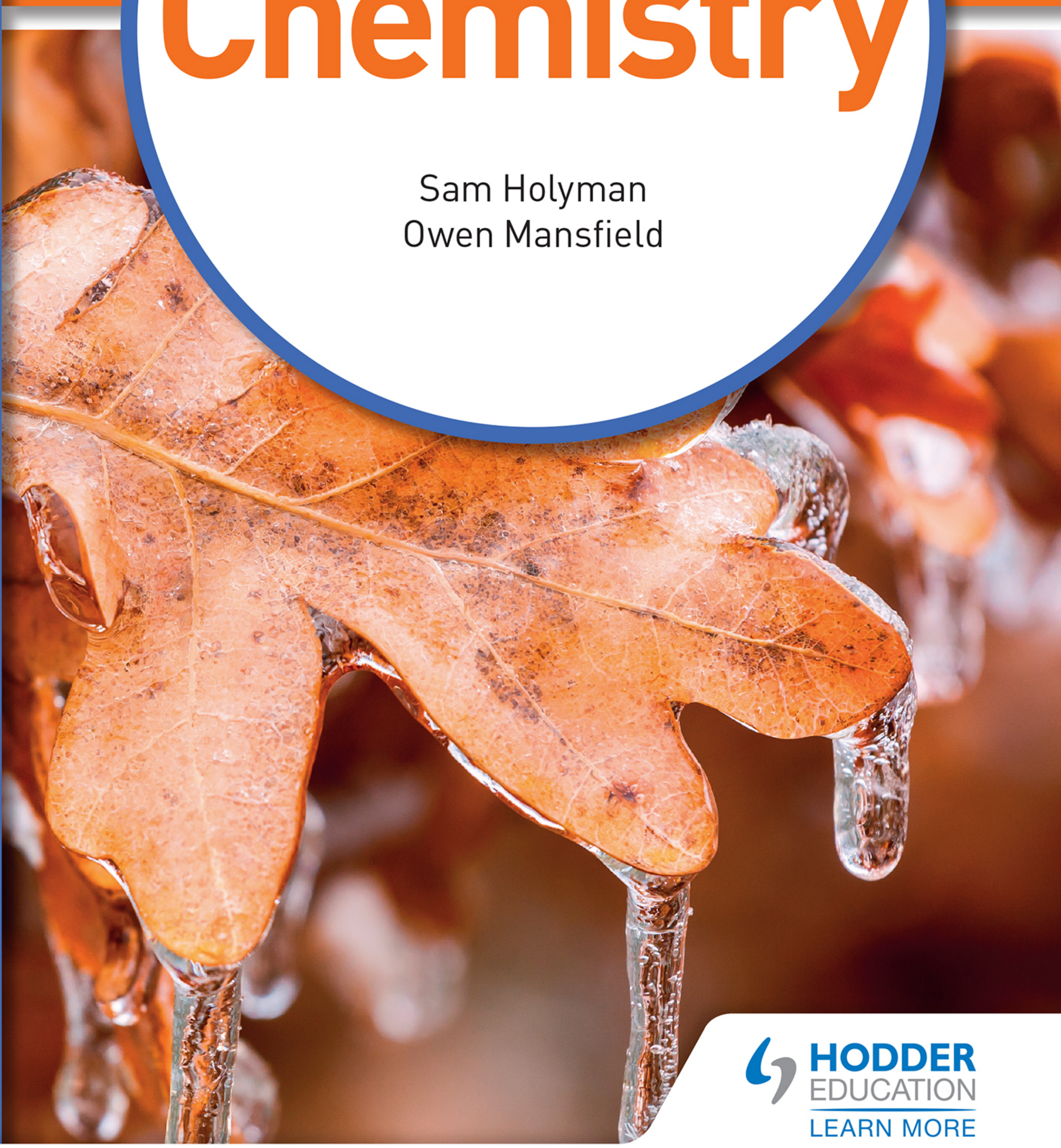
PRACTICE MAKES PERMANENT

350+ questions

AQA GCSE Chemistry

Sam Holyman
Owen Mansfield



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Introduction

Practice Makes Permanent is a series that advocates the benefits of answering lots and lots of questions. The more you practise, the more likely you are to remember key concepts; practice does make permanent. The aim is to provide you with a strong base of knowledge that you can automatically recall and apply when approaching more difficult ideas and contexts.

This book is designed to be a versatile resource that can be used in class, as home-work, or as a revision tool. The questions may be used in assessments, as extra practice, or as part of a SLOP (i.e. Shed Loads of Practice) teaching approach.

How to use this book

This book is suitable for the AQA GCSE Chemistry course, both at Higher and Foundation levels. It covers all the content that you will be expected to know for the final examination.

The content is arranged topic-by-topic in the order of the AQA specification, so areas can be practised as needed. Within each topic there are:

- **Quick questions** – short questions designed to introduce the topic.
- **Exam-style questions** – questions that replicate the types, wording and structure of real exam questions, but highly-targeted to each specification point.
- **Topic reviews** – sections of exam-style questions that test content from across the entirety of each topic more synoptically.

These topic questions are tagged with the following:

p64	page references for the accompanying Hodder Education Student Book: AQA GCSE (9-1) Chemistry, 978-1-4718-5134-6. This can be revisited before or after attempting the questions in a topic.
4.1.1.1	the AQA specification reference, which can be used if you want to practise specific areas.
H	indicates Higher-only content.
MS 5b	indicates where questions test Maths skills.
QWC	indicates where answers will also be marked on the quality of written communication.
WS 4.1	indicates where questions require you to work scientifically.
AT 1	indicates where questions ask you to use practical knowledge of apparatus and techniques.
RP 1	indicates where questions test understanding of required practicals.

At the end of the book there is a full set of **practice exam papers**. These have been carefully assembled to resemble typical AQA question papers in terms of coverage, marks and skills tested. We have also constructed each one to represent the typical range of demand in the GCSE Chemistry specification as closely as possible.

Full worked **answers** are included at the end of the book for quick reference, with awarded marks indicated where appropriate.

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1 Atomic structure and the periodic table

A simple model of the atom, symbols, relative atomic mass, electronic charge and isotopes

Quick questions

p2 4.1.1.1

1 What is the smallest particle of an element that can exist?

p11 & 278 4.1.1.1

2 For each of the following compounds, give their name and state what elements they contain: NaCl, MgO, H_2S, AlF_3, CuI_2.

p22–4 4.1.1.2 WS2.3

3 Separating mixtures relies on differences in physical properties of the components in the mixtures. For each of the following mixtures, choose either A or B to describe how the mixture is separated.

A: differences in solubility; B: differences in boiling points.

- a solid precipitate of lead iodide from a solution of potassium nitrate
- a mixture of cyclohexane (a liquid hydrocarbon) and water
- a solution of ethanol dissolved in water

p23–4 4.1.1.2 WS2.3

4 Name **one** technique used to separate a soluble solid from its solution.

p3–4 4.1.1.4

5 Define the terms 'atomic number' and 'mass number'.

p3 4.1.5 MS1b WS4.3

6 Give the approximate diameter of a typical atom, in metres, using standard form.

p4 4.1.1.5

7 Give the term for atoms with the same atomic number but different mass numbers.

p4 4.1.1.6

8 Define the term 'relative atomic mass'.

p5–6 4.1.1.7 MS5b WS1.2

9 The electronic structure of a sodium atom can be stated as 2.8.1. Give the electronic structures of an oxygen atom and a calcium atom.

Exam-style questions

p11&21 4.1.1.1

10 When sodium reacts with water, its products are sodium hydroxide and hydrogen gas.

10–1 Write a word equation for the reaction between sodium and water. *[1]*

10–2 The balanced symbol equation for the reaction between sodium and water is:

$2Na + 2H_2O \rightarrow 2NaOH + H_2$

Explain why hydrogen, H_2, is an element, whilst water, H_2O, is a compound. *[2]*

10–3 State the ratio of hydrogen to oxygen atoms in water. *[1]*

Total: 4

p23 | 4.1.1.2 | AT4 | WS2.3

11 **Figure 1** shows the apparatus used to separate a mixture of substances.

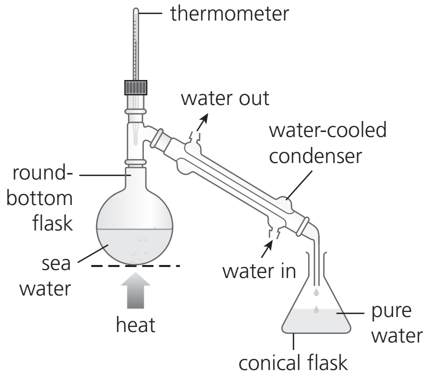


Figure 1

11–1 Name the separation process being carried out in **Figure 1**. *[1]*

11–2 Describe the role of the water-cooled condenser in this separation process. *[2]*

11–3 Ethanol and water are two liquids that could be separated using this technique. When a mixture of these two liquids is separated, the ethanol is the first liquid to arrive in the conical flask.

Explain why this occurs using ideas about boiling points. *[2]*

Total:5

p2–6 | 4.1.1.4 | WS1.2

12 Every element has its own type of atoms. For instance, neon atoms are different to magnesium atoms. Use a periodic table to answer the following questions.

12–1 Explain, in terms of sub-atomic particles, how neon atoms are different to magnesium atoms. *[3]*

12–2 Explain why a neon atom has no overall charge. *[3]*

12–3 When magnesium atoms react they lose their two outer-shell electrons. Determine the charge on the magnesium ions that form. Explain your answer. *[2]*

Total: 8

13 When magnesium burns in air, a small proportion of the magnesium reacts with nitrogen to form magnesium nitride. The ratio of magnesium to nitrogen particles in magnesium nitride is 3:2.

p278 | 4.1.1.1

13–1 Give the chemical formula for magnesium nitride. *[1]*

p6 | 4.1.1.7 | MS5b

13–2 Most nitrogen atoms have an atomic number of 7 and a mass number of 14. The nitrogen particle in magnesium nitride is the nitride ion, N^{3-}. Draw the electronic configuration of the nitride ion. *[2]*

p6 | 4.1.1.5

13–3 Give the number of protons, neutrons and electrons in a nitride ion, N^{3-}. *[3]*

Total: 6

p24 | 4.1.1.2 | WS2.7,3.5

14 Food colouring is often a mixture of different coloured substances dissolved in water. Some students separated a sample of brown food colouring. This is the method they used:

1. Draw a pencil line 1 cm from the base of a piece of chromatography paper.
2. Place a small spot of the brown food colouring onto the pencil line.
3. Fill a beaker with a 1.5 cm depth of water.
4. Stand the chromatography paper in the beaker.
5. Leave the paper in the beaker until the water has nearly reached the top of the paper.
6. Remove the paper and allow it to dry.

14–1 The students' experiment did not work. Suggest **one** improvement to step 3 that would ensure the experiment does work. *[1]*

14–2 Using an improved method, the students' chromatography experiment showed that the brown food colouring was made up of three different coloured substances. Describe how chromatography causes the brown food colouring to be separated into its three components. *[3]*

QWC

14–3 In another chromatography experiment, three more food colourings (A, B and C) were analysed. The chromatography paper is shown in **Figure 2**.

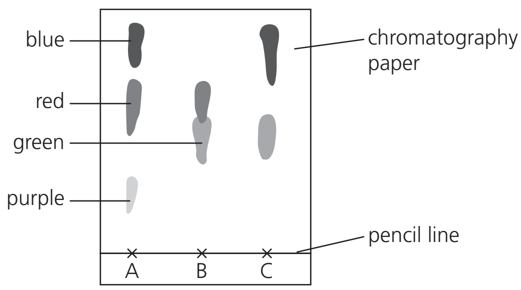


Figure 2

Use the information in **Figure 2** to compare the three food colourings. *[6]*

Total: 10

p4–8 | 4.1.1.5

15 The following particles all have the same electronic arrangement of 2.8.8:

- K^+
- Ca^{2+}
- Cl^-
- Ar
- S^{2-}

Use a periodic table to answer the following questions.

15–1 Which particle has the same number of protons as electrons? *[1]*

15–2 Which particle has 20 protons but only 18 electrons? *[1]*

15–3 Which particle has one more electron than protons? *[1]*

15–4 Which particle is an atom? *[1]*

15–5 An atom of phosphorus has an atomic number of 15 and a mass number of 31. Give the number of protons, neutrons and electrons in an atom of phosphorus. *[3]*

QWC

15–6 **Figure 3** shows a representation of an atom.

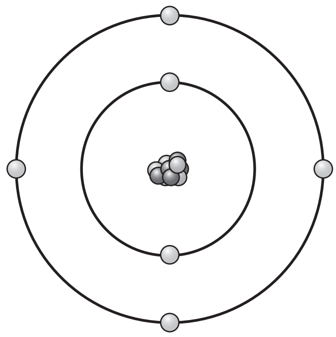


Figure 3

Describe how the work of Ernest Rutherford, Neils Bohr and James Chadwick changed the 'plum pudding' model of the atom to the model of the atom shown in **Figure 3**. *[6]*

Total: 13

p16&289 4.1.1.1

16 Potassium and fluorine combine in a ratio of 1 : 1 to form potassium fluoride.

16–1 Give the chemical formula for potassium fluoride. *[1]*

16–2 Write a balanced symbol equation for the reaction between potassium and fluorine (F_2). *[2]*

H 16–3 During this reaction, potassium atoms become potassium ions, K^+. Give the half equation, including electrons for this change. *[2]*

H 16–4 During this reaction fluorine molecules (F_2) become fluoride ions, F^-. Give the half equation, including electrons for this change. *[3]*

Total: 8

p4–5 4.1.1.6 MS1d,2a

17 Silicon's three most stable isotopes are ^{28}Si, ^{29}Si and ^{30}Si.

17–1 Define the term 'isotope'. *[3]*

17–2 Silicon has a relative atomic mass of 28.1. Choose the most abundant isotope from the list of silicon's three most stable isotopes. Explain your answer. *[3]*

17–3 Chlorine has two stable isotopes with mass numbers of 35 and 37. The relative atomic mass of chlorine is 35.5. Estimate the relative abundances of these two isotopes and choose the most accurate option from the list below: *[1]*

A 25% ^{35}Cl and 75% ^{37}Cl

B 50% ^{35}Cl and 50% ^{37}Cl

C 55% ^{35}Cl and 45% ^{37}Cl

D 75% ^{35}Cl and 25% ^{37}Cl

Total: 7

p4–5 4.1.1.6 MS1d,2a

18 The three most common isotopes of sulfur, and their relative abundance, are shown in **Table 1**.

Isotope	Relative abundance (%)
^{32}S	95.0
^{33}S	0.8
^{34}S	4.2

Table 1

18–1 ^{32}S and ^{34}S are the two most common isotopes of sulfur. Describe the similarities and differences between these two isotopes. *[3]*

18–2 Calculate the relative atomic mass of sulfur atoms using the information in **Table 1**. Give your answer to three significant figures. *[3]*

Total: 6

The periodic table

Quick questions

p9–12 4.1.2.1 WS1.2

1 Element X has an atomic number of 17.

- Which group of the periodic table will this element belong to?
- Give the atomic number of the element that would be directly above element X in the periodic table.

p19–21 4.1.2.1 WS1.1,1.6

2 Why did Mendeleev leave gaps in his version of the periodic table?

p19–21 4.1.2.2 WS1.6

3 How did Mendeleev's version of the periodic table benefit from the predictions he made about undiscovered elements?

p10 4.1.2.3

4 An element reacts by losing its outer-shell electrons to form positively charged ions. Is this element a metal or a non-metal?

p13 4.1.2.4 WS1.2

5 What do the Group 0 noble gases all have in common?

p14–15 4.1.2.5 WS1.2

6 Describe the trend in reactivity of the Group 1 metals.

p16–17 4.1.2.6 WS1.2

7 Describe the trend in boiling points of the Group 7 halogens.

Exam-style questions

8 The position of an element in the periodic table is determined by the element's atomic number and electron structure. Carbon has an atomic number of 6 and silicon has an atomic number of 14. Both elements are in Group 4 of the periodic table.

p9–12 | 4.1.2.1 | WS1.2

8–1 Explain why carbon is positioned in Group 4 of the periodic table. *[1]*

p5–6 | 4.1.1.7 | MS5b

8–2 **Figure 4** shows some electron structures. Choose which diagram, A–D, represents a silicon atom. *[1]*

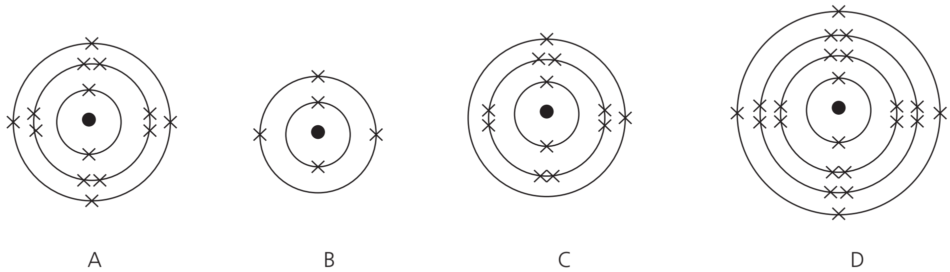


Figure 4

p9–12 | 4.1.2.1 | WS1.2

8–3 Carbon reacts with hydrogen to form the compound methane, CH_4. Predict the formula of the compound formed when silicon reacts with hydrogen. Explain your answer in terms of electron structure. *[3]*

Total: 5

p16–17 | 4.1.2.6

9 The halogens are in Group 7. Astatine is the fifth halogen, but is difficult to study because it is highly radioactive. Its properties can be predicted from the properties of the first four halogens. **Table 2** shows the appearances of the first four halogens.

Halogen	Appearance at room temperature
Fluorine (F_2)	pale yellow gas
Chlorine (Cl_2)	pale green gas
Bromine (Br_2)	dark brown liquid
Iodine (I_2)	grey solid

Table 2

WS1.2

9–1 Predict the appearance of astatine at room temperature. *[1]*

Figure 5 shows a graph of the boiling points of the halogens against their relative molecular mass:

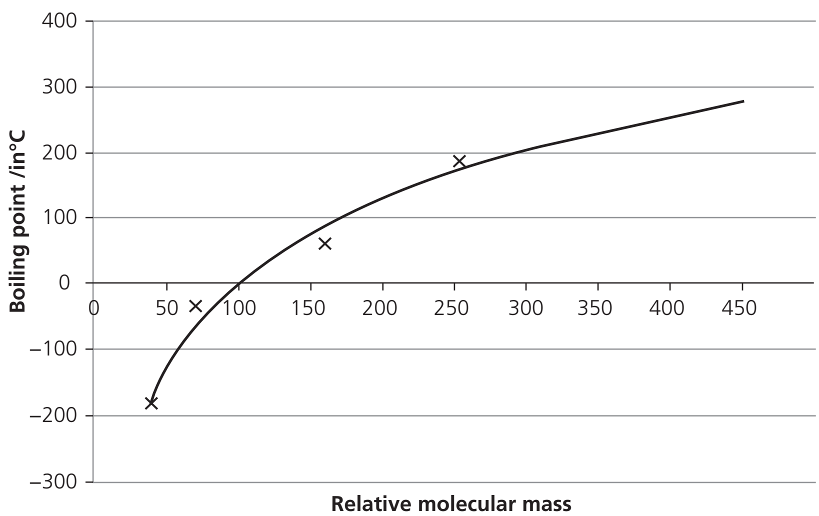


Figure 5

MS4a WS3.5 **9–2** Describe the trend in the boiling points of the halogens. *[1]*

MS4a WS3.2 **9–3** Use **Figure 5** to predict the boiling point of astatine. *[1]*

WS1.2 **9–4** The reactivity of the halogens decreases down the group. Predict the products formed when a solution of sodium astatide is mixed with a solution of chlorine. *[1]*

9–5 During the reaction between sodium astatide and bromine, astatide ions, At^-, would lose electrons to form astatine molecules, At_2. Show this change by writing the half equation, including electrons. *[3]*

9–6 Explain why chlorine would be more reactive than astatine. *[4]*

Total: 11

10 Lithium, an alkali metal in Group 1, reacts with chlorine, Cl_2, to form lithium chloride, LiCl.

p16–17 &283 4.1.1.1 **10–1** Write a balanced symbol equation for the reaction between lithium and chlorine. *[2]*

p289–90 4.1.1.1 H **10–2** During this reaction lithium atoms lose their outer-shell electron to form lithium ions. Show this change by writing and balancing the half equation, including electrons. *[3]*

p14–15 4.1.2.5 QWC WS1.2 **10–3** Potassium reacts with chlorine in a similar way to lithium, but the reaction releases more energy and occurs more rapidly. Explain why potassium reacts in a similar way to lithium, but is more reactive. Include the electronic arrangements of lithium and potassium in your explanation. *[6]*

Total: 11

11 **Table 3** shows how the boiling points of noble gases change with the relative masses of their atoms.

Element	Relative mass of atom	Boiling point in °C
Helium	4	−269
Neon	20	−246
Argon	40	−190
Krypton	84	X
Xenon	131	−111
Radon	222	−62

Table 3

p13 4.1.2.4 WS3.5 **11–1** Describe the trend in the boiling points of the noble gases. *[1]*

p13 4.1.2.4 WS3.2 **11–2** Use the data to predict the boiling point of krypton. *[1]*

p13& 16–17 4.1.2. 4/6 QWC WS1.2, 3.5 **11–3** The noble gases and the Group 7 halogens are both non-metal elements. Use your own knowledge to compare the difference in reactivity of the noble gases and halogen elements. Explain the difference in terms of the electronic arrangements of the two groups of elements. *[6]*

Total: 8

Properties of transition metals

Quick questions

p18–19 4.1.3.1/2

1 State **one** chemical property of transition metals that is not shared by Group 1 metals.

p18–19 4.1.3.1/2

2 State **one** physical property of transition metals that is not shared by Group 1 metals.

Exam-style questions

3 Sodium (Na) is a metal in Group 1 and reacts rapidly with oxygen gas at room temperature whilst nickel (Ni) is a transition metal which must be heated to high temperatures before it reacts.

p283 4.1.1.1

3–1 Nickel reacts with oxygen gas to form nickel oxide, NiO. Give the balanced symbol equation for this reaction. *[2]*

p6 4.1.2.3, 4.1.1.4

3–2 Metal atoms react by losing their outer-shell electrons to form positively charged ions. Explain why an atom becomes positively charged when it loses an electron. Answer in terms of sub-atomic particles. *[3]*

p15 4.1.2.5, 4.1.3.1

3–3 Explain why nickel is less reactive than sodium with oxygen. *[1]*

p289–90 4.1.1.1

H 3–4 When nickel reacts, its atoms form Ni^{2+} ions. Give the half-equation, including electrons, to show how a nickel atom forms a nickel ion. *[2]*

Total: 8

4 Manganese and rubidium are both metals. Manganese is a transition metal. Rubidium is a Group 1 alkali metal.

p18–19 4.1.3.1

4–1 Manganese and rubidium share several properties with each other. Choose **one** property from the list below that is common to both metals. *[1]*

- magnetism
- shiny when polished
- low reactivity
- catalytic behaviour

p18–19 4.1.3.2

4–2 Manganese is often used in glass making. Give **one** property of transition metals that makes manganese useful in glass making. *[1]*

Table 4 shows some of the properties of manganese and lithium.

	Manganese	**Lithium**
Density in g/cm³	7.21	0.53
Mohs hardness*	6.0	0.6
Formula of oxide	MnO MnO_2 Mn_2O_3 Mn_2O_7	Li_2O

**Mohs hardness is a quantitative scale of hardness. The greater the value, the harder the material.*

Table 4

p18–19 4.1.3.1/2 QWC WS3.5

4–3 Use the data in **Table 4** and your own knowledge to compare the chemical and physical properties of transition metals and Group 1 metals. *[6]*

Total: 8

Atomic structure and the periodic table topic review

1 Manganese sulfide is a naturally occurring compound of manganese. **Figure 6** shows the structure for manganese sulfide.

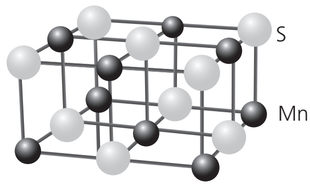


Figure 6

p277–9 4.1.1.1 WS1.2

1–1 Use **Figure 6** to determine the formula of manganese sulfide. *[1]*

p18–19 4.1.3.1

1–2 Manganese is found in the central block of the periodic table. It has an atomic number of 25 and a mass number of 55. What type of metal is manganese? *[1]*

p4–6 4.1.1.5 WS1.2

1–3 Give the number of protons, neutrons and electrons in an atom of manganese. *[3]*

p18–19 4.1.3.2

1–4 Give three typical **chemical** properties of manganese, based on its position in the periodic table. *[3]*

p5–6 4.1.1.7

1–5 Sulfur is a non-metal in Group 6 of the periodic table. The atomic number of sulfur is 16. Give the electron arrangement for an atom of sulfur. *[1]*

p6 4.1.1.4

1–6 When sulfur atoms react they tend to gain two electrons to form sulfide ions, S^{2-}. Explain why sulfide ions have a 2– charge. Answer in terms of sub-atomic particles. *[3]*

Total: 12

p7–8 4.1.1.3 WS1.1,1.2

2 Atomic theory has developed over time from a very simple model of the atom to the more complicated one that we use today.

The following list of terms are all related to atoms and sub-atomic particles:

- nucleus
- protons
- neutrons
- electrons
- electron shells

2–1 Which structure is most associated with the work of Neils Bohr? Choose **one** word from the list. *[1]*

2–2 Which particle is most associated with the work of James Chadwick? Choose **one** word from the list. *[1]*

2–3 Which structure is most associated with the work of Ernest Rutherford? Choose **one** word from the list. *[1]*

2–4 **Figure 7** is a representation of the alpha particle scattering experiment carried out by Hans Geiger and Ernest Marsden.

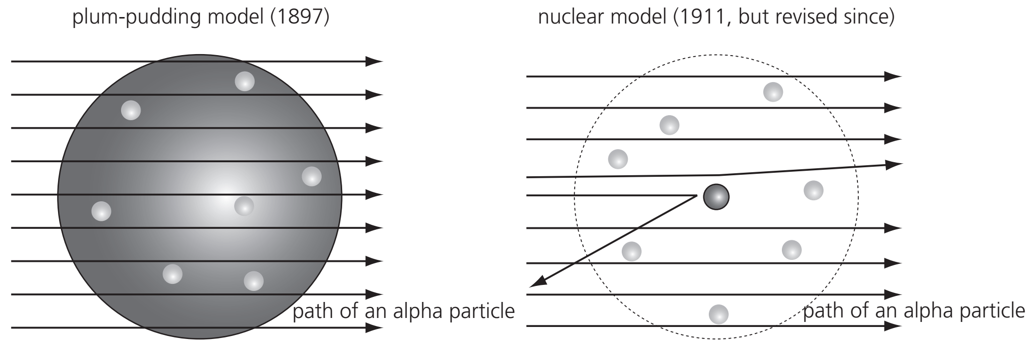


Figure 7

Use **Figure 7** and your own knowledge to describe how evidence from the scattering experiment led to a change in the atomic model. *[4]*

Total: 7

3 The halogens are elements in Group 7 of the periodic table.

p16 4.1.2.6 **3–1** Describe the relationship between the relative molecular mass and boiling points of the Group 7 elements. *[1]*

p283–4 4.1.1.1 **3–2** When bromine reacts with iron, the compound iron bromide, $FeBr_3$ is formed. Give the balanced equation for this reaction. *[2]*

p16–17 4.1.2.6 QWC **3–3** When iodine reacts with iron the reaction releases less energy and occurs much more slowly than the reaction of bromine and iron. Compare these two reactions. Use your knowledge to explain each of their similarities and differences. *[6]*

p17 4.1.2.6 **3–4** Name the products formed when chlorine reacts with a solution of iron bromide. *[1]*

p22–3 4.1.1.2 WS2.3 **3–5** Name a suitable method for separating soluble iron bromide from its solution. *[1]*

Total: 11

4 The alkali metals make up Group 1 of the periodic table.

p14 4.1.2.1 WS1.2 **4–1** Explain why the alkali metals are found in Group 1 of the periodic table. *[1]*

p10 4.1.2.3

4–2 Explain how it is possible to know that sodium is a metal and chlorine is a non-metal. You should make reference to the elements' positions in the periodic table, as well as the ions formed when the elements react. *[3]*

p14–15 4.1.2.5

4–3 The alkali metals react readily with water. Name the **two** products that form when sodium reacts with water. *[2]*

p14–15 4.1.2.5 WS1.2

4–4 Lithium metal reacts less vigorously with water than sodium. Explain why lithium is less reactive than sodium. *[4]*

Total: 10

p19–21 4.1.2.1

5 **Figure 8** shows part of Newlands' periodic table.

Column	1	2	3	4	5	6	7
	H	Li	Be	B	C	N	O
	F	Na	Mg	Al	Si	P	S
	Cl	K	Ca	Cr	Ti	Mn	Fe

Figure 8

5–1 Suggest why Newlands grouped the elements into columns. *[1]*

5–2 Explain why Newlands' periodic table was not widely accepted by other chemists of the time. Use your knowledge and **Figure 8**. *[3]*

5–3 **Figure 9** shows part of Mendeleev's periodic table.

H						
Li	Be	B	C	N	O	F
Na	Mg	Al	Si	P	S	Cl
K	Ca		Ti	V	Cr	Mn

Figure 9

Suggest why the Group 0 elements do not appear in either Newlands' or Mendeleev's periodic tables. *[2]*

WS1.6

5–4 Explain how Mendeleev's periodic table was an improvement upon Newlands' periodic table and why his ideas were more widely accepted. *[4]*

Total: 10

2 Bonding, structure and the properties of matter

Chemical bonds: ionic, covalent and metallic

Quick questions

p11 4.2.1.1

1 What type of chemical bond forms when a substance is formed from:
- two non-metals
- two metals
- a metal and a non-metal?

p47 4.2.1.1

2 For each of these types of chemical bond, state whether electrons are shared between atoms, or transferred from one atom to another.
- ionic bonding
- metallic bonding
- covalent bonding

p34 4.2.1.3 MS5b

3 Name the type of compound consisting of a lattice structure of oppositely charged ions.

p42 4.2.1.4 WS1.2

4 Name the type of bonding that occurs between nitrogen and hydrogen atoms in molecules of ammonia, NH_3.

p42 4.2.1.4 MS1c,5b

5 Draw a dot and cross diagram for each of the following molecules:
- hydrogen chloride, HCl
- water, H_2O
- ammonia, NH_3

p45 4.2.1.5 MS5b

6 Draw a simple two-dimensional diagram to represent metallic bonding. Label your diagram.

Exam-style questions

p47–8 4.2.1.1

7 Sodium, chlorine and hydrogen are three very reactive elements. These elements can react to form different types of compounds with different structures.

7–1 Give the type of bonding present in sodium. *[1]*

7–2 Give the type of bonding present in Cl_2 molecules. *[1]*

QWC

7–3 Sodium hydride, NaH, forms when sodium reacts with hydrogen. Hydrogen chloride, HCl, forms when chlorine reacts with hydrogen. Compare the bonding and structure of NaH to that of HCl. Explain your answer in terms of particles and electrons. *[6]*

Total: 8

8 The ionic compound magnesium fluoride forms when magnesium metal reacts with the non-metal fluorine. Magnesium is in Group 2 and fluorine is in Group 7 of the periodic table.

p35–6 | 4.2.1.1 | MS1c

8–1 Determine the formula of magnesium fluoride. *[2]*

p37–8 | 4.2.1.2

8–2 Both magnesium and fluorine atoms form ions with the electronic structure of a noble gas. Draw the electronic structure of both ions and give the name of the noble gas that has the same electronic structure. *[3]*

p34 | 4.2.1.3 | MS5b

8–3 **Figure 1** shows the structure of magnesium fluoride represented using the ball and stick model.

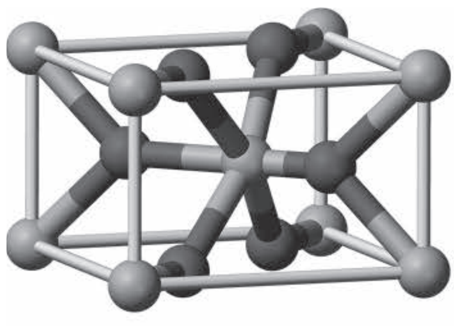


Figure 1

Give **one** limitation of representing this structure using this model. *[1]*

p33& 37–8 | 4.2.1.2

8–4 Describe the changes that take place in the magnesium and fluorine atoms when they react together to form ions. *[5]*

Total: 11

9 Silicon reacts with hydrogen to form a covalent compound made up of small molecules. Silicon reacts with oxygen to form a covalent compound with a giant structure.

p283–4 | 4.1.1.1 | MS1c

9–1 Write the balanced symbol equation for the reaction between silicon and hydrogen gas to form silane, SiH_4. You do not need to include state symbols. *[2]*

p42 | 4.2.1.4 | MS4a,5b

9–2 Copy and complete the dot and cross diagram in **Figure 2** to show the covalent bonding in a molecule of SiH_4. *[1]*

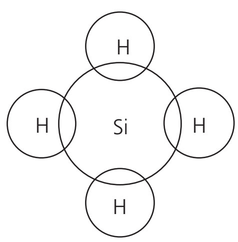


Figure 2

p39–40 | 4.2.1.4 | WS1.2

9–3 **Figure 3** shows the structure of silica, SiO_2.

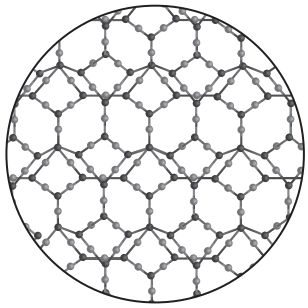


Figure 3

- Suggest what each small sphere represents in **Figure 3**. *[1]*
- Suggest what each line represents. *[1]*

Choose your answers from the options below.

an ion	an atom	a delocalised electron	a pair of electrons

p34 | 4.2.1.3 | MS5b

9–4 **Figure 4** shows two representations of a giant structure. One is a ball and stick structure and the other is a three-dimensional structure.

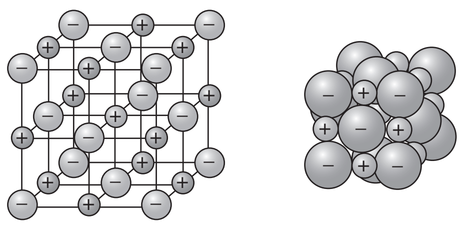


Figure 4

Give **one** advantage of a three-dimensional space filling representation over a ball and stick diagram. *[1]*

p41 | 4.2.1.4 | WS1.2

9–5 Pyrosilicic acid is a covalently bonded compound containing hydrogen, silicon and oxygen.

Figure 5 shows the stick diagram for a molecule of pyrosilicic acid.

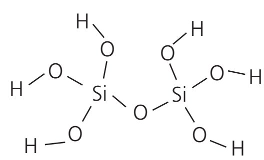


Figure 5

Give the molecular formula for pyrosilicic acid. *[1]*

Total: 7

10 **Figure 6** is a representation of the bonding in the element sodium.

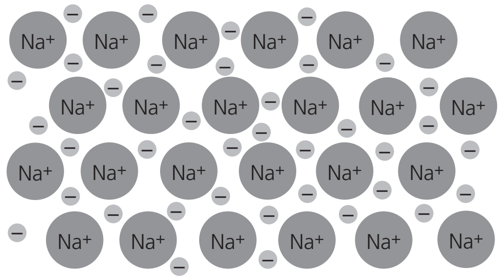


Figure 6

p45 | 4.2.1.5 | WS1.2 — **10–1** Describe fully the type of bonding represented by **Figure 6**. *[5]*

p283–4 | 4.1.1.1 | MS1c — **10–2** Sodium can be made by heating sodium carbonate, Na_2CO_3, with carbon, C. The products of the reaction are sodium and carbon monoxide, CO. Write the balanced symbol equation for this reaction. *[2]*

Total: 7

How bonding and structure are related to the properties of substances

Quick questions

p49 | 4.2.2.1 — 1 Which of the following compounds has the strongest forces of attraction between its particles?

- Methane, CH_4, boiling point = –162 °C
- Silica, SiO_2, boiling point = 2230 °C
- Sodium chloride, NaCl, boiling point = 1465 °C

p48 | 4.2.2.3 — 2 What type of particle is responsible for carrying the charge when molten sodium chloride conducts electricity?

p48 | 4.2.2.8 | MS5b — 3 What particles are responsible for carrying the charge when solid aluminium conducts electricity?

p40&44 | 4.2.2.4/6 | WS1.2 — 4 Why are most covalent substances unable to conduct electricity?

p34&46 | 4.2.2.3/7 | MS1c,5b — 5 Give **one** property of metallic substances that is not shared by ionic substances.

Exam-style questions

6 **Table 1** shows the melting and boiling points of three elements.

Element	Melting point in °C	Boiling point in °C
Nitrogen (N_2)	–210	–196
Oxygen (O_2)	–219	–183
Argon (Ar)	–189	–186

Table 1

6–1 Give the state of matter of argon at –190 °C. *[1]*

6–2 Which element is a liquid over the widest range of temperature? *[1]*

p49–50 4.2.2.1 WS1.2 MS5b

6–3 A mixture of these elements was cooled to −220 °C and then warmed slowly. Predict which element would boil first. Explain your answer. *[2]*

p39–40 4.2.2.4 WS1.2 MS5b

6–4 A student thinks that nitrogen and oxygen consist of very small molecules. Give evidence from **Table 1** to support this idea. *[3]*

Total: 7

7 Carbon dioxide, CO_2, is a gas at room temperature while silicon dioxide (silica), SiO_2, is a solid at room temperature. Silica has a very high melting point.

p39&48 4.2.1.1

7–1 Give the type of bonding that occurs between atoms of carbon and oxygen. Explain your answer. *[2]*

p48 4.2.2.4

7–2 Both carbon dioxide and silica have poor electrical conductivity. Suggest a reason for this. *[1]*

p39&44 4.2.2.4, 4.2.2.6 MS5b QWC

7–3 Explain in terms of structure and bonding why carbon dioxide and silica are different states at room temperature. *[6]*

Total: 9

8 Many metals and alloys are used for a wide variety of purposes. The bonding and structure of metals and alloys make them very useful materials. **Figure 7** is a representation of the bonding in a pure metal.

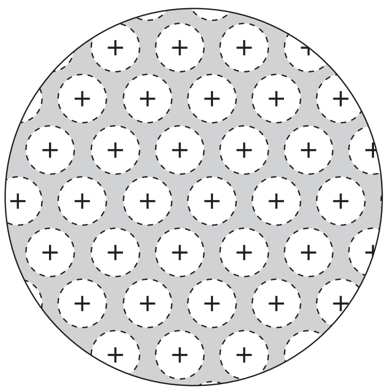


Figure 7

p46 4.2.2.8

8–1 Explain why metals are good electrical and thermal conductors. Give your answer in terms of structure and bonding. *[3]*

p46 4.2.2.7 WS1.2 MS5b QWC

8–2 Explain in terms of structure and bonding why metals are malleable while alloys tend to be harder than pure metals. Include a diagram in your answer. *[6]*

Total: 9

9 **Table 2** shows the properties of three different types of compound.

Compound	Melting point in °C	Boiling point in °C	Electrical conductivity as a solid	Electrical conductivity as a liquid
A	1085	2562	good	good
B	1326	2000	poor	good
C	1713	2950	poor	poor

Table 2

p49 | 4.2.2.1 | MS5b

9–1 Suggest a reason why all three substances have high melting and boiling points. *[1]*

p47–8 | 4.2.2.3/6/7 | WS1.2 | MS5b

9–2 Explain the difference in the electrical properties of these three compounds. Give your answer in terms of structure and bonding. *[6]*

Total: 7

10 Potassium bromide, KBr, can be made directly by reacting potassium with bromine. Potassium is in Group 1 of the periodic table. Bromine is in Group 7 of the periodic table.

Figure 8 shows a representation of how the particles are arranged in solid potassium bromide.

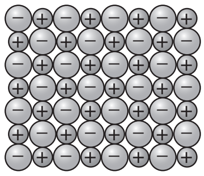


Figure 8

p34 | 4.2.1.3 | WS1.2 | MS5b

10–1 Describe the type of bonding and structure represented in **Figure 8**. *[3]*

p34–5 &48 | 4.2.2.3

10–2 Potassium bromide does not conduct electricity when solid, but does conduct electricity when molten or dissolved in water. Explain the electrical properties of potassium bromide. Answer in terms of structure and bonding. *[3]*

p49 | 4.2.2.3 | WS3.5, 1.2

10–3 **Table 3** shows the melting points of potassium bromide and some of the other Group 1 metal bromides.

A student says, "Larger ions form weaker forces of attraction than smaller ions". Do you agree? Use the data in Table 3 and your knowledge of Group 1 elements to justify your answer. *[4]*

Compound	Melting point in °C
Sodium bromide (NaBr)	747
Potassium bromide (KBr)	734
Rubidium bromide (RbBr)	693
Caesium bromide (CsBr)	636

Table 3

Total: 10

Structure and bonding of carbon

Quick questions

p54–6 | 4.2.3.1-3 | MS5b | WS1.2

1 **Figure 9** shows four structures of carbon.

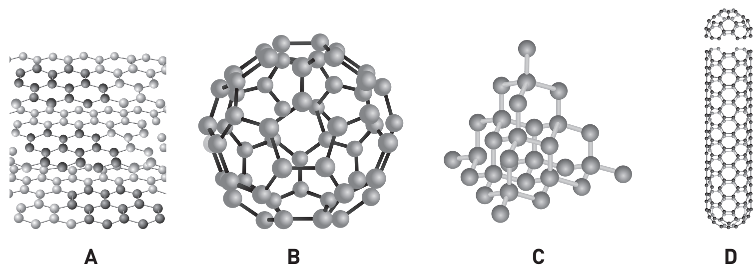


Figure 9

- Which structure represents a carbon nanotube?
- Which structure represents graphite?
- Which structure represents the first fullerene to be discovered?
- Which structure represents diamond?

p55 | 4.2.3.1 | WS1.2

2 Diamond is a very hard substance. Give **one** use of diamond that takes advantage of this property.

p56 | 4.2.3.3

3 State the property of carbon nanotubes that make them useful in sports equipment.

Exam-style questions

p54–6 | 4.2.3.1-3 | MS5b | WS1.2,1.4

4 Carbon exists in different forms.

4–1 Describe the structure and bonding in graphene. You may use a diagram in your answer. *[2]*

4–2 Explain why graphene and fullerene are able to conduct electricity. *[1]*

4–3 What structural feature of magnesium fluoride makes it suitable for delivering drugs to specific parts of the body? *[1]*

4–4 What structural feature of fullerenes makes them suitable for use as lubricants? *[1]*

4–5 Carbon nanotubes have high tensile strength, which means that they are very strong when pulled or stretched. Explain why this is. Write about structure and bonding in your answer. *[2]*

QWC

4–6 Diamond is a hard material that does not conduct electricity. Graphite is a soft material that does conduct electricity. Explain the differences in the properties of these two substances. Write about structure and bonding in your answer. *[6]*

Total: 13

Bulk and surface properties of matter including nanoparticles

Quick questions

p51–2 | 4.2.4.1 | MS1b,2h | WS4.2,4.3, 4.4,4.5

1 Nanoparticles are defined as particles with diameters between 1 and 100 nm. Write this range of diameters in metres in standard form.

p51–2 | 4.2.4.1 | MS1b, 2h,5c | WS4.2,4.3, 4.4,4.5

2 A cubic nanoparticle of zinc oxide has sides that are 15 nm long.

- Calculate the total surface area of this cube. Give your answer in nm^2.
- Calculate the total volume of this cube. Give your answer in nm^3.
- Calculate the surface area to volume ratio of this nanoparticle.

p52–3 | 4.2.4.2 | WS1.4

3 Give **three** uses of nanoparticles.

Exam-style questions

p51–2 | 4.2.4.1 | WS1.2,1.4

4 Silver nanoparticles are used in clothing due to their anti-bacterial properties. Using coarse particles of silver in clothing is not effective in killing the bacteria that cause unpleasant odours.

pxx | 4.2.4.1 | WS1.2,1.4

4–1 Explain why silver nanoparticles have different properties from those of larger silver particles. *[1]*

p52–3 | 4.2.4.2 | WS1.4

4–2 Explain why adding silver nanoparticles to clothing does not increase the cost of the clothing significantly. *[1]*

p51–2 | 4.2.4.1 | MS1b | WS4.3, 4.4,4.5

4–3 The diameter of a silver atom is 1.7×10^{-10} m. Calculate the number of silver atoms that fit side-by-side along the edge of a 20 nm wide silver nanoparticle. Give your answer to the nearest whole number. *[2]*

p51–2 | 4.2.4.1 | MS1b,1c, 2h,5c | WS4.3, 4.4,4.5

4–4 Show that the surface area to volume ratio of a cubic nanoparticle with sides of length 20 nm is ten times greater than for a cubic nanoparticle with sides of length 200 nm. *[4]*

p53 | 4.2.4.2 | WS1.3, 1.4,1.5

4–5 As a bulk material silver is an unreactive metal that has no known risks to human health. Suggest why some people are concerned about the possible health effects of using silver nanoparticles in clothing. *[2]*

Total: 10

Bonding, structure and the properties of matter topic review

1 Soft drink cans are made from sheets of the Group 3 metal aluminium. The sheets are cut and pressed into the shapes required to make the cans.

p45–6 | 4.2.1.5, 4.2.2.7

1–1 Explain why sheets of aluminium metal can be pressed into different shapes. Use ideas about the structure and bonding of aluminium in your answer. You may include a diagram. *[3]*

p45–6 | 4.2.2.8

1–2 Overhead power cables used in the National Grid are made from aluminium.

Explain why aluminium is able to conduct electricity. *[2]*

p37–8 4.2.1.2

1–3 The surface of an aluminium power cable is protected from corrosion by a layer of aluminium oxide that forms naturally. Aluminium oxide is an ionic compound with a very high melting point.

Describe how an ionic bond forms when aluminium reacts with oxygen. *[3]*

p34 4.2.2.3

1–4 Give the reason for the high melting point of aluminium oxide. *[1]*

p34 4.2.2.3

1–5 Explain why the aluminium oxide that protects the power cables does not conduct electricity. *[2]*

p51–2 4.2.4.1 WS1.2, 1.4,4.1

1–6 Aluminium oxide nanoparticles can be used to improve the properties of ceramic materials.

Explain why nanoparticles of aluminium oxide have different properties to those of the bulk material. *[1]*

p51–2 4.2.4.1 WS4.2,4.3, 4.4,4.5 MS1b,1c,5c

1–7 Calculate the surface area to volume ratio of a cubic nanoparticle with a side length of 30 nm. *[2]*

Total: 14

2 Carbon forms many different covalent substances with very different properties. Methane, CH_4, buckminsterfullerene, C_{60}, and silicon carbide, SiC, are three examples. Some properties of methane and silicon carbide are given in **Table 4**.

