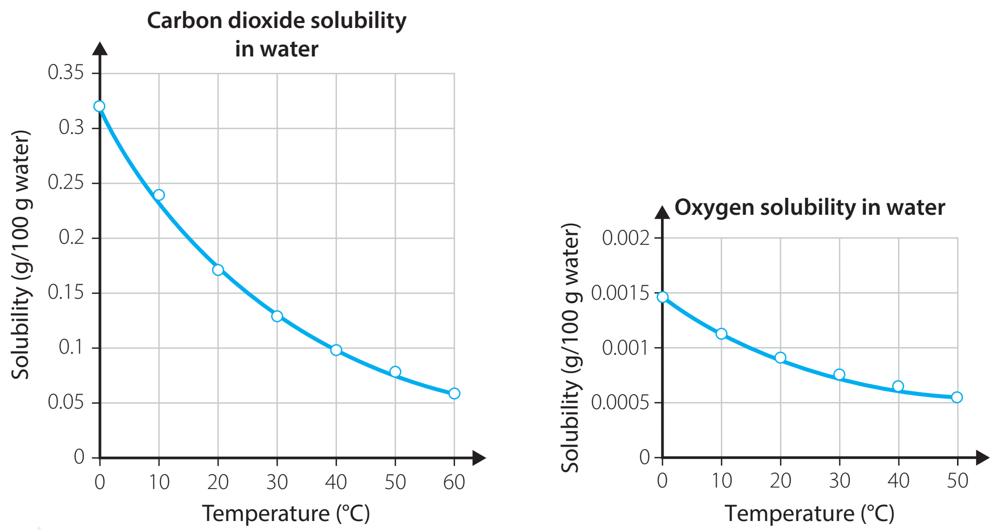


FIGURE 18.2.2 Solubility of carbon dioxide and oxygen in water

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This apparently contradictory trend is explained again by considering the forces that need to be overcome and the interactions that form when dissolution of gases takes place. The forces of attraction between the molecules of most gases are negligible. When the interactions with water molecules form, heat is released in an exothermic process. This means that, if the temperature is increased, it has the effect of reversing the exothermic dissolution process and reforming the gas.

Some of the most life-rich oceans of the world are off the coast of Antarctica, where the water is very cold, and the solubility of oxygen in the seawater is high. One current concern of climate scientists is that increasing atmospheric temperatures cause similar changes to the temperature of seawater, decreasing the amount of oxygen dissolved, with potentially large consequences for organisms living in the sea.

SECTION REVIEW 18.2

REMEMBERING

1 Define:

a temperature

b solubility curve.

UNDERSTANDING

2 Explain why the solubility of solids increases with temperature, yet the solubility of gases decreases with temperature.

APPLYING

3 Suggest reasons why scientists are concerned at the effect that rising sea temperatures, caused by climate change, might have on the Great Barrier Reef.

18.3 Interpreting and analysing solubility curves

Information from solubility curves can be used to determine predictive information about solutions, and the quantity of solute that will dissolve at a particular temperature.

WORKED EXAMPLE 18.3.1

Refer to the solubility curve in Figure 18.2.1 (page 348).

1 How much potassium dichromate ($K_2Cr_2O_7$) will dissolve at 20°C?

2 How much extra will dissolve at 80°C?

ANSWERS

1 At 20°C, approximately 12 g of $K_2Cr_2O_7$ will dissolve in 100 g of water.

a Find the curve for $K_2Cr_2O_7$ on the graph.

b Draw a line from 20°C up to the potassium dichromate curve.

c Draw a line from this point across to the *y*-axis. This indicates the amount that will dissolve at 20°C (Figure 18.3.1).

2 The extra amount that would dissolve at 80°C is 46 g.
Similarly, if the temperature is increased to 80°C, approximately 58 g of the substance will dissolve in 100 g (or 100 mL) of water.
58 – 12 = 46 g
This is also the amount that would crystallise out of solution if the temperature of the saturated solution was cooled from 80°C to 20°C.

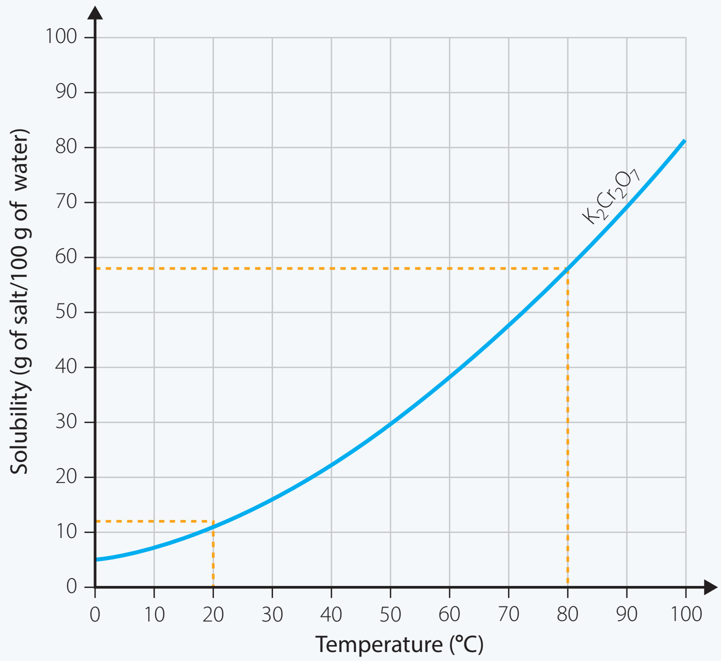

FIGURE 18.3.1 Calculating the amounts of potassium dichromate ($K_2Cr_2O_7$) that can dissolve at 20°C and at 80°C. The difference is how much extra or how much potassium dichromate will crystallise out of the solution on lowering the temperature from 80°C to 20°C.

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PRACTICAL ACTIVITY 18.3.1

Deriving the solubility curve of ammonium chloride

AIM

To gather data to draw a solubility curve of ammonium chloride

MATERIALS

- 250 mL beaker
- 10 mL measuring cylinder
- large test tubes
- Bunsen burner
- tripod
- retort stand
- bosshead and clamp
- 0–100°C thermometer
- stirring rod
- distilled water
- balance
- weighing bottle
- safety glasses

PROCEDURE

1 Copy and complete the risk assessment table. Add any more risks you can think of, as well as ways to manage them. In particular, expand on the three risks listed to identify specific risks involved with each of them.

WHAT ARE THE RISKS IN DOING THIS EXPERIMENT?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?
Ammonium chloride is slightly toxic.	Work in a well-ventilated area. Wear safety glasses and wash your hands after the experiment. Dispose of all chemicals according to your teacher's directions. Do not pour them down the sink.
The Bunsen burner will get hot.	Do not use the Bunsen burner if the gas tube is damaged. Ensure long hair is tied back and the flame is away from flammable material. If you burn yourself, place the affected area under cold running water for 10 minutes and inform your teacher.
Broken glassware will cut.	Inspect and discard any chipped or cracked glass wear, no matter how small the damage. Sweep up broken glass with a brush and dustpan; do not use your fingers.

2 Carefully weigh out the amount of ammonium chloride assigned to your group.

3 Place the weighed ammonium chloride into a test tube.

4 Add 10 mL distilled water.

5 Half-fill the beaker with tap water. Clamp the test tube securely into the beaker so that it is immersed in the water.

6 Warm the beaker, stirring the contents of the test tube with the stirring rod. Continue until the ammonium chloride has dissolved completely.

7 Allow the test tube to cool. Continue to stir it until tiny flakes of ammonium chloride start to appear. Note the temperature at which this occurs.

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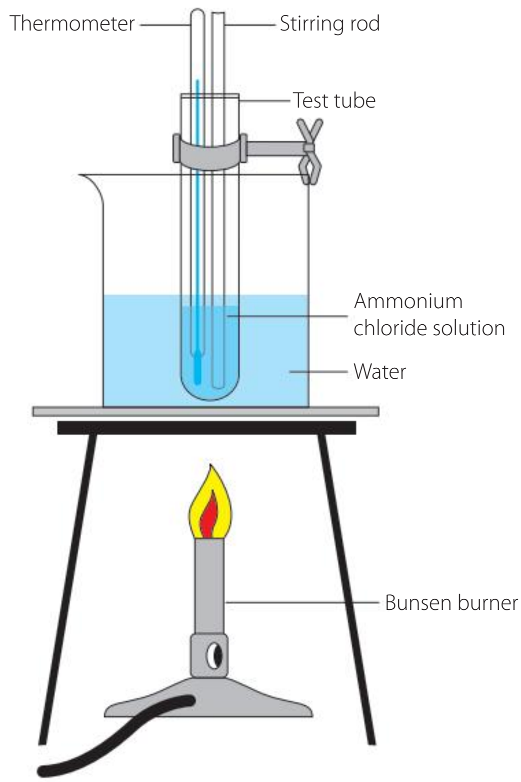
FIGURE 18.3.2 Experimental set-up for determination of a solubility curve



RESULTS

Copy and complete the following table.

GROUP	MASS OF AMMONIUM CHLORIDE (g)	MASS OF WATER (g)	TEMPERATURE OF RECRYSTALLISATION (°C) (*x*-AXIS)	SOLUBILITY (g/100 g H_2O) (*y*-AXIS)
1	4.00	10.00		
2	4.50	10.00		
3	5.00	10.00		
4	5.50	10.00		
5	6.00	10.00		
6	6.50	10.00		

ANALYSIS OF RESULTS

1. Use the class results to plot a graph of solubility against temperature. Plot the temperature along the horizontal *x*-axis (0–100°C).
2. From the graph, predict the solubility of ammonium chloride at 20°C, 40°C, 60°C and 80°C.

DISCUSSION

1. Describe what happens to the solubility of ammonium chloride as temperature increases.
2. At the point the crystals first appear, the solution is saturated. What does 'saturated' mean?
3. At what temperature would 10 g of ammonium chloride saturate 50 g of water?
4. If a saturated solution of ammonium chloride at 90°C is cooled to 10°C, how many grams of ammonium chloride would crystallise out of 100 g of water?

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5 If the theoretical value for the solubility of ammonium chloride at 50°C is 50 g/100 g, what percentage error does your experiment have?

$$\text{Percentage error} = \frac{\text{experimental value - true value}}{\text{true value}}$$

6 List possible sources of errors in your experiment.

TAKING IT FURTHER

Describe modifications to the experiment that could improve the accuracy.

SECTION REVIEW 18.3

UNDERSTANDING

Use the solubility curve in Figure 18.2.1 (page 348) to answer the following questions.

1 What mass of the following solutes will dissolve in 50 mL of water?
 - a $KClO_3$ at 65°C
 - b $CaCl_2$ at 25°C
 - c $Pb(NO_3)_2$ at 55°C

2 Is a solution of 45 g/100 g of KCl at 70°C unsaturated, saturated or supersaturated?

3 Determine which of $K_2Cr_2O_7$, KCl and $Ce_2(SO_4)_3$ is most soluble in water at 15°C.

4 What is the maximum mass of KNO_3 that could be dissolved in 25 mL of water at 40°C?

5 200 mL of a saturated solution of $K_2Cr_2O_7$ is made at 70°C. The solution is cooled to 30°C. Calculate the mass of solid that would be precipitated by the cooling of this solution.

APPLYING

6 What mass of the following solutes will dissolve in 100 mL of water?
 - a $K_2Cr_2O_7$ at 50°C
 - b NaCl at 100°C
 - c $NaNO_3$ at 10°C

7 Determine which of $K_2Cr_2O_7$, NaCl and $NaNO_3$ is most soluble in water at 15°C.

8 At 30°C, what mass of KCl and NaCl would just dissolve in 100 mL of water? For this mass at this temperature, would the solution be saturated, unsaturated or supersaturated?

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 State the relationship between solubility and temperature of most ionic solids.

2 State the relationship between the solubility and temperature of most gases.

CATEGORY QUESTIONS

3 Describe what happens at the molecular level when an ionic substance, such as potassium iodide, dissolves in water.

4 Suggest why cerium sulfate has a low solubility in water, despite being ionic.

ELABORATION QUESTIONS

5 Use Figure 18.2.1 (page 348) to determine the mass of potassium nitrate that will dissolve in:

- **a** 100 g of water to form a saturated solution at 20°C
- **b** 50 g of water to form a supersaturated solution at 30°C
- **c** 80 g of water to form a saturated solution at 10°C.

6 Figure 18.2.1 shows the solubility curve of potassium chloride, KCl.

- **a** Use diagrams to model how KCl dissolves in water.
- **b** Write a balanced equation for the dissolution of KCl in water.
- **c** How much KCl would dissolve in 50 mL of water at 80°C?
- **d** If the solution in part **c** were cooled to 10°C, would it be saturated, unsaturated or supersaturated? Justify your answer.
- **e** To what temperature do you need to heat 100 g of water so that 40 g of KCl dissolves completely? If you cooled this to 10°C, how much KCl would precipitate out of the solution?

EVIDENCE QUESTIONS

7 Steroids do not circulate in the blood system by themselves. Infer the polarity of steroid hormones. Dopamine can travel through the blood system easily. What does this indicate about the polarity of dopamine? Which of these is more likely to be injected directly into the veins?

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END-OF-CHAPTER EXAM

End-of-chapter test

CHEMISTRY

1 When sodium chloride dissolves in water, the water molecules will form:

A dipole-dipole bonds with the sodium chloride.

B hydrogen bonds with the sodium and chloride ions.

C covalent bonds with the sodium chloride.

D ion-dipole bonds with the sodium and chloride ions.

Questions 2 and 3 to refer to the following information.

The solubility of sodium nitrate ($NaNO_3$) at different temperatures is shown in the table below.

TEMPERATURE (°C)	SOLUBILITY (g/100 g)
10	21
30	48
50	88
70	138

2 210 g of sodium nitrate dissolves in 30 g of water at 70°C. The solution cools to 10°C without forming a supersaturated solution.

What is the mass that would crystallise out of the solution?

A 3.7 g

B 6.0 g

C 6.3 g

D 7.9 g

3 What would be the approximate number of nitrate ions dissolved in a 50.0 mL saturated solution of potassium nitrate at 50°C?

A 3.1×10^{23}

B 6.2×10^{23}

C 3.1×10^{25}

D 6.2×10^{25}

4 When the molecular compound ethanol, CH_3CH_2OH, is added to water, it:

A forms hydrogen bonds to water molecules and dissolves.

B forms ions by reaction with water molecules and dissolves.

C is immiscible with water and forms a separate layer.

D dissociates into ions and dissolves.

5 When ethanoic acid, CH_3COOH, is added to water, it:

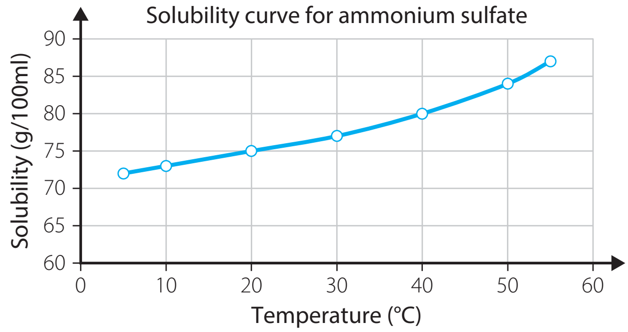
A forms hydrogen bonds to water molecules and dissolves.

B forms ions by reaction with water molecules and dissolves.

C is immiscible with water and forms a separate layer.

D dissociates into ions and dissolves.

6 The diagram below shows part of the solubility curve for ammonium sulfate, $(NH_4)_2SO_4$.



a A student places 205 g of solid ammonium sulfate in a beaker and adds 250 mL of deionised water. To what temperature will student have to heat the mixture to produce a saturated solution?

b When 75 mL of a saturated solution of ammonium sulfate at 50°C is cooled to 30°C, what mass of solid will be crystallised?

c A 6.5 M solution of ammonium sulfate is heated to a temperature of 40°C. Is this solution unsaturated, saturated or supersaturated? Explain your reasoning.

7 Potassium nitrate is a soluble substance with its solubility values listed in the table below.

SOLUBILITY (g/100 g WATER)	TEMPERATURE (°C)
8.5	0
31.5	20
105	60

a Using the data above, how much potassium nitrate would you need to dissolve in 60 mL of water to make a saturated solution at 60°C?

b A 90 g sample of potassium nitrate is dissolved in 350 mL at 20°C. Is the solution saturated? Explain your answer.

c Calculate the molarity of a 100 g saturated potassium nitrate solution at 0°C.

d Draw a labelled diagram to show how sodium chloride dissolves in water.

19 pH

Introduction

Acids and bases are all around us. At home, in industry and in our environment, acids and bases are a part of our lives. These substances vary greatly in their degree of acidity and basicity and it is important to be able to identify this. The Swedish scientist Svante Arrhenius recognised that the acidity of a solution was related to the concentration of hydrogen ions (H^+) in solution but the calculations required to compare solutions were cumbersome. The Danish biochemist Soren Sorensen devised a scale to make writing the hydrogen ion concentration in solution more manageable. This became known as the pH scale. pH is a measure of the degree of acidity or alkalinity of a solution.

Stimulus questions

How would a scale measuring the acidity or basicity of substances be useful, and possibly life-saving, to you at home?

How does the pH of a substance relate to the concentration of hydrogen ions it contains?

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19.1 pH

acid
a substance that produces hydrogen ions (H^+) in aqueous solution

base
a substance that either contains the oxide (O^{2-}) or hydroxide ion (OH^-), or produces hydroxide ions in an aqueous solution

When people hear the term **acid** they automatically think of something that is highly corrosive and should be avoided, yet many substances in the world are acidic. Rain is slightly acidic, as are many foods and drinks. Soft drink contains carbonic acid, citrus fruits contain citric acid and many foods contain vitamin C, which is ascorbic acid. Acids are also important components of many of the compounds that make up living things. Proteins are long chains of amino acids and the stomach produces hydrochloric acid to aid with digestion of food.

Bases are also part of the natural environment, although they are not as commonly recognised as acids. Ammonia is a base found in volcanic gases and produced by rotting plant and animal matter. Caffeine, found in coffee and tea, is a base, as is nicotine, which is in tobacco.

19.1.1 Acids and alkalis
19.1.2 Acids, bases, salts

FIGURE 19.1.1 Caffeine, found in the coffee bean (shown), is a base. Of the two species of coffee, *Coffea arabica* and *Coffea canephora*, the latter has a higher caffeine content.

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Definitions of acids and bases

Acids and bases are families of substances. Members of each family can be recognised by their common properties.

Acids have common properties such as a sour taste and the ability to:

- sting or burn the skin
- conduct electricity in solution
- turn blue litmus red.

Bases have common properties such as a bitter taste and the ability to:

- feel slippery in aqueous solution
- conduct electricity in solution (not all bases are soluble)
- turn red litmus blue.

The effects of acids and bases on vegetable dyes such as litmus were initially used to classify substances as acids and bases. While practical definitions are useful in the laboratory, chemists attempt to define substances by what they are as well as by what they do. The definition of an acid has changed several times over the last 250 years. In the 18th century, French scientist Antoine Lavoisier was one of the first to try to establish the chemical definition of an acid. In the late 19th century, Swedish chemist Svante Arrhenius attempted to explain acids and bases in terms of the particles they produced in aqueous solution. He proposed that acids were substances that ionised in solution to produce hydrogen ions (H^+) and that acids were strong if they ionised almost completely, and weak if they ionised only slightly. Arrhenius defined a base as a substance that in solution produced hydroxide ions (OH^-). Note that not all bases dissolve in water so this definition has been modified to include insoluble bases.

Bases have a potential for accepting hydrogen ions (H^+). Note that not all oxides are basic. Most oxides formed between a non-metal and oxygen such as carbon dioxide (CO_2) and sulfur trioxide (SO_3) are acidic or neutral, while those formed between a metal and oxygen such as calcium oxide (CaO) and magnesium oxide (MgO) are basic or **amphoteric**.

amphoteric substance
a substance that can act as either an acid or a base depending on the other reactants

Common acids are hydrochloric acid (HCl), sulfuric acid (H_2SO_4) and nitric acid (HNO_3). When these acids are added to water to form acidic solutions, the molecules dissociate, releasing H^+ ions, as shown in the following equations:

$$HCl \rightarrow H^+ + Cl^-$$

$$H_2SO_4 \rightarrow 2H^+ + SO_4{}^{2-}$$

$$HNO^3 \rightarrow H^+ + NO^{3-}$$

Common soluble bases (alkalis) are sodium hydroxide (NaOH), barium hydroxide $(Ba(OH)_2)$, potassium oxide (K_2O) and ammonia (NH_3). These substances either contain, or in aqueous solution release, the hydroxide ion (OH^-). For example:

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

$$K_2O + H_2O \rightarrow 2K^+ + 2OH^-$$

$$NH_3 + H_2O \rightarrow NH_4{}^+ + OH^-$$

The term 'pH' stands for 'power of hydrogen'. This name came about because the scale is based on the concentration of hydrogen ions in solution. Danish biochemist Soren Sorensen devised the scale to make writing the hydrogen ion concentration in solution more manageable. pH is a measure of the degree of acidity or alkalinity of a solution.

The strength of acids and bases

A strong acid is one that dissociates completely into its ions. Hydrochloric acid and sulfuric acid are examples of strong acids.

A weak acid is one that only partially dissociates into its ions. Ethanoic (acetic) acid (CH_3COOH) is an example of a weak acid:

$$CH_3COOH \rightarrow CH_3COO^- + H^+$$

Only a small number (fewer than 1%) of the ethanoic acid molecules dissociate into ions so that, for equal concentrations of hydrochloric acid and ethanoic acid solutions, the hydrochloric acid solution would contain far more hydrogen ions than ethanoic acid.

As with acids, so too can bases vary in strength. The strength of a base comes from the number of OH^- ions it produces in water. A strong base is one in which all of its molecules react to release OH^- ions. Sodium hydroxide (NaOH) and potassium hydroxide (K_2O) are examples of strong bases. A weak base is one in which only a small fraction of the molecules or ions react to release OH^- ions. Ammonia is an example of a weak base.

SECTION REVIEW

19.1

REMEMBERING

1 Define:

a acidic

b basic

c amphoteric

d pH.

UNDERSTANDING

2 Explain the meaning of the Arrhenius theory of acids and bases, with reference to a reaction between phosphoric acid and sodium hydroxide.

APPLYING

3 The table below gives the electrical conductivity values (a measure of the number of ions in solution) for a number of aqueous solutions.

AQUEOUS SOLUTION	ELECTRICAL CONDUCTIVITY SIEMENS PER METER ($S\,m^{-1}$)
HNO_3	108
NaOH	93.1
$Mg(OH)_2$	0.001
NaCl	30.2
NH_3	1.0
CH_3COOH	0.8

a How does electrical conductivity relate to acid and base strength? Justify your explanation by referring to the table.

b Explain why $Mg(OH)_2$ has a very low electrical conductivity.

19.2 pH scale

19.2.1 Indicators
19.2.2 The pH scale

Acids vary in strength depending on the concentration of H^+ ions that are produced when the acid dissolves in water. Similarly, bases vary in strength depending on the concentration of OH^- that are released when the base dissolves in water. Danish biochemist Soren Sorensen who first coined the term 'pH'. In his work on the effect of hydrogen ion concentration on the behaviour of acids, for simplicity, he recorded the H^+ ion concentrations as 5 or 6.7 instead of 10^{-5} or $10^{-6.7}$. This became common practice and influenced how scientists describe the acidity or alkalinity (another way of saying 'basic') of a substance.

The pH scale generally ranges from 0 to 14. The lower the pH, the more acidic a solution is; the higher the pH, the more basic the solution is. Acids have a pH of less than 7. Bases have a pH greater than 7. A substance that has a pH equal to 7 is neutral. The pH scale is logarithmic and, as a result, each whole pH value below 7 is 10 times more acidic than the next higher value. For example, pH 4 is 10 times more

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acidic than pH 5 and 100 times (10×10) more acidic than pH 6. For pH values above 7, each is 10 times more basic than the next lower whole value. For example, pH 10 is 10 times more basic than pH 9 and 100 times (10×10) more basic than pH 8 (Table 19.2.1).

TABLE 19.2.1 The relationship between pH and hydrogen ion concentration

pH	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
H^+ concentration ($mol\,L^{-1}$)	1	10^{-1} or 0.1	10^{-2}	10^{-3}	10^{-4}	10^{-5}	10^{-6}	10^{-7}	10^{-8}	10^{-9}	10^{-10}	10^{-11}	10^{-12}	10^{-13}	10^{-14}

Notice the relationship between the pH value and the superscript number in the hydrogen ion concentration (Table 19.2.1). They are the same number because pH is based on the hydrogen ion concentration of a solution.

An important aspect of the pH scale is that a change in 1 unit on the scale is equivalent to a 10 times change in the hydrogen ion concentration. For example, a change from pH 1 to 2 is equivalent to a change in the hydrogen ion concentration from 0.1 to $0.01\,mol\,L^{-1}$.

Although the pH scale is based on the hydrogen ion concentration, it can also be used to determine alkalinity and also the hydroxide ion (OH^-) concentration. A pH of 7 represents a neutral solution. If equal concentrations of H^+ and OH^- ions were to react, the resultant solution should be neutral. At pH 7, the concentration of H^+ is $10^{-7}\,mol\,L^{-1}$ so the concentration of OH^- must also be $10^{-7}\,mol\,L^{-1}$.

KEY FORMULA

The mathematical relationship between pH and hydrogen ion concentration (Table 19.2.1) is readily apparent:

pH = negative (log to base 10 of hydrogen ion concentration)

that is, if $[H^+] = 10^{-x}$

then pH = x

TABLE 19.2.2 The relationship between pH and hydroxide ion concentration

pH	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
H^+ concentration ($mol\,L^{-1}$)	1 or 10^0	10^{-1} or 0.1	10^{-2}	10^{-3}	10^{-4}	10^{-5}	10^{-6}	10^{-7}	10^{-8}	10^{-9}	10^{-10}	10^{-11}	10^{-12}	10^{-13}	10^{-14}
OH^- concentration ($mol\,L^{-1}$)	10^{-14}	10^{-13}	10^{-12}	10^{-11}	10^{-10}	10^{-9}	10^{-8}	10^{-7}	10^{-6}	10^{-5}	10^{-4}	10^{-3}	10^{-2}	10^{-1}	1 or 10^0

As the H^+ ion concentration decreases, the OH^- ion concentration increases. Also note that the concentration superscripts always add up to 14.

Measuring pH

There are many natural and synthetic indicators used to indicate whether a solution is acidic or alkaline, as well as its degree of acidity. Synthetic indicators include methyl violet, methyl orange, litmus, bromothymol blue and phenolphthalein. Their colour changes are shown in Figure 19.2.1. The actual range of acidity or alkalinity over which an indicator changes colour varies from one indicator to another. Chemists have developed a universal indicator that produces several colour changes depending on the pH of the solution. It is a mixture of several indicators (Table 19.2.3).

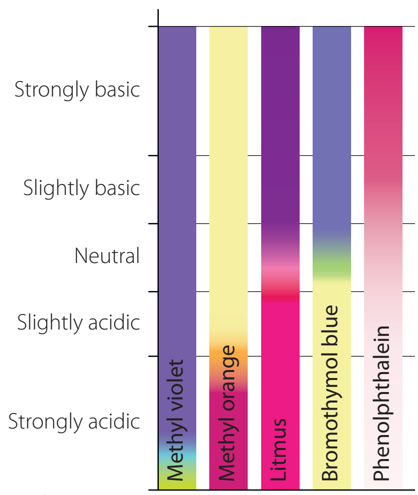


FIGURE 19.2.1 Colour changes in a range of indicators

TABLE 19.2.3 pH of some common substances

COLOUR OF UNIVERSAL INDICATOR	$[H_3O^+]$	pH	SUBSTANCE	CLASSIFICATION
	10	21	Concentrated hydrochloric acid	
	1	0	Car battery acid 1 mol L^{-1} hydrochloric acid	ACID
	10^{-1}	1	0.1 mol L^{-1} hydrochloric acid	
	10^{-2}	2	Stomach acid	
	10^{-3}	3	Vinegar Lemon juice	
	10^{-4}	4	Soft drinks Soda water	
	10^{-5}	5	Wine Black coffee	
	10^{-6}	6	Rain water Milk, saliva	
	10^{-7}	7	Pure water	NEUTRAL
	10^{-8}	8	Blood Sea water	ALKALINE
	10^{-9}	9	Bore water Baking soda solution	
	10^{-10}	10	Toilet soap	
	10^{-11}	11	Laundry detergents	
	10^{-12}	12	Household ammonia Dishwashing machine powders	
	10^{-13}	13	Chlorine bleach solutions	
	10^{-14}	14	Oven cleaners 1.0 mol L^{-1} sodium hydroxide	

SECTION REVIEW 19.2

REMEMBERING

1 Define:

a pH scale **b** indicator.

2 State the pH values for acidic substances, basic substances and neutral substances.

3 Use Figure 19.2.1 (page 361) to identify which of the following statements is correct.

a A substance that turns phenolphthalein pink is more acidic than one that turns methyl orange pink.

b A substance that turns methyl orange yellow is less acidic than one that turns methyl orange pink.

c A substance that turns phenolphthalein colourless is neutral.

d A substance that turns litmus red is more acidic than one that turns bromothymol blue a blue colour.

UNDERSTANDING

4 What is the relationship between the concentration of an acid, the concentration of a base and pH?

19.3 The Arrhenius model

The first modern definition of acids and bases was put forward by Svante Arrhenius, a Swedish scientist, in the late 19th century.

As suggested by Arrhenius: *an acid is a substance that ionises in water to produce H^+ ions.*

In other words, an acid increases the concentration of hydrogen ions in solution.

As suggested by Arrhenius: *a base is a substance that ionises in water to produce OH^- ions.*

In other words, a base is a substance that increases the concentration of hydroxyl ions in solution.

Acid strength

Not all compounds containing hydrogen are acids and not all hydrogens in an acid are released as H^+ ions.

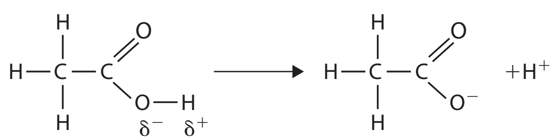
It is only the hydrogen atoms in very polar bonds that are released as H^+ ions. For example, the H—Cl bond is very polar due to the difference in electronegativity values of the hydrogen and chlorine atoms.

Due to the polar nature of this bond, it is easily ionisable and this explains why hydrochloric acid is considered a strong acid (Figure 19.3.1).

In a weak acid such as ethanoic (acetic) acid, there is only one ionisable hydrogen even though there are four hydrogen atoms in the molecule (Figure 19.3.2).

$$\overset{\delta+}{H}-\overset{\delta-}{Cl} \longrightarrow H^+(aq) + Cl^-(aq)$$

FIGURE 19.3.1 When hydrochloric acid dissociates, the polar H—Cl breaks, releasing the H^+ ion.



Ethanoic acid

Ethanoate ion

FIGURE 19.3.2 When ethanoic acid dissociates, only the polar O—H bond breaks, releasing the H^+ ion. The three H—C bonds are non-polar and these hydrogen atoms are not released.

With experience, it will be possible to recognise these ionisable hydrogens even in the most complex of molecules.

Base strength

A strong base, according to the Arrhenius theory, is one that dissociates completely in water to produce hydroxyl (OH^-) ions. An example of a strong base is sodium hydroxide (NaOH):

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

The Arrhenius definition of a weak base is one that does not dissociate completely in water to produce hydroxyl ions. An example of a weak base is ammonia (NH_3):

$$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$$

In this case, even though NH_3 does not contain OH^- ions, the NH_3 molecule reacts with water to produce OH^- ions, thereby conforming to the Arrhenius definition of a base.

Recognising strong and weak acids and bases

Strong acids

For the purposes of this unit, the strong acids are:

- hydrochloric acid (HCl)
- nitric acid (HNO_3)
- sulfuric acid (H_2SO_4).

Weak acids

Once the strong acids have been memorised, all other acids are to be considered weak! However, it is useful to be able to recognise a weak acid. As mentioned previously, a weak acid may have many hydrogen atoms but only one or, possibly, two ionisable hydrogen atoms. For example, potassium hydrogen phthalate ($C_8H_5KO_4$) or KHP for short, is an important acid used in quantitative analysis. Its structure may appear daunting at first but, on closer inspection, the ionisable hydrogen atom is easy to spot. It is the one attached to the oxygen atom (Figure 19.3.3).

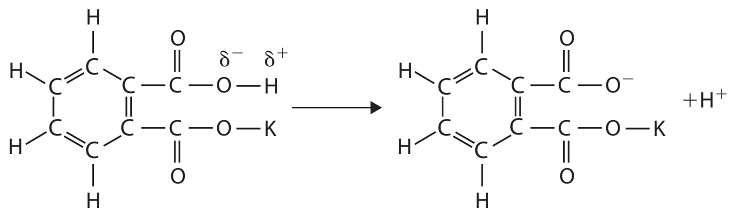


FIGURE 19.3.3 The ionisable hydrogen atom in potassium hydrogen phthalate (KHP) is the one attached to the oxygen atom. This O—H bond is very polar, making it relatively easy to break, thereby releasing the hydrogen ion.

Strong bases

Strong bases include the hydroxides of groups 1 and 2. For example, potassium oxide (K_2O) reacts in water to produce potassium (K^+) and hydroxyl (OH^-) ions:

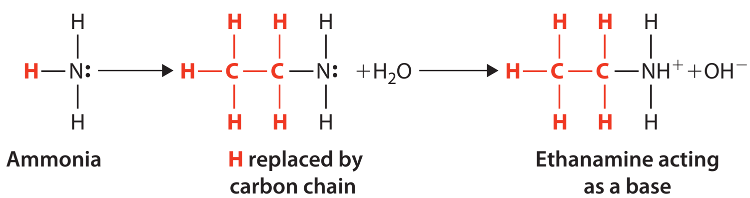
$$K_2O + H_2O \rightarrow 2K^+ + 2OH^-$$

Weak bases

Ammonia (NH_3) is a weak base and many weak bases are 'variations' of ammonia, where one or more of the hydrogen atoms has been replaced; for example, with a carbon chain. Ethanamine (ethylamine, C_2H_7N) is a weak base that is used extensively in the chemical industry in the manufacture of compounds, herbicides and pharmaceutical products (Figure 19.3.4).

concentrated solution
a large amount of solute dissolved in a given volume of solvent

FIGURE 19.3.4 Ethanamine (ethylamine) is a weak base and can be thought of as an ammonia molecule that has had one of its hydrogen atoms (in red) replaced by a carbon chain (CH_3CH_2–).



The difference between acid strength and concentration

The strength of an acid or base does not refer to its concentration. The terms **concentrated** or **dilute** are used to describe the concentration of an acid or base solution while the terms 'strong' and 'weak' refer to the degree of dissociation of an acid or base solution to produce either H^+ or OH^- ions.

dilute solution
a small amount of solute dissolved in a given volume of solvent

SECTION REVIEW 19.3

REMEMBERING

1 Define:

a Arrhenius acid

b Arrhenius base.

APPLYING

2 With the aid of balanced equations, show how propanoic (propionic) acid (CH_3CH_2COOH) behaves as a weak acid and how propanamine (propylamine, $CH_3CH_2CH_2NH_2$) behaves as a weak base.

3 Write a balanced equation for the reaction between propanoic acid and propanamine.

19.4 Mandatory practical

PRACTICAL ACTIVITY 19.4.1

Investigating the properties of strong and weak acids

In this activity you will examine the strength of acids in three different ways: indicators, pH meter and conductance.

AIM

To determine the relative strength of some strong and weak acids and to compare this with some common household substances

EQUIPMENT PER GROUP

- universal indicator solution and pH colour chart
- distilled water
- 2–3 mL of 0.1 mol L^{-1} solutions of hydrochloric acid (HCl), nitric acid (HNO_3), sulfuric acid (H_2SO_4), ethanoic (acetic) acid (CH_3COOH), citric acid ($C_6H_8O_7$) and ascorbic acid (vitamin C, $C_6H_8O_6$)
- various household substances such as vinegar, lemon juice, aspirin, swimming pool water
- test tubes and test-tube rack
- battery or power pack
- 1 small beaker (100 mL)
- electrical leads
- ammeter or multimeter
- carbon rod electrodes
- pH meter

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RISK ASSESSMENT

WHAT ARE THE RISKS IN DOING THIS EXPERIMENT?	HOW CAN YOU MANAGE RISKS TO STAY SAFE?
Dilute solutions of hydrochloric acid, sulfuric acid and nitric acid can cause irritation to skin and eyes.	Wear safety glasses and protective clothing. Do not allow contact with skin or clothing. If contact with skin occurs, wash with water.
Household chemicals can burn the skin.	Wear safety glasses and protective clothing. Take care when handling chemicals and tell your teacher when there is a spill.

Copy and complete the risk assessment table. Add any more risks you can think of, as well as ways to manage them.

PROCEDURE

PART A: TESTING THE SUBSTANCES WITH THE INDICATOR

1 Set up the test tubes in the rack and label each one. Place a few millilitres of each acid into its corresponding test tube. If the acid is a solid (such as ascorbic) add a little distilled water.
2 Add a drop of universal indicator to each acid and record the colour change.
3 Compare the colour of each test tube with the pH colour chart and try to assign a pH value for each acid.
4 Record the pH value.

PART B: TESTING THE SUBSTANCES WITH A pH METER

1 Empty the test tubes from the previous experiment and rinse them with distilled water.
2 Place a few millilitres of each acid into its corresponding test tube.
3 Place the tip of the pH meter into each acid (remembering to rinse the meter between each test) and record the reading.

PART C: CONDUCTANCE TEST OF EACH ACID

1 Set up the apparatus as shown in Figure 19.4.1.
2 Half-fill the beaker with the acid to be tested.
3 Switch on the power pack and set to 4 V, or connect the battery.
4 Record the reading for the electrical current.

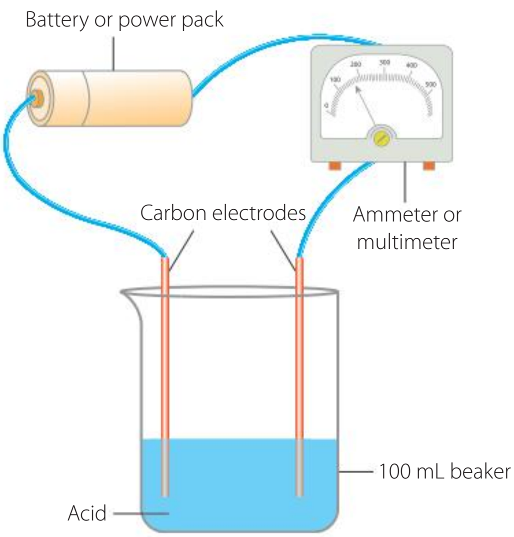

FIGURE 19.4.1 Experimental set-up for performing conductance experiments

RESULTS

Draw up a data table to record the colour, pH value from the pH colour chart, pH reading from the pH meter and the current value for each acid.

DISCUSSION

1 Identify any significant differences between the pH values obtained from the universal indicator and the pH meter. Suggest possible reasons for these differences.
2 Is there a trend in the current readings for the various acids? If so, what is it?
3 Explain why the conductance experiment was carried out by using your understanding of the Arrhenius theory of acids.
4 Explain your results of the conductance experiment using the Arrhenius theory.

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CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 Define the following terms.

- **a** Strong acid
- **b** Weak acid
- **c** Strong base
- **d** Weak base

CATEGORY QUESTIONS

2 Consider the pH ranges 1–3, 4–6, 8–10 and 11–13. Which pH range would apply to each of the following solutions?

- **a** Concentrated hydrochloric acid
- **b** Distilled water
- **c** Dilute ammonia solution
- **d** Oven cleaner

EVIDENCE QUESTIONS

3 You find an old first aid book that suggests that a blue bottle sting should be treated with vinegar. Provide evidence to support or refute this advice, as well as stating the current recommendations.

END-OF-CHAPTER EXAM

End-of-chapter test

1 According to the Arrhenius theory of acid and bases, a base is a substance that:

A tastes bitter.

B can donate a pair of electrons.

C has a pH above 7.

D ionises in water to produce hydroxide ions.

2 Which indicator in the following table could be used to distinguish between orange juice (pH = 3.5) and red wine vinegar (pH = 2.2)?

	INDICATOR	pH RANGE AND COLOUR CHANGE
A	Methyl orange	3.2–4.4 red → yellow
B	Bromocresol green	3.8–5.4 yellow → blue
C	Thymol blue	1.2–2.8 red → yellow
D	Cresol red	7.0–8.8 yellow → red

3 Two bases are compared:

- Base **i** is 0.1 mol L^{-1} sodium hydroxide
- Base **ii** is 1.0 mol L^{-1} ammonia

How does **i** compare with **ii**?

A Base **i** is weaker and more concentrated than base **ii**.

B Base **ii** is stronger and more concentrated than base **i**.

C Base **i** is stronger and more dilute than base **ii**.

D Base **i** is stronger and more concentrated than base **ii**.

4 If a solution has a pH of 2, how many times more acidic is it than a solution with a pH of 4?

5 If a solution has a pH of 11, how many times less acidic is it than a solution with a pH of 8?

6 In a weak acid such as ethanoic (acetic) acid (CH_3COOH), what kind of bond is present that enables it to release hydrogen ions?

7 What is superior about Arrhenius's theory compared with previous theories?

8 One major problem with Arrhenius's theory is that, if equal amounts of acid and base react, the resulting solution should be neutral with a pH = 7. Consider the reaction between equal amounts of hydrochloric acid and ammonia:

$$HCl(aq) + NH_3(aq) \rightarrow NH_4Cl(aq)$$

The pH of the resulting solution is about pH 4. How would you explain this?

