Workbook answers

Chapter 1

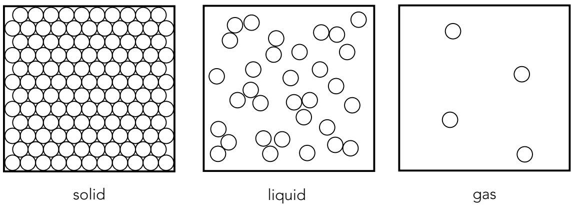
Exercise 1.1

1 A solid has a fixed **mass** and **volume** / **shape**.

A liquid has a fixed **mass** but its **shape** changes to that of the container in which it is placed.

A gas has no fixed **shape** or **volume**. A gas completely fills the container that it is in.

2



3
- **a** B
- **b** C
- **c** A
- **d** D

4 **A** freezing (solidification)

B melting

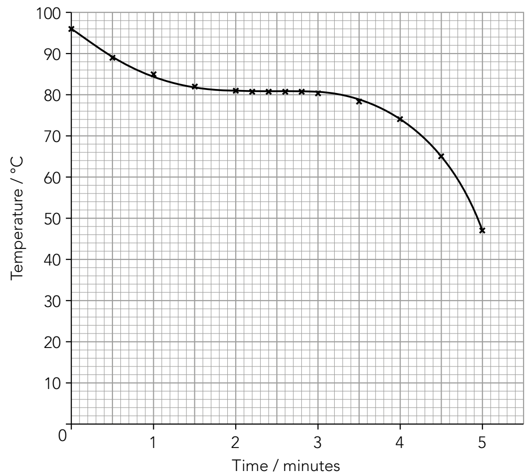
C condensation

D evaporation or boiling

5
- **a** radon
- **b** radon and nitrogen
- **c** nitrogen
- **d** cobalt
- **e** The sample of ethanoic acid is impure. The presence of impurities raises the boiling point of a substance.

Exercise 1.2

6



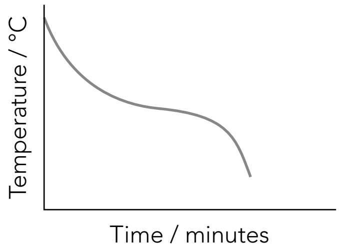
7 The student would need to use an oil bath (in place of the water bath) so that the higher temperature could be reached.

8 The substance is freezing (solidifying) / turning from liquid to solid.

9 The temperature stays constant because energy is being released as the substance solidifies / the molecules are giving out heat as they stop moving from place to place and become organised in a structured lattice arrangement / in the solid the molecules can only vibrate about fixed points / the heat released by the formation of new interactions (forces) between the particles keeps the temperature constant until all the substance is solid.

Chapter 1 continued

10 a



b The curve flattens but the temperature does not stay constant while the wax solidifies. This is because wax is a mixture of substances, not a pure compound.

11 a **A** solid

B solid and liquid (melting is taking place)

C liquid

D liquid and gas (boiling is taking place)

b 17 °C

c 115 °C

d The temperature remains constant until the change of state is complete.

e The melting point and boiling point are not those of water.

12 a The carbon dioxide is under pressure in the fire extinguisher.

b Hoar frost is a powdery **white** frost caused when solid **ice** forms from **humid** air. The solid surface on which it is formed must be **colder** than the **surrounding** air. Water vapour is deposited on a surface as fine ice **crystals** without going through the **liquid** phase.

c Evaporation: evaporation is a process that can take place at any temperature / it takes place at the surface of the liquid only / particles with enough energy can escape from the surface / these particles have sufficient kinetic energy to overcome the interactive forces between the particles and break free from the surface into the vapour (gas) phase / the rate of evaporation increases with temperature as the energy of the particles increases with temperature.

Boiling: boiling takes place at a specific temperature (the boiling point of the liquid) / bubbles of gas form throughout the liquid as particles gain sufficient energy to form a gas bubble / the gas bubbles then rise to escape the liquid and enter the gas phase (state) / the boiling point of a liquid changes with the atmospheric pressure / the lower the atmospheric pressure, the easier it is for the gas bubbles to form and the lower the boiling point.

Exercise 1.3

13 a The purple crystals are soluble in water, so the water begins to break up the crystals, and particles (ions) from the solid move into the water. This continues until all the solid dissolves. The particles then move through the liquid and spread out through the liquid until the solution is evenly coloured throughout.

b A shorter time – if the temperature were higher, the particles would be moving faster, as they would have more energy / the process of diffusion would take place more quickly.

14 a ammonium chloride → ammonia + hydrogen chloride

b NH_3(g) + **HCl(g)** → NH_4Cl(s)

15 a The smoke ring forms closer to the hydrochloric acid end of the tube because ammonia molecules diffuse (move) faster than hydrogen chloride molecules / this is because ammonia molecules have a lower relative molecular mass (M_r = 17) compared with hydrogen chloride molecules (M_r = 36.5) / ammonia molecules travel further in a given time.

b approximately 30 cm / approximately two-thirds of the way along the tube

c hydrogen > methane > oxygen > chlorine

d G must have a molecular mass greater than that of methane, but less than that of oxygen / it must have an M_r between 16 and 32.

Chapter 1 continued

16 The kinetic model states that the **particles** in a liquid and in a **gas** are constantly moving. In a gas, the particles are far apart from each other and their movement is said to be **random**. The particles in a solid are held in fixed positions in a regular **lattice**. In a solid, the particles can only **vibrate** about their fixed positions.

Liquids and gases are fluids. When particles move in a fluid, they can collide with each other. When they collide, they bounce off each other in **different** directions. If two gases or liquids are mixed, the different types of particle **spread** out and get mixed up. This process is called **diffusion**.

In gases at the same **temperature**, particles that have a lower mass move faster than those with higher mass. This means that the lighter particles will spread and mix more quickly. The lighter particles are said to **diffuse** faster than the heavier particles. When gaseous molecules diffuse, the rate at which they do so is **inversely** related to the relative **molecular** mass (M_r) of the gas.

Chapter 2

Exercise 2.1

1 Atoms are made up of three different particles:

- **protons**, which are positively charged
- **neutrons**, which have no charge
- **electrons**, which are negatively charged.

The negatively charged particles are arranged in different **shells** (energy levels) around the **nucleus** of the atom. The particles with a negligible mass are the **electrons**. All atoms of the same element contain the same number of **protons** and **neutrons**. Atoms of the same element with different numbers of **neutrons** are known as **isotopes**.

2 **a** 3

b 4

c 7

d $^{7}_{3}$Li

3 The electrons in an atom are arranged in a series of **shells** around the central nucleus. These shells are also called **energy** levels. In an atom, the shell **closest / nearest** to the nucleus fills first, then the next shell, and so on. There is room for:

- up to **two** electrons in the first shell
- up to **eight** electrons in the second shell
- up to **eight** electrons in the third shell.

(There are 18 electrons in total when the three shells are completely full.)

The elements in the Periodic Table are organised in the same way as the electrons fill the shells. Shells fill from **left** to **right** across the **rows** of the Periodic Table.

- The first shell fills up first, from **hydrogen** to helium.
- The second shell fills next, from lithium to **neon**.
- Eight **electrons** go into the third shell, from sodium to argon.
- Then the fourth shell starts to fill, from potassium.

4 **a** Mg (magnesium)

b F (fluorine)

c K (potassium)

5 **a** boron / 2,3

b phosphorus / 2,8,5

6

Beryllium	2,2
Magnesium	**2**,8,**2**
Calcium	2,8,**8**,**2**

Exercise 2.2

7 **a**

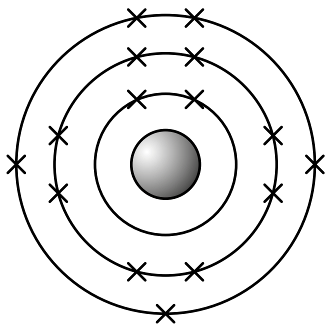
Atom	Proton number	Electronic configuration			
		1st shell	2nd shell	3rd shell	4th shell
A	2	**2**			
B	5	**2**	**3**		
C	13	**2**	**8**	**3**	
D	15	**2**	**8**	**5**	
E	19	**2**	**8**	**8**	**1**

b one (the atoms of element B)

c B and C

d 3

e



8 **a** 38

b 53

c 78

d 137 – 55 = 82

e Isotopes are different atoms of the same element that have the same proton number but different nucleon numbers.

Chapter 2 continued

9 **a** B

b E

c A and C

d B and D

10 **a** 2 protons, 2 neutrons, +2

b The gold atoms are packed together in a regular arrangement (lattice) / in layers / the atoms can only vibrate about fixed positions.

c This suggests that the atoms were largely empty space through which the α-particles passed.

d These α-particles had made direct hits on the nuclei of the gold atoms. They are repelled backwards because the nucleus of the atom is positively charged and so are the α-particles.

11 **a**

Isotope	Name of element	Atomic number	Mass (nucleon) number	Number of		
				protons	neutrons	electrons
$^{12}_{6}C$	Carbon	6	12	6	6	6
$^{14}_{6}C$	**Carbon**	**6**	**14**	**6**	**8**	**6**
$^{1}_{1}H$	**Hydrogen**	**1**	1	**1**	**0**	**1**
$^{3}_{1}H$	Hydrogen (tritium)	**1**	**3**	**1**	**2**	**1**
$^{31}_{15}P$	**Phosphorus**	15	31	**15**	**16**	**15**
$^{32}_{15}P$	**Phosphorus**	**15**	**32**	**15**	**17**	**15**
$^{127}_{53}I$	Iodine	**53**	**127**	53	**74**	53
$^{131}_{53}I$	**Iodine**	**53**	**131**	53	**78**	**53**

b The chemical properties of isotopes of the same element are the same because the number and arrangement of electrons in the isotopes are the same / the atoms of the isotopes all have the same number of outer electrons.

c percentage of rhenium-187 in natural sample = 100 – 37.4 = 62.6%

$$A_r \text{ of rhenium} = \frac{(185 \times 37.4) + (187 \times 62.6)}{100}$$

$$= \frac{6919 + 11\,706.2}{100} = \frac{18\,625.2}{100}$$

$$= 186.3$$

Chapter 3

Exercise 3.1

1

Compound	Mixture
The **elements** cannot be **separated** by **physical** methods.	The substances in it can be **separated** by **physical** methods.
The properties are **different** from those of the **elements** that went to make it.	The substances **present** still show the **properties** that they have when by themselves.
The elements are **combined** in a **definite** proportion by mass.	The substances can be **present** in **any** proportion by mass.

2 **a** compounds: distilled water, carbon dioxide, sodium chloride, copper sulfate

b mixtures: brass, lemonade, seawater, hydrochloric acid solution, air

3 **A** compound

B element

C mixture

D element

E compound

F mixture

4 sodium chloride is a white solid – different from the elements

sodium chloride dissolves in water – different from the elements, particularly sodium

sodium chloride is neutral in solution – again different from the two elements

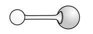
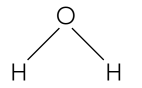
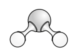
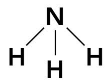
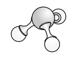
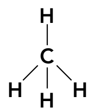
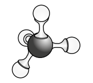
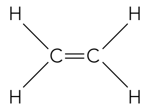
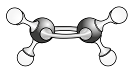
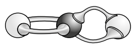
5 **a** iron is magnetic; sulfur is non-magnetic / iron is dark grey; sulfur is yellow / iron reacts with hydrochloric acid; sulfur does not react with acid

b the fact that the mixture continues to glow with heat shows that a chemical reaction is continuing to take place, and that it is an exothermic reaction

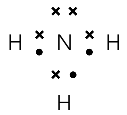
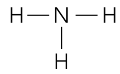
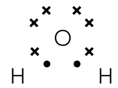
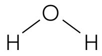
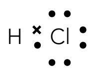
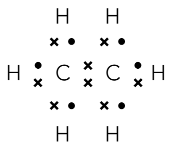
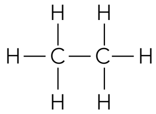
c the product is no longer magnetic / the product reacts with acid but gives a different product to that produced by the iron powder

Exercise 3.2

6

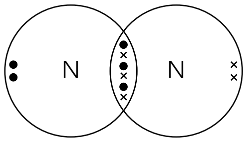
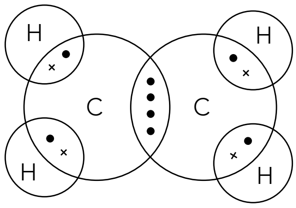
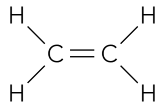
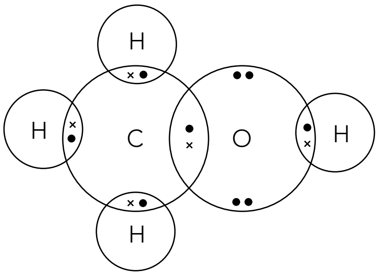
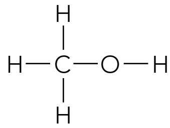
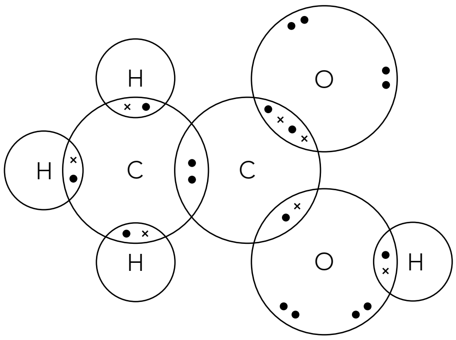
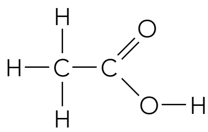
Name of compound	Formula	Displayed formula	Molecular model
Hydrogen chloride	**HCl**	H—Cl	
Water	H_2O		
Ammonia	$\mathbf{NH_3}$		
Methane	CH_4		
Ethene	$\mathbf{C_2H_4}$		
Carbon dioxide	$\mathbf{CO_2}$	O=C=O	

7

Molecule	Dot-and-cross diagram	Displayed formula
Ammonia (NH_3)		
Water (H_2O)		
Hydrogen chloride (HCl)		H—Cl
Ethane (C_2H_6)		

Chapter 3 continued

8

Molecule	Dot-and-cross diagram	Displayed formula
Nitrogen (N_2)		N≡N
Ethene (C_2H_4)		
Methanol (CH_3OH)		
Ethanoic acid (CH_3COOH)		

Note that in the answers to question **8**, the dot-and-cross diagrams are drawn differently to those in question **7**. Both methods are equally acceptable, but the drawings here have the advantage of showing the overlap between the atoms in covalent bonding.

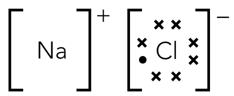
Chapter 3 continued

Exercise 3.3

9 **a** the sodium ion (2,8) has one less electron than a sodium atom (2,8,1)

b the chloride ion (2,8,8) has one more electron than a chlorine atom (2,8,7)

c



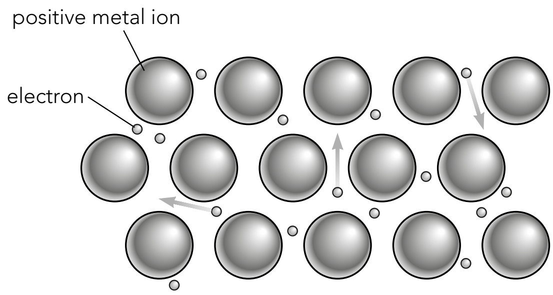
d electrostatic forces of attraction between ions with opposite charge

e NaCl has a high melting point because the forces between the ions are strong – it takes a large amount of heat to disrupt them

10

Property	Explanation
A solution of an ionic compound in water is a good conductor of electricity. These ionic substances are called electrolytes.	The ions in the giant ionic structure always have the same regular arrangement (see Figure 3.2).
Ionic crystals have a regular shape. The crystals of each solid ionic compound are the same shape. The angles between the faces of the crystal are always the same, whatever the size of the crystal.	Strong attraction between the positive and negative ions holds the giant ionic structure together. A lot of energy is needed to break down the regular arrangement of ions.
Ionic compounds have relatively high melting points.	In a molten ionic compound, the positive and negative ions can move – the ions can move to the electrodes when a voltage is applied.
A molten ionic compound (i.e. an ionic compound heated above its melting point) is a good conductor of electricity.	In a solution of an ionic compound, the positive metal ions and the negative non-metal ions can move – these ions can move to the electrodes when a voltage is applied.