Compound	Melting point in °C	Boiling point in °C	Electrical conductivity
Methane	–182	–164	poor
Silicon carbide	2830		semi-conductor

Table 4

p41–2 4.2.1.4 WS1.2

2–1 Draw the dot and cross diagram for a molecule of methane, CH_4. *[1]*

p39–40 4.2.2.4 WS1.2

2–2 Explain the properties of methane shown in **Table 4**. Give your answer in terms of structure and bonding. *[3]*

p55–6 4.2.3.3 WS1.2 MS5b

2–3 Recent research suggests that methane molecules can be stored inside fullerenes such as buckminsterfullerene, C_{60}. Describe the property of fullerenes that enables them to contain small molecules within their structures. *[1]*

p54 4.2.3.1 WS1.2

2–4 Silicon carbide has similar bonding and structure to diamond, but its melting point is significantly lower than for diamond. Describe the structure and bonding of diamond. *[2]*

p49 4.2.2.1 WS1.2

2–5 Which of the statements below explains the difference in melting points of silicon carbide and diamond? *[1]*

- The bonds between the atoms in silicon carbide are stronger than in diamond.
- The bonds between the ions in silicon carbide are stronger than in diamond.
- Diamond contains delocalised electrons that strengthen the forces of attraction between the carbon atoms.
- The bonds between the atoms in silicon carbide are weaker than in diamond.
- The bonds between the ions in silicon carbide are weaker than in diamond.

Total: 8

3 **Table 5** shows some properties of five different substances.

Substance	Melting point in °C	Boiling point in °C	Electrical conductivity as a solid	Electrical conductivity as a liquid
A	2852	3600	poor	good
B	−101	−35	poor	poor
C	3652	n/a	good	n/a
D	962	2212	good	good

Table 5

p48 4.2.2.4

3–1 Which substance is most likely to be chlorine, Cl_2? Explain your answer. *[2]*

p45 4.2.2.8

3–2 Identify the particles responsible for carrying the charge when substance D conducts electricity. *[1]*

p48 4.2.2.3

3–3 Determine which substance is most likely to consist of oppositely charged ions in a lattice structure. Justify your answer in terms of structure and bonding. *[4]*

p54–5 4.2.3.2 WS1.2

3–4 Explain why substance C could be graphite. Answer in terms of structure and bonding. *[3]*

Total: 10

3 Quantitative chemistry

Chemical measurements, conservation of mass and the quantitative interpretation of chemical equations

Quick questions

p277 4.3.1.1 WS1.2

1 State the total number of atoms present in the formula of ammonium sulfate, $(NH_4)_2SO_4$.

p283–4 4.3.1.1 WS1.2

2 Which of the following symbol equations is correctly balanced?

- $C_3H_8(g) + 4O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$
- $2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$
- $CuCO_3(s) \rightarrow CuO(s) + 2CO_2(g)$
- $Mg(s) + AgNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + Ag(s)$

p65–6 4.3.1.2

3 Use the periodic table to calculate the relative formula mass (M_r) of magnesium nitrate, $Mg(NO_3)_2$.

p68–9 4.3.1.3

4 Give **one** example of when the law of conservation of mass may appear not to be obeyed during a chemical reaction.

p126–7 4.3.1.4 WS3.4

5 Calculate the mean value and its uncertainty for the following mass measurements: 2.95 g, 2.89 g, 2.65 g, 2.99 g.

Exam-style questions

6 Sodium hydrogencarbonate, $NaHCO_3$, thermally decomposes to form sodium carbonate, Na_2CO_3, carbon dioxide gas and water. The equation for the reaction is:

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

Table 1 shows the relative formula masses (M_r) of the products of this reaction.

Product	**Relative formula mass (M_r)**
Na_2CO_3	106
CO_2	44
H_2O	18

Table 1

p283–4 4.3.1.1 WS1.2

6–1 Explain why the symbol equation above is balanced. *[1]*

p65–6 4.3.1.2

6–2 Use the periodic table to calculate the relative formula mass of sodium hydrogencarbonate. *[1]*

p68 | 4.3.1.1 | WS1.2

6–3 Use **Table 1** and your answer to 6–2 to show that mass is conserved during this reaction. *[3]*

p68–9 | 4.3.1.3 | MS1a | AT1,2&6

6–4 A student investigated the thermal decomposition of sodium hydrogencarbonate. They followed this method:

- Place 2.40 g of the solid into a crucible with a mass of 25.00 g.
- Heat the crucible strongly for 5 minutes and then allow it to cool.
- Reweigh the crucible and the solid it contains.

The total mass when reweighed was 26.51 g.

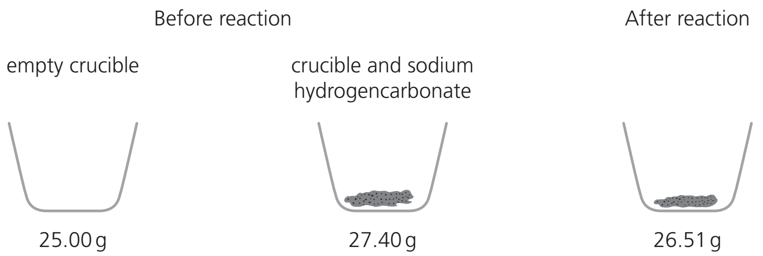


Figure 1

Calculate the mass of the products in the crucible. *[1]*

6–5 Suggest a reason why the mass of the products was less than the total mass of reactants used. *[2]*

Total: 8

7 Calcium metal and oxygen gas react to form calcium oxide, CaO.

p283–5 | 4.1.1.1

7–1 Give the balanced symbol equation for the reaction between calcium and oxygen. *[2]*

p126–7 | 4.3.1.4 | WS3.4 | MS1a,1c

7–2 **Table 2** shows the results from a series of experiments where 2.00 g of calcium was heated in a crucible with a mass of 28.00 g.

Mass of crucible and calcium in g	Mass of crucible and calcium oxide in g
30.00	30.81
30.00	30.79
30.00	30.41
30.00	30.84

Table 2

Calculate the mean mass of calcium oxide and its uncertainty produced in these four experiments. Give your answer to a suitable number of significant figures. *[5]*

p68–9 | 4.3.1.3 | AT1,2,6

7–3 Explain why these results do not appear to follow the law of conservation of mass. *[2]*

Total: 9

8 When lead nitrate, $Pb(NO_3)_2$, is heated it breaks down to form a mixture of products. The equation for the reaction is:

$2Pb(NO_3)_2(s) \rightarrow 2PbO(s) + 4NO_2(g) + O_2(g)$

Table 3 shows the relative formula masses of some of the substances in the reaction.

Substance	Relative formula mass (M_r)
$Pb(NO_3)_2$	331
PbO	223
O_2	32

Table 3

p68–9 4.3.1.3 AT1,2,6

8–1 When carried out as an experiment, the mass of the lead nitrate always seems to be greater than the mass of the three products. Explain why. *[2]*

p68 4.3.1.2 MS3b

8–2 Use **Table 3** and the balanced equation to show that the relative formula mass of nitrogen dioxide, NO_2, is equal to 46. Explain your answer. You must show your working. *[4]*

Total: 6

Use of amount of substance in relation to masses of pure substances

Quick questions

p85–6 4.3.2.5 WS4.3,4.5 MS1c,3c

1 Calculate the concentration of the solution formed when 0.34 g of potassium iodide is dissolved in 250 cm^3 of water. Give your answer in g/dm^3.

p65 4.3.2.1 WS4.1,4.2 H

2 Give the unit used to measure the amount of a chemical substance.

p65 4.3.2.1 WS4.1,4.2 MS1b H

3 Define the 'Avogadro constant'. What is its value?

p66 4.3.2.1 WS4.1,4.2 MS1a,1c

4 What is the mass of one mole of copper hydroxide, $Cu(OH)_2$?

p75–6 4.3.2.4 WS4.1 H

5 Describe what is meant by a 'limiting reactant'.

p70–1 4.3.2.2 WS4.2 MS1c,3c H

6 Methane, CH_4, reacts with steam, H_2O, to form carbon monoxide, CO, and hydrogen, H_2, as shown in the balanced symbol equation:

$CH_4 + H_2O \rightarrow CO + 3H_2$

If 1.50 moles of methane react, state the number of moles of hydrogen that will be formed.

Exam-style questions

7 At 20 °C, up to 70.1 g of sodium thiosulfate can be dissolved in 100 cm^3 of water.

p278–9 4.2.1.3 MS1c

7–1 Sodium thiosulfate is an ionic compound containing Na^+ and $S_2O_3^{2-}$ ions. Give the formula of sodium thiosulfate. *[1]*

p85–6 4.3.2.5 WS4.2 MS1c

7–2 A solution of sodium thiosulfate was prepared by dissolving 0.25 g of the solid into 50 cm^3 of water. Calculate the concentration of this solution. Give your answer in g/dm^3. *[2]*

p85–6 | 4.3.2.5 | WS4.2 | MS1c

H **7–3** Describe the effect on the concentration of the sodium thiosulfate solution if 100 cm^3 of water is used instead of 50 cm^3. Explain your answer. No calculation is necessary. *[2]*

p85–6 | 4.3.2.5 | WS4.2,4.6 | MS1a,1c, 2a, 2c

7–4 Another solution of sodium thiosulfate has a concentration of 1.85 g/dm^3. What mass of sodium thiosulfate is required to prepare 75.0 cm^3 of this solution? Give your answer to an appropriate number of significant figures. *[3]*

p85–6 | 4.3.2.5 | WS4.2,4.5 | MS1a,1c,3b

H **7–5** A student prepared a 2.75 g/dm^3 solution of sodium thiosulfate using only 0.40 g of the solid. Determine the volume of water the student dissolved the sodium thiosulfate into. Give your answer in cm^3. *[2]*

Total: 10

H **8** Propane, C_3H_8, is often used as the fuel for camping stoves.

p73–4 | 4.3.2.3 | WS4.2 | MS1a,1c,3c

8–1 During a combustion reaction 0.25 mol of propane reacted with 1.25 mol of oxygen, O_2, to form 0.75 mol of carbon dioxide, CO_2, and 1.00 mol of water, H_2O. Use this information to determine the balanced symbol equation for the reaction. *[2]*

p71–2 | 4.3.2.2 | WS4.2 | MS1a,1c, 3b,3c

8–2 In another reaction 3.52 g of propane were completely combusted. Use your answer to 8–1 to calculate the mass of carbon dioxide that was produced. *[3]*

Relative formula masses (M_r):
C_3H_8 = 44
O_2 = 32
CO_2 = 44
H_2O = 18

p75–7 | 4.3.2.4 | WS4.1,4.2 | MS1a,1c,3c

8–3 In a third reaction, 0.22 g of propane react with 0.40 g of oxygen. Show that the oxygen is the limiting reactant in this reaction. *[4]*

Total: 9

H **9** Aluminium is an important and useful metal. It is extracted from aluminium oxide, Al_2O_3, by electrolysis. The equation for this reaction is

$$2Al_2O_3 \rightarrow 4Al + 3O_2$$

p71–2 | 4.3.2.2 | WS4.2 | MS1a,1c, 3b,3c

9–1 In one extraction plant 400 kg of aluminium oxide is electrolysed each hour.

Relative formula masses (M_r): Al_2O_3 = 102 O_2 = 32

Relative atomic mass (A_r): Al = 27

Calculate the mass in kg of aluminium produced each hour by the plant. Give your answer to 1 decimal place. *[3]*

p71–2 | 4.3.2.2 | WS4.2 | MS1a,1c, 3b,3c

9–2 The plant runs continually. Calculate the mass of aluminium produced each week by the plant. Give your answer to three significant figures. *[2]*

p75–7 | 4.3.2.4 | WS4.1,4.2 | MS1a,1c,3c

9–3 Aluminium can be used as rocket fuel by reacting it with ammonium perchlorate, NH_4ClO_4, as follows:

$$10Al + 6NH_4ClO_4 \rightarrow 4Al_2O_3 + 2AlCl_3 + 12H_2O + 3N_2$$

Relative formula mass (M_r) NH_4ClO_4 = 117.5

In a small rocket engine 50.0 kg of aluminium is reacted with 150.0 kg of ammonium perchlorate. Explain why the aluminium is the limiting reactant. You must show your working. *[4]*

p73–4 | 4.3.2.3 | WS4.2 MS1a,1c,3c

9–4 Ammonium perchlorate thermally decomposes when heated.

2.35 g of ammonium perchlorate were heated and formed 0.73 g of hydrogen chloride, HCl, 0.28 g of nitrogen gas, N_2, 0.80 g of oxygen gas, O_2, and 0.54 g of water, H_2O.

Relative formula masses (M_r):

NH_4ClO_4 = 117.5
HCl = 36.5
N_2 = 28
O_2 = 32
H_2O = 18

Use this information to determine the balanced symbol equation for the thermal decomposition of ammonium perchlorate. You must show your working. *[4]*

Total: 13

Yield and atom economy of chemical reactions

Quick questions

p78 | 4.3.3.1 | WS4.1,4.2

1 Explain why reactions with a high percentage yield are preferred in industrial manufacturing processes.

p79 | 4.3.3.2 | WS4.1,4.2 MS1c

2 Write the equation used to determine the percentage atom economy for a desired product in a reaction.

p79 | 4.3.3.2 | WS4.2 MS1c

3 Explain why a reaction with only one product must have an atom economy of 100%.

Exam-style questions

4 Ammonia, NH_3, is manufactured using the Haber process. The reaction that occurs is

$N_2 + 3H_2 \rightarrow 2NH_3$

p78 | 4.3.3.1 | WS4.1,4.2, 4.6 MS1a,1c,2a

4–1 When 140 kg of nitrogen are reacted, the maximum theoretical yield of ammonia is 170 kg. One Haber process reactor produces an actual yield of ammonia of 31.7 kg.

Calculate the percentage yield for this reaction. Give your answer to three significant figures. *[2]*

p78 | 4.3.3.1 | WS4.1,4.2

4–2 Give **three** reasons why the percentage yield of this reaction is not 100%. *[3]*

p79–80 | 4.3.3.2 | WS4.1,4.2 MS1c

4–3 Explain why the manufacture of ammonia using the Haber process may be considered sustainable. Give your answer in terms of atom economy. *[2]*

Total: 7

5 Ethanol, CH_3CH_2OH, can be manufactured using two different processes.

Process 1 produces ethanol from the fermentation of glucose, $C_6H_{12}O_6$.

$C_6H_{12}O_6 \rightarrow 2CH_3CH_2OH + 2CO_2$

The glucose is derived from the plant sugar cane. The process has a relatively low energy requirement.

Process 2 produces ethanol by the hydration of ethene, C_2H_4.

$C_2H_4 + H_2O \rightarrow CH_3CH_2OH$

Ethene is a hydrocarbon produced from crude oil. Process 2 has a high energy requirement.

Relative formula masses (M_r):

$C_6H_{12}O_6 = 180$
$CH_3CH_2OH = 46$
$CO_2 = 44$
$H_2O = 18$

p79–80 | 4.3.3.2 | WS4.2,4.6 | MS1a,1c,2a

5–1 Calculate the percentage atom economy for the reaction to produce ethanol in Process 1. Give your answer to two significant figures. *[3]*

p78–9 | 4.3.3.1 | WS4.2,4.6 | MS1a,1c, 2a,3b

5–2 The overall percentage yield for Process 2 is 92.0%. A reactor has a maximum theoretical yield of 24.8 kg. Calculate the actual yield of ethanol for this reactor. Give your answer to three significant figures. *[3]*

p78–81 | 4.3.3.1-2 | WS4.1,4.2 | QWC

5–3 Compare the sustainability of the two processes for manufacturing ethanol. *[6]*

Total: 12

Ⓗ Using concentrations of solutions in mol/dm³

Quick questions

p86–7 | 4.3.4 | WS4.1,4.2, 4.5 | MS1a,1c

1 Calculate the concentration of the solution when 0.055 mol are dissolved in 25 cm³. Give your answer in mol/dm³.

p86–7 | 4.3.4 | WS4.1,4.2, 4.5 | MS1a,1c,3b

2 Calculate the number of moles of solute dissolved in 27.9 cm³ of a 0.780 mol/dm³ solution.

p86–7 | 4.3.4 | WS4.2,4.5, 4.6 | MS1a,1c,2a

3 What is the concentration of a 0.45 g/dm³ solution of copper sulfate in mol/dm³?

M_r $CuSO_4 = 159.5$

Exam-style questions

4 A student prepared a solution of potassium hydroxide, KOH, by dissolving 10.0 g of solid KOH into 250 cm³ of water.

p85 | 4.3.4 | WS4.2,4.5 | MS1a,1c

4–1 Calculate the concentration in g/dm³ of the potassium hydroxide solution prepared by the student. *[2]*

p85 | 4.3.4 | WS4.1 | MS1c

4–2 Describe how the concentration of this solution will change if the volume of water used is doubled to 500 cm³. Explain your answer. *[2]*

p88–9 | 4.3.4 | WS4.2,4.3,4.6 | MS1a,1c,2a,3b,3c

4–3 Another student titrated 25.0 cm^3 of dilute nitric acid, HNO_3, with a 0.110 mol/dm^3 solution of potassium hydroxide, KOH. 18.6 cm^3 of the potassium hydroxide solution reacted.

The equation for the reaction is:

$HNO_3 + KOH \rightarrow KNO_3 + H_2O$

Calculate the concentration of the dilute nitric acid. Give your answer in mol/dm^3 to three significant figures. *[4]*

p86–7 | 4.3.4 | WS4.2,4.5 | MS1a,1c,3b

4–4 Calculate the mass of potassium hydroxide dissolved in 18.6 cm^3 of the potassium hydroxide solution. *[2]*

Relative formula mass (M_r) KOH = 56

Total: 10

5 Oxalic acid is a naturally occurring acid found in many foods. A student carried out a titration in order to determine the concentration of oxalic acid in a solution. They titrated 25.0 cm^3 of oxalic acid solution with a 0.225 mol/dm^3 solution of sodium hydroxide.

The equation for the reaction is:

$C_2O_4H_2 + 2NaOH \rightarrow Na_2C_2O_4 + 2H_2O$

31.6 cm^3 of the sodium hydroxide solution reacted with the oxalic acid solution.

p88–9 | 4.3.4 | WS4.2,4.3,4.6 | MS1a,1c,2a,3c

5–1 Calculate the concentration of the oxalic acid in mol/dm^3. Give your answer to three significant figures. *[4]*

p86–7 | 4.3.4 | WS4.2,4.5 | MS1a,1c,3b

5–2 Use your answer to 5–1 to calculate the mass of oxalic acid dissolved in 25.0 cm^3 of the oxalic acid solution. *[2]*

Relative formula mass (M_r) $C_2O_4H_2$ = 90

Total: 6

H Use of amount of substance in relation to volumes of gases

Quick questions

p82 | 4.3.5 | WS1.2,4.2 | MS3b

1 Describe the relationship between the temperature of a gas and its volume.

p82 | 4.3.5 | WS4.2

2 Describe the relationship between the pressure of a gas and its volume.

p84 | 4.3.5 | WS4.2 | MS1a,1c,3c

3 Write down the volume that 1 mole of any gas occupies at room temperature and pressure.

p84 | 4.3.5

4 Pentane, C_5H_{12}, gas reacts with oxygen according to the equation

$C_5H_{12}(g) + 8O_2(g) \rightarrow 5CO_2(g) + 6H_2O(l)$

What volume of carbon dioxide will be produced if 0.25 dm^3 of pentane are reacted? Assume that both gases are at the same temperature and pressure.

Exam-style questions

5 The balanced equation for the thermal decomposition of copper carbonate, $CuCO_3$, is

$CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$

Figure 2 shows the apparatus a student used to investigate the thermal decomposition reaction.

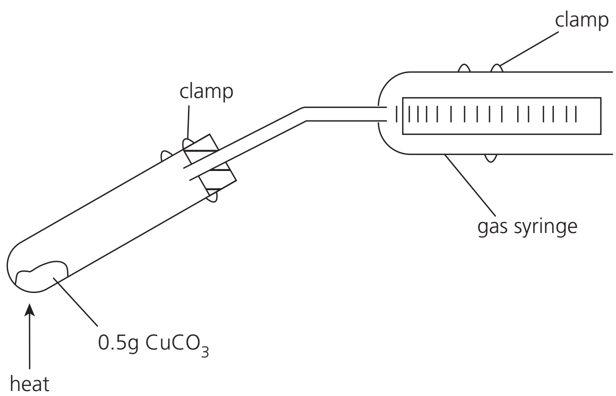


Figure 2

0.50 g of copper carbonate was heated until the thermal decomposition reaction began. The carbon dioxide gas was collected in a gas syringe and the volume measured. The student repeated the experiment three times.

Relative formula mass (M_r) $CuCO_3$ = 123.5

p70–2& 82–3 | 4.3.2.1, 4.3.2.2, 4.3.5 | WS4.1, 4.2,4.3,4.6 | MS1a,1c, 2a,3c

5–1 Calculate the maximum theoretical yield, in dm^3, of carbon dioxide gas that can be produced from 0.50 g of copper carbonate. Give your answer to two significant figures. (Note: The volume of one mole of any gas at room temperature and pressure is 24.0 dm^3.) *[4]*

Table 4 shows the student's results.

Experiment	Volume of CO_2 collected in cm^3
1	77.5
2	80.0
3	79.5

Table 4

p126–7 | 4.3.1.4 | WS3.4

5–2 Calculate the mean volume of CO_2. Give the uncertainty of the mean value. *[2]*

p78 | 4.3.3.1 | WS4.1,4.2, 4.5 | MS1a,1c

5–3 Use your answers to 5–1 and 5–2 to calculate the percentage yield for the reaction. *[2]*

p78 | 4.3.3.1 | WS4.1,4.2

5–4 Suggest a reason why the percentage yield of the reaction was not 100%. *[1]*

Total: 9

Quantitative chemistry topic review

1 Sulfur dioxide is an important chemical used in the manufacture of sulfuric acid by a process called the Contact process. Industrially, there are two methods for producing sulfur dioxide.

Method 1: burning sulfur in air

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

Method 2: roasting iron sulfide in air

$4FeS_2(s) + 11O_2(g) \rightarrow 2Fe_2O_3(s) + 8SO_2(g)$

p79 | 4.3.3.2 | WS4.1,4.2 | MS1c

1–1 Explain why the reaction for method 1 has a percentage atom economy of 100%. *[1]*

p79–80 | 4.3.3.2 | WS4.2 | MS1a,1c

1–2 Calculate the percentage atom economy for producing sulfur dioxide by method 2. *[3]*

Relative formula masses (M_r):
FeS_2 = 120
O_2 = 32
Fe_2O_3 = 160
SO_2 = 64

p71–2 | 4.3.2.2 | WS4.1,4.2, 4.6 | MS1a,c,2a, 3c

H 1–3 100 kg of iron sulfide, FeS_2, were roasted in an excess of air. Calculate the maximum theoretical yield of sulfur dioxide that this mass of FeS_2 can produce. Give your answer in kilograms to three significant figures. *[4]*

Relative formula masses (M_r): FeS_2 = 120, SO_2 = 64

p78 | 4.3.3.1 | WS4.1,4.2 | MS1a,1c

1–4 The actual yield of sulfur dioxide in method 2 is 95.0 kg. Calculate the percentage yield for this reaction. *[2]*

Total: 10

2 A student carried out a titration in order to determine the concentration of sulfuric acid in a solution. They titrated 25.0 cm^3 portions of the sulfuric acid solution with a 0.850 mol/dm^3 solution of potassium hydroxide.

The equation for the reaction is:

$H_2SO_4 + 2KOH \rightarrow K_2SO_4 + 2H_2O$

The student's results are shown in **Table 5**.

	Titration 1	Titration 2	Titration 3
Volume of potassium hydroxide solution added in cm^3	23.55	23.40	23.45

Table 5

p126–7 | 4.3.1.4 | WS3.4 | MS1a,1c,2b

2–1 Calculate the mean value for the volume of potassium hydroxide solution added. Give your answer to two decimal places. Write down the uncertainty in these three titration volumes. *[2]*

p88–9 | 4.3.4 | WS4.2,4.3, 4.6 | MS1a,1c, 2a,3c

H 2–2 Use your answer to 2–1 and the information in **Table 5** to calculate the concentration of the sulfuric acid solution in mol/dm^3. Give your answer to three significant figures. *[4]*

p86–7 | 4.3.4 | WS4.2,4.5 | MS1a,1c | H **2–3** Use your answer to 2–2 to calculate the mass of sulfuric acid dissolved in 25.0 cm^3 of the sulfuric acid solution. *[2]*

Relative formula mass (M_r) H_2SO_4 = 98

Total: 8

3 Urea has the formula CH_4N_2O, and is an important chemical used in agriculture, the chemical industry and in the manufacture of explosives.

Urea can be manufactured by reacting ammonia with carbon dioxide.

$2NH_3(g) + CO_2(g) \rightarrow CH_4N_2O(g) + H_2O(g)$

p84 | 4.3.5 | WS4.1,4.2 | MS1a,1c,3c | H **3–1** Give the number of moles of urea that will form when 0.46 mol of ammonia gas are reacted with an excess of carbon dioxide gas. *[2]*

p75–7 | 4.3.2.4 | WS4.1 | MS1a,1c | H **3–2** In one reaction, 5.00 g of ammonia reacts with 5.00 g of carbon dioxide. Show that the carbon dioxide is the limiting reactant in this reaction. You must show your working. *[4]*

Relative formula masses (M_r): NH_3 = 17, CO_2 = 44

p85–6 | 4.3.2.5 | WS4.5,4.6 | MS1c,2a,3b | **3–3** Urea is highly soluble in water. A solution of urea was prepared by dissolving 0.75 g of urea into 400 cm^3 of water. Calculate the concentration of this urea solution in g/dm^3. Give your answer to three significant figures. *[2]*

p86–7 | 4.3.4 | WS4.2,4.3, 4.5 | MS1a,1c | H **3–4** Convert the concentration of urea in 3–3 from g/dm^3 into mol/dm^3. *[1]*

Relative formula mass (M_r) CH_4N_2O = 70

Total: 9

4 When heated, ammonium dichromate, $(NH_4)_2Cr_2O_7$, decomposes to form chromium oxide, Cr_2O_3, nitrogen gas, N_2, and water.

A student heated 7.56 g of ammonium dichromate until it had completely decomposed. 4.56 g of chromium oxide, 0.84 g of nitrogen gas, and 2.16 g of water were formed.

p65–6 | 4.3.1.2 | **4–1** Calculate the relative formula mass of ammonium dichromate, $(NH_4)_2Cr_2O_7$. *[1]*

Relative atomic masses (A_r): N = 14, H = 1, Cr = 52, O = 16

p73–4 | 4.3.2.3 | MSb,3c | H **4–2** Use your answer to 4–1, and the information above, to determine the balanced equation for this reaction. You must show your working. *[4]*

Relative formula masses (M_r): Cr_2O_3 = 152, N_2 = 28, H_2O = 18

Total: 5

4 Chemical changes

Reactivity of metals

Quick questions

p278–9 | 4.4.1, 4.2.1.3 | MS1c

1 Determine the chemical formulae of the following two metal oxides from the charges of their ions.

- copper(I) oxide: Cu^{+} and O^{2-}
- barium oxide: Ba^{2+} and O^{2-}

p99 | 4.4.1.1

2 In the following reaction, identify the substance being oxidised. Explain your answer.

$4V + 5O_2 \rightarrow V_2O_5$

p101 | 4.4.1.2

3 Explain why the following reaction is a displacement reaction.

$Al + 3AgNO_3 \rightarrow Al(NO_3)_3 + 3Ag$

p106 | 4.4.1.3

4 The following reaction can be used to extract lead from lead oxide. Identify the substance being reduced. Explain your answer.

$2PbO + C \rightarrow 2Pb + CO_2$

p103 | 4.4.1.4

(H) 5 Define oxidation and reduction in terms of electrons.

Exam-style questions

6 Iron is extracted from iron(III) oxide, Fe_2O_3, in an industrial process which takes place inside a large reactor called a blast furnace. Iron(III) oxide is mixed with carbon and air at very high temperatures. The equation for one of the extraction reactions is

$2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$

p99 | 4.4.1.1, 4.4.1.3

6–1 Identify the substance being oxidised during this reaction. Explain your answer. *[1]*

p106 | 4.4.1.3

6–2 Explain why carbon can be used to extract iron from iron(III) oxide. *[2]*

p101 | 4.4.1.2

6–3 During this reaction Fe^{3+} ions in the Fe_2O_3 are reduced to form Fe atoms. Explain why this is a reduction reaction. *[1]*

p104& 288–9 | 4.4.1.4, 4.1.1.1

(H) 6–4 Give the half equation for this reduction reaction. *[1]*

p39–40 | 4.2.2.4 | WS1.2

6–5 The carbon dioxide formed is a gas at room temperature. Explain why. Answer in terms of structure and bonding. *[3]*

Total: 8

7 A student investigated the order of reactivity of a group of metals. They reacted each metal with dilute hydrochloric acid and measured the temperature change that took place. **Table 1** shows the results.

Metal	**Temperature increase in °C**
Iron (Fe)	10
Calcium (Ca)	32
Magnesium (Mg)	25
Copper (Cu)	0
Zinc (Zn)	14

Table 1

p100 4.4.1.1

7–1 Give the word equation for the reaction between calcium and hydrochloric acid. *[1]*

p101 4.4.1.1 WS1.2

7–2 Look at the following equations that show the behaviour of aluminium with zinc nitrate and magnesium nitrate.

$2Al + 3Zn(NO_3)_2 \rightarrow 2Al(NO_3)_3 + 3Zn$

$Al + Mg(NO_3)$ no reaction

Suggest the temperature increase for the reaction between aluminium and dilute hydrochloric acid. Assume the same conditions as for the student's investigation. *[3]*

p101 4.4.1.1

7–3 During the reaction between a metal and an acid, the metal atoms form positively charged ions. Identify the metal from the student's investigation that has the greatest tendency to form positive ions. Explain your answer. *[2]*

p104& 288-9 4.4.1.4, 4.1.1.

H 7–4 Explain why the formation of positive metal ions is an oxidation reaction. *[2]*

7–5 Give the half equation for the oxidation reaction of magnesium, Mg. *[1]*

Total: 9

8 Calcium is a reactive metal in Group 2 of the periodic table. It reacts readily with water.

p100 4.4.1.2

8–1 Give the word equation for the reaction between calcium and water. *[1]*

p104& 288-9 4.4.1.4

H 8–2 Explain why the reaction between calcium and water is an example of a displacement reaction. *[2]*

8–3 Calcium oxide reacts with sodium in another displacement reaction. The equation for the reaction is

$2Na + CaO \rightarrow Na_2O + Ca$

Explain why calcium is being reduced during this reaction whilst sodium is being oxidised. Give your answer in terms of electron transfer. You may include half equations in your answer. *[3]*

Total: 6

Reactions of acids

Quick questions

p111–13 4.4.2.1, 4.4.2.2

1 Give the names of the products formed in the following two chemical reactions.

- zinc metal added to sulfuric acid
- iron hydroxide added to nitric acid.

p115 4.4.2.3 RP1 AT4

2 Name the separation process used to:

- separate excess solid metal oxide from a solution of a soluble salt
- form crystals of a soluble salt from a solution of that salt.

p107–8 4.4.2.4

3 Describe the differences between an acid and an alkali.

p107–8& 111–12 4.4.2.4

4 Describe how the colour of universal indicator added to hydrochloric acid will change as sodium hydroxide is added until it is in excess.

p109–10 4.4.2.6

H 5 Describe the difference between:

- a strong acid and a weak acid
- a dilute acid and a concentrated acid.

Exam-style questions

6 Zinc chloride is a soluble salt that can be formed by reacting solid zinc carbonate with an acid.

p278–9 4.4.2.2

6–1 Zinc chloride contains the ions Zn^{2+} and Cl^{-}. Give the formula of zinc chloride. *[1]*

p113 4.1.1.1, 4.4.2.2

6–2 Give the word equation for the reaction of zinc carbonate to form zinc chloride. *[2]*

p107 4.4.2.4

6–3 Name the particle in the acid that makes it acidic. *[1]*

p115 4.4.2.3 RP1 AT2,3,4,6 QWC

6–4 Describe a method for producing pure, dry crystals of zinc chloride from solid zinc carbonate and your chosen acid. *[6]*

Total: 10

7 When magnesium reacts with hydrochloric acid the following reaction occurs.

$Mg + 2HCl \rightarrow MgCl_2 + H_2$

This reaction is a redox reaction involving both oxidation and reduction.

p111 4.4.2.1

H 7–1 Explain why magnesium is oxidised during this reaction. *[2]*

p71–2 4.3.2.2 WS4.2, 4.6 MS1a,1c, 2a,3b,3c

H 7–2 Explain why hydrogen is reduced during this reaction. *[2]*

1.25 g of magnesium were reacted with an excess of hydrochloric acid.

Relative atomic mass (A_r) Mg = 24

Relative formula mass (M_r) H_2 = 2

p82–3 4.3.5 WS4.5,4.6 MS1a,1c,2a

H 7–3 Calculate the mass of hydrogen gas formed. Give your answer to three significant figures. *[4]*

7–4 Use your answer to 7–3 to calculate the volume of hydrogen gas produced at room temperature and pressure, when 1.25 g of magnesium react.

The volume that one mole of any gas at room temperature and pressure is 24.0 dm^3. *[2]*

Total: 10

8 A student decided to monitor the changes in pH of the neutralisation reaction between nitric acid, HNO_3, and sodium hydroxide, NaOH, using universal indicator solution.

p111–12 4.4.2.2

8–1 Name the products of the reaction between nitric acid and sodium hydroxide. *[1]*

p111–12 4.4.2.4

8–2 Give the name of the particle that gives a solution a pH less than 7. *[1]*

p111–12 4.4.2.4

8–3 Give the name of the particle that gives a solution a pH greater than 7. *[1]*

p107–8 4.4.2.6 MS2h H

8–4 Give the ionic equation for the neutralisation reaction between nitric acid and sodium hydroxide. State symbols are not required. *[1]*

8–5 During the neutralisation reaction the student observed a change in colour of the universal indicator from red, to orange, to green and finally to purple.

Explain these colour changes. Give your answer in terms of pH and the concentrations of hydrogen and hydroxide ions in the reaction mixture. *[5]*

Total: 9

9 A student was asked to carry out a titration experiment to determine the concentration of some dilute potassium hydroxide using a solution of sulfuric acid with known concentration, and some phenolphthalein indicator.

The reaction between potassium hydroxide and sulfuric acid is

$2KOH(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$

Phenolphthalein indicator is pink in alkaline solutions and colourless in acidic solutions.

The student used 25.0 cm^3 of a 0.135 mol/dm^3 solution of sulfuric acid in their titration experiment. **Table 2** shows the results.

	Titration 1	Titration 2	Titration 3	Titration 4	Titration 5
Volume of potassium hydroxide added in cm^3	22.45	21.75	21.80	21.50	21.75

Table 2

p88 4.4.2.5 RP2 AT1,8 QWC

9–1 Describe a method for how the student could carry out a titration to determine the concentration of the potassium hydroxide solution. Do **not** include the calculations needed to determine the concentration. *[6]*

p88–9 4.4.2.5, 4.3.4 WS4.2, 4.3,4.6 MS1a,1c, 2a,3c

9–2 Calculate the concentration of the potassium hydroxide solution in mol/dm^3. Give your answer to three significant figures. Use only the student's concordant results. Concordant results are those within 0.10 cm^3 of each other. *[5]*

Total: 11

p108–10 4.4.2.6

H **10** A student was given unlabelled samples of two acids labelled acid **A** and acid **B**.

The student was told that one acid is hydrochloric acid, HCl, and the other is ethanoic acid, CH_3COOH, and that the concentration of both acids was 0.10 mol/dm^3.

The student carried out a series of experiments to determine which acid was a strong acid and which was a weak acid. **Table 3** shows the results.

	Colour with universal indicator	**pH**	**Observations on adding calcium carbonate to the acid**
Acid A	orange	3.0	fizzes very slowly
Acid B	red	1.0	fizzes rapidly

Table 3

10–1 Use **Table 3** to determine which acid, **A** or **B**, is hydrochloric acid. *[1]*

10–2 Determine how many times more concentrated the hydrogen ions are in acid **B** compared to acid **A**. *[1]*

10–3 Explain why both acids have the same concentration but different pH values. *[3]*

Total: 5

Electrolysis

Quick questions

p117–20 4.4.3.2, 4.4.3.4

H **1** Name the products formed when the following substances are electrolysed using inert electrodes:

- molten copper bromide, $CuBr_2$
- molten zinc oxide, ZnO
- aqueous sodium chloride, NaCl
- aqueous potassium nitrate, KNO_3.

p118–19 4.4.3.3

H **2** Give **one** reason why a metal would be extracted from its ore using electrolysis rather than heating with carbon.

p107–8& 289–91 4.4.3.5

H **3** Give the half equation for the oxidation of chloride ions to form chlorine gas.

p119–20 4.4.3.4

4 When a solution of sodium sulfate is electrolysed, hydrogen gas is formed at the cathode and oxygen gas is formed at the anode. Which two ions are discharged to form these two gases?

p289–91 4.4.3.5

H **5** Balance the following half equations. Write down whether each one shows oxidation or reduction.

$Cl^- \rightarrow Cl_2 + e^-$

$H^+ + e^- \rightarrow H_2$

$OH^- \rightarrow H_2O + O_2 + e^-$

Exam-style questions

6 This question is about the extraction of aluminium from its ore, bauxite. Bauxite contains aluminium oxide. During the electrolysis process carbon electrodes are used to pass electricity through the molten aluminium oxide, as shown in **Figure 1**.

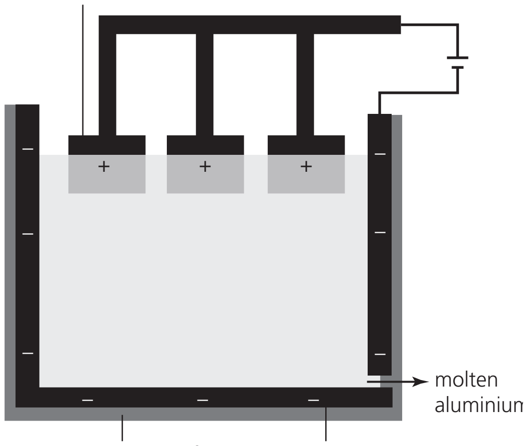


Figure 1

p34–5 4.2.2.3 **6–1** Explain why the aluminium oxide must be molten for electrolysis to occur. *[2]*

p119 4.4.3.3 **6–2** Explain why aluminium oxide is mixed with cryolite before electrolysis takes place. *[2]*

p119 4.4.3.1 QWC **6–3** Describe the changes that take place during the electrolysis of aluminium oxide to form aluminium and oxygen gas. *[6]*

p119 4.4.3.3 **6–4** Give the reason why the positive carbon electrodes need to be replaced regularly. *[1]*

Total: 11

7 A student investigated the electrolysis of aqueous copper sulfate. **Figure 2** shows the apparatus the student used to carry out the electrolysis reaction. **Table 4** shows the student's observations.

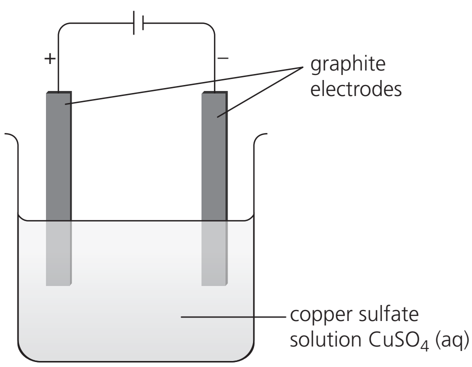


Figure 2

	Positive electrode (anode)	Negative electrode (cathode)
Aqueous $CuSO_4$	Bubbles of a colourless gas formed at the electrode	A red-brown solid formed on the electrode

Table 4

p34–5 4.2.2.3

7–1 Explain why the student did not need to heat the copper sulfate before it would conduct electricity. *[1]*

p54–5 4.4.3.2

7–2 Explain why graphite electrodes are able to conduct electricity. *[1]*

p119–20 4.4.3.4

7–3 The red-brown solid that formed at the negative electrode was copper metal. Explain why copper metal formed at this electrode. *[2]*

p119–20 4.4.3.4 WS1.2 RP3 AT3,7,8

7–4 Identify the gas formed at the positive electrode. *[1]*

7–5 Name the ions that are discharged at the positive electrode to form this gas. *[1]*

7–6 The student then decided to electrolyse a solution of sodium sulfate, $Na_2SO_4(aq)$. Describe the difference in the observation at the negative electrode compared with the electrolysis of copper sulfate solution. Explain your answer. *[4]*

Total: 10

8 **Figure 3** shows the apparatus used for the electrolysis of dilute sulfuric acid.

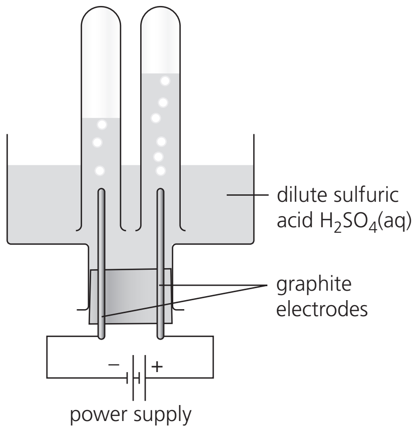


Figure 3

p119–20 4.4.3.4

8–1 Name the ions that are present in a solution of sulfuric acid. *[1]*

p120 4.4.3.5

H **8–2** The product at the negative electrode is a colourless gas.

Identify the colourless gas formed at the negative electrode. *[1]*

p119–20 4.4.3.4

8–3 Give the half equation to show the formation of this gas. *[2]*

p120 4.4.3.5

H **8–4** The product at the positive electrode is also a colourless gas.

Identify the colourless gas formed at the negative electrode. *[1]*

8–5 Give the half equation to show the formation of this gas. *[2]*

p82–3 4.3.5 WS4.5,4.6 MS1a,1c,2a,3b

8–6 4 cm^3 of gas was collected at the negative electrode at room temperature and pressure. Calculate the number of moles of this gas that were produced. Give your answer to three significant figures. *[3]*

The volume of one mole of any gas at room temperature and pressure is 24.0 dm^3.

Total: 10

Chemical changes topic review

1 Copper oxide can be formed by heating copper metal in air. The reaction is very slow compared with the same reaction using magnesium metal.

p99 4.4.1.1

1–1 Explain why the reaction between copper and oxygen is an oxidation reaction. *[1]*

p278–9 4.4.1.1, 4.2.1.3 MS1c

1–2 Copper oxide contains Cu^{2+} and O^{2-} ions. Give the formula of copper oxide. *[1]*

p283–5 4.4.1.1, 4.1.1.1

1–3 Give the balanced symbol equation for the reaction between copper and oxygen to form copper oxide. *[2]*

p101 4.4.1.2

1–4 Describe the reaction that will take place when magnesium powder is heated with a sample of copper oxide. Explain your answer. Include a word equation. *[3]*

p115 4.4.2.3 RP1 AT2,3,4,6 QWC

1–5 When solid copper oxide is warmed with dilute sulfuric acid, H_2SO_4, copper sulfate, $CuSO_4$, and water are produced. This reaction can be used to produce pure, dry crystals of copper sulfate.

Describe a method for producing pure, dry crystals of copper sulfate from solid copper oxide and dilute sulfuric acid. *[6]*

Total: 13

2 The Group 1 metal sodium reacts quickly with cold water. The reaction produces a colourless gas and an alkaline solution.

p107 4.4.2.4

2–1 Name the ion that causes the alkaline solution in the reaction described above. *[1]*

p100 4.1.1.1, 4.4.1.2

2–2 Give the word equation for the reaction between sodium and water. *[1]*

p103 4.4.1.4 H

2–3 During this reaction sodium atoms form sodium ions. Explain why this is an oxidation reaction. *[2]*

p107 4.4.2.4

2–4 A student reacted a small piece of lithium metal with water to produce a solution of lithium hydroxide, LiOH. They then added 10 cm^3 of this solution to some dilute nitric acid, HNO_3.

Suggest the pH values for the lithium hydroxide solution and the nitric acid solution. *[2]*

p111–12 4.4.2.2

2–5 Name the salt formed when lithium hydroxide is added to nitric acid. *[1]*

p112 4.4.2.4

2–6 Give the ionic equation for the reaction between lithium hydroxide and nitric acid. *[1]*

Total: 8

3 Hydrochloric acid, HCl, is a strong acid. Dilute hydrochloric acid is commonly used in school laboratories to carry out simple chemical reactions. Concentrated hydrochloric acid is used in many industrial processes.

p109–10 4.4.2.6 H

3–1 Explain the difference between the strength of an acid and how dilute or concentrated it is. *[2]*

p108 | 4.4.2.6 | MS2h | H

3–2 One sample of dilute hydrochloric acid has a pH of 2 while a sample of concentrated hydrochloric acid has a pH of –1. Determine how many times more concentrated the hydrogen ions are in the concentrated acid. *[1]*

p88–9 | 4.4.2.5, 4.3.4 | WS4.2, 4.3,4.6 | MS1a,1c, 2a,3c

3–3 A student was asked to carry out a titration experiment to determine the concentration of some dilute hydrochloric acid. They used 25.0 cm^3 portions of sodium hydroxide solution with a concentration of 0.200 mol/dm^3 and some phenolphthalein indicator. The equation for the reaction between sodium hydroxide and hydrochloric acid is

$KOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

The average volume of hydrochloric acid that was used in the titration experiment was 19.50 cm^3.

Calculate the concentration of the potassium hydroxide solution. Give your answer in mol/dm^3 to three significant figures. *[4]*

Total: 7

4 This is a question about magnesium and one of its compounds, magnesium nitrate. Magnesium reacts with acids to form a salt and hydrogen gas. Magnesium nitrate is a highly soluble salt.

p111 | 4.4.2.1

4–1 Name the acid required to react with magnesium to form magnesium nitrate. *[1]*

p278–9 | 4.4.1.1, 4.2.1.3 | MS1c

4–2 Magnesium nitrate contains the ions Mg^{2+} and NO_3^-. Give the formula for magnesium nitrate. *[1]*

p111 | 4.4.1.4 | QWC | H

4–3 During the reaction between magnesium and any acid both oxidation and reduction occur. The reaction is

$Mg + 2H^+ \rightarrow Mg^{2+} + H_2$

Explain why this reaction involves both oxidation and reduction reactions. Identify the substances being oxidised and reduced. Include half equations in your answer. *[6]*

p121 | 4.4.3.4 | WS1.2 | RP3 | AT3,7,8

4–4 A solution of magnesium nitrate was electrolysed using inert graphite electrodes. Predict the products of the electrolysis reaction at the positive electrode (anode). *[1]*

p121 | 4.4.3.4 | WS1.2 | RP3 | AT3,7,8

4–5 Predict the products of the electrolysis reaction at the negative electrode (cathode). *[1]*

Total: 10

5 Energy changes

Exothermic and endothermic reactions

Quick questions

p129–30 4.5.1.1

1 The reaction between zinc and copper sulfate solution causes the temperature of the solution to rise. Is the reaction exothermic or endothermic?

p130 4.5.1.1

2 Give **one** example of an application of an endothermic reaction.

p132 4.5.1.1 RP4 AT1,3,5,6

3 Give **one** variable that would affect the temperature change when hydrochloric acid reacts with magnesium metal.

p131 4.5.1.2

4 Define the term 'activation energy'.

p133–35 4.5.1.3 MS1a H

5 State which quantity is greater for an endothermic reaction: the amount of energy absorbed during bond breaking, or the amount of energy released during bond making.