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20 REACTIONS OF ACIDS AND BASES

Introduction

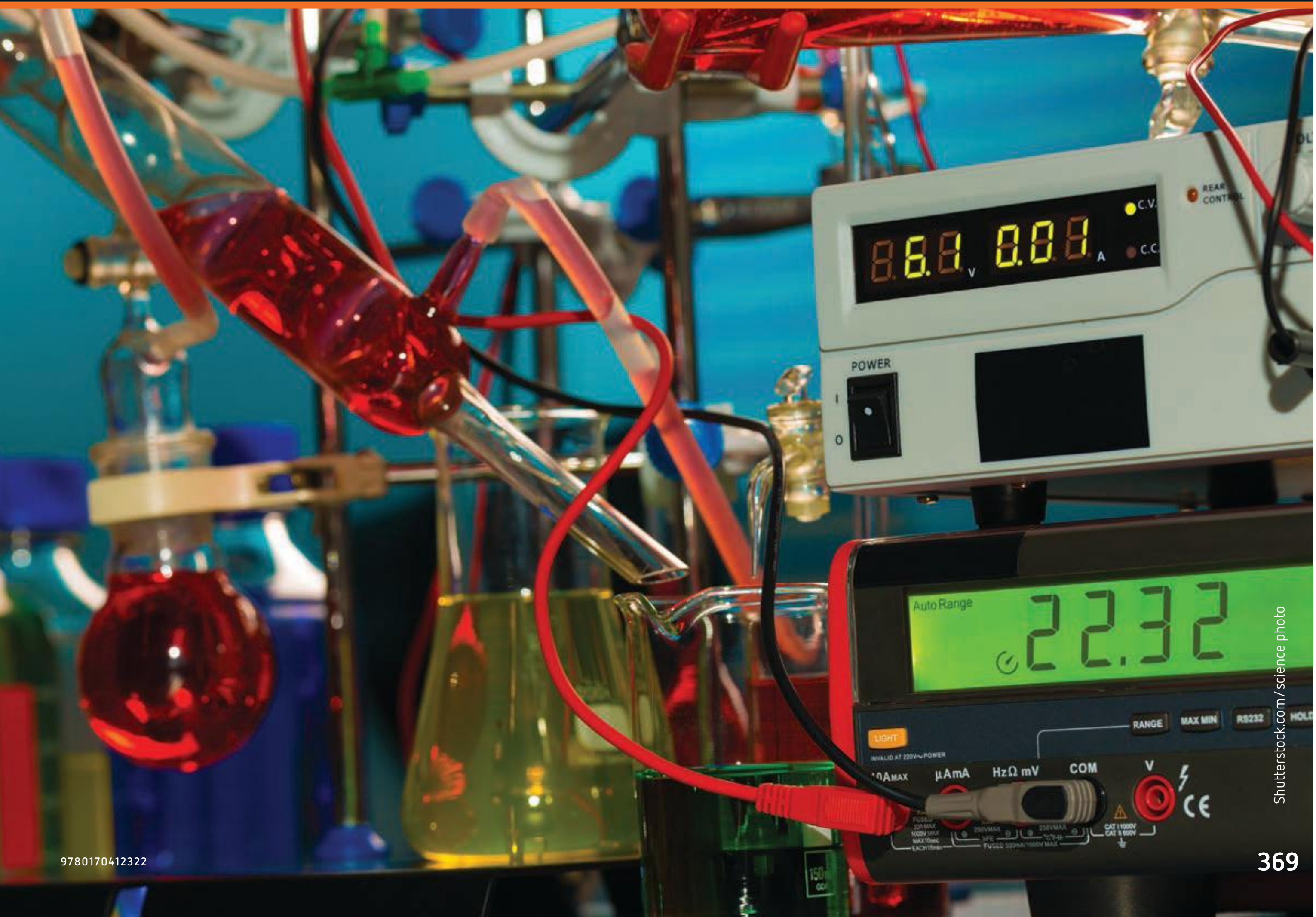
Acids and bases play an important part in our lives. Virtually all chemical reactions can be classified as acid–base reactions. Acid–base reactions follow patterns that scientists use to predict the products easily. Acids and bases affect our health, playing a crucial role in the chemical reactions in our blood. On a larger scale, an understanding of acid and base reactions is important in the regulation of our water supply.

The negative environmental impact of production of gases that cause acid rain associated with industrialisation on a global scale has forced scientists to investigate the chemistry of acid rain and how this problem can be solved.

Stimulus questions

What are examples of neutralisation reactions in daily life?

How can naturally occurring indicators be used to determine whether a substance at home is acidic or basic?



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20.1 Reactions of acids

20.1.1 Reactions of acids

Acids undergo several different types of reactions. With acid–base reactions, once the pattern of the reaction is understood, scientists can predict the products for most reactions of these types.

Reactions between acids and bases

Acids and bases react to form compounds called 'salts' and water. The name given to this reaction is **neutralisation**. If the right amounts of acid and base react, the resulting solution will be **neutral** (neither acidic nor basic). This generalised neutralisation reaction can be used to predict the products of most reactions between an acid and a base. The formula for the neutralisation reaction is:

neutralisation the reaction between an acid and a base to form a salt and water

neutral neither acidic nor basic – a pH of 7

KEY FORMULA

$$\text{Acid} + \text{base} \rightarrow \text{salt} + \text{water}$$

A typical reaction between acids and bases is the reaction between an acid and a hydroxide:

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

$$\text{Hydrochloric acid} + \text{sodium hydroxide} \rightarrow \text{sodium chloride} + \text{water}$$

Another typical reaction occurs between a basic or an amphoteric oxide and an acid:

KEY FORMULA

$$\text{Acid} + \text{oxide} \rightarrow \text{salt} + \text{water}$$

For example:

$$2HCl(aq) + MgO(s) \rightarrow MgCl_2(aq) + H_2O(l)$$

$$\text{Hydrochloric acid} + \text{magnesium oxide} \rightarrow \text{magnesium chloride} + \text{water}$$

Another common, but not typical, acid–base reaction is the reaction between an acid and ammonia. This reaction does not adhere to the generalised neutralisation reaction because the hydrogen ion attaches to the ammonia molecule (not H_2O) to produce the ammonium ion $\left(NH_4^+(aq)\right)$. For example:

$$HCl(aq) + NH_3(aq) \rightarrow NH_4Cl(aq)$$

$$\text{Hydrochloric acid} + \text{ammonia} \rightarrow \text{ammonium chloride}$$

However, when ammonia reacts with water, it produces hydroxide ions (OH^-) and is considered a weak base:

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

$$\text{Ammonia} + \text{water} \rightarrow \text{ammonium ion} + \text{hydroxide ion}$$

The ammonium ion is considered a weak acid. This means it will react with a base to produce a salt and water. However, ammonia gas will also be produced in the reaction as seen in the following example:

$$NaOH(aq) + NH_4Cl(aq) \rightarrow NaCl(aq) + NH_3(g) + H_2O(l)$$

$$\text{Sodium hydroxide} + \text{ammonium chloride} \rightarrow \text{sodium chloride} + \text{ammonia} + \text{water}$$

Ammonium salts such as ammonium sulfate or ammonium nitrate are commonly used as nitrogen fertilisers. Alkaline compounds (a soluble base) such as lime (CaO) are often used to prevent soils from becoming too acid. One reason why a farmer would not treat a field with an ammonium compound at the same time as using an alkali is that the ammonium salt will react with the base to give off ammonia. The ammonia will escape into the atmosphere as a gas, and not be available to fertilise the plants.

INQUIRING FURTHER

FIGURE 20.1.1 Antacids contain a base, which neutralises the excess acid in the stomach.

Neutralisation reactions are important in our bodies and in our lives. If too much hydrochloric acid is produced in our stomachs, then we could end up with heartburn or indigestion. This problem is commonly fixed by taking an antacid, which contains a base, such as magnesium hydroxide or aluminium hydroxide. The added base will neutralise the excess stomach acid.

The cost of antacid products varies substantially, with each brand claiming to be the most effective. Conduct research to answer the following questions.

1 Which substance in the stomach causes indigestion?
2 What is the active ingredient in the antacid product?
3 Is the amount of active ingredient related to the cost of the product?

Devise a chemical test to verify the manufacturers' claims concerning the amount of active ingredient in the antacid product. Compare the cost of the antacid product to the amount of its active ingredient.

Reactions between acids and metals

Acids react with metals. When rain becomes acidic because of pollutants in the air, it reacts with metals, causing them to corrode more quickly. Not all metals are reactive, but for those that are, the reaction between an acid and a metal can be represented by the general equation:

KEY FORMULA

Acid + metal → a salt + hydrogen gas

This reaction could also be called a corrosion reaction because the metal is 'eaten' away.

The following reactions are examples of reactions between an acid and magnesium metal:

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

Hydrochloric acid + magnesium → magnesium chloride + hydrogen

$$H_2SO_4(aq) + Mg(s) \rightarrow MgSO_4(aq) + H_2(g)$$

Sulfuric acid + magnesium → magnesium sulfate + hydrogen

Reactions between acids and carbonate or hydrogen carbonate

Food products such as bread and sponge cakes have a honeycomb structure, which contains bubbles. These bubbles are formed by carbon dioxide (CO_2) gas causing the mixture to rise during cooking. Carbon dioxide is formed by adding a carbonate (CO_3^{2-}) or hydrogen carbonate (HCO_3^-) and an acid to the cooking mixture.

When an acid and carbonate react, they produce a salt, carbon dioxide and water. The reaction between an acid and a carbonate can be represented by the general equation:

KEY FORMULA

Acid + carbonate → salt + carbon dioxide + water

The following two reactions are examples of reactions between an acid and calcium carbonate:

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

Hydrochloric acid + calcium carbonate → calcium chloride + water + carbon dioxide

$$H_2SO_4(aq) + CaCO_3(s) \rightarrow CaSO_4(aq) + H_2O(l) + CO_2(g)$$

Sulfuric acid + calcium carbonate → calcium sulfate + water + carbon dioxide

When an acid and a hydrogen carbonate react, they will also produce a salt, water and carbon dioxide according to the general equation:

KEY FORMULA

Acid + hydrogen carbonate → salt + carbon dioxide + water

The following reactions are examples of reactions between an acid and sodium hydrogen carbonate:

$$HCl(aq) + NaHCO_3(s) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g)$$

Hydrochloric acid + sodium hydrogen carbonate → sodium chloride + water + carbon dioxide

$$H_2SO_4(aq) + 2NaHCO_3(s) \rightarrow Na_2SO_4(aq) + 2H_2O(l) + 2CO_2(g)$$

Sulfuric acid + sodium hydrogen carbonate → sodium sulfate + water + carbon dioxide

In cooking, a relatively safe weak acid, such as acetic acid (the acidic component in vinegar) or citric acid (the acid in citrus fruits), is used, together with sodium bicarbonate, to release CO_2 during baking.

Acid rain

20.1.2 Acid rain

Acid rain is rain that has become acidic because gases from the atmosphere have dissolved in it. Approximately 70% of acid rain comes from sulfur dioxide (SO_2), which dissolves into the water in the atmosphere to form sulfuric acid. The remaining 30% of acid rain comes from various oxides of nitrogen, mainly nitrogen dioxide (NO_2) and nitrogen trioxide (NO_3), collectively called 'NO_x'.

Although the effect of acid rain on forests, lakes and rivers has been well publicised, the effects on people have not been as well documented. Many toxic metals are held in the ground in compounds. However, acid rain can break down some of these compounds, freeing the metals and washing them into water sources such as rivers. In Sweden, nearly 10 000 lakes now have such high mercury concentrations that people are advised not to eat fish caught in them.

FIGURE 20.1.2 Dead spruce trees killed by acid rain

Acid rain occurs naturally as a result of carbon dioxide dissolved in rainwater and from volcanic emissions of sulfur oxide gases. Carbon dioxide dissolves in rainwater to produce carbonic acid $(H_2CO_3(aq))$:

$$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$$

The carbonic acid then dissociates to produce hydrogen ions and hydrogen carbonate ions in solution:

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

This naturally occurring process accounts for only a tiny fraction of the acidity of rainwater – the vast amount of acid rain is caused by the production of carbon dioxide from fossil fuel combustion.

A similar process accounts for sulfuric acid in rainwater. Some is the result of natural processes, such as volcanic activity and biological decay, but most sulfuric acid is produced by human activity such as the combustion of sulfur-containing fossil fuels in power stations.

Sulfur in fossil fuels combusts, producing sulfur dioxide gas $(SO_2(g))$:

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

Sulfur dioxide reacts with oxygen in the atmosphere to produce sulfur trioxide $(SO_3(g))$:

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

Sulfur trioxide dissolves in water vapour in the atmosphere:

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

Environmental effects of acid rain

Acid rain is involved in several inorganic and biochemical reactions with serious environmental effects, making this a growing environmental problem worldwide.

- The pH of some aquatic systems such as lakes has dropped dramatically, to a level where they cannot support most marine life.
- Acid rain causes the breakdown of many minerals that contain toxic metal ions such as mercury and aluminium. These toxic metal ions make their way into the water supply, causing serious problems for domestic water consumption.
- Vital minerals such as calcium ions are lost from the soil, killing trees and damaging crops.
- Pollutants such as sulfur dioxide and the nitrogen oxide gases are easily moved around by wind currents and can cause problems far from their point of origin.
- Many stone monuments are made of limestone or marble, materials that contain calcium carbonate. These materials are highly susceptible to acid rain, resulting in the erosion of many of the world's historic monuments.

Science Source/Monica Schroeder

Science Source/Monica Schroeder

FIGURE 20.1.3 Over time, acid rain can damage buildings and statues

Acid rain prevention methods

The recognition that acid rain is a serious global problem has led to initiatives to reduce the emission of gases that cause acid rain. These initiatives can be grouped under three general themes: technical solutions, international treaties and emissions trading.

Technical solutions

Many coal-fired power stations use 'flue gas desulfurisation' (FGD) to remove sulfur-containing gases from their exhaust gases. A typical power station can remove up to 95% or more of the sulfur dioxide from its exhaust gases.

A commonly used method of FGD is the *wet scrubber*, which consists of a tower into which the flue (exhaust) gas is mixed with a limestone slurry (water mixed with calcium carbonate). The sulfur dioxide in the flue gas reacts with calcium carbonate to produce calcium sulfite ($CaSO_3$), according to the following equation:

$$CaCO_3(s) + SO_2(g) \rightarrow CaSO_3(s) + CO_2(g)$$

The calcium sulfite is then reacted with oxygen at around 100°C to produce gypsum $\left((CaSO_4 \cdot 2H_2O(s)\right)$, which can be sold as an industrially important material.

Another increasingly important technical solution to cater for acid rain formed as a result of oxides of nitrogen (such as NO_2) dissolving in rain and yielding nitric acid (HNO_3) is to fit exhaust gas recirculation devices to cars. These reduce the emission of nitrogen oxide gases from car exhausts.

International treaties

Several international treaties have been agreed to by many of the world's industrialised countries. These treaties seek, among other things, to curb the long-range transportation of atmospheric pollutants.

Emissions trading

In emissions trading, every polluting facility such as a power station or factory is given, or can purchase, an emission allowance for every 'unit' of pollution it produces. Each facility can install pollution-controlling equipment and then sell their part of their emission allowance they do not need. This would help recover some of the cost of installing the pollution-controlling equipment.

SCIENCE AS A HUMAN ENDEAVOUR

BLOOD CHEMISTRY

Blood is a complex **suspension** of plasma (fluid component) and cells. It is more viscous than water due to the dissolved salts and the biomolecules and cells being transported through the body. Polar substances can directly dissolve into the plasma. Non-polar substances, such as many hormones, have complex carrier proteins to transport them along the blood vessels. Blood is responsible for the delivery of oxygen and nutrients to all parts of the body, temperature regulation and waste removal.

Another role of blood is to help control the pH of the body. Despite all the chemicals and reactions occurring in the body, the pH of blood has to be maintained at a pH of close to 7.4. There are three chemical buffer systems for this: the carbonate/carbonic acid buffer, the phosphate buffer and the buffering of plasma proteins. The carbonate (CO_3^{2-})/carbonic acid (H_2CO_3) buffer is the most important.

Research the pH of the blood of a normal person and why it is important to maintain this value. Research the role of a buffer system.

suspension
a heterogenous mixture where the solid particles do not dissolve in the liquid medium

SCIENCE AS A HUMAN ENDEAVOUR

WATER QUALITY

Australia is a dry continent, so the supply of water is an important issue. In a drought region, sources of water, such as bore water, may be investigated. In other areas after a flood, the quality of the water can be a cause for concern. Clean sources of water are crucial to health, but not all water sources are pure. Scientists use different methods to analyse water depending on what they are looking for; for example, salt, organic compounds or gases. One proposed solution to the problem of drought and water supply is desalination.

Research desalination and the claims made by those in favour and against this technique as a proposed solution to drought in Australia. Based on the research, make a recommendation on the use of desalination in your local area.

SECTION REVIEW 20.1

REMEMBERING

1 Define:
 a acid
 b base
 c salt
 d neutralisation
 e amphoteric oxide.

UNDERSTANDING

2 Complete and balance the following equations.
 a $HNO_3(aq) + MgCO_3(s) \rightarrow$
 b $H_3PO_4(aq) + KOH(aq) \rightarrow$
 c $CH_3COOH(aq) + Mg(s) \rightarrow$
 d $H_2SO_4(aq) + CH_3NH_2(aq) \rightarrow$

3 Explain why you would not react an acid with sodium or potassium.

APPLYING

4 Write the dissociation reaction for the following acids in water.
 a Nitric acid (HNO_3)
 b Hydrochloric acid (HCl)
 c Phosphoric acid (H_3PO_4)

5 Why is acid rain regarded as a relatively recent problem when naturally occurring acid rain has been falling for billions of years?

20.2 Inquiry skills: constructing and using representations

Chemical reactions may be represented in several ways. This is especially true of reactions of acids and bases. It is important to distinguish between the different types of equations that chemists use.

SYLLABUS LINK
Chapter 8 revises representing chemical reactions.

Types of equations

The neutral species equation is the overall equation for the reaction. It shows all the reactants and products as neutral compounds. The complete ionic equation shows all the ions present in solution. When deriving the complete ionic equation from the neutral species equation, only the species with the (aq) state symbol are split into their ions. Some ions represented in the complete ionic equation do not take part in the reaction and are called **spectator ions**. The net equation shows only the reacting ions.

The various equation types are shown for the reactions between a strong acid and a base (Worked example 20.2.1), a strong acid with a weak base (Worked example 20.2.2), a metal and an acid (Worked example 20.2.3), an acid and a carbonate (Worked example 20.2.4), and an acid and a hydrogen carbonate (Worked example 20.2.5).

spectator ion
ions that remain in solution before and after a chemical reaction has taken place, which are not involved in the reaction

WORKED EXAMPLE 20.2.1

Write a neutral species equation, a complete ionic equation and a net ionic equation for the reaction between hydrochloric acid and sodium hydroxide (a strong acid and a strong base).

ANSWER

1 Neutral species equation:

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

Hydrochloric acid + sodium hydroxide $\rightarrow$ sodium chloride + water

2 Complete ionic equation:

$$H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(l)$$

Hydrogen ion + chloride ion + sodium ion + hydroxide ion → sodium ion + chloride ion + water

3 Net ionic equation:

The $Na^+(aq)$ and $Cl^-(aq)$ ions in the complete ionic equation are spectator ions.

This means that the reaction between an acid and a base is simply the reaction between the $H^+(aq)$ ions and $OH^-(aq)$ ions:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Hydrogen ion + hydroxide ion → water

WORKED EXAMPLE 20.2.2

Write a neutral species equation, a complete ionic equation and a net ionic equation for the reaction between sulfuric acid and ammonia (a strong acid and a weak base).

ANSWER

1 Neutral species equation:

$$H_2SO_4(aq) + 2NH_3(aq) \rightarrow (NH_4)_2SO_4(aq)$$

Sulfuric acid + ammonia → ammonium sulfate

2 Complete ionic equation:

$$2H^+(aq) + \cancel{SO_4}^{2-}(aq) + 2NH_3(aq) \rightarrow 2NH_4^+(aq) + \cancel{SO_4}^{2-}(aq)$$

Hydrogen ion + sulfate ion + ammonia → ammonium ion + sulfate ion

3 Net ionic equation:

$$2H^+(aq) + 2NH_3(aq) \rightarrow 2NH_4^+(aq)$$

Divide through by 2 to give:

$$H^+(aq) + NH_3(aq) \rightarrow NH_4^+(aq)$$

This method can be applied to all the acid reactions.

WORKED EXAMPLE 20.2.3

Write a neutral species equation, a complete ionic equation and a net ionic equation for the reaction between magnesium and nitric acid (a metal and an acid).

ANSWER

1 Neutral species equation:

$$Mg(s) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2(g)$$

2 Complete ionic equation:

$$Mg(s) + 2H^+(aq) + \cancel{2NO_3}^-(aq) \rightarrow Mg^{2+}(aq) + \cancel{2NO_3}^-(aq) + H_2(g)$$

Note that the solid magnesium, Mg(s), and the aqueous magnesium ion, $Mg^{2+}(aq)$, are not the same so they do not cancel out.

3 Net ionic equation:

$$Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$$

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WORKED EXAMPLE 20.2.4

Write a neutral species equation, a complete ionic equation and a net ionic equation for the reaction between hydrochloric acid and calcium carbonate (an acid and a carbonate).

ANSWER

1 Neutral species equation:

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

2 Complete ionic equation:

$$2H^+(aq) + \cancel{2Cl^-}(aq) + CaCO_3(s) \rightarrow Ca^{2+}(aq) + \cancel{2Cl^-}(aq) + H_2O(l) + CO_2(g)$$

3 Net ionic equation:

$$2H^+(aq) + CaCO_3(s) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$$

WORKED EXAMPLE 20.2.5

Write a neutral species equation, a complete ionic equation and a net ionic equation for the reaction between sulfuric acid and sodium hydrogen carbonate (an acid and a hydrogen carbonate).

ANSWER

1 Neutral species equation:

$$H_2SO_4(aq) + 2NaHCO_3(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l) + 2CO_2(g)$$

2 Complete ionic equation:

$$2H^+(aq) + \cancel{SO_4{}^{2-}}(aq) + \cancel{2Na^+}(aq) + 2HCO_3{}^-(aq) \rightarrow \cancel{2Na^+}(aq) + \cancel{SO_4{}^{2-}}(aq) + 2H_2O(l) + 2CO_2(g)$$

3 Net ionic equation:

$$2H^+(aq) + 2HCO_3{}^-(aq) \rightarrow 2H_2O(l) + 2CO_2(g)$$

Divide through by 2 to give:

$$H^+(aq) + HCO_3{}^-(aq) \rightarrow H_2O(l) + CO_2(g)$$

SECTION REVIEW 20.2

UNDERSTANDING

1 Write neutral species, complete ionic and net ionic equations for the following reactions.

- **a** Nitric acid solution and solid magnesium oxide
- **b** Sulfuric acid solution and aluminium hydrogen carbonate solution
- **c** Phosphoric acid solution and magnesium powder
- **d** Ethanamine (ethylamine) solution and sulfuric acid solution

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 Define the following terms.
   a Neutral species equation
   b Complete ionic equation
   c Net ionic equation
   d Spectator ion
   e Amphoteric oxide

CATEGORY QUESTIONS

2 List the three categories of acid reactions.

ELABORATION QUESTIONS

3 Identify the main industrial pollutants responsible for acid rain.

EVIDENCE QUESTIONS

4 Acid snow is becoming an increasingly serious problem in northern Europe, the north-eastern United States and, especially, the Scandinavian countries. Explain why acid snow could pose an even greater threat to the environment than acid rain.

5 Use the following table of different indicators showing colour changes over different pH ranges to answer the questions that follow.

INDICATOR	COLOUR CHANGE				
	HIGHLY ACIDIC	SLIGHTLY ACIDIC	NEUTRAL	SLIGHTLY ALKALINE	HIGHLY ALKALINE
Methyl orange	Red	Yellow	Yellow	Yellow	Yellow
Bromothymol blue	Yellow	Yellow	–	Blue	Blue
Litmus	Red	Red	Purple	Blue	Blue
Phenolphthalein	Colourless	Colourless	Colourless	Colourless	Red

   a You have two solutions. Solution A is red in methyl orange, while solution B is red in phenolphthalein. Determine which is the more acidic.
   b Four different solutions were tested with different indicators. Determine which of the solutions is neutral, using the table.
      i Colourless in phenolphthalein
      ii Red in litmus
      iii Yellow in methyl orange
      iv Blue in bromothymol blue

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END-OF-CHAPTER EXAM

End-of-chapter test

1 When acid reacts with a metal, it produces:

A salt + carbon dioxide + water

B salt + water

C salt

D salt + hydrogen.

2 In the following reaction, what salt would be formed?

$$H_3PO_4(aq) + Na_2CO_3(aq) \rightarrow$$

A $H_2CO_3(aq)$

B $Na_2PO_4(aq)$

C $Na_3PO_4(aq)$

D $NaOH(aq)$

3 Identify the term for the reaction between an acid and a base.

4 In the reaction between sulfuric acid and ammonia, what are the products?

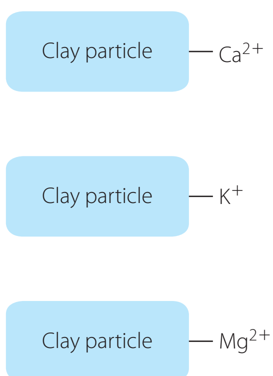
5 Identify the acid that needs to be reacted with potassium carbonate to produce potassium nitrate.

6 With the aid of neutral species and complete ionic equations, write the net ionic equation for the reaction between calcium hydrogen carbonate powder $\left(CaHCO_3(s)\right)$ and nitric acid $\left(HNO_3(aq)\right)$.

7 Explain why the reaction between an acid such as sulfuric acid $\left((H_2SO_4(aq)\right)$ and ammonia solution $\left(NH_3(aq)\right)$ does not produce water.

8 The effects of acid rain on soils is becoming an increasingly serious problem. One of the worst effects is that of mineral leaching, in which metal ions such as K^+, Ca^{2+} and Mg^{2+}, which are vital for the production of healthy crops, are removed.

Below is a diagram of how these minerals are present in soil:



The clay particles consist of a number of complex substances such as $Al_2Si_2O_5(OH)_4$ which are bound to the metal ions as shown in the diagram.

Suggest a way that an acid such as H_2SO_4 could cause the leaching of these metal ions from the soil.

» UNIT TWO

MOLECULAR INTERACTIONS AND REACTIONS

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Topic 3: Rates of chemical reactions

The rate of a chemical reaction is controlled by a number of key factors. This topic explores the relationship between reaction rates and reactant concentrations, temperature, pressure and surface area. The rate of reaction can be quantified by measuring the rate of formation of products or the rate of depletion of reactants. Prediction of the rate of reaction has become possible through the use of collision theory. The theory takes into account the distribution of energies in reacting atoms and molecules and identifies a minimum energy, the activation energy, required for a chemical reaction to occur. In this topic, energy profile diagrams will be employed to represent enthalpy changes and activation energies associated with chemical reactions. The mechanism by which catalysts, including enzymes and nanoparticles, control reaction rates will be investigated.

SCIENCE AS A HUMAN ENDEAVOUR

Students should be given opportunities to investigate how knowledge of enzymes enables scientists to investigate the ways catalysts work and help them to predict activation energies; how a knowledge of chemical reactions involving metals in an environment of moist air provides economic benefits through the development of technology to reduce corrosion; and how knowledge of collision theory enables scientists to predict a number of reaction rates under different contexts.

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21 RATES OF REACTIONS

Introduction

Just as some people run faster than others, some chemical reactions occur than others. As scientific understanding of what happens during a reaction has deepened, scientists have applied this knowledge to control the speed, or rate, of a reaction to suit our purposes.

Stimulus questions

- How can scientists measure the rate of a reaction?
- How can scientists maximise the rate at which they manufacture chemicals?

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21.1 The rate of a chemical reaction

FIGURE 21.1.1 These animals are travelling at different rates. How can you tell?

Rate of a reaction

Rate is how quickly one quantity changes compared with another quantity. The rate at which a car moves is how many kilometres the car travels in an hour. The rate at which someone loses weight might be the weight loss (often in kilograms) per week. Another common rate is rate of pay, which describes how much someone is paid per time period. In terms of reactions, the rate is how quickly a reaction occurs (that is, how the quantity or **concentration** of one or more reactants or products changes with time).

rate
how quickly one quantity (a dependent quantity) changes compared to another (independent quantity)

concentration
the number of particles in a given volume

FIGURE 21.1.2 Two beakers with the same reactants. Which one is reacting at the faster rate? How do you know?

FIGURE 21.1.3 Which sugar would dissolve the fastest? How could you measure it accurately to compare the three?

Measuring the rate of a reaction

To calculate any rate, divide the change in one quantity by the change in the other quantity.

KEY FORMULA

$$\text{Rate} = \frac{\text{change in quantity 1}}{\text{change in quantity 2}}$$

For example, at the 2012 London Olympic Games, Usain Bolt won the 200-metre sprint in 19.26 seconds. This means his average rate of running, or speed, was:

$$\text{Rate} = \frac{200}{19.26}$$
$$= 10.4\ \text{ms}^{-1}$$

To measure the rate of a reaction, you need something measurable to change, such as:

- mass
- colour
- volume
- pH
- concentration.

Limestone, a type of rock, is composed mainly of calcium carbonate ($CaCO_3$) and dissolves in acid. Limestone can be used to measure the rate or reaction between solid calcium carbonate ($CaCO_3(s)$) and a solution of hydrochloric acid ($HCl(aq)$). The mass of the solid will decrease as it reacts and forms the products calcium chloride in solution ($CaCl_2(aq)$), carbon dioxide gas ($CO_2(g)$) and liquid water ($H_2O(l)$), according to the equation:

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

To determine the rate of this reaction, the mass of the solid at the start of the reaction and at end of the reaction, and the time it takes for the reaction to occur, can be measured. You can then calculate the rate of the reaction. For example, if 2.00 g of calcium carbonate takes 48 s to completely react, the average rate of reaction over that time interval would be:

$$\begin{aligned}\text{Rate} &= \frac{\text{change in mass}}{\text{change in time}}\\ &= \frac{2.00}{48}\\ &= 4.17 \times 10^{-2}\ \text{g s}^{-1}\end{aligned}$$

However, a rate does not have to be calculated for the overall reaction. You can also calculate the rate at various intervals of the reaction by taking measurements throughout the reaction time. These rates may then be used to calculate the rate directly, or you can graph the changes that occur during the reaction.

Consider the reaction between magnesium solid and hydrochloric acid in solution:

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

The mass of the hydrogen gas produced during the reaction can be calculated by the mass lost from a flask (Figure 21.1.4 and Table 21.1.1).

Once the changes that occur during the reaction are graphed, such as the mass of hydrogen gas produced, the slope or **gradient** represents the rate of a reaction because it shows how quickly the reactants or products change. The graph in Figure 21.1.5 shows that rate of the reaction was the greatest at the start of the reaction and gradually slowed to zero at the end of the reaction. The raw data in Table 21.1.1 supports this, showing that the changes in mass are greatest at the start of the reaction.

gradient
the steepness of the line on a graph, which can be used as a measure of the rate of change of the quantity measured on the *y*-axis (the dependent variable) relative to that on the *x*-axis (the independent variable)

TABLE 21.1.1 Changes in mass during the reaction between magnesium and hydrochloric acid

TIME (s)	MASS OF FLASK AND CONTENTS (g)	MASS OF HYDROGEN GAS RELEASED (g)
0	20.00	0.00
10	18.60	1.40
20	18.30	1.70
30	18.20	1.80
40	18.15	1.85
50	18.10	1.90
60	18.06	1.94
70	18.04	1.96
80	18.02	1.98
90	18.01	1.99
100	18.00	2.00
110	18.00	2.00
120	18.00	2.00

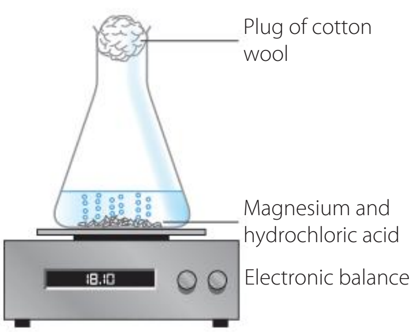

FIGURE 21.1.4 Measuring the mass of hydrogen gas lost during the reaction

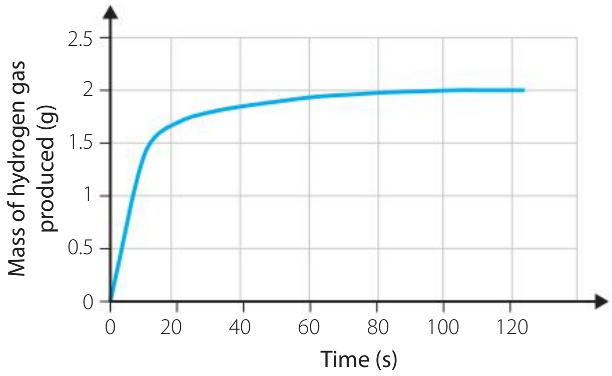

FIGURE 21.1.5 Mass of hydrogen gas produced

SECTION REVIEW 21.1

REMEMBERING

1 Define 'rate'.

2 What data would you need to record to calculate or compare the rates of different reactions?

3 During a reaction, when is the rate the fastest?

UNDERSTANDING

4 Outline how to calculate the average rate of a reaction over a given time interval.

5 Explain how the slope (or gradient) of a graph indicates the rate.

APPLYING

6 During a reaction, 20 g of product forms in 4 minutes. Calculate the reaction rate.

7 Copy the graph in Figure 21.1.6 and draw tangents to the curve at times 0, 1 and 5 to calculate the rate of reactions at each of these times.

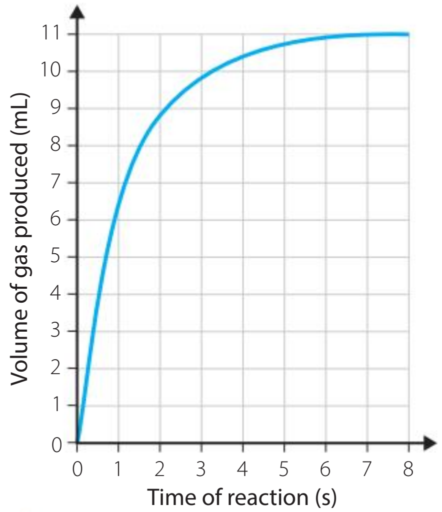

FIGURE 21.1.6 Production of carbon dioxide gas during the decomposition of copper(II) carbonate.

21.2 What is needed for a chemical reaction to occur?

21.2.1 Chemical reactions

To understand why some reactions occur faster than others, it is important to understand what happens to the reacting species during a reaction.

PRACTICAL ACTIVITY 21.2.1

Modelling a reaction

AIM

To use molecular models to represent a chemical reaction

EQUIPMENT PER GROUP

- Molecular model kit; for example, molymod®

PROCEDURE

1. Using the kit, make two molecules of hydrogen gas (H_2) and one molecule of oxygen gas (O_2) as shown in Figure 21.2.1.
2. Use these molecules to make two molecules of H_2O.
3. Describe what had to happen for the reactants (H_2 and O_2) to become the product (H_2O).

QUESTION

Based on the reaction you modelled, what has to happen for the reactants in a chemical reaction to become products?

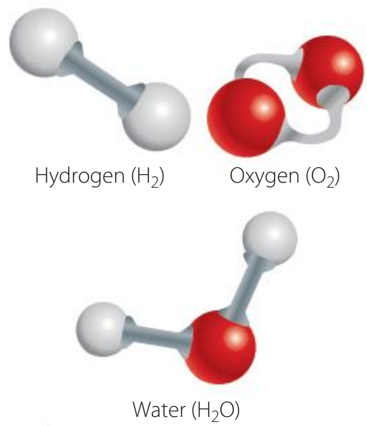

FIGURE 21.2.1 Modelling a chemical reaction

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PRACTICAL ACTIVITY 21.2.2

Dodgem car crash analogy

In this activity you will compare a 'successful' dodgem car crash with a chemical reaction.

AIM

To compare the components of a chemical reaction with those of dodgem cars when they crash

PROCEDURE

1 In small groups, brainstorm the following ideas.
   - Consider a dodgem car crash.
   - What would make a dodgem car crash 'successful'?
   - Are all dodgem car crashes successful?
   - What has to happen during a dodgem car crash for it to be successful?
2 Write a paragraph describing a dodgem car crash and what needs to happen for 'the crash' to be successful.
3 Save your paragraph to refer to later.

FIGURE 21.2.2 Is this a successful dodgem car crash?

Consider the reaction between hydrogen gas and oxygen gas that occurs to form water:

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

For this reaction to occur:

- bonds in hydrogen molecules must break to separate the hydrogen atoms
- bonds in oxygen molecules must break to separate the oxygen atoms
- new bonds must form between the hydrogen and oxygen atoms to make the water molecules.

This reaction requires enough energy to overcome the forces of attraction holding the atoms of the reactants in their original bonds. Enough energy is required to break the covalent bonds within the hydrogen molecules and oxygen molecules.

This reaction also requires the required atoms to contact each other to form the new bonds of the products. This means that two hydrogen atoms must contact one oxygen atom to form a water molecule.

Collision theory

To understand why reactions occur quickly or slowly, you need to know what happens during a chemical reaction. The **collision theory** explains what happens during a chemical reaction and the factors that affect it. The collision theory states that, for a reaction to occur, the reactants must collide with sufficient energy and at the required orientation. Collisions between reactant particles are classified as either 'successful' or 'unsuccessful' (Figure 21.2.3). A successful collision occurs with enough energy and at the required orientation to break old bonds and to make new bonds. In an unsuccessful reaction, the energy, the orientation or both are not satisfactory. The more often that successful collisions occur, the faster the reaction rate because more reactants are forming products.

collision theory
for a reaction to occur, the particles must collide with sufficient energy and in the required orientation

20.1.2 Collision theory

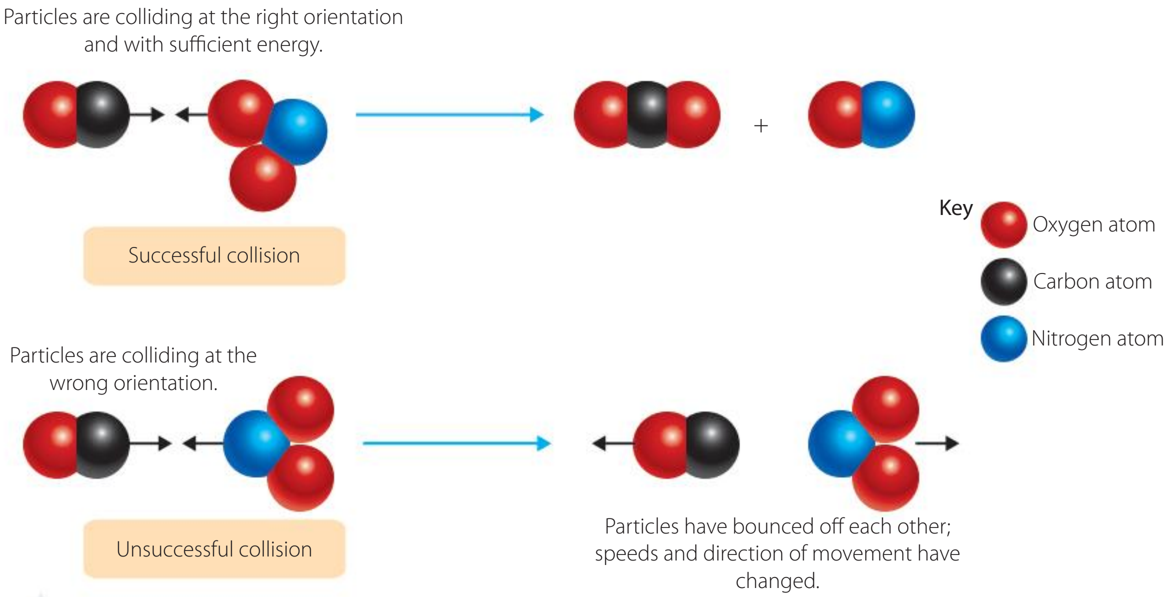


FIGURE 21.2.3 The correct orientation is needed for a successful collision.

PRACTICAL ACTIVITY 21.2.3

Dodgem car crash analogy – collision theory

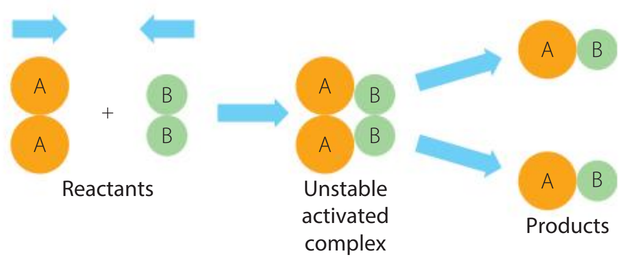
1. Review Practical activity 21.2.2: Dodgem car crash analogy
2. In a dodgem car crash, what would be the equivalent of these parts of the collision theory?
   a. The collision
   b. Enough energy
   c. Appropriate orientation
3. For each part of Question **2**, explain the relevance to a 'successful' dodgem car crash.

activated complex also known as the transition state, the intermediate stage of a chemical reaction where the reactant bonds have been broken but the product bonds are not yet formed

transition state see activated complex

During the reaction, as reactants form products, a brief intermediate stage occurs where the bonds of the reactants have broken, but the bonds of the products have not yet formed. This produces a temporary, highly unstable structure called the **activated complex** or **transition state** (Figure 21.2.4). Energy is needed to break the bonds of the reactants, so the activated complex has more energy than either reactants or products.

FIGURE 21.2.4 Events in a reaction ($A_2 + B_2 \rightarrow 2AB$)



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The activated complex will quickly form new bonds to produce the products, releasing energy. This means that the required bonding atoms must be positioned next to each other to form the new bonds. The orientation of the colliding particles needs to be correct.

Energy profile diagrams

The heat energy possessed by a chemical substance is called **enthalpy**, represented by the symbol *H*. In a reaction, the reactants, activated complex and products each have their own enthalpies. These can be represented in an energy profile diagram.

enthalpy the total energy possessed by a chemical substance, which changes as bonds are broken or made during a chemical reaction

Energy is needed to break the bonds of the reactants. The activated complex has more enthalpy than the reactants. Similarly, energy is lost from the activated complex when the bonds of the products are formed. The activated complex also has more enthalpy than the products (Figure 21.2.5).