11 **a**



b **i** the electrostatic forces between the metal ions and the delocalised sea of electrons are strong and so it needs a strong heat to overcome them

ii the fact that the layers can slide over each other without the bonding being broken means that a metal can be drawn out into wires

iii the layers of metal ions can move over each other without the metallic bonds being broken, therefore the delocalised electrons between the ions in the metallic structure are free to move and can be made to flow in one direction

Exercise 3.4

12

Observation	Explanation
Diamond is a very hard substance …	… because all the atoms in the structure are joined by strong covalent bonds.
Diamond does not conduct electricity …	… because **all the outer electrons of the atoms are involved in making bonds.**
Graphite is **slippery** …	… because the layers in the structure are only held together by weak forces.
Graphite conducts electricity …	… because there are some free electrons that are able to move between the layers to carry the current.

Chapter 3 continued

13 a Graphite conducts electricity because not all of the outer electrons of the carbon atoms are used in the covalent bonding that holds the atoms together in the layers. These 'free' electrons are able to move in between the layers. They can be made to move in one direction when a voltage is applied.

b Graphite acts as a lubricant because there are only weak forces between the layers of carbon atoms in the structure. The layers can be made to move over each other if a force is applied.

14 a The strong bonds between the atoms are **covalent** bonds.

b In the crystal, there are two oxygen atoms for every silicon atom, so the formula is $\mathbf{SiO_2}$.

c The atoms of the lattice are organised in a **tetrahedral** arrangement like diamond, with a silicon atom at the centre of each **tetrahedron (pyramid)**.

d This is an example of a **giant molecular (covalent)** structure.

e Each oxygen atom forms **two** covalent bonds.

f Each silicon atom forms **four** covalent bonds.

15 Silicon(IV) oxide occurs naturally as ~~**mud**~~ / **sand**. It has a giant **covalent** / ~~**electrostatic**~~ structure very similar to ~~**graphite**~~ /**diamond**. Such a structure can also be described as a giant **molecular** / ~~**ionic**~~ structure as all the atoms in the crystal are joined together by covalent bonds.

Each silicon atom is bonded to **four** / ~~**two**~~ oxygen atoms, while each oxygen atom is linked covalently to ~~**four**~~ / **two** silicon atoms. The oxygen atoms are arranged ~~**hexagonally**~~ / **tetrahedrally** around the silicon atoms.

The fact that all the atoms are bonded together in a ~~**two-dimensional**~~ / **three-dimensional** structure like ~~**graphite**~~ / **diamond** means that silicon(IV) oxide has similar physical properties to ~~**graphite**~~ / **diamond**. Silica is very **hard** / ~~**slippery**~~ and has a ~~**low**~~ / **high** melting point. All the outer electrons of the atoms are used in making the covalent bonds between the atoms. This means that silicon(IV) oxide ~~**does**~~ / **does not** conduct electricity. There are no electrons free to carry the current through the crystal.

Chapter 4

Exercise 4.1

1 **a** 1H; 1N; 3O
b 1Cu; 2N; 6O
c 2N; 8H; 1S; 4O
d 1K; 1Mn; 4O

2 **a**

			H 1							
Li 1					B 3	C 4	N 3	O 2	F 1	Ne 0
Na 1	Mg 2				Al 3			S 2	Cl 1	
K 1	Ca 2		transition elements variable valency	Zn 2					Br 1	

b Li, Na, K, Mg, Ca, Zn, Al
c N, O, F, S, Cl, Br
d B, C

3 **a** H_2S
b B_2O_3
c CS_2
d NH_3

4 12 carbon atoms + 22 hydrogen atoms + 11 oxygen atoms = 45

5 **a** PR_2 (or R_2P)
b QR (or RQ)
c Q_2
i ionic
ii ionic
iii covalent

6 **a** $MgBr_2$, $Mg^{2+}Br^-$
b Ca_3N_2, $Ca^{2+}N^{3-}$
c Al_2O_3, $Al^{3+}O^{2-}$

7 **a** molecular formula $C_2O_2H_4$
b empirical formula COH_2

8 The formula of a simple molecular **compound** shows exactly how many atoms are **bonded** together in each molecule. For example, ethane has two **carbon** and six **hydrogen** atoms, so its formula is C_2H_6. This is the **molecular** formula for ethane. This formula can be simplified to CH_3 by **dividing** through by 2. CH_3 is the **empirical** formula of ethane.

The formulae used for giant covalent and **ionic** compounds are the simplest **ratio** of the different **atoms** or ions in each compound. These formulae are known as the **empirical** formulae of these types of compounds.

Exercise 4.2

9 **a** methane + oxygen → carbon dioxide + water
b $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$

10 **a** zinc + copper(II) sulfate → zinc sulfate + copper
b ammonium chloride + sodium hydroxide → ammonia + sodium chloride + water
c iron + oxygen → iron(III) oxide

11 **a** $2Na + Cl_2 \rightarrow 2NaCl$
b $2SO_2 + 3O_2 \rightarrow 2SO_3$
c $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$
d $2PbO + C \rightarrow 2Pb + CO_2$

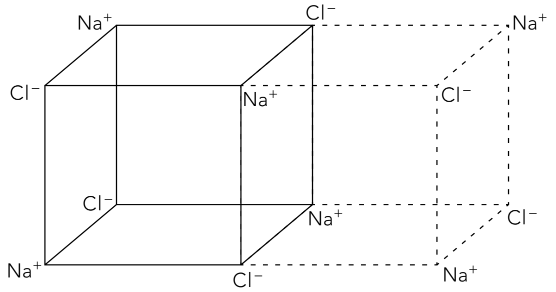
12 **a** CuO
b Na_2CO_3
c $ZnSO_4$
d $AgNO_3$
e $(NH_4)_2SO_4$
f K_3PO_4
g $Fe(OH)_3$
h $CrCl_3$

13 **a** 1 : 1
b 1 : 3: 3
c 2 : 8 : 1 : 4

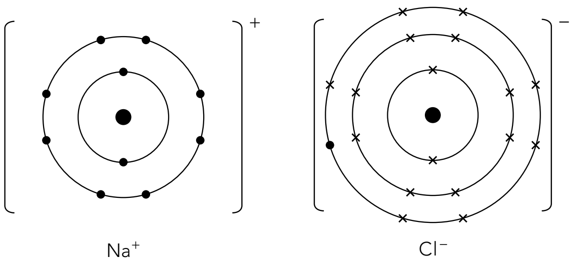
14 **a** 2 : 1
b K_2O

Chapter 4 continued

15 a



b



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

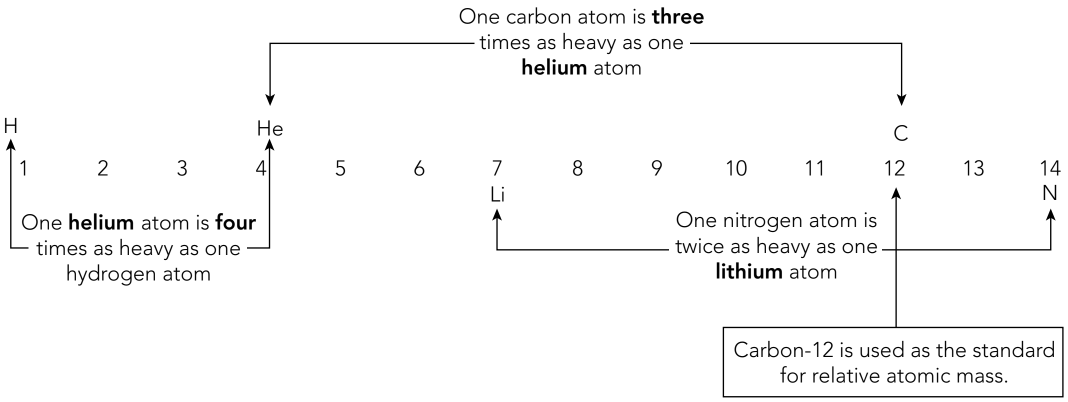
b $Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

c $Fe^{2+}(aq) + OH^-(aq) \rightarrow Fe(OH)_2(s)$

d $Zn^{2+}(aq) + Mg(s) \rightarrow Zn(s) + Mg^{2+}(aq)$

Exercise 4.3

17



18 5 tonnes zinc oxide ➞ 4 tonnes zinc

so 20 tonnes zinc oxide ➞ $4 \times \left(\frac{20}{5}\right)$
= 16 tonnes zinc

or $\frac{4}{5} = x / 20$

so $x = 4 \times \frac{20}{5}$ = 16 tonnes of zinc

19 a Mass is conserved during a reaction, so the mass of the products equals the mass of reactants:

mass of reactants = 68 + 96 = 164 g

so, mass of water is 164 – 56 = 108 g

b in the reaction shown, 68 g of ammonia is burnt. If the mass of ammonia burnt is 17 g, this is

$\frac{17}{68} = \frac{1}{4}$

so, mass of water formed is

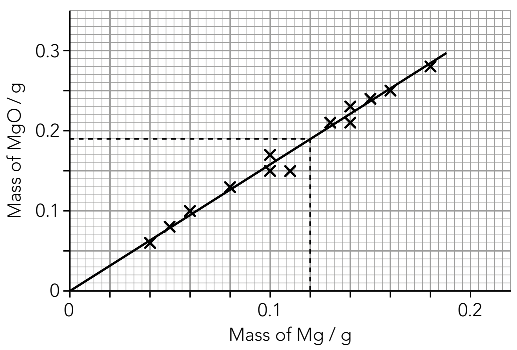
$\frac{1 \times 108}{4} = 27$ g

Chapter 4 continued

20

Molecule	Chemical formula	Number of atoms or ions involved	Relative molecular (or formula) mass
Oxygen	O_2	2O	2 × 16 = 32
Carbon dioxide	**CO_2**	1C and 2O	1 × 12 + 2 × 16 = **44**
Water	H_2O	2H and 1O	**2 × 1 + 16 = 18**
Ammonia	**NH_3**	1N and 3H	**14 + 3 × 1 = 17**
Calcium carbonate	**$CaCO_3$**	$1Ca^{2+}$ and $1CO_3^{2-}$	**40** + **12** + 3 × 16 = 100
Magnesium oxide	MgO	$1Mg^{2+}$ and $1O^{2-}$	1 × 24 + 1 × 16 = **40**
Ammonium nitrate	NH_4NO_3	$1NH_4^+$ and **$1NO_3^-$**	2 × 14 + **4 × 1** + **3 × 16** = 80
Propanol	C_3H_7OH	3C, **8H** and **1O**	3 × 12 + 8 × 1 + **1 × 16** = **60**

21 a



b The mass of magnesium oxide produced increases if more magnesium is used. The increase is linear (directly proportional).

c mass of MgO produced = 0.19 g

Note that your answer may differ slightly because your line of best fit may be slightly different.

d $0.19 - 0.12 = 0.07$ g

e $\frac{0.07}{0.12} \times 24 = 14$ g

f from experiment: 0.12 g magnesium gives 0.19 g magnesium oxide

1.00 g will give $\frac{0.19}{0.12}$ g magnesium oxide

therefore 3.5 g will produce $\frac{0.19}{0.12} \times 3.5 =$ 5.54 g of magnesium oxide

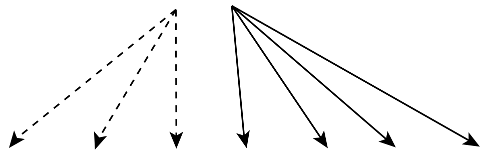
g Magnesium is sufficiently reactive to combine with nitrogen from the air to produce magnesium nitride, Mg_3N_2

Chapter 5

Exercise 5.1

1

The formula of magnetite is Fe_3O_4



The atoms present are Fe ....*Fe*.... Fe O ....*O*.... O ....*O*....

The relative atomic mass is ....*56*.... + ...*56*.... + ...*56*.... + ...*16*.... + ...*16*.... + ...*16*.... + ...*16*.... = ...*232*....

- The relative formula mass of the iron oxide (Fe_3O_4) = 232
- In this formula, there are three atoms of iron, Fe.
- The relative mass of 3Fe = 168.
- This means that in 232 g of Fe_3O_4 there are 168 g of iron.
- So 1 g of Fe_3O_4 contains $\frac{168}{232}$ = 0.724 g of iron.
- So 100 g of Fe_3O_4 contains 72.4 g of iron.
- In other words, the percentage (%) of iron in Fe_3O_4 = 72.4%.

2 a Empirical formula of oxalic acid = CHO_2

	C	H	O
% by mass	26.7	2.2	71.1
Moles in 100 g	$\frac{26.7}{12}$ = 2.23	$\frac{2.2}{1}$ = 2.2	$\frac{71.1}{16}$ = 4.44
Ratio of moles	2.23	2.2	4.4
Whole number ratio	1	1	2

b Mass of empirical formula = 12 + 1 + (2 × 16) = 45

But actual formula mass = 90

Therefore, actual molecular formula of oxalic acid = $C_2H_2O_4$

3 Empirical formula of compound = AsC_3H_9

	As	C	H
% by mass	62.5	30.0	7.5
Moles in 100 g	$\frac{62.5}{75}$ = 0.83	$\frac{30.0}{12}$ = 2.5	$\frac{7.5}{1}$ = 7.5
Ratio of moles	0.83	2.5	7.5
Whole number ratio	1	3	9

Exercise 5.2

4 a From equation: 1 mol Fe_2O_3 gives 2 mol Fe

100 g of Fe = $\frac{100}{56\,\text{mol}}$ = 1.79 mol

mol of Fe_2O_3 needed = $\frac{1.79}{2}$ = 0.895 mol

M_r of Fe_2O_3 = (56 × 2) + (16 × 3) = 160

mass of Fe_2O_3 needed = 0.895 × 160 = 143.2 g

100 g of iron is **1.79** moles of Fe, so **0.895** moles of Fe_2O_3 are needed for the reaction, or **143.2 g** of iron(III) oxide.

b From above: 143.2 g Fe_2O_3 gives 100 g of iron

so 143.2 tonnes Fe_2O_3 gives 100 tonnes of Fe

Therefore 71.6 tonnes of Fe_2O_3 are needed to produce 50 tonnes of Fe

Chapter 5 continued

c mol Fe_2O_3 in 100 tonnes = $100 \times \frac{10^6}{160}$ = 625 000 or 6.25×10^5 mol

mol Fe expected = 625 000 × 2 = 1 250 000 or 1.25×10^6 mol

mass of Fe expected = 1 250 000 × 56 = 70 000 000 g or 70 tonnes

Fe actual yield = 7 tonnes

Therefore, percentage yield = $\frac{7 \times 100}{70}$ = 10%

5 a $CaCO_3 \rightarrow CaO + CO_2$

b 1 mol $CaCO_3$ gives 1 mol CaO (quicklime)

100 g $CaCO_3$ gives 56 g CaO

or 100 tonnes $CaCO_3$ gives 56 tonnes CaO

1 tonne $CaCO_3$ gives $\frac{56}{100}$ tonnes CaO

$\frac{56}{100}$ tonnes CaO = 0.56 tonnes

c From above: 1 tonne $CaCO_3$ gives 0.56 tonnes CaO

expected: 2.5 tonnes $CaCO_3$ would give (0.56 × 2.5) = 1.4 tonnes CaO

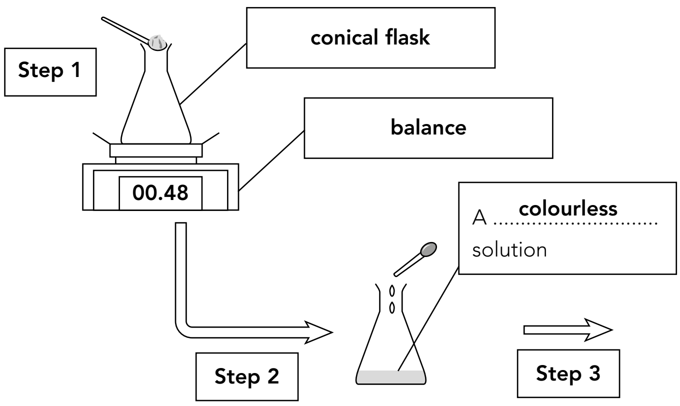
actual yield = 1.12 tonnes

therefore, percentage yield = $\frac{1.12 \times 100}{1.4}$ = 80%

The yield is not 100% because the limestone may not be 100% calcium carbonate – or other valid reason.