Exam-style questions

6 When ammonium nitrate, NH_4NO_3, is dissolved in water the temperature of the solution formed decreases.

$NH_4NO_3(s) + water \rightarrow NH_4NO_3(aq)$

A student investigated how the mass of ammonium nitrate dissolved affected the temperature change of the solution.

p130–1 4.5.1.1

6–1 Write down whether the dissolving of ammonium nitrate is exothermic or endothermic. Explain your answer. *[2]*

p131 4.5.1.2

6–2 Draw a fully labelled reaction profile for the dissolving of ammonium nitrate to form a solution. *[3]*

p132 4.5.1.1 RP4 AT1,3,5,6

6–3 Describe a method the student could use to investigate how the mass of ammonium nitrate dissolved affects the change in temperature of the water. Your method should give valid results. *[6]*

Total: 11

7 Alcohols are compounds made from carbon, hydrogen and oxygen atoms. The complete combustion of alcohols forms carbon dioxide and water. During the reaction, chemical bonds are broken and new chemical bonds form.

The equation for the combustion of the alcohol ethanol, CH_3CH_2OH, is

$CH_3CH_2OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$

p39 4.2.1.1

7–1 Give the type of chemical bonds that occur in alcohol molecules. Explain your answer. *[2]*

p131 4.5.1.2

7–2 All combustion reactions are exothermic. Draw a fully labelled reaction profile for the combustion of ethanol to form carbon dioxide and water. *[3]*

p133–5 4.5.1.3 MS1a H

7–3 **Figure 1** shows the displayed formulae for the complete combustion of ethanol.

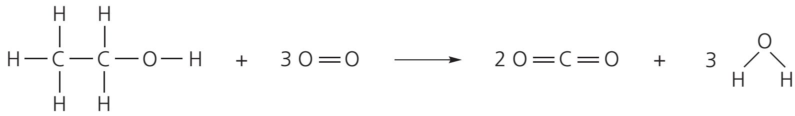


Figure 1

Table 1 shows the bond energies involved in the reaction.

	C–C	C–H	C–O	O–H	O=O	C=O
Energy in kJ/mol	348	412	360	463	496	743

Table 1

Calculate the overall energy change for this reaction. Use **Figure 1** and **Table 1**. *[4]*

p133–5 4.5.1.3 H

7–4 Explain why this reaction is exothermic overall. Give your answer in terms of bond breaking and bond making. *[2]*

Total: 11

8 Hydrogen peroxide, H_2O_2, is a very useful compound. It is very reactive due to the single covalent bond between the two oxygen atoms. Hydrogen peroxide decomposes to form water and oxygen.

Figure 2 shows the displayed formulae for the decomposition of hydrogen peroxide.

Figure 2

p133–5 4.5.1.3 MS1a H

8–1 **Table 2** shows the bond energies and the overall energy change for the decomposition of hydrogen peroxide. Calculate the bond energy **X** for the O**–**O bond in hydrogen peroxide. Use **Figure 2** and **Table 2**. *[4]*

	H–O	O–O	O=O	Overall energy change
Energy in kJ/mol	463	X	496	–208

Table 2

p86–7 4.3.4 MS1a,1c, 3b,3c WS4.2,4.3, 4.5

8–2 A solution of hydrogen peroxide has a concentration of 0.125 mol/dm^3. Use **Figure 2** to calculate the number of moles of oxygen produced when 50 cm^3 of this solution fully decomposes. *[2]*

p82–3 4.3.5 MS1a,1c, 2a,3b,3c WS1.2,4.2, 4.3,4.6

8–3 Calculate the volume, in cm^3, of this amount of oxygen gas. Give your answer to three significant figures. *[3]*

The volume of 1 mole of any gas at room temperature and pressure = 24.0 dm^3.

Total: 9

Chemical cells and fuel cells

Quick questions

p136 4.5.2.1 **1** Give **two** variables that affect the voltage produced by a chemical cell.

p137 4.5.2.1 **2** Explain the difference between a cell and a battery.

p137 4.5.2.1 **3** Explain why rechargeable cells and batteries can be recharged.

p138 4.5.2.2 **4** Describe the difference between a chemical cell and a fuel cell.

p139 4.5.2.2 **5** Give **one** advantage and **one** disadvantage of a hydrogen fuel cell over a rechargeable battery.

Exam-style questions

p136 4.5.2.1 **6** A student wanted to investigate how the type of electrodes used in a simple chemical cell affected the voltage produced. They kept a copper electrode attached to the positive terminal of the voltmeter and connected five different metals to the negative terminal. **Figure 3** shows the apparatus used.

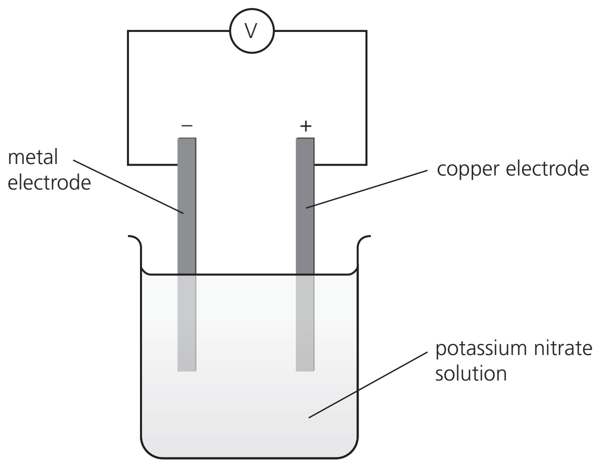


Figure 3

6–1 Describe the purpose of the potassium nitrate solution in the chemical cell. *[1]*

6–2 Identify the type of particles responsible for the flow of electricity through the potassium nitrate solution and through the electrical wires. *[2]*

6–3 Give **one** variable that the student must keep the same to ensure that their results are valid. *[1]*

Table 3 shows the student's results.

Metal electrode	Voltage measure in volts
Nickel	0.59
Vanadium	1.52
Chromium	1.08
Lead	0.47
Aluminium	2.00

Table 3

6–4 Predict the voltage produced if the student were to use two copper electrodes in their chemical cell. Explain your answer. *[2]*

6–5 Which of the five metals is the most reactive? Give a reason for your answer. *[2]*

6–6 An electrical device requires a voltage as close to 1.00 V as possible. Suggest **two** metals that would be the most suitable for the chemical cell for this device. Predict the voltage produced by this cell. *[2]*

Total: 10

7 Hydrogen gas is becoming an increasingly important fuel. Fuel cells react hydrogen with oxygen to produce electrical energy. Hydrogen gas can be produced by the electrolysis of water.

p103 4.4.1.4 7–1 At the anode of a hydrogen fuel cell, hydrogen gas is converted into hydrogen ions and electrons.

$2H_2 \rightarrow 4H^+ + 4e^-$

Explain why this is an oxidation reaction. *[1]*

p139 4.5.2.2 7–2 At the cathode of a hydrogen fuel cell, oxygen gas reacts with the hydrogen ions and electrons. Give the half equation to show this change. *[2]*

p133–5 4.5.1.3 MS1a H 7–3 Another source of hydrogen gas is from the reaction between natural gas, CH_4, and steam. Carbon monoxide, CO, is also produced.

Figure 4 shows the displayed formulae for this reaction.

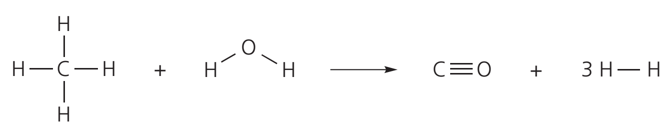


Figure 4

Table 4 shows the bond energies and the overall energy change for the reaction. Calculate the bond energy **X** for the C≡O bond in carbon monoxide. Use **Figure 4** and **Table 4**. *[4]*

	C–H	H–O	C≡O	H–H	Overall energy change
Energy in kJ/mol	412	463	X	436	+189

Table 4

p138–9 4.5.2.2 QWC 7–4 'Cars powered by hydrogen fuel cells are sustainable vehicles with zero carbon dioxide emissions.' Use the information in this question and your own knowledge to evaluate this statement. *[6]*

Total: 13

Energy changes topic review

1 Neutralisation reactions take place between acids and alkalis. A student investigated the temperature change during the reaction between sodium hydroxide and hydrochloric acid. They measured 20.0 cm^3 of sodium hydroxide solution and placed it in a polystyrene cup. They recorded the temperature of the solution and then added hydrochloric acid to it. The student's results are shown in **Table 5**.

Volume of acid added in cm^3	0.0	5.0	10.0	15.0	20.0	25.0	30.0	35.0	40.0
Temperature in °C	19.2	21.4	23.9	26.0	27.8	28.3	26.8	24.8	22.9

Table 5

p111–12 4.4.2.2 **1–1** Give the word equation for the reaction between sodium hydroxide and hydrochloric acid. *[1]*

p129–30 4.5.1.1 **1–2** Use **Table 5** to determine whether the reaction between sodium hydroxide and hydrochloric acid is exothermic or endothermic. Explain your answer. *[3]*

p132 4.5.1.1 RP4 AT1,3,5,6 **1–3** Use **Table 5** to suggest what volume of hydrochloric acid was required to neutralise the sodium hydroxide solution. Explain your answer. *[2]*

p132 4.5.1.1 RP4 AT1,3,5,6 **1–4** Another student wanted to investigate how the concentration of the hydrochloric acid solution affected the temperature change during the reaction between sodium hydroxide and hydrochloric acid.

Describe a method that the student could use to carry out their investigation. Your method should give valid results. *[6]*

Total: 12

2 One method of producing water is to react hydrogen gas with oxygen gas. **Figure 5** shows the displayed formula for this reaction. **Table 6** shows the bond energies involved in the reaction.

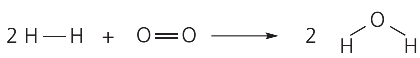


Figure 5

	H–H	O=O	H–O
Energy in kJ/mol	436	496	463

Table 6

p133–5 4.5.1.3 MS1a H **2–1** Calculate the overall energy change for this reaction. Use **Figure 5** and **Table 6**. *[4]*

p131 4.5.1.2 **2–2** Draw a fully labelled reaction profile for the reaction between hydrogen and oxygen to form water. *[3]*

p138–9 4.5.2.2 **2–3** Hydrogen fuel cells produce water by reacting hydrogen with oxygen. The fuel cell uses the chemical reaction to produce a voltage.

Give the **two** half equations that take place during the running of a hydrogen fuel cell. *[2]*

p137–8 4.5.2.1-4.5.2.2 **2–4** Explain why rechargeable batteries need to be recharged while fuel cells do not. *[3]*

Total: 12

6 The rate and extent of chemical change

Rate of reaction

Quick questions

p147	4.6.1.1	WS3	**1** Give a formula to measure the mean rate of reaction.
p151–2	4.6.1.2		**2** List **five** factors that can affect the rate of a reaction.
p153	4.6.1.3	WS4.1	**3** Define the term 'activation energy'.
p154	4.6.1.4		**4** Give the name of a catalyst used in a biological system.
p154	4.6.1.4		**5** Describe how catalysts increase the rate of reaction.
p147	4.6.1.1	WS4.3	**6** Give **two** units for measuring rate of reaction.
p151	4.6.1.3	WS1.2	**7** Use collision theory to explain how a chemical reaction happens.
p152	4.6.1.3		**8** Describe how increasing the concentration of a solution affects the rate of a reaction.

Exam-style questions

9 Magnesium reacts with hydrochloric acid to make hydrogen and magnesium chloride. A student investigated how changing the concentration affected the rate of reaction by collecting the gas that was produced. They plotted their results on a graph, shown in **Figure 1**.

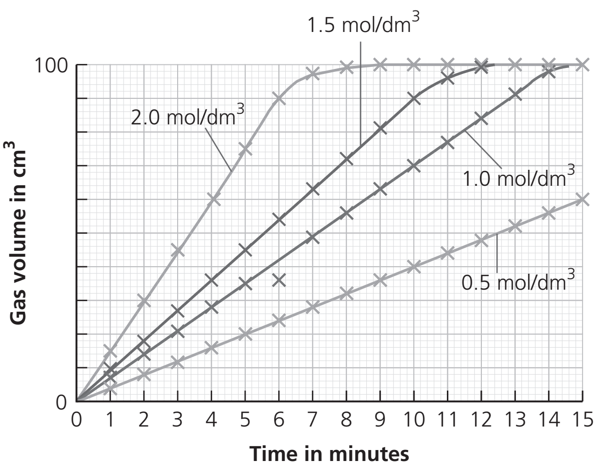


Figure 1

p281 4.1.1.1 — **9–1** Write a word equation for the reaction between magnesium and hydrochloric acid. *[1]*

p207 4.8.2.1 — **9–2** Describe the chemical test that can be used to show that the gas produced was hydrogen. *[2]*

p150 4.6.1.1 WS3.3,4.6 **9–3** Calculate the mean rate of reaction in the first 4 minutes for the 2.0 mol/dm^3 concentration of acid. Give your answer to two significant figures. *[4]*

Total: 7

10 Calcium carbonate is a white basic solid. Hydrochloric acid reacts with calcium carbonate to make water, calcium chloride and a gas. A student investigated how the rate of reaction was affected by the surface area of the calcium carbonate. They used lumps of calcium carbonate and powdered calcium carbonate.

p281 4.4.2.2 **10–1** Name the gas produced in this reaction. *[1]*

p149 4.3.1.3 **10–2** Suggest how the mass might change as the reaction progressed. Explain your answer. *[2]*

p152 4.6.1.3 WS1.2 **10–3** Use collision theory to predict how the rate of reaction would be different when using lumps of calcium carbonate compared with using powder. *[4]*

Total: 7

11 Sodium thiosulfate reacts with hydrochloric acid.

$Na_2S_2O_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + SO_2(g) + S(s)$

p284 4.2.2.2 **11–1** Name the chemical in this reaction that is a solid. *[1]*

p156 4.6.1.2 WS2.2 RP5 QWC **11–2** Describe a method to investigate how changing the concentration could affect the rate of reaction. Explain how you will make this a valid investigation. *[6]*

Total: 7

12 Hydrogen peroxide decomposes to form water and oxygen. This can be represented on an energy level diagram shown in **Figure 2**.

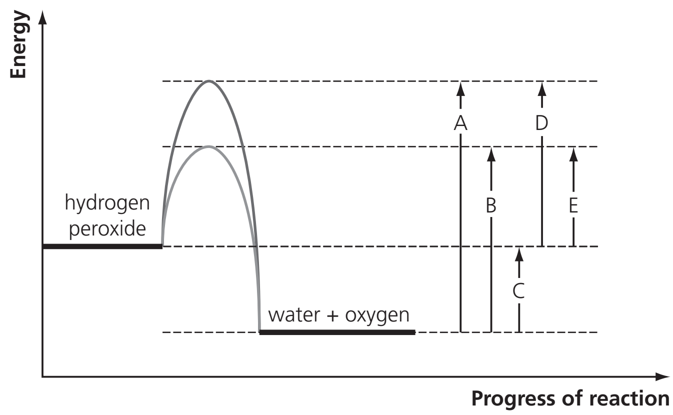


Figure 2

p283 4.3.2.3 MS3c H **12–1** Balance the symbol equation for the reaction. *[1]*

.... $H_2O_2 \rightarrow$ .... O_2 + ... H_2O

p153 4.6.1.4 **12–2** Give the letter on the diagram that represents the activation energy of the reaction **without** a catalyst. *[1]*

p153 4.6.1.4 WS1.2 **12–3** Give the letter on the diagram that represents the activation energy of the reaction **with** a catalyst. *[1]*

Total: 3

13 Lithium reacts with water to make lithium hydroxide and hydrogen.

p284 4.2.2.2

13–1 Copy and complete the balanced symbol equation by adding the state symbols. *[2]*

$2Li$ .... $+ 2H_2O$ .... $\rightarrow 2Li(OH)_2$ .... $+ H_2$ ....

p152 4.6.1 WS2.2 RP5

13–2 A student is investigating how the surface area of the lithium metal affected the rate of reaction. Explain why a method involving a change in colour or turbidity would not give valid results. *[2]*

p153 4.6.1 RP5

13–3 The student collected the gas produced in a measuring cylinder and recorded the volume after 1 minute. Explain why the results would be more accurate if a gas syringe rather than a measuring cylinder was used to collect and measure the volume of gas. *[2]*

p168 WS3.3

13–4 Suggest how the student could improve the reliability of their results. *[3]*

Total: 9

14 Sodium thiosulfate reacts with hydrochloric acid to make a sulfur suspension. A student investigated how the rate of reaction was affected by changing the concentration of sodium thiosulfate. **Table 1** shows the student's results.

Concentration of sodium thiosulfate in arbitrary units	Time in s	Rate of reaction in s^{-1}
0.4	105	0.00950
0.8	79	0.0127
1.2	54	
1.6	32	0.0313

Table 1

p158 4.6.1.1 MS4a,c,d

14–1 Copy and complete **Table 1** by calculating the rate of reaction for a concentration of 1.2 units.

Use the equation rate $= \dfrac{1}{\text{time}}$ *[1]*

p168–9 4.6.1.1

14–2 Plot a graph of concentration against time for the data in **Table 1**. Label the axes. *[4]*

pxxx 4.6.1.1

14–3 Why is it not appropriate to draw a line of best fit on this graph? Justify your answer. *[2]*

Total: 7

p144 4.6.1.1 WS2.2

15 A student investigated the volume of carbon dioxide gas produced when 7.3 g/dm^3 solution of hydrochloric acid reacted with excess powdered calcium carbonate. **Figure 3** shows the student's results.

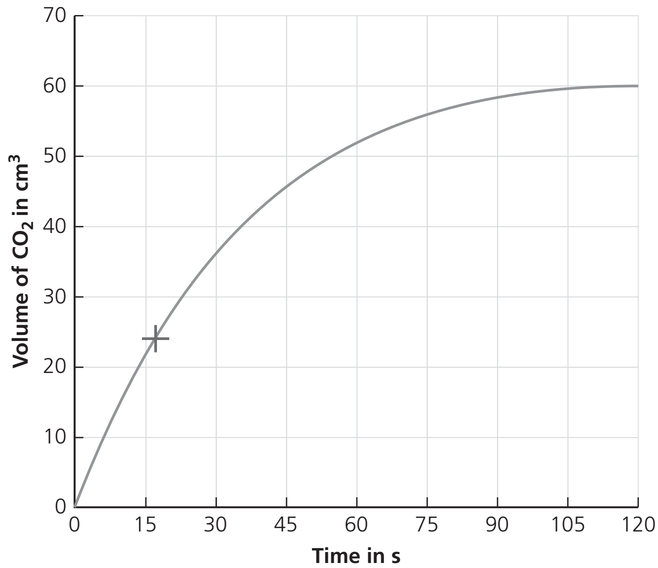


Figure 3

15–1 Identify the independent variable. *[1]*

WS2.2 **15–2** Give the units of the dependent variable. *[1]*

MS4e H **15–3** Calculate the rate of reaction at 18 seconds. *[5]*

15–4 Describe and explain how the graph would look different if a concentration of 3.65 g/dm^3 of hydrochloric acid was used. *[6]*

Total: 13

16 Zinc metal reacts with hydrochloric acid to make zinc chloride and hydrogen.

p285 4.4.1.4 H **16–1** Write an ionic equation for this reaction. *[3]*

p103 4.4.2.1 **16–2** Define a redox reaction and then justify how this reaction can be classified as a redox reaction. *[3]*

p152 4.6.1.3 WS1.2 **16–3** How will the surface area of the zinc affect the rate of the reaction? Use collision theory to explain your answer. *[3]*

Total: 9

Reversible reactions and dynamic equilibrium

Quick questions

p158 4.6.2.1 **1** Give the symbol for a reversible reaction.

p159 4.6.2.2 **2** If the forward reaction of a reversible reaction is exothermic, what is the energy change in the backwards reaction?

p159 4.6.2.3 **3** Describe what is meant by 'dynamic equilibrium'.

p160 4.6.2.4 H **4** Name the principle that can be used to predict the effect of changing conditions on the equilibrium position of a reaction.

p162 4.6.2.6 H **5** A reversible reaction has an exothermic forward reaction. Describe what happens to the yield of the products when the temperature is increased.

p162 &163 4.6.2.7 H **6** Describe what happens to the position of equilibrium when the pressure is reduced for an equilibrium system involving gases.

Exam-style questions

7 Ammonium chloride can thermally decompose to make hydrogen chloride gas and ammonia. This is a reversible reaction.

p159 4.1.1.1, 4.6.2.1 **7–1** Copy and complete the word equation for this reaction. *[1]*

.... ⇌ hydrogen chloride + ...

p130 4.5.1.1, 4.6.2.1 **7–2** Write down whether the forward reaction is exothermic or endothermic. *[1]*

p158 4.6.2.1 **7–3** Define the term 'reversible reaction'. *[1]*

Total: 3

8 Hydrated copper(II) sulfate can undergo a reversible reaction.

The following word equation summarises this reaction:

hydrated copper(II) sulfate ⇌ anhydrous copper(II) sulfate + water

p159 4.6.2.2 **8–1** Write down the colour change you would observe in this reaction. *[2]*

p34 4.2.1.1, 4.2.1.2 **8–2** Describe the type of bonding in anhydrous copper(II) sulfate. Write your answer in terms of forces. *[3]*

p159 4.6.2.2 AT8 **8–3** Suggest how anhydrous copper(II) sulfate could be used as a test for water. Describe any changes you would observe. *[3]*

Total: 8

9 Hydrogen chloride gas (HCl) and ammonia gas (NH_3) react in a reversible reaction.

p159 4.6.2.1 MS3c **9–1** Write a balanced symbol equation for this reaction. *[3]*

p159 4.6.2.1, 4.5.1.1 **9–2** The forward reaction has an energy change of –176 KJ/mol. Give the energy change for the reverse reaction. *[1]*

p160 &161 4.6.2.4, 4.6.2.6 H **9–3** Predict what would happen to the amount of reactants if the reaction was carried out in a sealed container and the temperature was increased. Explain your answer. *[4]*

Total: 8

10 In a chemical reaction, reactants A and B undergo a reversible reaction to make products C and D.

p158 4.6.2.1 **10–1** Write an equation to describe this reaction. *[2]*

p149 4.6.2.3, 4.6.1.1 **10–2** Sketch a graph to show the change in the concentration of the chemicals from the start of the reaction to when the system reaches dynamic equilibrium. *[5]*

Total: 7

11 The Haber process is an important method for manufacturing fertiliser. Nitrogen from the air is reacted with hydrogen to make ammonia.

p159 &283 | 4.10.4.1 | MS3c

11–1 Copy and complete the symbol equation by balancing. *[1]*

.... N_2 + ... $H_2 \rightleftharpoons$ .... NH_3

p162 &163 | 4.6.2.4, 4.6.2.6

(H) **11–2** The forward reaction is exothermic. When the system is at equilibrium, suggest the effect of increasing the temperature on the yield of ammonia. Explain your answer. *[3]*

p164 | 4.10.4.1, 4.6.2.4, 4.6.1.4

(H) **11–3** Name the catalyst used in this reaction. Explain the effect of the catalyst on the position of equilibrium. *[3]*

Total: 7

p161–3 | 4.6.2.4, 4.6.2.6, 4.6.2.7

12 The Haber process makes gaseous ammonia. **Figure 4** shows how the yield changes when the temperature and pressure are changed.

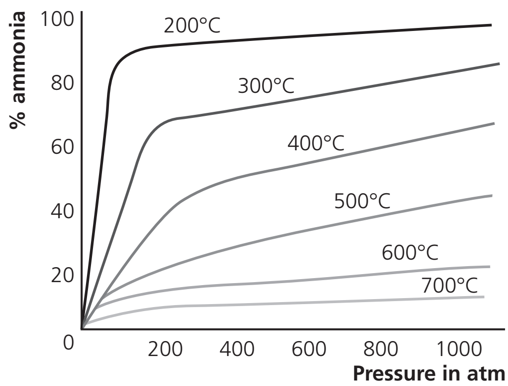


Figure 4

MS1b

12–1 One atmosphere pressure is 101 000 Pa. Convert 200 atmospheres into Pa. Give your answer in standard form. *[2]*

12–2 Give the conclusions that can be drawn from the graphs in **Figure 4**. *[2]*

QWC

(H) **12–3** Use **Figure 4** and your own knowledge to justify the use of a temperature of about 450°C and a pressure of about 200 atmospheres in the industrial Haber process. *[6]*

Total: 10

13 The Haber process is used to manufacture ammonia (NH_3), which can be used to produce nitrogen-based fertilisers. In a reversible reaction, nitrogen and hydrogen are reacted to make the ammonia. The ammonia gas is then condensed and run off.

p79&80 | 4.3.3.2 | MS1c

13–1 Calculate the atom economy of this reaction. *[2]*

p284 | 4.2.2.2

13–2 Write a balanced symbol equation to show ammonia gas condensing. *[1]*

p162 &163 | 4.6.2.4, 4.6.2.5

(H) **13–3** At equilibrium, suggest what would happen to the yield of ammonia if more hydrogen gas was added to the system. Explain your answer. *[3]*

Total: 6

14 The Contact process is used to make sulfuric acid. Sulfur dioxide (SO_2) is combusted to make sulfur trioxide (SO_3) in a reversible reaction.

p283 | 4.5.1.1 | MS3c

14–1 Write a balanced symbol equation for this reaction. You do not need to include state symbols. *[4]*

p129 | 4.5.1.1

14–2 Write down whether the forward reaction is exothermic or endothermic. *[1]*

p162 | 4.6.2.4, 4.6.2.6

(H) **14–3** When the system is in equilibrium, suggest the effect of decreasing the temperature on the yield of sulfur dioxide. Explain your answer. *[3]*

Total: 8

15 Ammonium chloride is a white solid. It breaks down when heated, forming ammonia and hydrogen chloride gases.

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

p159 | 4.4.2.1

15–1 Describe what you would observe when ammonium chloride is heated in a test tube. *[2]*

p162 &163 | 4.6.2.4

(H) **15–2** When the system is at equilibrium, suggest the effect of increasing the pressure of the system. Explain your answer. *[2]*

p199 &200 | 4.6.2.4, 4.6.2.6 | WS1.5

15–3 Suggest **one** safety precaution when carrying out this experiment. Explain why this is necessary. *[2]*

Total: 6

16 Hydrogen is a fuel and can be made by an endothermic reaction between carbon and steam.

p284 | 4.2.2.2

16–1 Copy and complete the equation for this reaction by adding the state symbols. *[2]*

C .... + H_2O .... $\rightleftharpoons$ H_2 .... + CO ....

p162 | 4.6.2.4, 4.6.2.6

(H) **16–2** When the system is in equilibrium, suggest the effect of increasing the temperature on the amount of hydrogen. Explain your answer. *[3]*

Total: 5

17 Iodine and hydrogen react in a reversible reaction to make hydrogen iodide. **Figure 5** is a sketch graph to show how the concentration of the chemicals change during the reaction.

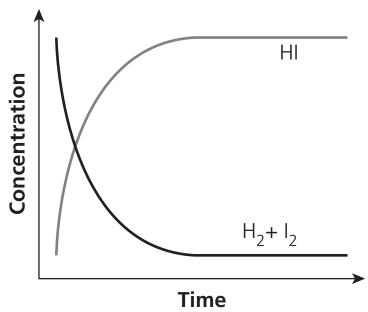


Figure 5

p85&86 | 4.3.2.5, 4.3.4

(H) **17–1** Suggest a suitable unit in which to measure concentration. *[1]*

p149& 159 | 4.6.2.3, 4.6.1.1

17–2 On a copy of the graph in **Figure 5**, show the time at which equilibrium was reached. *[1]*

p283 | MS3c

17–3 Write a balanced symbol equation for the reaction between iodine and hydrogen. *[3]*

Total: 5

18 Phosphorus pentachloride can thermally decompose into chlorine and phosphorus trichloride. This reaction is reversible.

$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$

p162 4.5.1.1, 4.6.2.4, 4.6.2.6 (H) **18–1** When the system is in equilibrium, suggest the effect of decreasing the temperature on the yield of chlorine. Explain your answer. *[3]*

p162,163 4.6.2.5, 4.6.2.6 **18–2** When the system is in equilibrium, suggest the effect of decreasing the pressure of the system. Explain your answer. *[3]*

Total: 6

19 Ethanol (C_2H_5OH) is an important fuel that can be made by reacting ethene (C_2H_4) with steam. The reaction is exothermic and reversible.

p283 MS3c **19–1** Write a balanced symbol equation for this reaction. *[3]*

p79&80 4.3.3.1 MS1c **19–2** Explain why the atom economy is 100%. *[3]*

p162 4.6.2.4, 4.6.2.6 **19–3** When the system is in equilibrium, suggest the effect of increasing the temperature on the yield of ethanol. Explain your answer. *[3]*

p162& 163 4.6.2.4, 4.6.2.7 (H) **19–4** When the system is in equilibrium, suggest the effect of increasing the pressure of the system. Explain your answer. *[3]*

Total: 12

20 Hydrogen and chlorine gases react together to make hydrogen chloride gas. This is a reversible reaction.

p284 4.2.2.2 **20–1** Copy and complete the equation for this reaction by adding the state symbols. *[1]*

H_2 .... + Cl_2 .... $\rightleftharpoons$ 2HCl ....

p162 &163 4.6.2.7 **20–2** Explain why changing pressure has no effect on the yield of hydrogen chloride. *[2]*

Total: 3

21 Methanol is a fuel and a solvent. It can be made by reacting carbon monoxide with hydrogen. The reaction vessel is kept at about 50 atmospheres pressure and 200 °C. This is an exothermic reaction.

p283 MS3c **21–1** Balance the symbol equation for this reaction. *[1]*

.... CO(g) + ... $H_2(g) \rightleftharpoons$ .... $CH_3OH(g)$

p160–4 4.5.1.1 QWC **21–2** Describe the conditions needed to make the maximum yield of methanol. Explain your answer. *[4]*

p267 4.6.2.4, 4.6.2.6, 4.6.2.7 QWC (H) **21–3** Justify the conditions used in industry for the production of methanol. Explain why they do not give the highest possible yield. *[6]*

Total: 11

22 In a reversible reaction, dinitrogen tetroxide (N_2O_4) will decompose to make nitrogen dioxide (NO_2).

p283 MS3c **22–1** Write a balanced symbol equation for this reaction. *[3]*

p149 | 4.6.2.3, 4.6.1.1 | MS4c

22–2 On **Figure 6**, sketch a graph to predict how the concentrations of dinitrogen tetroxide (N_2O_4) and nitrogen dioxide (NO_2) will change from the start of the reaction to when it reaches dynamic equilibrium. *[4]*

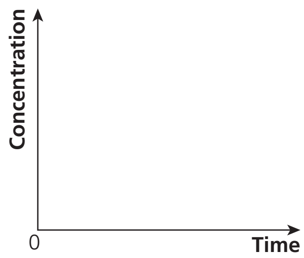


Figure 6

P149& 150 | 4.6.2.3, 4.6.1.1

22–3 Label your graph to show when equilibrium is reached. *[1]*

Total: 8

23 In the Contact process, sulfur dioxide is converted to sulfur trioxide by a reversible reaction with oxygen. **Figure 7** shows how the rate of conversion is affected by temperature.

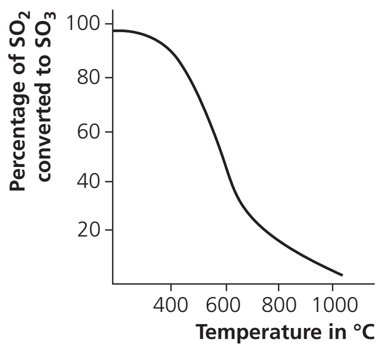


Figure 7

p129& 162 | 4.6.2.1

23–1 Use the graph and your own knowledge to determine whether the reaction is exothermic or endothermic. Justify your answer. *[3]*

Total: 3

24 The concentrations of each chemical in the Haber process can be monitored. **Figure 8** shows a graph of these results.

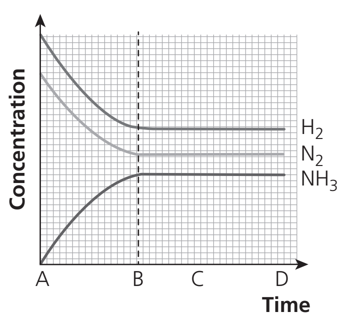


Figure 8

p149& 159 | 4.6.1.2

24–1 Use your own knowledge and **Figure 8** to describe what is happening at point B. Explain your answer. *[3]*

p161 4.10.4.1, 4.6.2.4

H **24–2** When this reaction is carried out in industry, the ammonia is constantly being removed from the reaction chamber. Suggest a reason for this. Justify your answer in terms of yield. *[3]*

Total: 6

25 The Haber process can be described by the following balanced symbol equation:

$N_2 + 3H_2 \rightleftharpoons 2NH_3$

The concentrations of chemicals in the Haber process were monitored. **Figure 9** shows the results.

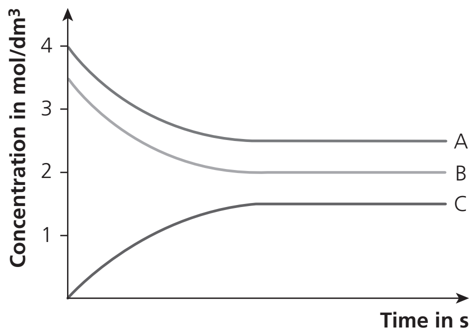


Figure 9

p149& 159 4.6.1.1, 4.10.4.1 MS1c

25–1 Use the graph and your own knowledge to suggest:

- which line is the product *[2]*
- which line is hydrogen. *[3]*

Explain your answers.

p149 4.6.1.1

25–2 Use the graph to estimate the equilibrium concentrations of each chemical. *[3]*

p149& 164 4.6.1.1, 4.6.1.4

25–3 Suggest how the graph would change if a catalyst is used in the reaction. *[3]*

Total: 11

The rate and extent of chemical change topic review

1 Hydrogen peroxide (H_2O_2) can decompose in UV light to produce water and a gas. The gas will relight a glowing splint.

p207 4.8.2.2 — 1–1 Identify the gas produced. *[1]*

p153 4.6.1.4 — 1–2 Manganese dioxide is a catalyst for this reaction. Explain the purpose of a catalyst. *[2]*

p221 4.6.1.4 AT5 — 1–3 Suggest **one** observation for this reaction. Explain your answer. *[2]*

p153 4.6.1.4 — 1–4 **Figure 10** is an incomplete energy level diagram for this reaction. Copy the diagram and label A, B, C and D. *[4]*

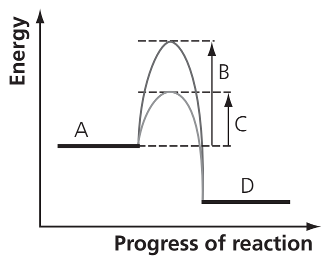


Figure 10

Total: 9

2 A student investigated the rate of reaction between calcium carbonate cubes and hydrochloric acid to produce a salt, water and carbon dioxide.

p113 4.8.3.3, 4.4.2.3, 4.1.1.1 — 2–1 Write a word equation for this reaction. *[2]*

p207 4.8.2.3 QWC — 2–2 Describe a chemical test to show that the gas produced was carbon dioxide. *[4]*

4.6.1.3 MS5a,5b — 2–3 The student carried out the investigation twice. First, she used a single cube of calcium carbonate. Then she used four smaller cubes of calcium carbonate. The dimensions of the cubes are shown in **Figure 11**.

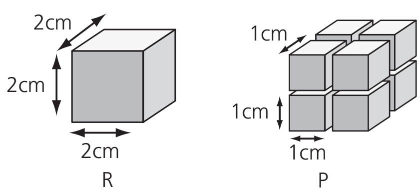


Figure 11

- Calculate the surface area of cube R. Give the unit with your answer. *[3]*
- Calculate the total volume of the cubes that make up P. Give the unit with your answer. *[3]*

4.6.1.2, 4.6.1.3 WS1.2 MS5c,MS1c QWC — 2–4 Use your own knowledge of collision theory and **Figure 11** to predict whether the cube (R) or the four smaller cubes (P) would have the fastest rate of reaction. Explain your answer. *[6]*

4.6.1.1 MS1d, MS4b RP5

2–5 Sketch a graph to show how the amount of gas collected would change throughout the reaction for the cube (R) and the four smaller cubes (P) that the student investigated. *[5]*

Total: 23

3 Methane (CH_4) can undergo a reversible reaction with steam to make hydrogen and carbon monoxide (CO). The reaction is endothermic. In a closed system, a dynamic equilibrium is established.

p162& 283 4.6.2.1

3–1 Write a balanced symbol equation for this reaction. *[4]*

p162 4.6.2.4, 4.6.2.6

3–2 Describe the effect of increasing the temperature on this reaction. Explain your answer. *[4]*

Total: 8

7 Organic chemistry

Carbon compounds as fuels and feedstock

Quick questions

p171 4.7.1.1 **1** What is 'crude oil'?

p172 4.7.1.1 **2** What is a 'hydrocarbon'?

p172 4.7.1.1 **3** What is a 'homologous series'?

p172 4.7.1.1 **4** What is the general formula of an alkane?

p172 4.7.1.1 **5** List the names of the first **four** members of the alkanes.

p173 4.7.1.2 **6** Name the process used to separate crude oil.

p174 4.7.1.3 **7** Write down **two** things that happen during the combustion of hydrocarbon fuels.

p174 4.7.1.3 **8** What are the products of the complete combustion of a hydrocarbon fuel?

p176 4.7.1.4 **9** What is 'cracking'?

p177 4.7.1.4 **10** Name the **two** types of cracking.

p177 4.7.1.4 **11** What are 'alkenes'?

Exam-style questions

12 Crude oil is a mixture of hydrocarbons. Methane is a hydrocarbon.

p41 4.1.1.1 **12–1** Write down the number of atoms in methane, CH_4. *[1]*

p41 4.2.1.4 **12–2** Describe the bonding in methane. *[3]*

Total: 4

13 Alkanes are a homologous series. **Figure 1** shows a diagram of two alkanes, ethane and butane.

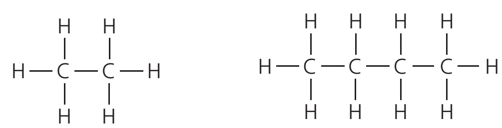


Figure 1

p41&172 4.1.1.1, 4.7.1.1 **13–1** Give the molecular formula of ethane. *[1]*

p173 4.1.1.2, 4.7.1.2 **13–2** Name the technique that could be used to separate ethane and butane. *[1]*

p39&175 4.2.2.4, 4.7.1.3 **13–3** Suggest which of these alkanes would have the highest boiling point. Explain your answer. *[4]*

p172 4.7.1.1 **13–4** Write down the molecular formula for an alkane with 10 carbon atoms. *[1]*

Total: 7

14 Crude oil is a mixture of hydrocarbons. It is separated using fractional distillation. Many of the products are used as fuels.

p174 | 4.7.1.3

14–1 Decane, $C_{10}H_{22}$, is used in petrol. Copy and complete the word equation for the combustion of decane. *[2]*

decane + oxygen → __________________ + __________________

p129 | 4.5.1.1, 4.7.1.3

14–2 Is this reaction exothermic or endothermic? Explain your answer. *[2]*

p175 | 4.7.1.3

14–3 Pentadecane ($C_{15}H_{32}$) is a different alkane found in petrol. Compare the physical properties of decane and pentadecane. *[3]*

Total: 7

15 Crude oil is a mixture of hydrocarbons. A teacher demonstrated how these can be separated in a laboratory using fractional distillation. **Figure 2** shows the equipment used.

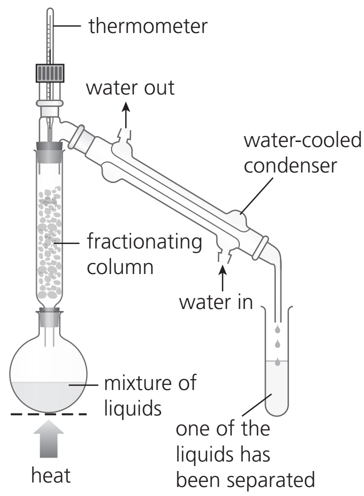


Figure 2

p24 | 4.7.1.2 | WS2.2

15–1 Explain why the bulb of the thermometer must be in line with the side arm. *[2]*

p199& 200 | 4.7.1.2 | AT2,6

15–2 Suggest a safer alternative to using a Bunsen burner. Explain your answer. *[3]*

Total: 5

16 A student is breaking down long chain hydrocarbons in a laboratory using catalytic cracking. **Figure 3** shows the equipment used.

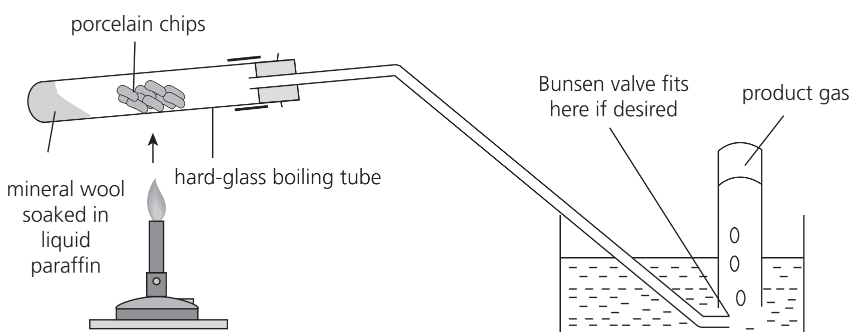


Figure 3

p199 4.7.1.4 AT2,6 **16–1** Identify a hazard in this experiment. Describe how the risk could be reduced. *[2]*

p179 4.7.1.4 QWC **16–2** Describe a chemical test to show that an alkene has been made. Include any observations with your answer. *[6]*

p153 4.6.1.4, 4.7.1.4 **16–3** The porcelain chips act as a catalyst. Explain the purpose of the catalyst in this reaction. *[3]*

Total: 11

17 Alkanes are a family of hydrocarbons.

Table 1 shows information about the first four members of the alkane homologous series.

Name of alkane	How many carbons?	Chemical formula	Boiling point in °C	State at 20 °C	Melting point in °C
Methane	1	**X**	–162	**Y**	–183
Ethane	2	C_2H_6	–89	gas	–172
Propane	**Z**	C_3H_8	–42	gas	–188
Butane	4	C_4H_{10}	0	gas	–138

Table 1

p168 4.7.1.3, 4.1.1.1 **17–1** Write down the missing data from boxes **X, Y** and **Z** in **Table 1**. *[3]*

p244 4.2.2.4, 4.7.1.3 **17–2** Describe the relationship between number of carbon atoms and boiling point. *[1]*

p39&175 4.2.2.4, 4.7.1.3 **17–3** Use the data in **Table 1** and your own knowledge to explain the effect of changing the number of carbon atoms on the melting point of an alkane. *[3]*

Total: 7

18 The petrochemical industry uses some of the heavier fractions to make petrol.

p177 4.5.1.1, 4.7.1.4 **18–1** Justify the classification of cracking as thermal decomposition. *[2]*

p176 4.7.1.4 QWC **18–2** Explain why steam cracking uses less energy than catalytic cracking. In your answer, you should outline the conditions used for both types of cracking. *[6]*

p283 4.7.1.4 MS3c **18–3** The equation shows the cracking of decane. Copy and balance the equation. *[1]*

.... $C_{10}H_{22} \rightarrow$ .... C_2H_4 + ... C_6H_{14}

Total: 9

19 Crude oil is an important source of organic chemicals.

p171 4.7.1.1 **19–1** Describe how crude oil is formed. *[4]*

p21&172 4.7.1.1 **19–2** Justify why crude oil cannot be described as pure. *[1]*

p172 4.7.1.1 **19–3** Explain why it is necessary to separate crude oil. *[2]*

p21,173 &204 4.7.1.1, 4.8.1.2 **19–4** Explain why petrol is a formulation but the petrol fraction from a fractionating column is not. *[4]*

Total: 11

20 Alkanes are a homologous series of chemicals.

p172 | 4.2.1.4, 4.7.1.1 | MS5b — 20–1 Draw the displayed formula of propane. *[2]*

p174& 283 | 4.7.1.3 | MS3c — 20–2 Write a balanced symbol equation for the complete combustion of propane. You do not need to write the state symbols. *[3]*

p174 | 4.7.1.3 | QWC — 20–3 Is this reaction reduction or oxidation? Justify your choice. *[6]*

Total: 11

21 Paraffin ($C_{10}H_{22}$) can undergo cracking to make ethene (C_2H_4) and one other product.

p283 | MS3c — 21–1 Write a balanced symbol equation for this reaction. *[1]*

p42 | 4.2.1.4, 4.7.2.1 | MS5b — 21–2 Draw a dot and cross diagram to show the bonding in ethene. *[3]*

p172& 177 | 4.2.1.4, 4.7.1.3, 4.7.1.4 | QWC — 21–3 Describe the similarities and differences between ethene and ethane. Give your answer in terms of structure and bonding. *[6]*

p176 | 4.7.1.4 — 21–4 Suggest a use for alkenes. *[1]*

Total: 11

22 In industry, crude oil is separated using fractional distillation. **Figure 4** shows an industrial fractionating column.

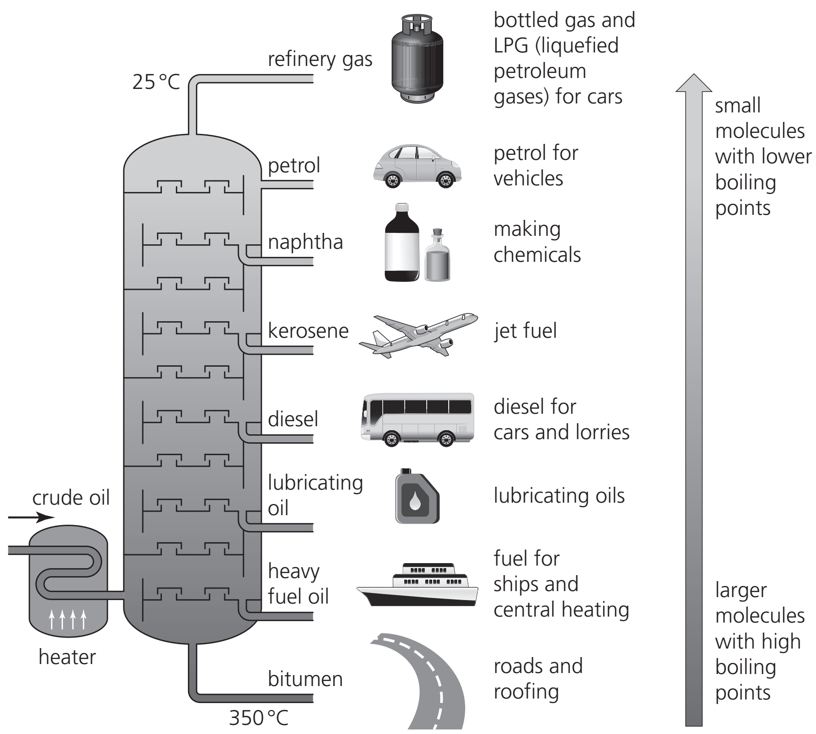


Figure 4

p173& 174 | 4.7.1.2 | QWC — 22–1 Use **Figure 4** and your knowledge to explain how crude oil is separated. *[6]*

Total: 6

23 Petrol and lubricating oil are both fractions of crude oil. A student has two unnamed samples which they want to identify.

p175 4.7.1.3 AT8 **23–1** Suggest how the student could determine which sample is petrol and which is lubricating oil by:

- using colour *[3]*
- using viscosity *[3]*
- using flammability. *[3]*

Your answers should describe how the student should carry out each test and the results they would obtain.