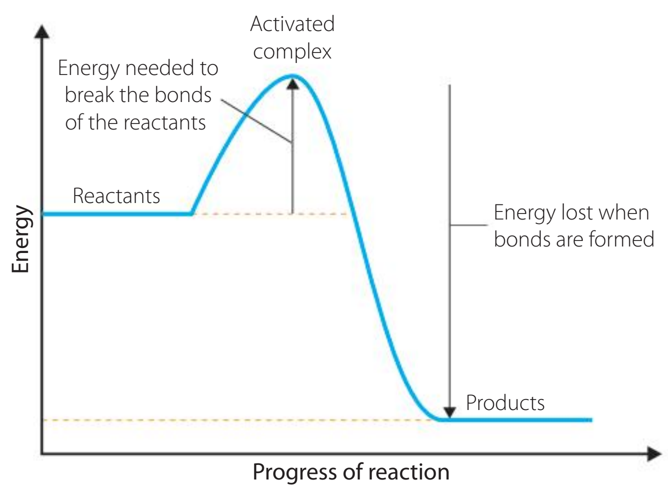


FIGURE 21.2.5 An energy profile diagram of a chemical reaction

If more energy is lost when products form than is gained to make the activated complex, then the enthalpy difference is lost from the system to the surroundings. This occurs in an exothermic reaction. In contrast, if more energy is gained than is lost, the excess must be gained from the surroundings. This occurs in an endothermic reaction (Figure 21.2.6).

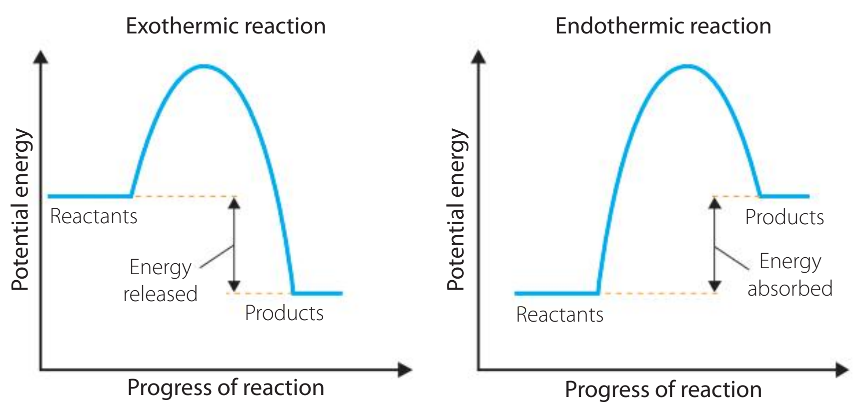


FIGURE 21.2.6 Energy profile diagrams for endothermic and exothermic reactions

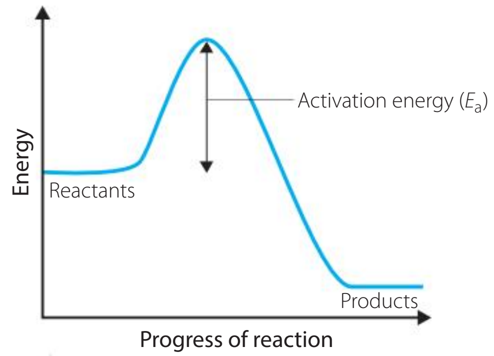


FIGURE 21.2.7 The activation energy of a reaction

activation energy
the minimum energy required for a chemical reaction to occur: it significantly influences the rate of a chemical reaction

temperature
a measure of the average kinetic energy of the particles that constitute a solid, liquid or gas

kinetic energy
the energy possessed by moving particles: a particle of mass *m*, moving with speed *v* has a translational kinetic energy equal to $\frac{1}{2}mv^2$; molecules also possess rotational and vibrational kinetic energy

Maxwell–Boltzmann distribution
a frequency distribution of the energy value of a sample of particles at a particular temperature

Activation energy

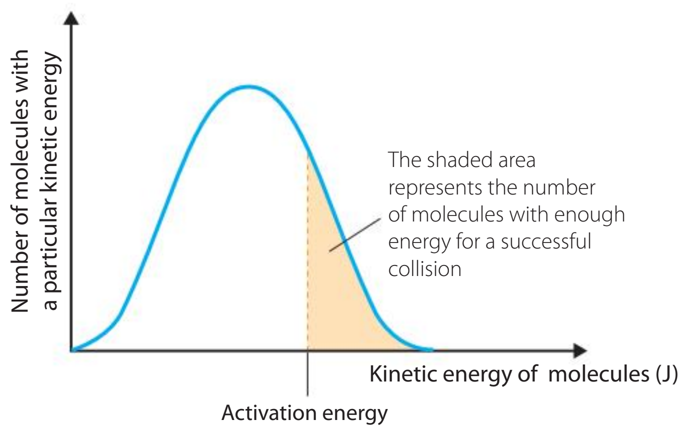
The amount of energy needed to form the activated complex is called the **activation energy** and is often represented as E_a. This energy is needed for the bonds of the reactants to break, so the activation energy is the minimum energy required for a reaction to occur. The activation energy can be determined from the difference in energy between the activated complex and the reactants (Figure 21.2.7).

The Maxwell–Boltzmann distribution

At a given **temperature**, particles will have different amounts of **kinetic energy**, with the average kinetic energy determining the temperature. This is represented in the **Maxwell–Boltzmann distribution**, which graphically shows the kinetic energies of the particles at a particular temperature.

In addition to kinetic energy, molecules also possess *internal* energy in the form of vibrational and rotational energy. If electrons are excited to levels above the ground state, the molecules can also have additional electronic energy. The Maxwell–Boltzmann distribution considers only kinetic energy (Figure 21.2.8).

FIGURE 21.2.8 The Maxwell–Boltzmann distribution shows the translational kinetic energy of molecules at a particular temperature.



The Maxwell–Boltzmann distribution shows that a few particles have low kinetic energy, and a few have high kinetic energy. Most particles have a kinetic energy value in the middle range.

Only those particles with the activation energy, or greater, will have enough energy to break the bonds of the reactants and have a successful collision. On the graph, the proportion of particles with that amount of energy is represented by the shaded yellow area under the curve greater than the activation energy.

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SECTION REVIEW 21.2

REMEMBERING

1 Outline collision theory.

2 Define:

a activation energy

b activated complex.

3 Draw a Maxwell–Boltzmann distribution.

UNDERSTANDING

4 Explain why a successful collision requires:

a sufficient energy

b the correct orientation.

5 What will happen to two particles that collide with energy less than the activation energy?

APPLYING

6 Draw and label an energy profile diagram for an:

a endothermic reaction with an activation energy of 1000 J

b exothermic reaction with an activation energy of 200 J.

21.3 Factors that affect the rate of a chemical reaction

Collision theory tells us that the rate of a reaction depends on the number of successful collisions between reactant particles in a given time. The number of successful collisions depends on the:

- total number of collisions
- proportion of collisions that are successful.

If either of these factors increases, the reaction rate will increase. Similarly, if either of these factors decreases, the reaction rate will decrease.

For example, if 1000 collisions between reactants occur per second and 60% are successful, this means there are 600 successful collisions per second (Table 21.3.1). If the number of collisions increases to 2000 per second, and 60% are successful, this means there are now 1200 successful collisions per second, meaning that the reaction rate has increased. Alternatively, with 1000 collisions per second and 80% of them are successful, this equates to 800 successful collisions per second; the reaction rate has still increased.

TABLE 21.3.1 Successful collisions

NUMBER OF COLLISIONS	PERCENTAGE OF COLLISIONS THAT ARE SUCCESSFUL	NUMBER OF SUCCESSFUL COLLISIONS
1000	60	600
1000	80	800
2000	60	1200
2000	80	1600

PRACTICAL ACTIVITY 21.3.1

Changing the colour of potassium permanganate

Rhubarb contains oxalic acid ($C_2H_2O_4$), which will react with acidified potassium permanganate ($KMnO_4$), causing it to decolourise. This can be represented by the following equation:

$$2MnO_4^- + 5C_2H_2O_4 + 6H_3O^+ \rightarrow 2Mn^{2+} + 10CO_2 + 14H_2O$$

Research to learn more about the reaction between oxalic acid in rhubarb and potassium permanganate. The time for this change in colour to occur can be measured and used to calculate the reaction rate. The rhubarb may be fresh, in frozen rhubarb sticks, or a puree that can be made by gently boiling rhubarb in water. Consider all the factors that may affect the time for the permanganate to decolourise. Choose one factor and design an investigation to determine its effect on the reaction rate.

AIM

Write an aim for your investigation, stating clearly what you are investigating

EQUIPMENT PER GROUP

List all the materials and equipment that you will need to carry out your investigation. Make sure that you are specific about the type, number and size of all the materials and apparatus needed.

RISK ASSESSMENT

WHAT ARE THE RISKS IN DOING THIS INVESTIGATION?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?

Copy and complete a risk assessment table. Identify specific risks involved in the investigation and how to manage them to avoid injuries or damage to equipment. Ask your teacher to check your risk assessment before you proceed.

PROCEDURE

Write clearly, in numbered steps, how you will conduct your investigation to achieve your aim. Make sure your steps show how you are going to control or measure all your variables.

QUESTIONS

1 What is your independent variable? How will you change this?
2 What is your dependent variable? How will you measure this?
3 Do you need to measure and record any other data?
4 Will you collect quantitative and/or qualitative results? List the sort of results you will be collecting and when you will collect them.

ANALYSIS

1 How will you calculate the rate of reaction from your data?
2 Explain how you analysed your results, and write up the analysis.

QUESTIONS

1 What did you find when you analysed your results?
2 Did you have any problems with your materials or method? How would you overcome this if you repeated this investigation?

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CONCLUSION

1 How can science be used to explain your conclusion?
2 Write a conclusion to your investigation based on your aim and your results.

TAKING IT FURTHER

Could you undertake any further investigation to find out more information about this topic? If yes, what would you want to find out, and how would you go about doing that?

PRACTICAL ACTIVITY 21.3.2

Dodgem car crash analogy – the nature of reactants

PROCEDURE

Review Practical activity 21.2.2: Dodgem car crash analogy

QUESTIONS

1 What is the equivalent of the nature of reactants for your dodgem car crash analogy?
2 How can you increase the 'activation energy' for your dodgem car crash?
3 How can you decrease the 'activation energy' for your dodgem car crash?
4 Do car manufacturers consider these factors to make cars safer?

CONCLUSION

Write a paragraph to summarise how the nature of reactants will affect the rate of dodgem car crashes.

Concentration and pressure

Concentration refers to the number of particles in a given volume. In a gas, the concentration of the gaseous particles determines its **gas pressure**. The greater the concentration, or pressure, the greater the number of collisions. Although the percentage of successful collisions remains the same, the total number of successful collisions increases because collisions increase. As the concentration, or pressure, of the reactants increases, the reaction rate increases.

gas pressure the force per unit area exerted by gas particles as they collide with the walls of their container


Mark Fergus Photography

FIGURE 21.3.1 Solutions of different concentrations of the same substance

PRACTICAL ACTIVITY 21.3.3

Effect of concentration on the rate of a reaction

Sodium thiosulfate ($Na_2S_2O_3$) reacts with dilute hydrochloric acid (HCl) to produce sodium chloride (NaCl), sulfur (S), sulfur dioxide (SO_2) and water (H_2O).

$$Na_2S_2O_3 + 2HCl \rightarrow 2NaCl + S + SO_2 + H_2O$$

When the concentration of sulfur is high enough, the solution will become cloudy. How quickly the solution becomes cloudy indicates the rate of the reaction.

AIM

To investigate the effect of concentration on the rate of the reaction between sodium thiosulfate and dilute hydrochloric acid

EQUIPMENT PER GROUP

For one trial:

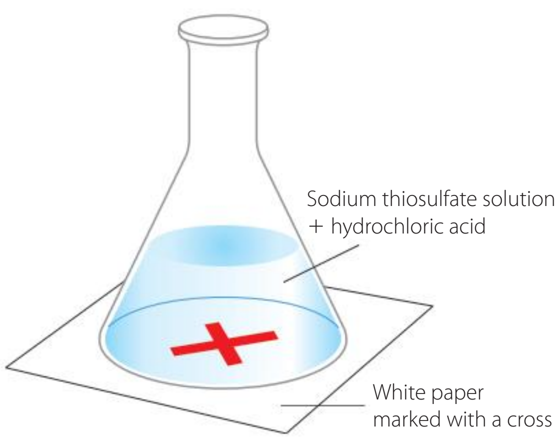
- 45 mL of 0.25 mol L^{-1} $Na_2S_2O_3$
- 45 mL of 0.125 mol L^{-1} $Na_2S_2O_3$
- 45 mL of 0.0625 mol L^{-1} $Na_2S_2O_3$
- 15 mL of 2 mol L^{-1} HCl
- stopwatch
- 3 × 100 mL conical flask
- 1 × 50 mL measuring cylinder
- 1 × 10 mL measuring cylinder
- white paper
- pen or marker pen

RISK ASSESSMENT

WHAT ARE THE RISKS IN DOING THIS EXPERIMENT?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?
Chemicals could splash in your eyes.	Wear safety glasses at all times.
2 mol L^{-1} HCl is corrosive to skin and clothes.	Wear gloves and an apron.
SO_2 gas irritates the nose, throat and airways.	Conduct the experiment in a well-ventilated area and do not inhale fumes.

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, as well as ways to manage them. Expand on the three risks listed to identify specific risks involved with each of them.

FIGURE 21.3.2 Experimental set-up for the reaction between sodium thiosulfate and dilute hydrochloric acid



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PROCEDURE

1. Draw a large cross in the centre of the piece of white paper.
2. Place the conical flask over the cross.
3. Accurately measure 45 mL of 0.0625 mol L^{-1} $Na_2S_2O_3$ in a measuring cylinder.
4. Pour the sodium thiosulfate into the conical flask.
5. Accurately measure 5 mL of 2 mol L^{-1} HCl in a measuring cylinder.
6. Have a stopwatch ready to start timing.
7. Pour the hydrochloric acid into the conical flask, starting the stopwatch immediately.
8. Quickly mix the two solutions by carefully swirling the conical flask.
9. Place the conical flask onto the cross and record the time it takes before you can no longer see the cross when looking from above. (Be careful not to breathe in any fumes from the reaction.)
10. Repeat steps **3**–**9** for the other two solutions.
11. If time and resources allow, repeat steps 3–10 for trials 2 and 3.
12. Place your solutions in the waste beaker, and rinse the conical flask.

RESULTS

Copy and complete your results in a results table similar to the one below.

CONCENTRATION OF SODIUM THIOSULFATE (mol L^{-1})	TIME TAKEN FOR THE CROSS TO DISAPPEAR (S)			
	TRIAL 1	TRIAL 2	TRIAL 3	AVERAGE
0.0625				
0.125				
0.250				

RESULTS

1. Graph the average time taken for the cross to disappear against the concentration of sodium thiosulfate.
2. How is the time taken for the cross to disappear related to the rate of the reaction?
3. What happened to the time taken for the cross to disappear as the concentration of sodium thiosulfate increased?
4. What is the relationship between the concentration of sodium thiosulfate and the rate of the reaction?
5. Do your data support the theoretical trend?
6. Were there any anomalies? If yes, why do you think they occurred?

DISCUSSION

1. Use collision theory to explain the relationship between concentration and the rate of the reaction.
2. Where might errors have occurred in the method that would have affected the accuracy of the results?
3. How could the method be improved to increase the accuracy and reliability of the results?

CONCLUSION

What conclusion can you make about the concentration of reactants and the rate of the reaction?

TAKING IT FURTHER

1. How could you further investigate the effect of concentration on the rate of the reaction in this reaction?
2. How could you further investigate the effect of concentration on the rate of reactions in general?

Changing the volume may also change the concentration. In a solution, adding water will dilute the solution, thereby decreasing the concentration. Alternatively, evaporating some of the water will decrease the volume and increase the concentration. In a gas, if the volume of the container is increased, then the concentration of all the gases will decrease. Alternatively, if the volume of the container is decreased, then the concentration of all the gases will increase.

During a reaction, reactants are being used to form the product, decreasing the concentration of the reactants. By applying the collision theory, scientists can predict that reaction rate will decrease as time progresses during the reaction. On a graph of the reactants or products versus time, the reaction rate is indicated by the slope of the graph. The slope will be steepest at the start of the reaction and will gradually decrease to zero at the completion of the reaction (Figure 21.1.6).

PRACTICAL ACTIVITY 21.3.4

Dodgem car crash analogy – concentration

AIM

To compare the effect of concentration on the rate of a reaction with dodgem car crashes

PROCEDURE

Review Practical activity 21.2.2: Dodgem car crash analogy.

QUESTIONS

1 In your dodgem car crash analogy, what would be the equivalent to concentration?

2 How would changing this concentration equivalent affect the rate of dodgem car crashes?

3 Would it matter if only one 'reactant' or all 'reactants' were changed?

CONCLUSION

Write a paragraph summarising how changes in concentration will affect the rate of dodgem car crashes.

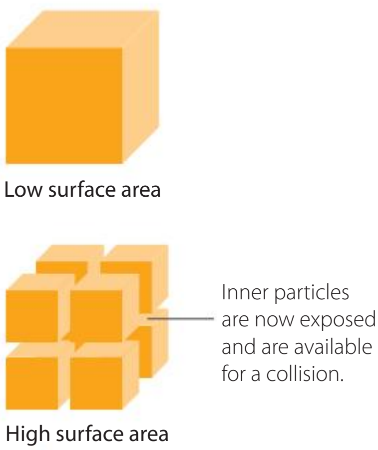

FIGURE 21.3.3 Dividing a solid will increase the surface area exposed for a reaction

surface area
the area for a successful collision between particles to occur that enables a reaction to take place

Surface area

For a reaction to occur, the reactant particles need to collide. When one of the reactants is a solid, only the particles on the surface of the solid are available for collision. Once these have reacted, then the particles beneath become exposed, and so on.

If the solid is divided or crushed, the inner particles become exposed. Although the amount of reactant remains the same, the total **surface area** has increased. More of the solid is then available to react (Figure 21.3.3). There are more collisions and the reaction rate increases.

People use this method of increasing the reaction rate in many everyday activities: chopping wood into smaller pieces of kindling to start the fire; cutting potatoes into small pieces to decrease cooking time; and using granulated sugar instead of sugar cubes in drinks.

PRACTICAL ACTIVITY 21.3.5

Effect of surface area on rate of reaction

Calcium carbonate ($CaCO_3(s)$) will react with dilute hydrochloric acid (HCl) according to the following equation:

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

Bubbles of carbon dioxide are visible during the reaction. When no more bubbles are produced, the reaction has stopped. The time it takes for the bubbles to stop can be measured to compare the rates of reactions.

In this experiment, you will change the surface area of a fixed mass of solid $CaCO_3$. By comparing the time taken for the reactions to occur, can determine the effect of surface area on the rate of the reaction.

AIM

To determine the effect of surface area on the rate of the reaction between calcium carbonate and dilute hydrochloric acid

EQUIPMENT PER GROUP

- 3 chips of calcium carbonate
- calcium carbonate powder
- 50 mL of 2 mol L^{-1} HCl
- electronic balance
- 10 mL measuring cylinder
- stopwatch
- 2 large test tubes

WHAT ARE THE RISKS IN DOING THIS EXPERIMENT?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?
Chemicals may splash in your eyes.	Wear safety glasses at all times.
2 mol L^{-1} HCl is corrosive to skin and clothes.	Wear gloves and an apron.

RISK ASSESSMENT

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, as well as ways to manage them. Expand on the two risks listed to identify specific risks involved with each of them.

PROCEDURE

1. Measure 8 mL of HCl in a measuring cylinder.
2. Weigh and record the mass of three calcium carbonate chips.
3. Place the calcium carbonate chips in one test tube.
4. Have a stopwatch ready to start timing.
5. Pour the hydrochloric acid into the test tube and start the stopwatch.
6. Stop the stopwatch when no more bubbles are produced, and record this time.
7. Carefully weigh the same mass of calcium carbonate powder as that of the calcium carbonate chips.
8. Repeat steps **3**–**6** for the powder.
9. If time permits, repeat steps **1**–**8** for several trials.
10. Place all the product solutions into the waste beakers as directed by your teacher.

RESULTS

Copy and complete your results in a results table similar to the one below.

FORM OF CALCIUM CARBONATE	TIME FOR THE REACTION TO COMPLETE (S)			
	TRIAL 1	TRIAL 2	TRIAL 3	AVERAGE
Chips				
Powder				

QUESTIONS

1 Which form of calcium carbonate took the least time to react?
2 Which form of calcium carbonate has the fastest rate of reaction?
3 Which form of calcium carbonate had the greatest surface area? Explain why.
4 What is the relationship between surface area and the rate of reaction?
5 Do your results support the theory about the effect of surface area on the rate of reaction?
6 Was this a fair test? Explain why or why not.
7 Were the results accurate and reliable? Explain why or why not.
8 How could the method be improved to more fairly and accurately test the effect of surface area on the rate of reaction between calcium carbonate and dilute hydrochloric acid?

CONCLUSION

What conclusion can you make from this experiment?

TAKING IT FURTHER

How could you further investigate the effect of surface area on the rates of reactions?

A similar effect to increasing the surface area is seen when the reactants are stirred. Stirring mixes the reactants together, exposing more of the reactants to each other. This increases the number of collisions and the reaction rate. This phenomenon is demonstrated when making a cup of tea or coffee. By stirring the sugar, the rate at which it dissolves increases.

FIGURE 21.3.4 Stirring will have the same effect as increasing the surface area as more particles are exposed to one another.

Mark Fergus Photography

PRACTICAL ACTIVITY 21.3.6

Dodgem car crash analogy – surface area

AIM

To compare the effect of surface area on the rate of a reaction with dodgem car crashes

PROCEDURE

Review Practical activity 21.2.2: Dodgem car crash analogy

QUESTIONS

1 In your dodgem car crash analogy, what would be the equivalent to surface area?

2 How would changing this surface area equivalent affect the rate of dodgem car crashes?

CONCLUSION

Write a paragraph summarising how the surface area will affect the rate of dodgem car crashes.

Temperature

The temperature of a substance is a measure of the average kinetic energy of the particles. This energy is due to movement and depends on the speed, or velocity, of the particles. As the temperature of a substance increases, so does the average kinetic energy and the average speed of the particles. For example, applying heat to a liquid or gas causes the particles to move faster, collide more energetically, increase the average kinetic energy and, in turn, the temperature (Figure 21.3.5).

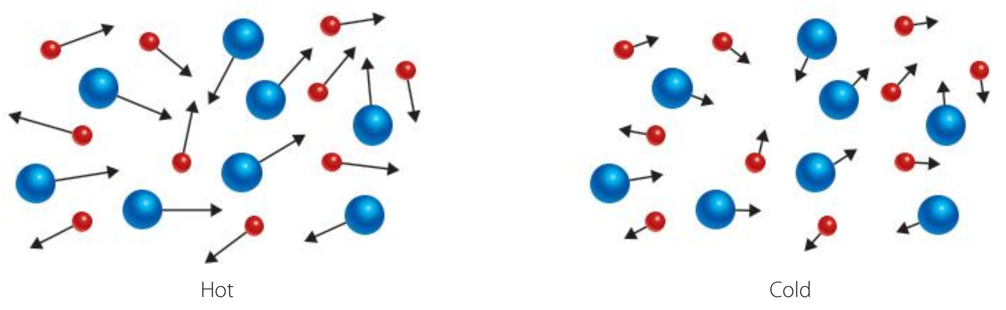

FIGURE 21.3.5 Temperature changes the speed at which particles move (as represented by the length of the arrows)

Kinetic energy can be calculated by using:

$$\text{Kinetic energy} = \frac{1}{2}\,\text{mass} \times \text{velocity}^2$$

In this calculation, velocity is the speed in a straight line. This means that if the kinetic energy of an object increases, then the velocity increases. This also means that if two objects at the same temperature (the same kinetic energy) have different masses, then their velocities will also differ. The object with the greater mass will have the lower velocity.

As the temperature of reactants increases, both the number of collisions and the percentage of successful collisions increase, because temperature has a dual effect on reaction rate.

Increased percentage of successful collisions

As the temperature increases, the average kinetic energy of the particles increases. More particles have enough energy to overcome the activation energy, form the activated complex and become products. A greater percentage of collisions are successful, increasing the reaction rate.

Figure 21.3.6 shows that at a higher temperature (T_2) the Maxwell–Boltzmann distribution moves to the right, increasing the area under the curve to the right of the activation energy. A greater proportion of particles with enough energy for a successful collision increases.

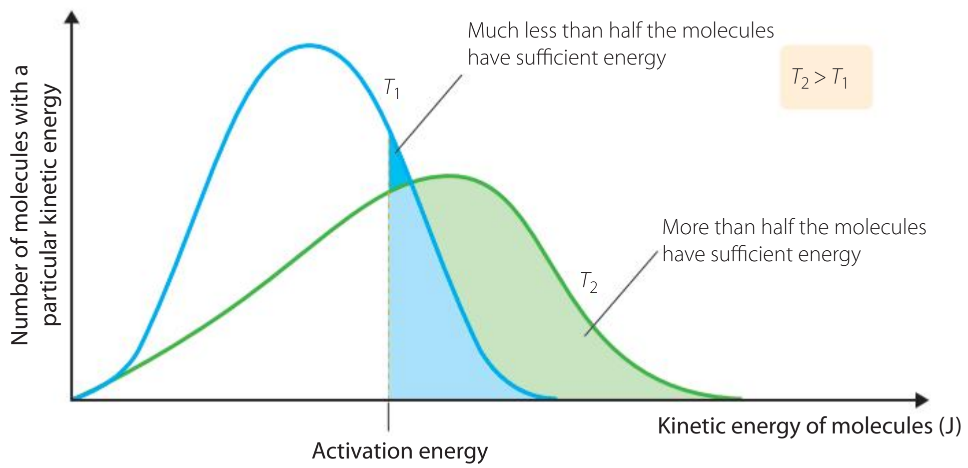

FIGURE 21.3.6 The Maxwell–Boltzmann distribution corresponding to different temperatures

Increased number of collisions

At a higher temperature, the particles have a greater average speed. They will collide with each other more frequently, increasing the number of collisions and the rate of the reaction. However, increasing the number of collisions does not affect the reaction rate as much as the effect of the increased energy, which enables a greater number of successful collisions as there are more particles with the energy equal to, or greater than, the activation energy.

PRACTICAL ACTIVITY 21.3.7

Dodgem car crash analogy – temperature

AIM

To consider the effect of temperature on reaction rates in the dodgem car crash analogy

PROCEDURE

Review Practical activity 21.2.2: Dodgem car crash analogy

QUESTIONS

1 In your analogy, what would be the equivalent of temperature? (Remember that temperature measures the average kinetic energy.)
2 How would increasing this affect the rate of dodgem car crashes?
3 Is this consistent with the effect of temperature on reaction rates?

CONCLUSION

Write a paragraph summarising how the equivalent of temperature will affect the rate of dodgem car crashes.

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Catalysts

A **catalyst** affects the rate of a reaction without being consumed in the reaction. Positive catalysts increase the rate of the reaction. Negative catalysts decrease the rate of the reaction. Scientists tend to refer to positive catalysts.

catalyst
a substance that affects the rate of certain reactions by providing an alternative reaction pathway

There are various ways that a catalyst can affect a reaction, but all positive catalysts provide an alternative pathway for the reaction where the alternative path has a lower activation energy. A lower activation energy means that, for a given temperature, more molecules will have enough energy for successful collisions. More successful collisions occur, so there is a faster rate of reaction.

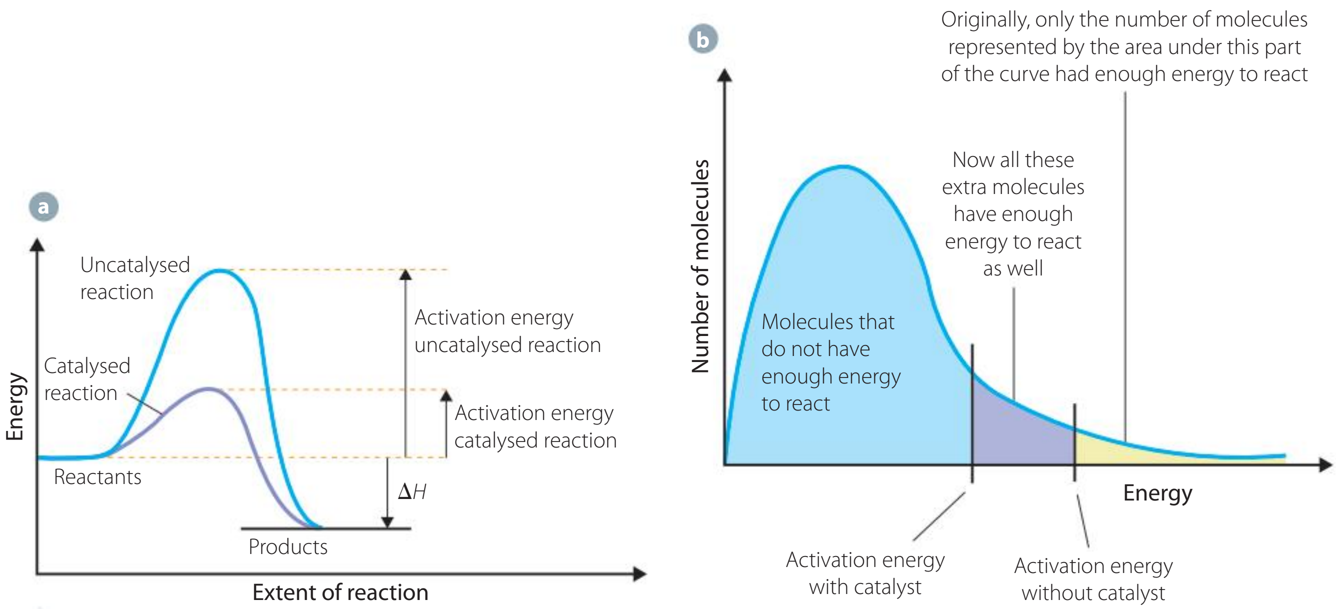

FIGURE 21.3.7 The effect of a catalyst on the number of molecules with sufficient energy for a successful collision: **a** energy profile diagram of a catalysed and an uncatalysed reaction; and **b** Maxwell–Boltzmann distribution for catalysed and uncatalysed reactions.

Cars are now equipped with catalytic converters. These structures are fitted so that the exhaust fumes pass over a layer of a metal catalyst, usually platinum, rhodium and/or palladium, which is incorporated in a heat-resistant ceramic core that provides a large surface area. As the gases pass over the catalyst, the harmful combustion products of nitrogen oxides, volatile organic compounds and carbon monoxide are converted to less harmful substances (mainly carbon dioxide, water, nitrogen and oxygen). This is why the catalytic converter is an important factor reducing toxic pollution emitted by cars.

Metal **nanoparticles** are being used more frequently as catalysts in scientific and industrial applications. A nanoparticle is a particle with a diameter less than 100 nm that behaves as a whole unit. Metal nanoparticles have an advantage over other metal catalysts because of their large surface area. Large amounts of reactants can access the catalyst and they can have a greater impact on the reaction rate. Scientists research different metal nanoparticles for a number of reactions, including gold and copper nanoparticles as catalysts for fuel cells.

nanoparticle
a particle with a diameter less than 100 nm that behaves as a whole unit

PRACTICAL ACTIVITY 21.3.8

Dodgem car crash analogy – catalyst

AIM

To consider the effect of a catalyst on reaction rates in the dodgem car crash analogy

PROCEDURE

Review Practical activity 21.2.2: Dodgem car crash analogy

QUESTIONS

1 In your analogy, what would be the equivalent of a catalyst? (Remember that a catalyst makes it easier for a collision to succeed because less energy is needed for a successful collision.)
2 How would adding this catalyst affect the rate of dodgem car crashes?
3 Is this consistent with the effect of catalysts on reaction rates?

CONCLUSION

Write a paragraph summarising how a catalyst will affect the rate of dodgem car crashes.

SECTION REVIEW 21.3

REMEMBERING

1 List the factors that affect the rate of a reaction.
2 Define 'catalyst'.
3 Draw a Maxwell–Boltzmann distribution to show the distribution of energy of particles at two different temperatures. Clearly identify which curve is for the higher temperature.

UNDERSTANDING

4 Explain how crushing a lump of solid will increase its surface area.
5 Compare the average kinetic energy of the particles of substances at 10°C and at 40°C.
6 Explain how the pressure of a gas is equivalent to its concentration.
7 Use collision theory to explain how each of the following factors affects the rate of a reaction.
   a Concentration or pressure
   b Temperature
   c Catalyst
   d Surface area

APPLYING

8 Ben and Amy both added a sugar cube to their cups of tea, which were at the same temperature. Ben stirred his tea, while Amy left her cup to sit. After 1 minute, they each drank their tea. Whose tea would have been sweeter? Justify your answer.

ANALYSING

9 Five test tubes each contained hydrochloric acid (HCl) and magnesium as shown in the following table. Which test tube would contain the fastest reaction? Justify your answer.

TEST TUBE	CONCENTRATION OF HCL ($mol\,L^{-1}$)	TEMPERATURE (°C)	FORM OF 2 g OF MAGNESIUM
1	1	25	Lump
2	2	25	Long strip
3	1	10	Strip cut into small pieces
4	0.5	40	Long strip
5	2	40	Strip cut into small pieces

REFLECTING

10 Explain how the development of collision theory has enabled us to understand why reactions occur at different rates.

21.4 Enzymes

Chemical reactions occur every day in your body to keep it healthy. Many of these reactions rely on **enzymes** to allow the reaction to occur at an adequate rate. An enzyme is a biological catalyst. This means that it is found in living organisms and, being a catalyst, can speed up a chemical reaction without being consumed.

enzyme
a biological catalyst

21.4.1 What are enzymes?

Enzymes are usually large **protein** molecules. Enzymes are highly specific and act on a particular reaction, type of reaction or bond type by making it easier for the reaction to occur. The enzyme has an **active site** with a shape that matches the reactant particles. When the enzyme and reactants (known as the **substrates**) bind, a change occurs which allows the reactants to form the products in a way that lowers the activation energy. Less energy is needed for a successful collision, so more particles have enough energy and more collisions will result in a chemical reaction.

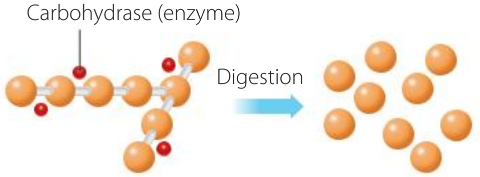

FIGURE 21.4.1 Enzymes help break complex molecules into smaller, simpler molecules that can be absorbed. Carbohydrase acts on starch molecules to produce sugar molecules.

Our digestive system provides the nutrients needed for growth, repair and energy. Our food is largely made up of **carbohydrates**, proteins and **fats**. These are large organic molecules that are made up of smaller units, as shown in Table 21.4.1. The digestive system breaks down the large molecules into smaller molecules that can be absorbed into the bloodstream. In the blood, the molecules are transported to where they are needed. However, if the large molecules cannot be broken down, they cannot be absorbed and the food cannot be used by the body.

protein
food group consisting of chains of amino acid molecules – they are essential for life

active site
area of the enzyme where the enzyme interacts directly with the substrate molecule

substrate
molecule that is taking part in the reaction which the enzyme is catalysing

carbohydrates
food group consisting of molecules such as glucose, sugars and starches

fat
food group consisting of chains of triglyceride molecules – they are essential for life

TABLE 21.4.1 Nutrients in food

NUTRIENT	SMALL UNITS THAT CAN BE ABSORBED
Proteins	Amino acids
Carbohydrates	Simple sugars; for example, glucose
Fats	Glycerol and fatty acids

Enzymes can be specific for a particular reaction, breaking a particular bond in a particular molecule, as, shown in Figure 21.4.2. This means that the larger molecules are split into smaller sections. Upon binding to the enzyme at the active site, the substrate is changed in such a way as to allow the bond to be broken with less energy. The products formed are then released and the enzyme is available for further reactions. Table 21.4.2 lists some features of some digestive enzymes.

FIGURE 21.4.2 A model for an enzyme

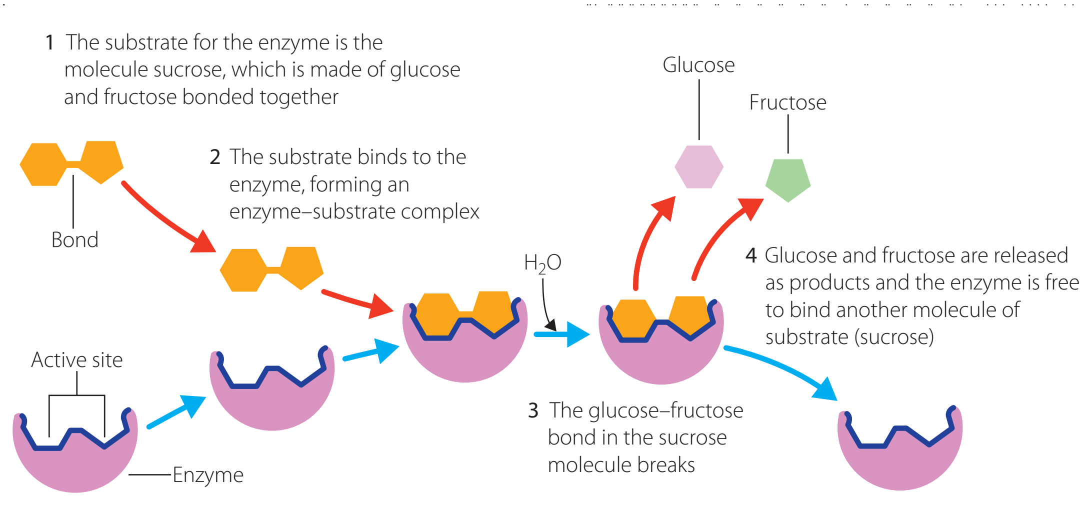

TABLE 21.4.2 Digestive enzymes

ENZYME	LOCATION	REACTANT	PRODUCT
Amylase	Saliva in the mouth Pancreas releases it into the small intestine	Starch (a large carbohydrate)	Maltose
Protease (pepsin)	Stomach	Proteins	Smaller proteins
Protease (trypsin)	Pancreas releases it into the small intestine	Proteins	Peptides and amino acids
Lipase	Pancreas releases it into the small intestine	Fats	Fatty acids and glycerol
Peptidase	Small intestine	Peptides (small sections of protein)	Amino acids
Sucrase	Small intestine	Sucrose (a small carbohydrate)	Glucose and fructose
Maltase	Small intestine	Maltose (a small carbohydrate)	Glucose
Lactase	Small intestine	Lactose (a small carbohydrate)	Glucose and galactose

There are thousands of different enzymes in the human body controlling the wide range of reactions that are occurring. These include:

- *carbonic anhydrase*, which catalyses the reaction between carbon dioxide (CO_2) and water (H_2O) to form carbonic acid (H_2CO_3), allowing carbon dioxide to travel through the blood:

$$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$$

- *DNA polymerase*, which is involved in DNA replication

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- *dehydrogenase*, an important enzyme in the first step of cellular respiration in mitochondria to produce energy in our bodies
- *cyclin dependent kinases*, which control cell division
- *tyrosinase*, which triggers melanin synthesis and is important in determining skin colour.

The activity or effectiveness of enzymes can be affected by the conditions or environment in which they are found. These conditions include:

- *concentration*: the higher the enzyme concentration, the more substrate that can bind to an enzyme at a given time and the greater the reaction rate
- *temperature*: enzymes are proteins, and as such many human proteins are **denatured** above about 45°C. Above this temperature, the structure of the protein changes, changing the shape of the active site. This permanently inactivates the enzyme as the substrate is unable to bind to the active site
- *pH*: different enzymes require different pH values for optimum activity; for example, the pH of the stomach is much less than the pH of the rest of the digestive system, so only specific enzymes will be activated in the stomach
- ***co-enzymes***: ions or non-protein molecules such as vitamins that change the shape of the active site for the enzyme to function.

denature
the permanent change than can occur to the structure of protein molecules when heated, or exposed to solutions of high acidity or alkalinity

co-enzymes
molecules that assist the operation of enzymes

SECTION REVIEW 21.4

REMEMBERING

1 In relation to enzymes, define:
  - a active site
  - b substrate
  - c denature
  - d co-enzyme.

2 List the factors that can affect the rate of a biological reaction.

UNDERSTANDING

3 Explain why enzymes are generally more effective than non-biological catalysts.

APPLYING

4 For a non-biological reaction, rate of the reaction increases as the temperature increases. This can be explained by collision theory. However, when a catalyst is involved, the relationship between temperature and rate of reaction is more complex. Explain why this is so.

21.5 Mandatory practical

PRACTICAL ACTIVITY 21.5.1

Investigating the rate of a reaction

The rate of a reaction is determined by measuring how much reactant is consumed per unit of time, or how much product is made per unit of time.

The reaction between calcium carbonate ($CaCO_3$) and hydrochloric acid (HCl) produces calcium chloride ($CaCl_2$), carbon dioxide (CO_2) and water:

$$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

Determine the rate of the reaction by measuring the amount of carbon dioxide produced at different time intervals.

AIM

To observe and calculate the rate of reaction between calcium carbonate and hydrochloric acid

EQUIPMENT PER GROUP

- 2 g calcium carbonate ($CaCO_3$) powder
- 250 mL of 0.25 mol L^{-1} HCl
- 100 mL beaker
- 500 mL conical flask
- 250 mL measuring cylinder
- spatula
- paper
- graph paper
- electronic balance
- stopwatch
- calculator

RISK ASSESSMENT

WHAT ARE THE RISKS IN DOING THIS EXPERIMENT?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?
Chemicals could splash in your eyes.	Wear safety glasses at all times.
0.25 mol L^{-1} HCl is corrosive to skin and clothing.	Wear safety glasses and protective clothing. Take care when pouring and clean up spills immediately. If HCl is spilt on your skin, wash the affected area with plenty of water and notify your teacher.

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, as well as ways to manage them.

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PROCEDURE

1. Measure 2 g of $CaCO_3$ powder into the clean, dry beaker.
2. Accurately measure 250 mL of 0.25 mol L^{-1} HCl.
3. Place the hydrochloric acid in a clean conical flask and place this on the electronic balance. Record the mass of the flask containing the acid.
4. Calculate the mass of the flask, hydrochloric acid and calcium carbonate by adding your values from steps **1** and **3**.
5. Place the calcium carbonate on a clean piece of paper, and roll the paper so that it can easily be inserted into the top of the flask.
6. Add the calcium carbonate to the flask (while it is on the electronic balance), immediately start the stopwatch and record the mass.
7. Record the mass (and time) as frequently as you can.
8. When the reaction is complete, remove the flask from the electronic balance.