Exercise 5.3

6 a



b

Final burette reading / cm^3	14.60
First burette reading / cm^3	0.20
Volume of NaOH(aq) added / cm^3	**14.40** (*P*)

c **Stage 1:**

- 14.40 cm^3 of NaOH(aq) containing 0.50 moles in 1000 cm^3 were used.
- Number of moles NaOH used = $\frac{0.50}{1000} \times 14.40 = 7.20 \times 10^{-3}$ moles (or 0.0072 moles)

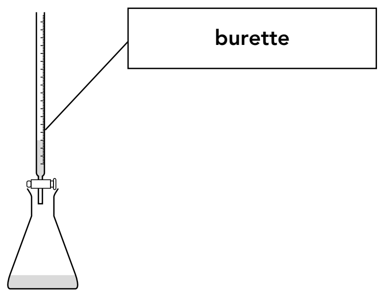
Stage 2:

- Note that 1 mole of citric acid reacts with 3 moles of sodium hydroxide.
- Number of moles citric acid in sample $= \frac{7.20 \times 10^{-3}}{3} = 2.40 \times 10^{-3}$ moles (or 0.0024 moles)

Stage 3:

- Relative formula mass of citric acid (M_r of $C_6H_8O_7$) = 192
- Mass of citric acid in sample = $2.40 \times 10^{-3} \times 192 = 0.46$ g
- Percentage purity of sample = $\frac{0.46}{0.48} \times 100 = 95.8\%$
- (Note that you have a clue that you are on the right lines in your calculation because your value for the mass of citric acid must be less than 0.48 g)

d The citric acid can be further purified by re-crystallisation.



Chapter 5 continued

7 Number of moles of H_2SO_4 in 25 cm^3 of 2 mol/dm^3 solution $= \frac{2}{1000} \times 25 = 0.05$ moles

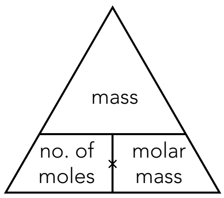
Maximum number of moles of $CuSO_4{\cdot}5H_2O$ that could be formed = 0.05 moles

Maximum mass of crystals, $CuSO_4{\cdot}5H_2O$, that could be formed $= 0.05 \times 250 = 12.5$ g (The mass of one mole of $CuSO_4{\cdot}5H_2O$ is 250 g.)

Percentage yield $= \frac{7.3}{12.5} \times 100 = 58.4\%$

Exercise 5.4

8



Substance	A_r or M_r	Number of moles	Mass / g
Cu	**64**	**2**	128
Mg	**24**	0.5	**12**
Cl_2	**71**	**0.5**	35.5
H_2	**2**	**2**	4
S_8	**256**	2	**512**
O_3	**48**	**0.033**	1.6
H_2SO_4	**98**	2.5	**245**
CO_2	**44**	0.4	**17.6**
NH_3	**17**	**1.5**	25.5
$CaCO_3$	**100**	**1**	100
$MgSO_4{\cdot}7H_2O$	**246**	**0.333**	82

9 **a** 128 g Cu = 2 mol Cu

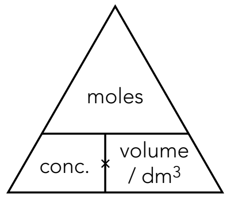
1 mol $= 6.02 \times 10^{23}$ atoms Cu

2 mol Cu $= 2 \times 6.02 \times 10^{23} = 1.20 \times 10^{24}$ atoms of copper

b 4 g H_2 = 2 mol H_2 molecules

2 mol H_2 molecules $= 2 \times 6.02 \times 10^{23} = 1.20 \times 10^{24}$ H_2 molecules $= 2.4 \times 10^{24}$ H atoms

10



Solute	Volume of solution	Molar concentration of solution / mol/dm^3	Moles of solute present
Sodium chloride	1 dm^3	0.5	**0.5**
Hydrochloric acid	500 cm^3	0.5	**0.25**
Sodium hydroxide	2 dm^3	**0.5**	1.00
Sulfuric acid	250 cm^3	**2**	0.50
Sodium thiosulfate	**200 cm^3**	2	0.40
Copper(II) sulfate	**7.5 dm^3**	0.1	0.75

Exercise 5.5

11

Experiment number	Volume of hydrogen collected / cm^3
1	85
2	79
3	82
Mean average	**82**

The three results are not equal because of the difficulty in cutting exactly equal lengths of magnesium ribbon. Also, the pieces of ribbon may not be exactly the same thickness or width, or gas may be lost as the magnesium is allowed to fall into the flask, or there may have been air in the measuring cylinder before starting.

Chapter 5 continued

12 From equation: 24 g of magnesium (1 mole) → 24 000 cm³ of hydrogen

so 1 cm³ of hydrogen produced from $\frac{24}{24000}$ = 0.001 g of magnesium

and 82 cm³ of hydrogen produced from 0.001 × 82 = 0.082 g

13 24 g of magnesium → 120 g of magnesium sulfate

so 0.082 g will give $\frac{120}{24}$ × 0.082 = 0.41 g

The answers to 12 and 13 could be calculated by other proportionality methods.

14 The key factor here is that 24 g of magnesium will produce 120 g of dried anhydrous magnesium sulfate ($MgSO_4$) (see the equation).

- Weigh out a known mass of magnesium ribbon.
- React it with excess dilute sulfuric acid until no more gas is given off and no magnesium remains.
- Transfer the solution to a beaker of known mass.
- Heat the solution to dryness, taking care to avoid it spitting.
- Allow to cool and weigh the beaker and residue.
- Filter, dry and weigh the crystals carefully.
- From the data above, calculate the mass of crystals that 5 cm would have given.

Exercise 5.6

15 75 cm³

16 25 cm³

17 50 cm³

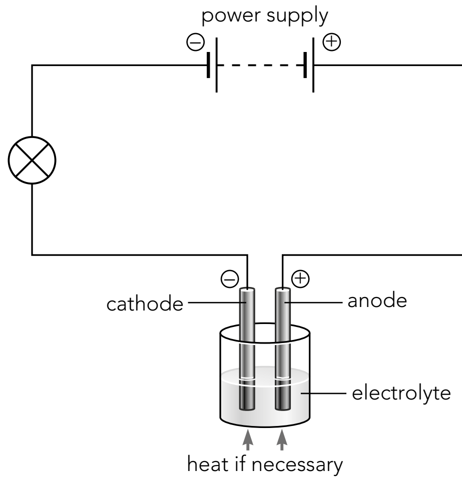
18

2NO	+	O_2	→	**2**NO_2
50 cm³		**25** cm³		**50** cm³

Chapter 6

Exercise 6.1

1 a



b graphite and platinum

c A–3, B–1, C–2, D–4

2

Molten electrolyte	Product at anode (+)	Product at cathode (–)	Observations of product at anode
Lead(II) iodide	iodine	lead	purple vapour given off
Magnesium chloride	chlorine	magnesium	green gas given off
Zinc bromide	bromine	zinc	red–brown vapour given off
Calcium oxide	oxygen	calcium	colourless gas given off

Exercise 6.2

3 a chromate ions

b Chromate ions are negatively charged and so move towards (are attracted to) the positive electrode.

c Anions: chromate (CrO_4^{2-}) and sulfate (SO_4^{2-})

Cations: potassium (K^+) and copper (Cu^{2+})

d A blue colour moves towards negative (Cu^{2+} ions), and a yellow colour moves towards positive (CrO_4^{2-} ions).

4 During electrolysis, ionic compounds are decomposed by the passage of an electric current. For this to happen, the compound must be either **molten** or in **solution** in water. Electrolysis can occur when an electric **current** passes through a molten **electrolyte**. The two rods dipping into the electrolyte are called the **electrodes**. In this situation, metals are deposited at the **cathode** and non-metals are formed at the **anode**.

When the ionic compound is dissolved in water, the electrolysis can be more complex. Generally, during electrolysis, **positive** ions move towards the **cathode** and negative ions move towards the **anode**. At the negative electrode (cathode) the metal or **hydrogen** ions gain electrons and form metal atoms or hydrogen **molecules**. At the positive electrode, (anode) non-metals are formed as their ions or **hydroxide** ions from the water lose electrons.

Exercise 6.3

5

Solution (electrolyte)	Gas given off at the anode	Gas given off or metal deposited at the cathode	Substance left in solution at the end of electrolysis
concentrated potassium bromide	**bromine**	hydrogen	**potassium hydroxide**
copper(II) nitrate	**oxygen**	copper	nitric acid
silver sulfate	oxygen	**silver**	**sulfuric acid**
sodium nitrate	**oxygen**	hydrogen	sodium nitrate

6 a $H_2(g) \rightarrow 2H^+(aq) + 2e^-$

b $O_2(g) + \mathbf{4}H^+(aq) + \mathbf{4}e^- \rightarrow \mathbf{2}H_2O(l)$

c $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

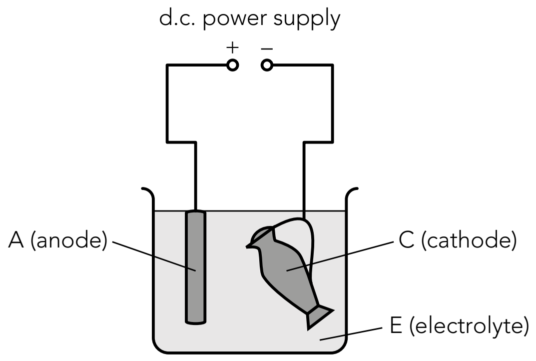
d The only waste product is non-polluting water / they do not need recharging / they are more efficient than a petrol engine.

e needs to be an increased availability of hydrogen refuelling stations / needs to be a move to 'greener' sustainable sources of hydrogen rather than from a fossil fuel source

Chapter 6 continued

Exercise 6.4

7 a



b copper

c copper(II) sulfate (or copper(II) chloride / copper(II) nitrate)

d The anode dissolves away, and a sludge forms below it.

e To protect the object from corrosion / for decorative purposes.

8 a copper(II) chloride ➞ copper + chlorine

$CuCl_2(aq) \rightarrow Cu(s) + Cl_2(g)$

b The gas will bleach moist litmus paper.

c The reaction is endothermic / electrical energy is being used to decompose the compound.

9 a i cathode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

ii anode: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$

b i Change in mass at cathode: mass increases as copper deposited

ii Change in mass at anode: mass decreases as copper dissolves away from anode

c The blue colour is due to the copper ions present in the solution / the copper ions discharged at the cathode are replaced by those dissolving into solution at the anode / so the colour does not fade.

Chapter 7

Exercise 7.1

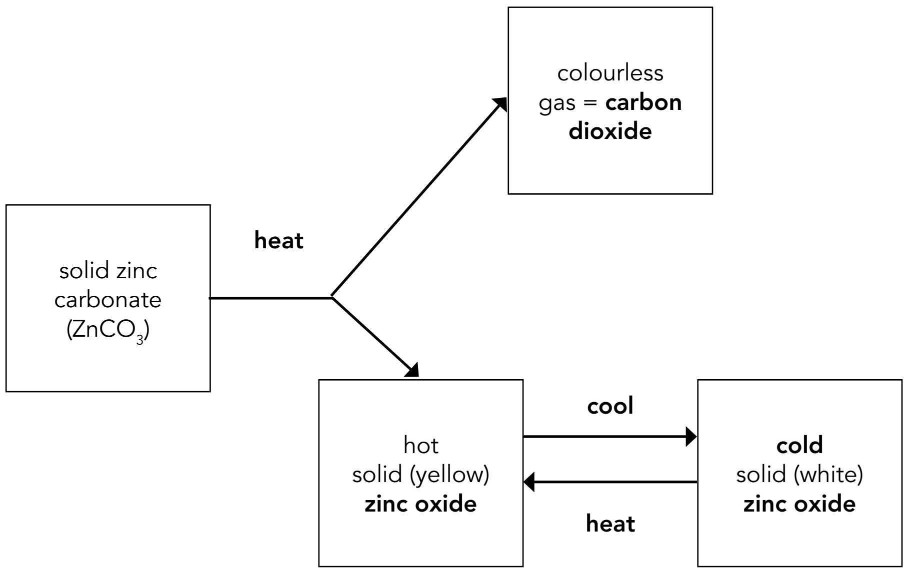
1 When a **physical** change takes place, the substance undergoing the change is unchanged **chemically**. When a **chemical** change takes place, the substance or substances formed are **different to** the starting substance. Physical changes are easy to **reverse** so that we can easily go back to the starting substance. **Chemical** changes are difficult to **reverse**, so it is more difficult to form the starting substance again.

2 **a** Ice melting is a physical change, because the ice and water are chemically the same / both H_2O. Melting is also easy to reverse.

b Magnesium burning in air is a chemical change because a new compound, magnesium oxide, is formed. Also, it is difficult to reverse the change and lots of energy is given out in the reaction.

c Salt dissolving in water is a physical change because the salt and water are unchanged chemically (the salt is just dissolved) – it is sodium chloride still. The dissolving can be reversed by evaporation of the water.

3 **a**



b $ZnCO_3 \rightarrow ZnO + CO_2$

c Chemical – new substances, zinc oxide and carbon dioxide, are formed from the zinc carbonate.

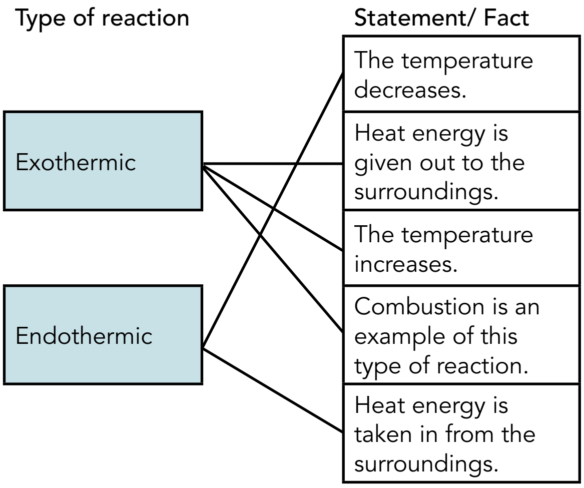
d The zinc carbonate is broken up (decomposition) using heat (thermal).

e When the zinc oxide cools down it changes colour from yellow to white, which is a physical change. This change is easily reversed because when it is heated it turns yellow again.