Total: 9

24 Long chain hydrocarbons may undergo cracking to make smaller molecules. Cracking is often random, and a mixture of products is made.

p172 4.7.1.1 MS1c **24–1** Write down the molecular formula for an alkane with 25 carbon atoms. *[1]*

p283 4.7.1.4 MS3c **24–2** Write a balanced symbol equation for the thermal cracking of octane (C_8H_{18}) to make two ethene (C_2H_4) molecules and one other product. *[3]*

p207 4.8.2.1 AT8 **24–3** A different cracking reaction makes a colourless, odourless gas which ignites with a squeaky pop. Identify the gas. *[1]*

p24 4.1.1.2 **24–4** Suggest how the products of cracking could be separated. *[1]*

Total: 6

Reactions of alkenes and alcohols

Quick questions

p177 4.7.2.1 1 What are 'alkenes'?

p178 4.7.2.1 2 Write down the general formula of alkenes.

p177 4.7.2.1 3 List the first **four** members of the alkene homologous series.

p180 4.7.2.2 4 What do alkenes react with in combustion reactions?

p178 4.7.2.2 5 How do alkenes react with hydrogen, water and halogens?

p181 4.7.2.3 6 Write down the functional group for the alcohol homologous series.

p181 4.7.2.3 7 List the first **four** members of the alcohol homologous series.

p181 4.7.2.3 8 Name the process used to make an aqueous solution of ethanol.

p183 4.7.2.4 9 List the first **four** members of the carboxylic acid homologous series.

p184 4.7.2.4 10 Why are carboxylic acids weak acids?

Exam-style questions

11 Propene is an alkene. The displayed formula of propene is shown in **Figure 5**.

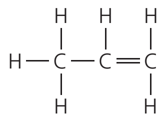


Figure 5

p178 4.7.2.1 **11–1** Write down the functional group of the alkene homologous series. *[1]*

p177 4.7.2.1 MS1c **11–2** Give the molecular formula of propene. *[1]*

p179 4.7.2.2 **11–3** Copy and complete the word equation for the reaction between steam and propene. *[1]*

propene + steam → ....

p178 4.7.2.2 **11–4** State the type of chemical reaction that happens when propene reacts with chlorine. *[1]*

Total: 4

12 Ethene is the first member of the homologous series of alkenes.

p177 4.7.2.1 MS5b **12–1** Draw the displayed formula of ethene. *[3]*

p172 4.7.2.1 **12–2** Explain why ethene can be described as a hydrocarbon. *[2]*

p177 4.7.2.1 **12–3** Justify how ethene can be described as unsaturated. *[1]*

Total: 6

13 Butene (C_4H_8) is an alkene.

p177 4.2.1.4 MS1c **13–1** Write down the number of atoms in a molecule of butene. *[1]*

p174, 175&180 4.7.1.3, 4.7.2.2 **13–2** Compare the combustion reaction of butene with the combustion reaction of butane. Your answer should include the names of any products formed. *[4]*

p180 4.7.2.2 **13–3** Write a word equation for the combustion of butene to make carbon dioxide, carbon and one other product. *[2]*

Total: 7

14 Pentene (C_5H_{10}) is a reactive hydrocarbon.

p177 4.7.2.1 **14–1** Identify the homologous series that pentene belongs to. *[1]*

p178& 283 4.7.2.2 MS1c **14–2** Pentene reacts with chlorine to make one product. Copy and complete the symbol equation for this reaction. *[1]*

$C_5H_{10} + Cl_2 \rightarrow$ ....

p179 4.7.2.2 **14–3** Pentene will react with hydrogen to make pentane. Identify this type of reaction. *[1]*

p181 4.7.2.2, 4.7.2.3 MS5b **14–4** Pentene reacts with steam. Draw the displayed formula of the product formed. Name the product. *[2]*

Total: 5

15 Alcohols are a homologous series and can be used as fuels.

p172 4.7.1.1 **15–1** Explain why alcohols are **not** classified as hydrocarbons. *[1]*

p181 4.7.2.3 **15–2** Copy and complete the word equation for the complete combustion of ethanol. *[2]*

ethanol + oxygen → .... + ...

p182 4.7.2.3 **15–3** Suggest what will happen if a sample of methanol is left open to the air. *[2]*

Total: 5

16 Ethanol is the second member of the alcohol homologous series. Ethanol can be made from a fermentation reaction using the apparatus in **Figure 6**.

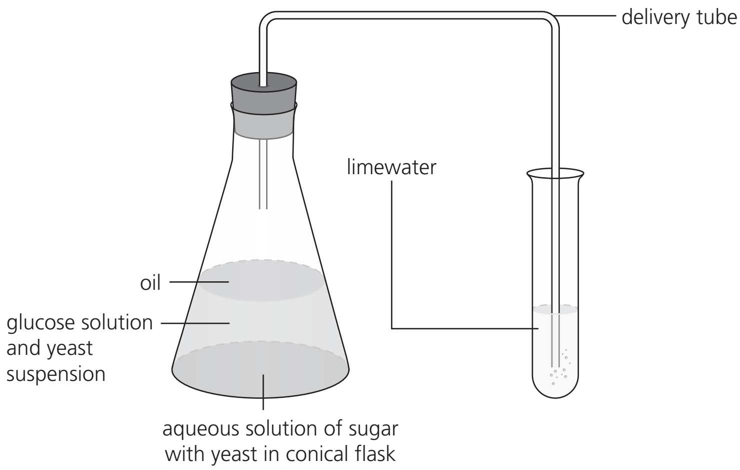


Figure 6

4.7.2.3 MS1c **16–1** Copy and balance the symbol equation for the fermentation of glucose using yeast. *[1]*

........$C_6H_{12}O_6 \rightarrow$ ....C_2H_5OH + ...CO_2

p181& 207 4.7.2.3, 4.8.2.3 AT5 **16–2** Suggest what observations you would make during this reaction. *[3]*

p258 4.7.2.3 **16–3** Suggest the purpose of the oil on the glucose solution. *[3]*

p23 4.1.1.2 AT4 **16–4** Ethanol made from fermentation is a mixture. Describe a method for collecting a pure sample of ethanol. *[4]*

Total: 11

17 Vinegar has a sharp taste and a pH less than 7. It is a carboxylic acid known as ethanoic acid.

p183 4.7.2.4 **17–1** Write down the functional group of ethanoic acid. *[1]*

p185 4.7.2.4 **17–2** Suggest the product when ethanoic acid is reacted with ethanol. *[1]*

p283 4.7.2.4 MS1c **17–3** Ethanoic acid ($C_2H_4O_2$) will react with calcium carbonate ($CaCO_3$) to make calcium ethanoate ($Ca(C_2H_3O_2)_2$), water (H_2O) and a gas. Write a balanced symbol equation for this reaction. *[3]*

p207 4.8.2.3 **17–4** Describe a method to identify the gas produced. *[2]*

Total: 7

18 Ethanoic acid can be used to pickle vegetables. The displayed formula of ethanoic acid is shown in **Figure 7**.

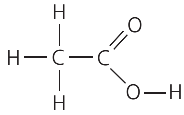


Figure 7

p40,41 &183 | 4.2.1.4 | MS5b — **18–1** Give the molecular formula of ethanoic acid. *[1]*

p64 &65 | 4.3.1.2 | MS3a, 3c — **18–2** Calculate the relative formula mass of ethanoic acid given that the relative atomic mass (A_r) of C=12, H=1 and O=16. *[2]*

p109& 184 | 4.4.2.6 | MS1c — (H) **18–3** Write down a balanced symbol equation that shows ethanoic acid behaving as a weak acid. *[3]*

p108 &109 | 4.4.2.6, 4.7.2.4 — (H) **18–4** Suggest the pH of a solution of ethanoic acid. *[1]*

Total: 7

19 **Figure 8** shows some examples of structures of organic molecules.

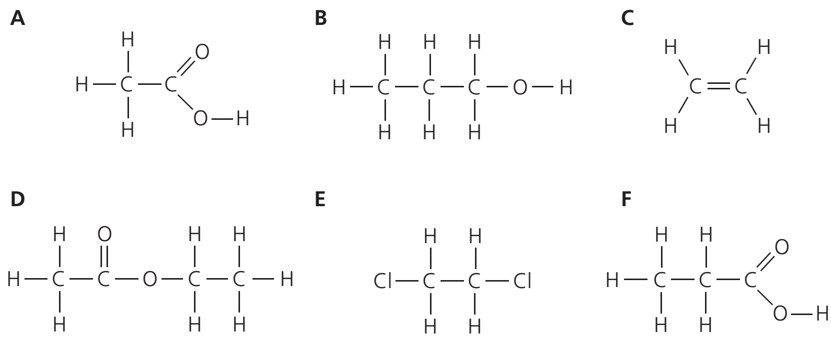


Figure 8

Write the letter of the compound that:

p172 | 4.7.1.1 — **19–1** is a hydrocarbon *[1]*

p187 | 4.7.3.1, 4.2.2.5 — **19–2** can be used to make addition polymers *[1]*

p178 &283 | 4.7.2.2 — **19–3** is the product of the reaction between ethene and a halogen *[1]*

p182 | 4.7.2.2 — **19–4** is the product of the reaction when ethanol is left open to the air *[1]*

p184 &185 | 4.7.2.2 — **19–5** is the product of the reaction between A and an alcohol. *[1]*

Total: 5

20 Ethene is the first member of the alkene homologous series.

p284 | 4.2.2.2 — **20–1** Copy the balanced symbol equation for the reaction of steam with ethene below and add the state symbols. *[2]*

C_2H_4 .... + H_2O .... → C_2H_6O (g)

p179, 283& 284 | 4.7.2.2 | MS1c — **20–2** Write the balanced symbol equation for the reaction of ethene with hydrogen. *[2]*

p178 | 4.7.2.2 — **20–3** Name the type of reaction when iodine reacts with ethene. *[1]*

Total: 5

21 Ethanol is an important fuel. Pure ethanol can be produced in industry by fermentation or by hydration of ethene. **Table 2** summarises the two methods.

	Fermentation	**Hydration of ethene**
Raw materials	Sugar from plants	Crude oil that has undergone fractional distillation and then cracking to make ethene
Energy cost	Low	High
Process	Batch	Continuous
Atom economy	51%	100%
Purity of product	Needs further processing such as distillation	Pure

Table 2

p79,80, 179&181 | 4.7.2.3, 4.3.3.2 | WS1.4 | MS1c | QWC

21–1 Use information from **Table 2** and your own knowledge to evaluate the two methods of production of ethanol. Your answer should include the balanced symbol equations for each reaction. *[6]*

Total: 6

22 Carboxylic acids are a homologous series of organic compounds. They can be made from alcohols in an oxidation reaction.

p183 | 4.7.2.4

22–1 Give the molecular formula of methanoic acid. *[1]*

p183 | 4.7.2.4 | MS5b

22–2 Draw the displayed formula of propanoic acid. *[2]*

p184 | 4.7.2.4

22–3 Ethanoic acid can react with ethanol to make water and another organic chemical. Give the name of the organic product and the homologous series that it belongs to. *[2]*

p109& 184 | 4.7.2.4, 4.4.2.6 | H

22–4 Describe what happens when butanoic acid is added to an equal volume of water and stirred. *[4]*

Total: 9

23 Ethene and ethane are both hydrocarbons. **Figure 9** shows the structure of these two molecules.

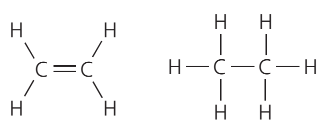


Figure 9

p174, 178& 179 | 4.7.2.2, 4.7.1.3, 4.7.1.4 | QWC

23–1 Describe the difference in reactions between ethene and ethane. Explain your answer in terms of bonding. *[6]*

Total: 6

24 Pentene is an organic molecule.

p177 | 4.7.2.1 | MS1c

24–1 Write down the molecular formula of pentene. *[1]*

p177 | 4.7.2.1

24–2 Explain why pentene is unsaturated whereas pentane is saturated. *[2]*

p178& 179 | 4.7.2.2 | MS1c

24–3 Write a balanced symbol equation for the reaction of pentene with bromine. *[2]*

Total: 5

25 Sodium is a Group 1 metal and can react with alcohols to make a salt and hydrogen.

p182, 283&284 | 4.7.2.3 | MS1c

25–1 Copy and balance the symbol equation for the reaction of methanol with sodium. *[1]*

..............Na + ...HCOOH → ....NaCOOH + ...H_2

p207 | 4.8.2.1 | AT6,3,8 | QWC

25–2 Describe a method to collect the gas produced in this reaction. Include the test and the observations that will identify the gas as hydrogen. Give a risk assessment with your answer. *[6]*

Total: 7

26 Esters can be made by reacting an alcohol with a carboxylic acid as shown in the equation in **Figure 10**.

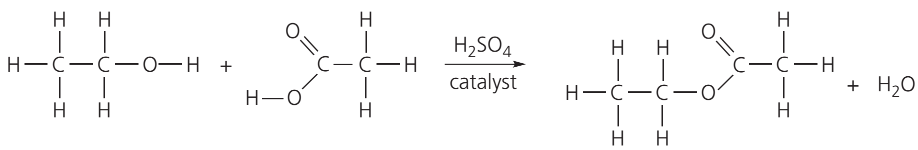


Figure 10

p181 | 4.7.2.3 | MS5b

26–1 Copy Figure 10 and circle the alcohol functional group. *[1]*

p184 | 4.7.2.4

26–2 Name the two products of this reaction. *[2]*

p153 | 4.6.1.4

26–3 Explain the role of the catalyst. *[3]*

Total: 6

27 A solution of ethanol was left open to the air and a chemical reaction took place. **Figure 11** shows the displayed formula of the organic product.

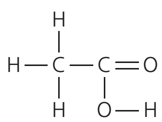


Figure 11

p39 | 4.2.1.4

27–1 Describe the structure and bonding in this molecule. *[3]*

p181 | 4.7.2.3

27–2 Give the name and homologous series of this chemical. *[2]*

p283& 284 | 4.7.2.3 | MS1c

27–3 Copy and then balance the symbol equation for the complete combustion of this chemical. *[1]*

..............$C_2H_4O_2$ + ...O_2 → ....CO_2 + ...H_2O

p23&181 | 4.1.1.2, 4.7.2.3 | QWC

27–4 Ethanol can be made by fermentation. Describe a method to make a sample of pure ethanol using yeast and a sugar solution. Include the reaction conditions in your answer. *[6]*

Total: 12

Synthetic and naturally occurring polymers

Quick questions

p187 4.7.3.1 — **1** Describe what happens in addition polymerisation reactions.

p187 4.7.3.1 — **2** Name the polymer made from butene.

p188 4.7.3.1 — **3** Name the monomer used to make poly(propene).

p189 4.7.3.1 — **4** Name the monomers needed to make a polyester.

p192 4.7.3.3 (H) — **5** What do amino acids make when they polymerise?

p193 4.7.3.3 (H) — **6** What is a protein made from?

p193 4.7.3.4 — **7** What does DNA stand for?

p193 4.7.3.4 — **8** Give **three** examples of natural polymers that are important to life.

p193 4.7.3.4 — **9** What is the shape of DNA?

Exam-style questions

10 Poly(ethene) is an addition polymer.

p187 4.7.3.1 — **10–1** Name the monomer used to make poly(ethene). *[1]*

p187 4.7.3.1 WS1.2 — **10–2** Draw the repeating unit of poly(ethene). *[3]*

p187 4.7.3.1 WS1.2 — **10–3** Explain why the repeating unit of the polymer will have the same relative formula mass as the monomer. *[2]*

Total: 6

11 Polyesters are often used to make fibres and fabrics. They are made from two different monomers.

p189 4.2.1.4, 4.7.3.2 MS5b (H) — **11–1** Ethane-diol is one of these monomers. It has the structural formula shown in **Figure 12**.

$HO—CH_2—CH_2—OH$

Figure 12

Give the molecular formula of ethane-diol. *[1]*

p189 4.7.2.4, 4.7.3. (H) — **11–2** Hexandioic acid is the second monomer that is used. Give its functional group and homologous series. *[2]*

The polyester can be represented in a block diagram as shown in **Figure 13**.

Figure 13

p189 4.7.3.2 WS1.2 MS5b (H) — **11–3** Draw the block structures for the two monomers. *[2]*

p189 4.2.1.4, 4.7.3.2 (H) — **11–4** Give the formula of the second product of this polymerisation reaction. *[1]*

Total: 6

12 Amino acids are the important building blocks of life. Glycine is the simplest amino acid. **Figure 14** shows the displayed formula of glycine.

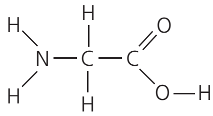


Figure 14

p192 | 4.7.3.3 | MS5b | H

12–1 Copy **Figure 14** and circle the two functional groups on the displayed formula of glycine. *[2]*

p192& 193 | 4.7.3.3 | WS1.2 MS5b | H

12–2 Glycine can react with itself to make a polymer. Write a symbol equation to show this reaction. *[3]*

p192& 193 | 4.7.3.2, 4.7.3.3 | H

12–3 Name the type of polymerisation that happens when glycine reacts with itself. Describe this type of polymerisation. *[4]*

p192& 193 | 4.7.3.3 | H

12–4 Identify the two products of this reaction. *[2]*

Total: 11

13 DNA is a natural polymer found in the nuclei of cells.

Figure 15

p193-4 | 4.7.3.4 | MS5b WS1.2 QWC

13–1 Use **Figure 15** and your own knowledge to describe the shape, structure and bonding in a molecule of DNA. *[6]*

Total: 6

14 Synthetic polymers can be made by addition polymerisation or condensation polymerisation.

p178& 191 | 4.7.3.1, 4.7.3.2 | QWC | H

14–1 Compare the two types of polymerisation. You should include examples of each type in your answer. *[6]*

Total: 6

15 Poly(ethene) is a widely manufactured polymer. The pie charts in **Figure 16** show the uses of low density polyethene (LDPE) and high density polyethene (HDPE).

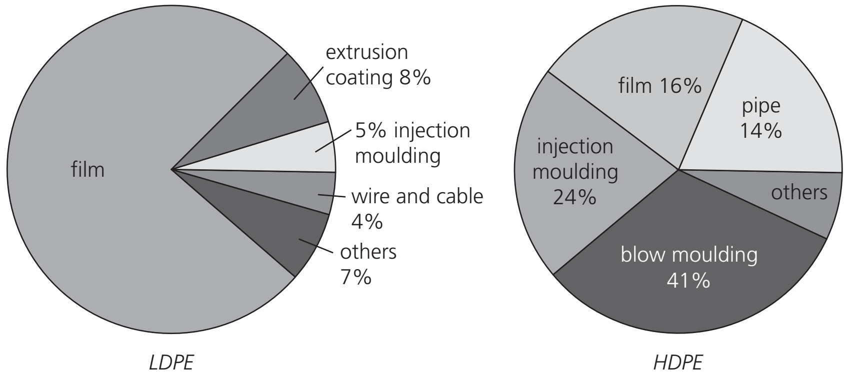


Figure 16

p250 | 4.7.3.1, 4.10.3.3 | MS1c,4a,2c | WS1.2,3.2

15–1 Calculate the percentage use of film for LDPE. *[2]*

p250 | 4.7.3.1, 4.10.3.3 | MS1c,4a,2c | WS1.2,3.2

15–2 Calculate the amount of other uses for HDPE. Give your answer as a fraction. *[1]*

p250 | 4.7.3.1, 4.10.3.3 | MS1c,4a,2c | WS1.2

15–3 There was 18.7 million tonnes of LDPE made in the world in 2014. Calculate the mass used in injection moulding. *[3]*

Total: 6

Organic chemistry topic review

1 Crude oil is an important raw material. It is made mainly of hydrocarbons.

p171 | 4.7.1.1

1–1 What does 'hydrocarbon' mean? Choose **one** answer. *[1]*

- contains carbon
- contains hydrogen
- contains carbon and hydrogen
- contains only carbon and hydrogen

p171 | 4.7.1.1

1–2 Describe how crude oil is made. *[2]*

p173 | 4.7.1.1

1–3 Describe how crude oil is separated. *[3]*

p175 | 4.7.1.3

1–4 Methane is collected in the refinery gas fraction and decane ($C_{10}H_{22}$) can be found in kerosene. Compare the properties of these two alkanes. *[4]*

p176, 283&284 | 4.7.1.4 | MS1c

1–5 Long chain hydrocarbons like decane are often cracked. Complete the equation for the cracking of decane to make ethene and one other alkane product. *[2]*

$C_{10}H_{22} \rightarrow$ ....

Total: 12

2 Ethane (C_2H_6) is a hydrocarbon that can be extracted from crude oil.

p173 | 4.7.1.2

2–1 Name the process used to obtain ethane from crude oil. [1]

p172 | 4.7.1.1 | MS5b

2–2 Draw the displayed formula of an ethane molecule. [2]

p172 | 4.7.1.1 | MS1c

2–3 What is the general formula of an alkane? Choose **one** answer. [1]

- C_nH_{2n}
- C_nH_{2n+1}
- C_nH_{2n+2}
- $C_{2n}H_{2n+2}$

p64&65 | 4.3.1.2 | MS3a,3c

2–4 Calculate the relative formula mass for one molecule of ethane.

Relative atomic mass (A_r): C = 12, H = 1 [2]

p174, 283&284 | 4.7.1.3 | MS1c

2–5 Ethane can be used as a fuel in camping stoves. Write a balanced symbol equation for the complete combustion of ethane. [3]

Total: 9

3 Propene (C_3H_6) is an alkene and can be made by cracking undecane.

p176, 283&284 | 4.7.1.4 | MS1c

3–1 Write a balanced symbol equation for the cracking of undecane ($C_{11}H_{24}$) to make three molecules of propene and one other product. [2]

p177 | 4.7.2.1

3–2 Explain why propene is unsaturated. [2]

p174& 180 | 4.7.1.3, 4.7.2.2

3–3 Compare the combustion of propene and propane in air. [5]

p178–9 | 4.7.1.4 | WS2.2, 2.3

3–4 A student was given two unlabelled colourless liquids. Describe a chemical test to determine which liquid was an alkane and which was an alkene. Include the observations that the student would make. [4]

Total: 13

4 A student is making ethanol by fermentation using the equipment shown in **Figure 17**.

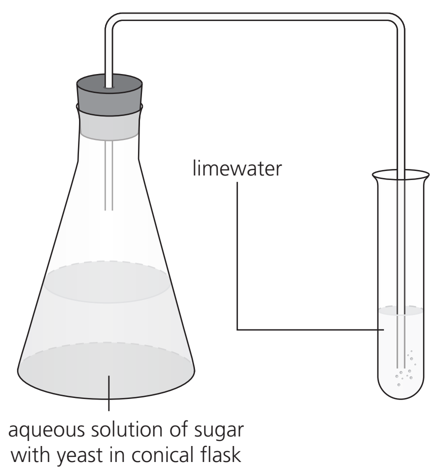


Figure 17

p181 | 4.7.2.3

4–1 Describe the process of fermentation to produce ethanol in terms of the reactions that occur. *[3]*

p181 | 4.7.2.3 | QWC

4–2 Describe any observations that the student makes. Explain your answer. *[6]*

p24 | 4.1.1.2

4–3 Suggest how the ethanol can be purified and collected. *[1]*

p182 | 4.7.2.4 | MS5b

4–4 Ethanol can oxidise in air to make ethanoic acid. Draw the displayed formula of this product. Name the homologous series it belongs to. *[2]*

p184 | 4.7.2.4

4–5 Ethanol can react with the product in question 4–4. Name the sweet-smelling substance that is made. *[1]*

Total: 13

5 Polymers are an important material in everyday life.

p187 | 4.7.3.1

5–1 Describe how addition polymers are made. *[2]*

p189 | 4.7.3.2

5–2 Describe how condensation polymers are made. *[3]*

p187, 189,192 &193 | 4.7.3.1, 4.7.3.2, 4.7.3.3

5–3 Polypeptides are an important class of natural polymers. Identify whether a polypeptide is an addition or a condensation polymer. Justify your answer. *[3]*

Total: 8

8 Chemical analysis

Purity, formulations and chromatography

Quick questions

p202 4.8.1.1 1 In chemistry, what does 'pure' mean?

p203 4.8.1.1 2 How can the melting and boiling points of a substance be used to show it is pure?

p204 4.8.1.2 3 What is a 'formulation'?

p204 4.8.1.2 4 How are formulations made?

p204 4.8.1.2 5 Give an example of a formulation.

p205 4.8.1.3 6 What is 'chromatography'?

p206 4.8.1.3 7 What is an 'R_f value'?

p206 4.8.1.3 8 Give the expression to calculate R_f values.

p205/6 4.8.1.3 9 What are the two phases used in chromatography?

p206 4.8.1.3 10 How does chromatography separate substances?

Exam-style questions

p202 4.8.1.1 **11–1** Most people use the term 'pure milk' in everyday speech. However, a scientist uses the word 'pure' in a different way and would not consider milk to be a pure substance. Explain why. *[4]*

MS1c,3a,3b **11–2** The composition of whole milk is 4.9% lactose (carbohydrate), 3.4% fat, 3.3% protein, 0.7% minerals, and water. Calculate the percentage amount of water in whole milk. *[2]*

p203 4.8.1.1 WS1.2 **11–3** The boiling point of water is exactly 100 °C. Predict the boiling point of whole milk and justify your prediction. *[4]*

Total: 10

p213 4.8.1.3 RP6 WS3.5

12 Chlorophyll was obtained by stirring a shredded green leaf in a suitable solvent. A fine line of the solution was drawn 2.0 cm from the bottom of a piece of chromatography paper and allowed to dry. The paper was placed in another solvent (mobile phase), making sure that the dried line of pigment was above the surface of the solvent. **Figure 1** shows the results of the chromatography.

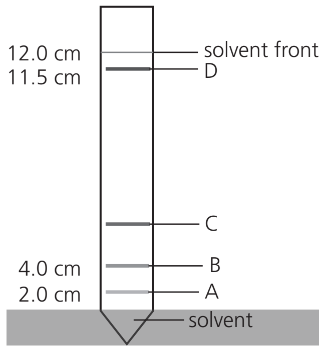


Figure 1

12–1 State the number of pigments in the sample of chlorophyll. *[1]*

12–2 State the pigment that was insoluble in the mobile phase. *[1]*

12–3 State the pigment was most attracted to the mobile phase. *[1]*

12–4 Calculate the R_f value of pigment D. *[3]*

12–5 State the pigment that had a R_f value of 0.33. *[3]*

Total: 9

13 Steel is an alloy made from iron with carbon added to it. **Figure 2** shows how the strength of steel is affected by the amount of carbon added.

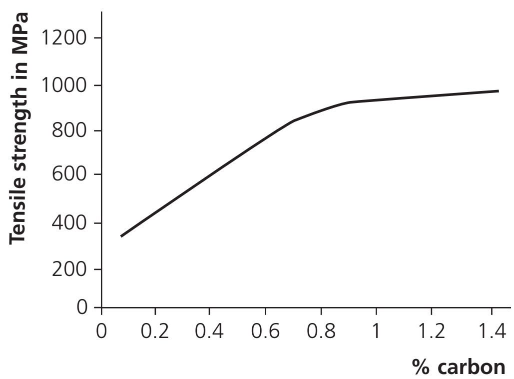


Figure 2

p204 4.8.1.2

13–1 Explain why steel can be considered to be a formulation. *[2]*

WS3.5

13–2 Describe what the graph in **Figure 2** shows. *[2]*

MS4a,1b WS3.2,4.3,4.4

13–3 State the tensile strength of steel with 0.6% carbon. Give your answer in Pa using standard form. *[2]*

Total: 6

14 Aspirin ($C_9H_8O_4$) can be made into a tablet and is a medicine that can help to relieve pain. A pure sample of aspirin has a melting point of 136 °C.

p203 4.8.1.1

14–1 Explain why the melting point of an aspirin tablet from a blister pack will not be exactly 136 °C. *[2]*

pXX 4.3.1.2 MS3c

14–2 Calculate the relative formula mass (M_r) of aspirin, given that the relative atomic masses (A_r) of H = 1, C = 12 and O =16. *[2]*

p204 4.8.1.2

14–3 A tablet of a painkilling medication contains aspirin (to relieve pain), chalk (to bulk out the tablet) and artificial sweeteners (to make the tablet taste better). Classify this substance. *[1]*

Total: 5

Identification of common gases

Quick questions

p207 4.8.2.1 AT8

1 What gas causes a pop sound when tested with a lighted splint?

p207 4.8.2.2 AT8

2 What gas causes a glowing splint to relight?

p207 4.8.2.3 AT8

3 What is the aqueous solution used to test for carbon dioxide?

p207 4.8.2.4 AT8

4 What happens to damp blue litmus paper when it is put into chlorine gas?

Exam-style questions

5 Calcium carbonate can undergo thermal decomposition to make calcium oxide and a gas.

p130 4.5.1.1

5–1 State if this reaction is exothermic or endothermic. *[1]*

p 42&69 4.2.1.4, 4.3.1.3

5–2 State the formula of the gas produced. *[1]*

p207 4.8.2.3 AT5,8

5–3 The gas can be tested with limewater. Describe the observation you would expect. *[2]*

p284 4.2.2.2, 4.8.2.3

5–4 Add the state symbols to the balanced symbol equation. *[4]*

CO_2... + $Ca(OH)_2$... → $CaCO_3$...+ H_2O...

Total: 8

6 Hydrogen peroxide (H_2O_2) can decompose to make water and oxygen gas.

p283 4.1.1.1

6–1 Write a balanced symbol equation for this reaction. *[3]*

p207 4.8.2.2

6–2 Outline the chemical test to show that oxygen has been made. *[2]*

p42 4.2.1.4

6–3 Oxygen forms a diatomic molecule. Draw the dot and cross diagram of oxygen. *[2]*

Total: 7

Identification of ions by chemical and spectroscopic means

Quick questions

p209 4.8.3.1 **1** What is the name of the ion that causes a crimson flame?

p209 4.8.3.1 **2** What is the formula of the ion that causes a yellow flame?

p209 4.8.3.1 **3** What is the colour of the flame caused by a potassium ion?

p209 4.8.3.1 **4** What is the colour of the flame caused by Ca^{2+}?

p209 4.8.3.1 **5** Which element would be present if a flame test showed a green colour?

p209 4.8.3.2 **6** What colour precipitate is made when sodium hydroxide is added to a solution containing calcium ions?

p209 4.8.3.2 **7** What ion causes a blue precipitate with sodium hydroxide?

p210 4.8.3.3 **8** What gas is made when carbonates react with acids?

p211 4.8.3.4 **9** Name the halide that causes a yellow precipitate with silver nitrate.

p211 4.8.3.4 **10** Give the formula of the halide that causes a cream precipitate with silver nitrate.

p211 4.8.3.5 **11** What is the name of the reagent used to show the presence of sulfate ions?

p214 4.8.3.6 **12** What are the advantages of using instrumental methods of analysis?

p214 4.8.3.7 **13** What is flame emission spectroscopy used for?

Exam-style questions

p209 4.8.3.2 **14** When sodium hydroxide is added to some solutions a coloured precipitate is formed. The colour indicates the cation present.

14–1 Define the term 'precipitate'. *[1]*

14–2 Balance the equation to show copper(II) sulfate reacting with sodium hydroxide. *[1]*

.....$CuSO_4$ + ...NaOH → ....$Cu(OH)_2$ + ...Na_2SO_4

14–3 Give the name of the blue precipitate formed in this reaction. *[1]*

WS2.2 AT8 QWC **14–4** Outline an experiment to determine whether a solution contained iron(II) or iron(III) ions. *[6]*

Total: 9

15 A solution of calcium chloride was analysed using different chemical analysis.

p208 4.8.3.1 **15–1** State the cation in calcium chloride solution. *[1]*

p208 4.8.3.2 **15–2** State the anion in calcium chloride solution. *[1]*

p209 4.8.3.1 AT8 **15–3** State the colour of the flame in a calcium chloride flame test. *[1]*

p210 4.8.3.2 AT8 **15–4** Describe the observations when calcium chloride solution was mixed with sodium hydroxide solution. *[2]*

p211 4.8.3.4 AT8 **15–5** Describe the observations when calcium chloride solution was mixed with acidified silver nitrate. *[2]*

Total: 7

16 A solution of magnesium sulfate ($MgSO_4$) was analysed using different chemical analysis.

p208 4.8.3.1 **16–1** Give the formula of the cation in magnesium sulfate solution. *[1]*

p208 4.8.3.2 **16–2** Give the formula of the anion in magnesium sulfate solution. *[1]*

p211 4.8.3.2 **16–3** Magnesium sulfate can react with sodium hydroxide to make a white precipitate. Add the state symbols to the balanced symbol equation. *[4]*

$MgSO_4$ + ... + 2NaOH..... → $Mg(OH)_2$ ..... + Na_2SO_4 .....

p210 4.8.3.5 WS2.2 AT8 QWC **16–4** Outline a chemical test to show that sulfate ions are present. *[6]*

Total: 12

17 A sample of a colourless liquid thought to be a solution of potassium carbonate (K_2CO_3) was found in a chemical store.

p209& 214 4.8.3.1, 4.8.3.7 **17–1** Justify the use of flame emission spectroscopy rather than a flame test to determine the cation present. *[4]*

p210 4.8.3.3 WS2.2 AT8 QWC **17–2** Outline a chemical test that can be used to show the identity of the anion. *[6]*

p211 4.8.3.4 AT8 **17–3** Suggest and explain the observations made if this solution was mixed with acidified silver nitrate solution. *[4]*

Total: 14

18 Flame emission spectroscopy can be used to analyse the composition of solutions. **Figure 3** shows the emission spectrums for lithium ions and sodium ions.

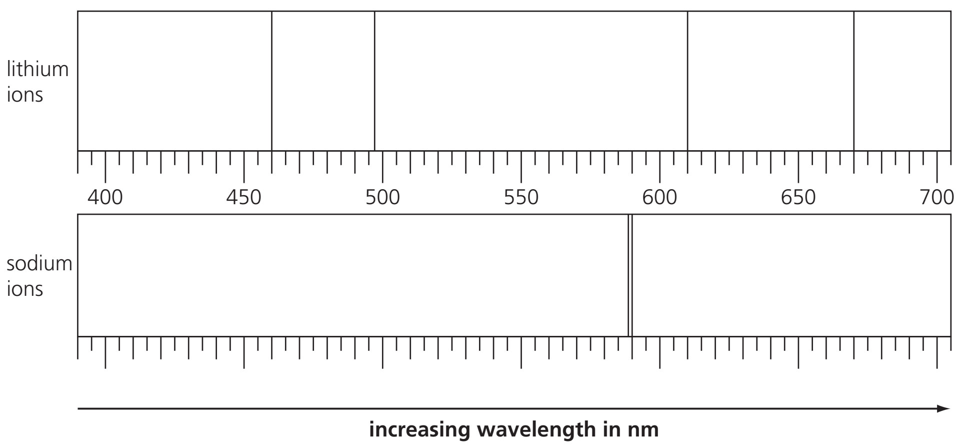


Figure 3

p62,95, 96&214 4.8.3.7 MS1a, 1b,2a WS4.3,4.4, 4.5,4.6 **18–1** Calculate the approximate wavelength of the two emission lines for the sodium ion. Give your answer in metres, to two significant figures and in standard form. *[3]*

p62,95, 96&214 4.8.3.7 **18–2** Sketch the emissions spectrum for a solution that contains a mixture of sodium iodide and lithium iodide. *[2]*

p210& 211 4.8.3.4 **18–3** When acidified silver nitrate is added to the solution, a yellow precipitate is formed. Explain why hydrochloric acid cannot be used to acidify the solution. *[2]*

Total: 7

19 A selection of chemicals has had their labels fall off in the chemical store. It is known that the solutions are: sodium carbonate, sodium chloride, magnesium chloride and aluminium iodide. Use the information below to determine the name of each solution.

p214, 215&211 | 4.8.3.4, 4.8.3.7 | AT8

19–1 Solution A has two characteristic lines on the flame emission spectrum at about 590 nm and forms a white precipitate on the addition of acidified silver nitrate solution. *[1]*

p209& 211 | 4.8.3.1, 4.8.3.7 | AT8

19–2 Solution B has a yellow flame test and fizzes on the addition of an acid. *[1]*

p209,210 &211 | 4.8.3.2, 4.8.3.7 | AT8

19–3 Solutions C and D produce a white precipitate with addition of sodium hydroxide. On addition of acidified silver nitrate solution both solutions make a precipitate. Solution C makes a yellow precipitate. *[2]*

Total: 4

Chemical analysis topic review

1 Gases are often made in a chemical reaction. Scientists may use indicator tests to determine the gas present.

p207 | 4.8.2

1–1 Match the name of the gas with its test. *[4]*

Gas	Test
Oxygen	Limewater turns from colourless to cloudy.
Chlorine	A lighted splint causes a squeaky pop sound.
Carbon dioxide	Glowing splint is re-lighted.
Hydrogen	Damp blue litmus paper turns red then bleaches white.

Total: 4

2 **Table 1** shows the data that a student collected on three different substances, A, B and C.

Substance	Melting point in °C	Boiling point in °C
A	–34	356
B	420	913
C	1425–1540	2530–2545

Table 1

p203–4 | 4.8.1.1 | AT8

2–1 Give the state of substance A at 0 °C. *[1]*

p203–4 | 4.8.1.1

2–2 Use **Table 1** to state and explain which substance(s) are pure. *[2]*

p203& 204 | 4.8.1.2

2–3 Justify why the data in **Table 1** is not enough to determine whether substance C is a formulation. *[3]*

p209 | 4.2.1.3, 4.8.3.1

2–4 A sample of substance C was put into a blue Bunsen flame and the flame turned crimson red. Give the formula of the ion that caused this. *[2]*

Total: 9

3 An unknown blue chemical was analysed by a student. When acid was added to the sample, bubbles were formed.

p207 | 4.8.2.3 | WS2.2

3–1 Describe an experiment to collect the gas, and then determine that the gas is carbon dioxide. *[4]*

p209, 285–8 | 4.8.3.2

3–2 The unknown blue chemical was added to sodium hydroxide solution. It made a blue precipitate. Write an ionic equation for this reaction. *[3]*

p209 | 4.8.3.1 | WS2.2,2.3 | QWC

3–3 Use the information in question 3–2 and your own knowledge to outline how to complete a flame test on the unknown chemical and the likely result. *[6]*

4.8.3 | WS3.5 | AT8

3–4 Use the information from questions 3–1, 3–2 and 3–3 to give the formula of the unknown compound. *[1]*

Total: 14

9 Chemistry of the atmosphere

The composition and evolution of the Earth's atmosphere

Quick questions

p224 4.9.1.1 MS1c
1 What is the fraction of nitrogen in the atmosphere today?

p224 4.9.1.1
2 Approximately how long has the proportion of gases in the atmosphere been the same as today?

p224 4.9.1.2
3 Approximately how old is the Earth?

p225–6 4.9.1.3
4 What made oxygen in the Earth's atmosphere?

Exam-style questions

p224 4.9.1.1
5 **Figure 1** shows the composition of the Earth's atmosphere.

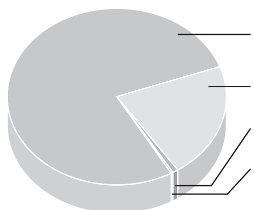


Figure 1

WS3.5
5–1 State the name of the gas that makes up most of the Earth's atmosphere. *[1]*

5–2 Give the molecular formula of the gas that makes up approximately ⅕ of the Earth's atmosphere. *[1]*

5–3 Suggest a gas that is found in the 0.1% of gases in our atmosphere. *[1]*

Total: 3

p224 4.9.1.2
6 The evolution of the Earth's atmosphere is not known. Evidence is collected from the Earth, as well as observing planets in our solar system to suggest how our own atmosphere may have been created. One theory is that the early atmosphere was like Venus or Mars today.

WS1.1, 1.2,3.6
6–1 Use your knowledge of this theory to explain how the concentration of carbon dioxide has naturally changed over the last 4.6 billion years. *[4]*

WS1.1,1.3
6–2 There is more than one theory about how the Earth's atmosphere formed. Explain why it is difficult to know exactly how the atmosphere developed. *[2]*

Total: 6

p224–6 4.9.1.2

7 Evidence collected from rocks and analysis of air bubbles trapped in Arctic ice are used by scientists to support their theory that the Earth's atmosphere has changed over time.

7–1 State the main gas in Earth's early atmosphere. *[1]*

7–2 Describe the theory of how nitrogen became part of Earth's atmosphere. *[2]*

7–3 Explain the role of the oceans in reducing the amount of carbon dioxide in Earth's atmosphere. *[4]*

Total: 7

p225 4.9.1.3

8 Oxygen is essential for life. Oxygen is released into the Earth's atmosphere by a process called photosynthesis.

WS4.1

8–1 Write a word equation for photosynthesis. *[2]*

8–2 Cyanobacteria use photosynthesis to make glucose to live. State two other organisms that perform photosynthesis. *[2]*

8–3 Air today is mainly made of nitrogen and oxygen. Give the ratio of these gases in the modern atmosphere. *[1]*

Total: 5

9 The Earth's atmosphere has evolved over about 4.6 billion years.

p225& 283 4.9.1.3 MS1c

9–1 Copy and balance the symbol equation to show how oxygen was formed in Earth's atmosphere. *[1]*

$$\ldots\ldots\ldots\ldots CO_2 + \ldots\ldots\ldots\ldots H_2O \rightarrow \ldots\ldots\ldots\ldots C_6H_{12}O_6 + \ldots\ldots\ldots\ldots O_2$$

p224, 225&284 4.9.1.2, 4.2.2.2

9–2 Add the state symbols to your balanced symbol equation from question 9–1 to show how the oceans were formed. *[1]*

p284, 109&110 4.9.1.2, 4.2.2.2

9–3 Carbon dioxide dissolved in the oceans to make carbonic acid (H_2CO_3). Write a balanced symbol equation for this reaction. Include state symbols. *[2]*

p284, 109&110 4.9.1.2, 4.2.2.2, 4.4.2.6

9–4 Carbonic acid is a weak acid. Write a balanced symbol equation to show how carbonic acid is a weak acid. *[3]*

p108& 110 4.4.2.6

9–5 Suggest what the effect is on the pH of the oceans as the carbon dioxide dissolved, and explain why this happens. *[2]*

Total: 9

Carbon dioxide and methane as greenhouse gases

Quick questions

p227 4.9.2.1

1 What do greenhouse gases do in the atmosphere?

p227 4.9.2.1

2 List **three** greenhouse gases found naturally in Earth's atmosphere.

p228& 229 4.9.2.2

3 Which **two** greenhouse gases are humans adding to the Earth's atmosphere?

p230 4.9.2.3 WS1.2

4 What is the major cause of global climate change?

p233 4.9.2.4 WS1.2

5 What is a 'carbon footprint'?

Exam-style questions

6 Greenhouse gases like carbon dioxide are found naturally in the Earth's atmosphere. These gases are important as they maintain the Earth's temperature high enough to support life. In a sample of dry air, approximately 0.0035 % of the gas is carbon dioxide.

p61&62 MS1b WS3.2 **6–1** Write the amount of carbon dioxide in dry air as a percentage in standard form. *[1]*

p42 4.2.1.4 MS5b **6–2** Draw the dot and cross diagram of carbon dioxide. *[3]*

p227& 228 4.9.2.1 WS1.2 **6–3** Describe the 'greenhouse effect'. *[4]*

p228& 229 4.9.2.2 QWC **6–4** State the general trend in the amount of carbon dioxide in the Earth's atmosphere over the last 100 years, and explain what has caused this trend. *[6]*

p234 4.9.2.4 **6–5** Describe **two** actions that can be taken to reduce the amount of carbon dioxide in our atmosphere. Explain your answer. *[4]*

Total: 18

7 Global warming is the rise in Earth's average surface temperature. Many scientists believe that this is the main cause of global climate change.

p231–3 4.9.2.3 **7–1** List **three** predicted effects of global climate change. *[3]*

p229 4.9.2.2 QWC **7–2** Explain how animal farming has contributed to global climate change. *[6]*

p244 4.9.2.2 WS1.3, 1.5,1.6 **7–3** Describe and explain the importance of peer-reviewed research into global climate change. *[5]*

Total: 14

p230 **8** Scientists monitor the global mean surface temperature change. A graph of these data from 1880 until the present day is shown below in **Figure 2**. This shows the change in global surface temperature relative to the average temperature over the period 1951–1980.

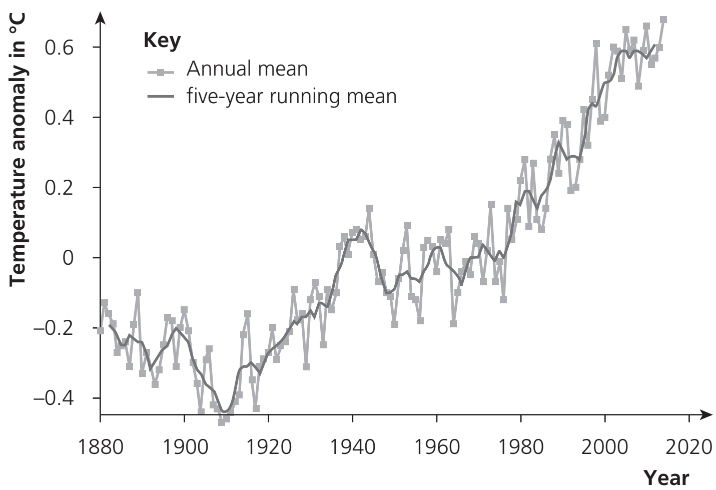


Figure 2

WS2.2 **8–1** State the units of the dependent variable. *[1]*

4.9.2.3 WS1.1,1.2, 3.2 **8–2** Describe the general trend shown in this graph. *[3]*

4.9.2.3 WS1.2,3.2 **8–3** State which was the hottest complete decade in these records. *[1]*

4.9.2.2, 4.9.2.3 WS1.4 **8–4** Suggest a cause for this change in global mean surface temperatures. *[3]*

4.9.2.2 WS1.4,1.5 **8–5** Suggest why a few scientists do not think climate change is due to human activity. *[2]*

Total: 10

Common atmospheric pollutants and their sources

Quick questions

p237 4.9.3.1 **1** What is a major source of atmospheric pollutants?

p237 4.9.3.1 **2** Which **three** elements are in most fuels?

p236 4.9.3.1 **3** What are particulates made from?

p238 4.9.3.2 **4** Which gas is colourless, odourless, toxic and results from the combustion of fuel?

p238 4.9.3.2 **5** What problems are caused by both sulfur dioxide and oxides of nitrogen?

p238 4.9.3.2 **6** What problems are caused by particulates?