RESULTS

Organise your data into an appropriate table, remembering to include a column to calculate the cumulative mass of carbon dioxide gas produced for each time.

QUESTIONS

1. What happened to the mass of the flask and its contents? Why did this occur?
2. How can you use the mass of the flask to determine the mass of the carbon dioxide produced?
3. For each time, calculate the cumulative mass of the carbon dioxide produced.
4. Graph the mass of carbon dioxide produced against time.
5. When was the rate of reaction the fastest? How does the graph show this?
6. When was the rate of reaction the slowest? How does the graph show this?
7. At what time was the reaction complete? How does the graph show this?
8. Calculate the average rate of reaction (average rate = $\frac{\text{total change in mass}}{\text{time for mass change}}$).
9. Draw tangents to the graph at three different times and calculate the gradients (or slopes) of the lines. These are the rates of the reaction at those times. Ask your teacher if you do not know how to draw a tangent.

QUESTIONS

1. What is the difference between qualitative data and quantitative data?
2. During the experiment, how could you observe the rate of the reaction?
3. Did your quantitative data support your qualitative data regarding the rate of the reaction?
4. What aspects of the experiment may have reduced the data's accuracy?
5. How could you improve the experiment to make the results more accurate and reliable?

CONCLUSION

What conclusion can you make about the rate of the reaction as the reaction proceeds?

PRACTICAL ACTIVITY 21.5.2

Investigate the effect of temperature on the rate of a reaction

In Practical activity 21.3.3, you investigated the reaction between sodium thiosulfate ($Na_2S_2O_3$) and dilute hydrochloric acid (HCl):

$$Na_2S_2O_3 + 2HCl \rightarrow 2NaCl + S + SO_2 + H_2O$$

In this experiment, you will change the temperature of the solutions to investigate the effect of temperature on the reaction rate.

AIM

To investigate the effect of temperature on the rate of the reaction between sodium thiosulfate and dilute hydrochloric acid

EQUIPMENT PER GROUP

For one trial:

- 150 mL of 0.25 mol L^{-1} $Na_2S_2O_3$
- 15 mL of 2 mol L^{-1} HCl
- crushed ice
- hot water
- stopwatch
- 2 × 1 L beakers (or a similar container for an ice bath and a hot water bath)
- 3 × 100 mL conical flask
- 50 mL measuring cylinder
- 10 mL measuring cylinder
- white paper
- pen or marker pen
- thermometer

RISK ASSESSMENT

WHAT ARE THE RISKS IN DOING THIS EXPERIMENT?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?
Chemicals may splash in your eyes.	Wear safety glasses at all times.
2 mol L^{-1} HCl is corrosive to skin and clothes.	Wear gloves and an apron.
SO_2 gas irritates the nose, throat and airways.	Conduct the experiment in a well-ventilated area; do not inhale fumes.

Copy and complete the risk assessment table in your write-up. Add any more risks you can think of, and how to manage them. Expand on the three risks listed to identify specific risks involved with each of them. »

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PROCEDURE

1. Draw a large cross in the centre of a blank piece of white paper.
2. Accurately measure 45 mL of 0.0625 mol L^{-1} $Na_2S_2O_3$ in a measuring cylinder.
3. Pour the sodium thiosulfate into the first conical flask.
4. Place some crushed ice and a small amount of water into the 1 L beaker, and carefully place the conical flask into the ice bath. Ensure that there is enough ice to reach the height of the solution in the flask.
5. Repeat steps **1**–**4** with the second conical flask, using hot water instead of ice to make a hot water bath.
6. Accurately measure 45 mL of 0.25 mol L^{-1} $Na_2S_2O_3$ in a measuring cylinder.
7. Pour the sodium thiosulfate into the third conical flask.
8. Measure and record the temperature of the sodium thiosulfate in all three conical flasks.
9. Accurately measure 5 mL of 2 mol L^{-1} HCl in a measuring cylinder.
10. Have a stopwatch ready to start timing.
11. Pour the hydrochloric acid into the first conical flask, starting the stopwatch immediately.
12. Quickly mix the two solutions by carefully swirling the conical flask.
13. Place the conical flask onto the cross and record the time it takes before you can no longer see the cross when looking from above. (Be careful not to breathe in any fumes from the reaction.)
14. Repeat steps **9**–**13** for the other two solutions.
15. If time and resources allow, repeat steps **2**–**13** for trials 2 and 3.
16. Place your solutions in the waste beaker, and rinse the conical flask.

RESULTS

Copy the following table and record your results.

TEMPERATURE OF SODIUM THIOSULFATE (°C)	TIME TAKEN FOR THE CROSS TO DISAPPEAR (s)			
	TRIAL 1	TRIAL 2	TRIAL 3	AVERAGE

1. Graph the average time taken for the cross to disappear against the temperature of thesodium thiosulfate.
2. What happened to the time taken for the cross to disappear as the temperature increased?
3. What is the relationship between temperature and the rate of the reaction?

QUESTIONS

1. Do your results support the theory about the effect of temperature on the rate of reactions?
2. Where might errors have occurred in the experiment and how would these have affected the results?
3. Were your results accurate and reliable? Explain why or why not.
4. How could you improve the experiment if you were to do it again?

CONCLUSION

What did you find out about the effect of temperature on the rate of the reaction between $Na_2S_2O_3$ and dilute hydrochloric acid?

TAKING IT FURTHER

How could you further investigate the effect of temperature on the rate of reactions?

PRACTICAL ACTIVITY 21.5.3

Effect of a catalyst on the rate of a reaction

In Practical activity 21.3.1, you investigated the reaction between oxalic acid ($C_2H_2O_4$) and acidified potassium permanganate ($KMnO_4$) to produce colourless products. The intense purple of the permanganate disappears as the reaction proceeds. You can measure the time for this to occur and use it to compare the rate of reactions.

Manganese(II) sulfate ($MnSO_4$) can act as a catalyst for this reaction, increasing the rate of the reaction and decreasing the time taken for the decolourisation to occur.

AIM

To observe the effect of a catalyst on the rate of the reaction between oxalic acid and acidified potassium permanganate

EQUIPMENT PER GROUP

- 6 mL of 0.02 mol L^{-1} of $KMnO_4$
- 50 mL of 2 mol L^{-1} sulfuric acid (H_2SO_4)
- 250 mL distilled water
- 6 mL of $C_2H_2O_4$
- 1 mL of $MnSO_4$
- stopwatch
- 2 × 250 mL beakers
- 3 × 10 mL, 1 × 25 mL and 1 × 250 mL measuring cylinders
- stirring rod

RISK ASSESSMENT

WHAT ARE THE RISKS IN DOING THIS EXPERIMENT?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?
Chemicals could splash in your eyes.	Wear safety glasses at all times.
2 mol L^{-1} H_2SO_4 is corrosive to skin and clothes.	Wear gloves and an apron.
$KMnO_4$ is an oxidising agent; it can irritate the skin and eyes.	Wear gloves and an apron.
Oxalic acid is harmful to the skin and eyes.	Wear gloves and an apron.

In your write-up, add any more risks you can think of, as well as ways to manage them. In particular, expand on the risks listed to identify specific risks involved with each of them.

»

PROCEDURE

1 Into the first beaker, place 25 mL of 2 mol L^{-1} H_2SO_4, 125 mL distilled water, 3 mL of 0.02 mol L^{-1} $KMnO_4$ and 3 mL of saturated $C_2H_2O_4$.
2 Start the stopwatch immediately. Record how long it takes for the solution to become colourless.
3 Into the second beaker place 25 mL of 2 mol L^{-1} H_2SO_4, 125 mL distilled water, 3 mL of 0.02 mol L^{-1} $KMnO_4$, 3 mL of saturated $C_2H_2O_4$ and 1 mL of saturated $MnSO_4$.
4 Start the stopwatch immediately. Record how long it takes for the solution to become colourless.
5 Follow your teacher's instructions regarding disposal of chemicals.

RESULTS

Copy and record your results in a table similar to the one below.

BEAKER	TIME TAKEN FOR DECOLORISATION (S)
No $MnSO_4$	
With $MnSO_4$	

RESULTS

1 What was the effect of adding manganese sulfate to the reaction mixture?
2 Given that manganese sulfate was acting as a catalyst, did your results support the theory about the effect of catalysts on the rate of reactions?

QUESTION

How could the experiment be improved to obtain more accurate and reliable data?

CONCLUSION

What did you determine about the effect of manganese sulfate as a catalyst on the reaction between oxalic acid and potassium permanganate?

TAKING IT FURTHER

1 How could you further investigate the effect of catalysts on the rate of reaction?
2 How could you determine if other factors influence the effect of catalysts on the rate of reactions?

PRACTICAL ACTIVITY 21.5.4

Effect of pH and temperature on the enzyme complex rennet

Rennet is a complex of enzymes naturally found in the stomachs of mammals. It helps the coagulation of milk – separating it into a solid and a liquid component. It is especially important in the digestion of milk in young animals.

Rennet is also utilised in milk desserts in the form of a setting agent known as 'junket'. The enzymes help the milk to set, forming a soft solid.

AIM

To investigate the effect of pH and temperature on the effectiveness of the enzymes in rennet

EQUIPMENT PER GROUP

- 1 packet of junket (This may vary depending on the quantity that can be made – you will need enough for 65 mL milk)
- 100 mL milk
- distilled water
- 1 mL of 1 mol L^{-1} hydrochloric acid (HCl)
- 1 mL of 1 mol L^{-1} ethanoic (acetic acid) (CH_3COOH)
- 1 mL of 1 mol L^{-1} sodium ethanoate (acetate) ($NaCH_3COO$)
- 1 mL of 1 mol L^{-1} sodium hydroxide (NaOH)
- 13 test tubes
- test-tube rack
- water baths at 10°C, 20°C, 30°C, 40°C, 50°C, 60°C, 70°C and 80°C
- 10 mL measuring cylinder
- thermometer
- stopwatch
- pH meter

RISK ASSESSMENT

WHAT ARE THE RISKS IN DOING THIS EXPERIMENT?	HOW CAN YOU MANAGE THESE RISKS TO STAY SAFE?
Chemicals could splash in your eyes.	Wear safety glasses at all times.
1 mol L^{-1} HCl is corrosive to skin and clothes.	Wear gloves and aprons.
1 mol L^{-1} NaOH is corrosive to skin and clothes.	Wear gloves and aprons.

In your write-up, add any more risks you can think of, as well as ways to manage them. In particular, expand on the risks listed to identify specific risks involved with each of them.

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PROCEDURE

PART A: EFFECT OF TEMPERATURE

1. Place 5 mL milk into each of eight test tubes.
2. Place one test tube in each water bath and leave it until the milk has reached the desired temperature.
3. Dissolve the junket in 15 mL water. (You will need to keep some of this mixture for part B.)
4. Place 1 mL junket into the first test tube, mix and record the temperature inside the test tube, and start the stopwatch immediately.
5. Stop the stopwatch when the milk has become firm. Record the time taken to set.
6. Repeat steps **4** and **5** for each of the milks at the other temperatures.

PART B: EFFECT OF pH

1. Place 5 mL milk into each of five test tubes.
2. Record the temperature of the milk – these should all be the same.
3. Into test tube 1, place 1 mL distilled water.
4. Into test tube 2, place 1 mL of 1 mol L^{-1} HCl.
5. Into test tube 3, place 1 mL of 1 mol L^{-1} CH_3COOH.
6. Into test tube 4, place 1 mL of 1 mol L^{-1} NaOH.
7. Into test tube 5, place 1 mL of 1 mol L^{-1} $NaCH_3COO$.
8. Use the pH meter to test the pH of the milk in each test tube. Ensure that the meters are cleaned thoroughly between test tubes. Record the pH.
9. Add 1 mL of the junket solution from part **A** to test tube 1 and time how long it takes the milk to become firm. Record the time taken.
10. Repeat step **9** for test tubes 2–5.

RESULTS

Construct a results table in your notebook such as one shown. Record your data in the table.

VARIABLE	TEST TUBE	TEMPERATURE (°C)	pH	TIME FOR MILK TO BECOME FIRM (min)
Temperature	1			
	2			
	3			
	4			
	5			
	6			
	7			
	8			
pH	1 (distilled water)			
	2 (HCl)			
	3 (CH_3COOH)			
	4 (NaOH)			
	5 ($NaCH_3COO$)			

RESULTS

1. Draw separate graphs for parts **A** and **B**.
2. What happened to the time taken for milk to become firm as the temperature increased?
3. What happened to the time taken for the milk to become firm as the pH increased?
4. Do these results support your understanding of the effect of temperature and pH on enzymes?

QUESTIONS

1. What aspects of the experiment increased the accuracy and reliability of the results?
2. What were the major difficulties in obtaining accurate results during the experiment?
3. How could the experiment be improved?

CONCLUSION

1. What can you conclude about the effect of temperature on the activity of enzymes in junket?
2. What can you conclude about the effect of pH on the activity of the enzymes in junket?

TAKING IT FURTHER

1. In what other ways could you investigate the action of enzymes?
2. Research the role of rennet (or rennin). Explain how your findings about the effect of temperature and pH may play a role in ensuring the enzyme's effectiveness.

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CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 Define the following terms.

- **a** Rate of reaction
- **b** Surface area
- **c** Catalyst
- **d** Enzyme
- **e** Collision theory
- **f** Activated complex
- **g** Activation energy
- **h** Exothermic
- **i** Endothermic
- **j** Active site

CATEGORY QUESTIONS

2 Explain how and why changing the concentration and pressure affects the reaction rate.

3 Explain how and why changing the temperature affects the reaction rate.

4 Explain how and why changing the size of particles of solid affects the reaction rate.

ELABORATION QUESTIONS

5 Explain why reaction rate does not always increase as the temperature increases for a biological reaction involving a catalyst.

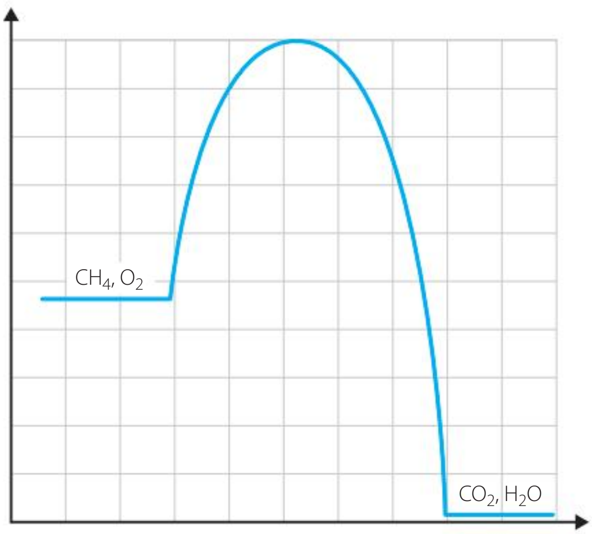
6 Use the Maxwell–Boltzmann distribution to explain why the following factors affect reaction rate.

- **a** Temperature
- **b** Presence of a catalyst

7 Copy the energy profile diagram shown (right) and label the axes, reactants, products, activation energy and the activated complex.

Describe the information that the energy profile provides about this reaction.

8 Why does the activated complex have more energy than the reactant or products?



EVIDENCE QUESTIONS

9 Explain why enzymes are critical to the functioning of the human body.

10 Explain why nanoparticles would be extremely effective catalysts.

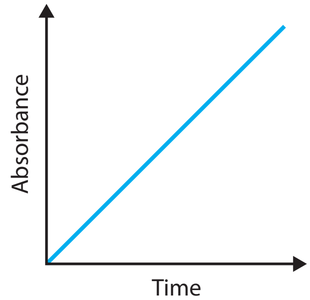
11 Carry out some research into the operation of catalytic converters in car exhaust systems. What is their purpose, and how do they utilise collision theory?

END-OF-CHAPTER EXAM

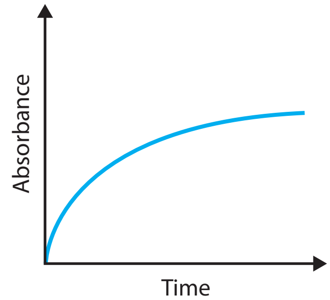
End-of-chapter test

1 Bromine liquid has a red–brown colour, and reacts with methanoic acid to form a colourless substance. The decrease in colour intensity can be readily measured by using a spectrometer. Which graph most accurately represents the changes in colour intensity as the reaction proceeds?

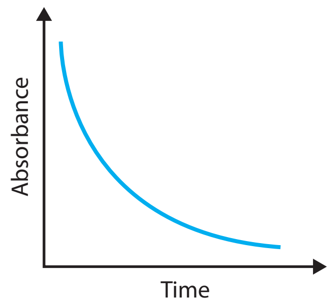
A



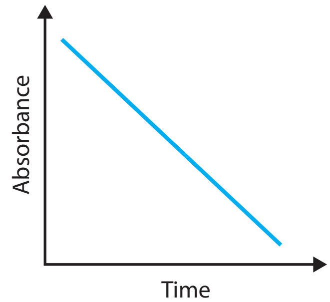
B



C



D



2 Magnesium metal reacts with dilute hydrochloric acid to produce a solution of magnesium chloride and hydrogen gas according to the following equation:

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

Which of the following steps would be least likely to increase the reaction rate?

A Using powdered zinc rather than zinc granules

B Increasing the gas pressure

C Increasing the concentration of HCl

D Using a suitable catalyst

3 An increase in temperature generally favours a higher reaction rate. Which of the following statements (i–iv) explains why this is so?

- **i** Because they are moving faster, so they collide more frequently
- **ii** Because they have more kinetic energy, so they collide with more force
- **iii** Because they are more likely to collide with the correct orientation
- **iv** Because they collide more frequently with the sides of the container

A i and iv

B All the statements are correct

C i and ii

D i, ii and iii

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4 Which of the following statements is correct, with respect to an endothermic chemical reaction?

A The rate is always faster than for an exothermic reaction.

B As the reaction proceeds, energy is released into the surrounding environment.

C The reaction causes the temperature of the surroundings to increase.

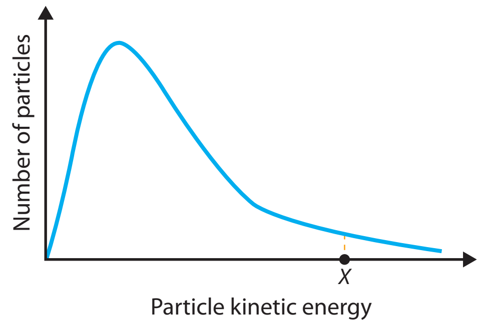
D The amount of energy required to break the bonds in the reactants is greater than the amount of energy released as the products form.

5 For a chemical reaction to proceed, particles must collide with sufficient energy and in the correct orientation. The graph below shows the range of average kinetic energies held by particles in a reaction mixture at a particular temperature.

a **i** Point *X* represents the activation energy for the reaction. Explain what this means.

ii On the axes, draw a new graph to represent the range of particle energies at a higher temperature.

iii Referring to the graph that you drew in part **ii**, explain why even a small increase in temperature can result in a significant increase in reaction rate.



b A series of experiments is conducted to investigate the reaction of barium with dilute hydrochloric acid. The equation that represents this reaction is:

$$Ba(s) + 2HCl(aq) \rightarrow BaCl_2(aq) + H_2(g)$$

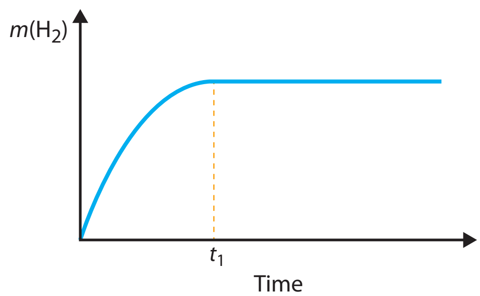
In the first experiment, 2.0 g of barium is dropped into a beaker containing 150 mL of 2.0 mol L^{-1} hydrochloric acid. The mass of hydrogen gas produced is measured at regular intervals and the results are shown on the following graph.

i At what time is the reaction proceeding at the greatest rate?

ii What event occurs at time t_1?

iii The experiment is repeated using 1.0 g of powdered barium and the results are recorded. On the axes to the right sketch the results of the second experiment.

iv List two ways (other than an increase in surface area of barium) by which the reaction rate could be increased.



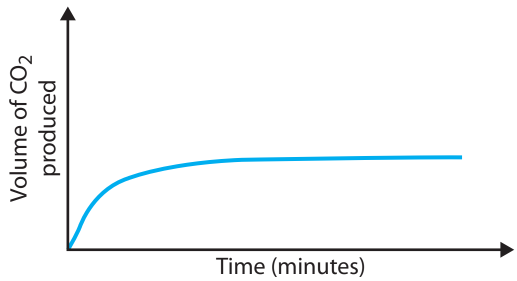
6 A series of experiments is undertaken to investigate the factors affecting the reaction rate of hydrochloric acid with marble, which is made from the chemical calcium carbonate ($CaCO_3$):

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

Flasks containing hydrochloric acid are set up under the conditions shown in the following table.

FLASK	CONCENTRATION OF HCL ($mol\,L^{-1}$)	VOLUME OF HCL (mL)	TEMPERATURE OF HCL (°C)
A	0.15	50.0	20
B	0.20	50.0	20
C	0.20	50.0	30
D	0.15	100.0	20

Each experiment is initiated by adding 0.080 g of calcium carbonate. The volume of the hydrogen gas produced is measured at SLC at set intervals. The result of the experiment for Flask **A** is shown in the following graph.



a **i** On the axes above, sketch the expected results for Flask **B** and label it 'B'.

ii On the axes above, sketch the expected results for Flask **C** and label it 'C'.

b To investigate the effect of the surface area of a solid reactant on the rate of the same reaction, a further three flasks are set up as shown in the following table. Each experiment is conducted as described in Question **6** except that the calcium carbonate is in the form shown in the following table.

FLASK	CONCENTRATION OF HCL ($mol\,L^{-1}$)	VOLUME OF HCL (mL)	TEMPERATURE OF HCL (°C)	FORM AND MASS OF $CaCO_3$
E	0.15	100	20	Powder (0.080 g)
F	0.15	50	30	Single lump (0.080 g)
G	0.20	50	20	Powder (0.20 g)

Which flask in the table above provides the most informative results on the effect of surface area on the rate of reaction when analysed together with the results from the table in Question **6**? Explain your choice.

c The number of successful collisions between the reactant particles is less than 1% of the total number of collisions. List two reasons why the vast majority of collisions between reactant particles do not produce a reaction.

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UNITS 1 & 2 PRACTICE EXAMINATION

MULTIPLE-CHOICE QUESTIONS

QUESTION 1

A substance made up of only one type of atom is called:

A a symbol.

B a nucleus.

C an electron cloud.

D an element.

QUESTION 2

Which of the following is the correct electronic configuration, using subshell notation, for silicon-14?

A $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$

B $1s^2\ 2s^2\ 2p^8\ 3s^2$

C $1s^2\ 2s^2\ 3s^2\ 2p^6\ 3p^2$

D 2, 8, 4

QUESTION 3

Which of the following statements about the ionisation energy of Al is correct?

A Al has a lower ionisation energy than Ga.

B Al has a higher ionisation energy than B.

C Al has a lower ionisation energy than Na.

D Al has a higher ionisation energy than Na.

QUESTION 4

Which of the following is correct for atomic absorption spectroscopy (AAS)?

A AAS is a quantitative technique only.

B The equipment used in AAS directly measures the concentration of an element in a sample.

C The sample being tested needs to be vaporised so that gaseous atoms are present.

D None of the above are true.

QUESTION 5

Look carefully at the graph below.

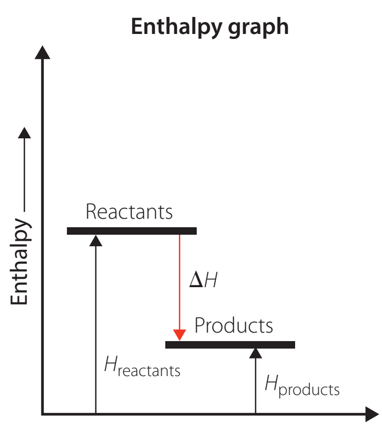
Which of the following is true?

A This represents an endothermic reaction.

B Energy is absorbed in this reaction.

C Energy is released in this reaction.

D The change in enthalpy is positive.



QUESTION 6

A student adds some white crystals to a polystyrene cup containing 100 mL of deionised water. The crystals dissolve when stirred. She uses a thermometer to measure the initial temperature of the water and final temperature of solution. They are found to be 17°C and 55°C respectively. Which of the following best describes the temperature change?

A Dissolution is an endothermic reaction.

B Dissolution is an exothermic reaction.

C The process of dissolution absorbed energy from the surroundings.

D The process of dissolution released energy. This caused the temperature of the solution to increase.

QUESTION 7

Which shows the three types of intermolecular force, in the correct order of decreasing strength?

A Covalent bond, hydrogen bond, dipole–dipole force

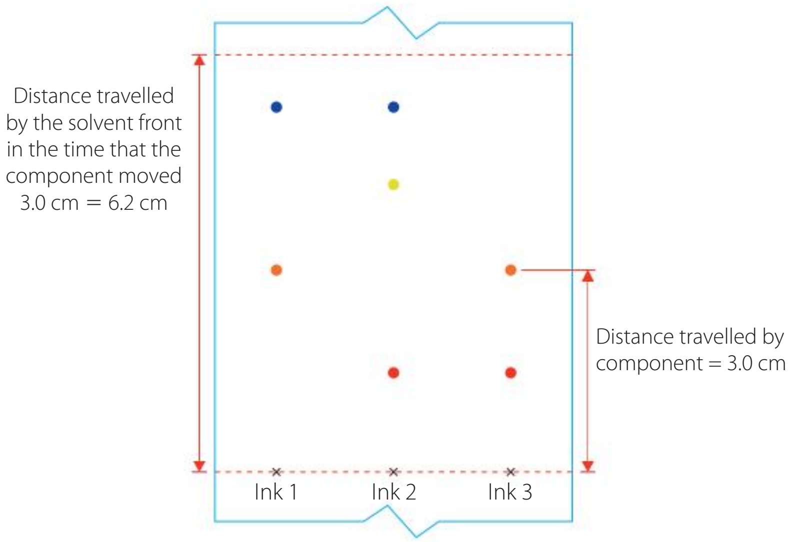
B Hydrogen bonds, dipole–dipole forces, dispersion forces

C Dispersion forces, dipole–dipole forces, hydrogen bonds

D Dispersion forces, dipole–dipole forces, covalent bonds

QUESTION 8

The figure below relates to both Questions 8 and 9.



Calculate the R_f value using the data on the diagram.

A 3.2

B 1.0

C 0.48

D 2.06

QUESTION 9

Which of the following statements is false?

A Ink 1 has two coloured components.

B Inks 1 and 2 share a coloured component.

C Ink 2 has two coloured components it does not share with any other ink.

D Inks 2 and 3 share a coloured component.

QUESTION 10

What volume of hydrogen gas is required to react with excess nitrogen to produce 15.5 L of ammonia at SLC?

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

A 7.75 L
B 15.5 L
C 23.25 L
D 31.0 L

SINGLE-ANSWER QUESTIONS

QUESTION 1

Write the correct number of protons, neutrons and electrons in an atom of sodium-24.

QUESTION 2

Calculate the molar mass of potassium hydroxide, KOH.

QUESTION 3

Which of the noble gases (group 18) has the lowest boiling point?

QUESTION 4

Convert 135K to degrees Celsius.

QUESTION 5

What is the percentage of oxygen in magnesium phosphate ($MgPO_4$)?

SHORT-ANSWER QUESTIONS

QUESTION 1

Outline how the electronegativity of the elements in the periodic table changes as we move across the table from left to right. Explain why this is so.

QUESTION 2

Explain what a polar molecule is, and justify why the molecule $CHCl_3$ is considered polar, yet CCl_4 is not polar.

QUESTION 3

Explain how atomic absorption spectroscopy can be used to selectively measure the concentration of a specific metal ion present in a mixture.

QUESTION 4

Explain how the properties of sodium oxide (Na_2O) and silicon dioxide (SiO_2) differ and explain why this is so.

QUESTION 5

Explain why increasing the temperature of the reactants can significantly increase the rate of a reaction.

HIGHER-ORDER QUESTIONS

QUESTION 1

The formulae and charges of some common ions are shown below.

chloride Cl^{1-}	sulfate SO_4^{2-}	oxide O^{2-}	phosphate PO_4^{3-}
calcium Ca^{2+}	sodium Na^{1+}	aluminium Al^{3+}	lithium Li^{1+}

a Select from the ions listed above to answer the following questions.

- **i** Write a formula for an ionic compound that will have the formula XY3.
- **ii** Write a formula for an ionic compound that will have the formula X3Y2.
- **iii** Write the formulas of two ionic compounds that have the formula XY.
- **iv** Name two ions that have the same electron configuration.

b The element magnesium has an electron configuration of $1s^2\ 2s^2\ 2p^6\ 3s^2$. Its melting point is 650°C.

- **i** Use a sketch to explain the structure of solid magnesium.
- **ii** Explain how knowledge of the electron configuration is relevant to the structure you have drawn.
- **iii** Explain what happens to a strip of magnesium when it is connected to an electric circuit and the power is switched on.

QUESTION 2

Compound A is a liquid at room temperature. It does not conduct electricity. Heated to 78°C, it turns into a gas. It can be easily condensed back to the original liquid.

a What evidence do you have that the liquid is not mercury metal?

b Provide a reason why the liquid cannot be an ionic material.

c Which forces are stronger in this compound: the forces between molecules or the forces within the molecule? Explain your answer.

d Butane and diamond both contain carbon, and both contain covalent bonds between neighbouring carbon atoms. The melting point of butane is –140°C, while diamond is still a solid at over 3000°C. Explain clearly why the melting points are so different.

QUESTION 3

Copper reacts with excess 8.0M nitric acid at SLC, producing nitrogen monoxide gas, according to the following equation:

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

a What volume of NO gas is produced at SLC when 20 g copper reacts with excess nitric acid?

b The volume of NO gas increases by 20% as the result of a temperature increase. If the gas pressure remains the same, calculate the size of this temperature rise.

c Write a balanced equation for the reaction of NO with oxygen gas to produce nitrogen dioxide (NO_2) gas.

d What mass of nitrogen dioxide occupies the same volume as the mass of NO produced in part a?

QUESTION 4

Beryllium and fluorine are located in the third period of the periodic table.

a Write the ground-state electronic configurations for:

- **i** beryllium
- **ii** fluorine.

b How does the atomic radius for fluorine compare with that of beryllium?

c The electronegativity for fluorine is 3.98. How would the electronegativity for beryllium compare with this value? Explain your statement.

d How would the first ionisation energy of fluorine compare to that of beryllium? Explain your statement.

e Write the formula of the compound that would form if beryllium and fluorine bonded together. List and describe three properties of the resulting compound.

QUESTION 5

Magnesium metal in the form of pieces of ribbon reacts with excess hydrochloric acid to form hydrogen gas and magnesium chloride.

a Write an equation to represent this reaction.

b An experiment is carried out to measure the rate of this reaction by measuring the volume of hydrogen gas produced at various time intervals during the reaction.

Time (s)	Volume of hydrogen gas evolved (mL)
0	0
2	310
4	607
6	799
8	907
10	977
12	1015
14	1020
16	1020
18	1020
20	1020

i Plot this graph on a sheet of graph paper.

ii Explain what the gradient of the graph represents and why it gradually decreases as the reaction proceeds.

c The experiment is repeated, using powdered magnesium instead of the magnesium ribbon used in the previous experiment.

i Draw a second line on the previous graph to indicate how the production would look for the second experiment.

ii Use collision theory to explain the differences between the results of the two experiments.

d Use data from experiment b to calculate the mass of magnesium that must have been used at SLC.

ANSWERS

CHAPTER 1: THE PERIODIC TABLE AND TRENDS

1.1 SECTION REVIEW

REMEMBERING

1 A substance made from one type of atom only

2 Mendeleev originally listed the elements in order of atomic mass and grouped them into rows and columns according to their properties. In the contemporary periodic table, the elements are now listed in order of atomic number.

3 A group is a vertical column of elements and a period is a horizontal row.

4 Group 1: alkali metals. Group 2: alkaline earth metals. Group 17: halogens. Group 18: noble gases.

UNDERSTANDING

5 The symbols for some elements are derived from their original name, which was not in English. The original Latin name for mercury was hydrargyrum, hence 'Hg'. The original name for tungsten was 'wolfram', derived from Swedish word for the metal, hence 'W'.

APPLYING

6 Na: sodium, Cl: chlorine and O: oxygen

7 Any group of three elements in a vertical column; for example, O, S and Se or B, Al and Ga

ANALYSING

8 Magnesium would react with chlorine to form magnesium chloride. This is because beryllium is in group 2, like magnesium, and chlorine in group 17, like fluorine. Therefore, the reactions would be very similar. Magnesium is below beryllium in group 2, therefore it is more reactive. Chlorine is lower than fluorine group 17, so it is less reactive. Therefore, it is not initially clear whether the overall reaction would be more or less vigorous.

9 Helium, neon, krypton, xenon and radon

10 Germanium

1.2 SECTION REVIEW

REMEMBERING

1 a Two
b Seven
c Five
d One
e Eight

2 The energy required to remove an electron from a gaseous atom

UNDERSTANDING

3 a 1
b 4
c 1

4 Lithium, because atomic radius decreases across the period

APPLYING

5 a Oxygen
b Beryllium
c Fluorine
d Carbon

ANALYSING

6 There is a greater number of protons in nucleus, so attraction between nucleus and electrons is greater. Thus, as you move from left to right across the period, the atomic radius decreases.

7 Although chlorine has more protons in the nucleus and is further to the right on the periodic table than oxygen, oxygen has one fewer electron shell than chlorine so its attraction for the external electrons is stronger, thus its electronegativity is greater.

1.3 SECTION REVIEW

REMEMBERING

1 An element that has some of the properties of both a metal and a non-metal

2 Metallic character decreases from left to right across the periodic table.

3 They react with water to form an alkaline solution.

UNDERSTANDING

4 It does not conduct electricity, it does not conduct heat, and it is not shiny.

5 Less vigorous: the reactivity of group 1 elements increases down the group, as the ionisation energy decreases and electrons can be more easily removed

APPLYING

6 They all have one electron in their valence shell and so lose that electron to form a positive ion

ANALYSING

7 Fluorine is to the right of the periodic table, which means it has a large number of protons in the nucleus and has only two electron shells. Therefore, there is a very strong attraction to external electrons.

8 Ionisation energy decreases down the periodic table. Group 1 elements react by losing their single valence electron and, so this is done more easily, there are a greater number of electron shells in the atoms. Thus, group 1 reactivity increases down the table.

However, group 17 reactivity decreases down the table, the reverse trend to ionisation energy. This is because the halogens react by gaining an additional electron rather than losing one.

ANALYSING

9 Magnesium is less reactive than sodium. It is further to the right in the periodic table, thus is more electronegative, with a higher ionisation energy. Therefore, it requires more energy to lose an electron, which means it reacts more slowly than sodium.

10 Tennessine will have a melting point and boiling point greater than iodine and astatine – a boiling point of possibly 300°C. It will therefore be a solid at room temperature and is unlikely to vaporise. It is likely to be toxic. It may well be coloured and may possibly be a metalloid, with some slightly conductive properties, as the metallic character increases down Group 17.

1.4 SECTION REVIEW

REMEMBERING

1 An oxide that reacts with both acid and alkaline solutions

2 As we move from left to right, the oxides become less acidic and more alkaline

UNDERSTANDING

3 Because it reacts with acid solutions and not with alkaline solutions and also dissolves sparingly in water to produce an alkaline solution

APPLYING

4 a Ionic b Covalent
c Covalent

5 a Alkaline b Acidic
c Acidic

ANALYSING

6 There is a greater proportion of alkaline oxides in period 4 than in period 3, because there are more metals – the transition metals are included in period 4. Therefore, the assumption is that they would form alkaline oxides.

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 a A sour-tasting substance that dissolves in water to produce a solution of low pH (dissociates in water, releasing hydrogen ions)
b An element in group 1 of the periodic table
c A substance that dissolves in water to produces a solution of high pH, which reacts with an acid solution
d A substance that reacts with both acids and bases
e The distance from the centre of the nucleus of an atom to the outermost electron shell
f The tendency for an atom to attract electrons from within a covalent bond
g A substance that is made from one type of atom only
h The amount of energy required to remove an electron from a gaseous atom
i A vertical column of elements in the periodic table
j An element in group 17 of the periodic table
k A substance that is shiny, malleable and conducts electricity when solid
l A substance that shows some of the properties of a metal, such as very low electrical conductivity when solid
m An element in group 18 of the periodic table
n A row of elements in the periodic table
o The section of the periodic table from groups 3–12, which consists of metals
p The number of other atoms that an atom tends to bond with

CATEGORY QUESTIONS

2 Elements are arranged in order of increasing atomic number. Elements with the same number of valence electrons are arrange in vertical columns.

3 a Increases, because the number of electron shells increases
b Decreases, because the nucleus gets further away from outermost electron shells
c Decreases, because attraction of nucleus to valence electrons becomes less as the atomic radius increases

4 a Decreases as nuclear charge increases and outer electrons are attracted more strongly to the nucleus
b Increases as nuclear charge increases and external electrons are attracted more strongly
c Increases as nuclear charge increases and electrons are more tightly bound to the atom

5 Sodium, magnesium and aluminium oxides are ionic substances with high melting points. Silicon dioxide is a covalent network with a very high melting point. Phosphorus, sulfur and chlorine oxides are small molecular substances with low melting points.

6 Sodium, magnesium and aluminium oxides are alkaline oxides. Silicon dioxide is amphoteric. Phosphorus, sulfur and chlorine oxides are acidic oxides.

ELABORATION QUESTIONS

7 Answers will vary. Elements 113, 115, 117 and 118 were named in 2016, so these may be good topics for research.

8 Electronegativity increases up and from left to right across the periodic table, so the relative position of two elements can be used to determine relative electronegativity.

EVIDENCE QUESTIONS

9 Electron shielding is the effect of inner electron shells experiencing the attraction of the nucleus to a greater extent than outer electron shells. Therefore, as the number of electron shells increases down a group of the periodic table, the atomic radius increases as the atom expands, and the first ionisation energy decreases, as the outer electrons are less tightly bound to the nucleus.

As we move across the periodic table, the atomic radius decreases, as there is no additional shielding of the outer electrons, so the increased nuclear charge causes the valence shell to be more strongly attracted to the nucleus. Therefore, the atom contracts. The first ionisation energy increases, again, as the valence electrons are more tightly bound to the atom, so require more energy to be removed.

10 The elements that he predicted the discovery of, and which were not known during his time, have indeed been discovered, largely with the properties he predicted.

Also, the grouping of elements in order of atomic number, rather than atomic mass, justifies Mendeleev's decision to swap the order of some elements, so that grouping them by properties took precedence over the atomic mass order.

11 Moseley looked at the interaction of X-rays with atoms. From the scattering patterns he observed, he could assign atomic numbers to elements, rather than atomic mass. As such, he could confirm that there were no additional elements with an atomic number less than 92 still to be discovered.

END-OF-CHAPTER EXAM

1 A
2 D
3 B
4 D
5 C
6 Greater attraction of the nucleus as number of protons increases, and no additional shielding of new electron shells, therefore outer electrons are harder to remove
7 Li, O, K, Ba
8

Radius of the K atom	>	Radius of the K^+ ion
Number of valence electrons in a Ca atom	=	Number of valence electrons in a Ba atom
Radius of an oxygen atom	<	Radius of a sulfur atom

9 a The ability of an atom to attract electrons (1) from within a covalent bond(1)

b Fluorine, element number 9

c • Relatively high nuclear charge given the small size of the atom, as it is to the right of the period
• Only two shells of electrons with little shielding, as it is at the top of the table
• Strong attractiveness to external electrons

d Data selected from three atoms in the same group, such as He, Ne, Ar, or Li, Na, K. Increased number of electron shells down the group. Decreased attraction of nucleus to external electrons due to greater shielding from the inner shells.

e The atoms with lowest electronegativity are group 18 elements (noble gases). The next noble gas will be Kr. Difference in atomic number between the noble gases increases due to the greater size of the third shell.

10a–b Up to four of the following points, (or similar) should be made:

- Mendeleev listed the elements in order of atomic number.
- He noticed that elements with the same properties reoccurred at regular intervals, although there were some apparent exceptions to this rule.
- Mendeleev was prepared to leave gaps if necessary, so that the elements lined up in columns where the elements in each column had very similar properties. He surmised that this meant that there were yet undiscovered elements.
- Mendeleev was also prepared to swap the order of some elements, according to the properties, even if this contradicted the order of increasing atomic mass.
- Mendeleev identified the patterns and trends in the existing elements and could extrapolate and interpolate as necessary to predict the properties of elements yet to be discovered according to their position on the periodic table.
- The table shows that Mendeleev's predictions were accurate and that the elements that were discovered after his death did have the properties he had predicted.

CHAPTER 2: ATOMIC STRUCTURE

2.1 SECTION REVIEW

REMEMBERING

1 Proton: positively charged. Electron: negatively charged. Neutron: uncharged.

2 Protons and neutrons are found in the nucleus. Electrons orbit the nucleus.

UNDERSTANDING

3 The protons in the nucleus repel each other because of their positive charge. The strong nuclear force holds the protons and neutrons in the nucleus together. If these two forces are balanced, the nucleus is stable.

4 Only a tiny proportion of the alpha particles were deflected by the nuclei, so they must have been very small indeed.

The alpha particles were large particles and a few were deflected back on their original path. The nuclei must have been very dense for this to occur.

The positively charged alpha particles that were passing close to the nucleus were still deflected. Therefore, the nuclei must also have been positively charged in order for this to occur.

APPLYING

5 Thomson proposed that the atom was of consistent density. No alpha particles would have been deflected, because the atom would not have had a dense nucleus, which caused the deflection.

ANALYSING

6 There are insufficient neutrons in the nucleus. This means that the size of the strong nuclear force is not sufficient to counteract the repulsive forces of the protons. Therefore, the nucleus breaks down quickly.

2.2 SECTION REVIEW

REMEMBERING

1 a The number of protons in the nucleus of an atom

b The number of protons *and* neutrons in the nucleus of an atom

UNDERSTANDING

2 The number of protons = atomic number. The number of electrons = atomic number (provided the atom is neutral and is not an ion). Therefore:

Number of neutrons = mass number − atomic number

3 Mass number is the number of protons and neutrons in the nucleus of an individual atom, which equates to the mass of an atom, because the mass of an electron is negligible in comparison with that of a proton or neutron.