Chapter 7 continued

Exercise 7.2

4



5 **a** exothermic

b The chemicals might burst the container and mix with the food/drink.

c By leaving an empty space above the chemicals.

d calcium oxide + water ➞ calcium hydroxide

$CaO + H_2O \rightarrow Ca(OH)_2$

6 **a** Advantage: can be used anywhere instantly.
Disadvantage: cannot be re-used.

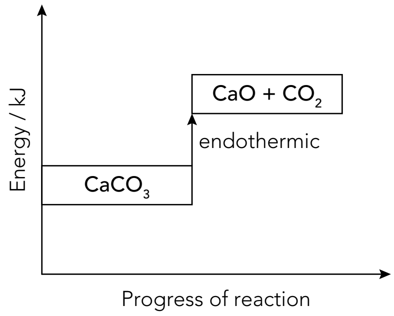
b Use a known amount of water, measure the temperature and then add a series of known masses of ammonium nitrate. Measure the temperature after each addition. Plot a graph to show temperature change against mass of ammonium nitrate added. Use the graph to determine the mass needed to reduce the temperature to 5 °C. Other methods are possible, but they must work.

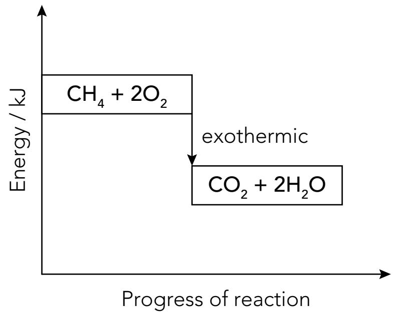
Exercise 7.3

7 **a** In an exothermic reaction, the **reactants** / ~~**products**~~ have more thermal energy than the ~~**reactants**~~ / **products**. This means that **thermal energy** / ~~**potential energy**~~ is transferred **to** / ~~**from**~~ the surroundings and therefore the temperature of the surroundings **increases** / ~~**decreases**~~.

b In an endothermic reaction, the ~~**reactants**~~ / **products** have more thermal energy than the **reactants** / ~~**products**~~. This means that energy is transferred ~~**to**~~ / **from** the surroundings and the temperature of the surroundings ~~**increases**~~ / **decreases**.

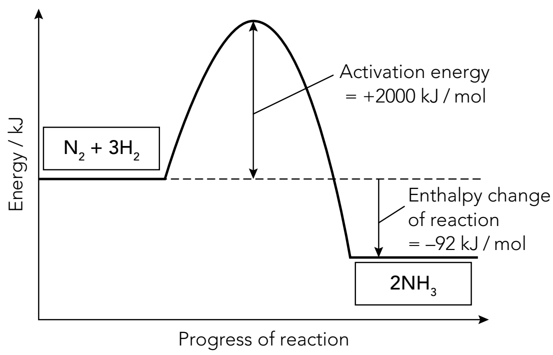
8





9 **a** **A** reactants
B activation energy
C enthalpy of reaction
D products

b



Chapter 7 continued

Exercise 7.4

10 The bond energy is the energy required to break one **mole** of covalent bonds. The energy is measured in **kilojoules**. When bonds are **broken** energy has to be added to the system and the sign for ΔH is **positive**. When bonds are **made** energy is given out by the system and the sign for ΔH is **negative**. The enthalpy change for the reaction is the difference between the energy **taken in** when bonds are broken and the energy **given out** when bonds are **made**. If the overall change in enthalpy is negative then the reaction is **exothermic**. If it is positive then the reaction is an **endothermic** one.

11 a i Weakest bond = C–C

ii Strongest bond = C=O

b i Number of moles of OH bonds = 4

ii Energy required = 4 × 464 = 1864 kJ

c i Bonds broken = 4C–H + 2O=O = 6 bonds

ii Bonds made = 2C=O + 4O–H = 6 bonds

12 a Number of moles of bonds broken: 2 × C–C, 8C–H and 5 × O=O

b Energy required to break moles
2 × C–C = 2 × 347 = 694

8 × C–H = 8 × 413 = 3304

5 × O=O = 5 × 498 = 2490

Total energy required for bond breaking = 6488 kJ / mol

c Bonds made: 6 × C=O and 8 × O–H

d Energy given out by making 6 × C=O bonds = 6 × 805 = 4830

Energy given out by making 8 × O–H bonds = 8 × 464 = 3712

Total energy given out = 8542 kJ / mol

e Enthalpy change for combustion of propane = (energy required to break bonds) − (energy given out when bonds formed) = 6488 − 8542 = –2054 kJ / mol

f i 0.2 mol of propane gives −2054 × 0.2 = –410.8 kJ

ii 4 mol of propane gives −2054 × 4 = –8216 kJ

iii 33 g of propane = $\frac{33}{44}$ mol = 0.75 mol

Therefore energy released = 0.75 × –2054 = –1540.5 kJ

Chapter 8

Exercise 8.1

1 **a** B has the greater surface area because more surfaces are exposed.

b calcium carbonate(s) + hydrochloric acid(aq) ➞ calcium chloride(aq) + water(l) + carbon dioxide(g)

2 The mass decreases because carbon dioxide is given off, leaves the solution and therefore the mass decreases.

3 temperature of acid, concentration of acid, volume of acid, mass of marble chips

4 Temperature: if it is increased the particles of acid move more quickly, collide with the solid surface more frequently and with greater force so that there are more collisions with an energy greater than the activation energy.

Concentration of acid: particles are more crowded and collide more frequently, so there is more chance of a reaction.

Volume of acid: if a greater volume of acid (of the same concentration) were present, then the reaction would take place for longer. As long as the acid covers the chips, it does not affect the initial rate of reaction.

Mass of solid (of same particle size): if more solid is present then reaction would continue for longer.

5 In the reaction between calcium carbonate and hydrochloric acid, the acid particles can only collide and **react** with the calcium carbonate particles on the **outside** of the pieces, the ones on the **inside** cannot react. In the smaller pieces, more of the **inside** particles are exposed to the acid and the acid particles can therefore **react** with more calcium carbonate particles and more reactions take place.

6 **a** For the large pieces of marble chips (flask B), readings (in grams) were:

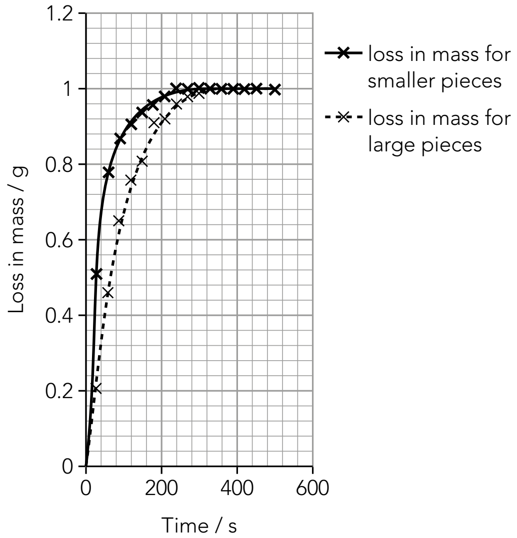
Time/s	0	30	60	90	120	150	180	210
Mass/g	240.86	240.65	240.40	240.21	240.10	240.05	239.95	239.94
Loss in mass/g	0.00	0.21	0.46	**0.65**	**0.76**	**0.81**	**0.91**	**0.92**
Time/s	240	270	300	330	360	390	420	450
Mass/g	239.90	239.88	239.87	239.86	239.86	239.86	239.86	239.86
Loss in mass/g	**0.96**	**0.98**	**0.99**	**1.00**	**1.00**	**1.00**	**1.00**	**1.00**

For the small pieces of marble chips (flask B), readings (in grams) were:

Time/s	0	30	60	90	120	150	180	210
Mass/g	240.86	240.35	240.08	239.99	239.95	239.92	239.90	239.88
Loss in mass/g	0.00	**0.51**	**0.78**	**0.87**	**0.91**	**0.94**	**0.96**	**0.98**
Time/s	240	270	300	330	360	390	420	450
Mass/g	239.87	239.86	239.86	239.86	239.86	239.86	239.86	239.86
Loss in mass/g	**0.99**	**1.00**	**1.00**	**1.00**	**1.00**	**1.00**	**1.00**	**1.00**

Chapter 8 continued

b



c Smaller pieces. The graph is steeper at the beginning and reaches the maximum amount of gas earlier.

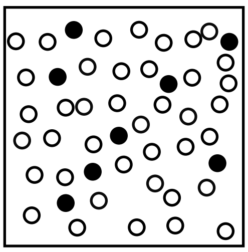
7 a The greater the surface area, the faster the rate.

b The same amount of acid was used.

Exercise 8.2

8 The concentration of a **solution** / ~~**solvent**~~ is a measure of the number of particles of **solute** / ~~**solvent**~~ per unit **volume** / ~~**area**~~.

9 The third box should have eight filled circles. The reason for this is that the concentration is equal to the first box which has two closed circles. But, the volume in the first box is 1 dm^3 and the third box is 4 dm^3 which means we need four times as many solute particles for the same concentration.



10 a $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

b A = delivery tube

B = clamp

C = gas syringe

c time, volume of gas

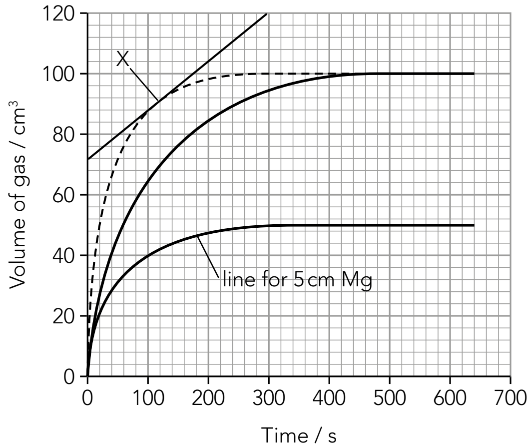
11 a 440 s

b average rate

$$= \frac{\textbf{maximum volume}}{\textbf{time taken to complete the reaction}}$$

$$= \frac{\mathbf{100}}{\mathbf{440}} = 0.23\,cm^3/s$$

12



a The line (X) should be steeper at first and level off to the same volume as the original line.

b The line should be as steep as the original line at first and then level off to 50 cm^3 because half the mass of magnesium was used.

c An example is shown. The slope of this line is $\frac{\mathbf{88-72}}{\mathbf{100}} = \frac{\mathbf{16}}{\mathbf{100}} = 0.16\,cm^3/s$

d As the reaction proceeds the acid is used up and therefore its concentration decreases and the rate slows down. This means that the slope of the graph decreases and the line gradually gets less steep.

Exercise 8.3

13 a $Na_2S_2O_3(aq) + \mathbf{2}HCl(aq) \rightarrow H_2O(l) + SO_2(g) + S(s) + \mathbf{2}NaCl(aq)$

b The sulfur formed is a solid and it makes the solution cloudy. Therefore as it increases the reaction mixture gets cloudier.

c The cross will be obscured when there is enough sulfur in the reaction mixture to make it so cloudy that it cannot be seen. This means that we are measuring the time it takes to make a certain amount of sulfur.

Chapter 8 continued

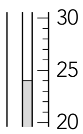
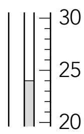
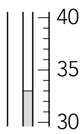
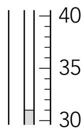
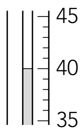
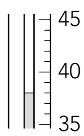
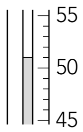
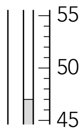
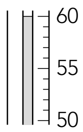
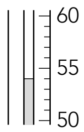
d The sulfur dioxide is toxic and causes respiratory distress, so we need a well-ventilated room.

Also when the reaction mixture is discarded, it should be discarded in a fume cupboard sink.

Safety specs are needed because of the acid used.

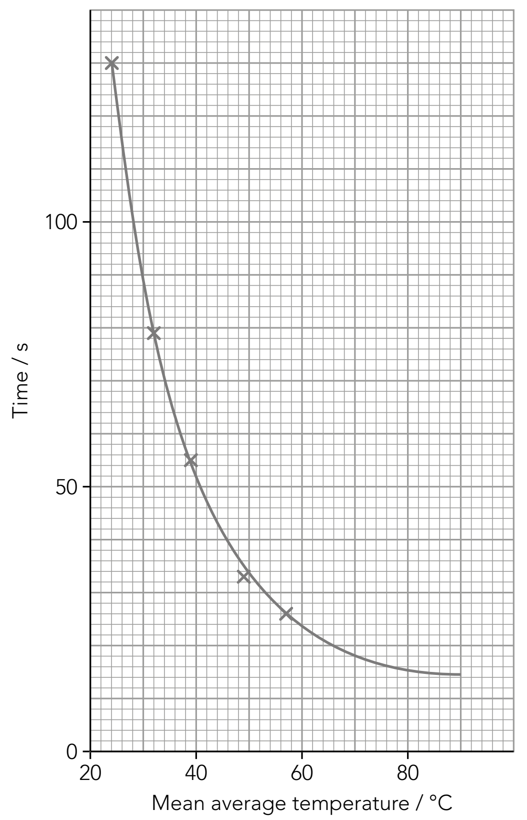
14 When the temperature of a reaction mixture is increased the particles move around more **quickly** and this increases their **kinetic** energy. Because they move around more **quickly** at higher temperatures, they collide more **frequently** and this **increases** the chance of a reaction taking place. More importantly, when the particles do **collide** the collisions are **more** efficient. This means that more collisions have an energy greater than the **activation** energy, which is the energy required for reaction to occur.

15 a

Experiment	Thermometer diagram	Initial temperature / °C	Thermometer diagram	Final temperature / °C	Mean average temperature / °C	Time for printed text to disappear / s
1		**24**		**24**	**24**	130
2		**33**		**31**	**32**	79
3		**40**		**38**	**39**	55
4		**51**		**47**	**49**	33
5		**60**		**54**	**57**	26

Chapter 8 continued

b

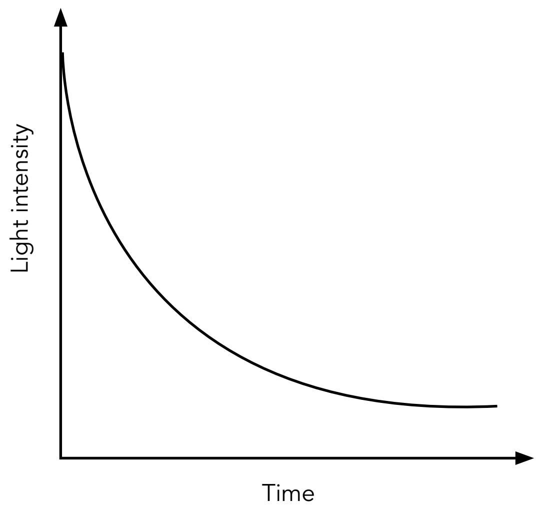


c Experiment 5

d This was the experiment carried out at the highest temperature. Increasing the temperature increases the rate of a reaction because the particles are moving faster and therefore collide more frequently. They also have more energy when they collide and so are more likely to react.

e As the reaction mixture gets cloudier, due to the production of sulfur, the intensity/brightness of the light reaching the light data-logger decreases.

f



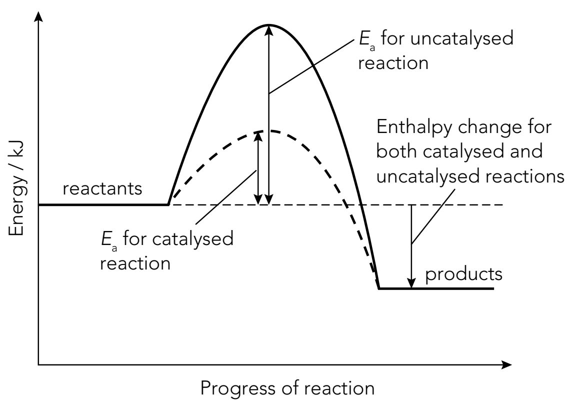
Exercise 8.4

16 The correct order is U W Q P V T R S

17 a Copper(II) sulfate. It speeds up the reaction because the formation of bubbles of hydrogen gas increase greatly but it is changed chemically. The red solid produced is copper (this newly-made copper acts as a catalyst and speeds up the reaction).

b Copper is the true catalyst because it speeds up the reaction and is unchanged and is still copper at the end.