Exam-style questions

7 Natural gas used to heat homes in the UK is mainly methane (CH_4).

p42&172 4.2.1.4, 4.7.1.1 **7–1** State the names of the elements in a molecule of methane. *[2]*

p41 4.2.1.4, 4.7.1.1 MS1c **7–2** State the number of atoms in one molecule of methane. *[1]*

p174 4.7.1.3 **7–3** Copy and complete the word equation for methane when completely combusted. *[2]*

methane + oxygen → .... + ...

p238 4.9.3.2 **7–4** If there is a limited amount of oxygen when methane is combusted then different products form. State **one** pollutant that can be made from the incomplete combustion of methane and describe its effect on human health. *[2]*

Total: 7

8 Many fuels are hydrocarbons that are made from the fractional distillation of crude oil.

p283 4.7.1.3 MS1c **8–1** Copy and balance the symbol equation for the combustion of propane fuel. *[1]*

$$\ldots\ldots C_3H_8 + \ldots\ldots O_2 \rightarrow \ldots\ldots CO_2 + \ldots\ldots H_2O$$

p238 4.9.3.1 WS1.4 **8–2** Describe how carbon particulates can be produced by the combustion of fossil fuels and the effect on the environment. *[4]*

p238 4.9.3.2 WS1.4 **8–3** Explain how carbon monoxide can be produced by combustion of propane and the effect on the human body. *[4]*

p238 4.9.3.1 WS1.4 QWC **8–4** Burning of fuels is a major source of atmospheric pollutants. Explain how combustion of fossil fuels can lead to the formation of acid rain. *[6]*

Total: 15

Chemistry of the atmosphere topic review

1 Theories about how the atmosphere was formed have changed and developed over time.

p224–7 | 4.9.1.2, 4.9.1.3, 4.9.1.4

1–1 Put the statements below in order so that they describe the main stages of atmospheric formation in one theory. *[3]*

A	Carbon dioxide levels reduced as plants used the gas, oceans dissolved the gas and carbonate rocks were formed.
B	Nitrogen from volcanic activity built up.
C	Volcanic activity made an atmosphere of mainly carbon dioxide.
D	As the Earth cooled water vapour condensed to form the oceans.
E	Oxygen was made by algae and early plants.

p225 | 4.9.1.3

1–2 State the name of the process that made oxygen in the Earth's atmosphere. *[1]*

p227 | 4.9.2.1

1–3 Carbon dioxide is a greenhouse gas. Define the term 'greenhouse gas'. *[2]*

p229 | 4.9.2.2

1–4 Give **one** way in which humans have increased the amount of carbon dioxide gas in our atmosphere. *[1]*

Total: 7

2 Coal, oil and natural gas are all examples of fossil fuels. These resources are widely used in industry as well as our homes for energy, heating and electricity production.

p238 | 4.9.3.1

2–1 Describe how acid rain can form due to combustion of coal. *[3]*

p238 | 4.9.3.1, 4.9.3.2 | QWC

2–2 Explain how carbon monoxide can be formed from the combustion of fossil fuels. Describe the health problems that it can cause. *[6]*

p225–30 | 4.9.1.3, 4.9.2.1 | WS1.2 | QWC

2–3 Justify the importance of having some carbon dioxide in our atmosphere. *[6]*

p233–4 | 4.9.2.4 | WS1.3

2–4 Explain carbon neutral fuels could reduce carbon footprints. *[2]*

Total: 17

3 Peer-reviewed scientific research indicates that humans are adding more greenhouse gases to the atmosphere and this is likely to be the main cause of climate change.

p230 | 4.9.2.2

3–1 Explain the importance of peer-reviewed scientific research. *[2]*

p230 | 4.9.2.2

3–2 Suggest why almost all scientists now agree that humans are contributing to global warming, but some people do not think that global warming is due to human activity and that climate change will continue. Suggest why this is. *[2]*

p231–2 | 4.9.2.3 | WS1.5 | QWC

3–3 Describe the potential effects of global climate change. *[6]*

p234 | 4.9.2.4 | WS1.3

3–4 Suggest ways to reduce global climate change. *[3]*

Total: 13

10 Using resources

Using the Earth's resources and sustainable development

Quick questions

p248 4.10.1.1 — 1 What do humans use Earth's resources for?

p248 4.10.1.1 — 2 What are the Earth's finite resources processed to provide?

p248 4.10.1.1 — 3 What is 'sustainable development'?

p252 4.10.1.2 — 4 What is 'potable water'?

p253 4.10.1.2 — 5 How can salty water be made into potable water?

p255 4.10.1.3 — 6 What should be removed from industrial waste water before it is released into the environment?

p261 4.10.1.4 — H 7 What is 'bioleaching'?

Exam-style questions

p248 4.10.1.1 — 8 Natural resources are found all around the world. They are used by humans to provide warmth, shelter, food and transport.

8–1 State **one** finite resource that is used as a fuel. *[1]*

8–2 State **one** renewable resource used as a fuel. *[1]*

8–3 Polyester is a synthetic fibre that can be used to make clothes. Give **one** natural product that polyester could replace. *[1]*

8–4 Aluminium metal is used to make drinks cans. Once they are used, they can be thrown away into landfill. Suggest how this metal can be used in a more sustainable way. *[1]*

8–5 Describe, with at least **one** example, the importance of chemistry in sustainable development. *[4]*

Total: 8

p248 4.10.1.1 WS3.2

9 **Figure 1** shows a pie chart that shows the sources of global energy.

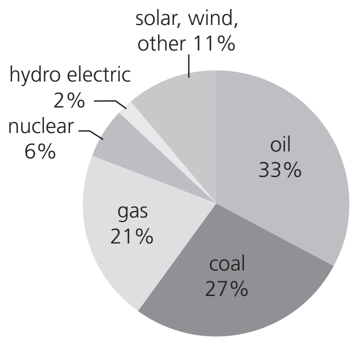


Figure 1

MS2c,4a
9–1 Which renewable energy source contributes the least amount of global energy? *[1]*

MS2c,4a
9–2 Calculate the percentage of energy resources that are finite. *[2]*

MS2c, 4a,2h
9–3 Calculate how many times greater the contribution of nuclear power to global energy is compared to hydroelectric power. *[1]*

MS2c, 4a,2h
9–4 Give the ratio of global energy for gas compared to nuclear. *[1]*

MS2c,4a,2h QWC
9–5 Use **Figure 1** and your own knowledge to evaluate whether global energy sources are sustainable. *[6]*

Total: 11

p252 4.10.1.2

10 Safe, reliable drinking water is essential for humans and often water will need to be processed to make it safe to drink.

10–1 Describe the properties of drinking water. *[2]*

QWC
10–2 Pure water is also potable water. But not all sources of potable water are pure. Explain the difference between potable water and pure water. *[6]*

QWC
10–3 Describe and explain how potable water is produced in the UK. *[6]*

Total: 14

11 **Table 1** shows the results of a student's investigation of different water samples. Scientists sometimes use Kelvin (K) as the unit of temperature. To convert Kelvin to Celsius (^{o}C) subtract 273.

Water	**pH**	**Boiling point in K**
Distilled water	7.0	373
Tap water	6.8	372–374
River water	7.4	371–378

Table 1

p108& 254 4.10.1.2
11–1 Describe **two** ways the student could test the pH of the water sample. *[2]*

p203& 252 4.10.1.2
11–2 Use the data in the table to justify the classification of each water sample as pure water or a mixture of water and other substances. *[4]*

p252 4.10.1.2
11–3 Outline a method to determine which water sample had the greatest mass of dissolved solids. *[4]*

Total: 10

12 In the UK there are approximately 11 billion litres (11 000 000 000 litres) of waste water produced daily. The waste water is processed by 9000 sewage treatment works. It must be treated before it is released into streams and rivers.

p254 4.10.1.3 — **12–1** State what is removed from sewage. *[2]*

p254 4.10.1.3 — **12–2** List the contaminants that may be found in industrial waste water. *[2]*

p255 4.10.1.3 QWC — **12–3** Describe how sewage is treated in the UK. *[6]*

MS1a,1b,2a — **12–4** Calculate, on average, how many litres of water are processed by each sewage treatment works. Give your answer in standard form and to two significant figures. *[3]*

Total: 13

13 **Figure 2** shows copper demand by source. There are approximately 18 million metric tonnes of copper produced each year.

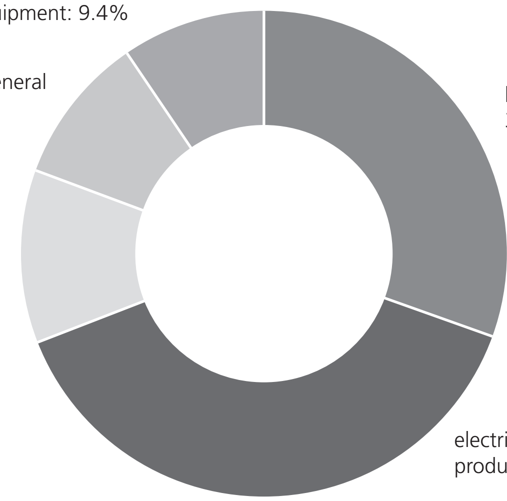


Figure 2

p45&46 4.2.1.5 WS3.2 MS4a QWC — **13–1** Use the information in **Figure 2** and your knowledge of structure and bonding to justify the main use of copper. *[6]*

WS3.2 MS1a,1c,2a — **13–2** Calculate the mass of copper used in transportation equipment each year. Give your answer to the nearest million metric tonnes. *[3]*

p260& 261 4.2.1.5 — H **13–3** Explain why we now extract copper from low grade copper ores. *[2]*

p260& 261 4.2.1.5 WS1.1, 1.4 — H **13–4** Justify the use of phytomining for obtaining copper. *[3]*

Total: 14

H **14** Copper ores are becoming more scarce. New technologies, such as bioleaching, allow lower grade ores to be used to obtain copper for electronics.

p260–1 4.2.1.5 — **14–1** Describe the process of bioleaching. *[3]*

p261& 283–4 4.2.1.5 — **14–2** Displacement can be used to extract copper metal from the leachate. Write a balanced symbol equation to show how scrap iron can be used to make copper from leachate containing copper(II) sulfate. *[2]*

p261, 289&290 | 4.10.1.4

14–3 Electrolysis can be used to extract the copper metal from the leachate. Write a half equation to show the production of copper. Include state symbols in your equation. *[3]*

p261& 290 | 4.10.1.4

14–4 Explain in terms of electrons, whether the copper in the leachate is being oxidised or reduced during these methods of extraction. *[2]*

Total: 10

Life cycle assessment and recycling

Quick questions

p250 | 4.10.2.1

1 What is the purpose of a life cycle assessment?

p249 | 4.10.2.2

2 How are metals recycled?

p249 | 4.10.2.2

3 How do we obtain raw materials from the Earth to make metals and building materials?

Exam-style questions

p250–1 | 4.10.1.1

4 Shopping bags can be made from several different materials including plastic, paper and cotton.

4–1 State the finite raw material used to make plastic shopping bags. *[1]*

4–2 Give the raw material used to make paper shopping bags. *[1]*

QWC

4–3 Explain how shopping bags can be made more sustainably. *[6]*

Total: 8

5 **Figure 3** shows what happens to plastic shopping bags after they have been purchased.

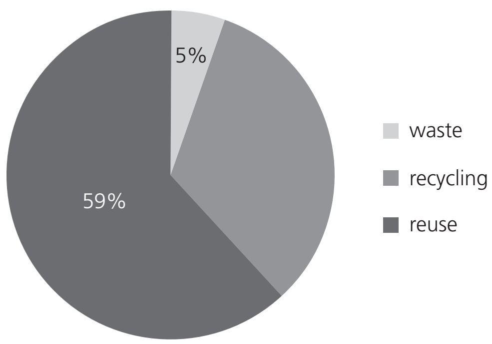


Figure 3

MS1c,1d, 4a

5–1 Calculate the percentage of shopping bags that are recycled. *[2]*

MS1a,c, 1d,2a,4a | WS4.6

5–2 Calculate how many times more likely you are to reuse than put into landfill the shopping bag. Give your answer to two significant figures. *[3]*

p250& 251 | 4.10.2.1

5–3 Explain how putting a plastic shopping bag into landfill can cause environmental damage. *[3]*

p250 | 4.10.2.1 | WS1.3

5–4 Explain how life cycle assessments may or may not be biased. *[6]*

Total: 14

p249 4.10.2.2

6 Glass is a type of ceramic and can be used to make bottles.

6–1 Describe how glass bottles can be reused. *[3]*

6–2 Describe how glass can be recycled. *[3]*

6–3 State where most of the energy to make glass has come from. *[1]*

MS1c,1d WS4.3,4.4,4.5

6–4 The energy saved by producing a bottle from recycled glass rather than from raw materials is enough to power a computer for 25 minutes. A desk top computer uses 30 J/s. Calculate the energy saved. Give your answer in kJ. *[3]*

Total: 10

Using materials

Quick questions

p258 4.10.3.1 1 What is 'corrosion'?

p257 4.10.3.2 2 What **two** metals make up bronze?

p257 4.10.3.2 3 Which metal alloy has a low density?

p264 4.10.3.3 4 What is a 'thermosoftening polymer'?

p264 4.10.3.3 5 What is borosilicate glass made from?

Exam-style questions

6 Iron is an important metal in the construction industry. It is often used in the form of steel to reinforce structures and make them stronger.

p257 4.10.3.2 6–1 Define the term 'alloy'. *[1]*

p258 4.10.3.2 6–2 Compare the properties of high carbon and low carbon steels. *[2]*

p259 4.10.3.1 6–3 Steel, like iron, can corrode. Name the type of corrosion found on steel. *[1]*

p259 4.10.3.1 6–4 Write a word equation for rusting. *[3]*

Total: 7

p259 4.10.3.1 7 **Figure 4** shows three test tubes that were used to investigate rusting.

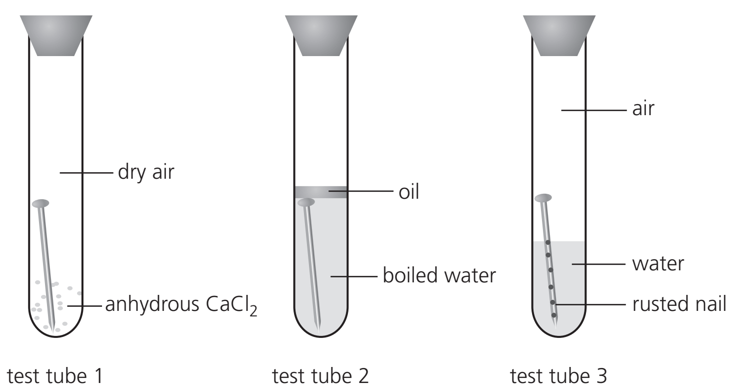


Figure 4

WS3.7 **7–1** Explain the purpose of adding anhydrous calcium chloride to test tube 1. *[2]*

7–2 Explain the purpose of oil in test tube 2. *[2]*

7–3 Explain why test tube 3 is the only one which shows evidence of rusting. *[2]*

WS3.7 **7–4** Suggest **three** ways that the results from this experiment could be made more reliable. *[3]*

WS3.7 **7–5** Suggest how this experiment on rusting could be improved to obtain quantitative results. *[3]*

Total: 12

8 **Table 2** shows the composition of gold used in jewellery.

Carat rating	Gold (%)	Other metals (%)
24	100	**A**
22	91.6	8.4
18	75	25
14	**B**	41.5

Table 2

p257 4.10.3.2 MS1c **8–1** Write down the missing values, **A** and **B**, from **Table 2**. *[2]*

p257&202 4.10.3.2 MS1c **8–2** Explain why 24 carat gold is defined as pure. *[2]*

p257 4.10.3.2 **8–3** List the names of the metals used to alloy with gold. *[3]*

p257 4.10.3.2 WS3.1,3.3 MS1a, 2a,2h **8–4** A ring was made of 2.5 g of metals. Calculate how much more gold would be in a 24 carat ring compared to a 18 carat gold ring. Give your answer to three significant figures. *[3]*

p46,51, 98&257 4.10.3.2 WS1.5,4.6 QWC **8–5** Alloys of gold can be used in dental fillings. Justify the use of gold in dental fillings. *[6]*

Total: 16

p264 4.10.3.3 **9** Glass is a type of ceramic material made of mainly silicon dioxide.

9–1 Name the main raw material used to make glass. *[1]*

9–2 Describe how soda-lime glass is made. *[3]*

9–3 Give **one** difference of borosilicate glass over soda-lime glass. *[1]*

WS1.2, 1.4,1.5 **9–4** A contact lens is a prescription medical device that can rest on the surface of the eyeball and improve a person’s sight. Glass or polymers can be used to make a contact lens. Evaluate the use of contact lenses and the material used to make them. *[4]*

WS1.2, 1.4,3.8 QWC **9–5** Modern contact lenses are a mixture of polymers, which include thermosoftening polycarbonates. Explain the difference between thermosoftening and thermosetting polymers. Include a labelled diagram in your answer. *[6]*

Total: 15

p265 4.10.3.3

10 Composite materials are made from two or more materials.

10–1 Give an example of a synthetic composite and describe its structure. [3]

10–2 Justify natural wood being classified as a composite. [2]

Total: 5

The Haber process and the use of NPK fertilisers

Quick questions

p266 4.10.4.1 1 What does the Haber process make?

p266 4.10.4.1 2 What are the raw materials for the Haber process?

p266 4.10.4.1 3 What is the catalyst in the Haber process?

p268 4.10.4.2 4 What elements are found in NPK fertilisers?

p268 4.10.4.2 5 What is phosphate rock treated with to make soluble salts?

Exam-style questions

p266 4.10.4.1 6 **Figure 5** shows a flow chart to describe the Haber process.

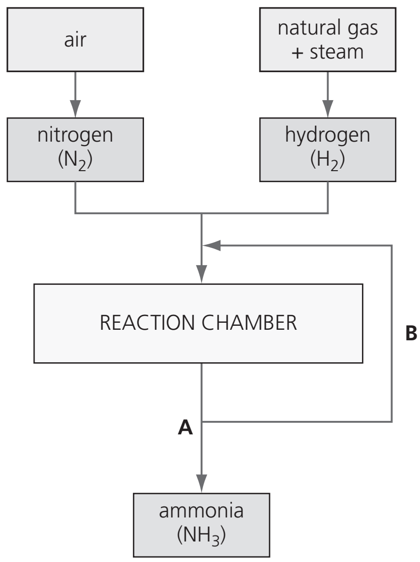


Figure 5

6–1 Describe the processes happening at point A in **Figure 5**. [3]

6–2 Explain in terms of economics the importance of the process marked B on the flow chart. [4]

6–3 Balance the symbol equation describing the Haber process. [1]

$N_2 + H_2 \rightleftharpoons NH_3$

6–4 State the conditions in the reaction chamber during the Haber process. [3]

Total: 11

p268 4.10.4.2

7 **Figure 6** shows a diagram of a bag of fertiliser. Fertilisers are important for farmers.

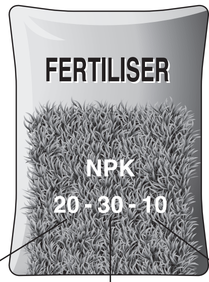


Figure 6

7–1 Name the three elements that are labelled on the fertiliser bag shown in **Figure 6**. *[3]*

MS1c 7–2 Calculate the percentage of the fertiliser **not** made of these three elements. *[1]*

7–3 Describe an NPK fertiliser. *[3]*

7–4 Copy and complete the word equation to show how ammonium nitrate can be made. *[2]*

.... + ... → ammonium nitrate

Total: 9

p268 4.10.4.2

8 Phosphate rock and two other mined resources are used as raw material used to make fertilisers.

8–1 Give the name of **one** other raw material used to make fertilisers. *[1]*

8–2 State the name of the desired product when phosphate rock is treated with sulfuric acid. *[1]*

QWC 8–3 Describe how phosphate rock is processed to make NPK fertilisers. *[4]*

Total: 6

9 **Figure 7** shows a titration set up to make a fertiliser.

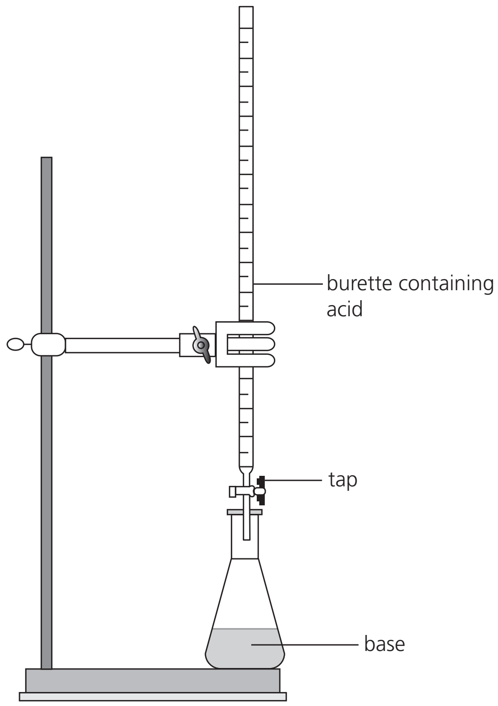


Figure 7

p88	4.10.4.2, 4.4.2.5	WS2.3 AT4	9–1	State the chemical that would be in the burette.	[1]
p88	4.10.4.2, 4.4.2.5	WS2.3 AT4	9–2	Suggest the chemical name of the fertiliser made in this process.	[1]
p88	4.10.4.2, 4.4.2.5	WS2.3,2.5 AT4	9–3	Outline how you would know that the reaction is complete.	[4]

Total: 6

10 The flow chart in **Figure 8** shows how fertilisers are made in industry.

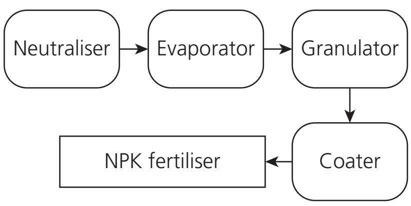


Figure 8

p268&269 | 4.10.4.2 | QWC

10–1 Use the flow chart and your own knowledge to explain the stages of making an NPK fertiliser in industry. [6]

Total: 6

11 The graph in **Figure 9** shows how the yield of ammonia changes in the Haber process under different temperatures and pressures.

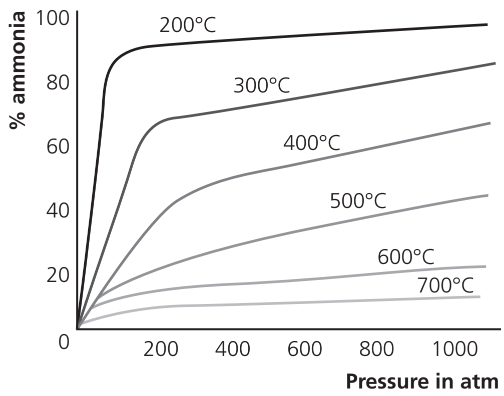


Figure 9

WS4.5,4.6 MS1a

11–1 101 kPa is the same as 1 atmosphere in pressure. Convert 200 atmospheres into pascal. Give your answer to three significant figures and in standard form. *[2]*

p266 4.10.4.1

11–2 Copy and complete the balanced symbol equation describing the Haber process by adding state symbols. *[1]*

$N_2 + 3H_2 \rightleftharpoons 2NH_3$

p267 4.10.4.1 WS3.5,3.8 MS1a QWC

11–3 Use **Figure 9** and your own knowledge to explain why pressures of greater than 200 atm are not used in the Haber process. *[6]*

p129 4.10.4.1 WS3.5,3.8 MS1a QWC

11–4 Use the graph to explain whether the Haber process is exothermic or endothermic. *[4]*

Total: 13

Using resources topic review

1 The Haber process is an important industrial reaction used to make ammonia (NH_3).

p42 4.2.1.4

1–1 State the number of atoms in one molecule of ammonia (NH_3). *[1]*

p266 4.10.4.1

1–2 Write a word equation for the formation of ammonia from its elements. *[2]*

p266 4.10.4.1

1–3 State where the nitrogen for the reaction comes from. *[1]*

p266 4.10.4.1

1–4 Describe where the hydrogen for the reaction comes from. *[2]*

Total: 6

2 Water is essential for human life. In the UK, rain water is treated, and this is the main source of potable water.

p252 4.10.1.2

2–1 What does potable water mean? Choose **one** answer. *[1]*

- pure water
- water that is safe to drink
- rain water
- mineral water

p252 4.10.1.2

2–2 In the UK water treatment process, state how microbes are removed from rain water. [1]

p254& 255 4.10.1.3

2–3 Explain why waste water must be treated before it can be potable. [3]

Total: 5

3 There are approximately 2.1 billion (2 100 000 000) shopping bags purchased in the UK. **Table 3** has some information about shopping bags.

	Plastic shopping bags	Paper shopping bags
Source of raw material	Crude oil	Trees
Manufacture of bags	High energy needed to crack the crude oil and make the plastic.	Uses a lot of water to make the paper.
Transportation of bags	Light weight and cheap.	Heavier and uses more fuel to transport.
Re-use	Potentially many times, as plastic is durable and waterproof.	Single or few uses as paper is not as durable and is not waterproof.
Disposal	Can be reused or recycled many times and can be incinerated and the energy used to make electricity. Does not biodegrade.	Can be recycled up to 7 times, can be incinerated and energy used to make electricity. Biodegrades.

Table 3

MS1a

3–1 The population of the UK is about 66 million (66 000 000). Calculate the number of shopping bags purchased per person. Give your answer to the nearest bag. [3]

p250–1 4.10.2.1

3–2 Life cycle assessment is a useful tool to evaluate the impact on the environment throughout its life. List the four stages of the life cycle assessment. [4]

p250–1 4.10.2.1, 4.10.2.2 WS1.2,1.4 QWC

3–3 Use the information in **Table 3**, the stages of life cycle assessment and your own knowledge to evaluate the use of plastic and paper shopping bags. [6]

p264 4.10.3.3

3–4 Justify the use of thermosoftening plastic shopping bags if recycling is to occur. [3]

Total: 16

4 The Haber process can be described by this balanced symbol equation:

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

The energy change for the forward reaction is −92 kJ/mol.

p134 MS1c

4–1 Calculate the energy change when 2 moles of ammonia are made. [1]

p131 4.5.1.1

4–2 State whether the forward reaction is exothermic or endothermic. Explain why. [2]

p267 4.10.4.1 WS3.5,3.8 QWC

4–3 Justify the use of a temperature of 450 °C for the Haber process. [6]

p76 4.3.2.2 MS1c

4–4 Calculate the mass in tonnes of nitrogen needed to make 17 tonnes of ammonia. 1 tonne = 1000 kg. [4]

Total: 13

Practice exam papers

Paper 1

1 Potassium chloride is a soluble salt that can be formed in a variety of chemical reactions.

1–1 Potassium chloride contains the ions K^+ and Cl^-. Give the formula of potassium chloride. *[1 mark]*

1–2 Give the name of **two** reactants that can be used to make a sample of potassium chloride in a neutralisation reaction. *[2 marks]*

1–3 Potassium has two common isotopes. Their masses numbers are 39 and 41. The percentage abundance of each isotope is 93.3 % of ^{39}K and 6.7 % ^{41}K. Calculate the relative atomic mass (A_r) of potassium. Give your answer to three significant figures. *[3 marks]*

1–4 Describe and explain the properties of potassium chloride. Answer in terms of structure and bonding. *[5 marks]*

Total: 11

2 Early forms of the periodic table arranged the elements in order of their atomic weights. Modern periodic tables arrange the elements in order of their atomic numbers.

2–1 Based on atomic weight, argon should be placed after potassium and not before it. Explain why Mendeleev placed these two elements in their correct positions in his periodic table. *[2 marks]*

2–2 Explain why lithium, sodium and potassium are in Group 1 of the periodic table. *[1 mark]*

Table 1 shows the observations made when three Group 1 metals react with water.

Group 1 metal	Observations
Lithium	The metal fizzes on the surface of the water.
Sodium	The metal melts, fizzes and moves quickly on the surface of the water.
Potassium	The metal melts, fizzes, and burns with a lilac flame on the surface of the water.

Table 1

2–3 Describe the trend in the reactivity of the Group 1 metals. Use **Table 1**. *[1 mark]*

2–4 Name the gas produced during the reaction between lithium and water. *[1 mark]*

2–5 Suggest a value for the pH of the solution formed during the reaction between potassium and water. Explain your answer. *[2 marks]*

2–6 During the reaction between potassium and water, potassium atoms react to form potassium ions, K^+. Explain why this is an example of oxidation. *[2 marks]*

2–7 Give the half equation for the oxidation of potassium atoms. *[1 mark]*

2–8 The halogens make up Group 7 of the periodic table. Describe the trend in the reactivity of the Group 7 elements. Explain your answer. *[5 marks]*

Total: 15

3 Hydrazine, N_2H_4, is a highly reactive compound. Hydrazine can be made in the following reaction.

$2NH_3 + NaClO \rightarrow N_2H_4 + NaCl + H_2O$

3–1 Calculate the percentage atom economy for this reaction to produce hydrazine.

Relative formula masses (M_r):

$NH_3 = 17$, $NaClO = 74.5$, $N_2H_4 = 32$, $NaCl = 58.5$, $H_2O = 18$

Give your answer to three significant figures. *[3 marks]*

3–2 When 1.0 kg of ammonia, NH_3, reacts the maximum theoretical mass of hydrazine is 0.94 kg. The percentage yield of the reaction is 71%.

Calculate the actual mass of hydrazine formed in the reaction. *[2 marks]*

3–3 The reaction between nitrogen and hydrogen to form hydrazine is:

$N_2 + 2H_2 \rightarrow N_2H_4$

This reaction is endothermic.

Draw a fully labelled reaction profile for this reaction between nitrogen and hydrogen. The reaction profile shows how the energy changes during the progress of the reaction. *[3 marks]*

Total: 8

4 Methane, CH_4, is a flammable gas that is used in combustion reactions to provide heat for cooking and heating.

4–1 State whether the combustion of methane is exothermic or endothermic. Give **one** reason for your answer. *[1 mark]*

4–2 Explain why methane is a gas at room temperature. Answer in terms of structure and bonding. *[3 marks]*

4–3 **Figure 1** shows the displayed formulae for the combustion reaction between methane and oxygen.

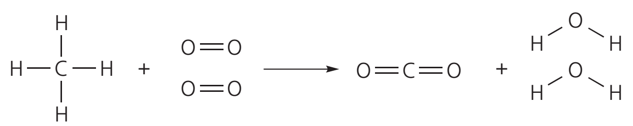


Figure 1

Table 2 shows the bond energies in the reaction. Calculate the overall energy change for the combustion of methane. Use **Figure 1** and **Table 2**. *[4 marks]*

	C–H	O=O	C=O	H–O
Energy in kJ/mol	413	496	743	463

Table 2

Total: 8

5 Titanium dioxide, TiO_2, nanoparticles are used in sunscreen.

5–1 Choose **one** property of TiO_2 nanoparticles that make them suitable for use as a sunscreen. *[1 mark]*

- They have a high surface area to volume ratio.
- They filter out ultraviolet light.
- They have diameters of less than 100 nm.
- They act as catalysts for many reactions.

5–2 Cubic nanoparticles used in sunscreen have a width of 50 nm. Calculate the surface area to volume ratio of these nanoparticles. *[3 marks]*

5–3 Give the surface area to volume ratio of a cubic nanoparticle with a width of 5 nm. *[1 mark]*

5–4 Explain why some people are concerned about the possible health effects of using TiO_2 nanoparticles in sunscreens. *[2 marks]*

5–5 Titanium is a transition metal. Give **two** differences between the properties of titanium and potassium. *[2 marks]*

5–6 The most common isotope of titanium has a mass number of 48 and an atomic number of 22. Give the number of protons, neutrons and electrons in this isotope of titanium. *[3 marks]*

Total: 12

6 Fuel cells oxidise fuels electrochemically to produce electrical energy. Hydrogen fuel cells use hydrogen as the fuel. Direct methanol fuel cells use methanol as the fuel.

6–1 Give one advantage of fuel cells over rechargeable batteries. *[1 mark]*

6–2 Give the overall balanced equation for the reaction that occurs in a hydrogen fuel cell. *[2 marks]*

6–3 Write the balanced symbol equation for the reaction between methanol and oxygen. Use the information:

In a direct methanol fuel cell, 5.12 g of methanol (CH_3OH) reacted with 10.24 g of oxygen to form 7.04 g of carbon dioxide and 5.76 g of water.

Relative formula masses (M_r): $CH_3OH = 32$; $O_2 = 32$; $CO_2 = 44$; $H_2O = 18$

You must show your working. *[4 marks]*

6–4 **Table 3** gives information about hydrogen and direct methanol fuel cells.

Evaluate the use of hydrogen fuel cells compared with direct methanol fuel cells. Use **Table 3** and your own knowledge. *[6 marks]*

	Hydrogen fuel cell	**Direct methanol fuel cell**
Energy released per litre of fuel, in kJ.	9000	16 000
Sources of fuel	Most hydrogen is produced from natural gas, but can also be made by the electrolysis of water.	Most methanol is produced from natural gas, but can also be made sustainably using fermentation.
Refuelling issues	High-pressure storage and pumping infrastructure needed.	Can use existing petrol refuelling infrastructure.
Fuel cell efficiency	Approximately 50%	Approximately 10%

Table 3

Total: 13

7 **Figure 2** shows apparatus that can be used to electrolyse molten magnesium chloride.

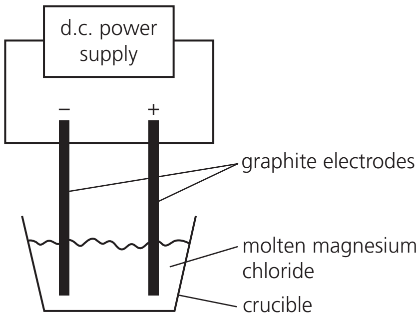


Figure 2

7–1 Explain why magnesium chloride must be molten for electrolysis to occur. *[2 marks]*

7–2 Explain why graphite is a suitable material to use for the electrodes. Answer in terms of structure and bonding. *[3 marks]*

7–3 Give the product formed at the negative electrode. *[1 mark]*

7–4 Give the products formed at the positive electrode and negative electrode when a solution of magnesium chloride is electrolysed. *[2 marks]*

7–5 A student tested the hypothesis: 'Hydrogen is always produced at the negative electrode when the metal in solution is more reactive than hydrogen.'

Table 4 shows the student's results. Suggest what the student must do next to determine whether their hypothesis is correct or not. *[2 marks]*

Aqueous solution	Observation at the negative electrode	Observation at the positive electrode
Copper iodide	Red-brown solid forms	Brown solution forms
Potassium nitrate	Colourless gas	Colourless gas
Silver sulfate	Grey solid forms	Colourless gas forms
Calcium bromide	Colourless gas forms	Yellow-orange solution forms

Table 4

Total: 10

8 A student investigated how the concentration of iron sulfate solution affected the temperature change of the displacement reaction between magnesium powder and iron sulfate.

8–1 Describe a method for this investigation. Your method should give valid results. *[5 marks]*

8–2 The student used 0.10 g of magnesium powder in each experiment. **Figure 3** shows their results. Describe the trend shown in the student's results. *[2 marks]*

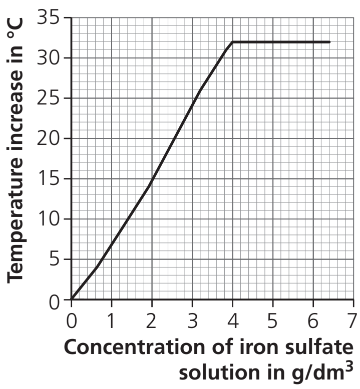


Figure 3

8–3 Explain why magnesium was the limiting reactant in the displacement reactions where the concentration of iron sulfate solution was greater than 4 g/dm³. *[3 marks]*

8–4 Calculate the concentration of a 4.00 g/dm³ solution of iron sulfate in mol/dm³.

Relative formula mass (M_r) $FeSO_4$ = 152 *[1 mark]*

Total: 11

9 This question is about a titration between sodium hydroxide, NaOH, and sulfuric acid, H_2SO_4.

9–1 Give the type of reaction that occurs between sodium hydroxide and sulfuric acid. *[1 mark]*

9–2 Sulfuric acid is a strong acid. Explain the difference between the strength of an acid and the concentration of an acid. *[2 marks]*

9–3 Describe a suitable method for carrying out a titration between sodium hydroxide solution and sulfuric acid, using phenolphthalein indicator. *[6 marks]*

9–4 The reaction between sodium hydroxide and sulfuric acid is:

$$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$$

A student titrated 25.0 cm³ portions of sodium hydroxide solution with a 0.215 mol/dm³ solution of sulfuric acid. The average titre was 20.65 cm³.

Calculate the concentration of the sodium hydroxide solution in mol/dm³. *[3 marks]*

Total: 12

Total marks for Paper 1: 100

Paper 2

1 A student investigated how the surface area of a piece of calcium carbonate affected the rate of reaction with hydrochloric acid.

1–1 Copy and complete this sentence. Choose your answer from the options below. *[1 mark]*

increases	stays the same	decreases

As the surface area increases, the rate of reaction …………………………

1–2 What are the units of rate of reaction? Choose **one** answer from the options below. *[1 mark]*

g/cm^3	cm^3/g	g/s

1–3 Give a reason why the student observed bubbles. *[1 mark]*

1–4 **Figure 1** shows a sketch of how mass changes for the reaction between hydrochloric acid and small pieces of calcium carbonate. Copy **Figure 1** and sketch the line of best fit for the same reaction with larger calcium carbonate chips. *[3 marks]*

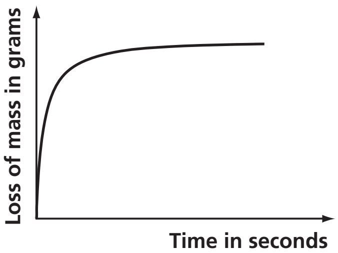


Figure 1

Total: 6

2 Chromatography is a separation technique that can be used to separate colours in washable ink pens. **Figure 2** is the set up for a chromatography experiment to separate the colours of the ink.

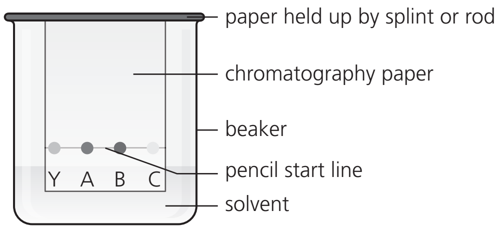


Figure 2

2–1 Explain why pencil is used to draw the start line and mark the solvent front. *[2 marks]*

2–2 Explain why it is important that the solvent level is below the base line. *[2 marks]*

2–3 Explain why it is important to mark the solvent front as soon as the chromatogram is taken from the developing tank. *[4 marks]*

2–4 Suggest **one** improvement to this experiment that would give a more reproducible chromatogram. Explain your answer. *[3 marks]*

Total: 11

3 Ammonium chloride can reversibly decompose into ammonia and hydrogen chloride.

3–1 Give the symbol that shows this is a reversible reaction. *[1 mark]*

3–2 State condition needed for the forward reaction to happen. *[1 mark]*

3–3 Describe the observations for the forward reaction. *[2 marks]*

Total: 4

4 Copper is a transition metal that can form compounds as a 2^{+} ion.

4–1 Give the formula of the copper(II) ion. *[1 mark]*

4–2 State the colour of a copper compound in a flame test. *[1 mark]*

4–3 Copper (II) sulfate will react with sodium hydroxide to make sodium sulfate and a copper(II) salt. Write a word equation for this reaction. *[1 mark]*

4–4 State the colour of the precipitate of the copper(II) compound made in this reaction. *[1 mark]*

Total: 4

5 Flame tests are an analytical technique used to show that certain metal ions are present in a sample.

5–1 Describe how a flame test can be used to identify a sodium compound. *[6 marks]*

5–2 Define the term ‘cation’. *[1 mark]*

5–3 Explain why it would be difficult to use flame tests to determine the composition of a mixture of cations. *[1 mark]*

5–4 Explain why it is difficult to use a flame test to determine if a solution has lithium or calcium ions in it. *[3 marks]*

5–5 Explain why it is important to ensure that the wire used for the flame test is clean. *[2 marks]*

Total: 13

6 When sodium is put into water it reacts to form hydrogen gas and one other product.

6–1 Write the word equation for this reaction. *[2 marks]*

6–2 Describe the observations that you would make if you put sodium into water. *[3 marks]*

6–3 Describe how the gas could be collected and tested to show it is hydrogen. *[6 marks]*

Total: 11

7 Crude oil is a finite resource found in rocks. Products of crude oil are often used as fuels.

7–1 Explain why crude oil is an example of a finite resource. *[1 mark]*

7–2 Explain how crude oil is separated. *[6 marks]*

7–3 Methane (CH_4) can be extracted from crude oil and used in camping gas. Balance the symbol equation for the complete combustion of methane. *[1 mark]*

$CH_4 + O_2 \rightarrow CO_2 + H_2O$

7–4 Large hydrocarbons can undergo a chemical reaction known as cracking to make smaller more useful hydrocarbons. Write a balanced symbol equation for the cracking of decane ($C_{10}H_{22}$) to form ethene and one other product. *[2 marks]*

7–5 Ethene can react with steam to make an alcohol. Draw the displayed formula of this alcohol. *[2 marks]*

Total: 12

8 **Figure 3** shows the composition of gases in dry air.

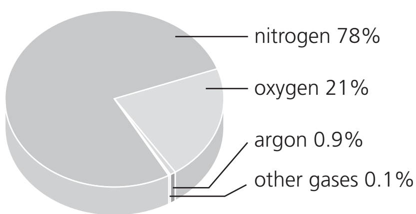


Figure 3

8–1 Give the approximate fraction of dry air that is nitrogen. *[1 mark]*

8–2 Describe how to test for the gas that makes up about 20% of dry air. *[3 marks]*

8–3 Describe how carbon dioxide decreased from its level in Earth's early atmosphere to the present level. *[3 marks]*

Total: 7

9 Flame emission spectroscopy is an instrumental method that can be used to analyse a substance.

9–1 Describe how flame emission spectroscopy can be used to identify metal ions and give their concentrations. *[4 marks]*

9–2 The flame emission spectroscope must be calibrated. The intensity of a flame is measured compared to known concentrations of the solutions. **Figure 4** shows a graph of these data. Draw a line of best fit. *[1 mark]*

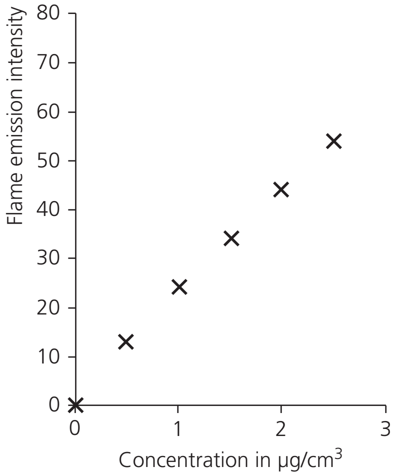


Figure 4

Total: 5

10 Brine is a solution of sodium chloride. It can undergo electrolysis to make hydrogen, chlorine and sodium hydroxide.

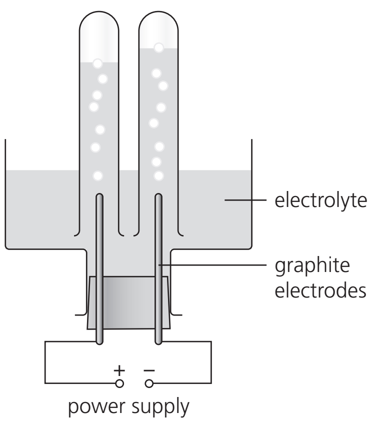


Figure 8

10–1 Write a half equation for hydrogen being made at the cathode. *[3 marks]*

10–2 Describe a chemical test that can be used on the gas collected at the anode to show it is chlorine. *[3 marks]*

10–3 Explain why the electrolysis of sodium chloride only happens if the chemical is molten or in solution. *[2 marks]*

10–4 Suggest what would be observed if universal indicator was added to the electrolyte. *[4 marks]*

Total: 12

11 Copper(II) chloride is widely used in industry as a catalyst.

11–1 Sketch a reaction profile of an exothermic reaction to show the effect of adding a catalyst. *[3 marks]*

11–2 Explain why copper makes a good catalyst. *[2 marks]*

11–3 Write a balanced ionic equation for the formation of a blue precipitate when a solution of copper(II) chloride reacts with sodium hydroxide. *[3 marks]*

11–4 Copy the equation and then add the state symbols to the ionic equation for the formation of a white precipitate when a solution of copper(II) chloride reacts with acidified silver nitrate. *[1 mark]*

Ag^{+} ………… + Cl^{-} ……..…… → AgCl…………

11–5 The copper(II) chloride solution could be analysed using instrumental analysis. Describe the advantages of using instrumental analysis. Explain your answers. *[3 marks]*

Total: 12

12 A student analysed an unknown compound, compound A. The compound gave a green flame in a flame test. When the student added hydrochloric acid followed by barium chloride solution to a solution of A, a white precipitate was produced.

12–1 Identify the positive ion in compound A. *[1 mark]*

12–2 Identify the negative ion in compound A. *[1 mark]*

12–3 Write the name and formula of compound A. *[1 mark]*

Total: 3

Total marks for Paper 2: 100

Answers

A simple model of the atom, symbols, relative atomic mass, electronic charge and isotopes (p. 1)

Quick questions

1 An atom.

2 NaCl: sodium chloride: sodium and chlorine; MgO: magnesium oxide: magnesium and oxygen; H_2S: hydrogen sulfide: hydrogen and sulfur; AlF_3: aluminium fluoride: aluminium and fluorine; CuI_2: copper iodide: copper and iodine;

3 I: A; II: A, III: B

4 Evaporation or crystallisation

5 Atomic number is the number of protons in the nucleus of an atom. Mass number is the number of protons and neutrons in the nucleus of an atom.

6 1×10^{-9} m (1 nm)

7 Isotopes

8 The average mass of atoms of an element that takes into account the mass and amounts of each isotope the element contains.

9 oxygen atom: 2.6; calcium atom: 2.8.8.2

Exam-style questions

10–1 sodium + water → sodium hydroxide + hydrogen [1]

10–2 Hydrogen contains only one type of atom [1] whilst water contains two types of atoms (hydrogen and oxygen), chemically bonded together in a fixed ratio [1]

10–3 H:O ratio is 2:1 [1]

11–1 Distillation. [1]

11–2 The condenser causes the water vapour to condense [1] so that it can be collected as a liquid [1].

11–3 As the mixture of ethanol and water is heated the substance with the lowest boiling point will boil and leave the mixture first [1]. Ethanol must have a lower boiling point than water [1].

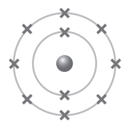
12–1 Ne atoms have 10 protons, whilst Mg atoms have 12 [1]; Ne atoms have 10 neutrons whilst Mg atoms have 12 [1]; and Ne atoms have 10 electrons, whilst magnesium atoms have 12 [1].

12–2 Ne atoms have the same number of protons as electrons [1]. Protons are positively charged and electrons are negatively charged [1]. Because there is the same number of each, their charges cancel [1].

12–3 $^{2+}$ [1]. If a Mg atom loses two electrons it will have two more protons than electrons, so their charges will no longer cancel out [1]

13–1 Mg_3N_2 [1]

13–2



Marks for: Two electrons in inner shell [1], *8 electrons in outer shell* [1]

13–3 protons: 7 [1], neutrons: 7 [1], electrons: 10 [1]

14–1 Fill the beaker with water to a depth less than 1 cm [1]

14–2 During chromatography the three substances move up the paper at different speeds [1]. Substances which are more soluble in the solvent travel further up the paper [1] so the substances are separated across the paper [1].