In contrast, relative atomic mass is the *average* mass of one atom of an element within a sample of the element, given that, for most elements, there are several different isotopes (atoms of the same element with different mass numbers).

APPLYING

4 a Fluorine: 9 protons, 10 neutrons and 9 electrons

b Bromine: 35 protons, 45 neutrons and 35 electrons

5 a $^{27}_{13}Al$ b $^{9}_{4}Be$

6 Answers are highlighted in italics in the table below.

ELEMENT	ATOMIC NUMBER	MASS NUMBER	NO. OF PROTONS	NO. OF NEUTRONS	NO.OF ELECTRONS
Hydrogen	1	1	*1*	*0*	*1*
Magnesium	*12*	*24*	12	12	*12*
Boron	5	*11*	*5*	6	*5*
Chlorine	*17*	35	*17*	*18*	17
Nickel	*28*	59	*28*	31	28

2.3 SECTION REVIEW

REMEMBERING

1 1st shell: 2. 2nd shell: 8. 3rd shell: 18.

2 The specific arrangement of electrons in the energy levels around the nucleus of an atom, which is unique to each element

3 The Aufbau principle states that the electrons fill from the lowest energy shell first and, once a shell is at capacity, they enter the next lowest energy shell.

4 a K: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

K^+: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

b S: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

S^{2-}: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

c N: $1s^2\ 2s^2\ 2p^3$

N^{3-}: $1s^2\ 2s^2\ 2p^6$

d Ca: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

Ca^{2+}: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

UNDERSTANDING

5 a 2,2 b 2,8,6

c 2,8,8,2 d 2

6 a $1s^2\ 2s^2\ 2p^3$

b $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

c $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^1$

d $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$

e $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^2$

7 a Mg b He

c Fe

APPLYING

8 a Sodium: Incorrect because maximum of 8 electrons in the second shell, so configuration should be $1s^2\ 2s^2\ 2p^6\ 3s^1$.

b Calcium: 2 electrons enter the fourth shell before after the 3p is complete, so configuration should be $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$.

ANALYSING

9 They both have the electron configuration of $1s^2\ 2s^2\ 2p^6$. Hence being iso-electronic is to have the same electron configuration. Note: The number of protons in the nucleus of these ions is different, so the charges are different.

10 Copper has the configuration $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^{10}$, rather than the expected $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^9$. This is because the energy level of the completely filled 3d subshell is below that of the 4s, so therefore the 4s is left unfilled. For the other transition metals, with an unfilled 3d subshell, the 4s fills first.

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 a The number of protons in an atom

b The region of space around an atom that has a specific shape and may contain a maximum of two electrons

c The total number of protons and neutrons in the nucleus of an atom

d A group of electrons with the same energy level in orbit around the nucleus of an atom

e A section of an electron shell that contains orbitals of the same energy

2 Thomson proposed the 'plum pudding' model of negatively charged electrons randomly distributed in atom. Rutherford proposed a positively charged nucleus, with electrons in orbit

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around it. Bohr proposed that electrons exist in energy levels or shells around the nucleus.

3 Atomic number = number of protons (= number of electrons if the atom is neutral)

Number of neutrons = mass number − atomic number

4 In the periodic table, the s-block: groups 1 and 2, with highest energy electrons in s-orbital; the p-block: groups 13–18, with highest energy electrons in p-orbital; d-block = transition element: highest energy electrons in d-orbital; f-block: lanthanides and actinides, with highest energy electrons in f-orbital.

5 a Electrons shells fill, lowest energy first

b No two electrons can occupy the same space – the position of each electron can be uniquely described

c Electrons filling orbitals within a subshell will not pair until each orbital has a single electron

CATEGORY QUESTIONS

6

NAME	MASS	CHARGE
Proton	1	+1
Neutron	1	0
Electron	0	−1

7 a The number of electrons in the outermost shell

b The number of shells the atom has

8 Eight neutrons and seven electrons

9 a Six b Seven

c Two d Six

10 a $1s^2\ 2s^2\ 2p^1$

b $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

c $1s^2\ 2s^1$

d $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^3$

e $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5$

f $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}$

g $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$

ELABORATION QUESTIONS

11 Mass number = number of protons + neutrons in a single atom

Relative atomic mass = average mass of one atom of an element

12 $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$

EVIDENCE QUESTIONS

13 It is the first element with an electron in the d-orbital. No d-orbitals exist in the first two shells.

14 By observing the pattern from previous elements, its configuration *should* be [Ar] $4s^2\ 3d^4$. However, it is actually [Ar] $4s^1\ 3d^5$, and the 4s subshell is not filled before the 3d, due to the additional stability of the half-filled 3d subshell.

END-OF-CHAPTER EXAM

1 C 2 C

3 D 4 B

5 D

6 a $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

b $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

7 a B → A and C → D

b Positively charged, very small and very dense

c Electrons exist in energy levels surrounding the nucleus. They cannot exist between the levels but can transition from one level to another if sufficient energy is supplied. Thus, absorption spectra show that atoms absorb only specific frequencies of light, which are equivalent to the energy difference between two levels during a transition.

CHAPTER 3: INTRODUCTION TO BONDING

3.1 SECTION REVIEW

REMEMBERING

1 Ionic, covalent and metallic bonding

2 Ionic bonding occurs between a metal and a non-metal. Covalent bonding occurs between two non-metals. Metallic bonding occurs between two metals.

UNDERSTANDING

3 Chemical bonds allow the electrons to exist in a lower energy state than in the elements.

4 They have the same number of valence electrons.

5 a Ionic b Covalent

c Metallic

3.2 SECTION REVIEW

REMEMBERING

1 They are loosely held, as the metals have low ionisation energies

2 A series of positively charged ions arranged in a lattice, surrounded by delocalised electrons

UNDERSTANDING

3 They are attracted by the nuclei of the other atoms that are present and so are drawn into the space between the atoms, becoming no longer directly attached to any individual atom

4 a Metals are highly malleable and ductile.

b Metals are good conductors of electricity.

c Metals are good conductors of heat.

ANALYSING

5 Melting point appears to increase from left to right across the periodic table and is also much higher for copper, which is a transition metal. The melting point correlates to the strength of the metallic bonding that exists within the metal. The greater the number of valence electrons, the more electrons there are to contribute to the metallic bonding, thus the higher the melting point.

3.3 SECTION REVIEW

REMEMBERING

1 Electrostatic attraction between positive ions (atoms that have lost electrons) and negative ions (atoms that have gained electrons)

2 Zero: the total charge of the positive ions is exactly countered by the total charge of the negative ions.

UNDERSTANDING

3 Metals have a less than half-full valence shell, so it is energetically favourable to lose electrons. Non-metals have a valence shell that is more than half-full, so it is energetically favourable to gain electrons.

4 a Bromide Br^-

b Sulfur S^{2-}

c Barium Ba^{2+}

d Potassium K^+

e Nitrogen N^{3-}

5 There are no molecules formed, and the ionic bond is non-directional. Hence, an ionic compound is an array of positive and negative ions in fixed proportions.

APPLYING

6 a Potassium iodide

b Barium chloride

c Calcium hydride

d Lead oxide

e Sodium phosphate

f Zinc nitrate

7 a Mg_3N_2 b Al_2O_3

c HgS d NH_4OH

e $CuCO_3$ f $Zn_3(PO_4)_2$

8 a 2+ b 3−

9 a Sodium charge is 1+, bromide charge is 1−. Therefore, a ratio of 1:1 is required for an overall charge of 0.

b Aluminium charge is 3+, fluoride charge is 1−. Therefore, a ratio of 1:3 is required for an overall charge of 0.

10 a Mg^{2+} and O^{2-}

b Ag^+ and Cl^-

c Li^+ and I^-

ANALYSING

11 It has one valence electron in the first shell, so it can lose or gain an electron to form H^+ or H^-, respectively.

12 a +1 b +2

c +1 d +1

e −1 f −1

g −2

3.4 SECTION REVIEW

REMEMBERING

1 The element that is furthest to the left in the periodic table is always written first in a covalent compound.

2 a A pair of electrons shared between two atoms

b i 2 ii 4

iii 6

UNDERSTANDING

3 a Carbon: 8. Hydrogen: 2. Chlorine: 8.

b Carbon: Neon. Hydrogen: Helium. Chlorine: Argon.

4 a F_2O b BH_3

5 Carbon has 4 valence electrons, so shares one each with 4 hydrogen atoms, so that it has a total of 8 electrons in its valence shell.

APPLYING

6 a SO_2 b N_2O_5

c CCl_4 d NF_3

e $SiBr_4$

7 a Dinitrogen monoxide

b Nitrogen trichloride

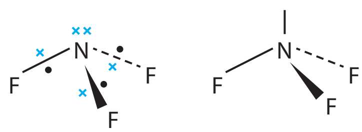
c Sulfur trioxide

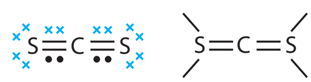
d Dihydrogen sulfide

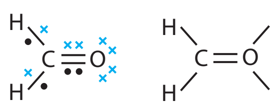
e Dinitrogen tetroxide

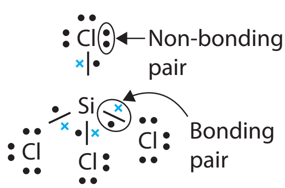
8 a 

b 

c 

d 

e 

9 

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CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 a Bonding that occurs between two or more atoms, where electron pairs are shared between the valence shells

b Electrostatic attraction between a metal atom that has lost an electron to form a positively charged ion and a non-metal atom that has gained an electron to form a negatively charged ion

c A lattice of positively charged metal cations, surrounded by a 'sea' of delocalised negatively charged electrons

2 a Example: methane

b Example: sodium chloride

c Example: iron

Diagrams will vary depending on the answers given.

CATEGORY QUESTIONS

3 a Metallic b Covalent

c Covalent d Covalent

e Ionic

4 a Silver (i) oxide

b Aluminium chloride

c Potassium sulfate

d Ammonium bromide

e Magnesium hydroxide

f Iron (ii) carbonate

5 a $MgCl_2$ b $CaCO_3$

c $Cu(NO_3)_2$ d NH_4Cl

e K_2S f PbO_2

6 a Phosphorous trichloride

b Sulfur tetrafluoride

c Dinitrogen trioxide

d Nitrogen monoxide

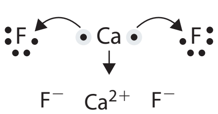
e Dichlorine heptoxide

7 a BCl_3 b PI_5

c HBr d N_2O_4

e SiO_2

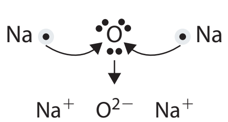
CaF_2

8 a 

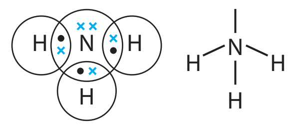
HCl

b 

Na_2O

c 

NH_3

d 

ELABORATION QUESTIONS

9 a +3 b +2

10 a −1 b −3

EVIDENCE QUESTIONS

11 We know that different elements take part in different types of bonding due to electron configuration, and how a stable electron configuration can be reached: metals lose electrons to form positive ions while non-metals attempt to gain electrons to either form a negative ion or to bond covalently.

12 Research should show that an alloy is not a compound – a no new substance is formed – the alloy has properties of its constituent elements. The properties of a compound are totally independent from that of its elements.

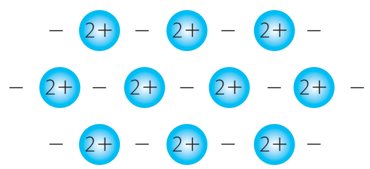
END-OF-CHAPTER EXAM

1 C 2 D

3 D 4 C

5 a i FeI_3 ii $Ca_3(PO_4)_2$

iii LiI and CaS iv Ca^{2+} and S^{2-}

b 

Lattice of positively charged Ca^{2+} ions are surrounded by delocalised electrons. Two 4s valence electrons for Ca, so two delocalised electrons per ion. Strong electrostatic attraction between ions and electrons, so a high melting point.

6

	COMPOUND A	COMPOUND B	COMPOUND C
To which group in the periodic table does X belong?	15	14	16
Suggest an identity for X.	Nitrogen	Carbon	Oxygen
Suggest a possible molecule with this structure.	Ammonia NH_3	Methane CH_4	Water H_2O

CHAPTER 4: ISOTOPES

4.1 SECTION REVIEW

REMEMBERING

1 Two atoms of the same element with the same number of protons but with different masses, resulting from their different number of neutrons

UNDERSTANDING

2 Both isotopes have seven protons and seven electrons, but nitrogen-14 has seven neutrons, while nitrogen-13 has six neutrons.

3 Isotopes of the same element have the same electron configuration, so they take part in chemical reactions in the same way.

4.2 SECTION REVIEW

UNDERSTANDING

1 Some isotopes are unstable because the strong nuclear force holding protons and neutrons together and the electrostatic repulsion between protons is not balanced, so the nucleus breaks up and decomposes.

2 Isotopes have the same chemical properties but differing physical properties. This is useful because an atom of a radioactive isotope can be substituted for another in a molecule without affecting the chemical reactions that the molecule is involved with, but it can still be detected because of its radioactivity.

4.3 SECTION REVIEW

REMEMBERING

1 The average mass of one atom of an element, relative to the mass of carbon-12

UNDERSTANDING

2 First, the masses of each naturally occurring isotope are multiplied by their abundances. Then, the sum of these products is divided by the total of the abundances.

APPLYING

3 $3 \times 12 = 36$

4 $\text{RAM} = \dfrac{(63 \times 69.17) + (65 \times 30.83)}{100} = 63.6$

5 Vanadium-51 is more abundant, because the average is closer to 51 than to 50.

6 $$\text{Relative atomic mass of V} = 50.94 = \frac{(x \times 50) + ((100 - x) \times 51)}{100}$$
$$5094 = (50x + 5100 - 51x)$$
$$x = 5100 - 5094 = 6$$
$$x = 6$$

Percentage abundance of $^{50}\text{V} = x = 6.0\%$

Percentage abundance of $^{51}\text{V} = (100 - x) = 94.0\%$

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 a The time taken for half of the nuclei in a sample to decay

b Atoms of the same element with different numbers of neutrons

c A nucleus that is unstable and decays, forming a smaller nucleus and releasing radiation

d The average mass of one atom of an element

CATEGORY QUESTIONS

2 Answers in italics

ATOM	ATOMIC NUMBER	MASS NO.	NO. OF PROTONS	NO. OF NEUTRONS	NO. OF ELECTRONS
A	6	15	*6*	9	*6*
B	*7*	15	*7*	8	7
C	12	25	*12*	*13*	*12*
D	*6*	12	6	*6*	*6*
E	*11*	*13*	11	14	*11*
F	*12*	*14*	12	16	12
G	*11*	22	*11*	11	*11*

ELABORATION QUESTIONS

3 $$\text{RAM of Si} = \frac{(92.23 \times 28) + (4.68 \times 29) + (3.09 \times 30)}{100}$$
$$= \frac{2582.44 + 135.72 + 92.7}{100}$$
$$= 28.1086 \text{ (rounded to 28.11)}$$

4
$$\text{Relative atomic mass of Br} = 79.90 = \frac{(x \times 79) + ((100 - x) \times 81)}{100}$$
$$7990 = (79x + 8100 - 81x)$$
$$2x = 8100 - 7900 = 110$$
$$x = 55$$

Percentage abundance of $^{79}\text{Br} = x = 55.0\%$

Percentage abundance of $^{81}\text{Br} = (100 - x) = 45.0\%$

EVIDENCE QUESTIONS

5 Answers will vary based on the research conducted.

END-OF-CHAPTER EXAM

1 A

2 B

3 D

4 a Two atoms of an element with the same number of protons but different numbers of neutrons

b Isotopes have the same chemical properties but differing physical properties, such as radioactivity, so they can be used in a number of applications, such as in labelling experiments.

c Answers will vary depending on the example chosen in part b.

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CHAPTER 5: ANALYTICAL TECHNIQUES

5.1 SECTION REVIEW

REMEMBERING

1 A mass spectrum

UNDERSTANDING

2 They have different masses, so they are accelerated by different amounts by the electric field and so are moving at different speeds through the magnetic field. Therefore, they will be deflected by different amounts.

ANALYSING

3 a 90, 91, 92, 94, 96

b

RELATIVE MASS	ABUNDANCE
90	52
91	12
92	18
94	15
96	3

c 91.3 – zirconium

5.2 SECTION REVIEW

REMEMBERING

1 They move into a higher energy level, further away from the nucleus.

UNDERSTANDING

2 Every element has a unique electron configuration and, for that reason, a unique sequence of absorptions that can take place.

3 Black is the absence of light, and this shows the frequencies of light that have been absorbed rather than transmitted.

5.3 SECTION REVIEW

REMEMBERING

1 An atom in the excited state will emit energy as the electrons return to their ground state. The energy emitted will be a specific frequency, equal to the energy difference between the levels.

2 An emission spectroscope uses a prism to separate the different frequencies of light emitted by a source, and then shines them onto a photographic film, so that they can be viewed. Diagrams should resemble Figure 5.3.1.

UNDERSTANDING

3 The ground state is the lowest energy configuration for the electrons in an atom. The excited state is where one or more electrons have moved into a higher energy level because of the absorption of energy. Diagrams should resemble Figure 5.2.1.

4 The configuration for every sodium atom is identical, so the same frequencies of light will be absorbed and emitted when heated.

5 Their electron configurations are different, so the energy levels are different and different frequencies of light will be emitted.

APPLYING

6 Element is sodium – the emission lines clearly match those for sodium in the standards supplied.

5.4 SECTION REVIEW

REMEMBERING

1 AAS measures the concentration of a particular metal ion. It does so by measuring a sample's absorption of light, as produced from a lamp that is made from the same metal being detected.

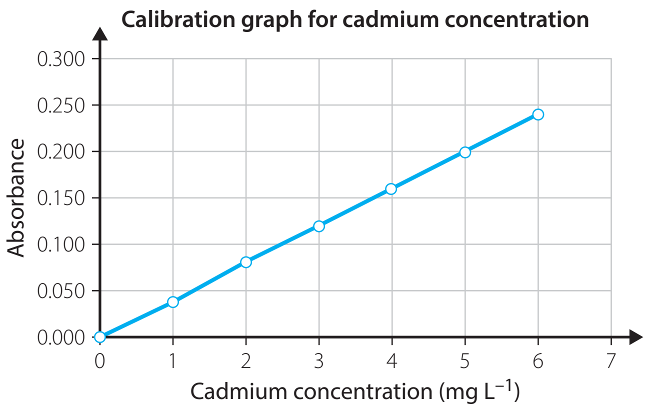
2 Measure the absorption of a series of known standard solutions. Plot a graph of absorption against concentration. Measure the absorption of the solution of unknown concentration and then use the graph to interpolate to find the concentration.

UNDERSTANDING

3 To ensure that only the required metal absorbs the light; although there may be a range of different metal ions in the sample, only the required metal will be measured.

ANALYSING

4 a



b $2.7\,mg\,L^{-1}$

REFLECTING

5 Atomic absorption enables scientists to determine the concentration of metal ions in sample solutions to a high level of accuracy. Examples can be used at student discretion.

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 a A quantitative analysis that measures the concentration of a metal ion in a solution, by measuring the absorption of light. The light is emitted from an element made from the target metal, so therefore only the absorption of the target metal is measured.

b Identifying the presence of an element from its emission spectrum – the series of frequencies of light emitted after it has absorbed energy

c Identifying a metal from the distinctive colour of the flame it emits when burning

d Ionising the particles in a mixture and separating them according to their mass. Then measuring the quantity of each particle present according to their mass:charge ratio

e Finding the identity of a substance in a sample

f Finding the amount of a substance present in a sample

CATEGORY QUESTIONS

2 a Quantitative only – more information is needed in order to determine the identity of the sample

b Qualitative only

c Qualitative only

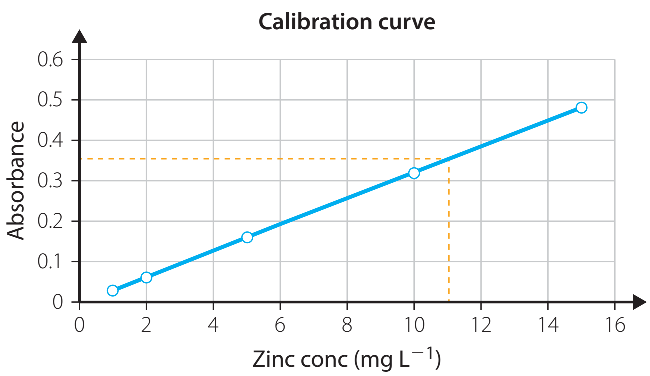
d Quantitative only

ELABORATION QUESTIONS

3 a Zinc

b The light emitted by the lamp would be the specific frequencies only emitted by zinc, so would not have been absorbed by any other metal.

c



d See red line on graph – 11 mgL^{-1}

e False – concentration appears to be higher than this from this experiment

4 a Different isotopes have different masses and therefore are accelerated by different amounts as they move through the tube. Hence, they are separated in terms of the speed at which they move.

b 5

c $\frac{m}{z} = 44$

EVIDENCE QUESTIONS

5 Helium – the frequencies emitted match those of the unknown sample

6 This is a research task, so answers will vary based on the research conducted by each student. However, lines present in the emission spectrum of light from the Sun could not be attributed to any known element at the time, and therefore was proof that a new element had been discovered.

END-OF-CHAPTER EXAM

1 B

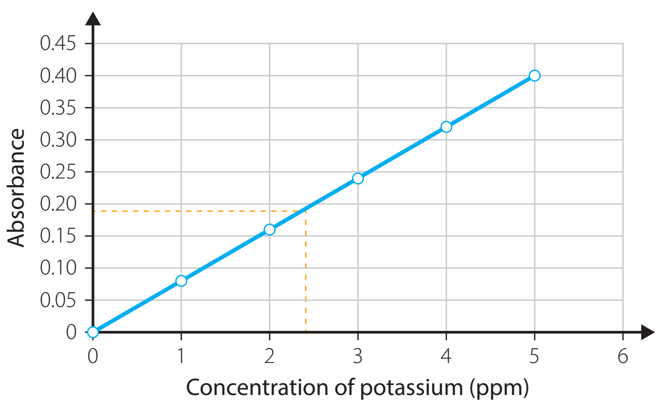
2 B

3 a $\text{RAM} = \frac{(20 \times 10) + (80 \times 11)}{100} = 10.8$

b It ionises particles and then accelerates them to different speeds. The smaller particles will move more quickly than bigger particles, so they will be separated and reach the detector at the end of the tube at different times. As the ions hit the detector, a burst of current is produced, the size of which is proportional to the number of ions in the sample. Hence the abundance can be determined.

4 a AAS can only measure the relative absorption of two different samples, so a comparison with different known standards is required.

b



(see red line) – approx. 2.4 ppm of potassium

c $\text{Concentration} = 2.4 \times \frac{100}{20} = 12\,\text{ppm}$

CHAPTER 6: COMPOUNDS AND MIXTURES

6.1 SECTION REVIEW

REMEMBERING

1 a A substance consisting of atoms of two or more elements chemically bonded together in a fixed ratio

b A substance consisting of only one type of atom with the same atomic number

c A substance that is made up of one type of particle and has a constant composition

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UNDERSTANDING

2 Helium shares the same physical property of low density with hydrogen, but helium has low flammability, while hydrogen is highly flammable.

SIMILARITIES	DIFFERENCES
Both are elements	Hydrogen is diatomic, helium is monatomic
Both are colourless, odourless gases	Helium is an inert, highly unreactive element
Both occur naturally	Hydrogen is reactive and highly flammable
Both are less dense than air	

3 The low chemical reactivity of gold, silver and copper means they will stay intact long-term, making coins and metal bars better long-term investments than paper currency, which do not share this property. However, the value of gold, silver and copper is prone to market fluctuation, while paper currency remains as printed.

ADVANTAGES	DISADVANTAGES
COINS	
Low reactivity ensures that coins made from precious and semi – precious metals will last for a long time	Expensive to produce
Coins made of metal have a high durability	Bulky
Greater exchange flexibility	
METAL BARS	
Low reactivity ensures that bars made from precious and semi – precious metals will last for a long time	Not as flexible in exchange
Bars made of metal have a high durability	Bulky
Cheaper than coins made of the same metal	
PAPER CURRENCY	
Easy and cheap to produce	Low durability
Portable	Low exchange value
Economically stable	Low intrinsic value

4 Elements that appear in nature in their elemental form often to be discovered thousands of years ago, such as those with low chemical reactivity, such as gold, silver and copper. The perfecting of chemical technologies led to discovery of elements such as hydrogen and oxygen. Other technologies, such as electrolytic techniques, beginning in the 1800s led to the discovery of more elements because elements could be isolated from each other, as with potassium and calcium. As technology has developed further (particularly nuclear technologies), elements have been synthesised by fusing elements together, such in the case of technetium and tennessine.

APPLYING

5 Answers will vary, but should include:

- Add a small amount of copper turnings or granules to a conical flask.
- Add a small amount of concentrated nitric acid to the flask. Care should be taken due to the highly corrosive nature of concentrated nitric acid. Lab coats, goggles and gloves should be worn.
- The experiment should be carried out in a fume cupboard due to the rapid evolution of toxic nitrogen dioxide gas.
- Care should be taken due to the highly exothermic nature of the reaction. Do NOT handle the flask during reaction.

6.2 SECTION REVIEW

REMEMBERING

1 a The vapour that condenses back to a liquid (solvent) and is kept after distillation

b A technique for separating the solvent from a solution, when the solvent is required to be kept

c A liquid that does not mix with another liquid

d The substance that is dissolved, or is the smaller component, in a solution

e The substance in which the solute of a solution is dissolved, or is the greater part of the solution

f A technique for separating the solvent from a solution, when the solute is required to be kept, and is based on boiling point

UNDERSTANDING

2 Both vaporisation and distillation separate the solvent from the solution.

In distillation, the technique separates the solvent from a solution when the solvent needs to be kept, and uses a round-bottomed flask over a tripod and Bunsen burner. A condenser is used so that the distillate – the solvent – can be kept. The liquid-to-gas phase change takes place at the boiling point of the liquid, and occurs throughout the whole of the liquid.

However, in vaporisation, the technique separates the solvent from the solution when the solute is what needs to be kept. It also uses a tripod and Bunsen burner, but an evaporating basin rather than a flask, and no condenser, so that the solvent is vaporised and the solute, which needs to be kept, remains in the basin. The liquid-to-gas phase change takes place below the boiling point of the liquid, and occurs at the surface of the liquid.

6.3 SECTION REVIEW

REMEMBERING

1 a A non-crystalline solid, lacking any clearly defined shape or form

b Having a non-uniform composition throughout

c Having a uniform composition throughout d An alloy where the dopant occupies the space between main atoms

e An alloy where the dopant has been substituted in the crystalline structure with the main atom

UNDERSTANDING

2 Water and salt water are both colourless, transparent liquids, while carbon dioxide gas is easily visible in soda water, which makes it easily classified as a heterogeneous mixture. However, while water is correctly classified as a pure substance, the compound H_2O, salt water's appearance is misleading, because it is actually a heterogeneous mixture, of solute salt dissolved in solvent water. Water and salt water can be distinguished also by measuring boiling points. At room temperature and pressure, water's boiling point is always 100°C. Salt water's boiling point rises as the concentration of salt in the water increases.

SIMILARITIES	DIFFERENCES
Water, soda water and salt water are all colourless liquids	Soda water and salt water are mixtures, water is a pure substance
Water and salt water are homogenous mixtures	Soda water is a heterogeneous mixture

APPLYING

3 For going out to dinner, jewellery consisting of 75% gold, 12.5% silver and 12.5% copper should be used. This makes the gold jewellery durable.

For swimming in a pool, plastic, resin, enamel or even stainless steel jewellery should be used. These substances do not react with the chlorine in the pool water. Chlorine will attack the silver and copper in gold jewellery.

6.4 SECTION REVIEW

REMEMBERING

1 a A rolled up sheet of graphene capped at one or both ends with fullerene(s)

b A mixture in which tiny dusters of particles are dispersed through another substance

c A nearly spherical arrangement of 20–84 covalently bound carbon atoms

d A sheet of covalently bound carbon atoms

e A substance that is made up of or incorporates particles in the range of 1–100 nm

f A branch of science dealing with particles in the range 1–100 nm

UNDERSTANDING

2 Graphite: Carbon atoms arranged in two-dimensional sheets with each sheet attached to others by weak intermolecular forces. High melting and boiling points, good conductor of electricity. Soft.

Diamond: Carbon atoms arranged in a three-dimensional crystal lattice. High melting and boiling points, non-conductor of electricity. Extremely hard.

Charcoal: Carbon atoms are not arranged in any discernible structure (amorphous). High melting and boiling points, non-conductor of electricity. Soft.

Fullerene: Similar in structure to graphite but are arranged in hexagonal and pentagonal rings which prevent the two-dimensional sheets being formed. Relatively low melting and boiling points. Non-conductor of electricity. Soft.

APPLYING

3 Medicinal applications: Carbon nanoparticles have extremely strong intramolecular bonds, which means they will be chemically stable in the body and their generally non-polar nature ensures they will not dissolve away in the predominantly aqueous environment of the body.

Electronics: Their high electrical conductivity and ability to be made into nano-wires have led to their incorporation in miniaturisation technologies.

4 Gold has low chemical reactivity. It does not dissolve in most acids or combine with oxygen. It does not react with the halogens, such as bromine or chlorine, very easily. This makes it harder for gold to corrode (rust) or tarnish very easily. Gold can be used as an alloy in dental fillings with little risk of corrosion from food acids ingested.

Many surgical devices, electronic equipment and life-support devices are made using small amounts of gold because it is non-reactive in the instruments; for example, it will not corrode or tarnish at contact points it can carry electricity efficiently.

Gold is considered nonallergenic, making it more suitable for implantation into the human body to treat certain medical conditions, because it is inert.

6.5 SECTION REVIEW

APPLYING

1 Repeat steps 1–15 by replacing ascorbic acid with retinol.

2 Long-chain vitamins will tend to be hydrophobic, so will lead to high-partition coefficients.

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 Hydrogen gas is highly flammable, and like any flammable gas, it is susceptible to combustion. It is also of low density, which enables it to spread and mix quickly.

2 Distillation: condenser, flask, beaker and heat source.
Vaporisation: evaporating basin and heat source (optional)

3 A chemical process occurs when a substance or substances changes chemically to produce new substances. Chemical bonds are broken and then new bonds form between different atoms, forming new substances.

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CATEGORY QUESTIONS

4 Answers will vary, but should include the following techniques: sieving, filtration, recrystallisation, distillation, vaporisation and immiscible liquid separation.

5 Places associated with the desalination of salt water are those that tend to have long coastlines.

6 Pure substances have fixed physical properties, such as melting and boiling points, density, electrical conductivity and hardness. These fixed properties are due to the fixed composition of the pure element or compound.

Impure substances, due to their varying composition, have different physical properties from their constituent components.

ELABORATION QUESTIONS

7 Answers will vary, but should include electrodes in batteries, drug delivery systems, improved tumour diagnosis, stronger materials, electronics miniaturisation, etc.

8 Rose gold is typically a substitutional alloy made with 75% gold and 23% copper, and tends to contain small amounts of silver or zinc (2%) to adjust the colour.

White gold is typically a substitutional alloy made with 90% gold and 10% of at least one white metal, usually nickel, palladium or manganese, though platinum and silver are also possible.

EVIDENCE QUESTIONS

9 Today's airships are filled with helium because it has low density (physical property) as well as low flammability (chemical property).

10 Answers will vary, but the partition coefficients should be the focus. The illness-preventative properties of vitamins are widely known. If the particle coefficients of the vitamins are explained adequately, then the relative virtues of their uptake and retention by the body can be a selling point.

END-OF-CHAPTER EXAM

1 D

2 A

3 B

4 D

5 Elements

6 Compounds

7 Mixtures

8 Graphite is carbon's most stable form, and is a crystalline allotrope. It is made up of layers of graphene weakly held together by intermolecular forces. Amorphous forms of elemental carbon turns to graphite at 2500°C and graphite turns to diamond at 100 000 atm and 3000°C. It is a form of coal. Graphene is a sheet of covalently bound carbon atoms; that is, a repeating hexagonal lattice, forming a flat sheet. Graphene is mostly used in industry producing semiconductors and electronics.

9 $P = 0.11$

CHAPTER 7: BONDING AND PROPERTIES

7.1 SECTION REVIEW

REMEMBERING

1 A type of bonding that involves the electrostatic forces of attraction between positively charged cations and negatively charged anions; results in an ionic compound with a three-dimensional lattice structure

2 High melting point, hard but brittle, and do not conduct electricity

UNDERSTANDING

3 High melting point: There is strong attraction between ions, so a lot of energy is needed to move the ions out of their fixed position in the lattice.

Hard but brittle: Ions are held rigidly in place in the lattice but the application of stress brings ions of the same charge close together so the crystal can shatter.

Do not conduct electricity: Electrical conduction depends on the presence of mobile, charged particles, such as electrons or ions. In the solid state the ions are fixed in the crystal lattice, so they are unable to move.

Mostly soluble in water: Ions are attracted to the water molecules and move out of position.

4 The empirical formula of a compound is the simplest whole number ratio of atoms present in an ionic compound. An ionic compound is a giant lattice structure composed of huge numbers of ions so a ratio of one atom to another is the simplest way of representing the compound.

5 Each ion in an ionic compound is surrounded by oppositely charged ions, so each ion can be attracted to another oppositely charged ion in any direction.

ANALYSING

6 The ionic radii of halogens increase down the group. The ionic radius of Na^+ is much smaller than that of Pb^{2+}, so the attraction for Na^+ ions for Cl^- ions is much stronger than the attraction of Pb^{2+} ions for I^- ions. This means the bonding in NaCl is much stronger than PbI_2, so much more thermal energy is required to melt NaCl than PbI_2.

7.2 SECTION REVIEW

REMEMBERING

1 Metallic lustre, good conductors of heat and electricity, malleable/ductile

2 a A substance that is hard but can be easily broken

b The ability of a substance to be drawn into a wire

c The ability of a substance to be hammered or pressed into shape

3 Density, hardness and melting point

UNDERSTANDING

4 a Metals can be drawn out into a wire because the mobile delocalised valence electrons can move to offset the positive charge build-up caused by bringing metal ions close together.

b Metals are good conductors of electricity, so a metal fork stuck into a power socket will conduct a large amount of electricity to the hand holding it.

c Metals are good conductors of heat, so a metal spoon that has been sitting in a cup of hot water will have easily conducted the heat from the water into itself and will be hot.

APPLYING

5 a Copper, being a metal, is a good conductor of electricity and can be used in electrical wiring.

b Iron is a particularly hard metal, so it is very difficult to deform the metal lattice, making it a very useful material in construction.

c Aluminium is a particularly good conductor of heat so it can be used to make a saucepan where the efficient transfer of heat from the heat source to the food in the saucepan is required.

d Gold and silver exhibit a high lustre, meaning that the delocalised electrons at the surface of these metals reflect most of the light that hits their surface. The relative non-reactivity of gold and silver make them ideal as jewellery metals.

e Tungsten, being a metal, conducts electricity and its ductility means that it can be drawn into a thin wire. This is useful in a light filament because a very thin wire has a high electrical resistance, meaning that when electricity is passed through it, the wire become so hot it glows.

ANALYSING

6 a As atomic mass increases, thermal and electrical conductivities also increase.

b As atomic mass increases from sodium to copper, the number of delocalised increases, meaning there are more delocalised electrons to conduct electricity and heat.

c As atomic mass increases, melting point increases. More importantly, as atomic mass increases, the number of delocalised electrons increases. This means that the metals that are higher in their respective groups or those that are to the left-hand side of the periodic table will have fewer delocalised electrons. The more delocalised electrons that are present, the more electrostatic attractions will be present, so the thermal energy required to melt the substance will be higher.

7.3 SECTION REVIEW

REMEMBERING

1 Covalent network solids: very high melting and boiling points, very hard, brittle solids, chemically unreactive

Covalent molecular substances: low melting and boiling points, solids are generally soft, variable reactivity

2 Group 14 (group IV) elements such as carbon and silicon are most likely to form covalent network substances due to their ability to form four covalent bonds, which means that they can form three-dimensional networks of covalently bonded atoms.

UNDERSTANDING

3 Covalent molecular substances exist as discrete molecules consisting of intramolecular covalent bonding. The molecules are held together by weak intermolecular forces. This explains why covalent molecular substances have low melting points; the solids are usually soft and they exhibit variable chemical reactivity.

However, the atoms in covalent network substances form three-dimensional networks of covalently bonded atoms in which the covalent bonding extends indefinitely throughout the whole crystal. This explains the very high melting points of these substances as well as their hardness, and their non-reactivity.

4 In graphite, the carbon atoms are arranged in flat parallel layers. Within each layer, each carbon atom is covalently bonded to three other carbon atoms, which means only three of carbon's four electrons are involved in covalent bonding. The fourth electron is delocalised, which means it is free to move along the layer occupied by its 'parent' atom. This explains the ability of graphite to conduct electricity.

In diamond, each carbon atom is bonded to four other carbon atoms, so all the electrons are involved in forming covalent bonds. There are no free electrons to conduct electricity.

5 In graphite, the carbon atoms are arranged in flat parallel layers. Within each layer, each carbon atom is covalently bonded to three other carbon atoms, so each layer is a two-dimensional network of carbon atoms. Weak intermolecular forces hold the layers together. These sheets can easily be rubbed off, giving graphite a soft, slippery feel.

In diamond, each carbon atom is bonded to four other carbon atoms. It is very difficult to dislodge these atoms from the lattice. This makes diamond a very hard substance.

6 As with graphite, fullerenes consist of carbon atoms covalently bonded to three other carbon atoms, the fourth electron of each carbon atom is delocalised. Unlike graphite, in which the carbon atoms form a hexagonal arrangement, the carbon atoms in fullerenes are arranged in hexagonal and pentagonal rings, which prevent them from forming flat sheets.

As such, fullerenes, like graphite, are soft due to the weak forces of attraction between molecules but, unlike graphite, do not conduct electricity and have relatively low melting points.

ANALYSING

7 Covalent molecular substances are discrete molecules held together by weak intermolecular forces. These require very little thermal energy to overcome. This means that these substances commonly exist as gases and the liquids can easily evaporate. As such, these substances can easily be detected as odours.

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7.4 SECTION REVIEW

REMEMBERING

1 a C_nH_{2n+2}

b C_nH_{2n}

c C_6H_6

2 Saturated compounds contain carbon atoms that are bonded to other carbon atoms by single bonds only. Unsaturated compounds contain carbon atoms that are bonded to other carbon atoms by double or triple bonds.

3 Benzene's stability is a result of the cyclic nature of the delocalised electrons within the benzene ring.

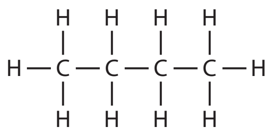
UNDERSTANDING

4 a C_6H_{12} is the alkene (hexene) – but it could be cyclohexane.

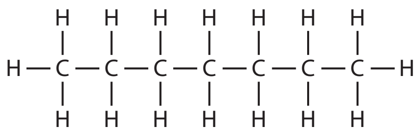
b C_6H_{14} is the alkane (hexane)

c C_6H_{10} could be cyclohexene

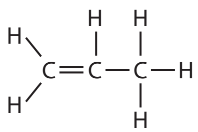
Butane

5 a 

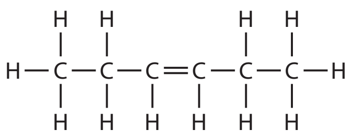
Octane

b 

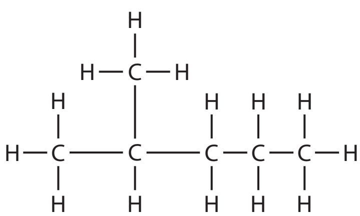
1-propene

c 

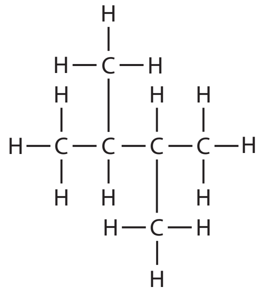
3-hexene

d 

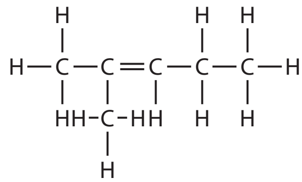
2-methylpentane

e 

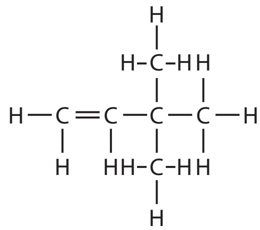
2, 3-dimethylbutane

f 

2-methyl-2-pentene

g 

3, 3-dimethyl-1-butene

h 

6 a Alkanes and alkenes are covalent molecular substances, so they share similar physical properties, such as low melting and boiling points, low solubility in water and are non-conductors of electricity. However, due to the presence of the double bond, alkenes are much more reactive than alkanes.

b There are no polar areas in a benzene molecule, so the molecule has no no hydrophilic parts that would enable it to dissolve in water.

7 a Butane

b Nonane

c 2-pentene

d 1-butene

e 3-methylhexane

f 4-ethyl-3-methylheptane

g 3-methyl-2-pentene

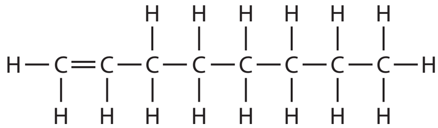
h 2,3-dimethyl-1-pentene

8 In a substitution reaction, an atom in a molecule is replaced by another atom. Alkanes typically undergo this type of reaction.

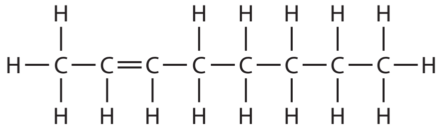
In an addition reaction, two or more molecules combine to form a larger one. Alkenes readily undergo reactions in which a molecule adds across the double bond; that is, one part of a molecule attaches to one of the carbon atoms in a double bond, while the rest of the molecule adds to the other carbon atom in the double bond. The double bond is broken, converting it into a single bond.