18



Chapter 9

Exercise 9.1

1 a endothermic

b By adding water drop-wise to the powder.

c If the forward reaction is endothermic then the reverse reaction must be exothermic / because the water being added becomes chemically combined (with the cobalt ions) in the solid and bonds are formed.

d anhydrous

2 a $CuSO_4{\cdot}5H_2O \rightleftharpoons \mathbf{CuSO_4} + 5H_2O$
blue white

b The reaction is reversible.

c The white powder would turn blue, with heat given out / the test detects the presence of water and the water does not have to be pure / the water can still react even though there is some ethanol there.

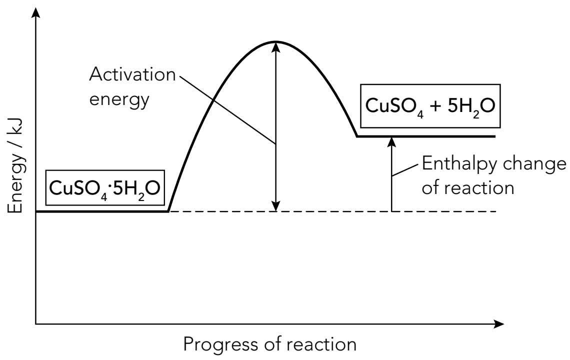
d Test its boiling point (should be 100 °C at sea level) / or its freezing point.

3 a –75.3 kJ / mol

b The temperature increases.

c The backward reaction is exothermic which means that heat is released to the surroundings and the temperature increases.

d



4

	True	False
All reactions are easily reversible.		✓
Reversible reactions are represented by the arrows ⇌.	✓	
When the concentrations of reactants and products are equal, reversible reactions stop.		✓
If the forward reaction is endothermic, the backward reaction is exothermic.	✓	
Heating always favours the endothermic reaction in a reversible reaction.	✓	
When equilibrium is reached, reversible reactions stop.		✓
Reversible reactions can go forwards or backwards depending on the conditions.	✓	

Exercise 9.2

5 Crop plants are found to need three elements for healthy growth: nitrogen (N), **phosphorus** (P) and potassium (K). Plants take up these elements in the form of **salts** such as **ammonium** nitrate, potassium nitrate, and potassium or ammonium **phosphate**. These **elements** are needed for the new plants to make the **proteins** needed for growth. Farmers add **fertilisers** to the soil to replace the **nutrients** that previous crops have absorbed during growth.

6 a NH_3

b KCl: potassium chloride
HNO_3: nitric acid
H_3PO_4: phosphoric acid
NH_4NO_3: ammonium nitrate

c potassium phosphate: phosphoric acid and potassium hydroxide

ammonium sulfate: ammonia solution (ammonium hydroxide) and sulfuric acid

d ammonia + nitric acid ⟶ ammonium nitrate

$NH_3(aq) + HNO_3(aq) \longrightarrow NH_4NO_3(aq)$

Chapter 9 continued

7 **a** forwards

b **i** The rate of the forward reaction equals the rate of backward reaction.

ii Equilibrium shifts towards the ammonia and hydrogen chloride.

iii The system resists the decrease in pressure by making more gas molecules.

Exercise 9.3

8

Change in conditions	Effect on the position of the equilibrium
Increase in pressure for reactions involving gases	Equilibrium shifts to favour the side of the reaction with ~~**more**~~ / **fewer** gas molecules (the side which occupies ~~**more**~~ / **less** space).
Decrease in pressure for reactions involving gases	Equilibrium shifts to favour the side of the reaction with **more** / ~~**fewer**~~ gas molecules (the side which occupies **more** / ~~**less**~~ space).
Increase in temperature for any reaction	Equilibrium shifts to favour the ~~**exothermic**~~ / **endothermic** reaction.
Decrease in temperature for any reaction	Equilibrium shifts to favour the **exothermic** / ~~**endothermic**~~ reaction.
Addition of a catalyst	Forward and reverse reactions both sped up so no effect on equilibrium position, but the reaction reaches equilibrium ~~**slower**~~ / **faster**.

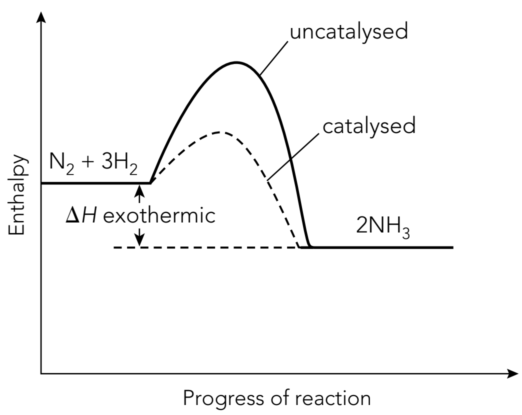
9 **a** vanadium(V) oxide, V_2O_5, or vanadium pentoxide

b The products side of the equation (the right-hand side) has fewer moles of gas (occupies a smaller volume) – so a higher pressure will favour this side of the equation. As pressure increases, the equilibrium mixture contains more sulfur trioxide.

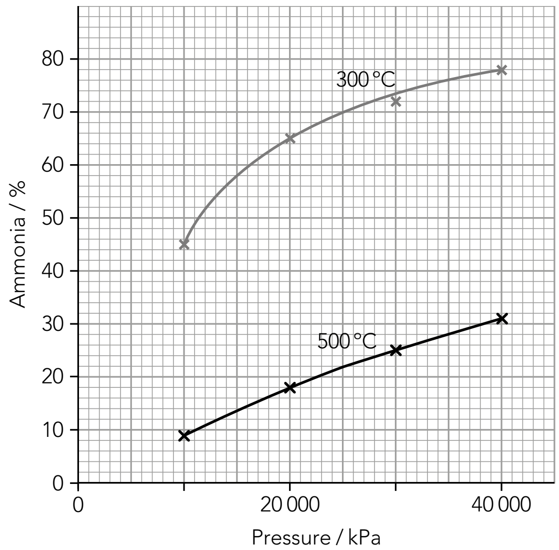
c The yield is good enough and reaction vessels that withstand high pressures are expensive to build. Also high pressures need very strong pumps and a lot of energy.

10 **a** 450 °C, 20 000 kPa pressure (and a powdered iron catalyst)

b



11 **a**



b 70% ammonia

c 46% ammonia (possible range is 42–50%)

d The disadvantage of using a low temperature would be that the ammonia would be produced at a slower rate. This slower rate could make the process uneconomic at the lower temperature.

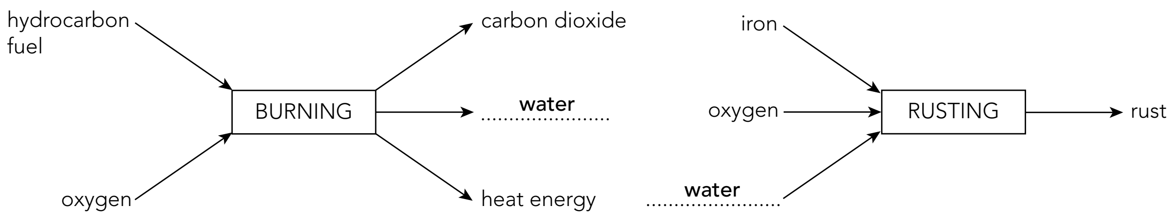
e Using a high pressure:

- favours the production of ammonia / moves the equilibrium position to the right (because there is a smaller volume of gas on the products side of the equation)
- increases the rate of production of ammonia (because the reacting molecules are therefore closer together, and collide more often).

Chapter 10

Exercise 10.1

1 **a**



b oxidation

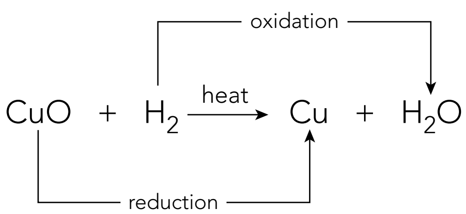
2 **a** If a substance gains oxygen during a reaction, it is **oxidised**.

b If a substance **loses** oxygen during a reaction, it is reduced.

c The substance that gives oxygen to another substance in a chemical reaction is the **oxidising** agent.

d The substance that accepts oxygen from another substance in a chemical reaction is the **reducing** agent.

3 **a**



b reducing

c **i** $\underline{Zn}(s) + Ag_2O(s) \rightarrow ZnO(s) + 2Ag(s)$

ii $Fe_2O_3(s) + 2\underline{Al}(s) \rightarrow 2Fe(s) + Al_2O_3(s)$

iii $3\underline{Mg}(s) + Al_2O_3(s) \rightarrow 3MgO(s) + 2Al(s)$

iv $CO_2(g) + 2\underline{Mg}(s) \rightarrow C(s) + 2MgO(s)$

4 **a** Carbon monoxide is the reducing agent, and oxygen is the oxidising agent. The carbon monoxide gains oxygen. Therefore oxidation and reduction have taken place.

b Nitrogen monoxide is the oxidising agent, and carbon monoxide is the reducing agent. The carbon monoxide gains oxygen. Therefore oxidation and reduction have taken place.

c $C_7H_{16} + 11O_2 \rightarrow 7CO_2 + 8H_2O$

The carbon and hydrogen both gain oxygen and are therefore oxidised. The heptane is therefore the reducing agent and oxygen is the oxidising agent.

Exercise 10.2

5 **a** **i** Oxidation is the **loss** of electrons.

ii Reduction is the **gain** of electrons.

b Transfer __1__ e^-

i Sodium (Na) 2,8,1 | Chlorine (Cl) 2,8,7

ii Sodium ion (Na^+) 2,8 | Chloride ion (Cl^-) 2,8,8

Explanation: Sodium has lost one electron and is therefore oxidised; chlorine has gained one electron and is therefore reduced.

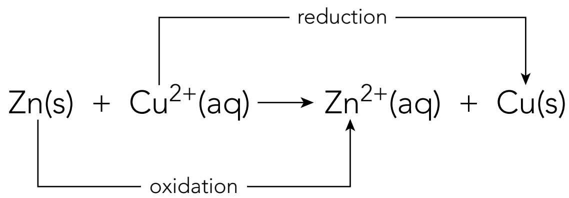
Transfer __2__ e^-

i Ca (Ca) 2,8,8,2 | Oxygen (O) 2,6

ii Calcium ion (Ca^{2+}) 2,8 | Oxide ion (O^{2-}) 2,8

Explanation: Calcium has lost two electrons and is therefore oxidised; oxygen has gained two electrons and is therefore reduced.

6 **a**



Copper(II) ions are acting as oxidising agents.

Chapter 10 continued

b Reducing agent = RA;
Oxidising agent = OA

i $Cu(s) + 2Ag^{+}(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$
RA OA

ii $3Mg(s) + 2Al^{3+}(aq) \rightarrow 3Mg^{2+}(aq) + 2Al(s)$
RA OA

c i $Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$
Electrode = cathode
Type of reaction = reduction

ii $2Cl^{-}(aq) \rightarrow Cl_2(g) + 2e^{-}$
Electrode = anode
Type of reaction = oxidation

iii $Al^{3+}(l) + 3e^{-} \rightarrow Al(l)$
Electrode = cathode
Type of reaction = reduction

7 a $Cl_2(g) + 2KI(aq) \rightarrow 2KCl(aq) + I_2(aq)$
b $Cl_2(g) + 2I^{-}(aq) \rightarrow 2Cl^{-}(aq) + I_2(aq)$
c chlorine
d $Cl_2 + 2e^{-} \rightarrow 2Cl^{-}$

8 $Cl_2 + 2Br– \rightarrow 2Cl^{-} + Br_2$

Exercise 10.3

9 a +2 **b** +3 **c** + 2 **d** +4

10 a +3 **b** 0 **c** +3 **d** +6
e +6 **f** +1

11 a $2Mg(s) + CO_2(g) \rightarrow 2MgO(s) + C(s)$
ON . 0 +4 +2 0

b The magnesium has been oxidised, because its oxidation number has increased from 0 to +2.
The carbon has been reduced, because its oxidation number has decreased from +4 to 0.
(Oxygen remains as –2 throughout.)

12

	MnO_4^{-}	$\rightarrow$	Mn^{2+}
oxidation number:	+7		+2
colour:	purple		pink (almost colourless)

13 a $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$
b an oxidation reaction / electrons have been lost; oxygen has been gained by the hydrogen
c $O_2(g) + 2H_2O(l) + 4e^{-} \rightarrow 4OH^{-}(aq)$
d $H_2(g) + 2OH^{-}(aq) \rightarrow H_2O(l) + 2e^{-}$
(multiply this equation by 2 in order to cancel out the electrons)
$O_2(g) + 2H_2O(l) + 4e^{-} \rightarrow 4OH^{-}(aq)$ / add equations together
$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$ / electrons and OH^{-} ions cancel out
e Hydrogen has greater energy efficiency / more energy produced for a given mass of fuel; water is the only product of burning hydrogen.
f Hydrogen must be transported and stored safely as it is highly explosive / hydrogen is currently expensive to produce / as it is a gas, hydrogen is difficult to transport.

Chapter 11

Exercise 11.1

1 Acids are substances that dissolve in **water** to give a solution with a pH **less** than 7. **Hydrochloric** acid has the formula HCl and is a strong acid. **Sulfuric** acid (formula = H_2SO_4) and **nitric** acid (formula = HNO_3) are also strong **acids**. In acidic solutions the concentration of the **hydrogen** ions is greater than the concentration of **hydroxide** ions.

Bases are the **oxides** and hydroxides of metals and ammonia. A base will neutralise an acid to form a salt and **water**.

Most bases are **insoluble** in water, but alkalis are bases that are **soluble** in water and their solutions have pH values **greater** than 7. KOH (**potassium** hydroxide) and NaOH (**sodium** hydroxide) are both strong **alkalis**. In alkaline solutions the concentration of **hydroxide** ions is greater than the concentration of **hydrogen** ions.