14–3 Colouring A contains three different substances [1]. B and C contain only two different substances [1]. A and B contain the same red substance [1]. B and C contain the same green substance [1]. A and C contain the same blue substance [1]. Only A contains the purple substance [1].

15–1 Ar [1]

15–2 Ca^{2+} [1]

15–3 Cl^- [1]

15–4 Ar [1]

15–5 protons: 15 [1], neutrons 16 [1], electrons 15 [1]

15–6 The plum-pudding model of the atom described atoms as spheres of positive charge with negatively charged electrons embedded throughout them [1]. Ernest Rutherford's work proved that the atom consisted of a small central nucleus [1] that was positively charged [1], with electrons around the nucleus (the nuclear model). Neils Bohr's work later showed that the electrons orbited the nucleus at certain distances from it [1], called energy levels (or shells) [1]. Finally, James Chadwick discovered neutrons [1] that were the last sub-atomic particle to be discovered.

16–1 KF [1]

16–2 $2K + F_2 \rightarrow 2KF$
correct reactants and products [1], *correct balancing* [1]

16–3 $K \rightarrow K^+ + e^-$
correct reactants and products [1], *electron on right hand side* [1]

16–4 $F_2 + 2e^- \rightarrow 2F^-$
correct reactants and products [1], *correct balancing* [1], *electrons on left hand side* [1]

17–1 An isotope is an atom with the same number of protons [1] and electrons [1], but a different number of neutrons [1] to other atoms of the same element.

17–2 ^{28}Si [1]. Relative atomic mass is the average mass of isotopes taking into account their relative abundances [1]. The relative atomic mass is 28.1, so ^{28}Si must be the most abundant to cause the average mass to be so close to 28 [1].

17–3 D [1]

18–2 Both isotopes have the same number of protons [1] and electrons [1]. The ^{34}S isotope has two more neutrons than the ^{32}S isotope [1].

18–3 $A_r = [(32 \times 95) + (33 \times 0.8) + (34 \times 4.2)] \div 100$ [1] = 32.092 [1] = 32.1 (3 sf) [1]

The periodic table (p. 5)

Quick questions

1 Group 7 (if the atomic number is 17, then there will be 17 protons and 17 electrons in the atoms. 17 electrons will be arranged as 2.8.7, so with 7 electrons in the outer shell, the element is in Group 7). Atomic number = 9 (moving up the Group we lose one shell of electrons, so the electron arrangement must be 2.7)

2 He believed that there were still elements to be discovered, so left gaps for them.

3 When Mendeleev's predictions were proven correct his model of the periodic table gained more support from other chemists of the time.

4 A metal

5 All Group 0 elements have a full outer shell of electrons/they are all unreactive.

6 Reactivity increases down the Group.

7 Boiling point increases down the Group.

Exam-style questions

8–1 Carbon atoms have four electrons in their outer shell [1]

8–2 A (2.8.4) [1]

8–3 SiH_4 [1]. The electron structures of carbon and silicon are 2.4 and 2.8.4 respectively/ they both have 4 electrons in their outer shells [1] so they have the same chemical properties and react similarly [1].

9–1 A dark/black coloured solid [1]

9–2 As the relative molecular mass of the halogens increases, the boiling point increases. [1]

9–3 260 to 270 °C [1]

9–4 sodium chloride and astatine [1]

9–5 $2At^- \rightarrow At_2 + 2e^-$
Correct reactants and products [1], *correct balancing* [1], *2e– on right hand side* [1]

9–6 When halogens react they gain one electron to fill their outer shell of electrons [1]. Chlorine atoms will have fewer electron shells than astatine atoms, so the gained electron is closer to the nucleus in chlorine[1]. So the electron is more strongly attracted to the chlorine nucleus [1] than in astatine, and is easier to gain [1]. This makes chlorine more reactive.

10–1 $2Li + Cl_2 \rightarrow 2LiCl$ (correct reactants and products [1], correct balancing [1])

10–2 $Li \rightarrow Li^+ + e^-$ *correct reactants and products* [1], *correct balancing* [1], 1e– *on right hand side* [1]

10–3 Potassium and lithium both have one outer-shell electron in their atoms [1] so they have similar chemical properties [1]. When Group 1 metals react they lose their one outer-shell electron [1]. Potassium has more electron shells than lithium so the outer-shell electron is further from the nucleus [1] and less strongly attracted to it [1]. This makes it easier for the electron to be lost [1], so potassium is more reactive.

11–1 As the relative mass of the noble gas atoms increases, boiling point increases. [1]

11–2 – 160 °C to – 140 °C [1]

11–3 Noble gases are unreactive/inert elements [1] whereas the halogens are very reactive elements [1]. Noble gases are unreactive because they all have a stable electron arrangement/full outer shell of electrons [1] so they do not easily react. Halogens are reactive because they all have seven electrons in their outer shells [1]. They react by gaining one electron to complete their outer shell [1]. The halogens get less reactive the further down Group 7 they are [1].

Properties of transition metals (p. 8)

Quick questions

1 *Any one from:* catalytic properties/they form ions with different charges/they form coloured compounds.

2 *Any one from:* they tend to have higher melting points/higher densities/greater strength/hardness.

Exam-style questions

3–1 $2Ni + O_2 \rightarrow 2NiO$ correct reactants and products [1], correct balancing [1]

3–2 Protons are positively charged and electrons are negatively charged [1]. In atoms there are the same number of protons as electrons, so their charges cancel out [1]. When an atom loses one electron there will be one more proton than electrons, so the charges will no longer cancel and the ion will be positively charged overall [1]

3–3 It is harder for nickel atoms to lose their outer-shell electrons than sodium atoms [1]

3–4 $Ni \rightarrow Ni^{2+} + 2e^-$ correct reactants and products [1], 2e– on right hand side [1]

4–1 Shiny when polished [1]

4–2 Transition metals form coloured compounds [1]

4–3 **Chemical properties**: transition metals (and their compounds) tend to have catalytic properties/can be used as catalysts [1], form ions with different charges [1] and form coloured compounds [1] whilst Group 1 metals do not.
Physical properties: transition metals tend to be harder/stronger [1], have higher densities [1] and higher melting points [1] than Group 1 metals

Atomic structure and the periodic table topic review (p. 9)

1–1 MnS [1]

1–2 Manganese is a transition metal. [1]

1–3 25 protons [1], 30 neutrons [1], 25 electrons [1]

1–4 *Award a maximum of* [3] *marks for any three from*: reacts to form ions with different charges [1]; forms coloured compounds [1]; it (or its compounds) has catalytic properties [1]; generally low reactivity [1].

1–5 2.8.6 [1]

1–6 Sulfide ions have 16 protons and 18 electrons [1]. The total positive charge is 16+ and the total negative charge is 18– [1]. So the overall charge is 2– [1].

2–1 electron shells [1]

2–2 neutrons [1]

2–3 nucleus [1]

2–4 The plum-pudding model of the atom described atoms as balls of positive charge with negatively charged electrons spread through them [1]. The results from the scattering experiment could not be explained by the plum-pudding model [1]. The results proved that atoms had a tiny, positively charged nucleus [1]. This led to the development of the nuclear model of the atom [1]

3–1 As the relative molecular mass of the elements increases, their boiling point increases [1].

3–2 $2Fe + 3Br_2 \rightarrow 2FeBr_3$ *correct reactants and products* [1], *correct balancing* [1].

3–3 Iodine reacts in a similar way to bromine because it has the same number of outer-shell electrons as bromine [1]. Iodine is less reactive than bromine [1]. When halogens react they gain one electron to fill their outer-shell of electrons [1]. Iodine atoms will have more electron shells than bromine atoms, so the gained electron is further from the nucleus in iodine [1]. This means the electron is less strongly attracted to the iodine nucleus [1], and is harder to gain [1]. This makes iodine less reactive.

3–4 iron chloride and bromine [1]

3–5 crystallisation OR evaporation [1]

4–1 They all have one electron in their outer shells. [1]

4–2 Sodium is on the left-hand side of the periodic table, whilst chlorine is on the right-hand side [1]. When sodium atoms react they form positive ions (by losing their one outer-shell electron) [1]. When chlorine atoms react they form negative ions (by gaining one outer-shell electron) [1].

4–3 sodium hydroxide [1]; hydrogen [1]

4–4 When Group one metals react they lose their one outer-shell electron [1]. Lithium has fewer electron shells than sodium so the outer-shell electron is closer to the nucleus [1] and more strongly attracted to it [1]. This makes it harder for the electron to be lost [1], so lithium is less reactive.

5–1 Because the elements had similar properties. [1]

5–2 Newlands' groupings stopped working after the first 20 elements or so [1]. For instance, he placed iron, a metal [1], together with oxygen and sulfur, which are non-metals [1] (*accept any other correct metal with non-metals grouping from Newland's table*)

5–3 Group 0 elements are very unreactive [1], so had not been discovered at that time [1].

5–4 Mendeleev left gaps in his periodic table for elements that had not been discovered yet [1]. He also swapped the positions of some elements to better suit their properties [1]. Mendeleev made predictions about the undiscovered elements [1]. When new elements were discovered that matched Mendeleev's predictions his periodic table became widely accepted [1].

Chemical bonds: ionic, covalent and metallic (p. 12)

Quick questions

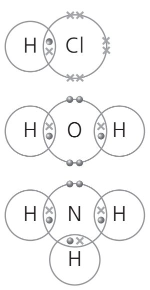
1 two non-metals: covalent bonding; two metals: metallic bonding; a metal and a non-metal: ionic bonding

2 ionic bonding: transferred; metallic bonding: shared; covalent bonding: shared

3 An ionic compound

4 Covalent bonding

5



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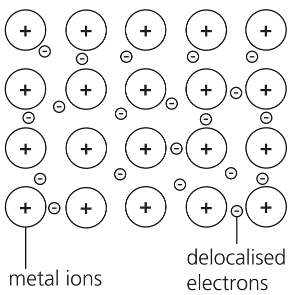


Exam-style questions

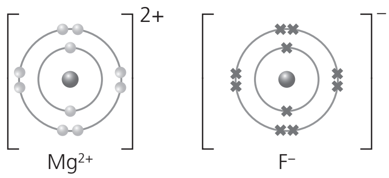
7–1 metallic bonding [1]

7–2 covalent bonding [1]

7–3 NaH is formed from a metal reacting with a non-metal [1]. Electrons are transferred from the sodium atoms to the hydrogen atoms [1]. The particles are ions, Na^+ and H^-, and the bonding is ionic with a giant lattice structure [1]. HCl is formed from two non-metallic elements reacting together [1]. Electrons are shared between the atoms [1]. The particles are molecules and the bonding is covalent [1].

8–1 Mg forms 2+ ions. F forms 1– ions. [1] Formula is MgF_2. [1]

8–2



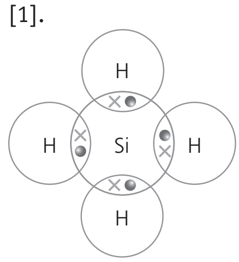
[1] mark for **each** correct ion (must include the charge of each ion). Both ions have the same electronic structure as neon atoms [1].

8–3 The sticks in the image may be confused with covalent bonds but there are no covalent bonds present, *or*: the model suggests the ions are far apart, and they are not. [1]

8–4 Each magnesium atom loses its two outer-shell electrons [1] to form magnesium ions with a 2+ charge (Mg^{2+}) [1]. Each fluorine atom gains one electron [1] to get a full outer shell, forming a fluoride ion with a 1- charge (F^-) [1]. The oppositely charged ions attract each other and form a giant lattice structure [1].

9–1 $Si + 2H_2 \rightarrow SiH_4$ Correct reactants and products [1], correct balancing [1].

9–2 [1]



9–3 Each circle represents an atom [1]. Each line represents a pair of electrons [1].

9–4 3D space-filling models give a better representation of how close the atoms are *or* 3D space-filling models show how the atoms merge together when they bond [1].

9–5 $H_6Si_2O_7$ [1] *order of the elements is not important*

10–1 Figure 6 represents metallic bonding [1]. There is a giant lattice of metal atoms [1]. The outer-shell electron from each atom becomes delocalised [1], which means they can move throughout the whole structure [1]. There is a strong attraction between the delocalised electrons and the positive nuclei of the metal atoms (this is the metallic bond) [1].

10–2 $Na_2CO_3 + 2C \rightarrow 2Na + 3CO$ *correct reactants and products* [1], *correct balancing (can include multiples of balancing shown)* [1].

How bonding and structure are related to the properties of substances (p. 15)

Quick questions

1 silica

2 ions

3 delocalised electrons

4 They do not contain charged particles that are free to move.

5 Metals are malleable/metals conduct electricity when solid.

Exam-style questions

6–1 Solid [1]

6–2 Oxygen [1]

6–3 Nitrogen [1] because it has the lowest boiling point [1].

6–4 Nitrogen and oxygen have low melting and boiling points [1]. Small molecules only have weak intermolecular forces between them [1] that do not need a lot of energy to overcome, so they have low melting and boiling points [1].

7–1 Covalent bonding [1] because carbon and oxygen are both non-metal elements [1].

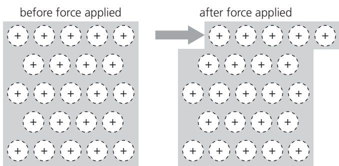
7–2 The particles/molecules/atoms are neutrally charged/they do not contain ions/delocalised electrons which are free to move [1]

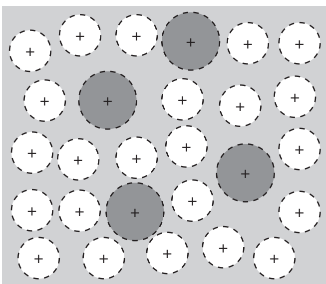
7–3 Carbon dioxide is made of small molecules [1]. Between these molecules are only weak intermolecular forces of attraction [1]. Very little energy is needed to overcome these forces, so at room temperature carbon dioxide is a gas [1]. Silica has a giant covalent structure [1]. Between the atoms are strong covalent bonds [1]. A lot of heat energy is required to overcome these covalent bonds, so silica has a high melting point [1].

8–1 Metals have delocalised electrons in their structure [1]. Electricity is the flow of charged particles, and the delocalised electrons are free to move, so metals conduct electricity [1].

The delocalised electrons also transfer thermal energy [1] so metals conduct heat.

8–2 Metals are malleable because their atoms are arranged in layers [1]. These layers are able to slide over each other easily [1] so metals can be bent or hammered into shape. Alloys are metals with other elements added to them [1]. These different atoms disrupt the lattice structure [1] preventing the layers from sliding over each other [1]. This makes the alloy harder. *Either of the following two diagrams could be used to support the answer* [1].





9–1 They have strong forces of attraction between their particles [1]

9–2 Substance A is a metal [1]. Metals contain delocalised electrons that are free to move and carry the charge when solid and liquid [1], so metals conduct when solid and liquid. Substance B is ionic [1]. The ions can only flow when the substance is a liquid so it doesn't conduct when solid. [1] Substance C is giant covalent [1]. There are no charged particles that can flow, so it doesn't conduct electricity [1].

10–1 Ionic bonding [1]. Oppositely charged ions are held in a giant lattice structure [1] by strong forces of attraction [1] between them.

10–2 Electricity is the flow of charged particles. In potassium bromide the ions carry the charge [1]. When solid the ions are not free to move so charge cannot flow [1]. When molten or dissolved the ions are free to move, so charge can flow [1].

10–3 Going down the Group 1 bromides the melting point decreases [1]. The force of attraction between the ions must therefore be getting weaker down Group 1 [1]. Going down Group 1 we add an extra shell of electrons to each ion, so they get bigger [1]. This matches the statement that bigger ions have weaker forces of attraction [1].

Structure and bonding of carbon (p. 18)

Quick questions

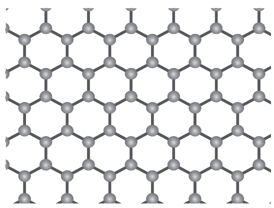
1 D; A; B; C

2 Drill and/or saw tips.

3 They have high tensile strength.

Exam-style questions

4–1 Graphene contains carbon atoms bonded in a single flat layer/graphene is a single layer of graphite [1]. Each carbon atom is bonded to three other carbon atoms in the layer [1].



4–2 They both contain delocalised electrons [1].

4–3 Fullerenes have hollow parts at the centre of the molecule that can be used as a cage/carry the drug molecule [1].

4–4 Fullerenes have roughly spherical shapes so they can roll past each other [1].

4–5 Carbon nanotubes are tubes of carbon atoms bonded together by strong covalent bonds [1]. The strength of these bonds means it is hard to pull the carbon atoms apart, so the nanotube has a high tensile strength [1].

4–6 Diamond has a giant covalent structure in which each carbon atom is bonded to four other carbon atoms by covalent bonds [1]. These bonds are very strong, (holding the atoms rigidly in place) so diamond is a hard substance [1]. Diamond does not contain any delocalised electrons (as all the electrons are used in bonding) so it does not conduct electricity [1]. Graphite has a giant covalent structure in which each carbon atom is bonded to only three other carbon atoms in flat layers [1]. The layers themselves are not strongly bonded together and can slide over each other, making graphite soft [1]. Graphite does contain delocalised electrons (as each carbon atom has one un-bonded electron) so it does conduct electricity [1].

Bulk and surface properties of matter including nanoparticles (p. 19)

Quick questions

1 1×10^{-9} m and 1×10^{-7} m

- 15 nm × 15 nm × 6 = 1350 nm^2 (*six sides of the cube*)
- 15 nm × 15 nm × 15 nm = 3375 nm^3
- 1350 : 3375 = 1 : 2.5 = 0.4

3 Any three from: in fuel cells [1]; medicine OR drug delivery OR synthetic skin [1]; sun cream [1]; cosmetics [1]; clothing [1]; deodorants [1]; electronics [1].

Exam-style questions

4–1 Nanoparticles are very small so have very high surface area to volume ratios compared with larger particles [1] (this means that a much higher fraction of atoms are at the surface of the nanoparticle, and this gives the nanoparticles different properties).

4–2 Due to the high surface area to volume ratio, only small amounts of silver nanoparticles are necessary to be effective. [1]

4–3 Diameter of Ag atom in nm = 0.17 nm [1]
Note that the width of the cube could be converted into metres for this mark.
number of Ag atoms = 20 nm ÷ 0.17 nm = 117.6 atoms = 118 atoms [1]

4–4 20 nm cube: surface area = 6 × 20 nm × 20 nm = 2400 nm^2 [1]; volume = 20 nm × 20 nm × 20 nm = 8000 nm^3; SA : V = 2400 : 8000 = 1 : $3\frac{1}{3}$ [1]
200 nm cube: surface area = 6 × 200 nm × 200 nm = 240 000 nm^2 [1]; volume = 200 nm × 200 nm × 200 nm = 8 000 000 nm^3 ; SA : V = 240 000 : 8 000 000 = 1 : $33\frac{1}{3}$ [1]; 1 : $3\frac{1}{3}$ is 10 times greater than 1 : $33\frac{1}{3}$.

4–5 Because nanoparticles have different properties to those of the bulk material it is not safe to assume that they are harmless [1]. Silver nanoparticles may be able to enter into cells and cause damage, unlike bulk silver [1].

Bonding, structure and the properties of matter topic review (p. 19)

1–1 Aluminium has a lattice structure with layers of metal atoms surrounded by delocalised electrons [1]. The layers of atoms are able to slide past each other easily when a force is applied [1] so the metal is malleable and can be shaped [1].

1–2 Electricity is the flow of charged particles [1]. Aluminium contains delocalised electrons that can move and carry the charge [1].

1–3 Each aluminium atom loses three electrons to form Al^{3+} ions [1]. Each oxygen atom gains two electrons to form O^{2-} ions. [1] The oppositely charged ions attract each other strongly, and this attraction is the ionic bond [1].

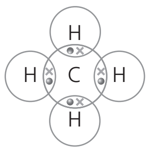
1–4 The ionic bond is very strong so a lot of heat energy is required to overcome it [1].

1–5 Electricity is the flow of charged particles. [1] The aluminium oxide is solid so the ions are unable to move and carry the charge [1].

1–6 Nanoparticles have a much higher surface area to volume ratio [1].

1–7 surface area = 6 × 30 nm ×30 nm = 5400 nm² [1]; volume = 30 nm × 30 nm × 30 nm = 27 000 nm³; SA : V ratio = 5400 : 27000 = 1 : 5 [1]

2–1 [1]



2–2 Methane is made up of small molecules with weak intermolecular forces between them [1]. These forces do not require a lot of heat energy to overcome, so methane has a low melting and boiling point [1]. Electricity is the flow of charged particles, but the molecules do not have an overall electric charge, so methane does not conduct electricity [1].

2–3 Fullerenes have a cage-like/hollow structure [1] that smaller molecules can be put into.

2–4 Diamond has a giant covalent structure [1] where every carbon atom is bonded to four other carbon atoms [1].

2–5 The bonds between the atoms in silicon carbide are weaker than in diamond. [1]

3–1 Substance B [1] as it has the lowest melting and boiling point [1]. OR Substance B [1] as it is a poor conductor of electricity as a solid and as a liquid, so must be simple covalent/small molecules [1].

3–2 Delocalised electrons [1] (*substance D must be a metal*)

3–3 Substance A [1]. The structure is ionic and the bonds between the oppositely charged ions are strong [1] so a lot of heat energy is required to overcome them (giving high melting and boiling points). When the substance is solid the ions are unable to move and carry the charge so it cannot conduct electricity [1], but when it is a liquid the ions are free to move and carry the charge, so it can conduct electricity [1].

3–4 Substance C has a very high melting point. Graphite has a giant covalent structure with many strong covalent bonds between the carbon atoms [1]. A lot of heat energy is required to overcome these bonds so it has a very high melting point [1]. Graphite contains delocalised electrons [1] that can carry the charge, so it conducts electricity.

Chemical measurements, conservation of mass and the quantitative interpretation of chemical equations (p. 22)

Quick questions

1 15 atoms

2 $2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$

3 24 + (2 × 14) + (6 × 16) = 148

4 When metals react with oxygen gas (or chlorine gas), or during thermal decomposition reactions.

5 Anomalous value (2.65 g) discarded. Average = (2.95 + 2.89 + 2.99) ÷ 3 = 2.94 g The average value is no more than 0.05 g from the actual values, so the uncertainty = ± 0.05 g

Exam-style questions

6–1 There is the same number of atoms of each element on both sides of the equation [1].

6–2 M_r = 23 + 1 + 12 + (3 × 16) = 84 [1]

6–3 Relative mass of the reactants = 2 × 84 = 168 [1]; Relative mass of the products = 106 + 44 + 18 = 168 [1]. The sum of the relative formula masses of the reactants is the same as for the products, so mass is conserved [1].

6–4 Mass of products = 26.51 g − 25.00 g = 1.51 g [1]

6–5 The water and carbon dioxide will have escaped from the crucible [1], so their masses will not have been measured/only the mass of the solid (sodium carbonate) has been measured. [1]

7–1 $2Ca + O_2 \rightarrow 2CaO$
Correct reactants and products [1]. Correct balancing [1]

7–2 Masses of calcium oxide: 2.81 g, 2.79 g, 2.41 g, 2.84 g [1]
Identification and removal of anomalous value (2.41 g) [1]
Correct mean value calculated: 2.81 g [1]
Mean given to three significant figures [1]
Correct uncertainty stated: ± 0.03 g [1]

7–3 The calcium gains oxygen as calcium oxide forms [1], but this mass of oxygen is not measured/only the mass of calcium is measured as a reactant by the mass balance [1].

8–1 When the mass of the products is measured it will not include the mass of NO_2 and O_2 as these are both gases [1]. This makes the mass of reactants appear to be greater than the mass of products [1].

8–2 Mass is conserved during a reaction, so the relative masses of the reactants = the relative mass of all the products [1]. 2(331) = 2(223) + 32 + 4(NO_2) [1]; 662 = 478 + 4(NO_2); 4(NO_2) = 184 [1]; NO_2 = 46 [1]

Use of amount of substance in relation to masses of pure substances (p. 24)

Quick questions

1 1.36 g/dm³

2 The mole, or mol.

3 The number of particles in one mole of a substance. Its value is 6.02×10^{23}.

4 The mass of one mole of a substance is the same as its relative formula mass.
63.5 + (2 × 16) + (2 × 1) = 97.5 g for $Cu(OH)_2$.

5 The limiting reactant is the reactant that gets completely used up (while other reactants are usually in excess).

6 The molar ratio for $CH_4 : H_2$ is 1 : 3. 1.50 mol of CH_4 will produce 4.50 mol of H_2.

Exam-style questions

7–1 $Na_2S_2O_3$ [1]

7–2 volume in dm³ = 50 cm³ ÷ 1000 = 0.050 dm³ [1]; concentration = mass ÷ volume = 0.25 g ÷ 0.050 dm³ = 5.0 g/dm³ [1]

7–3 The concentration of the solution will halve [1] as doubling the volume halves the concentration [1].

7–4 volume in dm³ = 75.0 cm³ ÷ 1000 = 0.075 dm³ [1]
mass = concentration × volume = 1.85 g/dm³ × 0.075 dm³ = 0.13875 g [1] = 0.139 g (3 sf) [1]

7–5 volume = mass ÷ concentration = 0.40 g ÷ 2.75 g/dm³ = 0.145 dm³

(3 sf) [1]; volume in cm^3 = 0.145 dm^3 × 1000 = 145 cm^3 (3 sf) [1]

8–1 *Correct whole number ratio of reactants and products*: $C_3H_8 : O_2 : CO_2 : H_2O$ = 1 : 5 : 3 : 4 [1] *Correct symbol equation*: $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$ [1].

8–2 mol C_3H_8 = mass ÷ M_r = 5.52 ÷ 44 = 0.08 mol [1]; mol CO_2 = 3 × mol C_3H_8 = 3 × 0.08 mol = 0.24 × mol [1]; mass CO_2 = mol × M_r = 0.24 × 44 = 10.56 g [1]

8–3 mol C_3H_8 = 0.22 ÷ 44 = 0.005 mol [1]; mol O_2 = 0.40 ÷32 = 0.0125 mol [1]; O_2 reacts with C_3H_8 in a 5 mol: 1 mol ratio, so 0.0125 mol O_2 requires 0.0025 mol C_3H_8, or 0.005 mol of C_3H_8 requires 0.025 mol O_2 [1] There is twice this amount, so propane is in excess/oxygen is the limiting reactant, or there is only half this amount of O_2 so propane is in excess/oxygen is the limiting reactant [1].

9–1 mol Al_2O_3 = mass (in grams) ÷ M_r = 400 000 ÷ 102 = 3922 mol [1]; mol Al = 2 × mol Al_2O_3 (*2:4 ratio from the balanced equation*) = 7844 mol (*allow 7843 if 3922 not rounded*) [1]; mass Al = mol × M_r = 7844 × 27 = 211 788 g = 211.8 kg [1]

9–2 211.8 kg × (24 hours × 7 days) = 35582 kg per week [1] = 35 600 kg per week (3 sf) [1]

9–3 mol Al = 50 000 ÷ 27 = 1852 mol [1]; mol NH_4ClO_4 = 150 000 ÷ 117.5 = 1277 mol [1] 10 mol Al will react with 6 mol of NH_4ClO_4, so 1 mol Al will react with 0.6 mol of NH_4ClO_4. 1852 mol Al will react with 1111 mol of NH_4ClO_4 [1]. There is 1277 mol NH_4ClO_4, so it is in excess/Al is the limiting reactant. [1]

9–4 mol NH_4ClO_4 = mass ÷ M_r = 2.35 ÷ 117.5 = 0.02 mol; mol HCl = 0.73 ÷ 36.5 = 0.02 mol; mol N_2 = 0.28 ÷ 28 = 0.01 mol; mol O_2 = 0.80 ÷ 32 = 0.025 mol; mol H_2O = 0.54 ÷ 18 = 0.03 mol *Correct amounts of reactants and products in mol* [1] *Correct whole number ratio of reactants by dividing each number of moles by 0.01 mol* [1] *multiplying each result by 2* [1] to give $NH_4ClO_4 : HCl : N_2 : O_2 : H_2O$ = 4 : 4 : 2 : 5 : 6 Balanced equation: $4NH_4ClO_4 \rightarrow 4HCl + 2N_2 + 5O_2 + 6H_2O$ [1]

Yield and atom economy of chemical reactions (p. 26)

Quick questions

1 The higher the percentage yield of a reaction, the more efficient the process is/the less waste is produced/the greater the profits for the manufacture.

2 Atom economy = 100 × (sum of the M_r of desired products/sum of the M_r of all the reactants)

3 If there is only one product then all the atoms of the reactants must have formed the product, so the M_r of the desired product is equal to the M_r of all the reactants.

Exam-style questions

4–1 Percentage yield = 100 × (31.7 ÷ 170) = 18.647% [1] = 18.6% (3 sf) [1]

4–2 The reaction does not go to completion/the reaction is reversible. [1] Some of the product may be lost during separation from the reaction mixture. [1] The reactants may react in a different way to the desired reaction/the reactants may carry out different reactions. [1]

4–3 The percentage atom economy for the Haber process is 100% as there is only one product. [1] This means there is less waste of the Earth's resources in undesired products. [1]

5–1 Percentage atom economy = 100 × ((2 × 46) ÷ 180); *Correct doubling of M_r of ethanol* [1] = 51.111 % [1] = 51 % (2 sf) [1]

5–2 Actual yield = maximum theoretical yield × percentage yield [1] = 24.8 kg × (92 ÷ 100) = 22.816 kg [1] = 22.8 kg (3 sf) [1]

5–3 *At least one direct comparison must be made between the two processes.* [1] The source of the glucose in Process 1 is plant matter which is renewable so sustainable [1], while in Process 2 the ethene from crude oil is non-renewable, so not sustainable (as crude oil will eventually run out) [1] Process 2 is more sustainable in terms of atom economy. The atom economy for Process 2 is 100% (as there is only one product) and much greater than for Process 1 [1]. This means less waste in Process 2 [1]. Process 1 has a low energy requirement compared to Process 2. This means fewer energy resources are required for Process 1, making it more sustainable than Process 2 [1].

Using concentrations of solutions in mol/dm³ (p. 27)

Quick questions

1 2.2 mol/dm^3

2 0.0218 mol

3 0.00282 mol/dm^3

Exam-style questions

4–1 Concentration = mass ÷ volume in dm^3 = 10.0 g ÷ 0.250 dm^3 [1] = 40.0 g/dm^3 [1] *Correct conversion of 250 cm^3 into 0.250 dm^3 by dividing by 1000* [1] *Correct answer* [1]

4–2 The concentration will halve to 20.0 g/dm^3 [1]. Doubling the volume halves the concentration [1].

4–3 mol KOH = concentration × volume = 0.110 mol/dm^3 × (18.6 ÷ 1000) = 0.00205 mol [1]; mol HNO_3 = mol KOH (*1:1 ratio in balanced equation*) [1]; concentration HNO_3 = mol ÷ volume = 0.00205 ÷ (25.0 ÷ 1000) = 0.08184 mol/dm^3 [1] = 0.0818 mol/dm^3 (3 sf) [1]

4–4 mol KOH = concentration × volume = 0.110 mol/dm^3 × (18.6 ÷ 1000) = 0.00205 mol [1]; mass KOH = moles × M_r = 0.00205 × 56 = 0.115 g [1]

5–1 mol NaOH = concentration × volume = 0.225 mol/dm^3 × (31.6 ÷ 1000) = 0.00711 mol [1] mol $C_2O_4H_2$ = $\frac{1}{2}$ × mol NaOH (*1:2 ratio in balanced equation*) [1] = 0.00356 mol; concentration $C_2O_4H_2$ = mol ÷ volume = 0.00356 ÷ (25.0 ÷ 1000) = 0.1422 mol/dm^3 [1] = 0.142 mol/dm^3 (3 sf) [1]

5–2 mol $C_2O_4H_2$ = concentration × volume = 0.142 mol/dm^3 × (25.0 ÷ 1000) = 0.00356 mol [1]; mass $C_2O_4H_2$ = moles × M_r = 0.00356 × 90 = 0.320 g [1]

Use of amount of substance in relation to volumes of gases (p. 28)

Quick questions

1 The greater the temperature of a gas, the greater its volume.

2 The greater the pressure of a gas, the smaller its volume.

3 24.0 dm^3

4 1 mole of C_5H_{12} forms 5 moles of CO_2. 5 × 0.25 dm^3 = 1.25 dm^3 of CO_2

Exam-style questions

5–1 mol $CuCO_3$ = mass ÷ M_r = 0.50 g ÷ 123.5 = 0.00405 mol [1]; mol CO_2 = mol $CuCO_3$ (*1:1 ratio in balanced equation*) [1]; volume CO_2 = moles × 24.0 = 0.00405 × 24.0 = 0.0972 dm^3 [1] = 0.097 dm^3 (2 sf) [1]

5–2 Mean volume = 79 cm³ [1] ± 1.5 cm³ [1]

5–3 % yield = 100 × (actual volume ÷ maximum theoretical volume) [1] = 100 × (0.079 ÷ 0.0972) = 81.3 % [1]

5–4 *Any one from:* not all of the copper carbonate thermally decomposed/ the copper carbonate cooled down and the reaction stopped/the reaction didn't go to completion/ the reaction was reversible [1]

Quantitative chemistry topic review (p. 30)

1–1 There is only one product/no undesired product, so none of the reactants are wasted [1]

1–2 Atom economy = 100 × (sum of the M_r of desired products ÷ sum of the M_r of all the reactants) = 100 × (8 × 64) ÷ ((4 × 120) + (11 × 32)) = 61.5 % [1]
Correct multiplication of M_r of SO_2 [1], correct multiplication of M_rs of reactants [1]

1–3 mol FeS_2 = mass ÷ M_r = 100 000 ÷ 120 = 833.3 mol [1]; mol SO_2 = 2 × mol FeS_2 (from the 4:8 ratio in the balanced equation) = 1666.7 mol SO_2 [1]; mass SO_2 = mol × M_r = 1666.7 × 64 = 106666 g = 106.7 kg [1] = 107 kg (3 sf) [1]

1–4 % yield = 100 × (actual yield ÷ max. theoretical yield) [1] = 100 × (95.0 ÷ 107) = 88.8 % [1]

2–1 mean volume = (23.55 + 23.40 + 23.45) ÷ 3 = 23.47 cm³ [1] ± 0.08 cm³ [1]

2–2 mol KOH = concentration × volume = 0.850 mol/dm³ × (23.47 ÷ 1000) = 0.01995 mol [1]; mol H_2SO_4 = $\frac{1}{2}$ × mol KOH (from 1:2 ratio in balanced equation) = 0.009975 mol [1]; conc. H_2SO_4 = mol ÷ volume = 0.009975 ÷ (25.0 ÷ 1000) = 0.399 mol/dm³ [1]

2–3 mol H_2SO_4 in 25.0 cm³ = 0.009975 mol [1]; mass H_2SO_4 = mol × M_r = 0.009975 × 98 = 0.978 g [1]

3–1 Volume ratio of NH_3 : CH_4N_2O = 2 : 1 [1]; mol CH_4N_2O = $\frac{1}{2}$ × mol NH_3 = 0.23 mol [1]

3–2 mol NH_3 = mass ÷ M_r = 5.00 g ÷ 17 = 0.294 mol [1]; mol CO_2 = mass ÷ M_r = 5.00 g ÷ 44 = 0.114 mol [1]; so, 0.294 mol of NH_3 will react with 0.147 mol of CO_2 [1].
There are fewer moles of CO_2, so CO_2 will run out before the NH_3. CO_2 is the limiting reactant [1]

3–3 concentration = mass ÷ volume = 0.75 g ÷ (400 ÷ 1000) [1] = 1.88 g/dm³ [1]

3–4 mol/dm³ = g/dm³ ÷ M_r = 1.88 ÷ 70 = 0.0269 mol/dm³ [1]

4–1 M_r $(NH_4)_2Cr_2O_7$ = (2 × 14) + (8 × 1) + (2 × 52) + (7 × 16) = 252 [1]

4–2 mol $(NH_4)_2Cr_2O_7$ = mass ÷ M_r = 7.56 ÷ 252 = 0.03 mol; mol Cr_2O_3 = 4.56 ÷ 152 = 0.03 mol; mol N_2 = 0.84 ÷ 28 = 0.03 mol; mol H_2O = 2.16 ÷ 18 = 0.12 mol
Correct calculation of all amounts [1]
Dividing each number of moles by 0.03 mol to give the whole number ratio [1]
Ratio $(NH_4)_2Cr_2O_7 : Cr_2O_3\, N_2 : H_2O$ = 1 : 1 : 1 : 4 [1]
Correct symbol equation: $(NH_4)_2Cr_2O_7 \rightarrow Cr_2O_3 + N_2 + 4\,H_2O$ [1]

Reactivity of metals (p. 32)

Quick questions

1
- Cu_2O
- BaO

2 Vanadium/V is oxidised as it gains oxygen.

3 The more reactive metal (Al) takes the place of the less reactive metal (Ag) in a compound.

4 Lead has been reduced as it has lost oxygen.

5 Oxidation is the loss of electrons. Reduction is the gain of electrons (OIL RIG)

Exam-style questions

6–1 Carbon, C, is being oxidised as it is gaining oxygen [1]

6–2 Carbon is more reactive/higher in the reactivity series than iron [1] so is able to <u>displace</u> iron from its compound/iron oxide [1].

6–3 The Fe^{3+} ions gain electrons/ reduction is the gain of electrons [1]

6–4 $Fe^{3+} + 3e^- \rightarrow Fe$ [1]

6–5 Carbon dioxide is made up of small molecules [1] with only weak intermolecular forces/forces of attraction between the molecules [1]. These forces do not require a lot of energy to be overcome/at room temperature the molecules have enough energy to overcome these forces [1].

7–1 calcium + hydrochloric acid → calcium chloride + hydrogen [1]

7–2 Aluminium can displace zinc, so is more reactive than zinc [1]. Aluminium cannot displace magnesium so is less reactive than magnesium [1]. Any temperature increase between 14 and 25 °C [1].

7–3 Calcium has the greatest tendency to form positive ions [1] as it is the most reactive metal [1]

7–4 Metal atoms lose electrons to form positive ions [1]. Oxidation is the loss of electrons [1].

7–5 $Mg \rightarrow Mg^{2+} + 2e^-$ [1]

8–1 calcium + water → calcium hydroxide + hydrogen [1]

8–2 Calcium is more reactive than hydrogen [1] so it is able to displace hydrogen.

8–3 During the reaction calcium ions in CaO <u>gain electrons</u> to form calcium atoms / $Ca^{2+} + 2e^- \rightarrow Ca$ [1] (*this mark can be gained by the statement or the correct half-equation*).
Reduction is the gain of electrons so calcium is reduced [1]. Sodium atoms <u>lose electrons</u> to form sodium ions in sodium oxide / $Na \rightarrow Na^+ + e^-$ [1] (*this mark can be gained by the statement or the correct half equation*). Oxidation is the loss of electrons, so sodium is oxidised [1].

Reactions of acids (p. 34)

Quick questions

1 zinc sulfate and hydrogen gas; iron nitrate and water

2 filtration; crystallisation

3 An acid releases hydrogen ions in solution whilst an alkali releases hydroxide ions in solution. Acids have a pH less than 7, alkalis have a pH greater than 7.

4 Universal indicator will start off red. As sodium hydroxide is added to the hydrochloric acid the colour will change to orange, then green, then to blue and finally to purple.

5 A strong acid completely ionises when dissolved in water, but a weak acid only partially ionises when dissolved in water. A dilute acid has fewer acid molecules dissolved in a given volume of water, whilst a concentrated acid has more acid molecules dissolved in a given volume of water.

Exam-style questions

6–1 $ZnCl_2$ [1]

6–2 Zinc carbonate + <u>hydrochloric acid</u> [1] → zinc chloride + carbon dioxide + water [1].
Correct acid [1], *correct products* [1]

6–3 The hydrogen <u>ion/H^+ ion</u> [1]

6–4 Add hydrochloric acid to a beaker and add the zinc carbonate until it is in excess [1]. Once the reaction/ fizzing has stopped filter out the excess zinc carbonate [1] using a filter paper and funnel [1]. Pour the solution of zinc chloride into an evaporating basin [1]. Heat the solution until crystals start to form [1]. Leave the solution to cool, filter out the crystals and allow them to dry [1].

7–1 Magnesium atoms lose two electrons to form Mg^{2+} ions/$Mg \rightarrow Mg^{2+} + 2e^-$ [1]. Oxidation is the loss of electrons [1].

7–2 Hydrogen ions gain electrons to form hydrogen gas/$2H^+ + 2e^- \rightarrow H_2$ [1]. Reduction is the gain of electrons [1]

7–3 mol Mg = mass ÷ A_r

$= \frac{1.25}{24}$

= 0.05208 mol [1]

mol H_2 = mol Mg (1:1 ratio in the balancedvequation) [1]

mass H_2 = mol × M_r

= 0.05208 × 2

= 0.1042 g [1]

= 0.104 g (3 sf) [1]

7–4 volume = mol × 24.0

= 1.25 dm^3 [1]

8–1 sodium nitrate and water [1]

8–2 Hydrogen ion/H^+ ion [1]

8–3 Hydroxide ion/OH^- ion [1]

8–4 $H^+ + OH^- \rightarrow H_2O(l)$ [1] (or $H_3O^+ + OH^- \rightarrow 2H_2O$)

8–5 Red colour: At the start the pH = 1-2 as the concentration of H^+ ions is at its greatest (no OH^- has been added) [1]. As hydroxide ions are added they react with the hydrogen ions to form water / $H^+ + OH^- \rightarrow H_2O$ [1]. The concentration of H^+ ions decreases so the pH increases and the colour changes to orange [1]. Eventually all the H^+ ions have been used up/neutralised and the concentration of H^+ and OH^- are equal. pH = 7/the solution is neutral, and the colour is green [1]. As more hydroxide ions are added in excess their concentration increases and the solution becomes alkaline. The colour changes to purple [1].

9–1 Add sulfuric acid to a burette [1]. Use a glass pipette to measure out 25.0 cm^3 of potassium hydroxide solution into a conical flask [1]. Add 3 drops of the indicator to the potassium hydroxide (it will turn pink) [1]. Add the acid to the alkali and swirl until the indicator changes colour/goes colourless [1]. Record the volume added [1]. Repeat and find the mean volume of sulfuric acid added [1].
Allow method with potassium hydroxide in the burette and sulfuric acid pipetted into the conical flask.

9–2 Correct choice of titrations 2, 3 and 5 [1]; Average titration volume = 21.77 (cm^3) [1]; mol H_2SO_4 = concentration × volume = 0.135 × (25.0 ÷ 1000) = 0.003375 mol [1]; mol KOH = 2 × mol H_2SO_4 (*from 2:1 ratio in balanced equation*) = 0.00675 mol [1]; concentration KOH = mol ÷ volume = 0.00675 ÷ (21.77 ÷ 1000) = 0.3101 mol/dm^3 [1] = 0.310 mol/dm^3 (3 sf) [1]

10–1 Acid B is hydrochloric acid [1].

10–2 100 times more concentrated [1]. (for every 1 unit fall on pH scale, concentration increases by a factor of 10)

10–3 Acid A/ethanoic acid is a weak acid that only partially ionises when dissolved in water [1].
Acid B/hydrochloric acid is a strong acid that completely ionises in water [1].
This means the concentration of H^+ ions in the weak acid is less than the concentration in the strong acid, so the pH is higher (or vice versa) [1].

Electrolysis (p. 36)

Quick questions

1 copper and bromine;
zinc and oxygen;
hydrogen and chlorine;
hydrogen and oxygen

2 It is more reactive than carbon (so carbon is unable to displace it from its compound). OR The metal would react with carbon (in an unwanted reaction).

3 $2Cl^- \rightarrow Cl_2 + 2e^-$
Note that $2Cl^- - 2e^- \rightarrow Cl_2$ is also an acceptable way of representing this oxidation reaction.

4 Hydrogen ions/H^+ are discharged to form hydrogen gas, and hydroxide ions/OH^- are discharged to form oxygen gas.

5 $2Cl^- \rightarrow Cl_2 + 2e^-$ oxidation
Note that $2Cl^- - 2e^- \rightarrow Cl_2$ is also an acceptable way of representing this oxidation reaction.
$2H^+ + 2e^- \rightarrow H_2$ reduction
$4OH^- \rightarrow 2H_2O + O_2 + 4e^-$ oxidation
Note that $4OH^- - 4e^- \rightarrow O_2 + 2H_2O$ *is also an acceptable way of representing this oxidation reaction.*

Exam-style questions

6–1 Electricity is the flow of charged particles [1]. When molten, the ions are free to move and carry the charge [1].

6–2 Mixing aluminium oxide with cryolite lowers the melting point of the aluminium oxide [1]. Less energy is needed to melt the mixture, so the cost is less [1].

6–3 Aluminium ions/Al^{3+} ions are attracted to the negative electrode/cathode [1]. The aluminium ions gain electrons [1] to be discharged at the cathode as aluminium atoms [1] /$Al^{3+} + 3e^- \rightarrow Al$ oxide ions/O^{2-} ions are attracted to the positive electrode/anode [1]. The oxide ions lose electrons [1] to be discharged at the anode as oxygen gas [1] /$2O^{2-} \rightarrow O_2 + 4e^-$

6–4 They gradually burn away [1].

7–1 When dissolved in water the ions are already free to move and carry the charge [1].

7–2 Graphite contains delocalised electrons [1].

7–3 Copper ions/Cu^{2+} ions are attracted to the negative electrode [1] where they gain electrons to form copper atoms/metal [1].

7–4 Oxygen gas [1]

7–5 Hydroxide ions/OH^- ions [1]

7–6 A colourless gas would be formed at the negative electrode (instead of a red-brown metal) [1]. Sodium ions/Na^+ ions are more reactive than hydrogen so harder to discharge (or vice versa) [1]. Instead, hydrogen ions/H^+ from water are discharged. [1]. Hydrogen ions gain electrons to form hydrogen gas [1] /$2H^+ + 2e^- \rightarrow H_2$

8–1 Hydrogen ions/H^+ and sulfate ions/SO_4^{2-} [1]

8–2 hydrogen gas/H_2 [1]

8–3 $2H^+ + 2e^- \rightarrow H_2$
Correct reactants and products [1], correct balancing [1]

8–4 oxygen gas/O_2 [1]

8–5 $4OH^- \rightarrow O_2 + 2H_2O + 4e^-$ /
$4OH^- - 4e^- \rightarrow O_2 + 2H_2O$
Correct reactants and products [1], correct balancing [1]

8–6 moles = volume (in dm^3) ÷ 24.0 = (4 ÷ 1000) ÷ 24.0 [1] = 0.00016667 mol [1] = 0.000167 mol (3 sf) or 1.67×10^{-4} mol [1]

Chemical changes topic review (p. 39)

1–1 Copper gains oxygen during the reaction [1]
Note that 'the copper atoms lose electrons to form Cu^{2+} ions' would also be correct.