APPLYING

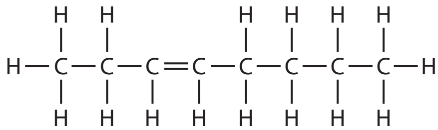
1-octene

9 

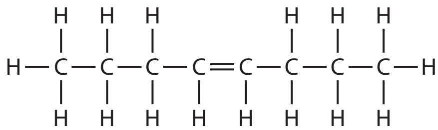
2-octene



3-octene



4-octene



10 a $C_5H_{12} + 8O_2 \rightarrow 5CO_2 + 6H_2O$

b $C_4H_{10} + Cl_2 \xrightarrow{\text{UV light}} C_4H_9Cl + HCl$

c $C_3H_6 + H_2O \rightarrow C_3H_7OH$

d $C_2H_4 + Br_2 \rightarrow C_2H_4Br_2$

e $C_6H_6 + Cl_2 \rightarrow C_6H_5Cl + HCl$

f $C_6H_6 + CH_3CH_2CH_2Cl \rightarrow C_6H_5CH_2CH_2CH_3 + HCl$

ANALYSING

11 Decolourisation of bromine indicates that a reaction has occurred. There has been no mention of the presence of UV light, so it can be concluded that the unknown hydrocarbon is an alkene.

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1

	ELECTRICAL CONDUCTIVITY	MELTING/ BOILING POINT	HARDNESS
Metallic	Good	Generally high	Generally hard
Ionic	Conduct in liquid and aqueous state only	High	Hard
Covalent molecular	Non-conductors	Low	Solids are soft
Covalent network	Non-conductors (except graphite)	Very high	Hard

2 Diamond consists of carbon atoms covalently bonded to four other carbon atoms in a giant crystal lattice. Each covalent bond requires a large amount of thermal energy to break leading to a very high melting point. Carbon atoms are dislodged from the lattice only with difficulty, making diamond a very hard substance. All electrons are involved in bonding so there are no delocalised electrons meaning that diamond does not conduct electricity.

Graphite consists of carbon atoms bonded to three other carbon atoms forming hexagonal rings arranged in large flat sheets, one on top of another. The sheets are held together by weak intermolecular forces. Each carbon atom has one spare electron not involved in bonding giving rise to many delocalised electrons giving graphite its good electrical conductivity. Each covalent bond requires a large amount of thermal energy to break leading to a very high melting point. The carbon sheets are easily rubbed off from one another giving graphite its soft slippery feel.

Fullerenes consist of carbon atoms covalently bonded to three other carbons with the fourth electron delocalised and free to move throughout the structure. The carbon atoms form pentagonal and hexagonal rings that prevent them being planar. Fullerenes are soft and slippery due to weak forces between molecules. They do not conduct electricity because there is no movement of electrons from one molecule to the next.

3 Alkanes, alkenes and benzene are covalent molecular hydrocarbons. They have low melting points, do not conduct electricity and are insoluble in water. Alkanes and benzene are relatively unreactive, undergoing combustion and substitution reactions. Due to the presence of the double bond, alkenes are reactive, readily undergoing substitution reactions.

4 A compound with the same molecular formula but a different structural formula from another isomer. Butane and methyl propane both have the same molecular formula (C_4H_{10}), but have different structures.

CATEGORY QUESTIONS

5 While phosphorus trichloride is a covalent molecular substance – low melting point, non-conductor – calcium chloride is an ionic substance, with a high melting point, and it conducts electricity when molten.

6 a Metallic

b Covalent molecular

c Ionic

d Covalent molecular

e Covalent network

f Ionic

g Covalent molecular

h Covalent network

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ELABORATION QUESTIONS

7 a Butane

b Hexane

c 1-pentene

d 2-butene

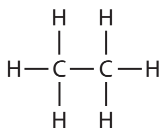
e 2-pentene

f 2,3-dimethyl butane

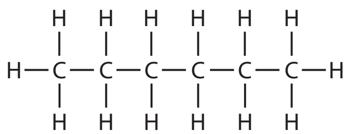
g 2-ethyl-2-hexene

h 2,3-dimethyl-2-pentene

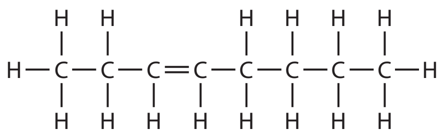
8 a **Ethane**



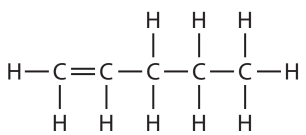
b **Hexane**



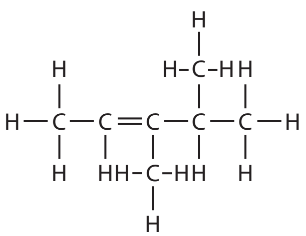
c **3-octene**



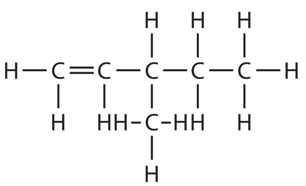
d **1-pentene**



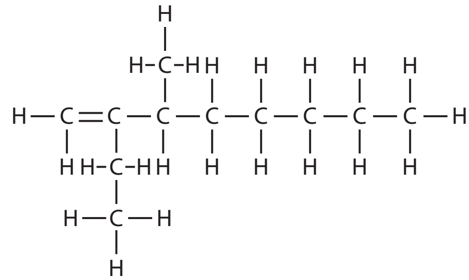
e **3, 4-dimethyl-2-pentene**



f **3-methyl-1-pentene**



g **2-ethyl-3-methyl-1-hexene**



9 a 5-hexene should be named 1-hexene. The carbon chain is numbered from the side that gives the C atom from which the double bond originates the smallest number

b 4-heptene should be 3-heptene.

10 a i $C_5H_{12} + 8O_2 \rightarrow 5CO_2 + 6H_2O$

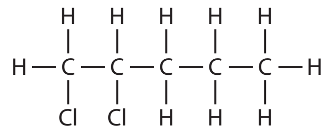
ii $2C_7H_{14} + 21O_2 \rightarrow 14CO_2 + 14H_2O$

b i $C_3H_8(g) + Br_2(aq) \rightarrow C_3H_7Br(g) + HBr(aq)$

ii $C_8H_{18}(g) + Cl_2(aq) \rightarrow C_8H_{17}Cl(g) + HCl(aq)$

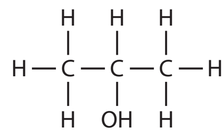
11 a $C_5H_{10}(l) + Cl_2(aq) \rightarrow C_5H_{10}Cl_2(l)$

1-pentene with Cl_2 reaction



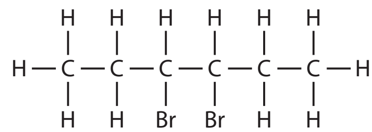
b $C_3H_6(g) + H_2O(l) \rightarrow C_3H_8O(l)$

Propene with H_2O (with H_2SO_4 catalyst)



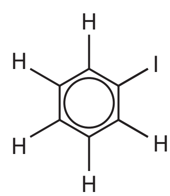
c $C_6H_{12}(l) + Br_2(aq) \rightarrow C_6H_{12}Br_2(l)$

3-hexene with bromine water



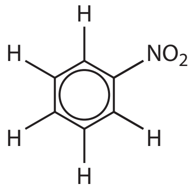
d $C_6H_6(l) + I_2(aq) \rightarrow C_6H_5I(l) + HI(aq)$

Benzene with iodine



e $C_6H_6(l) + HNO_3(aq) \rightarrow C_6H_5NO_2(l) + H_2O(l)$

Benzene with nitric acid



EVIDENCE QUESTIONS

12 a Metallic
b Covalent network
c Covalent molecular
d Ionic
e Covalent network

13 Bottle A is water, because it is solid, non-flammable and does not react with bromine water. Bottle B is hexane, because it is insoluble, flammable and reacts with bromine water. Bottle C is pentane, which is insoluble, flammable and has no reaction with bromine water. Bottle D is methylated spirits, which is soluble, flammable, and has no reaction with bromine water.

END-OF-CHAPTER EXAM

1 B
2 C
3 C
4 C
5 B

6 a Allotrope: carbon (diamond) and carbon (graphite). Isotope: carbon-12 (six protons and six neutrons) and carbon-13 (six protons and seven neutrons
b Covalent molecular: discrete molecules such as water. Covalent network: giant lattice of covalently bonded atoms, such as silicon dioxide.

7 a It does not conduct electricity
b The melting and boiling points are too low.
c Forces within molecules are stronger. Low melting and boiling points indicate that the forces between the molecules are weak.

8 a Answers may vary, but butane is a covalent molecular substance with weak intermolecular forces and little thermal energy is required to separate them. Diamond is a covalent network substance. A large amount of thermal energy required to separate all the carbon atoms.
b Alkanes are reactive due to the presence of double bonds.

9 Pentane, methyl butane, and dimethyl propane

CHAPTER 8: CHEMICAL REACTIONS

8.1 SECTION REVIEW

REMEMBERING

1 a The creation of new substances and energy transformations
b The transformation of a substance between a solid, liquid or gaseous state

UNDERSTANDING

2 a Iron + oxygen → iron(iii)oxide
$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

b Pentane + oxygen → carbon dioxide + water
$C_5H_{12}(l) + 8O_2(g) \rightarrow 5CO_2(g) + 6H_2O(g)$

In this instance, combustion, therefore water, will be gaseous till it condenses.

8.2 SECTION REVIEW

REMEMBERING

1 Chemical equations are used to summarise information about chemical reactions. These equations communicate information concisely about reactants, products and numbers of atoms of each element present.

2 Refer to Worked example 8.2.1.

UNDERSTANDING

3 a $Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$
b $2NaHCO_3(s) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2CO_2(g) + 2H_2O(l)$
c $2CuFeS_2(s) + 2SiO_2(s) + 4O_2(g) \rightarrow Cu_2S(s) + 2FeSiO_3(s) + 3SO_2(g)$
d $2PbS(s) + 3O_2(g) \rightarrow 2PbO(s) + 2SO_2(g)$
e $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$

APPLYING

4 a $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$. Note that the water will be gaseous after the reaction until it condenses, so the equation could also be written with $2H_2O(g)$ as the product.
b Balanced
c $C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2CO_2(g)$
d Balanced
e Balanced

8.3 SECTION REVIEW

REMEMBERING

1 a A reaction that occurs without the input of additional energy such as heat, light or electricity
b Corrosion is the spontaneous reaction of a metal with oxygen in a moist environment

UNDERSTANDING

2 Reactivity relates to how readily a chemical will undergo a chemical reaction, so it is clearly a chemical property.

8.4 SECTION REVIEW

REMEMBERING

1 Balanced chemical equation and net ionic equation

APPLYING

2 a $Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$
b $Mg^{2+}(aq) + CO_3^{2-}(aq) \rightarrow MgCO_3(s)$

3 a $Ca(s) + 2H^+ \rightarrow Ca^{2+}(aq) + H_2(g)$

b $Mg(s) + 2H^+ \rightarrow Mg^{2+}(aq) + H_2(g)$

c $Zn(s) + 2H^+ \rightarrow Zn^{2+}(aq) + H_2(g)$

d $Pb^{2+} + 2I^- \rightarrow PbI_2(s)$

CHAPTER REVIEW QUESTIONS

ELABORATION QUESTIONS

1 Rainwater seeps through cracks in limestone rock. Rainwater contains dissolved carbon dioxide, which gives rise to carbonic acid (H_2CO_3). This weak acid passes through joints and cracks in limestone. The mineral calcite (calcium carbonate, $CaCO_3$) is dissolved from the limestone rock to form the soluble calcium bicarbonate $Ca(HCO_3)_2$. When this seeping water containing rock is exposed to the air in the cave, it evaporates, releasing CO_2. As the carbon dioxide is released, some bicarbonate ions also dissociate to yield carbonate ions and $CaCO_3$ precipitates out of solution from the water droplets on the ceilings and from droplets on the cave walls and floors. As the re-deposited $CaCO_3$ (and small amounts other related minerals that undergo similar reactions) build up over time, stalactites are formed on ceilings of caves. The water that drops onto the floor of the cave, after evaporation and release of CO_2, forms a stalagmite.

$CO_2(g)$	+	$H_2O(l)$	→	$HCO_3^-(aq)$	+	$H^+(aq)$	CO_2 dissolving in rainwater
$CaCO_3(s)$	+	$H^+(aq)$	→	$Ca^{2+}(aq)$	+	$HCO_3^-(aq)$	Calcite dissolved
$HCO_3^-(aq)$	+	$H^+(aq)$	→	$CO_2(g)$	+	$H_2O(l)$	Evaporation yielding CO_2
$HCO_3^-(aq)$	+	$CO_3^{2-}(aq)$	→		+	$H^+(aq)$	Bicarbonate dissociation to yield $CO_3^{2-}(aq)$
$Ca^{2+}(aq)$	+	$CO_3^{2-}(aq)$	→	$CaCO_3(s)$			Re-precipitation of calcite to form stalagmite or stalactite

EVIDENCE QUESTIONS

2 a $C_6H_{12}(l) + 9O_2(g) \rightarrow 6CO_2(g) + 6H_2O(g)$

b $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

c $C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$

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

b $Mg(OH)_2(aq) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$

c $Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$

4 a $2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$

b $Cu(s) + 4H^+(aq) + 2NO_3^-(aq) \rightarrow Cu^{2+}(aq) + 2NO_2(g) + 2H_2O(l)$

END-OF-CHAPTER EXAM

1 D

2 A

3 D

4 C

5 B and D are false

6 D

7 B

8 Au

9 K

10 C

11 C and D

12 $2Cu(NO_3)_2(s) \rightarrow 2CuO(s) + 4NO_2(g) + O_2(g)$

13 a $2NaOH(aq) + MgCl_2(aq) \rightarrow 2NaCl(aq) + Mg(OH)_2(s)$

b $2Na^+(aq) + 2OH^-(aq) + Mg^{2+}(aq) + 2Cl^-(aq) \rightarrow$
$2Na^+(aq) + 2Cl^-(aq) + Mg(OH)_2(s)$

c $Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$

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

b $2FeSO_4(s) \rightarrow Fe_2O_3(s) + SO_2(g) + SO_3(g)$

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

d $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$

e $6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$

CHAPTER 9: EXOTHERMIC AND ENDOTHERMIC REACTIONS

9.1 SECTION REVIEW

REMEMBERING

1 a A reaction in which energy is absorbed from the surroundings, causing a temperature decrease

b A reaction in which energy is released to the surroundings, causing a temperature increase

2 a True b False

APPLYING

3 a $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$ $\Delta H = -242$ kJ

b $2NH_3(g) \rightarrow 3H_2(g) + N_2(g)$ $\Delta H = +92$ kJ

c $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ $\Delta H = +180$ kJ

d $4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$ $\Delta H = -905$ kJ

e $H_2O(l) \rightarrow H_2O(s)$ $\Delta H = -6.02$ kJ

9.2 SECTION REVIEW

REMEMBERING

1 a Also known as bond enthalpy; the amount of energy required to break one mole of the stated bond to give separated atoms, usually expressed in units of kJ mol^{-1}

b A measure of the average kinetic energy of the particles that constitute a solid, liquid or gas

c Temperature (°C) = temperature (K) −273. Absolute zero is the lowest temperature that is theoretically possible, at which the motion of particles, which constitutes heat is at a minimum. It equals −273.15°C.

UNDERSTANDING

2 a 298 K b 196 K

3 a 362°C b −135°C

4 Translational, rotational and vibrational kinetic energy contribute to the total kinetic of a molecule. Translational kinetic energy (KE_T) is given by $KE_T = \frac{1}{2}mv^2$. Rotational kinetic energy (KE_R) is the kinetic energy contained in rotational motion, determined by the mass and angular velocity. Vibrational kinetic energy (KE_V) is associated with the oscillation of atoms within a molecule or solid lattice as they vibrate. Even at absolute zero (0K), there is vibrational energy, termed 'zero-point vibrational motion' or Brownian motion.

5 In solids, which are tightly bound lattices of atoms or molecules, essentially all the *KE* is contained in vibrational *KE* as the atoms oscillate backwards and forwards.

In liquids, which are more 'ordered', much of the *KE* can be due to translational motion, but *KE* from both rotational and vibrational motion also contribute.

For gases, most of their *KE* is due to random translational motion but for molecular gases, rotational and vibrational motion will also contribute.

APPLYING

6 292.2 K

9.3 SECTION REVIEW

APPLYING

1 16 720 J or 16.72 kJ

2 35 112 J or 35.11 kJ

9.4 SECTION REVIEW

APPLYING

1 a 1476 J or 1.48 kJ

b 39.1 g

9.5 SECTION REVIEW

REMEMBERING

1 a An enthalpy change accompanying a chemical change is independent of the route by which the chemical change occurs

b Exothermic chemical reaction between a fuel and an oxidising agent (usually oxygen) that produces energy

c Heat released when a fuel undergoes complete combustion at standard atmospheric pressure

d Atmospheric pressure on Earth's surface, described as 1 atm, or 101.325 kPa

APPLYING

2 −511 kJ

3 −746.6 kJ

4 +51 kJ

5 −1789 kJ

9.6 SECTION REVIEW

APPLYING

1 Heat absorbed by water = 12 540 J or 12.54 kJ. So, if 0.372 g of 1-propanol released this quantity of heat then 1 g of 1-propanol would release $\frac{1}{0.372} \times 12.45 = 33.71$ kJ.

2 a 30.44 kJ g^{-1}

b The main cause of discrepancy will derive from the loss of heat from the spirit burner to the surroundings. In other words, not all the heat released by the burning butanol will be absorbed by the water in the calorimeter. This may be due to air currents disturbing the flame as well as poor insulation of the calorimeter leading to its own heat loss to the surroundings, which yields a lower measurement for the temperature rise of the water. This will yield a lower measured estimate for ΔH_c for the butanol sample.

3 −68.9 K

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 a The total energy possessed by a chemical substance (changes as bonds are broken or made during reactions)

b A measure of the average *KE* of the particles in a system

c Energy is absorbed from the surroundings, causing a temperature decrease

d The heat needed to increase the temperature of 1 g of a substance by 1 K

e *KE*, the energy possessed by moving particles: a particle of mass *m*, moving with speed *v* as a translational kinetic energy equal to $\frac{1}{2}mv^2$, molecules also possess rotational and vibrational kinetic energy

f Also known as bond enthalpy, the amount of energy required to break 1 mole of the stated bond to give separated atoms, usually expressed in units of kJ mol^{-1}

g An enthalpy change accompanying a chemical change is independent of the route by which the chemical change occurs

h Energy is released to the surroundings, causing a temperature increase

CATEGORY QUESTIONS

2 Temperature is a measure of heat energy. A change in temperature occurs when heat is gained or lost. In terms of atoms and molecules, the kinetic energy that we can associate with temperature changes are contained in translational, rotational and vibrational motion. Translational

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motion refers to atoms and molecules moving through space. This occurs at a variety of speeds. Temperature provides a scale for describing the average total *KE* in a system.

3 Temperature (°C) = temperature (K) – 273

4 a Exothermic: heat is lost to the surroundings as the water temperature decreases

b Endothermic: heat is absorbed and employed in providing the latent heat of sublimation, which converts solid CO_2 to gaseous CO_2

c Exothermic: $\Delta H = -58$ kJ is a negative quantity, hence heat is lost to the surroundings

d Photosynthesis: the reaction can be summarised as:

$$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2 \quad \Delta H = +2803 \text{ kJ}$$

$\Delta H = +2803$ kJ is positive, so the reaction is endothermic.

ELABORATION QUESTIONS

5 Biodiesel is a renewable fuel consisting of long-chain organic molecules called fatty acid methyl esters, containing carbon, hydrogen and oxygen; diesel is a product of crude oil refining, and is a mixture of hydrocarbons (carbon, hydrogen only) that typically contain between 8 and 21 carbon atoms per molecule.

Since diesel and biodiesel contain very different mixtures of chemical compounds, each with their own specific heats of combustion, the measured heat of combustion for them will be different.

6 Answers will vary.

EVIDENCE QUESTIONS

7 Answers will vary.

8 Answers will vary.

END-OF-CHAPTER EXAM

1 B

2 C

3 C

4 A

5 A

6 C

7 A

8 a $2Cu(NO_3)_2 \rightarrow 2CuO + 4NO_2 + O_2$

b Endothermic, because the reaction requires strong heating for it to proceed. In addition, if the enthalpy of reaction is calculated from heats of formation, one obtains $\Delta H = +220$ kJ mol^{-1}.

9 Refer to Figure 9.1.3 from Chapter 9.

10 a Specific heat capacity is the amount of heat energy required to raise the temperature of 1 gram of a substance by 1 K.

b 4.18 J K^{-1} (Water)

11 143.6 kJ

12 –5471 kJ mol^{-1} (exothermic reaction)

13 Fractional numbers of moles make the calculation easier. This is a valid strategy to use, so write the contributing reactions as follows. Note that signs for enthalpy changes are reversed when the reactions are reversed:

$2O_3 \rightarrow 3O_2 \quad \Delta H = -427$ kJ

Reverse the reaction and divide by 2:

$$\frac{3}{2}O_2 \rightarrow O_3 \qquad \therefore \Delta H = \frac{1}{2} \times -(-427) \qquad = 213.5 \text{ kJ}$$

$$NO + O_3 \rightarrow NO_2 + O_2 \qquad \Delta H = -199 \text{ kJ}$$

$$O_2 \rightarrow 2O \qquad \Delta H = -427 \text{ kJ}$$

Reverse the reaction and divide by 2:

$$O \rightarrow \frac{1}{2}O_2 \qquad \therefore \Delta H = \frac{1}{2} \times -(+495) \qquad = -247.5 \text{ kJ}$$

$$NO + O \rightarrow NO_2 \quad \Delta H = 213.5 - 199 - 247.5 \qquad = -233 \text{ kJ}$$

CHAPTER 10: SCIENTIFIC MEASUREMENTS – UNCERTAINTY AND ERROR

10.1 SECTION REVIEW

REMEMBERING

1 a Using a measurement procedure together with a measuring standard to determine a measurement

b A defined strategy for making a measurement using a measuring device

c A calibrated device for making measurements

UNDERSTANDING

2 Calibration is the procedure of comparing a measurement device with a measurement standard (usually a secondary standard), to establish the precision and accuracy of the measuring device.

3 Secondary standards are measuring devices calibrated against a primary standard and which provide a convenient basis for calibrating other devices. Primary standards are maintained by government institutions. For example, the National Measurement Institute is responsible for maintaining Australia's official standard of length.

10.2 SECTION REVIEW

REMEMBERING

1 a The closeness of agreement between the average value obtained from a series of test results and the accepted true value

b The closeness of agreement between independent measurements

c Uncertainty characterises the range of values within which the true value is asserted to lie with some level of confidence

d When a number of measuremenets of the same quantity are repeated, they will lie within a range determined by the measurement uncertainty. if the range is within an acceptable uncertainty interval, the measurements are considered repeatable.

UNDERSTANDING

2 When you take repeat measurements, the best estimate of the measured quantity is the average value. If you have taken fewer than 10 measurements, then the best estimate of the uncertainty is half the range. If you have more than 10 measurements, then the best estimate of the uncertainty is the standard deviation.

3 Uncertainty identifies the range of values within which the true value of a measurement is asserted to lie with some level of confidence (e.g. X ± y kg). Error is the difference between the true value of the measured quantity and the measured value (which is an estimate for the true value).

APPLYING

4 Mean = 1.10, standard deviation = 0.05

10.3 SECTION REVIEW

REMEMBERING

1 a Numerical data acquired through counting or measuring

b Data described by descriptive values that cannot be specified using a numeric scale

c Variables that have an infinite number of possible values within a selected range

d Variables that represent qualitative information that can be categorised into a classification or is based on counts, with a finite number of possible values; such values cannot be meaningfully subdivided

UNDERSTANDING

2 Quantitative data provide numerical measures of quantities that can be manipulated using mathematical operations. The data can be graphed and functional analyses applied to provide the means to generate mathematical relationships. This can offer quantitative predictions from trends that might be observed.

10.4 SECTION REVIEW

REMEMBERING

1 a A variation that affects a measurement in a random way so that the measurement is as likely to change in any one direction as in any other

b An error that acts to give a consistent offset in data; for example, a zero error

UNDERSTANDING

2 Human error is a term that describes a mistake. A mistake can and should be rectified. However, even mistake-free lab measurements have an inherent uncertainty or error.

10.5 SECTION REVIEW

REMEMBERING

1 a All the non-zero digits of a number, and the zeros that are included between them, or that are final zeros that signify accuracy

b The numbers after a decimal point and before the first non-zero integer. For numbers lower than the whole numbers (that is, numbers containing a decimal point), we must use zeroes as place holders after the decimal point. For instance, the fraction $\frac{9}{10}$ is written in decimal form as 0.09. If the zero is not included between the dot and the 9, the number would be $\frac{9}{10}$, which is not the fraction $\frac{9}{100}$. The number of decimal places is the number of digits, i ncluding the place holders, to the right of the decimal point.

UNDERSTANDING

2 The number of significant figures that should be displayed when quoting a measurement requires estimating the uncertainty and estimating the consequences of any error propagation through calculations.

10.6 SECTION REVIEW

REMEMBERING

1 a The magnitude of the difference between the observed/measured value and the true value

b The spread or interval of measured values within which the true value is expected to lie

c $\text{percentage error} = \frac{\text{accepted value} - \text{measured value}}{\text{accepted value}} \times \frac{100}{1}\%$

d $\text{percentage uncertainty} = \frac{\text{absolute uncertainty}}{\text{measured quantity}} \times \frac{100}{1}\%$

UNDERSTANDING

2 a–b The number of significant figures that should be displayed when quoting a measurement requires estimating the uncertainty and estimating the consequences of any error propagation through calculations.

10.7 SECTION REVIEW

REMEMBERING

1 a A graph that uses horizontal (x) and vertical (y) axes to plot a set of data points represented by (x, y), without a line joining the points; used to demonstrate or determine a mathematical relationship between variables

b A data point that does not fit the pattern shown by other measured data points; sometimes they are defined quantitatively as lying more than 1.5 times the interquartile range above the third quartile or below the first quartile

c Reading or constructing a new data point that has not been measured but is within the range of measured data

d Estimating a value for a data point that lies outside the range of measured values using a mathematical relationship based on the measured data, or extending a perceived sequence of values, and assuming that the trend identified will continue

UNDERSTANDING

2 For a straight-line relationship, a minimum of three data points is required.

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3 Six data points are considered a bare minimum. More are preferable.

4 Mean, standard deviation, fitting of trendlines and establishing uncertainties in trendlines and associated parameters are best established using technology available on graphics calculators or computer spreadsheets. Standard deviation is a preferred method of estimating the statistical spread of data when approximately 10 or more data points are available.

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 a 8.33 m

b 40.68 g

c 6.022×10^{22} g

d 5.5 g

e 8.0252 cm

f 16.397 s

2 Absolute error = 0.11 g. Percentage error = 5.61%

3 a ± 2.6%

b Consistent

c 2.4%

4 a Enthalpy (ΔH): experimental result agrees (within certainty) with accepted value. Time: experimental result disagrees with accepted value. Mass: experimental result disagrees with accepted value.

b % Uncertainties: ΔH: ±0.5%. t: ±0.4%. m: ±2.9%

c Systematic error

d % error ΔH: +0.42%. t: +2.3%. m: +5% (but well within uncertainty for accepted value)

5 a Quantitative, continuous

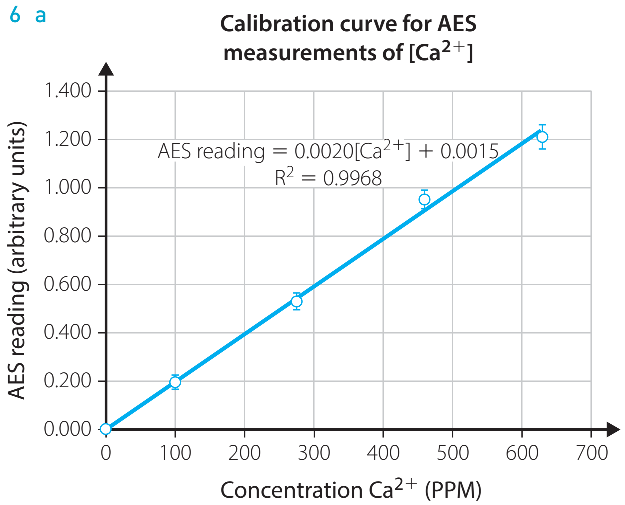
b Quantitative, continuous

c Quantitative, discrete

d Qualitative

e Quantitative, discrete

6 a



From LINEST: slope = 0.0020 ± 0.00006
intercept = 0.0015 ± 0.02

b AES reading = 0.0020 [Ca^{2+}] + 0.0015, so:

$$[Ca^{2+}] = \frac{\text{AES reading} - 0.0015}{0.0020}$$

This means that the [Ca^{2+}] in the river water sample is:

$$[Ca^{2+}] = \frac{0.832 - 0.0015}{0.0020}$$
$$= 415 \text{ PPM}$$

c Uncertainty in slope = $\pm \frac{0.00006}{0.0020} \times 100\% = 3$

The intercept (0.0015) is small compared with the AES reading of 0.832. However, it has a high absolute uncertainty (0.02), which will not contribute significantly to the uncertainty in the interpolated value. Taking the upper and lower possible values for the difference (AES reading) − (0.0015±0.02), the following is obtained:

$$[Ca^{2+}] = \frac{0.832 - 0.0015 \pm 0.02}{0.0020}$$
$$= 425 \text{ (upper value) and } 405 \text{ (lower value)}$$

This corresponds to percentage uncertainty of ±2.4%, making the total uncertainty:
$\pm(3 + 2.4 = 5.4)$ % of the interpolated value (415 PPM); that is:

$$\frac{(5.4 \times 415)}{100} = 22 \text{ PPM}$$

So, the total concentration of the river water sample is reported as: 415 ± 22 PPM.

7 a Mean value = 0.322 mol L^{-1}

b No (refer to answer for part c for the justification)

c Standard deviation (σ) = 0.011; all data lie within 2σ of the mean, so no data points can be classified as outliers. The two values 0.308 and 0.309 could possibly be classified as anomalously low relative to the other values, but the sample set is too small to justify their exclusion.

d Given an uncertainty of 0.011, or rounded to 0.01, it would be sufficient to quote the readings with two significant figures only.

e With no information concerning the instrument's calibration, systematic errors may be present. The observed spread of values is likely to due to random uncertainty.

END-OF-CHAPTER EXAM

1 B (accuracy)

2 B (precision)

3 D (precise but not accurate)

4 a 1 cup ± 1 tablespoon

b 1 m^2 ± 1 cm^2

c 1 day ± 3 hours

d 395.2 ± 0.5 kg

5 a 4
   b 3
   c 4
   d 4
   e 3
   f 4
   g 3
6 a 0.000355
   b 0.0000753
   c 9.09×10^6
7 a Quantitative, continuous
   b Qualitative
   c Quantitative discrete
   d Quantitative, continuous
   e Quantitative, continuous
   f Quantitative discrete
   g Quantitative, continuous
   h Quantitative discrete
8 a Second student states uncertainty correctly; first statement should be 0.77 ± 0.01 g
   b First student's measurement
   c Calibration error (systematic error)
   d No
9 a 0.389
   b ±0.0005 JK–1 g^{-1}
   c ±0.024, hence estimate of specific heat should be quoted as 0.39 ±0.03 JK-1g-1 since only two significant figures are justified and it is preferable to round up the uncertainty rather than round down.
   d Standard deviation is calculated as σ = ± 0.017. This can be rounded up to 0.02 and hence the estimate of specific heat may be quoted as 0.39 ±0.02 $JK^{-1}g^{-1}$ based on the standard deviation σ.
   e The uncertainty based on range is preferable if the sample set is small. For sample sets of 10 or more values, standard deviation σ, provides a statistically defensible method of reporting uncertainty more meaningfully.
   f Given the uncertainty; that is, specific heat = 0.39 ±0.03 $JK^{-1}g^{-1}$, i.e. a range 0.36–0.42, the reported specific heats of zinc, copper and brass are all consistent with the measured value. The specific identity of the metal cannot be established.
   g Iron and nickel may be excluded as possible identities of the metal since their reported specific heats are inconsistent with the measured value.
   h Random errors include weighing, temperature change measurements (initial and final temperatures), allowing sufficient time for temperature equilibration before taking final temperature reading, heat loss during transfer of heated metal to calorimeter, imperfect insulation of calorimeter (hence heat loss).
   Systematic errors include calibration errors in mass balance, thermometer, and systematic flaws in procedures.

CHAPTER 11: FUELS

11.1 SECTION REVIEW

REMEMBERING

1 a An energy source, such as crude oil and coal, produced from the decayed remains of animals and plants
   b Unable to be regenerated at a rate similar to which it is consumed
   c A reaction with oxygen to form the oxides of each of the elements present; with adequate oxygen, combustion of a hydrocarbon will produce carbon dioxide and water
   d Heating of Earth's near-surface atmosphere by the trapping of out-going infrared radiation
   e Gas that traps outgoing infrared radiation

UNDERSTANDING

2 Coal: carbon dioxide, sulfur dioxide. Oil: carbon dioxide, nitrogen dioxide. natural gas: carbon dioxide
3 Energy–type trends moved from heat and light to locomotion (steam) to locomotion (internal combustion engine) to electricity generation.

APPLYING

4 Answers will vary.

11.2 SECTION REVIEW

REMEMBERING

1 a Splitting of heavy nuclei
   b Joining of small nuclei

UNDERSTANDING

2 Advantages: Low CO_2 emissions have less of an impact on the greenhouse effect and it is more capable of meeting modern energy demands. Disadvantages: There is a risk of radiation contamination during operation, high initial cost and the waste products are dangerous.
3 Advantages: There are virtually no waste products (He), the fuel is limitless and, once operational, it is cheap to run. Disadvantages: the set-up costs are extremely high, the temperatures generated are also extreme and difficult to contain, and the technology has not yet been perfected.

APPLYING

4 Answers will vary.

11.3 SECTION REVIEW

REMEMBERING

1 a A fuel that is primarily made of biomass
   b Types of energies that have the potential to be replenished and are essentially limitless

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2 Supplementing coal powerplants involves using renewable energy sources such as wind, solar, hydroelectric power to supply energy in addition to the coal. Replacing coal powerplants involves removing the plant and replacing with renewable energies of even other fossil fuel-powered plants such as oil and natural gas.

APPLYING

3 Answers will vary greatly. This is more of a research task. Finding out the prevailing weather conditions in particular countries, and so on.

11.4 SECTION REVIEW

APPLYING

1 Precision: The trials should be consistent ($\pm 0.05\,\text{kJ}\,\text{g}^{-1}$)

Accuracy: How close the experimental answer is to the literature value.

The percentage error should be calculated:

$$\text{Percentage error} = \frac{\text{literature value} - \text{experimental value}}{\text{literature value}} \times 100$$

2 Modifications could include:

Using insulation (and a heat-proof material) around the outside of the soft-drink can

Put a lid (with a hole to allow access by the temperature probe) on the soft-drink can

Lowering the soft drink can so that it is just above the spirit burner flame

Any modifications should address the loss of heat from the flame and water in the soft drink can to the surroundings.

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 There are three main dangers: in extraction, emission of alpha radiation, which can enter water and agricultural systems, and emission of gamma radiation. In waste storage, the waste products are toxic for many years so storage locations need to be geologically and politically stable. Finally, the potential for meltdown of fission reactors is of concern, because it can cause radiation fallout over a large area, and local areas may remain toxic for many years.

CATEGORY QUESTIONS

2 Experimentally: fuel burner, heat-conducting containers of water, thermometer, and mass balance

Industry: Bomb calorimeter, which is a modification of the equipment listed above where the fuel is burnt in a sealed container, with none lost to the environment

3 a $C(s) + O_2(g) \rightarrow CO_2(g)$

b $2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(l)$

4 $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$

5 a–c Answers will vary.

6 Coal, natural gas, oil and ethanol

ELABORATION QUESTIONS

7 Deuterium and tritium are used in nuclear fusion reactions.

8 Low density and high flammability

EVIDENCE QUESTIONS

9 Electric trains are more efficient and produce fewer polluting gases.

10 Answers will vary but should include: supplements, because it is relatively easy to supplement with other forms of energy. Replacements are possibly unrealistic and very expensive.

END-OF-CHAPTER EXAM

1 D

2 A

3 D

4 C

5 Non-renewable

6 Renewable

7 Large, fission

8 Small, fusion

9 The Great Smog of 1952 resulted from burning coal, which produced sulfur dioxide. Photochemical smog is due to the burning of petrol, which produces nitrogen dioxide.

10 The natural greenhouse gas effect caused by greenhouse gases such as water vapour, carbon dioxide, methane and ozone, which occur due to natural biological and physical processes.

Anthropogenic greenhouse effect is due to the production, by humans, of greenhouse gases such as carbon dioxide.

11 $5566\,\text{J}\,\text{g}^{-1}$

CHAPTER 12: MOLE CONCEPT AND THE LAW OF CONSERVATION OF MASS

12.1 SECTION REVIEW

REMEMBERING

1 a $6.02214078 \times 10^{23}$ (usually approximated as 6.022×10^{23}) – the number of atoms in exactly 12 grams of the carbon-12 isotope

b The amount of a substance containing N_A particles of that substance

UNDERSTANDING

2 The relationship between the number of moles (n) of a substances and the number of particles (atoms, ions or molecules) is given by:

$$\text{number of moles}(n) = \frac{\text{number of particles}}{\text{number of particles per mole}} = \frac{\text{number of particles}}{N_A}$$

APPLYING

3 a 6.64

b 0.017

c 0.20

d 0.014

4 a 1.51×10^{24}

b 1.99×10^{23}

c 9.64×10^{23}

d 2.11×10^{22}

5 a 3

b 3

c 9

d 5.42×10^{24}

6 a 498

b 3985

c 8967

ANALYSING

7 a 10.0

b 0.40

c 45.1

12.2 SECTION REVIEW

REMEMBERING

1 The relationship between the relative quantities of substances taking part in a reaction or forming a compound, typically a ratio of whole integers

2 Mass is neither created nor destroyed by chemical or nuclear reactions, radioactive decay or by physical transformations

3 Chemical reactions must obey the law of the conservation of mass. This means that, in chemical reactions, the combined mass of the products must equal the combined mass of the reactants. Atoms cannot be created or destroyed in a chemical reaction, so the number of individual atoms must be the same on both sides of a chemical equation (balanced in terms of the number of atoms).

UNDERSTANDING

4

FE_2O_3	CO	FE	CO_2
1	3	2	3
1.5	4.5	3	4.5
0.2	0.6	0.4	0.6
0.03	0.09	0.06	0.09

12.3 SECTION REVIEW

REMEMBERING

1 The relative molecular mass (atomic mass units) is calculated by adding the relative atomic masses of all the component atoms of the molecule. The mass of 1 mole of a substance is called the molar mass (M). The mass of 1 mole of a particular element or compound is numerically equal to its relative atomic, molecular or formula mass, but the units of molar mass are grams per mole (g/mol or g mol^{-1}).

2 Empirical formula is the chemical formula showing the simplest ratio of elements in a compound, while the molecular formula specifies the total number of each of the atoms that are contained in the molecule.

APPLYING

3 a 34.02 g mol^{-1}

b 184.11 g mol^{-1}

c 100.08 g mol^{-1}

d 72.10 g mol^{-1}

4 a 23.8 g

b 1125 g

c 646 g

d 148.8 g

e 95.9 g

5 a $COCl_2$

b C_2H_5

c NO_2

d Na_2PO_3

6 a 0.555

b 0.167

c 0.333

d 0.0975

e 0.0579

7 ClO_2

8 a C_5H_7N

b $C_{10}H_{14}N_2$

ANALYSING

9 a 294.30 g

b 0.0211 mol

c 72.10 g

d 8.76×10^{18} molecules

10 a 1.51×10^{23} molecules

b 3.01×10^{23} atoms

11 a 3552 moles

b 2.14×10^{27} water molecules

12.4 SECTION REVIEW

REMEMBERING

1 The reactant that is totally consumed when the chemical reactions reaches completion

2 The amount (mass) of a product that is produced from a complete reaction of a limiting reagent

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3 The experimental yield is usually less than the theoretical yield is usually less than the theoretical yield because:
- Competing reactions may not have been considered in the balanced equation we have written
- The mechanism of the reaction may differ because the experimental conditions are different than what is required for the assumed reaction.
- The conditions required for the reaction to take place, such as temperature or pressure, may not be perfect.
- The reaction mixture may contain impurities, which influence the reaction.

UNDERSTANDING

4 a Air
  b Ethanol
  c 1:2

APPLYING

5 a $2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$
  b Butane
  c $\frac{13}{2}$ or 6.5 moles O_2
  d 18.2 g CO_2
  e 9.3 g H_2O
  f 59.1%

ANALYSING

6 a $(CH_3)_2NNH_2(l) + 2N_2O_4(g) \rightarrow 2CO_2(g) + 4H_2O(g) + 3N_2(g)$
  b From mole ratios of $(CH_3)_2NNH_2$ and $2N_2O_4$ (2496 : 4999 = 1 : 1.002), we determine that N_2O_4 is slightly in excess, so $(CH_3)_2NNH_2$ is the limiting reagent. This means the theoretical yield of N_2 is calculated as 3 × 2496 moles = 7487 moles, which converts to 104.6 kg.
  c 19 kg N_2O_4 yielding 30 kg N_2 corresponds to 68% yield
7 a 10.24 g
  b 54.9 g

12.5 SECTION REVIEW

APPLYING

1 a 54.8
  b 0.020
  c 0.011
  d 5.00
  e 1.50
2 a 6×10^{22}
  b 3.6×10^{24}
  c 3×10^{23}
  d 4.94×10^{22}
  e 1.3×10^{24}
3 a 3 moles K^+, 3 moles Cl^- ions
  b 72 moles of H atoms
  c 20 moles of atoms
  d 4.5×10^{24} individual oxygen atoms
  e 6.77×10^{24} individual atoms ($9 \times N_A$ atoms in each mole of C_2H_5OH)
4 a $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$
    i 2.25 moles O_2
    ii 4.5 moles MgO
  b i 0.75 moles Fe
    ii 1.5 moles H_2
    iii 4 moles Fe
  c $2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$
    i 35.75 moles O_2
    ii 22 moles CO_2
    iii 27.5 moles H_2O

12.6 SECTION REVIEW

APPLYING

1 a 123.88
  b 44.01
  c 63.02
  d 180.16
2 a 40.99
  b 261.32
  c 105.99
  d 159.61
3 a 74.87% C and 25.13% H
  b 39.34% Na and 60.66% Cl
  c 43.51% Ca, 26.07% C and 30.42% N
  d 22.58% N, 6.50% H, 19.35% C and 51.57% O
4 a i CCl_4
    ii $C_4H_{10}O$
    iii C_2H_5NO
  b Molecular formula for (ii) is $C_4H_{10}O$

12.7 SECTION REVIEW

APPLYING

1 a i 14 g CaO
    ii 11 g CO_2
  b 75% yield of CaO
2 a i 5.42 g $ZnCl_2$
    ii 0.080 g H_2
  b 96%
3 a $2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$
    i 132 kg
    ii 116 kg

b 90.9% yield

Hydrocarbon combustion involves competing reactions that lead to production of CO and C. These reactions, called 'incomplete combustion', reduce the CO_2 produced according to the stoichiometry in the equation written above.