2

Practical observation	Acid	Base	Alkali
A solution of the substance has a pH of 8.			✓
A solution of the substance turns litmus paper blue.			✓
A solution of the substance turns litmus paper red.	✓		
A substance that neutralises an acid but is insoluble in water.		✓	
A substance that neutralises an acid and is soluble in water.			✓
A substance that is an insoluble oxide or hydroxide of a metal.		✓	
A substance that is a soluble hydroxide of a metal.			✓
A solution with a pH of 9 that is produced when ammonia is dissolved in water.			✓
A solution of the substance has a pH of 3.	✓		
A solution of the substance has a pH of 13.			✓

3 a Colour at A = red Colour at B = orange Colour at C = purple

b Colour of thymolphthalein = blue Colour of methyl orange = yellow

c N placed at $25\,cm^3$, pH 7 on graph

d $25\,cm^3$

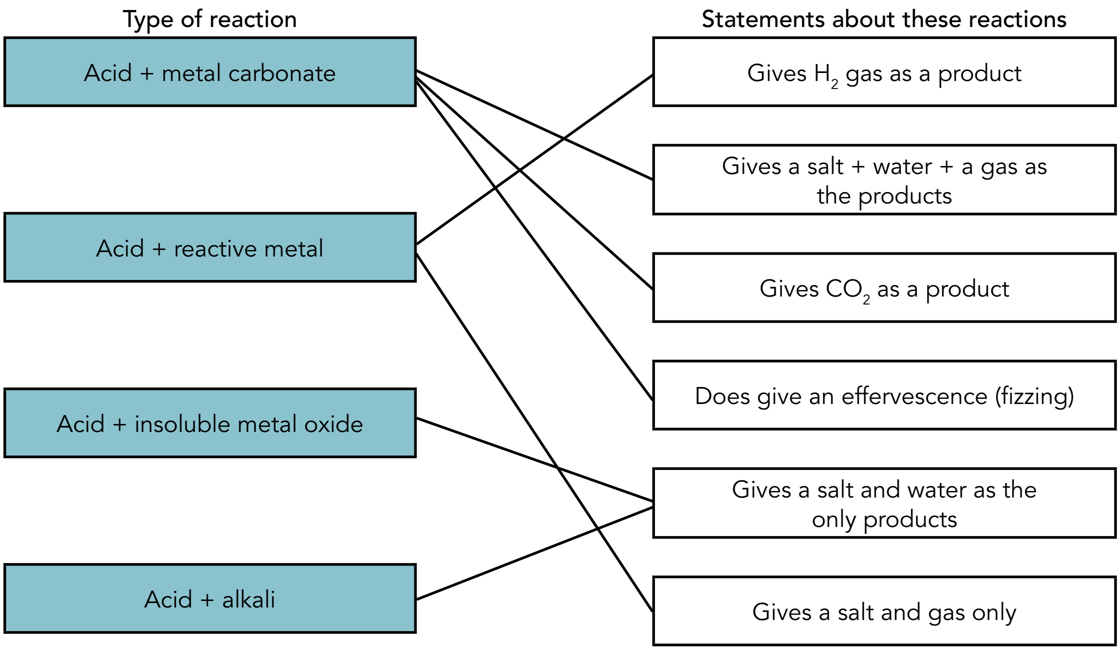
e a burette

f pH = 1.6

Chapter 11 continued

Exercise 11.2

4



5 a word equation: magnesium + hydrochloric acid → magnesium chloride + hydrogen

symbol equation:
$Mg + 2HCl \rightarrow MgCl_2 + H_2$

b word equation: calcium hydroxide + hydrochloric acid → calcium chloride + water

symbol equation:
$Ca(OH)_2 + 2HCl \rightarrow CaCl_2 + 2H_2O$

c word equation: calcium oxide + nitric acid → calcium nitrate + water

symbol equation:
$CaO + 2HNO_3 \rightarrow Ca(NO_3)_2 + H_2O$

d word equation: sodium hydroxide + sulfuric acid → sodium sulfate + water

symbol equation:
$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$

e word equation: sodium carbonate + nitric acid → sodium nitrate + carbon dioxide + water

symbol equation: $Na_2CO_3 + 2HNO_3 \rightarrow 2NaNO_3 + CO_2 + H_2O$

f word equation: magnesium + nitric acid → magnesium nitrate + hydrogen

symbol equation:
$Mg + 2HNO_3 \rightarrow Mg(NO_3)_2 + H_2$

6 a $Mg(OH)_2 + 2HCl \rightarrow MgCl_2 + 2H_2O$

b $MgCO_3 + 2HCl \rightarrow MgCl_2 + H_2O + CO_2$

c Take the sample of the antacid and place in a conical flask. Add some distilled water and make sure this is well mixed.

Add the methyl orange indicator to this mixture and place on the white tile.

Fill the burette to the zero mark with the 1 mol/dm^3 hydrochloric acid.

Slowly add the acid to the antacid/mixture until the methyl orange changes colour yellow to orange.

Note the volume of acid used and repeat with the other antacid. The antacid that requires the greater volume of hydrochloric acid is the more effective antacid.

d The magnesium carbonate reacts with the acid to form carbon dioxide, which will cause bloating and burping.

Chapter 11 continued

Exercise 11.3

7 a Experiment 1

Using a ~~**beaker**~~ / **pipette**, 10 cm^3 of the sodium hydroxide solution was placed in a **conical** / ~~**round-bottomed**~~ flask. Thymolphthalein indicator was added to the flask turning the solution ~~**red**~~ / **blue** because of the ~~**acidic**~~ / **alkaline** conditions. A **burette** / ~~**volumetric pipette**~~ was filled to the **0.0** / ~~**50.0**~~ cm^3 mark with hydrochloric acid (solution P). The acid was then run into the flask until the colour changed to ~~**red**~~ / **colourless** showing that the alkali in the flask had been **neutralised** / ~~**naturalised**~~ by the acid.

The volume of **acid** / ~~**alkali**~~ added was noted. The flask was washed thoroughly with ~~**tap**~~ / **distilled** water.

b Experiment 2

The experiment was repeated using the same volume of sodium hydroxide ~~**acid**~~ / **alkali** in the flask but a different solution of acid (solution Q) in the **burette** / ~~**volumetric pipette**~~.

c The white tile gives a background that makes it easier to distinguish colour changes.

8 a and b

Burette readings / cm^3	Experiment 1	Experiment 2
Final reading	10.6	21.4
Initial reading	0.0	0.2
Difference	10.6	21.2

9 a sodium hydroxide + hydrochloric acid → sodium chloride + water

b $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

c blue to colourless

d experiment 2

e The acid used in experiment 1 (P) was the more concentrated. Less of it was required to neutralise the same amount of NaOH.

f 10.6 cm^3. The same volume because although it has been diluted, the amount of NaOH is the same.

Exercise 11.4

10 a pure water: $H^+ = OH^-$

b acid solution: $H^+ > OH^-$

c alkali solution: $OH^- > H^+$

11

	Acid	Base	Alkali
A solution of the substance contains an excess of hydroxide (OH^-) ions.			✓
A solution of the substance contains an excess of hydrogen (H^+) ions.	✓		
The substance is a proton acceptor.		✓	✓
The substance is a proton donor.	✓		

Chapter 11 continued

12 a Hydrochloric acid is a ~~**weak**~~ / **strong** acid. Hydrogen chloride gas consists of **covalent** / ~~**ionic**~~ molecules. When they dissolve in water these molecules ~~**partly**~~ / **completely** dissociate into hydrogen and ~~**chlorine**~~ / **chloride** ions, producing as many hydrogen ions in the solution as possible.

b Ethanoic acid is a **weak** / ~~**strong**~~ acid. When it dissolves in water **some** / ~~**all**~~ of the molecules dissociate into ions. The majority of the molecules remain intact. This means that the concentration of hydrogen ions is ~~**more**~~ / **less** than it could be if all the molecules had dissociated into ions.

c Two solutions of hydrochloric acid and ethanoic acid have the same concentration. The hydrochloric acid solution will have the ~~**higher**~~ / **lower** pH value. It has the **higher** / ~~**lower**~~ concentration of hydrogen ions and is the **more** / ~~**less**~~ acidic solution.

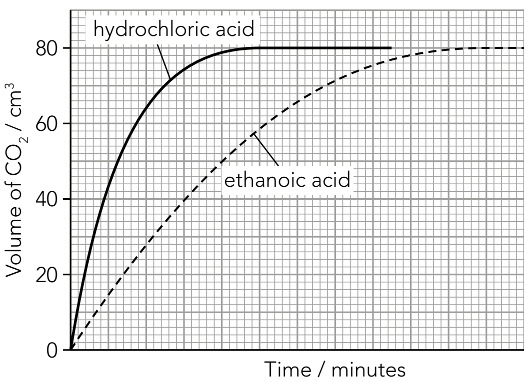
13 a i ethanoic acid:
$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$

ii hydrochloric acid:
$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$

iii sulfuric acid:
$H_2SO_4(aq) \rightarrow HSO_4^-(aq) + H^+(aq)$

b The reaction is reversible and will reach an equilibrium when the rate of the reverse reaction is equal to that of the forward reaction.

14 a



b It is the hydrogen ions from the acid that react with the calcium carbonate. The hydrochloric acid is the strong acid and therefore its concentration of hydrogen ions is greater than the weaker ethanoic acid. This means that the reaction with hydrochloric acid will be faster and end more quickly. This is shown by the steeper graph for hydrochloric acid.

c Hydrochloric acid will have the greater conductivity because it is a strong acid, completely dissociated and therefore has a higher concentration of ions which will conduct electricity.

d The hydrochloric acid loses a proton to the ammonia and is therefore a proton donor; the ammonia gains a proton and is therefore a proton acceptor.

Chapter 12

Exercise 12.1

1 a Hydrochloric acid always produces a **chloride**.

b Nitric acid always produces a **nitrate**.

c Phosphoric acid always produces a **phosphate**.

2

Substances reacted together		Salt produced	Other products of the reaction
dilute hydrochloric acid	zinc oxide	**zinc chloride**	**water**
dilute sulfuric acid	**copper carbonate**	copper sulfate	water and carbon dioxide
dilute sulfuric acid	**magnesium carbonate**	magnesium sulfate	water and carbon dioxide
dilute hydrochloric acid	**magnesium**	magnesium chloride	hydrogen
dilute nitric acid	copper oxide	**copper nitrate**	**water**
dilute ethanoic acid	**sodium hydroxide**	sodium ethanoate	water
dilute phosphoric acid	potassium hydroxide	potassium phosphate	**water**

3 All salts are ionic compounds. Salts are produced when an alkali neutralises an **acid**. In this reaction, the salt is formed when a **metal** ion or an ammonium ion from the alkali replaces one or more **hydrogen** ions of the acid.

Salts can be crystallised from the solution produced by the neutralisation reaction. The salt crystals formed often contain chemically **combined** water. These salts are called **hydrated** salts. The salt crystals can be heated to drive off this **water**. The salt remaining is said to be **anhydrous**.

Salts can be made by other reactions of acids. Magnesium sulfate can be made by reacting magnesium carbonate with **sulfuric** acid. The gas given off is **carbon dioxide**. Water is also formed in this reaction.

All **sodium** and potassium salts are soluble in water. Insoluble salts are usually prepared by **precipitation**.

4 a Sulfuric acid

Magnesium oxide/magnesium hydroxide/magnesium carbonate

Warm the acid. Add the base until it stops disappearing/dissolving.

- Filter.
- Warm the filtrate until crystals start to appear.
- Leave to cool so that the rest of the crystals are formed.

b Hydrochloric acid

Potassium hydroxide/potassium carbonate

Place accurate volume of solution of base in conical flask.

- Add a few drops of indicator.
- Add hydrochloric acid from burette until the indicator changes colour.
- Note volume of acid used and repeat same experiment without indicator.
- Warm solution formed until crystals start to appear.
- Leave to cool so that the rest of the crystals are formed.

Chapter 12 continued

5 **a** To make sure that all of the acid is reacted and none carries through into the stages of concentration and crystallisation.

b by filtration – it is the filtrate

c Heat the filtrate in an evaporating basin to concentrate the solution. Then allow to cool slowly to form crystals. Filter off the crystals and dry between filter papers.

d zinc + nitric acid ➞ zinc nitrate + hydrogen

$Zn(s) + 2HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + H_2(g)$

6 **a** E C B A F D

b Because heating strongly would risk driving off any water of crystallisation present which would destroy the structure of the crystals, giving a powder.

c water of crystallisation

Exercise 12.2

7 **a** soluble

b insoluble

c insoluble

d soluble

e insoluble

f soluble

8 **a** Reagents used copper(II) nitrate and potassium or sodium carbonate

$Cu(NO_3)_2(aq) + K_2CO_3 \rightarrow CuCO_3(s) + 2KNO_3(aq)$

b Reagents used silver nitrate and potassium or sodium iodide

$AgNO_3(aq) + KI(aq) \rightarrow AgI(s) + KNO_3(aq)$

c Reagents used silver nitrate and potassium or sodium chloride

$AgNO_3(aq) + KCl(aq) \rightarrow AgCl(s) + KNO_3(aq)$

d Barium nitrate and potassium or sodium sulfate

$Ba(NO_3)_2(aq) + K_2SO_4 \rightarrow BaSO_4(s) + 2KNO_3(aq)$

e One of the following equations

$Cu^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CuCO_3(s)$

$Ag^+(aq) + I^-(aq) \rightarrow AgI(s)$

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

$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

9 **a** No further precipitate is formed after $4\,cm^3$ of sodium phosphate solution is added.

Therefore the reacting ratio of the two solutions is 6 : 4 or 3 : 2.

The two solutions have the same concentration, so the molar ratio of metal X ions and phosphate ions is 3 : 2.

Formula of the phosphate of metal X is therefore $X_3(PO_4)_2$.

b $3X^{3+}(aq) + 2PO_4^{3-}(aq) \rightarrow X_3(PO_4)_2(s)$

c filter off the precipitate as the residue / wash the precipitate / dry carefully in warm oven

Chapter 13

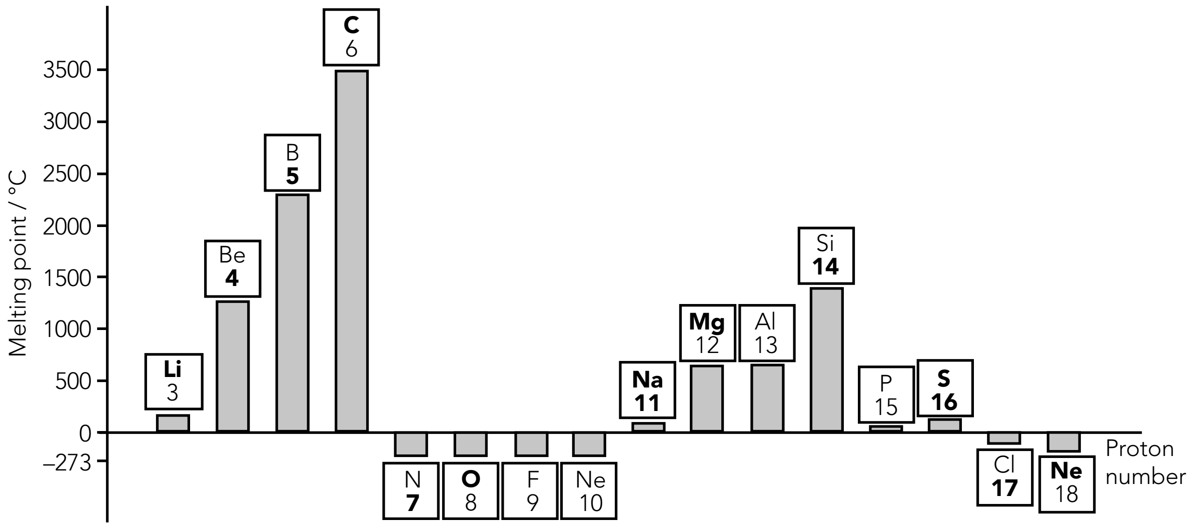
Exercise 13.1

1 The Periodic Table is a way of arranging the elements according to their properties. They are arranged in order of their **proton** number. Elements with similar properties are placed together in **vertical columns** called **groups**. Periods are **horizontal rows** of the elements. The table shows trends down the **groups** and patterns across the **periods**.