1–2 CuO [1]

1–3 $2Cu + O_2 \rightarrow 2CuO$
correct formulae [1], *correct balancing* [1]

1–4 Magnesium is more reactive than copper [1]. The reaction will be a displacement reaction [1]
magnesium + copper oxide → magnesium oxide + copper [1]

1–5 Add dilute sulfuric acid to a beaker and add the copper oxide until it is in excess [1]. Once the reaction has stopped filter out the excess copper oxide [1] using filter paper and a funnel [1]. Pour the solution of copper sulfate into an evaporating basin [1]. Heat the solution until crystals start to form [1]. Leave the solution to cool, filter out the crystals and allow them to dry [1].

2–1 hydroxide ion/OH^- [1]

2–2 sodium + water → sodium hydroxide + hydrogen [1]

2–3 Oxidation is the loss of electrons [1]. Sodium atoms lose an electron to form the sodium ion [1].

2–4 Lithium hydroxide solution any value greater than 7 [1]; nitric acid any value between 1 and 3 [1]

2–5 Lithium nitrate [1]

2–6 $H^+ + OH^- \rightarrow H_2O$ [1]

3–1 The strength of an acid describes how much the molecules ionise in water [1]. How dilute or concentrated an acid is describes the number of acid molecules dissolved in a certain volume of water [1].

3–2 1000 times more concentrated (*for every 1 unit decrease in pH the concentration of H^+ ions increases by a factor of 10*) [1]

3–3 mol NaOH = concentration × volume = 0.200 mol/dm^3 × (25.0 ÷ 1000) = 0.005 mol [1]; mol HCl = mol NaOH (*1:1 ratio in balanced equation*) [1]; concentration HCl = mol ÷ volume = 0.005 mol ÷ (19.5 ÷ 1000) = 0.25641 mol/dm^3 [1] = 0.256 mol/dm^3 (3 sf) [1]

4–1 nitric acid [1]

4–2 $Mg(NO_3)_2$ [1]

4–3 During the reaction Mg atoms lose electrons to form Mg^{2+} ions [1]. Oxidation is the loss of electrons, so magnesium is oxidised [1]. $Mg \rightarrow Mg^{2+} + 2e^-$ or $Mg - 2e^- \rightarrow Mg^{2+}$ [1]. Hydrogen ions/H^+ ions gain electrons to form H_2 molecules [1]. Reduction is the gain of electrons, so hydrogen is reduced [1]. $2H^+ + 2e^- \rightarrow H_2$ [1]

4–4 oxygen gas [1]

4–5 hydrogen gas [1]

Exothermic and endothermic reactions (p. 41)

Quick questions

1 Exothermic

2 Cold packs for sports injuries

3 Any from: the concentration of the acid, the volume of the acid, the amount/mass of magnesium added, the surface area of the magnesium, the starting temperature of the acid.

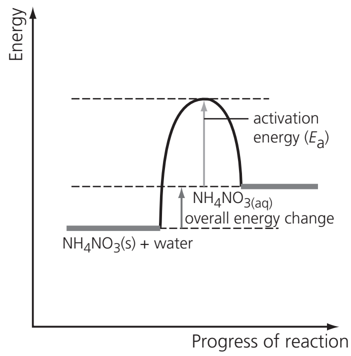
4 The minimum energy needed for particles to react.

5 The amount of (heat) energy absorbed during bond breaking is greater.

Exam-style questions

6–1 The process is endothermic [1] because the temperature decreases [1].

6–2

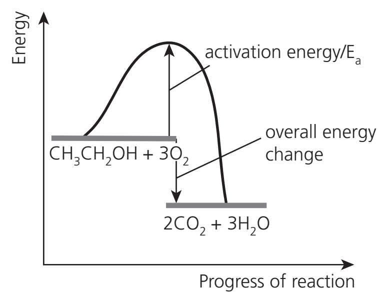


Correct overall shape for endothermic reaction [1]; *Activation energy correctly labelled* [1]; *Overall energy change correctly labelled* [1]. *Reactants and product lines do not need to be labelled.*

6–3 Measure out a given volume of water (10–50 cm^3) [1].
Measure the temperature of the water with a thermometer [1].
Measure out a given mass of ammonium nitrate using a mass balance [1].
Add the ammonium nitrate to the water and stir (with a glass rod) [1].
Measure the lowest or final temperature of the water [1].
Repeat using a different mass of ammonium nitrate [1].

7–1 Covalent bonds [1] because the elements bonded are all non-metals [1].

7–2



Correct overall shape for exothermic reaction [1]; *Activation energy correctly labelled* [1]; *Overall energy change correctly labelled* [1]; Reactants and product lines do not need to be labelled.

7–3 bonds broken = 5(412) + 348 + 360 + 463 + 3(496) = 4719 [1]; bonds formed = 4(743) + 6(463) = 5750 [1]; energy change = bonds broken – bonds formed [1] = 4719 – 5750 = –1031 kJ/mol [1]

7–4 More energy is released during bond formation than is absorbed during bond breaking [1], so there is an overall release of energy into the surroundings [1].

8–1 bonds broken = 4(463) + 2**X** = 1852 + 2**X** [1]; bonds formed = 4(463) + 496 = 2348 [1]; –208 = 1852 + 2**X** – 2348 [1]; so, **X** = 144 kJ/mol [1]
*Note that the answer can be achieved by ignoring the 4 O–H bonds that get broken and then reformed: –208 = 2**X** – 496, so **X** = 144.*

8–2 mol H_2O_2 = concentration × volume = 0.125 × (50 ÷ 1000) = 0.00625 mol [1]; mol $O_2 = \frac{1}{2}$ × mol H_2O_2 (2:1 ratio from balanced equation in **Figure 2**) = 0.003125 mol [1]

8–3 volume = mol × 24.0 = 0.003125 × 24.0 = 0.075 dm^3 [1] = 75 cm^3 [1] = 75.0 cm^3 (3 sf) [1]

Chemical cells and fuel cells (p. 43)

Quick questions

1 The electrodes used and the electrolyte used.

2 A cell carries out a chemical reaction to produce a voltage.
A battery is two or more cells connected together (in order to increase the voltage supplied).

3 Rechargeable cells and batteries use chemical reactions that can be reversed to reform the reactants.

4 A chemical cell will eventually run out of reactants and stop producing a voltage. A fuel cell has a continuous supply of reactants so can produce a continuous voltage supply.

5 Advantages: they produce a continuous supply of electricity; the only waste product from the fuel cell is water.
Disadvantages: they are very expensive; hydrogen is a flammable gas; hydrogen is often made from natural gas so is non-renewable; there are not many places to buy hydrogen gas from.

Exam-style questions

6–1 Potassium nitrate is the electrolyte. [1].

6–2 ions [1]; delocalised electrons [1]

6–3 The type of electrolyte or the concentration of the electrolyte [1]

6–4 0.00 V [1]. There is no difference in reactivity between the two electrodes [1], so the cell would not work.

6–5 Aluminium is the most reactive [1]. The greater the voltage reading the bigger the difference in reactivity of the two metals [1].

6–6 Lead and vanadium [1]
1.52 V – 0.47 V = 1.05 V [1]

7–1 Oxidation is the loss of electrons [1] and H_2 is losing electrons.

7–2 $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$
H_2O as product [1], correct balancing [1]

7–3 bonds broken = 4(412) + 2(463) = 2574 [1]; bonds formed = 3(436) + **X** = 1308 + **X** [1]; so 189 = 2574 – (1308 + **X**) [1]; and **X** = 1077 kJ/mol [1]

7–4 **For sustainable:** sustainable hydrogen can be produced by the electrolysis of water using renewable sources of electricity such as wind or solar power [1].
Against sustainable: if the electricity for the electrolysis of water is generated by burning fossil fuels, then this is not sustainable [1]. Natural gas is a fossil fuel, so producing hydrogen from natural gas is also unsustainable [1].
For zero carbon emissions: the only waste product from the fuel cell is water [1].
Against zero carbon emissions: a lot of carbon emissions are likely to be caused by the production of hydrogen gas [1].
A final conclusion regarding the sustainability and carbon emissions based on the arguments above is made [1].

Energy changes topic review (p. 45)

1–1 sodium hydroxide + hydrochloric acid → sodium chloride + water [1]

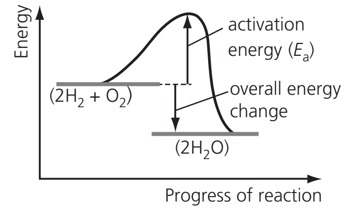
1–2 The reaction is exothermic [1] because the temperature of the solution rises [1]. Therefore, energy is being released into the solution during the reaction [1].

1–3 25.0 cm^3 of hydrochloric acid [1]. After this volume the temperature stops rising [1] so the neutralisation reaction is over.

1–4 A suitable method described, including at least one control variable for valid results:
Use a measuring cylinder to measure out a given volume of hydrochloric acid of a certain concentration [1].
Record the temperature of the hydrochloric acid solution using a thermometer [1].
Measure out and add a given volume of sodium hydroxide solution [1].
Stir the mixture with the thermometer and record the (maximum) temperature rise [1].
Repeat these steps using different concentrations of hydrochloric acid solution [1].
Control variables: the concentration of the sodium hydroxide solution; the volumes of the hydrochloric acid and sodium hydroxide solutions [1].

2–1 bonds broken = 2(436) + 496 = 1368 [1]; bonds formed = 4(463) = 1852 [1]; energy change = bonds broken – bonds formed [1] = 1368 – 1852 = –484 kJ/mol [1]

2–2



[1] *mark for showing exothermic reaction*; [1] *mark for activation energy*; [1] *mark for reactants and products.*

2–3 $2H_2 \rightarrow 4H^+ + 4e^-$ [1] and $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$ [1]

2–4 In rechargeable batteries the reactants eventually get used up [1]. Recharging the battery reverses the chemical reactions to reform the reactants [1]. Fuel cells have a continuous supply of reactants [1] so do not need recharging.

Rate of reaction (p. 46)

Quick questions

1 Either of the following:

$$\text{mean rate of reaction} = \frac{\text{quantity of reactant used}}{\text{time taken}}$$

$$\text{mean rate of reaction} = \frac{\text{quantity of product formed}}{\text{time taken}}$$

2 concentrations of reactants in solution; the pressure of reacting gases; the surface area of solid reactants; the temperature; the presence of catalysts.

3 The minimum amount of energy that particles must have to react.

4 Enzyme

5 Provide an alternative reaction pathway with a lower activation energy.

6 g/s or cm^3/s

7 Chemical reactions can occur only when reacting particles collide with each other and with sufficient energy.

8 Increases the frequency of collisions and so increases the rate of reaction.

Exam-style questions

9–1 magnesium + hydrochloric acid → hydrogen + magnesium chloride [1].

9–2 Use a lighted splint [1] and you will hear a (squeaky) pop if hydrogen is present [1].

9–3 mean rate of reaction = quantity of product formed/time [1]; mean rate of reaction = 60 ÷ 4.0 = 15 [1]; mean rate of reaction = 15 [1] cm^3/min [1]

10–1 carbon dioxide [1]

10–2 The mass would decrease [1] as carbon dioxide is lost to the air [1].

10–3 Powdered calcium carbonate will have a faster rate of reaction than lumps [1].
Powdered calcium carbonate has a higher surface area/higher surface area to volume ratio than lumps [1].
In powdered calcium carbonate there are more particles exposed at any time, so there is more chance of a collision in the same amount of time [1].
There would be more successful collisions in the same time for the powder than the lumps [1] and so a faster rate of reaction.

11–1 sulfur [1]

11–2 Use a measuring cylinder/volumetric pipette to measure out 25 cm^3 of sodium thiosulfate solution [1] into a conical flask.
Add 10cm^3 of hydrochloric acid [1].
Then either of the following methods for [2] *more marks*.
Method 1:
Put the conical flask on top of a cross and start the stopwatch. Time how long it takes until the cross can no longer be seen through the solution [1]. Repeat with different concentrations of sodium thiosulfate [1].
Method 2: Use a light meter [1] to monitor the amount of light transmitted through the sample [1].
Award [2] *more marks for ways of making the investigation valid:*
To make the investigation valid, you should use the same volume of acid and thiosulfate [1], the same conical flask/same depth of solution, complete the investigation at the same temperature and the same cross [1].

12–1 $2H_2O_2 \rightarrow O_2 + 2H_2O$ [1]

12–2 D [1]

12–3 E [1]

13–1 $2Li(s) + 2H_2O(l)$
$\rightarrow 2Li(OH)_2(aq) + H_2(g)$
Award [2] *for all four correct,* [1] *if two are correct*

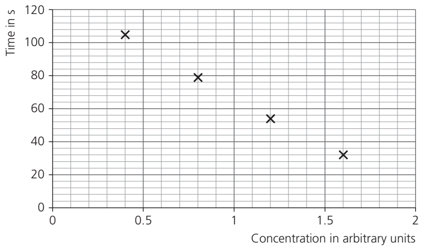
13–2 No turbidity/no solid made/no cloudiness [1] so cross would always be visible [1].

13–3 Gas syringe is more precise/more sensitive measurements than a measuring cylinder [1] Less gas is likely to escape using a gas syringe [1].

13–4 Repeat three times [1]; use only similar results/ignore outliers [1]; calculate a mean/average [1].

14–1 $1 \div 54 = 0.0185$ [1]

14–2



[1] *for correct* x*-axis and units,* [1] *for correct y-axis and units,* [1] *for sensible scale so that at least half of the graph paper is used for the plot,* [1] *correct plotting of the data*

14–3 There needs to be at least five data points for a trend [1]; therefore, you cannot be sure of the pattern in the data and so a line of best fit cannot be drawn with any certainty [1].

15–1 time [1]

15–2 volume of CO_2 in cm^3 [1]

15–3 draw a tangent to the curve at 18 s [1]; calculate change in volume; e.g. 55 – 5 = 50 cm ^{3} [1]; calculate change in time; e.g. 45 – 0 = 45 s [1]; calculate the gradient = 50 ÷ 45 = 1.11 [1]; units cm^3/s [1]

15–4 The line would be less steep [1] as the rate of reaction would be slower as there are less acid particles available for a collision at any one time [1]. They would have the same shape as the original line [1] as the rate of reaction will slow as time goes on. This is due to the reactant particles being used up during the reaction and so reducing the number of successful collisions in any one time [1]. The reaction would finish (shown as a horizontal line) with half the volume of gas compared to the original [1]. This is because there are half as many acid particles available for collision (calcium carbonate is in excess so does not limit the amount of gas made) [1].

16–1 $Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$
State symbols are optional.
correct formulae of reactants [1]; *correct formulae of products* [1]; *correct balancing* [1].

16–2 Zinc is losing electrons and is being oxidised [1]; The acid is gaining electrons and being reduced [1]; Both oxidation and reduction are happening in the same chemical reaction. As zinc is being oxidised while the acid is being reduced in the same reaction, it must be an example of a REDOX reaction [1].

16–3 Increasing the surface area (surface area to volume ratio) increases the rate of reaction [1]. There are more reactant particles available for a collision at any one time [1]. There will be more collisions and therefore more successful collisions in the same time [1].

Reversible reactions and dynamic equilibrium (p. 49)

Quick questions

1 $\rightleftharpoons$

2 Endothermic

3 When a reversible reaction happens in a closed container and the forward and reverse reactions happen at the same rate.

4 Le Chatelier's Principle

5 Yield decreases.

6 Moves to the side with the lowest number of moles of gas/lowest pressure.

Exam-style questions

7–1 Ammonium chloride $\rightleftharpoons$ hydrogen chloride + ammonia [1]

7–2 Endothermic [1]

7–3 A chemical reaction where the products of the reaction can react to produce the original reactants [1].

8–1 From blue [1] to white [1].

8–2 Ionic bonding [1] electrostatic forces of attraction [1] between oppositely charged ions [1].

8–3 White anhydrous copper sulfate [1] could be put into a liquid [1]. If there is a colour change from white to blue, this shows that water is present [1].

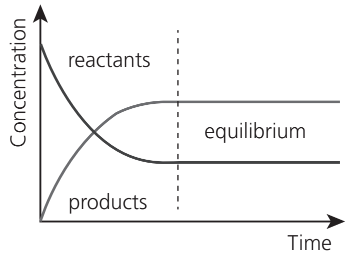
9–1 $HCl + NH_3 \rightleftharpoons NH_4Cl$
Award marks as follows: correct reactants [1]; correct products [1]; reversible arrows [1].

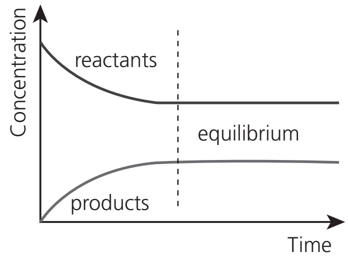
9–2 +176 KJ/mol [1]

9–3 The system would oppose the change [1]. It would try to reduce the temperature [1], favouring the backwards reaction [1] and increasing the amount of reactants [1].

10–1 $A + B \rightleftharpoons C + D$
Award marks as follows: correct reactants AND products [1]; reversible arrows [1].

10–2 Either of these graphs:





Award marks as follows:
concentration on *y*-axis [1]; time on *x*-axis [1]; products starting at (0,0) and then increasing [1]; reactants starting from more than 0 and decreasing [1]; the concentration of reactants and products not changing at the same time [1].

11–1 $N_2 + 3H_2 \rightleftharpoons 2NH_3$ [1]

11–2 The system will oppose the change/try to reduce the temperature [1]. It will favour the endothermic/backwards reaction [1] and reduce the yield of ammonia [1].

11–3 Fe/iron [1]. No effect on position of equilibrium/yield of ammonia [1] as the rate of both the forwards and backwards reactions are increased by the same amount [1].

12–1 $101\,000 \times 200 = 20\,200\,000$ Pa [1] = 2.02×10^7 Pa [1]

12–2 As temperature increases, percentage of ammonia/yield of ammonia decreases (or inverse) [1]. As pressure increases, percentage of ammonia/yield of ammonia increases (or inverse) [1].

12–3 High temperatures favour the endothermic reaction and would reduce the yield of ammonia [1]. So, for the highest yield, low temperatures should be used [1]. However, low temperatures cause a slow rate of reaction [1]. So, a compromise temperature is used so that an economical amount of ammonia is made in a reasonable amount of time [1].
High pressures favour the product side as there are the fewest moles of gas [1]. So, for the highest yield, high pressure should be

used. However, high pressures are expensive to maintain and so a compromise pressure is used [1].

13–1 atom economy = relative formula mass of desired product from the equation/sum of the relative formula masses of all the reactants in the equation × 100 [1]
100 % [1]

13–2 $NH_3(g) \rightarrow NH_3(l)$ [1]

13–3 The system will oppose change/try to use up the reactants [1], favouring the forward reaction [1]. This would increase the yield of ammonia [1].

14–1 $2SO_2 + O_2 \rightleftharpoons 2SO_3$
Award marks as follows: correct reactants [1]; correct products [1]; correct balancing [1]; reversible arrows [1].

14–2 Exothermic [1]

14–3 The system will oppose the change/try to increase the temperature [1]. It will favour the exothermic/forwards reaction [1] and increase the yield of sulfur trioxide [1].

15–1 White solid at the bottom of the tube 'disappearing' [1], white solid appearing further up the test-tube [1].

15–2 The system will oppose the change/move to the left [1]. This is the side with the fewest number of moles of gas [1].

15–3 Carry out in a fume cupboard/well ventilated room [1] Hydrogen chloride gas and ammonia gas are irritants [1].

16–1 $C(s) + H_2O(g) \rightleftharpoons H_2(g) + CO(g)$
Award marks as follows: correct state symbols for the reactants [1]; correct state symbols for the products [1].

16–2 The system will oppose the change/try to decrease the temperature [1]. It will favour the endothermic/forwards reaction [1] and increase the yield of hydrogen [1].

17–1 g/dm^3 or mol/dm^3 [1]

17–2 Students should draw a line vertically at the point where both lines level out [1].

17–3 $H_2 + I_2 \rightleftharpoons 2HI$
Award marks as follows: correct reactants [1]; correct products [1]; correct balancing [1].

18–1 The system will oppose the change/try to increase the temperature [1]. It will favour the exothermic/backwards reaction [1] and decrease the yield of chlorine [1].

18–2 The system will oppose the change/try to increase the pressure [1]. It will favour the products side [1] as there are more moles of gas [1].

19–1 $C_2H_4 + H_2O \rightleftharpoons C_2H_5OH$
Award marks as follows: correct reactants [1]; correct product [1]; reversible arrows [1].

19–2 atom economy = relative formula mass of desired product from the equation ÷ sum of the relative formula mass of desired product from the equation × 100 [1]
There is only one product [1] so all the atoms from the reactants are in the product [1].

19–3 The system will oppose the change/try to decrease the temperature [1]. It will favour the endothermic/backwards reaction [1] and decrease the yield of ethanol [1].

19–4 The system will oppose the change/try to decrease the pressure [1]. It will favour the product side [1] as there are fewer moles of gas [1].

20–1 $H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g)$ [1]

20–2 The system will oppose any change [1] but as there are the same number of moles of gas on each side of the equation the equilibrium position will not change [1].

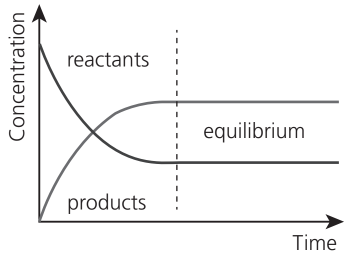
21–1 $CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$ [1]

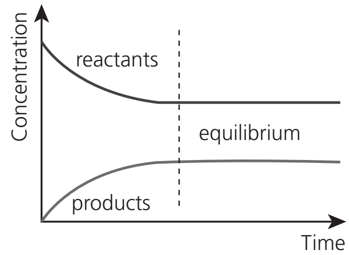
21–2 The reaction is in equilibrium. The forward reaction is exothermic, therefore high temperatures [1] favour the forward reaction and the reaction mixture contains a higher proportion of product molecules [1]. Since the forward reaction produces fewer molecules low pressures favour it [1] and the proportion of product molecules in the reaction mixture increases [1].

21–3 Although low temperatures give the highest yield [1] the rate of reaction is slower [1]. It takes a longer time [1] to make the methanol. A compromise is reached in which the equilibrium mixture contains an acceptably high proportion of methanol molecules (yield) in a suitable time. High pressures increases the rate of both forward and back reactions [1]. However, high pressures reduce the proportion of methanol in the reaction mixture and are expensive and difficult to maintain [1]. A compromise pressure is used to give a good yield in a reasonable time [1].

22–1 $N_2O_4 \rightleftharpoons 2NO_2$
Award marks as follows: correct reactants and products [1]; correct balancing [1]; reversible arrows [1].

22–2 *Accept either of these graphs:*





Award marks as follows: correct labels of the axis [1]; NO_2 starting at (0,0) and then increasing [1]; N_2O_4 starting from more than 0 and decreasing [1]; the concentration of reactants and products not changing at the same time [1].

22–3 See graphs in 22–2 [1]

23–1 As temperature increases, the percentage of SO_2/yield decreases [1]. A reversible reaction at equilibrium, will oppose change [1]. So, the reaction must be exothermic in the forward direction [1].

24–1 Equilibrium has been reached [1] the concentration of each chemical is constant/stays the same [1] rate of the forward reaction is the same as the rate of the reverse reaction [1].

24–2 A system at equilibrium will oppose change [1] by removing the ammonia, the forward reaction will be favoured [1] so more ammonia/greater yield of ammonia will be made [1].

25–1 C [1] as there is 0 concentration at the start [1]; A [1] as there is some hydrogen at the start of the reaction this must be A or B [1]; there is a greater amount of hydrogen/more moles of hydrogen in the balanced symbol equation so it must be A [1].

25–2 A = 2.75 mol/dm^3 [1]; B = 2 mol/dm^3 [1]; C = 1.5 mol/dm^3 [1]

25–3 The concentration of each chemical at equilibrium would be the same [1]. Equilibrium would be reached quicker [1] initial gradients of the lines would be steeper as the rate of reaction would be quicker [1].

The rate and extent of chemical change topic review (p. 56)

1–1 Oxygen can relight a glowing splint. [1]

1–2 A catalyst is a chemical that speeds up a reaction [1] but does not get used up itself [1].

1–3 In this reaction you would observe bubbles/effervesces/fizzing [1] as a gas is being made [1].

1–4 A = reactants [1]; B = activation energy without catalyst; C = activation energy with catalyst [1]; D = products [1].

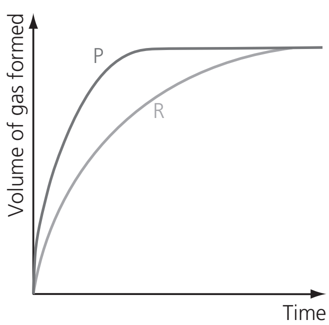
2–1 calcium carbonate + hydrochloric acid → calcium chloride + water + carbon dioxide
Award [1] *mark for correct reactants;* [1] *mark for correct products.*

2–2 To test for carbon dioxide, pass the gas through [1] limewater/calcium hydroxide solution [1] if the solution turns from colourless [1] to cloudy [1] it is carbon dioxide.

2–3 $2 \times 2 \times 6$ [1] = 24 [1] cm^2 [1]
$2 \times 2 \times 2$ [1] = 8 [1] cm^3 [1]

2–4 Both cubes have the same volume [1] and therefore the same amount of calcium carbonate and will eventually make the same amount of products. [1] However, P has a bigger surface area compared to cube R. [1] This means that there are more particles available for collision at any one time in P [1]. So, there will be more successful collisions in a given time [1] and this will give a faster rate of reaction for P [1].

2–5



[1] for correct label of *y*-axis; [1] correct label of *x*-axis. [1] for both lines having the same maximum volume of gas formed; [1] for P having a bigger gradient than R, both lines starting from the origin and not crossing each other [1].

3–1 $CH_4 + H_2O \rightleftharpoons 3H_2 + CO$
Award: [1] correct reactants; [1] correct products; [1] balancing; [1] reversible arrows

3–2 The system will oppose the change [1]. It will follow Le Chatelier's principle and try to reduce the temperature [1]. This will favour the forwards endothermic reaction [1] and the position of equilibrium will move to the right and so increase the yield of the products [1].

Carbon compounds as fuels and feedstock (p. 58)

Quick questions

1 Crude oil is a finite resource found in rocks. It is a mixture of hydrocarbons.

2 A compound containing ONLY hydrogen and carbon atoms.

3 A family of (organic) similar chemicals. Each successive member differs by $-CH_2-$ and they have similar chemical properties as they have the same functional group.

4 C_nH_{2n+2}

5 Methane, ethane, propane and butane.

6 <u>Fractional</u> distillation.

7 Release of energy (exothermic); the carbon and hydrogen are oxidised.

8 Carbon dioxide and water

9 Breaking down of long chain hydrocarbons to form smaller (more useful) shorter chain molecules.

10 Catalytic cracking and steam cracking

11 A homologous series of hydrocarbons containing at least one C=C (unsaturated).

Exam-style questions

12–1 5 [1]

12–2 Covalent bonding [1] made from a shared [1] pair of electrons [1].

13–1 C_2H_6 [1]

13–2 Fractional distillation [1]

13–3 Butane [1] as it's a large molecule/ has a higher relative formula mass [1] this means larger/bigger/greater forces of attraction/intermolecular forces/forces between molecules [1] so more energy is needed to overcome these forces. [1]

13–4 $C_{10}H_{22}$ [1]

14–1 Decane + oxygen → carbon dioxide [1] + water [1]
Each product can be in any order to obtain the marks.

14–2 Exothermic [1] heat lost to the surrounding/temperature increases/ energy is released in the form of heat and light [1]

14–3 Pentadecane has a higher boiling point [1], more viscous [1] and ignites less easily [1]. OR decane has a lower boiling point [1], less viscous [1] and ignites more easily [1].

15–1 To measure the temperature of the vapour [1] and use this to help identify the fraction being collected. [1]

15–2 Use another source of heat that is not a naked flame e.g. a heating mantle/water or oil bath [1]. This is safer because crude oil/fractions of crude oil are flammable/fuels [1] and exposure to naked flames can cause a fire [1].

16–1 *Any one of the following*:
Naked flame [1] Use a heating mantle [1]
Suck back [1] Remove the water trough before stopping heating/add a safety valve on the delivery tube [1]

16–2 Take a sample of the reactant and the product [1]. Add bromine (water) [1] to each and shake [1]. It will decolourise with the product [1] showing it was an alkene [1] but there will be no colour change with the paraffin [1].

16–3 *Award marks for any three of:*
Does not get used up [1]; Creates an alternative pathway with lower activation energy [1]; Speeds up the reaction [1]; Allows the reaction to happen at lower temperatures [1].

17–1 **X** = CH_4 [1]; **Y** = gas [1]; **Z** = 3 [1]

17–2 As the number of carbon atoms increases, so does the boiling point. [1]

17–3 Overall, The greater the number of carbon atoms, the higher the melting point [1]. There is a larger/ bigger/greater forces of attraction/ intermolecular forces /forces between molecules [1] so more energy is needed to overcome these forces (to melt the compound). [1]

18–1 Heat is used/endothermic reaction [1] to break down a chemical into simpler substances. [1]

18–2 In steam cracking the alkane is heated to high temperatures [1] to vaporise it/form a gas [1] and it is mixed with steam [1]. This requires a lot of energy.
In catalytic cracking, the alkane is heated to lower temperatures [1] and passed over a hot catalyst. [1] It therefore requires more energy to make steam and get the high temperatures needed for steam cracking than for catalytic cracking. [1]

18–3 $C_{10}H_{22} \rightarrow \underline{2}C_2H_4 + C_6H_{14}$ [1]

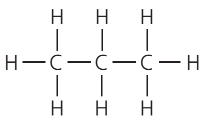
19–1 Plankton and other organisms died [1] in ancient times [1] these were buried in mud [1] and over time/ thousands of years became crude oil [1].

19–2 It is more than one chemical. [1]

19–3 Crude oil is a mixture of hydrocarbons [1]. The mixture is more useful when it is separated into smaller ranges or hydrocarbons [1].

19–4 A petrol fraction is a mixture of hydrocarbons [1] but the ratios of the mixture may be different for each fractionating column/crude oil separated [1]. Petrol is a mixture [1] that has been designed to be a useful product [1] and so is a formulation.

20–1



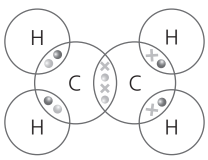
Award marks as follows: [1] for correct number of carbon and hydrogen atoms; [1] for correct number and arrangement of bonds

20–2 $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$
Award marks as follows: [1] for correct formula of reactants; [1] for correct formula of products; [1] for correct balancing
Allow multiples.

20–3 Oxidation [1]; Propane is a hydrocarbon, containing only hydrogen and carbon atoms [1]. The hydrogen gains oxygen [1] to make water [1] while carbon atoms gain oxygen [1] to make carbon dioxide [1].

21–1 $C_{10}H_{22} \rightarrow C_2H_4 + C_8H_{18}$ [1]

21–2



Award marks as follows: a shared pair of electrons between each hydrogen and a carbon atom [1]; two shared pair of electrons between the carbon atoms [1]. Hydrogen has 2 t in their outer shell and each carbon atom has 8 electrons in the outer shell [1].

21–3 Both ethane and ethene have covalent bonds [1], they are hydrocarbons/contain only hydrogen and carbon [1], they have the same number/2 carbon atoms [1], and they have a simple molecular structure [1]. Ethane only contains single bonds but ethene has one double bond [1]. Ethane has two more hydrogen atoms than ethene/ethane has 6 hydrogen atoms but ethene has 4 hydrogen atoms/less [1].

21–4 Polymers/medicine/starting point for many other chemicals. [1]

22–1 Crude oil is heated and turned to vapour/gas/evaporated [1]. The fractionating column is hot at the bottom and colder at the top [1]. The gases rise and when they reach their boiling point they condense [1] and are tapped/tapped off at different levels [1]. The smaller/lighter hydrocarbons are collected at the top and the heavier/larger hydrocarbons are collected at the bottom [1]. Each fraction is still a mixture but has a smaller range of carbon chain length [1].

23–1 Petrol would be lighter in colour than lubricating oil [1] as the carbon chain length/number of carbons is less. [1]
Any sensible/valid outline for comparing this colour. For example: take a sample of each, ensure they are the same depth and compare the colour of the two samples [1].
Petrol would be less viscous than lubricating oil [1] as the carbon chain length/number of carbons is less [1]
Any sensible/valid outline for comparing this colour. For example: take a sample of each and drop it onto a slope, the one that runs down the slope the quickest is petrol [1].
Petrol would be more flammable than lubricating oil [1] as the carbon chain length/number of carbons is less [1]
Any sensible/valid outline for comparing this colour. For example: take a small amount of each and try to ignite them. [1]

24–1 $C_{25}H_{52}$ [1]

24–2 $C_8H_{18} \rightarrow 2C_2H_4 + C_4H_{10}$
Correct reactant [1], correct products [1], balanced equation [1].

24–3 Hydrogen [1]

24–4 Fractional distillation [1]

Reactions of alkenes and alcohols (p. 62)

Quick questions

1 Hydrocarbons with carbon–carbon double bonds (C=C).
2 C_nH_{2n}
3 ethene, propene, butene and pentene.
4 Oxygen
5 By the addition of atoms across the carbon–carbon double bond (so that the double bond becomes a single carbon–arbon bond).
6 –OH
7 Methanol, ethanol, propanol and butanol
8 Fermentation
9 methanoic acid, ethanoic acid, propanoic acid and butanoic acid.
10 They only partially ionise in solution.

Exam-style questions

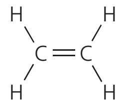
11–1 C=C [1]

11–2 C_3H_6 [1]

11–3 propene + steam → propane [1]

11–4 Addition [1]

12–1



Award marks as follows:
[1] for correct number of carbon atoms; [1] for correct number of hydrogen atoms; [1] for correct arrangement of bonds

12–2 As ethene contains only [1] hydrogen and carbon atoms [1]
Accept: By definition.

12–3 Contains C=C/contains a double bond/contains two fewer hydrogen atoms than the alkane with the same number of carbon atoms [1]

13–1 12 [1]

13–2 Butane will be more likely to undergo complete combustion whilst butene will be more likely to undergo incomplete combustion [1]. The butene flame is smokier than the butane flame [1]. Combustion of butene is more likely to make pollutants such as carbon/soot [1] and carbon monoxide. [1]
Accept reverse argument.

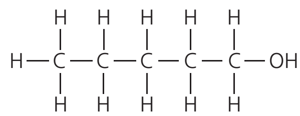
13–3 butene + oxygen → carbon dioxide + carbon + water
Award marks as follows: [1] correct reactants; [1] correct products

14–1 Alkene [1]

14–2 $C_5H_{10} + Cl_2 \rightarrow C_5H_{10}Cl_2$ [1]
Accept symbols of the product in any order e.g. $Cl_2C_5H_{10}$.

14–3 Hydrogenation/addition reaction [1]

14–4 [1]



Pentanol [1]

15–1 They contain oxygen/they contain more than just hydrogen and carbon. [1]

15–2 Carbon dioxide [1] + Water [1]
Accept these in any order

15–3 Oxidise [1] to form a carboxylic acid/methanoic acid [1]

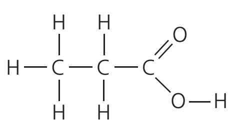
16–1 $C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$ [1]

16–2 *Any three of the following:*
Bubbles going through the limewater. [1]; Limewater would go from colourless [1] to cloudy [1]; Glucose solution and yeast could get foamy. [1]

16–3 To prevent any oxygen getting into the solution [1] to ensure anaerobic respiration [1] so that ethanol is made. [1]

16–4 This could be done through distillation [1]. Heat [1] the reaction mixture until the ethanol evaporates/it reaches 78 °C [1]. Collect the vapours and allow them to condense [1]

17–1 –COOH [1]
17–2 Ester/ethyl ethanoate [1]
17–3 $2C_2H_4O_2 + CaCO_3 \rightarrow Ca(C_2H_3O_2)_2 + H_2O + CO_2$
Award marks as follows:
[1] correct reactants; [1] correct products; [1] correct balancing
17–4 Gas to be bubbled through limewater [1], which will turn cloudy/milky/white if the gas is carbon dioxide [1]
18–1 $C_2H_4O_2$ [1]
18–2 $(2 \times 12) + (4 \times 1) + (2 \times 16)$ [1] = 60 [1]
Award both marks if correct answer is given without working.
18–3 $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$
Award marks as follows:
[1] correct reactant; [1] correct products; [1] equilibrium arrows
18–4 Any number between 3 and 6 [1]
19–1 C [1]
19–2 C [1]
19–3 E [1]
19–4 A [1]
19–5 D [1]
20–1 $C_2H_4\ (g) + H_2O\ (g) \rightarrow C_2H_6O\ (g)$
Award marks as follows: [1] for each correct state symbol
20–2 $C_2H_4 + H_2 \rightarrow C_2H_6$
Award marks as follows: [1] correct reactants; [1] correct products
20–3 Addition [1]
21–1 *Award marks as follows:*
Fermentation is: $C_6H_{12}O_6 \rightarrow 2CO_2 + 2C_2H_6O$ [1]
Hydration of ethene: $C_2H_4 + H_2O \rightarrow C_2H_6O$ [1]
Then any 3 of the following:
Fermentation is a more sustainable way to make ethanol as plants are a renewable resource [1] and less energy is required as the reaction mixture is only heated to a maximum of 40 °C [1]. However fermentation has a lower atom economy [1] and further processing is needed to get a pure product [1]. In contrast, hydration of ethene is quicker [1] and produces 100% pure product [1] with less waste as it has a 100% atom economy [1].
A judgement must be made for the final [1] mark. For example, to make pure ethanol hydration of ethene is best.
22–1 CH_2O_2 or HCOOH [1]
22–2



Award marks as follows:
Correct number of carbon, hydrogen and oxygen atoms [1]
Correct bonds [1]
22–3 Ethyl ethanoate [1] ester [1]
22–4 Butanoic acid will dissolve [1]. The acid group/-COOH partially ionises [1] and releases $H^+(aq)$ [1] making the solution acidic/having a pH of less than 7/given pH value from 3 – 6 [1].
23–1 Ethene is unsaturated/has C=C [1] this makes it more reactive than ethane [1]. Both molecules will undergo combustion, but ethene is more likely to undergo incomplete combustion [1] and have a smokier flame [1]. Ethene will undergo addition reactions/named addition reaction [1] whereas ethane will not. Ethene can be a monomer/make polymers/make poly(ethene) whereas ethane cannot [1].
24–1 C_5H_{10} [1]
24–2 Pentene contains two fewer hydrogen atoms than the pentane with the same number of carbon atoms. [1] Pentene has a double carbon carbon bond/C=C whereas pentane does not/only has single bonds. [1]
24–3 $C_5H_{10} + Br_2 \rightarrow C_5H_{10}Br_2$
Award marks as follows: [1] correct reactants; [1] correct products
25–1 $2Na + 2HCOOH \rightarrow 2NaCOOH + H_2$ [1]
25–2 Fill a test tube with the alcohol and trap the sodium beneath it/collect the gas via displacement [1] then use a lighted splint [1] you will hear a (squeaky) pop if the gas is hydrogen [1]. Keep the naked flame from the ethanol and sodium as they are both flammable [1]. Wear gloves because sodium and ethanol can damage the skin [1] and wear eye protection to prevent any splashes of chemicals into the eyes [1].
26–1 *Circle around O—H* [1]
26–2 Ethyl ethanoate [1] and water [1]
26–3 To speed up the chemical reaction [1] by providing an alternative reaction pathway [1] with a lower activation energy. [1]
27–1 Covalent bonding [1] between the atoms. Each bond is made from a shared pair of electrons [1]. The product has a simple molecular structure or two shared pairs [1].
27–2 ethanoic acid [1]; homologous series: carboxylic acid [1]
27–3 $C_2H_4O_2 + 3O_2 \rightarrow 2CO_2 + 2H_2O$ [1]
27–4 Mix yeast and sugar solution together [1]. Warm to about 30 °C and remove air [1] leave for a few days [1]. Distil the mixture [1] by heating and evaporating the ethanol [1] and then condense and collect the distillate [1]
N.B. accept labelled diagrams of fermentation (must include the conditions) and distillation (must include evaporation and condensation).

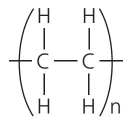
Synthetic and naturally occurring polymers (p. 68)

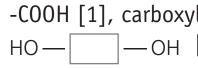
Quick questions

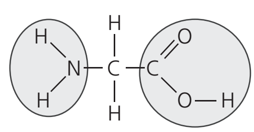
1 Many small molecules/monomers join together to form very large molecules/polymers.
2 Poly(butene)
3 Propene
4 Diol and a dicarboxylic acid
5 Polypeptide
6 Different amino acids in the same polymer chain.
7 Deoxyribonucleic acid
8 Proteins, starch and cellulose
9 Double helix

Exam-style questions

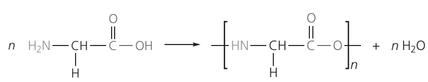
10–1 Ethene [1]
10–2



Award marks as follows:
[1] for single bonds only and correct number and type of atoms; [1] correct structure; [1] for bonds extending outside the brackets
10–3 The repeating unit has the same atoms as the monomer [1] because no other molecule is formed in the reaction. [1]
11–1 $C_2H_6O_2$ [1]
11–2 -COOH [1], carboxylic acid [1]
11–3  [1]
 [1]
11–4 H_2O [1] or H-O-H
Only accept molecular formula as displayed, and not 'water'.
12–1



Award [1] mark for each correct circling.
12–2



Award marks as follows: [1] correct formula of reactant; [1] correct formula of polymer; [1] correct balancing
Note: bond should extend through the brackets

12–3 Condensation polymerisation [1] as there are two products [1] one being the polymer/long chained molecule/polypeptide [1] and a small molecule/water [1].

12–4 Polypeptide [1] and water [1]

13–1 The atoms are held by covalent bonds [1] made of shared pairs of electrons [1]. The DNA molecule is a giant molecule/macromolecule/ made of many atoms [1]. Each polymer chain is made from different monomers/nucleotides [1]. The DNA molecule is two polymer chains [1], which are coiled together to make a double helix [1].

14–1 Both forms of polymerisation make a very large molecule [1]. Addition polymerisation uses one monomer but condensation polymerisation uses two monomers [1]. Addition polymerisation uses a monomer that is unsaturated/contains C=C. Condensation polymerisation monomers do not have C=C [1]. Addition polymerisation has only one product but condensation polymerisation has two products [1]. Examples of addition polymers include: poly(ethene), poly(propene), PVC [1]. Examples of condensation polymers include: polyamide, nylon, polyester [1]. *Do not accept natural polymers of starch, DNA or protein due to wording in the question.*

15–1 100 – (8 + 5 + 4 + 7) [1] = 76 % [1]

15–2 $\frac{5}{100} = \frac{1}{20}$ [1]

15–3 18.7 × 0.05 = 0.935 million tonnes
Award marks as follows:
[1] for using the injection moulding value from the LDPE pie chart; [1] for the calculation; [1] for correct answer

Organic chemistry topic review (p. 70)

1–1 Hydrocarbons contain only carbon and hydrogen [1].

1–2 Crude oil is made by ancient biomass/ plankton [1] being trapped in mud. [1]

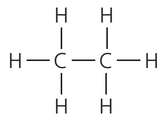
1–3 Crude oil is heated and evaporated [1]. The vapour enters a fractionating column [1]. The vapour rises up the column and each fraction will condense at a different height/temperature [1].

1–4 Methane has a lower boiling point than decane [1]. So, methane is a gas at room temperature, but decane is a liquid [1]. Methane is colourless whereas decane is brown/darker colour [1]. Methane is more flammable than decane [1].

1–5 $C_{10}H_{22} \rightarrow C_2H_4 + C_8H_{18}$ [2]

2–1 Fractional distillation [1]

2–2



[2]

2–3 C_nH_{2n+2} [1]

2–4 (2 × 12) + (6 ×1) [1] = 30 [1]

2–5 $2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$
[1] correct reactant; [1] correct products; [1] correct balancing *N.B. accept multiples*

3–1 $C_{11}H_{24} \rightarrow 3C_3H_6 + C_2H_6$ [2]
[1] for C_2H_6; [1] balancing

3–2 Propene contains a carbon–carbon double bond/C=C [1] and has two fewer hydrogen atoms compared to the similar alkane. [1]

3–3 Both propene and propane are hydrocarbons and will react with air in an exothermic reaction [1]. Propene is more likely to undergo incomplete combustion than propane [1]. This means that the propene flame is smokier [1] as carbon (soot) is made [1]. The propene flame releases less energy/ is a cooler flame as incomplete combustion is less efficient than complete combustion. [1]

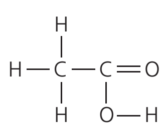
3–4 Add a few drops of bromine water to a sample of the liquid to be tested [1]. Shake well [1]. If the solution turns from red/orange/ brown to colourless then it was unsaturated and is the alkene [1]. If there is no visible change/colour remains, then the liquid was the alkane [1].

4–1 A mixture of glucose solution and yeast is added to the conical flask [1]. This is warmed to about 40 °C [1] and fermentation occurs. The enzymes in the yeast break down the sugar to make ethanol and carbon dioxide. [1]

4–2 There would be bubbles/froth in the conical flask [1] as the sugar reacted and made a gas [1]. There would be bubbles seen going through the limewater [1]. The limewater would go from colourless to cloudy [1] as it reacted with the gas showing that carbon dioxide was made [1]. As the reaction progresses, there would be a yeasty/bread smell [1].

4–3 Fractional distillation [1]

4–4 Carboxylic acid [1]



[1]

4–5 Ethyl ethanoate [1]

5–1 Many small molecules (monomers) [1] join together to form very large molecules (polymers) [1].

5–2 Condensation polymerisation involves monomers with two different functional groups [1]. When these types of monomers chemically react, they join together [1], they also lose small molecules such as water [1].

5–3 They are examples of condensation polymers [1] as they are made from monomers with two different functional groups [1] that chemically react to join them and they release a small molecule of water [1].

Purity, formulations and chromatography (p. 73)

Quick questions

1 A pure substance is a single element or compound, not mixed with any other substance.

2 Pure elements or compounds melt and boil at specific temperatures, they do not melt or boil over a range of temperatures.

3 A formulation is a mixture that has been designed as a useful product.

4 Mixing the components in carefully measured quantities to ensure that the product has the required properties.

5 Any from: fuels, cleaning agents, paints, medicines, alloys, fertilisers, foods.

6 Chromatography can be used to separate mixtures and can give information to help identify substances.

7 Ratio of the distance a substance moves compared to the distance moved by the solvent. Accept retention factor.

8 R_f = distance moved by substance ÷ distance moved by solvent

9 Mobile phase and stationary phase

10 It depends on the distribution of substances between the phases (mobile and stationary). The more attracted the substance is to the stationary phase, the less it moves and the smaller the R_f value. The more attracted the substance is to the mobile phase, the more it moves and the larger the R_f value.

Exam-style questions

11–1 In everyday language, a pure substance can mean a substance that has had nothing added to it [1]. It is unadulterated and in its natural state [1]. But chemically pure means that there is only one element or compound in the sample [1] and <u>not mixed</u> with any other substance [1]. So, pure milk is not chemically pure.