A typical 60 L tank of petrol (density ~ 0.75 kg L^{-1}) will have a mass of 45 kg. Thus a single tank of fuel will produce 116 kg of CO_2.

If a motorist travels 16000 km annually at a fuel economy of 10 L/100 km, this will correspond to a total of 1600 L of fuel used. Hence assuming that 116 kg of CO_2 is produced per 60 L, this yields an estimate of $\frac{116 \times 1600}{60} = 3097$ kg or 3.1 tonnes CO_2 annually for the motorist.

4 a $C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2CO_2(g)$

b 23 g C_2H_5OH

c 60% experimental yield. Glucose fermentation is usually carried out under oxygen-free (anaerobic) conditions with yeast as a catalyst. Yeast population depends on many factors, including temperature, pH (acidity) and potentially toxic impurities. Competing reactions are also present. This means that ideal conditions may not prevail for the fermentation process. Finally, the simple overall equation written above does not include the complex pathways involved in glucose fermentation to yield ethanol.

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 a One-twelfth of the mass of on atom of carbon-12

b $6.02214078 \times 10^{23}$ (usually approximated as 6.022×10^{23})

c Mass is neither created nor destroyed by chemical or nuclear reactions, radioactive decay or by physical transformations

d The amount of a substance (g) containing $N_A = 6.022 \times 10^{23}$ particles (atoms, ions or molecules) of that substance

e Mass of 1 mole of a substance (equals mass of $N_A = 6.022 \times 10^{23}$ atoms, molecules or ions of a substance)

f The percentage by mass of each of the different elements in the compound

g The mass of one formula unit of an ionic compound on a scale in which the mass of an atom of the carbon-12 isotope is exactly 12

h The mass of one molecule of a molecular substance on a scale in which the mass of an atom of the carbon-12 is exactly 12

i The relationship between the relative quantities of substances taking part in a reaction or forming a compound, typically a ratio of whole integers

CATEGORY QUESTIONS

2 The relative atomic mass is used to describe the mass of an atom on a scale in which the mass of an atom of the carbon-12 isotope is exactly 12 amu. The relative molecular mass (M_r) of a substance is the mass of one molecule of the substance on a scale in which the mass of an atom of the carbon-12 isotope is exactly 12 amu.

3 The relative formula mass is used for ionic compounds that have cations and anions in a defined ratio rather than existing as individual molecules. The relative formula mass is the mass of one formula unit of an ionic compound on a scale in which the mass of an atom of the carbon-12 isotope is exactly 12 amu.

4 As an example, percentage composition can be used to determine the mineral ores that contain the higher percentage of a specific metal, providing advice for mining and metal extraction industries. Percentage composition of alloys and composites, such as stainless steel, solder, carbon fibre and ceramics, can be critical in determining the properties of the material.

5 A mole is the amount of substance containing N_A particles of that substance ($N_A = 6.022 \times 10^{23}$).

6 Relative molecular mass is the mass of a single molecule (relative to the mass of the carbon-12 isotope taken as 12 amu) and is expressed in atomic mass units. Molar mass is the mass of one mole of the substance and is expressed in units of g mol^{-1}.

7 a The empirical formula of a compound is the simplest whole number ratio in which the atoms of the elements are present in the compound. The molecular formula states the actual number of each type of atom present in the molecule.

b For an ionic compound, cations and anions exist in a defined ratio rather than being combined as individual molecules. This ratio is equivalent to the empirical formula; that is, the simplest whole number ratio in which the atoms of the elements are present. This means the relative formula mass and the empirical formula are equivalent for an ionic compound.

8 The law of conservation of mass

ELABORATION QUESTIONS

9 a 99.41 b 46.07

10 a 54.09% b 61.54%

11 a 5.00 b 20000

12 a 175.21 g b 90.08 g

EVIDENCE QUESTIONS

13 a 8 b 2 c 2.41×10^{24}

14 CuO

15 1.1 kg

16 64.87 g

17 a $FeSO_{11}H_{14}$

b The presence of 14 hydrogens implies $7H_2O$, so the compound in hydrate form must be $FeSO_4 \cdot 7H_2O$.

18 50.02% Ag in the coin.

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END-OF-CHAPTER EXAM

1 B

2 B

3 B

4 C

5 A

6 A

7 D

8 A

9 D

10 B

11 A

12 D

13 D

14 B

15 36.77% Fe, 21.11% S and 42.13% O

16 a 180.16 g b 6 c 6.022 × 1023

17 a $2H_2S(g) + 3O_2(g) \rightarrow 2SO_2(g) + 2H_2O(l)$

b 1.88 kg

18 a $C_4H_5NO_2$ b $C_8H_{10}N_2O_4$

19 a 0.0200 moles Zn; 0.055 moles HCl

b HCl is in excess by 0.0149 moles

c 0.0403 g H_2

20 a $3AgNO_3(aq) + FeCl_3(aq) \rightarrow 3AgCl(aq) + Fe(NO_3)_3(aq)$

b $AgNO_3$

c 37 grams of $FeCl_3$ remaining

21 Chlorine is assumed to be in excess

	C_6H_6	Cl_2	C_6H_5Cl
Molar mass (g mol^{-1})	78.05	70.90	112.49
Mass (g) product (C_6H_5Cl) required			100
Number of moles (n) product required			0.889
Minimum number of moles (n) C_6H_6 required (if yield is 100%)	0.889		
Number of moles (n) C_6H_6 required if yield is 65%	1.37		
Hence mass (g) C_6H_6 required if yield is 65%	107g		

CHAPTER 13: INTERMOLECULAR FORCES

13.1 SECTION REVIEW

REMEMBERING

1 a A pair of electrons shared by two atoms to form a covalent bond

b A valence pair of electrons present in a covalent substance that is not involved in the covalent bond

c Electrons in the outermost shell of an atom in its ground state

UNDERSTANDING

2 CO_2 is linear while SO_2 is bent.

13.2 SECTION REVIEW

REMEMBERING

1 a A permanent build-up of negative charge at one end, and positive charge at another end, of a covalent bond or molecule

b A covalent bond that has a separation of charge

c A molecule with a separation of charge; one end positive' another region negative

UNDERSTANDING

2 a O—H: O more negative δ—, H more positive δ+

Cl—H: Cl δ—, H δ+

C—H: very small difference, C δ—, H δ+

P—H: no difference

N—O: O δ—, N δ+

N—Cl: no difference

NH: N δ—, H δ+

S—H: S δ—, H δ+

C—S: no difference

F—O: F δ—, O δ+

Cl—O: O δ—, Cl δ+

C—O: O δ—, C δ+

P—O: O δ—, P δ+

C—Cl: Cl δ—, C δ+

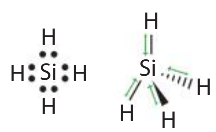
b Highly polar: O—H, N—H, Cl—H, C—O, P—O, C—Cl

Some polarity: N—O, N—Cl, S—H, F—O, Cl—O,

Non-polar: N—Cl, S—H, C—H, P—H, C—S

APPLYING

3 a i SiH_4



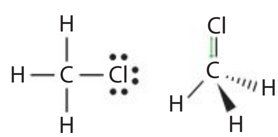
ii Tetrahedral

iii No polar bonds

iv As shown above

v Non-polar

b i CH_3Cl



ii Tetrahedral

iii Polar bonds present

iv As shown above

v Polar

c i CS_2

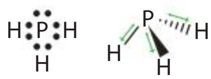
ii Linear

iii No polar bonds

iv As shown above

v Non-polar

d i PH_3



ii Pyramid

iii No polar bonds

iv As shown above

v Non-polar

13.3 SECTION REVIEW

REMEMBERING

1 Metallic, ionic and covalent

2 Dipole-dipole, H-bonding and dispersion forces

UNDERSTANDING

3 In liquid water, the hydrogen bonding intermolecular forces do not affect the overall volume. When the water freezes, the water molecules are held apart from each other by the hydrogen bonds, causing the solid to have a greater volume than the liquid.

4 Dispersion, dipole-dipole, hydrogen

5 Substances dissolve best in solvents that have the same intermolecular forces as them

APPLYING

6 Hydrogen disulfide exhibits dipole–dipole intermolecular forces. Water exhibits hydrogen bonding, a much stronger intermolecular force. This requires a much greater thermal energy to overcome, meaning that water has a much higher boiling point.

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 a Electrons shared by two atoms to form a covalent bond

b A valence pair of electrons present in a covalent substance that is not involved in the covalent bond

c The temperature at which solids melt

d The temperature at which liquid boils and turns to a vapour

e The amount of solute that can dissolve in 100 g of solvent

f Pressure exerted down by a vapour onto its liquid form

g The weak attractive force between atoms and molecules caused by an instantaneous temporary change in dipole moment, arising from the movement of orbiting electrons

h The attraction between molecules with permanent dipoles

i The intermolecular attraction between hydrogen in 2OH, 2NH or 2FH and the lone pairs on nitrogen, oxygen or fluorine on an adjacent molecule

CATEGORY QUESTIONS

2 a Polar

b Polar

c Slightly polar

d Polar

e Polar

3 Any group that contains a hydrogen atom directly bonded to an F, O or N atom

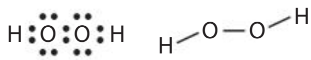
ELABORATION QUESTIONS

4 Answers will vary.

EVIDENCE QUESTIONS

5 The Lewis structure of hydrogen peroxide suggests a linear shape but the lone pairs of electrons on the oxygen atoms move so that they are as far apart as possible, forcing the molecule to adopt a 'skewed' shape. The polar O–H bonds are not symmetrically oriented and do not cancel out, resulting in the polarity of H_2O_2.

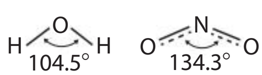
H_2O_2



6 When non-bonding pairs of electrons are present around the central atom of a molecule, the negative charge due to the non-bonding pair of electrons is concentrated much closer to the central atom. This will exert a much greater repulsion on nearby bonding pairs.

In water, there are four pairs of electrons: two bonding and two non-bonding. This gives rise to a tetrahedral arrangement, which should have bond angles of 109.5°. However, the presence of the two non-bonding pairs changes this angle to 104.5°. NO_2 has two oxygen atoms around a central nitrogen atom. This should give a linear shape but the presence of the temporary non-bonding pair of electrons on the nitrogen atom causes the two N=O bonds to be pushed away.

Water vs NO_2



7 Ethyne contains a C≡C triple bond and two C–H bonds. The C≡C triple bond is counted as one bonding pair, giving a total of three bonding pairs. This gives a linear shape.

END-OF-CHAPTER EXAM

1 C

2 D

3 B

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4 A

5 Trigonal pyramidal

6 F–P

7 Vapour pressure

8 Both molecules have dispersion forces but CCl_4 contains more electrons, so its dispersion forces will be stronger and more energy will be required to move from liquid to gaseous state, and the boiling point will be higher.

9 Using the rule 'like dissolves like', H_2O is a hydrogen-bonded solvent. The NH_3 solute also has hydrogen bonding, which is the reason for its high solubility. Phosphine has dipole–dipole forces only, so it is insoluble in water.

10 Dipoles result from a difference in electronegativity between atoms in a covalent bond such as that between H and Cl in a hydrogen chloride molecule. The dipoles present in a Cl2 molecule would be temporary, and virtually non-existent.

11 CCl_4 would evaporate more quickly. Dispersion forces between molecules means that those at the surface of the liquid are only weakly held by the molecules beneath them. As a result, less energy is required from the surroundings to cause the surface molecules to evaporate.

H_2O molecules at the surface are held much more strongly to those beneath them by hydrogen bonding.

CHAPTER 14: CHROMATOGRAPHY

14.1 SECTION REVIEW

REMEMBERING

1 Paper chromatography, thin-layer chromatography, high performance liquid chromatography and gas chromatography

2 a Organic mixtures such as pigments and drugs

b Small, heat stable compounds

c Large organic compounds

APPLYING

3 A gas chromatogram is produced when a mixture of substances, injected into the column, have been separated due to their relative differences in solubility in the liquid phase in the column and then been detected. These show as peaks on the chromatogram. These peaks can be used to identify the substances in the mixture and the area of the peaks can give information on the quantities of the substances in the mixture.

4 A large quantity of material would soon clog up a TLC plate

14.2 SECTION REVIEW

REMEMBERING

1 a The substance in which the solute dissolves or the greater part of a solution

b The substance that is dissolved or the smaller component of a solution

c The intermolecular attraction between hydrogen in 2OH, 2NH or 2FH and the lone pairs on nitrogen, oxygen or fluorine on an adjacent molecule

d A covalent bond that has a separation of charge

e A covalent bond that has little or no separation of charge

UNDERSTANDING

2 a In a non-polar stationary phase on a TLC plate, hydrogen-bonding substances would be insoluble, so would spend most of their time in the mobile phase. They would appear at the top of the plate close to the solvent front. Dipole–dipole substances would appear next, followed by the dispersion-only substances.

b In a polar stationary phase, the situation described in part a would be reversed.

APPLYING

3 Compound 1. Compound 5 must the most non-polar. Its retention time is the longest, so it spends the most time dissolved in the non-polar stationary phase.

4 Changing the polarity of the stationary phase can make chromatography more effective but great care needs to be taken to choose a stationary phase that does not chemically interfere with the components to be separated.

14.3 SECTION REVIEW

UNDERSTANDING

1 While each spot represents a particular component, it is possible that very small quantities of other substances travel with each component and will spend very slightly different amounts of time in the stationary phase

APPLYING

2 a To provide known samples with which to compare samples from police

b Y. The sample produces a spot with an R_f value of $\frac{14.5}{20} = 0.725$, which is the same as the R_f value for the drug Y.

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 a The attraction to the surface of a material; the opposite of 'desorption'

b The action of the substance moving from the stationary phase into the mobile phase

c The substance being analysed

d The phase that carries solutes through the stationary phase

e The substance that is fixed in place during chromatography, to which the solute absorbs

f Retardation factor; the distance solute moves compared to the distance the solvent has moved

g Retention time; the time it takes for a solute to elute from the column in gas chromatography and high-performance liquid chromatography

CATEGORY QUESTIONS

2 The different types of polar groups, the amount of charged and polar chemical groups present, molecular weight, geometry, the positions and numbers of carbon–carbon double bonds.

ELABORATION QUESTIONS

3 HPLC can be used to analyse heat-sensitive compounds.

4 Area under a peak indicates the amount of substance present. A calibration graph must be plotted using known amounts of a substance. The inverse of the slope of this line is the proportionality constant by which an area on the chromatogram must be multiplied.

EVIDENCE QUESTIONS

5 a D, because it had the longest retention time.

b 22%

c Honey decomposes when heated over 40°C

d Put a sample of pure fructose through the column, measure its retention time and compare this with the retention time for peak C.

END-OF-CHAPTER EXAM

1 A

2 A

3 B

4 Stationary phase

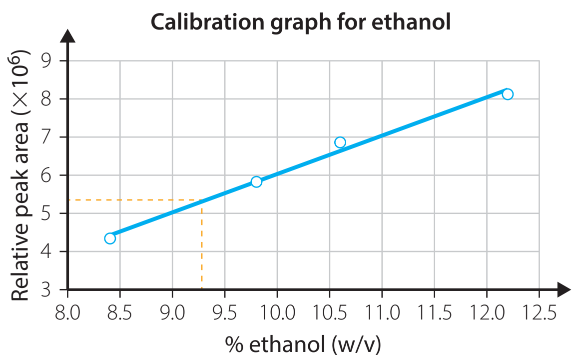
5 Retention time

6 Carrier gas should be inert

7 No oven means that it can analyse heat – sensitive substances.

Small size of particles in stationary phase ensures better separation allowing for shorter columns

8 a



The percentage of ethanol in the champagne is 9.3% (w/v).

b 11.9% v/v

c A calibration graph enables quantitative measurements to be made where the % ethanol of an unknown sample is compared to those of unknown samples

d The calibration graph uses known samples of ethanol ranging from 8.4–12.2% ethanol.

For low alcohol samples the calibration graph would need to be extrapolated down to around 7%, which would create an unacceptable error.

CHAPTER 15: GASES

15.1 SECTION REVIEW

REMEMBERING

1 a A state of matter in which particles flow easily to take the shape and volume of their container

b Equal volumes of any gas, measured at the same temperature and pressure, contain the same number of particles

c The volume occupied by 1 mole of a gas at known temperature and pressure

APPLYING

2 Pressure

3 a 2.13 L

b i 0.02 mol

ii 0.64 g

c i 112 L

ii 5.1 L

iii There are more moles of H_2 in 10 g than there are moles of CO_2. The greater number of moles of gas present, the greater the volume occupied by the gas.

15.2 SECTION REVIEW

REMEMBERING

1 Gases consist of particles that move in continual, random straight-line motion.

The average distance between gas molecules is very large compared to the size of the molecule.

Intermolecular forces between molecules are negligible.

All collisions of gas molecules are perfectly elastic collisions, which means there is no net energy loss during these collisions.

Pressure is due to collisions of the molecules with the walls of the container.

Temperature is a measure of the average kinetic energy of the molecules.

UNDERSTANDING

2 At higher temperatures, the gas particles move faster, which creates greater pressure on the inside of the balloon, causing it to burst.

3 Hydrogen gas, H_2 has a molar mass of $2\,g\,mol^{-1}$.

Oxygen gas, O_2, has a molar mass of $32\,g\,mol^{-1}$.

Hydrogen is lighter, so it moves faster.

APPLYING

4 Molar mass of $HCl = 36.5\,g\,mol^{-1}$, molar mass of $NH_3 = 17\,g\,mol^{-1}$.

The white 'smoke' will appear closer to where the HCl is introduced.

NH_3, being a lighter molecule, will travel fastest and, therefore, furthest.

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5 a 4.11 L
   b 1.4 atm
   c 283 mL
   d 25 kPa
6 a 4.2 mL
   b i 2.1 L
     ii 140K = –132°C
   c 53 mL
7 a 14 L
   b 0.736 g
   c 336 kPa

15.3 SECTION REVIEW

APPLYING

1 1.1 L
2 20.4 L
3 278 g

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 a Conditions of a gas under which temperature = 0°C and pressure = 100 kPa
   b Conditions of a gas under which temperature = 25°C and pressure = 100 kPa

CATEGORY QUESTIONS

2 a There are, on average, great distances between the particles in a gas. These can be pushed quite close together before particle repulsion becomes significant.
   b Under normal conditions of temperature and pressure, gas particles are in constant random motion and there are great distances between particles, meaning that they can move easily past one another.
   c The great distances between gas particles means that, in any given volume, there are relatively few gas particles.
3 The general gas equation is based on the assumptions made by the kinetic theory of gases. At high pressures and low temperatures, the gas particles are forced closer together. Under these conditions intermolecular forces and particle volume become significant, contravening two of the main assumptions made by the kinetic theory.

EVIDENCE QUESTIONS

4 A cyclone is a region of intense low pressure (approximately 79 kPa). If all windows are closed and the inside of the house has a pressure of approximately 100 kPa, the air would rush out of the living room to equalise the pressure. Any closed windows could smash as a result.
5 In winter, the air in the tyres is colder, so it will have a lower pressure.
6 a The various gases that that make up air could behave differently under the high pressures inside a car tyre, particularly water vapour, which could condense.
   b Water makes up a tiny proportion of air so any condensation would have virtually no effect on the performance of car tyres.
7 Air rushing over the top of the straw draws paint up the straw suggesting that the instantaneous pressure at the top of the straw caused by air rushing over it is less than at the bottom of the straw that is dipped into the paint.
   Bernoulli's principle could be that the velocity of a fluid moving at high density is greater than that of a fluid moving at low density.

ELABORATION QUESTIONS

8 a 3200 kg
   b 3.2×10^{12} kg
   c 3.5×10^{18} g
   d 9.3×10^{24}
   e 1×10^{10} trees
9 An ideal gas is one that conforms to the assumptions made by the kinetic theory of gases. A real gas is one that takes into account the intermolecular forces present in the gas and the volume occupied by the gas particles.

END-OF-CHAPTER EXAM

1 B
2 B
3 C
4 Pa, atm
5 22.71 L
6 Pressure
7 A gas takes the volume of its container.
8 The pressure of the vapour decreases as it cools overnight. The air pressure outside the can is greater and so the walls of the can are crushed.
9 Ne
10 C
11 A
12 C
13 A
14 A
15 A
16 B
17 A

18 a Correct

b Correct

c Incorrect. At constant temperature and pressure, the volume of a gas increases as the number of moles of the gas increase.

d Incorrect. $T_2 = \frac{P_2V_2T_1}{P_1V_1}$

19 6 L

20 83.3 kPa

21 32 g

22 23.5 g mol^{-1}

23 12.2 L

24 98.5 atm

CHAPTER 16: AQUEOUS SOLUTIONS AND MOLARITY

16.1 SECTION REVIEW

REMEMBERING

1 a The mass per unit volume of a substance

b The tension of the surface film of a liquid, due to the attraction between particles in the surface layer

c The pressure of a substance in the gaseous state above its liquid

d The ability of a liquid to flow through narrow spaces in opposition to gravity

e Water molecules present in a fixed ratio that form an essential part of the crystal structure of some compounds

UNDERSTANDING

2 Water freezes at 0°C and boils at 100°C. It can act as a solvent over this entire temperature range, dissolving and transporting materials across a cell or over oceans. It is Earth's only substance that exists naturally in its three states of matter: solid, liquid and gas. And, unlike other substances, solid water is less dense than the liquid form, which explains ice's capacity to float. Boiling point, density in solid and liquid phases, surface tension and ability to act as a solvent are among water's unique properties.

3 Water's surface tension is high because hydrogen bonding occurs in all directions and is relatively strong. It holds the molecules together. Water's cohesive nature helps plants absorb nutrients from soil, and capillary action, where liquid can flow through narrow spaces in opposition to gravity, helps the water to move up the xylem of the plant to the leaf's surface. It relies on the adhesion of water molecules to the sides of the xylem vessel, and water's cohesive nature itself. Cohesive forces, surface tension and adhesive forces all assist in this phenomena.

4 Hydrated copper(II) sulfate ($CuSO_4 \cdot 5H_2O$) has five water molecules of crystallisation per formula unit. The bonds between the ions and water are called ion–dipole bonds, and they form a weak bond in the water molecules. The water ligands take positions around the central ion, following VSEPR theory, giving $CuSO_4 \cdot 5H_2O$ its octahedral shape.

APPLYING

5 Water expands when it freezes. If the bottle is completely full it can shatter, due to the expansion of the ice as it forms.

16.2 SECTION REVIEW

REMEMBERING

1 High heat capacity; high boiling point; solid is less dense than liquid; high surface tension

2 Good solvent, so materials can be dissolved and transported; good heat sink, so temperature can be regulated; relatively chemically inert

3 Water molecules are attracted to a salt by ion–dipole bonds and are held in a fixed configuration. They contribute to the array of particles that form a crystal lattice.

UNDERSTANDING

4 More energy is released when the ion–dipole bonds are formed than is required to break the ionic bonds in the crystal lattice of salt

5 Water expands when it freezes. If the bottle is completely full it can shatter, due to the expansion of the ice as it forms.

APPLYING

6 The energy required to break the ionic bonds in the crystal lattice of barium sulfate is more than is released when the ion-dipole bonds with water are formed. This means the dissolution cannot happen.

7 If water were still present in the sample, then the mass would decrease with each subsequent heating and weighing. Once the mass stops changing, it can be assumed that the sample is anhydrous

ANALYSING

8 Water moved up through the xylem cells in the trunk and stems through capillary action. The narrower the capillary, the greater the height that the water rises to. So, at the top of the tree, where the gravitational potential energy is at its greatest, the tubes are at their narrowest.

16.3 SECTION REVIEW

APPLYING

Answers will depend on student experiments.

1 $\frac{6\text{g}}{100\text{mL}} \times 200\,\text{mL} = 12\,\text{g sugar}$

2 $\frac{36}{300} \times 100 = 12\%\,\text{v/v}$

3 500 ppm = 500 mg L^{-1}

Convert L to mL:

$500\,\text{mg}\,\text{L}^{-1} = \frac{500\text{ mg}}{1000\text{ mL}} = 0.5\text{ mg mL}^{-1}$

Calculate for 100 mL: 0.5 × 100 = 50 mg. 50 mg would be present in 100 mL.

16.4 SECTION REVIEW

REMEMBERING

1 Saturated; unsaturated

UNDERSTANDING

2 a 1.5 g

b 9 g

3 8% w/v

4 a 0.06% w/v

b Yes

5 15 g

6 a No

b Saturated

c Insufficient water

d Crystals would form

7 2.5% w/w; 25 000 ppm

APPLYING

8 Answers will vary.

16.5 SECTION REVIEW

REMEMBERING

1 Concentration is the ratio of solute per amount of solution, while molarity is a unit of concentration. Molarity specifically relates the number of moles of a solute per litre of solution.

UNDERSTANDING

2 a i 0.75 mol L^{-1}

ii 89.25 g L^{-1}

b i 0.0100 mol L^{-1}

ii 1.50 g L^{-1}

3 a 0.040 mol

b 0.075 mol

c 0.064 mol

APPLYING

4 a 10.5 g

b 10.7 g

c 8.6 g

5 a 0.182 M

b 0.0044 M

c 0.67 M

6 a 0.053 g

b 0.51 g

7 0.04 L

8 a $c = \frac{n}{V} = \frac{0.2}{1.5} = 0.133 \text{ M}$

b $c = \frac{n}{V} = \frac{0.04}{0.45} = 0.089 \text{ M}$

9 a $n = \frac{m}{M} = \frac{8}{74.6} = 0.107 \text{ mol}$

$c = \frac{n}{V} = \frac{0.107}{0.25} = 0.43 \text{ M}$

b $n = \frac{m}{M} = \frac{1.46}{151.8} = 0.0096 \text{ mol}$

$c = \frac{n}{V} = \frac{0.0096}{0.100} = 0.096 \text{ M}$

c $n = \frac{m}{M} = \frac{7.5}{169.9} = 0.044 \text{ mol}$

$c = \frac{n}{V} = \frac{0.0044}{0.500} = 0.088 \text{ M}$

10 $2NaOH(aq) + CuSO_4(aq) \rightarrow Na_2SO_4(aq) + Cu(OH)_2(s)$

$n(NaOH) = CV = 0.450 \times 0.100 = 0.0450 \text{ mol}$

$n(Cu(OH)_2) = \frac{1}{2}n(NaOH) = 0.0225 \text{ mol}$

$m = nM = 0.0225 \times 97.6 = 2.20 \text{ g}$

11 $ZnCl_2(aq) + 2NaOH(aq) \rightarrow Zn(OH)_2(s) + 2NaCl(aq)$

$n(ZnCl_2) = CV = 0.500 \times 0.010 = 0.0050 \text{ mol}$

$n(Zn(OH)_2) = n(ZnCl_2) = 0.0050 \text{ mol}$

$m = nM = 0.0050 \times 99.4 = 0.50 \text{ g}$

ANALYSING

12 0.88 g

13 a 0.034 M

b 1.38 g in 1 L

c Yes

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 a The tension of the surface film of a liquid, due to the attraction between particles in the surface layer

b The ability of a liquid to flow through narrow spaces in opposition to gravity

c The amount of solute that can dissolve in 100 g water (solvent) at a particular temperature

d A substance that is dissolved in a solvent to form a solution

e A substance that does not easily dissolve in water

f A substance that mixes readily with water

g The number of particles in a given volume

h The number of moles of solute per litre of solution

CATEGORY QUESTIONS

2 Surface tension arises due to the hydrogen bonding interactions between water molecules

3 a 0.001 492 moles

b 0.001 92 moles

c 0.003 84 moles

4 a 3 moles

b 0.7 moles

5 a 2.5 M

b 5.0 M

6 a 0.0242 M
   b 0.0019 M
   c 0.00456 M
7 a 10.01 ml
   b 0.1 L
   c 0.943 L

ELABORATION QUESTIONS

8 4% w/v or 40 ppm
9 The formamide would rise up the inside of the tube much higher than the level in the container
10 31.88 ml
11 0.0015 ng

END-OF-CHAPTER EXAM

1 C
2 B
3 D
4 B
5 A
6 a $8.0 \times 10^3\ \text{mg L}^{-1}$
   b 0.80 % (m/V)
   c 0.137 M
   d 8.0×10^3 ppm
7 a Sugar is a polar molecular substance and so is soluble in water, due to the hydrogen bonding interactions with water molecules. Flour is a non-polar molecule and cannot form these interactions, so is insoluble.
   b The heat capacity of water is much greater than that of metals, due to the hydrogen bonding interactions that exist between the water molecules. As such, the amount of heat required to raise the temperature of water is much greater than that of metals. Consequently, more heat is lost when cooling, and hence the cooling process of water is slower than for a metal.
   c Ice is less dense than water, due to the crystal structure of ice where the water molecules are held apart, by the hydrogen bonding interactions between the molecules
   d Capillary action due to the hydrogen bonding interactions between water molecules can cause water to rise up the straw due to its narrow width.
8 a $n = CV = 0.200 \times 2.00 = 0.400$ mol
   $M(\text{NaI}) = 149.9\ \text{g mol}^{-1}$
   $m = nM = 0.400 \times 149.9 = 60.0$ g
   b $n = CV = 0.01 \times 0.250 = 0.0025$ mol
   $M(K_3PO_4) = 213.3\ \text{g mol}^{-1}$
   $m = nM = 0.0025 \times 213.3 = 0.533$ g
   c $n = CV = 0.0075 \times 0.500 = 0.00375$ mol
   $M(Ba(NO_3)_2) = 261.3\ \text{g mol}^{-1}$
   $m = nM = 0.00375 \times 261.3 = 0.980$ g
9 $Na_2CO_3(aq) + MgCl_2(aq) \rightarrow 2NaCl(aq) + MgCO_3(s)$
   $n(MgCl_2) = CV = 0.08 \times 0.075 = 0.0060$ mol
   $n(MgCO_3) = n(MgCl_2) = 0.0060$ mol
   $m = nM = 0.0060 \times 84.3 = 0.51$ g

CHAPTER 17: IDENTIFYING IONS IN SOLUTION

17.1 SECTION REVIEW

REMEMBERING

1 a An insoluble solid formed from a chemical reaction between two solutions
   b An ion that remains in solution before and after a chemical reaction, so it does not take part in the reaction

UNDERSTANDING

2 Sodium hydroxide, copper(ii) sulfate, lithium chloride
3 a Potassium nitrate and zinc chloride
   b Ammonium carbonate and sodium sulfate
   c Magnesium sulfate and copper(ii) bromide

APPLYING

4 a Magnesium hydroxide and sodium carbonate
   $Mg(OH)_2(aq) + Na_2CO_3(aq) \rightarrow MgCO_3(s) + 2NaOH(aq)$
   b Sodium sulfate and lead nitrate
   $Na_2SO_4(aq) + PbNO_3(aq) \rightarrow 2NaNO_3(aq) + PbSO_4(s)$
   c Silver nitrate and potassium bromide
   $AgNO_3(aq) + KBr(aq) \rightarrow AgBr(s) + KNO_3(aq)$

ANALYSING

5 D
6 a Add sodium hydroxide – lead will produce a precipitate but barium will not.
   b Add sodium hydroxide – copper will produce a blue precipitate, and iron will produce a green precipitate.
7 $K_2S(aq) + Co(CH_3COO)_2(aq) \rightarrow 2CH_3COOK(aq) + CoS(s)$
   $S^{2-}(aq) + Co^{2+}(aq) \rightarrow CoS(s)$
8 a $Cu(OH)_2$ copper hydroxide
   b None produced
   c $PbSO_4$ lead sulfide

17.3 SECTION REVIEW

REMEMBERING

1 a When the icons an ionic compound separate from one another to form a solution in water
   b A solution that contains salts, which will therefore conduct electricity
   c An equation showing all ions present in a chemical reaction, both those reacting and spectator ions
   d A simple equation showing how the formulae and ratio of all compounds taking part in a chemical reaction, showing no ions

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e An equation that shows only the species that are taking part in the chemical reaction; spectator ions are not included

f Ions that remain in a solution before and after a chemical reaction has taken place, which are not involved in the reaction

UNDERSTANDING

2 a $NaCl(aq) + AgNO_3(aq) \rightarrow NaNO_3(aq) + AgCl(s)$

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

b $CuSO_4(aq) + 2KOH(aq) \rightarrow Cu(OH)_2(s) + K_2SO_4(aq)$

$Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s)$

c NR

d $Na_2CO_3(aq) + FeSO_4(aq) \rightarrow Na_2SO_4(aq) + FeCO_3(s)$

$Fe^{2+}(aq) + CO_3^{\ 2-}(aq) \rightarrow FeCO_3(s)$

e $Zn(NO_3)_2(aq) + (NH_4)_2S(aq) \rightarrow ZnS(s) + 2NH_4NO_3(aq)$

$NO_3^{\ -}(aq) + NH_4^{\ +}(aq) \rightarrow NH_4NO_3(aq)$

f $K_2CO_3(aq) + CaCl_2(aq) \rightarrow CaCO_3(s) + 2KCl(aq)$

$CO_3^{\ 2-}(aq) + Ca^{2+}(aq) \rightarrow CaCO_3(s)$

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 a Ions that remain in a solution before and after a chemical reaction has taken place, which are not involved in the reaction

b An insoluble solid formed from a chemical reaction between two solutions

c A mixture of a liquid and an insoluble solid, where solid distributes evenly throughout the liquid

d An equation that shows only the species that are taking part in the chemical reaction; spectator ions are not included

e A solution that contains ions, so it will conduct electricity

ELABORATION QUESTIONS

2 a No precipitation

b $2NH_4Br(aq) + Pb(NO_3)_2(aq) \rightarrow PbBr_2(s) + 2NH_4NO_3(aq)$

$2Br^-(aq) + Pb^{2+}(aq) \rightarrow PbBr_2(s)$

c $(NH_4)_2S(aq) + Mg(CH_3COO)_2(aq) \rightarrow MgS(s) + 2NH_4CH_3COO(aq)$

$Mg^{2+}(aq) + S^{2-}(aq) \rightarrow MgS(s)$

d $SrCl_2(aq) + ZnSO_4(aq) \rightarrow SrSO_4(s) + ZnCl_2(aq)$

$Sr^{2+}(aq) + SO_4^{\ 2-}(aq) \rightarrow SrSO_4(s)$

3 a Lead nitrate and sodium sulfate

b Iron nitrate and ammonium sulphide

c Magnesium nitrate and potassium hydroxide

4 Add some ammonium chloride solution – if lead, a precipitate of lead chloride will form; if calcium, no precipitate will form

EVIDENCE QUESTIONS

5 Sodium carbonate

6 Barium

7 A – lead; B – iron(iii); C – barium; D – calcium

8 Answers will vary.

END-OF-CHAPTER EXAM

1 D

2 C

3 B

4 A

5 A

6 Since all common nitrate compounds are soluble then the precipitate would be barium sulfate. This is supported by the colour as chromium(III) compounds tend to be coloured because chromium is a transition metal element.

$$Ba^{2+}(aq) + SO_4^{\ 2-}(aq) \rightarrow BaSO_4(s)$$

7 The target ion is the calcium ion. A solution containing sulfite, carbonate, sulfate or phosphate ions should precipitate the calcium ion.

8 a Ionic solids will have a higher solubility at higher temperatures

b Cost of heating water makes this prohibitive

c Any soluble substance with an anion that will form a precipitate with magnesium, eg ammonium hydroxide

d $NH_4^{\ +}(aq) + OH^-(aq) \rightarrow NH_4OH(s)$

CHAPTER 18: SOLUBILITY

18.1 SECTION REVIEW

REMEMBERING

1 Refer to Figure 18.1.2.

2 a $LiBr\ 2H_2O$

b $Ca(NO_3)_2\ 4H_2O$

c $Mg(NO_3)_2\ 6H_2O$

UNDERSTANDING

3 HBr is polar, so it can form dipole–dipole attractions with the water molecules.

4 Sugar becomes the solute and the coffee the solvent

5 a The ionic solid splits into positive and negative ions, which form ion-dipole bonds with the water

b $KBr(s) \rightarrow K^+(aq) + Br^-(aq)$

c Yes – presence of dissociated ions means that it will conduct electricity.

6 $CH_3CH_2OH(l) \rightarrow CH_3CH_2OH(aq)$. Hydrogen bonds between ethanol molecules break and new hydrogen bonds form between ethanol molecules and water molecules.

APPLYING

7 a NaCl
   b Covalent polar molecular
   c Sparingly soluble
   d Hydrocarbons
   e Cellulose
   f Soluble due to significant number of polar bonds
   g Macromolecular
   h Insoluble
   i Insoluble
   j Pb, Cu
8 a Soluble – polar covalent molecule
   b Soluble – ionic
   c Insoluble
   d Soluble
   e Soluble
   f Soluble
   g Soluble
   h Insoluble
   i Soluble

ANALYSING

8 Answers will vary.
9 Answers will vary.

18.2 SECTION REVIEW

REMEMBERING

1 a A measure of the average kinetic energy of the particles that constitute a solid, liquid or gas
   b A graph showing the variation of the solubility of a substance in water with temperature

UNDERSTANDING

2 Greater kinetic energy causes solid structure to be broken down more easily, so the higher the temp, the better the solubility of an ionic solid. For gases, there are no bonds to break and the formation of ion-dipole bonds is exothermic, which is reversed by raised temperature.

APPLYING

3 Raising sea level temperatures can lower the solubility of oxygen and CO_2 in water, which can affect the survival of marine animals and plants.

18.3 SECTION REVIEW

UNDERSTANDING

1 a 13.5 g
   b 46 g
   c 45 g
2 unsaturated
3 $K_2Cr_2O_7$
4 15.5 g
5 26 g

APPLYING

6 a 30 g
   b 40 g
   c 80 g
7 $NaNO_3$, because it has the highest value.
8 At 30°C, there would be 36 g of KCl and 36 g of NaCl. The solutions would be saturated because the curve represents where the solutions are saturated.

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 The higher the temperature, the greater the solubility
2 The higher the temperature, the lower the solubility

CATEGORY QUESTIONS

3 It dissociates, forming ion–dipole bonds with the water molecules
4 Energy released when ion–dipole bonds are formed is less than the energy required to break the ionic bonds in the solid.

ELABORATION QUESTIONS

5 a 30 g
   b 23 g or more
   c 8 g
6 a Answers will vary.
   b $KCl(s) \rightarrow K^+(aq) + Cl^-(aq)$
   c 26 g
   d Supersaturated
   e 43°C; 11 g

EVIDENCE QUESTIONS

7 Answers will vary.

END-OF-CHAPTER EXAM

1 D
2 A
3 A
4 A
5 D
6 a 82 g/100 mL. This would require heating the solution to approximately 43°C to prepare a saturated solution.
   b 5.25 g
   c Supersaturated – 6.5 M = 858 g L^{-1} = 85.8 g/100 g. Saturation limit of 80 g/100 g is exceeded.

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7 a 63 g
  b Unsaturated
  c 0.84 M
  d Diagram should show formation of ion–dipole bonds between sodium and chloride ions and water molecules

CHAPTER 19: PH

19.1 SECTION REVIEW

REMEMBERING

1 a A substance that produces hydrogen ions, H^+, in aqueous solution
  b A substance that either contains the oxide (O^{-2}) of hydroxide ion (OH^2), or produces an hydroxide ion in aqueous solution
  c A substance that can act as either an acid or a base depending on the other reactants
  d A measure of the acidity or alkalinity of a solution, based on the concentration of hydrogen ions in a solution

UNDERSTANDING

2 $H_3PO_4(aq) + 3NaOH(aq) \rightarrow Na_3PO_4(aq) + 3H_2O(l)$

The Arrhenius theory states that an acid is one that releases H^+ ions into solution and a base is one that donates OH^- ions into solution.

In the example given, the phosphoric acid produced the H^+ ions, which combined with the OH^- ions produced by the sodium hydroxide to form water.

APPLYING

3 a The stronger the acid or base is, the greater its electrical conductivity.
  b $Mg(OH)_2$ should have a high electrical conductivity. The low value must be due to its very low solubility.

19.2 SECTION REVIEW

REMEMBERING

1 a A scale that measures the acidity and alkalinity of a solution, based on the concentration of hydrogen ions in a solution
  b A substance that changes colour in solution depending on whether the solution is acidic or basic

2 Acidic: pH 1–6.9

Basic: pH 7.1–14

Neutral: pH 7

3 B and D

UNDERSTANDING

4 There is not necessarily a relationship between concentration of an acid or a base and pH. The pH depends on the concentration of H^+ ions or OH^- ions.

19.3 SECTION REVIEW

REMEMBERING

1 a A substance that produces H^+ ions in solution
  b A substance that produces OH^- ions in solution

APPLYING

2 $CH_3CH_2COOH(aq) \rightarrow CH_3CH_2COO^-(aq) + H^+(aq)$

$CH_3CH_2CH_2NH_2(aq) + H_2O(l) \rightarrow CH_3CH_2CH_2NH_3^+(aq) + OH^-(aq)$

3 $CH_3CH_2COOH(aq) + CH_3CH_2CH_2NH_2(aq) \rightarrow CH_3CH_2CH_2NH_3^+(aq) + CH_3CH_2COO^-(aq)$

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 a A substance that produces a relatively large number of hydrogen ions, H^+, in aqueous solution
  b A substance that produces a relatively small number of hydrogen ions, H^+, in aqueous solution
  c A substance that produces a relatively large number of hydroxide ions, OH^-, in aqueous solution
  d Author to supply answer

CATEGORY QUESTIONS

2 a 1–3
  b 4–6
  c 8–10
  d 11–13

EVIDENCE QUESTIONS

3 Vinegar has been used because it was mistakenly believed that vinegar neutralised the basic bluebottle venom. In fact, vinegar aggravates the area around the wound and does not do anything to relieve the pain. The current consensus is that warm water should be used. Recent research has shown that applying warm water can reduce the pain by half in 10 minutes.