The placing of the elements in the table also corresponds to their **electron** arrangements (electronic configurations). The number of **electrons** in the outer **electron** shell is the same as its **group** number in the table. The number of occupied **electron** shells of the element is the **period** in which it is placed.

2 a Cu and Cr
b He
c Br
d C
e S
f He, Ne and Kr
g Cu
h Ca
i Mg

3 a



b C and Si (carbon and silicon)
c Group 4 (IV)
d When the melting points of elements are plotted against **atomic (or proton) number** they show a repeating **pattern**. The highest melting points are given by the elements of Group number **4 / IV** and the lowest by the **noble** gases.

Exercise 13.2

4 a Cs
b Na
c one electron in outer shell
d conductor, lustrous, reactive, soft

5 To stop them reacting with water and air.

6 a two of: float on surface of water; move around; fizz; both melt; disappear at end of reaction
b The potassium bursts into (lilac) flame. The sodium does not.
c $2Na + 2H_2O \rightarrow 2NaOH + H_2$

7 a two of: grey solid; conducts electricity; soft/easily cut; lustrous/shiny
b 1

Chapter 13 continued

8

Group I metal	Density / g per cm^3	Radius of metal atom / nm	Boiling point / °C	Reactivity with water
sodium	0.97	0.191	883	floats and fizzes quickly on the surface, disappears gradually and does not burst into flame
potassium	0.86	0.235	760	**reacts instantly, fizzes and bursts into flame, may spit violently**
rubidium	1.53	0.250	686	reacts instantaneously, fizzes and bursts into flame then spits violently and may explode
caesium	1.88	**0.255–0.265** **(actual value 0.260)**	**620–650** **(actual value 671)**	**reacts instantly and explosively**

Exercise 13.3

9 a The halogens are ~~**metals**~~ / **non-metals** and their vapours are **coloured** / ~~**colourless**~~.

b The halogens are **toxic** / ~~**non-toxic**~~ to humans.

c Halogen molecules are each made of ~~**one**~~ / **two** atoms; they are ~~**monatomic**~~ / **diatomic**.

d Halogens react with **metal** / ~~**non-metal**~~ elements to form crystalline compounds that are salts.

e Halogens can ~~**colour**~~ / **bleach** vegetable dyes and kill bacteria.

10 a

Halogen	Reaction with iron wool
Chlorine	When iron wool was lowered into a gas jar, a very exothermic reaction could be seen and dark red solid formed.
Bromine	The iron wool had to be heated at first, but there was a very **exothermic** reaction and a **dark red** solid was formed.
Iodine	Lots of heat was needed and a small amount of **heat** was given out. A **dark red** solid formed.

b The elements get less reactive as we descend the group.

c $\mathbf{2}Fe(s) + \mathbf{3}Cl_2(g) \rightarrow \mathbf{2}FeCl_3(s)$

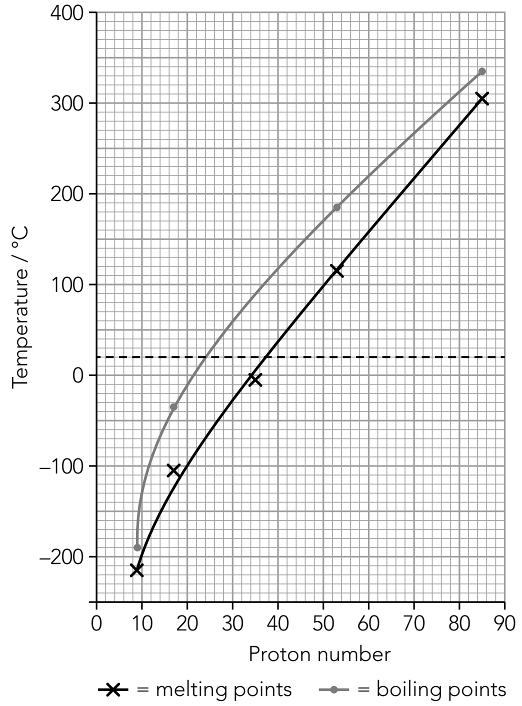
d $2Fe(s) + 3Br_2(g) \rightarrow 2FeBr_3(s)$

11 a

Element	Proton number	Melting point / °C	Boiling point / °C	Colour
fluorine	9	–219	–18	pale yellow
chlorine	17	–101	–34	pale green
bromine	35	–6		**red/dark red**
iodine	53	114	185	grey–black
astatine	85	303	337	**black**

Chapter 13 continued

b



c estimated boiling point = 90–115 °C (actual value 114 °C)

colour: red–brown

physical state: liquid

d fluorine and chlorine

e solid, black

f The melting points increase as you go down the group.

Exercise 13.4

12

	True	False
Some transition elements are non-metals.		✓
Transition elements have high densities.	✓	
Some transition elements make good catalysts.	✓	
Transition elements form only white compounds.		✓
All transition elements are magnetic.		✓

13 **a** copper

b platinum

c copper

d iron

e manganese

f chromium

g silver, gold

14 **a** oxidation number

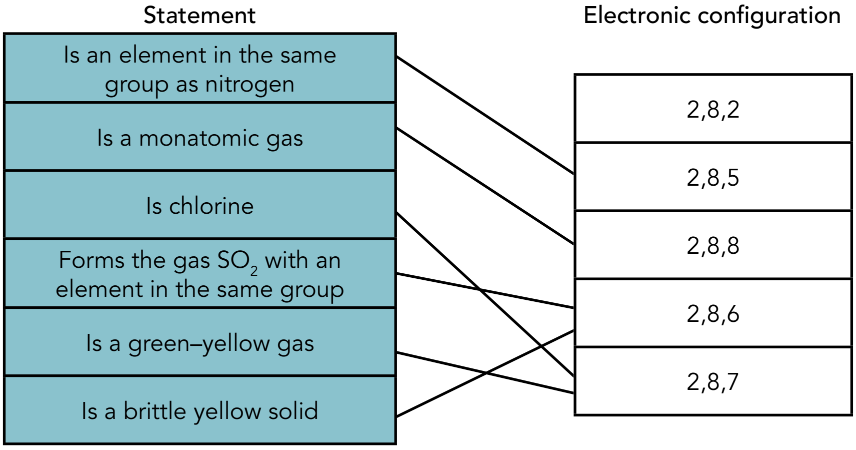
b variable oxidation number

c iron(II) oxide: FeO
iron(III) oxide: Fe_2O_3

Chapter 13 continued

Exercise 13.5

15



16 a

Noble gas	Density g/dm³	Atomic number	Period	Electronic configuration
Helium	0.18	2	1	**2**
Neon	0.90	10	2	2,8
Argon	1.78	18	**3**	2,8,**8**
Krypton	3.75	36	**4**	2,8,18,8
Xenon	5.89	54	5	2,8,18,18,8

b The period number (row number) is equal to the number of occupied shells (energy levels) in the atom of the element.

c Helium is used in balloons because it is (a lot) less dense than air.

d It will sink because argon is more dense than air.

17

Name of element	Sulfur	Selenium	Tellurium
Density / g per cm³	2.07	**4.79**	6.24
Melting point / °C	115	221	**450**
Boiling point / °C	445	**685**	988
Ionic radius / nm	0.184	**0.198**	**0.221**

Chapter 14

Exercise 14.1

1 1 = strong 2 = high 3 = sonorous 4 = they can be beaten into shape 5 = ductile

2

Description of physical property	Name for this property	Use that depends on this property
They can be moulded or bent into shape.	**malleability**	**car bodies**
They transfer heat well.	thermal conductivity	**cooking pots and pans**
They can be drawn into wires.	**they are ductile**	**cabling and wiring**
They conduct electricity.	**electrical conductivity**	**electrical wiring**
They are heavy for their volume.	**high density**	**building construction**
They can bear weight and are not broken easily.	**strength**	**bridges**
They make a ringing sound when struck by a hammer.	**they are sonorous**	**bells**

3 **a** Metals show good electrical conductivity whereas non-metals are **poor conductors of electricity**.

b Metals are malleable and ductile but non-metals are **brittle**.

c Metals are good conductors of heat but most non-metals are **thermal insulators**.

d Metals are usually grey in colour and can be polished whereas non-metals are **a wide range of different colours and cannot be polished**.

4 **a** Aluminium has a high electrical conductivity.

b The low density of aluminium means that the overhead cables do not sag too much between pylons.

c Copper is an even better conductor of electricity and the weight of the wiring is not a factor.

d Titanium is light and strong – it is also expensive, so can only be used for specialist purposes.

5 **a** magnesium > zinc > iron > copper
Magnesium produced the most vigorous bubbling of gas when added to the acid, copper produced no gas at all. Zinc produced more effervescence than iron.

b hydrogen

c zinc + hydrochloric acid → zinc chloride + hydrogen

6 **a** magnesium oxide

b B is sodium / sodium hydroxide is the solution produced

c A < C < B

d B

e magnesium oxide, MgO

Chapter 14 continued

Exercise 14.2

7 **a** iii **b** vi **c** v **d** ii **e** i **f** iv

8

Alloy	Composition	Use	Useful property
mild steel	iron: >99.75% carbon: <0.25%	**car bodies**	**can be beaten into shape (malleable)**
stainless steel	iron: 74% **chromium**: 18% nickel: 8%	**cutlery**, surgical instruments, chemical vessels for industry	resistant to corrosion (does not rust easily)
brass	copper: 70% **zinc**: 30%	**musical** instruments, ornaments	'gold' colour, harder than copper
bronze	copper: 95% **tin**: 5%	statues, church bells	hard, does not **corrode**
aerospace aluminium	aluminium: 90.25% zinc: 6% magnesium: 2.5% copper: 1.25%	aircraft construction	**light (low density) but strong**
solder	tin: 60% lead: 40%	**joining wires in electrical circuitry**	low melting point
tungsten steel	iron: 95% tungsten: 5%	cutting edges of drill bits	**very hard**

9 **a** In a pure metal the layers can slide over each other so the metal can be shaped without breaking.

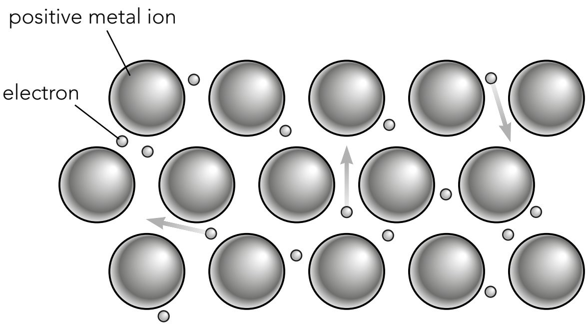
b Solder has a lower melting point and so it is easier to melt. It is also stronger and so the joints will be stronger.

c chromium and nickel / they are transition metals

d The presence of the different sized atoms of the added metal in the alloy means that the layers cannot slide over each other as easily as in the pure metal / the alloy is stronger and less malleable.

e The presence of the added iron and silicon atoms make the commercial aluminium stronger and harder.

f



The copper used in wiring needs to be very pure so that the delocalised electrons can move as freely as possible through the structure. The presence of impurities would interfere with this flow.

g **i** The pins of the electric plugs have to be strong, so brass of composition Cu60:Zn40 is used.

ii Brass instruments have to be shaped into tubes, so a softer alloy is needed. Cu70:Zn30 is therefore used.

Chapter 15

Exercise 15.1

1 a All the metals in the group react with cold water to give an alkaline solution of the metal hydroxide.

b sodium: 2,8,1

potassium: 2,8,8,1

c They each contain one electron in their outer (valency) shell.

d Potassium is the more reactive, as the energy generated by the reaction is sufficient to ignite the hydrogen produced.

e potassium + water ➞ potassium hydroxide + hydrogen

f $2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$

2 a The sodium melts into a ball.

b Yellow, because of the sodium ions present / yellow is the flame test colour for sodium.

c magnesium + water (steam) ➞ magnesium oxide + hydrogen

$Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$

3 a two outer electrons

b

Group II metal	Density / g per cm^3	Radius of metal atom / nm	Boiling point / °C	Reactivity with water and steam
magnesium	1.74	0.173	1090	reacts very slowly with cold water, but reacts strongly with steam
calcium	1.54	0.231	1484	reacts strongly with cold water, unsafe to react with steam
strontium	2.64	0.249	1377	**reacts strongly with cold water, unsafe to react with steam**
barium	3.62	0.268	1845	reacts strongly with cold water, unsafe to react with steam

c Atomic size increases as you descend the group.

d Calcium appears to show values for density and boiling point that do not follow the general trends in the group.

e calcium + water ➞ calcium hydroxide + hydrogen

$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$

f In Groups I and II the metals become more reactive with water as you descend the groups.

Chapter 15 continued

Exercise 15.2

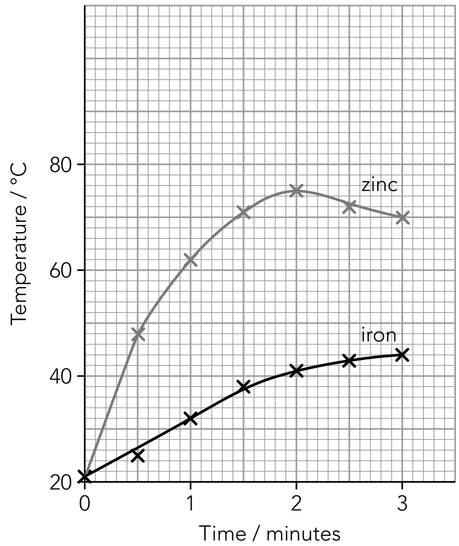
4 a When metals react with **cold** water, the products are a metal hydroxide and **hydrogen**. The hydroxides formed are **alkaline** and they will turn **red** litmus blue. Some metals do not react with cold water but do with **steam**. The products in this reaction are the metal **oxide** and hydrogen. Copper does not react with water or with dilute **acids**. This is because it is **lower** in the reactivity series than hydrogen.

5 a zinc + steam (water) ➞ zinc oxide + hydrogen

b copper or silver (or another metal low in the series)

c iron, zinc or magnesium (not calcium or sodium, etc., because these are too reactive to be safe)

6 a



b the reaction between zinc and copper(II) sulfate

c because the reactions have finished and no more heat is given out

d zinc + copper sulfate ➞ zinc sulfate + copper

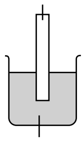
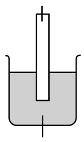
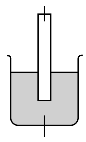
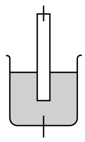
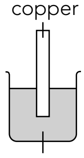
$Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$

e Zinc gave the higher temperature rise because it is the more reactive metal.

f This experiment seems to be a 'fair test' in that most of the variables except the choice of metal are controlled. One difficulty is whether the two metals were powdered to the same extent so the surface area available for the reaction will be different in each case. Note that, although 5 g is not an equal number of moles of the two metals, it is an excess in both cases.