11–2 100 – (4.9 + 3.4 + 3.3 + 0.7) [1] = 87.7% [1]

11–3 Accept a value of 100.1 – 110 °C [1]; Milk is mainly water so will have similar boiling point to water [1] but as it is a mixture it will be slightly higher than water [1]. Milk is a mixture so it will boil over a range. [1]

12–1 Four (4) [1]

12–2 A [1]

12–3 D [1]

12–4 (11.5 – 2.0) ÷ (12 – 2.0) [1] = 0.95 [1]

12–5 If pigment moved *x* cm, then *x* ÷ 10 = 0.33, and *x* = 3.3 cm [1]. Since the sample started at the 2.0 cm line on the paper, the pigment would be at 5.3 cm [1]. This is pigment C [1].

13–1 Steel is an alloy, which is a mixture [1] each chemical has a particular purpose [1] and so is a special type of mixture called a formulation.

13–2 As the percentage carbon increases, so does the strength [1]. From about 0.8% there is very little change in strength as more carbon is added. [1]

13–3 800 MPa [1]
8×10^{10} Pa [1]

14–1 The tablet is not made just from aspirin and instead is a formulation/mixture [1] so will not have a definite melting point **or** will melt over a range. [1].

14–2 (9 × 12) + (8 × 1) + (4 × 16) [1] = 180 [1]

14–3 Formulation [1]

Identification of common gases (p. 75)

Quick questions

1 Hydrogen
2 Oxygen
3 Limewater
4 Turns red then white/is bleached.

Exam-style questions

5–1 endothermic [1]

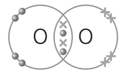
5–2 CO_2 [1]

5–3 Limewater goes from colourless [1] to cloudy. [1]

5–4 $CO_2(g) + Ca(OH)_2(aq) \rightarrow CaCO_3(s) + H_2O(l)$; [1] for each correct state symbol

6–1 $2H_2O_2 \rightarrow 2H_2O + O_2$
[1] correct reactants; [1] correct products; [1] correct balancing, accept multiples

6–2 Glowing splint [1] re-lights. [1]

6–3 

[1] for 4 electrons between the atoms; [1] for correct number of electrons in each atom

Identification of ions by chemical and spectroscopic means (p. 76)

Quick questions

1 Lithium
2 Na^+
3 Lilac
4 Orange-red
5 Copper
6 White
7 Copper(II)/Cu^{2+}
8 Carbon dioxide/CO_2
9 Iodide
10 Br^-
11 (Acidified) barium chloride (solution)
12 They are accurate, sensitive and rapid.
13 To analyse metal ions in solution.

Exam-style questions

14–1 A solid formed when two solutions mix. [1]

14–2 $CuSO_4 + 2NaOH \rightarrow Cu(OH)_2 + Na_2SO_4$ [1]

14–3 copper (II) hydroxide
OR copper hydroxide [1]

14–4 Put a sample of each solution to be tested in separate test tubes. [1]
Add sodium hydroxide solution. [1]
Mix the solutions. [1]
Make a note of the colour of the precipitate. [1]
If the precipitate is green then iron(II) is present. [1]
If the precipitate is brown then iron(III) is present. [1]

15–1 calcium/Ca^{2+} [1]

15–2 chloride/Cl^- [1]

15–3 Red-orange [1]

15–4 White [1] precipitate [1] formed.

15–5 White [1] precipitate [1] formed.

16–1 Mg^{2+} [1]

16–2 SO_4^{2-} [1]

16–3 $MgSO_4(aq) + 2NaOH(aq) \rightarrow Mg(OH)_2(s) + Na_2SO_4(aq)$
[1] for each correct state symbol.

16–4 Put a sample of the sample solution in a test-tube. [1]
Add dilute hydrochloric acid solution. [1]
Add dilute barium chloride solution. [1]
Mix the solutions. [1]
Make a note of the colour of the precipitate. [1]
If the precipitate is white then sulfate ions are present. [1]

17–1 Flame emission spectroscopy is more accurate than a flame test [1] as it can be difficult to judge the colour of a flame in a flame test [1]. Flame tests can only identify some metal ions [1] whereas the flame emission spectroscopy can identify the metal ion and give the concentration of the ion. [1]

17–2 Put a sample of the sample solution in a test tube. [1]
Add dilute acid solution. [1]
Observe [1]
If bubbles/fizzing/effervesces is observed then a gas is made. [1]
Collect and test the gas with limewater. [1]
If the limewater goes cloudy then the gas was carbon dioxide [1] and a carbonate was present.

17–3 No observable change/solution remains colourless, [1]
Acidified silver nitrate solution tests for halide ions, [1]
No halide ions are present. [1]
So, no chemical reaction will occur. [1]

18–1 Read off the spectrum: 590 nm [1] = 590×10^{-9} m [1] = 5.9×10^{-11} m to 2 s.f. [1] (*allow all three marks for correct answer to 2 s.f.*)

18–2 A line at: 460, 495, 610 and 670 nm [1]
Two lines close together at 590 nm [1]

18–3 The chloride ions from the acid would cause a white precipitate [1] and mask the accurate result. [1]

19–1 Solution A is sodium chloride. [1]

19–2 Solution B is sodium carbonate. [1]

19–3 Solution C is aluminium iodide. [1]
Solution D is magnesium chloride. [1]

Chemical analysis topic review (p. 78)

1–1 Oxygen: Glowing splint is re-lighted [1].
Chlorine: Damp blue litmus paper turns red then bleaches white [1].
Carbon dioxide: Limewater turns from colourless to cloudy [1].
Hydrogen: A lighted splint causes a squeaky pop sound [1].

2–1 Liquid [1]

2–2 A and B [1] are pure substances as each changes state/melts and boils at a specific temperature. [1]

2–3 A formulation is a mixture that has been designed to be a useful product [1]. The data in the table show that C changes state/melts and boils over a range and so it is a mixture [1]. Without knowing the name or use of substance C we cannot decide if it is a mixture that has been created for a specific purpose [1].

2–4 Li^+ [1] for symbol, [1] for charge

3–1 Mix a sample of the blue chemical with an acid in a test tube [1]. Put a bung with a delivery tube on the end of the test tube and allow the gas to blow through limewater [1]. If the limewater turns from colourless [1] to cloudy then the gas is carbon dioxide [1].

3–2 $Cu^{2+} + 2OH^- \rightarrow Cu(OH)_2$
[1] correct reactants
[1] correct products
[1] correct balancing

3–3 Get a clean, dry piece of nichrome wire (or damp splint) [1]. Put it into the sample for testing [1]. Turn the air hole on the collar of the Bunsen flame to give a blue/roaring/heating flame [1]. Put the sample into the blue flame [1] and observe the colour [1]. As this chemical contains copper ions the expected flame colour would be green. [1].

3–4 $CuCO_3$ [1]

The composition and evolution of the Earth's atmosphere (p. 80)

Quick questions

1 $\frac{4}{5}$

2 200 million years

3 4.6 billion years

4 Plants and/or algae produce oxygen by photosynthesis.

Exam-style questions

5–1 Nitrogen [1]

5–2 O_2 [1]

5–3 Any one of the following: carbon dioxide, water vapour and any noble gas i.e. He, Ne, Ar, Kr, Xe or Rn. [1]

6–1 Intense volcanic activity created the early atmosphere made of mainly carbon dioxide [1]. This level of carbon dioxide reduced, as the Earth cooled, and the oceans formed some carbon dioxide dissolved [1] and formed carbonate rocks [1]. Then as algae/plants developed photosynthesis occurred and the carbon dioxide levels reduced to present levels/0.04% [1].

6–2 It happened a very long time ago [1] and it is difficult to collect evidence. [1]

7–1 Carbon dioxide [1]

7–2 Volcanoes released gases which included nitrogen [1]. The levels of nitrogen gradually increased [1].

7–3 Carbon dioxide dissolved in the water [1], this increased the amount of dissolved carbon dioxide [1]. Animals that contain shells take in carbon dioxide in the air to make carbonate. When they die they sink and form sedimentary rocks [1]. These processes reduce and store the carbon dioxide [1].

8–1 carbon dioxide + water $\rightarrow$ glucose + oxygen
[1] correct reactants [1] correct products

8–2 Algae [1] and plants [1]

8–3 4:1 [1]

9–1 $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$ [1]

9–2 $6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$ [1]

9–3 $H_2O(l) + CO_2(g) \rightarrow H_2CO_3(aq)$; *award* [1] *mark for correct reactants and products;* [1] *mark for correct state symbols.*

9–4 $H_2CO_3 \rightleftharpoons 2H^+ + CO_3^{2-}$; *award* [1] *mark for correct reactant;* [1] *mark for correct product;* [1] *mark for equilibrium sign.*

9–5 pH decreased [1]; because carbonic acid released $H^+(aq)$ [1]

Carbon dioxide and methane as greenhouse gases (p. 81)

Quick questions

1 Maintain temperature on Earth high enough to support life.

2 Water vapour, carbon dioxide and methane.

3 Carbon dioxide and methane.

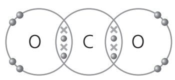
4 Increase in average global temperatures.

5 The total amount of carbon dioxide and other greenhouse gases emitted over the full life cycle of a product, service or event.

Exam-style questions

6–1 3.5×10^{-3} % [1]

6–2



Award [1] mark for two pairs of electrons between the O and C atoms; [1] mark for each double bond correctly drawn; [1] mark for 8 electrons in each outer shell.

6–3 Water vapour, carbon dioxide and methane gases [1] absorb longer wavelength/infrared radiation reflected from the Earth [1], but they do not absorb short wavelength radiation/UV/visible light from the Sun [1]. This keeps the Earth warm enough to support life as we know it [1].

6–4 The general trend is increasing amounts of carbon dioxide in the atmosphere [1]. During and after the industrial revolution more fossil fuels were combusted [1] and this has led to more carbon dioxide being released into the air. In addition, deforestation [1] has occurred to make way for farming and settlements/towns/cities/houses [1]. This has resulted in fewer plants photosynthesising [1] and removing carbon dioxide from the atmosphere [1].

6–5 *Any two of the following:*
Increasing the use of alternative energy sources [1] this will reduce the amount of carbon dioxide released into the air [1].
Use less energy [1]. This will reduce the amount of carbon dioxide released into the air [1].
Carbon capture and storage/collect carbon dioxide emissions and store underground in rocks [1], reducing the amount of carbon dioxide put into the atmosphere [1].
Carbon off-setting/plant trees [1]. This increases the amount of photosynthesis and removes more carbon dioxide from the air [1].
Use carbon neutral fuels such as bioethanol and biodiesel [1] where the carbon dioxide produced when they are being combusted matches the amount of carbon dioxide taken in by the plants as they grow [1].

7–1 *Any three of the following:*
Sea level rise [1]; More extreme weather, such as drought, storms [1]; The amount, timing and distribution of rainfall [1]; Reduction in availability of fresh water [1]; Changes to ecosystems [1]; Capacity of some regions to produce food due to changes in rainfall patterns, drought, flooding, higher temperatures, or type and number of pests in the region [1].

7–2 Animal farming produces a lot of methane [1]. This is from the digestive systems of cattle [1] as well as the decomposition of their manure [1]. Methane is a greenhouse gas [1]. This means that it traps the energy from the reflected/longer wavelength radiation from earth [1]. This leads to increase in global temperatures and this is a major contributor to global climate change [1].

7–3 Peer-review means that the work is looked at by other scientists/experts in the same field [1] before it is published in scientific journals or released to the public [1]. This is an important step as it checks that the science is valid [1] and not biased [1]. As there is still debate that climate change is down to human activities and the changes suggested have a big impact (economic and social), so we need to be confident that the science is valid [1].

8–1 °C [1]
8–2 The temperature of the Earth's surface has increased [1]. Except for the levelling off in 1940s [1] and 1970s [1].
8–3 2000–2009 [1]
8–4 Increase in greenhouses gases/ carbon dioxide/methane [1] by human activity [1] leading to increase greenhouse effect [1].
8–5 Evidence is difficult to collect over a long enough period to be sure that humans are the reason [1]. Climate models have a high degree of uncertainty [1].

Common atmospheric pollutants and their sources (p. 83)

Quick questions

1 Combustion of fuels.
2 Carbon, hydrogen and sulfur.
3 Solid particles and unburned hydrocarbons.
4 Carbon monoxide.
5 Respiratory problems and acid rain.
6 Global dimming and health problems for humans.

Exam-style questions

7–1 Carbon [1] and hydrogen [1]
7–2 5 [1]
7–3 methane + oxygen → carbon dioxide [1] + water [1]
7–4 Carbon monoxide [1] – toxic/ leads to death/reduces oxygen carrying capacity of the blood. [1] *or* Carbon/soot/particulate [1] – breathing problems [1]
8–1 $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$ [1]
8–2 During incomplete combustion/ limited supply of oxygen [1] solid particles [1] and unburnt hydrocarbons [1] can be released into the atmosphere. These can cause global dimming [1].
8–3 During incomplete combustion/ limited supply of oxygen [1] carbon is not fully oxidised and carbon monoxide is formed [1]. This gas reduces the oxygen carrying capacity of the blood/binds to red blood cells and stops them from carrying oxygen [1] and is toxic/ leads to death [1].
8–4 Fossil fuels can have some impurities of sulfur in them [1]. When the fossil fuel combusts, the sulfur also combusts/oxidises [1] to form sulfur dioxide [1]. During some combustion reactions, the nitrogen and oxygen in the air can be forced to react [1]. The oxides of nitrogen [1] and sulfur dioxide/ sulfuric acid dissolve into rain water and form acid rain [1].

Chemistry of the atmosphere topic review (p. 84)

1–1 The order is C – D – B – E – A. [3] all correct; [2] one in an incorrect order; [1] two in an incorrect order
1–2 Photosynthesis [1]
1–3 A gas that absorbs longwave infrared radiation reflected by the Earth [1]. This maintains the temperature high enough on Earth to support life as we know it [1].
1–4 Burning fossil fuels releases carbon dioxide into the atmosphere. *or* OR Deforestation has reduced the amount of photosynthesis and this has led to a build-up of carbon dioxide in the atmosphere [1].
2–1 Coal contains sulfur impurities [1]. When coal is used and combusted, the sulfur will also burn [1]. Oxides of sulfur are acidic and combine with natural rain water to make acid rain [1].
2–2 Fossil fuels are made of mainly <u>hydrocarbons</u> [1]. When they are used, they are combusted. If oxygen is limited, then <u>incomplete combustion</u> happens which produces carbon monoxide [1]. Carbon monoxide is a colourless, odourless <u>toxic</u> gas [1]. Carbon monoxide attaches to the haemoglobin/red blood cells [1] and reduces their ability to carry oxygen around the body [1]. This causes headaches, confusion and can cause death [1].
2–3 Carbon dioxide is needed by plants for photosynthesis [1]. Without this there would be no oxygen made in our atmosphere [1] and life that needs oxygen from the air would die out [1]. Carbon dioxide is a greenhouse gas [1]. This means that it traps longer wave radiation reflected from the Earth [1]. This natural process keeps the temperatures at a level to support life [1].
2–4 Carbon neutral fuels have a zero-net release of greenhouse gases into the atmosphere [1]. This would reduce the carbon footprint of many products and services as fuels are essential for producing goods and services [1].
3–1 Peer-reviewing means that the research is likely to be unbiased [1]. The research is likely to be of a high standard/draw accurate conclusions/and can be trusted [1].
3–2 Models that predict the climate are simplified [1]. Leading to a level of uncertainty of future global temperatures [1].
3–3 As temperatures rise, the polar ice caps will melt, and this will cause the sea level to rise [1]. Low lying areas will be flooded and coastal erosion will be increased [1]. More extreme weather events like drought and storms will increase in frequency [1]. The amount, timing and frequency of rainfall affected with some areas of the world getting more rain and some less [1]. The availability of fresh water is likely to reduce as droughts occur [1]. These changes will affect habitats and wildlife as well as the types of crops that can be grown in different parts of the world [1].
3–4 Global climate change can be reduced by:
Reducing emissions of greenhouse gases (carbon dioxide and methane) [1]. Carbon capture and storage [1]. Carbon neutral fuels [1].

Using the Earth's resources and sustainable development (p. 85)

Quick questions

1 We use Earth's resources for warmth, shelter, food and transport.
2 Processed finite resources provide energy and materials.
3 Development that meets the needs of current generations without compromising the ability of future generations to meet their own needs.
4 Water that is safe to drink.
5 Salty water can be made safe to drink by desalination, using techniques such as by distillation or reverse osmosis.
6 Industrial waste water must be treated to remove organic matter and harmful chemicals.
7 Bioleaching uses bacteria to produce leachate solutions that contain metal compounds.

Exam-style questions

8–1 One of: coal, oil, natural gas, fossil fuels. [1]

8–2 Any biomass such as wood, charcoal, biogas, biodiesel and bioethanol. [1]

8–3 Polyester could replace natural fibres such as cotton, silk or linen. [1]

8–4 Recycle the metal and use the same metal again to make aluminium products. [1]

8–5 Chemistry plays an important role in improving agricultural [1] and industrial processes [1] to provide new products [1] which could be focused on being sustainable.
Any one from:
Metals/Recycle metals. Fuels for transport/Use biofuels (e.g. biodiesel and ethanol made from crops). Electricity generation/Use renewable energy sources. [1]

9–1 Hydroelectric [1]

9–2 33+27+21 [1] = 81% [1]

9–3 Three times greater. [1]

9–4 21:6 [1]; 7:2 [1]

9–5 Sustainability is using resources to meet the needs of people today [1] without preventing people in the future using them [1]. Therefore, finite resources should be used responsibly as they can only be used once [1]. Only 13% of energy is from renewable resources. [1]. The majority of the energy resources are fossil fuels which are finite [1]. Therefore at the moment global energy use is not sustainable by the methods that we are using to generate the energy [1].

10–1 Drinking water should have sufficiently low levels of dissolved salts [1] and microbes [1].

10–2 Potable water is safe to drink [1]. Pure water only contains water molecules [1]. Although pure water is potable [1] other water is safe to drink and so potable [1]. Tap water is a mixture/formulation and is safe to drink [1], so is potable but as it contains other substances, it is not pure [1].

10–3 Rain water/fresh water is collected/ collected from the ground, lakes or rivers [1] as it has low levels of dissolved substances [1]. The water is passed through filter beds [1] to remove insoluble substances [1] and then sterilised/treated with chlorine/ozone/UV light [1] to kill the microbes [1].

11–1 Use a pH probe [1]; Use universal indicator solution/dip a universal indicator test paper into the sample of the water [1].

11–2 In pure samples, the boiling point is sharp, whereas mixtures boil over a range [1]. So, the tap water and river water are both mixtures [1]. Pure water will boil at exactly 100 °C which is 273K [1] and so the distilled water must be pure water [1].

11–3 Measure the mass of the water sample [1]. Boil away/evaporate all the water [1] and take the mass of any residue [1]. Calculate the percentage mass of the residue of each sample using: mass of residue ÷ mass of water × 100 [1] and compare the values of the samples.

12–1 Sewage must have organic matter [1] and harmful microbes [1] removed.

12–2 In industrial waste water there may be organic matter [1] and harmful chemicals [1].

12–3 Waste water from the sewage system is screened [1] and grit is removed [1]. It then undergoes sedimentation [1] to make sewage sludge and effluent [1]. The sewage sludge goes to the anaerobic digestor [1] and the effluent goes to the aerobic biological treatment [1].

12–4 11 000 000 000 ÷ 9000 [1] = 1 222 222 litres = 1.2×10^6 litres [2]

13–1 The main use of copper is electrical and electronic products [1]. Copper is an electrical conductor [1] as there are free moving [1] delocalised electrons [1] that can carry the charge. Copper is also ductile and malleable [1] as the planes/layers of atoms easily slide over each other [1] allowing many different shapes and wires to be made.

13–2 (11.4 ÷ 100) [1] × 18 000 000 = 2 052 000 [1] = 2 million metric tonnes [1]

13–3 Copper has a variety of uses and so we use large amounts of copper [1]. As a result, high-grade copper ores have all been used and we now have to extract copper from low grade ores and copper ores are therefore becoming scarce [1].

13–4 Uses lower grade ore [1] and avoids traditional mining methods of digging, moving and disposing of large amounts of rock, which is better for the environment [1]. It also uses less energy than traditional methods [1].

14–1 Uses bacteria [1] to produce leachate solutions/soluble metal solutions [1] that contain copper compounds [1].

14–2 $Fe + CuSO_4 \rightarrow FeSO_4 + Cu$
Accept:
$2Fe + 3CuSO_4 \rightarrow Fe_2(SO_4)_3 + 3Cu$
[1] correct reactants; [1] correct products

14–3 $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
Accept: $Cu^+(aq) + e^- \rightarrow Cu(s)$
[1] correct reactants; [1] correct products; [1] correct state symbols

14–4 Copper ions in the leachate are being reduced [1] as electrons are being gained [1] to form copper atoms.

Life cycle assessment and recycling (p. 88)

Quick questions

1 To examine the impact of a product on the environment throughout its life.

2 Metals are recycled by melting and recasting or reforming into new products.

3 Raw materials for metals and building materials are obtained by quarrying or mining.

Exam-style questions

4–1 Crude oil is the finite raw material used to make plastic shopping bags. [1]

4–2 Wood/trees are used to make paper shopping bags. [1]

4–3 There are three ways to use resources more sustainably: reduce, reuse and recycle [1].
When we reduce, we use less material for products. This reduces the amount of material used [1] and the cost of transport as it is lighter [1].
If we reuse shopping bags rather than just using them once and then throwing them away landfill will be reduced [1]. As we need to make fewer bags, the energy cost to make the product is reduced [1].
Recycling products uses less energy and reduces use of resources from the Earth [1].

5–1 100 − (59+5) [1] = 36% [1]

5–2 $\frac{59}{5}$ [1] = 11.8 [1] = 12 [1]

5–3 Plastic is non-biodegradable and will not break down in the environment [1], but it may break into smaller pieces and be eaten by animals/which can kill animals [1]. We are using more land for landfill. [1]

5–4 Use of water, resources, energy sources and production of some wastes can be quantified [1]. So, this aspect of the LCA is an objective process without bias [1]. Allocating numerical values to the impact of pollutant is less straightforward and requires value judgements/opinions [1], so LCAs can be biased as the process is not completely objective [1]. Selective or abbreviated LCAs can be made which evaluate a product [1]. These can be misused to support pre-determined conclusions, for example in support of claims for advertising purposes [1].

6–1 Glass bottles can be collected [1], cleaned and sterilised [1] and re-filled [1].

6–2 Glass can be recycled by crushing [1], melting [1] and reforming into a new shape [1].

6–3 Most of the energy to make glass comes from finite resources/limited resources/fossil fuels. [1]

6–4 $60 \times 25 \times 30$ [1] = 45 000 J [1] = 45 kJ [1]

Using materials (p. 89)

Quick questions

1 Corrosion is the destruction of materials by chemical reactions with substances in the environment.

2 Bronze is made of copper and tin.

3 Aluminium alloys have low density.

4 A polymer that softens and melts when heated.

5 Borosilicate glass is made from sand and boron trioxide

Exam style questions

6–1 Alloys are a mixture of a metal with small amounts of other metals or carbon. [1]

6–2 High carbon steel is strong but brittle [1]. Low carbon steel is softer and more easily shaped [1].

6–3 Rusting [1]

6–4 iron + water + oxygen → rust
Accept: iron + water + oxygen → hydrated iron(III) oxide
Award [1] mark for iron and oxygen, [1] mark for water, [1] mark for reactants on the left of an arrow and correct product on the right.

7–1 Anhydrous calcium chloride is a drying agent [1] that removes any water from the test tube [1].

7–2 Boiled water has no dissolved air/oxygen in it [1]. The oil floats on the surface of the water and stops any other air from dissolving back into the water [1].

7–3 For rusting to happen there needs to be iron, oxygen and water [1]. Test tube 3 is the only test tube with all the reactants needed for rusting [1].

7–4 Repeat the experiment [1]. Compare the results with other groups who did the same experiment [1]. To see if similar results were collected [1].

7–5 The mass of the iron nail could be taken at the start of the experiment [1]. Then the mass of the iron measured at the end of the experiment [1]/remove the rust and measure the iron at the end of the experiment. Then calculate the % mass change to allow direct comparison between the conditions without any error from initial mass differences for each nail [1].

8–1 A = 0 [1]; B = 58.5 [1]

8–2 Pure means containing only one type of substance [1]. 24 carat gold contains only gold atoms and so is pure [1].

8–3 Gold is usually alloyed with silver [1], copper [1] and zinc [1].

8–4 24 carat gold ring would have 2.5 g of gold [1]. 18 carat gold ring would have 1.875 g gold [1]. Therefore, there is 2.5 – 1.875 = 0.625 g [1] more gold in the 24 carat ring than the 18 carat ring.

8–5 Gold is an unreactive metal and therefore will not chemically react with the body or food and drinks [1]. This means that it is long lasting and durable [1]. As pure gold is so soft, it must be turned into an alloy [1] to make it strong and wear resistant as teeth will wear away pure gold [1]. Gold alloys will have a lower melting point [1] than gold and can be easily cast into any shape to make a bespoke filling and is easier to put into position in the mouth [1].

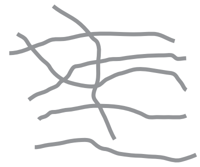
9–1 Sand is the raw material used to make glass. [1]

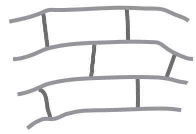
9–2 Soda-lime glass is made by heating [1] a mixture of sand, sodium carbonate [1] and limestone [1].

9–3 Borosilicate glass melts at a higher temperature than soda-lime glass. [1]

9–4 Contact lens must be transparent and bend the light/make a lens; both glass and some colourless polymers can do this [1]. Glass is more expensive than plastic [1]. Glass is more brittle than plastic and therefore more likely to break and be less comfortable to wear [1]. Therefore, polymers are a better choice of material to make contact lens [1].

9–5 In thermosoftening polymers chains are not joined to each other, but they are tangled up [1]. This means they soften and melt when heated [1]. In thermosetting polymers, chains are joined together by covalent bonds known as crosslinks [1]. This means they do not soften or melt when heated [1].





Thermosoftening Thermosetting

[1] for each diagram

10–1 Fibre glass [1] with the matrix or binder made from a polymer [1] and the reinforcement or fragments made from glass fibres [1].
OR Concrete [1] with the matrix or binder made from a cement and water [1] and the reinforcement or fragments made from sand and crushed rock [1].
OR Composite wood [1] with the matrix or binder made from adhesives [1] and the reinforcement or fragments made from wood fibres [1].
OR Carbon fibre composites [1] with the matrix or binder made from a polymer [1] and the reinforcement or fragments made from carbon fibres and carbon nanotubes [1].

10–2 Natural wood is made from more than one material [1]. It is made from a matrix (lignin) with fragments (cellulose fibres) in it [1].

The Haber process and the use of NPK fertilisers (p. 91)

Quick questions

1 Ammonia is made in the Haber process.

2 The raw materials for the Haber process are nitrogen and hydrogen.

3 Iron is the catalyst used in the Haber process.

4 NPK fertilisers contain nitrogen, phosphorus and potassium.

5 Phosphate rock is treated with acid such as nitric acid, sulfuric acid or phosphoric acid to make soluble salts.

Exam-style questions

6–1 Gas mixture is cooled [1] so that the ammonia condenses [1] into a liquid, but the nitrogen and hydrogen do not condense. [1]

6–2 It takes energy and time, and therefore money [1] to make hydrogen and purify air to make nitrogen [1]. The Haber process does not make a high yield [1] and the unreacted hydrogen and nitrogen can be recycled, and used again to make ammonia [1]. This is cheaper than creating nitrogen and hydrogen from raw materials/new reactants and so money is saved.

6–3 $N_2 + 3H_2 \rightleftharpoons 2NH_3$ [1]

6–4 High temperatures (450 °C) [1]; High pressure (200 atmospheres) [1]; Iron based catalyst [1]

7–1 This fertiliser contains nitrogen [1], phosphorus [1] and potassium [1].

7–2 100 − (20 + 30 + 10) = 40% [1]

7–3 An NPK fertiliser is a formulation [1] of various salts [1] containing appropriate proportions/ percentages of the elements N, P, and K. [1]

7–4 Ammonia [1] + nitric acid [1]

8–1 Potassium chloride or potassium sulfate are also raw materials used to make fertilisers. [1]

8–2 Potassium sulfate is made when phosphate rock is reacted with sulfuric acid. [1]

8–3 Phosphate rock does not have all the elements in it that are needed to make fertilisers [1]. The phosphate rock does not have soluble phosphate compounds in it/ compounds are insoluble in water [1]. When the rock is treated with acids, such as nitric acid or sulfuric acid, then soluble phosphate compounds are made [1]. These soluble phosphate compounds can be combined with other chemicals to make a formulation and an NPK fertiliser [1].

9–1 Sulfuric acid [1]

9–2 Ammonium sulfate [1]

9–3 Use a pH probe (with a data logger) and monitor the pH [1]. Slowly add the acid [1] and swirl [1]. When it is pH 7 [1] then the neutralisation reaction is complete.
Alternatively, add 1 cm^3 amounts of the acid [1], and use a glass rod to take a small sample of the mixture [1] onto indicator paper [1]. Note the colour change and stop when the pH is 7/colour indicates neutral [1].

10–1 In industry, fertilisers are made on a much larger scale [1]. In the neutraliser, the ammonia which contains nitrogen [1] reacts with the phosphoric acid [1], which contains the phosphorus. Potassium chloride is added as a source of K [1]. Water is removed in the evaporator to form a solid [1]. The NPK formulation is turned into pellets in the granulator and coater to make the NPK fertiliser easier to store and use [1].

11–1 101 000 = 20 200 000 Pa [1] = 2.02×10^7 Pa [1]

11–2 $N_2\,(g) + 3H_2\,(g) \rightleftharpoons 2NH_3(g)$ [1]

11–3 The Haber process has less moles of gas on the product side [1]. By increasing the pressure the system will oppose the change [1] and favour the products, shifting the position of equilibrium to the products [1] and increasing the yield [1]. The graph shows that after 200 atms pressure, increasing the pressure has very little effect on the yield [1]. High pressures are more expensive due to high cost of energy and apparatus that can withstand high pressures, and also more dangerous than lower pressures [1]. So, a compromise pressure of 200 atmospheres is used.

11–4 As temperature increases, yield decreases [1]. The equilibrium system will oppose the change [1] and favour the endothermic reaction [1]. Therefore, the backwards reaction is favoured, reducing the yield and therefore must be endothermic. So, the forward reaction must be opposite and exothermic [1].

Using resources topic review (p. 94)

1–1 4 [1]

1–2 nitrogen + hydrogen $\rightleftharpoons$ ammonia [1] for correct reactants and products; [1] for $\rightleftharpoons$

1–3 The air provides the nitrogen for the Haber process. [1]

1–4 Hydrogen is made from a chemical reaction between natural gas/ methane [1] and steam [1].

2–1 water that is safe to drink [1]

2–2 Microbes are removed from rainwater to make it potable by sterilising/adding chlorine/adding ozone/treating with UV. [1]

2–3 Industrial waste water might have harmful chemicals in it [1], sewage and agricultural waste water could have microbes in it [1]; and both types of water may contain organic matter. These pollutants could make the water unsafe to drink [1].

3–1 2 100 000 000 ÷ 66 000 000 [1] = 31.8 [1] = 32 bags [1]

3–2 The life cycle assessment has four stages: extracting and processing raw materials [1], manufacturing and packaging [1], use and operation during its lifetime [1], disposal at the end of its useful life [1].

3–3 In the extracting and processing, plastic bags have a higher impact on the environment as the crude oil is finite and the extraction and processing of crude oil to make plastic can cause environmental problems [1]. However, paper bags are made from renewable trees and if the trees are farmed sustainably there will be replacement of the trees that are cut down [1].
In the manufacture of the bags, plastic bags require a lot of energy, which can cause pollution [1], whereas paper bags use a lot of water, which can cause water pollution [1]. So both have a negative impact on the environment.
Plastic bags weigh less, and therefore take less energy to transport, which is better for the environment. As plastic is waterproof and more durable it is more likely to be reused more often, and at the end of its life it can be recycled indefinitely [1]. Paper bags are more likely to be single use and there is a limit to their recycling. [1]. However, if the bags are put in landfill as paper can biodegrade this is better for the environment than plastic.
Overall plastic shopping bags have the lowest impact on the environment.

3–4 Thermosoftening plastics melt when they are heated [1]. Whereas thermosetting plastics do not melt when they are heated [1]. It is essential that the plastic used to make the plastic bag can melt so the plastic can be recycled and remoulded into new products [1].

4–1 $-92 \times 2 = -184$ KJ [1]

4–2 Exothermic [1]. As energy is released/the energy change is negative [1].

4–3 As temperature is increased, an equilibrium system will oppose the change [1]. For the Haber process this will favour the reverse, endothermic reaction and reduce the yield [1]. However, at low temperatures the rate of reaction is low [1]. This is because not enough collisions are favourable [1] and have enough energy to reach the activation energy level [1]. So, a compromise of 450 °C is used which gives an acceptable yield in an acceptable time[1].

4–4 17 tonnes of ammonia = $17 \times 1000 \times 1000$ = 17 000 000 g
17 000 000 ÷ 17 = 1 000 000 moles of ammonia [1]
Ratio 1 nitrogen : 2 ammonia [1]
So, 1 000 000 ÷ 2 = 500 000 moles of nitrogen [1]
500 000 × (14+14) = 14 000 000 g nitrogen = 14 tonnes of nitrogen [1]

Practice exam papers (p. 96)

Paper 1

1–1 KCl [1]

1–2 hydrochloric acid [1] and any one of potassium hydroxide, potassium carbonate, or potassium oxide [1]

1–3 $A_r = ((39 \times 93.3) + (41 \times 6.7)) \div 100$ [1] = 39.134 [1] = 39.1 [1]

1–4 Potassium chloride is an ionic compound [1] with strong forces of attraction between the oppositely charged ions in the lattice structure [1]. So it has high melting and boiling points [1] as a lot of energy is needed break the strong forces of attraction [1]. It conducts electricity when molten (or dissolved) as the ions are free to move and carry the charge [1].

2–1 Mendeleev positioned the elements based on their properties [1]. By placing argon before potassium, the properties of argon matched those of the other elements in the same column/Group. [1]. *or*: If argon was placed after potassium, its properties did not match those of the elements in the same column/ Group [1].

2–2 They all have 1 electron in their outer shells. [1]

2–3 Reactivity increases down Group 1. [1]

2–4 Hydrogen gas [1]

2–5 Any pH value greater than 7 [1] as the solution will contain hydroxide ions, making it alkaline [1]

2–6 The potassium atoms lose one electron to form K^+ ions [1]. Oxidation is the loss of electrons [1].

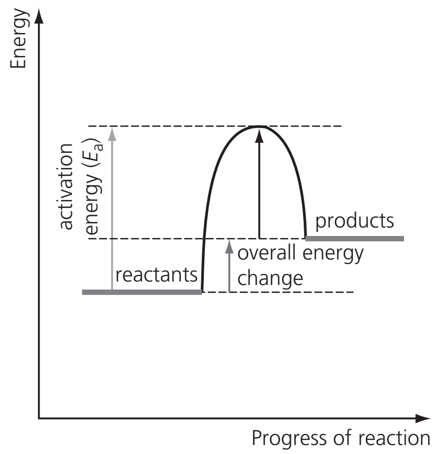
2–7 $K \rightarrow K^+ + e^-$ *or* $K - e^- \rightarrow K^+$

2–8 Reactivity decreases down the Group 7 [1]. When halogens react they gain one electron to fill their outer shell of electrons [1]. Going down the Group the number of electron shells increases, so the gained electron is further from the nucleus [1]. So the electron is less strongly attracted to the halogen nucleus [1], and it is harder for halogens further down the Group to gain an electron [1].

3–1 Total M_r of reactants = 108.5 [1]
% atom economy = $(32 \div 108.5) \times 100$ = 29.49 % [1] = 29.5 % (3 sf) [1]

3–2 actual mass = $(71 \div 100) \times 0.94$ kg [1] = 0.667 (kg) [1]

3–3



Correct overall shape for endothermic reaction profile [1] (*reactant and product lines need not be labelled*)
Labelled activation energy [1]
Labelled overall energy change [1]

4–1 Exothermic as heat energy is released [1]

4–2 Methane is made up of small molecules/is simple molecule [1]. It has weak intermolecular forces [1] that require little energy to overcome [1], so at room temperature all the molecules have enough energy to spread out and become a gas.

4–3 bonds broken = $(4 \times 413) + (2 \times 496)$ = 2644 kJ [1]; bonds made = $(2 \times 743) + (4 \times 463)$ = 3338 kJ [1]; energy change = bonds broken − bonds made [1] = 2644 − 3338 = −694 kJ [1]

5–1 They filter out ultraviolet light [1]

5–2 surface area = $50 \times 50 \times 6$ = 15 000 (nm^2) [1]; volume = $50 \times 50 \times 50$ nm = 125 000 (nm^3) [1]; SA : V ratio = 15 000 : 125 000 = 0.12 [1]

5–3 1.2 [1] (*ratio increases by a factor of 10 when width decreases by a factor or 10*)

5–4 Nanoparticles have different properties to those of the bulk material and it is not safe to assume that they are harmless [1]. TiO_2 nanoparticles may be able to enter into cells and cause damage, unlike bulk TiO_2 [1].

5–5 Any two from: titanium has a higher melting/boiling point; is denser; is harder; is stronger; is less reactive; forms coloured compounds; can act as a catalyst; can form ions with more than one charge [2]. *Allow the converse statements for potassium.*

5–6 protons: 22 [1]; electrons: 22 [1]; neutrons: 26 [1].

6–1 *Any one from:*
Fuel cells can run continually (as long as the fuel is supplied) [1].
Rechargeable batteries will go flat and need to be recharged [1].
Rechargeable batteries may release toxic chemicals when disposed of [1]. The only waste product from the fuel cell is water [1].

6–2 $2H_2 + O_2 \rightarrow 2H_2O$
Correct reactants and products [1], *correct balancing* [1]

6–3 mol CH_3OH = mass ÷ M_r = 5.12 ÷ 32 = 0.16 mol; mol O_2 = 10.24 ÷ 32 = 0.32 mol; mol CO_2 = 7.04 ÷ 44 = 0.16 mol; mol H_2O = 5.76 ÷ 18 = 0.32 mol
Correct amounts of reactants and products in mol [1]; Dividing through by 0.16 [1] gives 1 : 2 : 1 : 2 [1]; Balanced equation: $CH_3OH + 2O_2 \rightarrow CO_2 + 2H_2O$ [1]

6–4 *Answer should contain the following points; maximum* [6] *marks:*
Hydrogen fuel cells are more efficient (than direct methanol fuel cells) [1]. Hydrogen can be made using renewable energy sources [1]. Methanol can be made sustainably from fermentation [1]. Electrolysis requires a source of energy, whereas fermentation does not (or, at least, very little) [1]. Methanol can use existing infrastructure for refuelling, whereas hydrogen cannot [1]. Methanol does not need high pressure storage tanks to be used [1]. Hydrogen cells do not produce greenhouse gases (they are less polluting) [1].

7–1 When molten the ions are free to move [1] so can carry the charge [1]

7–2 In graphite each carbon atom forms three bonds [1]. One electron per carbon atom is delocalised [1]. These electrons can move and carry the charge through the graphite [1]

7–3 Magnesium [1]

7–4 Negative electrode: hydrogen (gas) [1]; positive electrode: chlorine (gas) [1]

7–5 Collect the gases produced [1] and test to identify them [1]

8–1 Suitable method:
Measure out 20 cm^3 of iron sulfate solution using a measuring cylinder and pour into a beaker/polystyrene cup [1]. Record the starting temperature of the solution using a thermometer [1]. Add 2 g of magnesium and stir with a glass rod; record the highest temperature reached [1]. Repeat with a different concentration of iron sulfate solution [1]. Use the same mass of magnesium and the same volume of iron sulfate solution in each experiment [1].

8–2 Initially as the concentration of iron sulfate increases, the temperature rise also increases [1]. After 4 g/dm^3 the temperature rise is constant [1].

8–3 The limiting reactant is completely used up before the other reactant [1]. After 4 g/dm^3 increasing the concentration of iron sulfate does not cause the temperature to rise further [1]. There must not be enough magnesium present to react with the extra iron sulfate to cause a further increase in temperature [1], so it has been used up.

8–4 concentration = 4.00 g/dm^3 ÷ 152 = 0.0263 mol/dm^3 [1]

9–1 Neutralisation [1]

9–2 Strength is a measure of how completely the acid is ionised in solution [1]. Concentration is a measure of the amount of acid dissolved in a given volume [1]

9–3 Suitable method including:
Fill a burette with sulfuric acid solution [1]. Using a glass pipette, measure out 25.0 cm^3 of sodium hydroxide solution into a conical flask [1]. Add 3 drops of the indicator to the sodium hydroxide (it will turn pink) [1]. Add the acid to the alkali and swirl until the indicator changes colour/goes colourless [1]. Record the volume added [1]. Repeat and find the mean volume of sulfuric acid added [1].
Allow method with sodium hydroxide in the burette and sulfuric acid pipetted into the conical flask.

9–4 mol H_2SO_4 = concentration × volume = 0.215 × (20.65 ÷ 1000) = 0.00444 [1] mol NaOH = 2 × mol H_2SO_4 = 0.00888 [1] Concentration NaOH = moles ÷ volume = 0.00888 ÷ (25 ÷ 1000) = 0.355 mol/dm^3 [1]

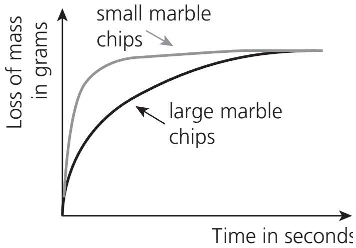
Paper 2

1–1 increases [1]

1–2 g/s [1]

1–3 A gas was made [1]

1–4



Starting from the origin and not crossing the small marble chip line of best fit [1] Smaller gradient [1] Line finishes at the same loss of mass [1]

2–1 Pencil is insoluble/will not run [1] and so will not interfere with the separation [1]

2–2 So that the colours dissolve into the solvent as it travels up the paper [1] rather than dissolving directly into the solvent in the jar [1].

2–3 To mark where the solvent got to [1] to calculate the R_f value [1]. If it is not done straight away, it might evaporate and you cannot see where it travelled to [1] or it might carry on travelling upwards once out of the tank [1].

2–4 Add a lid [1] to reduce the evaporation rate of the solvent [1], which makes the separation clearer [1]

3–1 [1]

$$\rightleftharpoons$$

3–2 Heat is needed to decompose ammonium chloride [1].

3–3 White solid [1] changes to colourless gas [1].

4–1 Cu^{2+} [1]

4–2 Green [1]

4–3 Copper(II) sulfate + sodium hydroxide → sodium sulfate + copper(II) hydroxide [1]

4–4 Blue [1]

5–1 Use a nichrome/platinum/inert wire [1]. The wire must be clean/dip the wire in concentrated acid and burn the acid off in a blue flame [1]. Put a sample of the chemical on the wire [1]. Hold the sample into a blue flame [1]. Note the colour of the flame [1]. If the flame is yellow/orange then sodium ions are present [1].

5–2 Positive ion [1]

5–3 Some flame colours can be masked. [1]

5–4 Lithium makes a crimson-red flame [1] and calcium makes a red-orange flame [1]. These look very similar in colour and it is difficult to tell them apart [1].

5–5 To prevent contamination [1]. To prevent false positives [1].

6–1 sodium + water → hydrogen + sodium hydroxide
[1] *correct reactants;* [1] *correct products*

6–2 *Any* [3] *from:* Metal floats [1]; Metal melts into a ball [1]; Metal moves across the surface [1]; Metal gets smaller [1]; Metal appears to disappear [1]; Head fizzing/see bubbles/effervescence [1].

6–3 Fill a test tube with water [1]. Trap the metal under the test tube [1]. The gas pushes out the water/collect the gas by displacement [1]. When the test-tube is full, put a bung on it [1]. Remove the tube from the water. Then remove the bung and quickly hold a lighted splint into the gas [1]. If you hear a pop the gas is hydrogen [1].

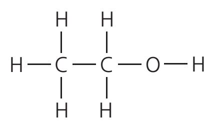
7–1 Crude oil cannot be replaced once it has been used. [1]

7–2 Crude oil is heated and evaporated until it forms a vapour [1]. It then is put into a fractionating column that is cooler at the top than at the bottom [1]. The vapour rises and condenses when it reaches its boiling point [1]. Each fraction of crude oil has a different boiling range [1]. Fractions have hydrocarbons with similar chain length. The longer the chain length, the greater the intermolecular forces and the higher the boiling point [1] so longer or heavier hydrocarbon fractions are collected at the bottom of the fractionating column as they have high boiling points [1].

7–3 $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ [1]

7–4 $C_{10}H_{22} \rightarrow C_2H_4 + C_8H_{18}$
[1] *for each product*

7–5



[1] *correct number and type of atoms;* [1] *correct structure*

8–1 $\frac{4}{5}$ [1]

8–2 To test for oxygen [1] use a glowing splint [1] that should re-light [1].

8–3 The early atmosphere was mainly carbon dioxide [1]. Then the percentage of carbon dioxide in the atmosphere decreased due to photosynthesis by algae and plants. [1]. Carbon dioxide also decreased due to the formation of sedimentary rocks and fossil fuels that contain carbon which captured/locked in the carbon [1].

9–1 The sample is put into a flame [1]. Light given out [1] is passed through a spectroscope [1]. The output is a line spectrum [1].

9–2 *Straight line from (0,0) passing through data points [1]*

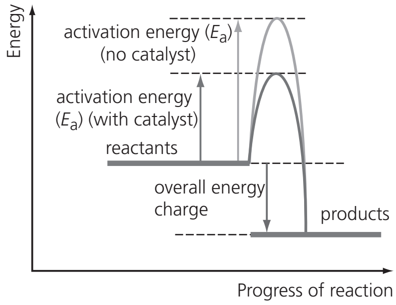
10–1 $2H^+ + 2e^- \rightarrow H_2$
[1] *correct reactants;* [1] *correct product;* [1] *correct balancing*

10–2 Damp [1] blue litmus paper [1] goes red then white/bleached [1]

10–3 Ions [1] must be free to move [1]

10–4 At the start, the universal indicator would be green as the salt solution/brine would be neutral [1]. As the electrolysis happens, OH^- ions/sodium hydroxide would be made [1] this will increase the pH [1] and the solution will turn blue/purple [1].

11–1



Award [1] mark for correct labels of *x* and *y* axes; [1] mark for reactants have higher energy than products; [1] mark for activation energy lower with catalyst than without

11–2 It is a transition metal [1] and has ions with different charges [1].

11–3 $Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s)$
[1] *reactants;* [1] *products;* [1] *balancing*
State symbols are not required.

11–4 $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$ [1]

11–5 Sensitive – small amounts of chemical can be analysed [1]. Informative – provide more information than traditional chemical tests such as concentration [1]. Rapid – much faster than traditional chemical tests [1].
Explanation is needed to get the mark.

12–1 copper ions (Cu^+) [1]

12–2 sulfate ions (SO_4^{2-}) [1]

12-3 copper sulfate ($CuSO_4$) [1]

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