END-OF-CHAPTER EXAM

1 D
2 C
3 C
4 100
5 1000
6 Polar covalent bond
7 Can provide information on the strength of acids and bases
8 The NH_4^+ ion produced is itself a weak acid. If the HCl and NH_3 have been neutralised, all that is left is the NH_4^+ and Cl^- ions. The Cl^- ion is neutral but the NH_4^+ ion has a pH of about 4.

CHAPTER 20: REACTION OF ACIDS AND BASES

20.1 SECTION REVIEW

REMEMBERING

1 a A substance that produces hydrogen ions, H^+, in aqueous solution

b A substance that either contains the oxide (O^{2-}) or hydroxide ion (OH^-), or produces an hydroxide ion in aqueous solution

c Chemical compound formed from the reaction of an acid with a base, carbonate or metal with all or part of the hydrogen of the acid replaced by a metal or other cation.

d A reaction between an acid and a base to form a salt and water

e An oxide that can act as either an acid or a base depending on the other reactants

UNDERSTANDING

2 a $2HNO_3(aq) + MgCO_3(s) \rightarrow Mg(NO_3)_s(aq) + CO_2(g) + H_2O(l)$

b $H_3PO_4(aq) + 3KOH(aq) \rightarrow K_3PO_4(aq) + 3H_2O(l)$

c $2CH_3COOH(aq) + Mg(s) \rightarrow (CH_3COO)_2Mg(aq) + H_2(g)$

d $H_2SO_4(aq) + CH_3NH_2(aq) \rightarrow (CH_3NH_3)_2SO_4(aq)$

3 The reaction is extremely exothermic, so much so that the heat generated would ignite the hydrogen gas produced.

APPLYING

4 a $HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$

b $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$

c $H_3PO_4(aq) \rightarrow H^+(aq) + PO_4^{3-}(aq)$

5 Although naturally occurring acid rain has been falling for billions of years, the amount of acid (H^+ ions) is relatively small due to small amount of $CO_2(g)$ in the atmosphere. In recent times the amount of atmospheric CO_2 has risen dramatically due to the burning of fossil fuels.

20.2 SECTION REVIEW

UNDERSTANDING

1 a $2HNO_3(aq) + MgO(s) \rightarrow Mg(NO_3)_2(aq) + H_2O(l)$

$2H^+(aq) + 2NO_3^-(aq) + MgO(s) \rightarrow Mg^{2+}(aq) + 2NO_3^-(aq) + H_2O(l)$

$2H^+(aq) + MgO(s) \rightarrow Mg^{2+}(aq) + H_2O(l)$

b $3H_2SO_4(aq) + 2Al(HCO_3)_3(aq) \rightarrow Al_2(SO_4)_3(aq) + 6CO_2(g) + 6H_2O(l)$

$6H^+(aq) + 3SO_4^{2-}(aq) + 2Al^{3+}(aq) + 6HCO_3^-(aq) \rightarrow 2Al^{3+}(aq) + 3SO_4^{2-}(aq) + 6CO_2(g) + 6H_2O(l)$

$6H^+(aq) + 6HCO_3^-(aq) \rightarrow 6CO_2(g) + 6H_2O(l)$

$H^+(aq) + HCO_3^-(aq) \rightarrow CO_2(g) + H_2O(l)$

c $2H_3PO_4(aq) + 3Mg(s) \rightarrow Mg_3(PO_4)_2(aq) + 6H_2(g)$

$6H^+(aq) + 2PO_4^{3-}(aq) + 3Mg(s) \rightarrow 3Mg^{2+}(aq) + 2PO_4^{3-}(aq) + 6H_2(g)$

$6H^+(aq) + 3Mg(s) \rightarrow 3Mg^{2+}(aq) + 6H_2(g)$

$2H^+(aq) + Mg(s) \rightarrow Mg^{2+}(aq) + 2H_2(g)$

d $H_2SO_4(aq) + 2CH_3CH_2NH_2(aq) \rightarrow (CH_3CH_2NH_3)_2SO_4(aq)$

$2H^+(aq) + SO_4^{2-}(aq) + CH_3CH_2NH_2(aq) \rightarrow 2CH_3CH_2NH_3^+(aq) + SO_4^{2-}(aq)$

$2H^+(aq) + 2CH_3CH_2NH_2(aq) \rightarrow 2CH_3CH_2NH_3^+(aq)$

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 a An equation showing all of the reactants and products as neutral compounds or elements

b An equation that shows all the ions present in solution

c An equation that only contains the ions that take part in the chemical reaction; that is, spectator ions are excluded

d An ion that is present in the same form on each side of the equation

e A substance that can act as either an acid or a base depending on the other reactants

CATEGORY QUESTIONS

2 Acid–base, acid–metal, acid–carbonate/hydrogencarbonate and acid–ammonia

ELABORATION QUESTIONS

3 Carbon dioxide and nitrogen oxides

EVIDENCE QUESTIONS

4 Snow falls through the air more slowly than rain, giving it more time to pick up pollutants such as CO_2, SO_2 and NO_2. Once the snow settles on the ground it may not thaw immediately but may build up. When the acid snow melts, large amounts of pollutants can be released immediately, causing significant problems.

5 a Solution A is more acidic

b Solution iii is neutral

END-OF-CHAPTER EXAM

1 D

2 C

3 Neutralisation

4 Ammonium sulfate

5 Nitric acid

6 $Ca(HCO_3)_2(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + 2CO_2(g) + 2H_2O(l)$

$Ca(HCO_3)_2(s) + 2H^+(aq) + 2NO_3^-(aq) \rightarrow Ca^{2+}(aq) + 2NO_3^-(aq) + CO_2(g) + H_2O(l)$

$Ca(HCO_3)_2(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + 2CO_2(g) + 2H_2O(l)$

7 The ammonium ion adds directly to the sulphate ion

8 H_2SO_4 can react with the clay particle which has the metal attached to it.

$H_2SO_4(aq) + Na_2[Al_2Si_2O_5(OH)_4](s) \rightarrow Na_2SO_4(aq) + H_2[Al_2Si_2O_5(OH)_4](s)$

The $Na^+(aq)$ ions are soluble in water and are washed away.

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CHAPTER 21: RATES OF REACTIONS

21.1 SECTION REVIEW

REMEMBERING

1 Change in a quantity (such as mass or volume), or change in time

2 For example, you could measure the changing variable (such as mass or volume) for a series of measured regular time periods.

Alternatively, you could measure the time taken for a particular change in the variable to occur (such as the time taken for 100 mL of gas to be produced).

3 At the beginning of the reaction, when the concentration of the reactants is highest

UNDERSTANDING

4 Measure a change in quantity, and divide the magnitude of the change by the time taken for it to occur.

5 The greater the gradient – the steeper the slope – the faster the reaction

APPLYING

6 Rate = 20/4 = 5 $g\,min^{-1}$

7 Time = 0, rate = 11 $mL\,s^{-1}$

Time = 1, rate = 3.5 $mL\,s^{-1}$

Time = 5, rate = 0.25 $mL\,s^{-1}$

21.2 SECTION REVIEW

REMEMBERING

1 A reaction will occur when two particles collide with sufficient energy, and in the appropriate orientation.

2 a The minimum energy needed for a reaction to occur

b The species formed when the reactant bonds have been broken, but the product bonds not yet been formed

3 Answers will vary.

UNDERSTANDING

4 a If the collision is too gentle, then the reactant bonds will not be broken, and the reaction will not occur

b Certain sections of the molecules reacting may need to come into contact with each other for the reaction to occur. Hence if the molecules are not oriented correctly, the reaction cannot occur

5 The reaction will not occur

APPLYING

6 a Products higher in enthalpy than reactants, and position of activated complex should be 1000 J higher than that of reactants.

b Products lower in enthalpy than reactants, and position of activated complex should be 200 J higher than that of reactants.

21.3 SECTION REVIEW

REMEMBERING

1 Concentration, gas pressure, surface area, temperature, catalysts

2 A substance which changes the rate of a reaction, but which is not used up in the reaction

3 Author to supply artwork brief for answer

UNDERSTANDING

4 With a greater number of smaller-sized pieces, more of the solid is in contact with the other reactant

5 Temperature is a measure of kinetic energy, so the average kinetic energy will be greater at 40°C than at 10°C.

6 Both pressure and concentration are a measure of a number of particles per unit volume.

7 a Greater concentration or pressure means that a collision between reactant molecules is more likely, so the reaction rate will be faster.

b The higher the temperature, the fast the molecules move, so, the greater the frequency of collisions and the greater the energy of the collisions, and the more likely the collisions are to be successful,

c The catalyst provides an alternative pathway for the reaction, with lower activation energy. Hence a greater proportion of collisions will be successful and the rate of the reaction will be faster

d Smaller particle size means that a greater amount of a reactant is in contact with the other reactant. This means that more collisions can occur and the reaction can be faster.

APPLYING

8 Ben's tea would have been sweeter. Stirring would promote mixing of the sugar and water particles, and therefore increase the rate of the dissolving process.

ANALYSING

9 Tube 5 – high concentration, high temperature and also small particle size

REFLECTING

10 There are various factors that affect the rate of a reaction, including temperature, particle size, concentration and the presence of a catalyst. Collision theory enables us to explain how each of these affects the rate of a reaction and also to predict the effect of changing any one of these factors, so it has a number of extremely useful applications.

21.4 SECTION REVIEW

REMEMBERING

1 a Area of the enzyme where the enzyme interacts directly with the substrate molecule

b Molecule that is taking part in the reaction that the enzyme is catalysing

c The permanent change that can occur due to the structure of protein molecules when heated, or exposed to extremes of pH conditions

d Molecules that assist the operation of enzymes

2 Temperature, surface area contact, concentration and enzyme presence

UNDERSTANDING

3 Their complex and specific shape enables them to bind selectively with molecules, hold them in the correct conformation to minimise the activation energy required for a reaction, so maximising the rate of a reaction.

APPLYING

4 A catalyst, being a protein molecule, will denature when the temperature rises beyond a maximum value, or when pH varies too much. This means that the complex shape of the catalyst is altered permanently, meaning that it can no longer bind effectively to the substrate and that it can no longer function as a catalyst.

CHAPTER REVIEW QUESTIONS

DETAIL QUESTIONS

1 a Change in measured quantity per unit time

b Area of solid in contact with other reactant

c Substance which changed the rate of a reaction without itself being used up

d Biological catalyst

e A reaction will occur when two particles collide with sufficient energy in the correct orientation

f An reaction intermediate or transition state formed when the reactant bonds have been broken, but the product bonds not yet formed

g Minimum amount of energy required by a molecule for a reaction to occur

h A reaction that gives out heat

i A reaction that absorbs heat

j The section of an enzyme where it binds to the substrate

CATEGORY QUESTIONS

2 With a greater number of collisions, there is a faster reaction.

3 With a greater number of collisions, and harder collisions, which are more likely to cause a reaction.

4 Smaller particle size = greater surface area where reaction can occur, so more collisions are possible

ELABORATION QUESTIONS

5 If temperature is too high, enzyme can become denatured – permanently altered, so unable to bind to substrate and carry out reaction

6 a Greater proportion of particles with more energy than the activation energy

b Activation energy value is lowered

7 Author to supply artwork brief for answer

8 Because it has absorbed the energy required to break the reactant molecule bonds, but has not released the energy from the new bond formation

EVIDENCE QUESTIONS

9 They catalyse many reactions, such as digestion, which are critical to life.

10 Extremely high surface area

11 This is a research question, so answers will vary, however:

- Catalytic converters are tubes filled with a honeycomb-like substance, coated with catalyst materials.
- As the exhaust gases pass through these tubes, the nature of the honeycomb structure maximises surface area contact with the gases.
- The catalysts include metals, such as platinum, rhodium or palladium.
- They catalyse reactions that convert potentially harmful gases from the exhaust into less harmful gases, such as carbon monoxide into carbon dioxide and nitrogen oxides into nitrogen and oxygen.

END-OF-CHAPTER EXAM

1 C

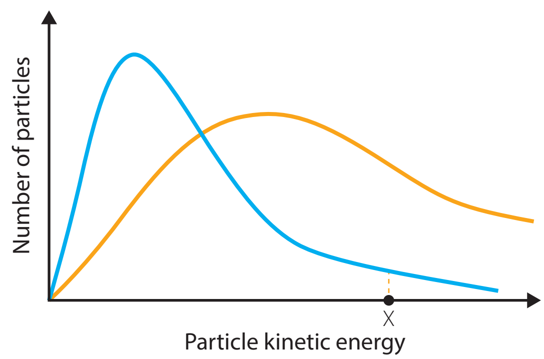
2 B

3 C

4 D

5 a i Minimum amount of energy required for a reaction to occur

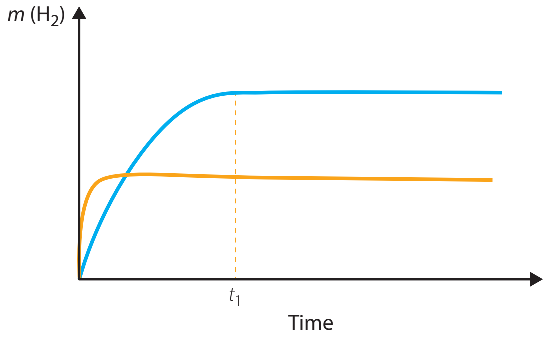
ii



iii Significantly more particles have an energy greater than the activation energy b i At the start

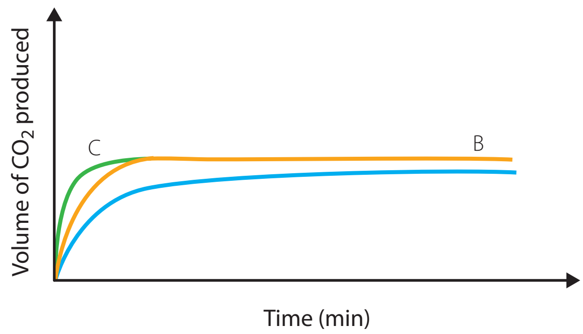
ii Reaction is over

iii



iv Higher temperature; greater concentration of HCl

6 a i–ii



b E – same mass as A, but powder instead of lump

c They have an energy level lower than the required activation energy for that reaction

UNITS 1 & 2 PRACTICE EXAM ANSWERS

MULTIPLE-CHOICE QUESTIONS

1 D
2 A
3 D
4 A
5 C
6 D
7 B
8 C
9 C
10 C

SINGLE-ANSWER QUESTIONS

1 Protons 11; Neutrons 13; Electrons 11
2 56
3 Helium
4 −138°C
5 53.64%

SHORT-ANSWER QUESTIONS

1 Electronegativity is the ability of an element to attract electrons when bonded covalently. It increases from left to right across the periodic table – the number of protons in the nucleus increases, the atomic radius decreases, and ultimately the attraction to electrons outside the atom increases.

2 A polar molecule is one with a permanent dipole – where one end of the molecule is more electronegative than the other and so the electron density is not evenly distributed around the molecule. In both molecules, the carbon–chlorine bond is polar because there is a significant difference in electronegativity between these two elements. However, in CCl_4, the molecule is symmetrical and therefore the effect of the polar bonds is not seen across the full molecule. The overall molecule is non-polar. $CHCl_3$ is a polar molecule because it does not have the symmetry of CCl_4.

3 AAS measures the absorption of a beam of light. The beam of light is created by using a sample of the metal being investigated as the filament of a cathode lamp. Consequently, it emits the unique frequencies of that particular metal which therefore are only absorbed by that metal. This means that AAS will only measure the absorption of the metal ion present in the sample which is the same metal as the cathode is made from.

4 Sodium oxide has a lower melting and boiling point than silicon dioxide. It is soluble in water and conducts electricity when molten, both of which silicon dioxide does not do. Sodium oxide is an ionic compound, made of positively charged sodium ions and negatively charged oxide ions held together in a lattice by electrostatic attraction. Silicon dioxide is a covalent network structure with no charged particles. Hence it has a very high melting and boiling point and does not conduct electricity in any state.

5 Raising the temperature increases the kinetic energy of the particles. This results in increased frequency of collisions and also increased impact of collisions. This means there are more collisions occurring and also a greater likelihood that the collisions will have energy greater than the activation energy for the reaction and so will be successful. Hence the rate of the reaction will be faster when the reactants are heated.

HIGHER-ORDER QUESTIONS

1 a i $AlCl_3$
ii $Ca_3(PO_4)_2$
iii From NaCl, CaO, $CaSO_4$, LiCl or $AlPO_4$
iv Any two from O^{2-}, Na^+, Al^{3+}, alternatively also Cl^{1-} and Ca^{2+}
b i Sketch showing a lattice of positively charged ions surrounded by delocalised electrons
ii 2 3s electrons per atom are delocalised and contributing to the metallic bonding
iii there is a net movement of electrons through the magnesium, moving from the negative terminal of the battery towards the positive terminal

2 a Does not conduct electricity, so it cannot be metal
b Does not conduct electricity when molten, so it cannot be an ionic substance.
c Forces within the molecule are stronger than forces between the molecule – it has a relatively low melting point of 78°C where the intermolecular forces are overcome. The original liquid then re-forms easily, indicating that the intramolecular forces have not been affected.
d Butane is a simple covalent molecule with weak intermolecular forces. Diamond is a covalent network structure with strong forces between the atoms. Hence its melting point is significantly higher

3 a 5.91 L
b The temperature rise is 59.6 K
c $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$
d 9.66 g

4 a i Be: $1s^2 2s^2$
ii F: $1s^2 2s^2 2p^6$
b Atomic radius of fluorine is less than that of beryllium
c Electronegativity of Be is less than that of F – it is a larger atom with a lower nuclear charge. Therefore, it is less attracted to external electrons.

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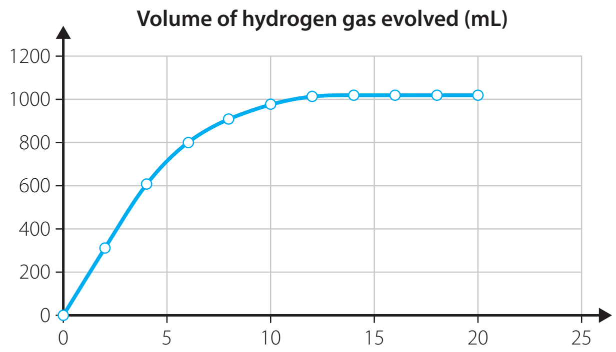
d The first ionisation of F would be greater than that of Be. F is a smaller atom with a greater nuclear charge, so its attraction to its outer shell electrons is greater, and more energy would be required to remove an electron.

e BeF_2 compound would be soluble in water, brittle and conduct electricity when molten.

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

b i



ii Gradient represents the rate of the reaction. The gradient falls as reaction proceeds, as reactant molecules are used up and concentration also falls.

c i Draw in an additional line – steeper at the start, but finishing at the same point.

ii Smaller pieces of magnesium means that the number of collisions between magnesium and acid is greater, so the rate of reaction is greater.

d 0.980 g Mg

GLOSSARY

A

absolute error the magnitude of the difference between the observed/measured value and the true value

absolute uncertainty the spread or interval of measured values within which the true value is expected to lie

absolute zero precisely 0 K on the Kelvin scale, which is a thermodynamic (absolute) temperature scale and −273.15 degrees Celsius on the Celsius scale

absorbance value a measurement of the relative amount of radiation which is absorbed by a sample

absorption spectrum a spectrum of electromagnetic radiation transmitted through a substance, showing dark lines or bands due to the absorption at specific wavelengths corresponding to transitions from excited energy levels to higher energy levels

absorption transition an electron moving to a higher energy level, when absorbing additional energy

accuracy the degree to which the result of a measurement, calculation, or specification conforms to the correct value or a standard

acid a substance that causes the concentration of hydrogen ions in a solution to increase: has a sharp, sour taste, are usually corrosive and have a pH value of less than 7

activated complex also known as the transition state, the intermediate stage of a chemical reaction where the reactant bonds have been broken but the product bonds are not yet formed

activation energy the minimum energy required for a chemical reaction to occur: it significantly influences the rate of a chemical reaction

active site area of the enzyme where the enzyme interacts directly with the substrate molecule

addition reaction a reaction where two molecules combine, resulting in a larger molecule being formed

alkali a substance that causes the concentration of hydrogen ions in a solution to decrease: reacts with acidic substances to form water, which is neutral, are often corrosive and have a pH value of greater than 7

allotrope different forms of the same element, with different physical properties. For example, diamond and graphite are both different forms of carbon

alpha particle a nucleus of a helium atom that is emitted by some radioactive substances as alpha radiation

amorphous shape a non-crystalline solid

amphoteric substance reacts with both alkaline and acidic solutions, so it is classified as neither an alkali nor an acid

anhydrous ionic solids with no water present

anion a non-metallic atom that has become negatively charged through gaining one or more electrons

aromatic hydrocarbons hydrocarbon containing one or more benzene rings

atom the smallest stable unit of matter that can exist by itself electron a subatomic particle with a negative charge that is found in all atoms

atomic mass unit one twelfth of the mass of on atom of carbon-12

atomic number the number of protons contained within one atom of an element – each element has a unique atomic number

atomic orbital the region of space around an atom that has a specific shape and may contain a maximum of two electrons

atomic radius a measure of the size of an atom - measured as the distance from the nucleus to the boundary of the cloud of electrons surrounding it – the distances are very small and are measured in picometres (pm) ($1\,pm = 1 \times 10^{-12}\,m$)

atomised a substance is heated until it exists as a vapour of gaseous atoms

Aufbau principle electrons around an atom will fill the lowest available energy level first, before filling higher energy levels

Avogadro's constant $6.02214078 \times 10^{23}$ (usually approximated as 6.022×10^{23})

Avogadro's hypothesis equal volumes of any gas, measured at the same temperature and pressure, contain the same number of particles

B

base a substance that either contains the oxide (O^{2-}) or hydroxide ion (OH^{-}), or produces hydroxide ions in an aqueous solution

biofuels fuels that are primarily made of biomass

bonding pair a pair of electrons shared between two different atoms

brittle a substance which will shatter rather than bend, when a force is applied

C

calibration a comparison, and if required, adjustment of a measuring device against a known primary standard

calibration curve a graph of absorbance values of a series of samples of known concentrations used to determine the concentration of an unknown sample

Capillary Action the ability of a liquid to flow through narrow spaces in opposition to gravity

carbohydrates food group consisting of molecules such as glucose, sugars and starches

carbon nanotube a rolled-up sheet of graphene capped with fullerene(s)

catalyst a substance that affects the rate of certain reactions by providing an alternative reaction pathway

cation a metallic atom which has become positively charged through losing one or more electrons

chemical bond an interaction between the electron shells of one or more atoms causing them to combine to form a new substance called a compound

chemical bond energy also known as bond enthalpy, is the amount of energy required

chemical properties properties of a substance such as acidity and flammability that relate to the chemical reactions that the substance takes part in

chemical reaction the creation of new substances and associated energy transformations

co-enzymes molecules that assist the operation of enzymes

collision theory for a reaction to occur, the particles must collide with sufficient energy and in the required orientation

colloid a mixture in which tiny clusters of particles are dispersed through another substance

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combustion exothermic chemical reaction between a fuel and an oxidizing agent (usually oxygen) that produces energy

complete ionic equation an equation showing all ions present in a chemical reaction – both those reacting and spectator ions

compound a substance consisting of atoms of two or more elements chemically bonded together in a fixed ratio

concentrated solution a large amount of solute dissolved in a given volume of solvent

concentration the number of particles in a given volume

continuous variables quantitative data for which there are an infinite number of possible values within a selected range

covalent bond a bond that is formed when the valence shells of two atoms overlap to share electrons

covalent bonding bonding that occurs between two or more atoms, where electron pairs are shared between the valence shells

covalent molecular substances composed of discrete molecules whose atoms are held together by intramolecular bonds

covalent network substances carbon and silicon compounds with structures that are three-dimensional networks of covalently bonded atoms

D

delocalised electrons electrons that are not held in a fixed orbit around one nucleus, but are shared by many atoms

denature the permanent change than can occur to the structure of protein molecules when heated, or exposed to solutions of high acidity or alkalinity

density the mass per unit volume of a substance

dependent variable the variable that changes as a result of changes to the independent variable

dilute solution a small amount of solute dissolved in a given volume of solvent

dilution adding water to a solution to reduce the concentration

dipole–dipole force the attraction between molecules with permanent dipoles

discrete variables quantitative or qualitative information that can be categorised into a classification, or is based on counts, with a finite number of values possible; the values cannot be subdivided meaningfully

dispersion force the weak attractive force between atoms and molecules caused by an instantaneous temporary change in dipole moment, arising from the movement of orbiting electrons

dissociated when the ions from an ionic compound separate from one another and form a solution in water

distillate the solvent kept after distillation

distillation a technique for separating the solvent from a solution, when the solvent is required to be kept

double covalent bond a double bond consists of two shared pairs of electrons

ductile a substance that can be drawn out to form a thin wire

ductility ability to be drawn out into a wire

E

elastic collision a collision with no net loss of energy

electrolytic solution a solution that contains ions, so it will conduct electricity

electron configuration the specific arrangement of electrons within an individual atom

electron energy levels, or shells a group of electrons existing at the same distance from the nucleus of an atom; all the electrons in each shell have the same amount of energy

electron shells electrons exist in orbit around an atom in energy levels called shells; the greater the energy level, the further away the shell is from the nucleus

electronegativity the ability of an atom to attract electrons to itself, while being bonded to another atom

electrostatic attraction a force that occurs between particles with opposite charges

electrostatic repulsion a force that occurs between particles with the same charge

element a substance consisting of only one type of atom

emission spectroscopy a technique where the discrete frequencies of radiation emitted are analysed to provide qualitative and quantitative information about the sample

emission spectrum the distribution of the frequencies of electromagnetic radiation emitted by an atom which has been heated or excited

empirical formula a chemical formula showing the simplest ratio of elements in a compound rather than the total number of atoms in the molecule

endothermic reaction energy is absorbed from the surroundings, so the temperature falls

enthalpy change for a chemical reaction is the heat gained or lost per mole of reactant. Enthalpy change = (enthalpy of products) – (enthalpy of reactants)

enthalpy the total energy possessed by a chemical substance, which changes as bonds are broken or made during a chemical reaction

enzyme a biological catalyst

error propagation uncertainty produced as a result of combining measurements, each with uncertainties, in a calculation

estimate the outcome of a measurement

excited state an altered electron configuration from the ground state that occurs when an atom absorbs additional energy

exothermic reaction energy is released to the surroundings, so the temperature increases

experimental yield the actual yield (mass) of a product obtained for a reaction conducted in the laboratory

extrapolation to estimate a value for a data point that lies outside the range of measured values using a mathematical relationship based on the measured data, or extending a perceived sequence of values, and assuming that the trend identified will continue

F

fat food group consisting of chains of triglyceride molecules – they are essential for life

first ionisation energy the amount of energy required to remove a single electron from an atom in the gaseous form

fossil fuels decomposed organic matter under high temperature and pressure for a long time

fullerene a nearly spherical arrangement of 20–84 covalently bound carbon atoms

G

gas a substance (element, compound or mixture) that exists in a state with no defined volume or shape

gas chromatography (GC) a separation technique for small organic molecules that can withstand relatively high temperatures

gas pressure the force per unit area exerted by gas particles as they collide with the walls of their container

giant covalent network a substance made from a series of covalently bonded atoms extending indefinitely throughout a whole crystal; very hard substances with high melting points (e.g. diamond is a giant covalent network formed from the element carbon)

gradient the steepness of the line on a graph, which can be used as a measure of the rate of change of the quantity measured on the y-axis (the dependent variable) relative to that on the x-axis (the independent variable)

graphene a sheet of covalently bound carbon atoms

greenhouse effect trapping of out-going infrared radiation heating Earth's near-surface atmosphere

greenhouse gas gas that traps out- going infrared radiation

ground state the lowest energy electron configuration of an atom

group elements contained in a single vertical column in the periodic table

H

half-life the time taken for half of the number of radioactive nuclei in a sample to undergo decay

heat of combustion heat released when a fuel undergoes complete combustion at standard atmospheric pressure

heterogeneous mixtures mixtures of nonuniform composition

high-performance liquid chromatography a separation technique for large organic molecules that are unstable at high temperatures

homogeneous outwardly appearing to be made up of only one substance

homogeneous mixtures mixtures of uniform composition

Hund's rule of maximum multiplicity every orbital in a subshell is singly occupied with one electron before any one orbital is doubly occupied

hydrated ionic crystals with a specific number of water molecules in their chemical formulas

hydrocarbons compounds containing only hydrogen and carbon

hydrogen bonds occurs when an H atom that is bonded to an O, N or F atom in one molecule becomes attracted to the lone pair of electrons of N, O or F of an adjacent molecule

hypoxia deficiency in the amount of oxygen

I

ideal gas a gas that perfectly obeys all the proposals of the kinetic theory of gases and the gas laws

immiscible liquid a liquid that does not mix with another liquid

independent variable a variable upon which another variable depends

intermolecular bonding interactions that exist between molecules – these are much weaker than intramolecular bonds but are important in determining the physical properties of the molecule such as melting and boiling points and solubility

intermolecular forces attractive forces that cause molecules to aggregate together to form solids; they are significantly weaker than intramolecular forces, which are the ionic or covalent bonds that bond the atoms together inside a molecule

interpolation to read or construct a new data point that has not been measured but is within the range of measured data

interstitial alloy an alloy where the dopant occupies the space between main atoms

intramolecular bonding bonds that occur between atoms inside a molecule

ion an atom that has become charged through the gain or loss of an electron

ionic bond the mutual, nondirectional, electrostatic attraction of positively and negatively charged ions, which holds them together in a lattice in a compound

ionic compounds compounds consisting of atoms that have lost and gained electrons to form charged ions that are electrostatically attracted to one another

ionic radius the distance between the centre of the nucleus and the outermost electron shell in an ion

isotope two atoms of the same element with the same number of protons but with different masses, resulting from their different number of neutrons

isotopic composition a list of the different isotopes present in a sample of an element and the relative amount of each isotope that is present

IUPAC the International Union of Pure and Applied Chemistry

K

kinetic energy (KE) the energy possessed by moving particles: a particle of mass m, moving with speed v has a translational kinetic energy equal to $\frac{1}{2}mv^2$; molecules also possess rotational and vibrational energy.

L

latent heat the energy required to convert a solid into a liquid, or a liquid into a gas, without a change in temperature

lattice a regular repeating three-dimensional arrangement of ions, which results in the formation of crystals

level of precision the resolution of a measurement, typically defined by the number of decimal digits given for a measurement

Lewis structure a representation of the electron arrangement of atoms, using dots or crosses to represent electrons in their shells

limiting reagent the reactant that it is totally consumed when the chemical reaction is complete

M

malleability ability to be beaten into another shape or flattened into a thin sheet with a hammer without breaking

malleable a substance which can be repeatedly deformed, or hammered, without shattering

mass number the number of protons and neutrons in the nucleus of an atom

mass spectrometry a technique used to determine the atomic or molecular composition of an element and its isotopic composition

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mass spectrum a graph of the information from a mass spectrometer, showing the number of particles of different mass (or more precisely, mass:charge ratio) that are present in a sample and the relative abundance of each

Maxwell–Boltzmann distribution a frequency distribution of the energy value of a sample of particles at a particular temperature

measurement using a measurement procedure together with a measuring standard to determine a measurement

measurement procedure a defined strategy for making a measurement using a measuring device

measuring standard calibrated device for making measurements

metallic bonding the electrostatic forces of attraction that occur between cations and the electrons that are free to move within the regular lattice structure of a metal

molar mass mass of 1 mole of a substance (equals mass of $N_A = 6.022 \times 10^{23}$ atoms, molecules or ions of a substance)

molar volume of a gas the volume occupied by 1 mole of gas at a known temperature and pressure

molarity the number of moles of solute per litre of solution

mole the amount of a substance (g) containing $N_A = 6.022 \times 10^{23}$ particles (atoms, ions or molecules) of that substance

molecule a group of atoms covalently bonded together representing the smallest unit of an element or compound that can take part in a chemical reaction

monochromator filter used to ensure that only radiation of a specific frequency is shone at a sample

multiple covalent bond two atoms share more than one pair of electrons: a double bond consists of two shared pairs of electrons and a triple bond of three shared pairs

N

nanomaterial a substance that is made up of or incorporates particles in the range of 1–100 nm

nanoparticle a particle with a diameter less than 100 nm that behaves as a whole unit.

nanotechnology a branch of science dealing with particles in the range 1–100 nm

net ionic equation shows only the species that are taking part in the chemical reaction; spectator ions are not included

neutralisation a reaction that occurs when an acid is mixed with an alkali (or base), resulting in the formation of water

neutron an uncharged particle that exists in the nucleus of all atoms except hydrogen

noble gas electron configuration an arrangement of electrons that is identical to one of the elements in group 18, known as the noble gases

non-bonding or lone pair electrons a pair of electrons located on an individual atom and that are not shared

non-renewable fuels unable to be regenerated at a rate similar to which is consumed

nuclear fission splitting of heavy nuclei

nuclear fusion joining of small nuclei

nucleus the positively charged centre of an atom, containing positively charged protons and electrically neutral neutrons, to which the orbiting electrons are attracted

O

organic chemistry the study of carbon containing compounds

outlier a data point that does not fit the pattern shown by other measured data points; sometimes they are defined quantitatively as lying more than 1.5 times the interquartile range above the third quartile or below the first quartile

overall equation a simple equation showing the formulae and ratio of all compounds taking part in a chemical reaction, showing no ions

oxide a compound that forms when an element combines with oxygen

P

parallax uncertainty the uncertainty which arises from the change in the apparent position of an object (e.g. scale reading for a liquid level) when viewed from different points

Pascal (Pa) the SI unit of gas pressure, equivalent to a force of 1 newton per square metre

Pauli exclusion principle no two electrons can occupy the same quantum state; hence, two electrons in the same orbital will have opposing spins

percentage composition the percentage by mass of each of the different elements in the compound

percentage error

$$\text{percentage error} = \frac{(\text{accepted value} - \text{measured value})}{(\text{accepted value})} \times \frac{100}{1}\%$$

percentage uncertainty

$$\text{percentage uncertainty} = \frac{(\text{absolute uncertainty of measurement})}{(\text{measurement value})} \times \frac{100}{1}\%$$

period elements contained in a single horizontal row in the periodic table

periodic table of elements a table listing the known elements in order of atomic number, based on their chemical and structural properties, so that patterns and trends are easily recognisable

pH a scale used to measure the acidity of a solution: a solution of pH = 1 is strongly acidic; pH = 7 is neutral; and pH = 14 is strongly alkaline

phase change transformation of a substance between solid, liquid or gaseous forms

photon a particle representing a quantum of light

physical properties properties of a substance such as melting point and electrical conductivity that do not involve a chemical reaction

polyatomic ion groups of non-metallic atoms that chemically bond together and have an overall charge, which can be positive or negative: they can take part in ionic bonding

precipitate the solid that forms when two different liquids react as a result of the insolubility of at least one product

precipitation reaction a reaction between two solutions that results in the formation of a solid

precision the closeness of several independent measurements of the same quantity to each other

protein food group consisting of chains of amino acid molecules – they are essential for life

proton a positively charged particle that exists in the nucleus of all atoms

pure substance composed of only one substance

Q

qualitative analysis a technique that determines the identity of the sample, rather than the quantity

qualitative data/measurement quantities that are described by descriptive values and cannot be specified using a numeric scale

quantitative analysis a measurement of the amount of a substance present in a sample

qualitative data/measurement numerical data acquired through counting or measuring

R

radioactive decay the process by which an unstable nucleus breaks down, releasing energy and matter

random error a variation that affects a measurement in a random way so that the measurement is as likely to change in any one direction as in any other

rate how quickly one quantity (a dependent quantity) changes compared to another (independent quantity)

reaction time the time taken for a person or system to respond following a given stimulus or event

relative atomic mass the average mass of one atom of an element, relative to the mass of carbon-12

relative formula mass (for ionic compounds) the mass of one formula unit of an ionic compound on a scale in which the mass of an atom of the carbon-12 isotope is exactly 12

relative molecular mass the mass of one molecule of a molecular substance on a scale in which the mass of an atom of the carbon-12 isotope is exactly 12

relative standard uncertainty the standard uncertainty of a measurement result is the standard deviation of the measurement (as defined for a statistical normal distribution). The relative standard uncertainty is the ratio of the standard uncertainty and the measurement itself, i.e. the fractional standard uncertainty

renewable fuel have the potential to be replenished and are essentially limitless

reproducible obtaining the same result, within the limits of experimental uncertainty, after carrying out exactly the same measurement more than once

retardation factor (R_f) the distance travelled by the sample from its origin divided by the distance of the solvent front from the sample origin

rotational kinetic energy kinetic energy contained in rotational motion - determined by the mass and angular velocity

S

saturated compound an organic compound where each carbon is bonded to four different atoms, without any double bonds present

saturated solution a solution in which no more of a particular solute can dissolve in a given quantity of solvent

scatter graph a graph that uses horizontal (x) and vertical (y) axes to plot a set of data points represented by (x, y), without a line joining the points; used to demonstrate or determine a mathematical relationship between variables

significant figures all the nonzero digits of a number and the zeros that are included between them or that are final zeros and signify accuracy

solubility the amount of solute that can dissolve in 100 g water (solvent) at a particular temperature

solubility curve a graph showing the variation of the solubility of a substance in water with temperature

solute the substance that is dissolved in a solution

solution a homogenous mixture with a solute dissolved in a solvent, where the particles of the solute cannot be observed and the solution may be distinct in colour

solvent a substance (usually a liquid) in which another substance is dissolved

specific heat capacity heat energy required to raise the temperature of 1 gram of a substance by 1 kelvin

spectator ions ions that remain in a solution before and after a chemical reaction has taken place, which are not involved in the reaction

spectroscopy a method of analysis performed by subjecting a sample to electromagnetic radiation and observing the frequencies absorbed or emitted

spontaneous chemical reaction a reaction that occurs without the input of additional energy such as heat, light or electricity

standard atmospheric pressure atmospheric pressure on the surface of the earth, described as 1 atm, or 101.3 kPa

standard laboratory conditions (SLC) 25°C and 100 kPa

standard of measure an object, system, or experiment that bears a defined relationship to a unit of measurement of a physical quantity

standard temperature and pressure (STP) 0°C and 100 kPa

standards the fundamental reference for weights and measures, against which all other measuring devices are compared - administered by government agencies

stoichiometry the relationship between the relative quantities of substances taking part in a reaction or forming a compound, typically a ratio of whole integers

stoichiometry ratio where the mole ratios for reactant atoms or molecules present in a chemical equation are in the exact proportion for the reaction to occur with no reactant in excess

strong nuclear force an attractive force that occurs between all particles in the nucleus, regardless of charge

structural isomers two compounds that contain the same number of each atom, but are arranged in a different way

sublimation the process of a solid turning immediately into a gas when heated, without forming a liquid

subshell a part of an electron shell that contains orbitals of the same energy

substituent a carbon atom (or group of carbon atoms) that is not part of the longest chain of carbon atoms in an organic molecule

substitution a reaction where one atom of a compound is replaced by an atom of a different element

substitution alloy an alloy where the dopant has been substituted in the crystalline structure with the main atom

substrate molecule that is taking part in the reaction which the enzyme is catalysing

successive ionisation energies the energies required to remove the electrons from an atom, in sequence, starting with the outermost

supersaturated solution an unstable solution that has more than the maximum amount of solute dissolved in a given quantity of solvent; it has a higher amount of solute than the saturated solution

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surface area the area for a successful collision between particles to occur that enables a reaction to take place

surface tension the tension of the surface film of a liquid, due to the attraction between particles in the surface layer

suspension a mixture of a liquid and an insoluble solid, where the solid distributes evenly throughout the liquid

systematic error an error that acts to give a consistent offset in data; for example, a zero error

T

temperature a measure of the average kinetic energy of the particles that constitute a solid, liquid or gas

theoretical yield the amount (mass) of a product that is produced from the complete reaction of the limiting reagent

transition metal the section of the periodic table in groups 3–12 containing metallic elements

transition state see activated complex

translational kinetic energy given by KE (translational) = $\frac{1}{2}mv^2$

triple covalent bond a triple bond consists of three shared pairs of electrons

U

unsaturated compounds organic compounds containing at least one double bond between carbon atoms

unsaturated solution a solution with less than the maximum amount of solute dissolved in a given quantity of solvent

V

valence shell the outermost electron energy level, or shell, in an atom

valency the number of hydrogen atoms with which a single atom can bond when forming a compound (e.g. hydrogen has a valency of 1)

vaporisation a technique for separating the solvent from a solution, when the solute is required to be kept

vapour pressure the pressure of a substance in the gaseous state above its liquid

variable something that can change or be changed, as distinct from a constant, which does not

vibrational kinetic energy kinetic energy associated with the oscillation of atoms within a molecule or solid lattice as they vibrate.

W

Water of crystallisation water molecules present in a fixed ratio that form an essential part of the crystal structure of some compounds

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Key

element name
atomic number
symbol
atomic weight*

* standard atomic weight based on 12 C.
( ) indicates mass number of longest-lived isotope
For higher-precision values for atomic masses, visit ciaaw.org

- s-block
- p-block
- d-block transition metals
- d-block lanthanides and actinides

He gas at room temperature
Br liquid at room temperature
Tc synthetic (does not occur naturally)
Li solid at room temperature

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
hydrogen 1 **H** 1.008																	helium 2 **He** 4.003
lithium 3 **Li** 6.941	beryllium 4 **Be** 9.012											boron 5 **B** 10.81	carbon 6 **C** 12.01	nitrogen 7 **N** 14.01	oxygen 8 **O** 16.00	fluorine 9 **F** 19.00	neon 10 **Ne** 20.18
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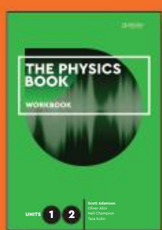

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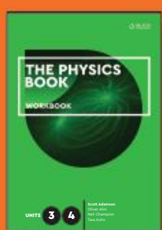

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