Chapter 15 continued

7 a

		zinc  tin(II) chloride solution	zinc  copper(II) sulfate solution	tin  copper(II) sulfate solution	silver  copper(II) sulfate solution	copper  silver nitrate solution
At start	Colour of metal	grey	**grey**	silver coloured	silver coloured	**brown**
	Colour of solution	colourless	**blue**	blue	blue	colourless
At finish	Colour of metal	coated with silver-coloured crystals	**coated with brown solid**	coated with brown solid	silver coloured	coated with silver-coloured crystals
	Colour of solution	colourless	**colourless**	colourless	blue	**blue**

b zinc > tin > copper > silver

c The first equation is correct as lanthanum is more reactive than aluminium and will displace aluminium from aluminium sulfate (reduce Al^{3+} ions to Al atoms).

d transition metals

e They can be good catalysts; they form coloured compounds.

f Aluminium becomes coated with a very thin layer of aluminium oxide which masks the true reactivity of the metal.

8 a Magnesium is less ready to form a positive ion than sodium because there are more protons in the nucleus attracting the electrons.

b Potassium forms a positive ion more easily because the outer electrons are further from the nucleus and this outweighs the fact that there are more protons in the nucleus / there is also one more inner shell of electrons in potassium.

Chapter 16

Exercise 16.1

1 a–c

Metal	Reactivity	Method of extraction	Energy needed to extract the metal	Cost of extracting the metal
sodium	most ↑	extraction by electrolysis	most ↓	highest ↓
calcium				
magnesium				
aluminium				
carbon				
zinc		reduction by heating with carbon		
lead				
copper				
silver				
gold	least		least	lowest

2 copper and silver

3 zinc

4 no, zinc is more reactive than hydrogen

5 Sodium can be obtained by electrolysis of the molten ore (energy costs high).

6 a Sodium is a good reducing agent, as it has a strong tendency to form positive ions; it does this more readily than aluminium and so can be used to extract aluminium from aluminium chloride.

b i $3Na(s) + AlCl_3(s) \rightarrow 3NaCl(s) + Al(s)$

ii $3Na(s) + Al^{3+}(s) \rightarrow 3Na^+(s) + Al(s)$

7 a exothermic reactions

b aluminium + iron(III) oxide → aluminium oxide + iron

$2Al(s) + Fe_2O_3(s) \rightarrow Al_2O_3(s) + 2Fe(s)$

c iron < aluminium < sodium

Exercise 16.2

8 a E **b** C **c** A **d** B **e** D

9 blast furnace: because hot air is blown (blasted) into the furnace at the base

10 a hematite

b at the centre of the furnace, half-way up

c carbon monoxide (CO) produced from burning the coke

d iron(III) oxide + carbon monoxide → iron + carbon dioxide

e The iron is liquid at the temperature in the furnace and so it flows down to the bottom.

11 a $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(l) + 3CO_2(g)$ [state symbols as apply in the furnace]

b $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

12 a SiO_2

b CaO

c $CaO(s) + SiO_2(s) \rightarrow CaSiO_3(l)$ [state symbols as apply in the furnace]

13 a Zinc has a lower boiling point than iron and it boils at the temperatures in the furnace. The zinc vapour is collected and condensed at the top of the furnace.

b The liquid zinc collected at the top of the furnace will be pure as it passed through the furnace as a gas. Iron remains as a liquid at the furnace temperature and flows down to the bottom. The liquid iron dissolves impurities as it flows down through the furnace.

Exercise 16.3

14 a bauxite

b Al_2O_3, Al^{3+} and O^{2-}

c The aluminium oxide must be molten so that the ions are able to move to the electrodes and be discharged.

d i oxygen

ii aluminium

e any one correct, e.g. aircraft, food containers, overhead power cables

Chapter 16 continued

15 a B

b A

c C outer casing, D molten aluminium

d Cryolite is a solvent in which the aluminium oxide is dissolved; it lowers the melting point of the aluminium oxide and so reduces the energy costs of the extraction.

e They burn away in the oxygen discharged by the electrolysis.

f $Al^{3+}(l) + 3e^- \rightarrow Al(l)$

$2O^{2-}(l) \rightarrow O_2(g) + 4e^-$

16 a Aluminium oxide is amphoteric and so reacts with the sodium hydroxide; iron oxide is basic and does not react.

b (thermal) decomposition

Exercise 16.4

17 a Anhydrous calcium chloride is a drying agent to dry the air in the tube.

b To boil any dissolved air (oxygen) out of the water.

c To prevent any air dissolving in the water.

d tube 4

e Dissolved salt would speed up the rusting process (seawater makes rusting take place faster).

18 a air (oxygen) and water

b Two of: painting, coating in plastic, oiling, plating with a metal.

19 a The can could rust all the way through.

b The zinc, being more reactive, would corrode in preference to the steel.

c It is reactive and could react with acids in the food.

d Tin is unreactive and does not react with acids.

20 a Lead is poisonous, and a little could dissolve and contaminate the food.

b If the layer of tin is broken it causes the iron beneath to corrode more rapidly. The tin coating is less reactive and no longer protects the can.

21 Aluminium is not as strong and is more expensive.

22 a galvanisation

b Zinc is more reactive than iron and so will corrode away in preference to the iron / the zinc provides sacrificial protection / the zinc forms positive ions more readily than the iron.

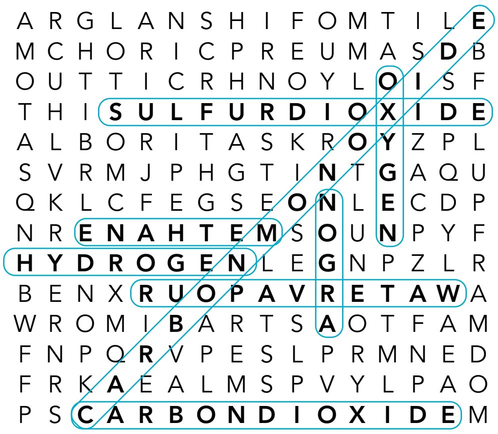
23 a magnesium

b $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

Chapter 17

Exercise 17.1

1 a



Formula	Name of gas	Found in clean dry air	Considered a pollutant
Ar	**argon**	✓	
CO_2	**carbon dioxide**	✓	✓
CO	**carbon monoxide**		✓
H_2	**hydrogen**		
CH_4	**methane**		✓
N_2	**nitrogen**	✓	
NO_2	**nitrogen dioxide**		✓
O_2	**oxygen**	✓	
SO_2	**sulfur dioxide**		✓
H_2O	**water vapour**		

b hydrogen

c oxygen / 21%

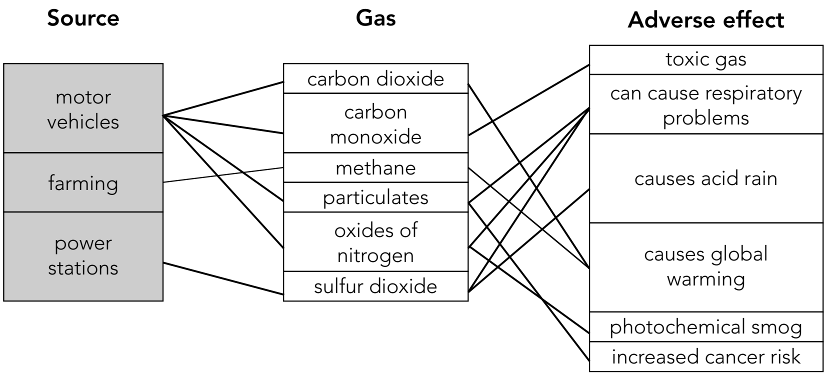
d nitrogen and nitrogen dioxide

e nitrogen and argon

f See completed table in **a**.

g water

2 a



b Carbon monoxide decreases the capacity of the blood to carry oxygen.

c sulfuric acid, nitric acid (accept sulfurous acid, nitrous acid also)

d methane

e (anaerobic) decay of vegetation in marshes and paddy fields / (anaerobic) decay of organic material in landfill sites / digestion in livestock

3 a A and B: both are major contributors of carbon dioxide. Possibly D: a major contributor to methane. Carbon dioxide and methane are greenhouse gases.

b C and F: both can be done without impacting on the human population. The technology is already available.

c The easiest methods to implement do not impact on everyday human life but do not make a major difference to levels of pollution. The more difficult methods must be attempted, but will have a wider range of impacts on society.

4 a Harmful nitrogen oxides (NO_x) arise from nitrogen and oxygen from the air reacting in the hot engine of the vehicle. (Diesel engines produce higher levels of NO_x as they have a higher operating temperature.)

b $2CO(g) + 2NO(g) \rightarrow \mathbf{2CO_2(g) + N_2(g)}$

c Nitrogen(II) oxide is acting as an oxidising agent.

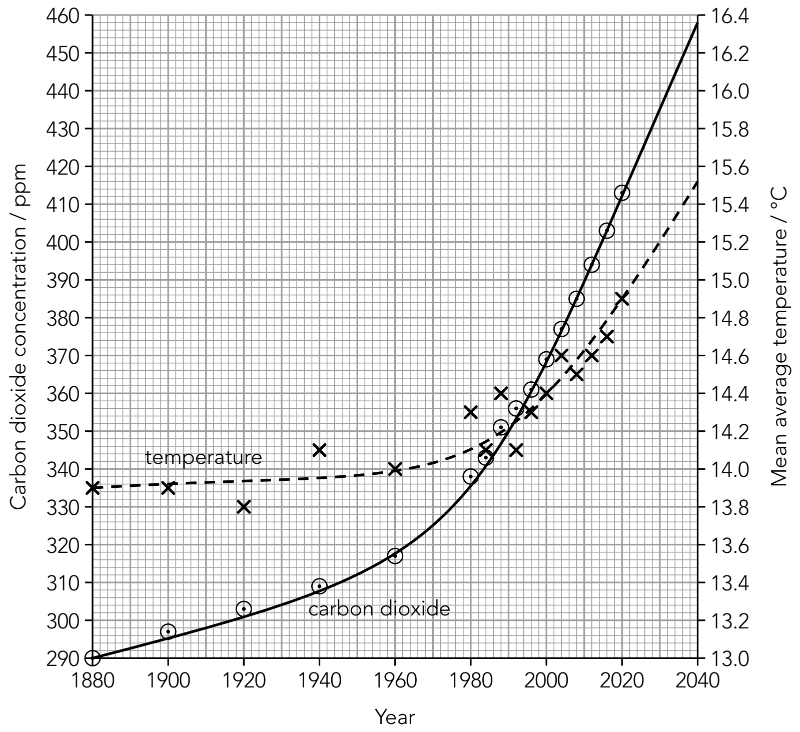
Chapter 17 continued

Exercise 17.2

5 The **greenhouse** effect has kept the Earth at a temperature suitable for the development of **life** for many thousands of years. As **industrial** activity has increased during the 20th century, more and more greenhouse gases have been released into the **atmosphere**. Carbon dioxide and methane are the two **gases** causing the greatest problem. **Methane** is 25 times more potent as a greenhouse gas but **carbon dioxide** is present in greater quantities.

Because of these gases, more of the heat from the sun is kept within the Earth's atmosphere and this causes **global warming**. Much of this heat warms **water** in the oceans, which cover about 70% of the Earth's surface. Increased temperatures in the oceans and in the atmosphere have an effect on the Earth's climate. This effect means that extreme **weather** events are more likely.

6 a



b i Carbon dioxide: there has been a steady increase in the level of carbon dioxide in the atmosphere since 1880. However, the curve has increased more sharply since the 1960s. This rate of increase has remained steady over recent decades.

ii Mean temperature: the change in mean temperature is more variable, showing more peaks and troughs. However, since the 1940s, the broad trend is for the mean temperature to increase.

Chapter 17 continued

c No. It was only a minor fluctuation as the trend is still upwards.

d carbon dioxide = 450–460 ppm / mean temperature = 15.4–15.5 °C (values depend on extrapolated line drawn)

7 a glaciers melt: sea level rises / coastal flooding more likely

b lack of rain: vegetation dries out / fire spreads more rapidly

c Dependent on student answers:

Problems: increased land temperatures / droughts, etc.

Causes: industrialisation / burning fossil fuels / intensive livestock farming / deforestation, etc.

Solutions: increasing use of hydrogen and renewable energy / planting trees / decreasing reliance on fossil fuels / reduction in livestock farming, etc.

8 a Photosynthesis

b carbon dioxide + water → glucose + oxygen

c $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$

9 The greenhouse effect is a **natural** phenomenon that warms the surface of the Earth.

When thermal **energy** (**infrared** radiation) from the **Sun** reaches the Earth's **atmosphere**, around 30% is **reflected** back to space and around 70% is **absorbed** by the oceans and land to heat the planet. Some of this heat is then **radiated** back up into the atmosphere.

Greenhouse **gases** in the atmosphere such as carbon dioxide and **methane** can absorb this infrared radiation and then **re-emit** it back toward the Earth. This reduces heat **loss** to space and keeps the Earth's **temperature** warm enough to sustain life.

Human **activities** are increasing the amount of greenhouse gases released into the atmosphere, **trapping** extra heat and causing **global** temperatures to rise. This warming of both the atmosphere and the **oceans** gives rise to the different aspects of **climate** change.

Chapter 17 continued

Exercise 17.3

10

Stage of treatment	What is removed
passed through a coarse screen	**A, B**
sedimentation and filtration	**E, H**
chlorination	**F**
not removed in water treatment	**C, D, G**

11

Test	Seawater	River water
Add to anhydrous copper sulfate	**Changed from white to blue**	**Changed from white to blue**
Test with cobalt chloride paper	**Changed from blue to pink**	**Changed from blue to pink**
Measure the boiling point	**Boiled at 103–105 °C**	**Boiled at slightly over 100 °C**
Boil away a sample of water	**There was a lot of white solid remaining**	**There was very little solid remaining**

12 Distilled water contains no dissolved substances. Even very small amounts of dissolved impurities in water can interfere with science experiments.

13 a It has been filtered as it flows through the ground / it has not flowed through fields and towns.

b No: only water that is consumed by people needs to be free from microbes.

c Some minerals are essential for health, and spring (well) water can provide some of these minerals.

d Boiling will kill the harmful microbes in the water.

e Plastic bottles can be discarded, causing environmental pollution.

f Sterilising tablets can be added to water / water can be boiled before consumption / filter straws, etc.

Chapter 18

Exercise 18.1

1 The chief source of organic compounds is the naturally occurring mixture of hydrocarbons known as **petroleum**. Hydrocarbons are compounds that contain carbon and **hydrogen** only. There are many hydrocarbons because of the ability of carbon atoms to join together to form long **chains**. There is a series of hydrocarbons with just single covalent bonds between the carbon atoms in the molecule. These are saturated hydrocarbons, and they are called **alkanes**. The simplest of these saturated hydrocarbons has the formula CH_4 and is called **methane**. Unsaturated hydrocarbons can also occur. These molecules contain at least one carbon–carbon **double** bond. These compounds belong to the **alkenes**, a second series of hydrocarbons. The simplest of this 'family' of unsaturated hydrocarbons has the formula C_2H_4 and is known as **ethene**.

The test for an unsaturated hydrocarbon is to add the sample to **bromine** water. It changes colour from orange–brown to **colourless** if the hydrocarbon is unsaturated.

2 a

Name	Formula	Boiling point / °C
ethene	C_2H_4	–102
propene	C_3H_6	–48
butene	C_4H_8	–7
pentene	C_5H_{10}	30
hexene	**C_6H_{12}**	**60 (58–62)**

b $C_{12}H_{24}$

c C_nH_{2n}

d A homologous series is a series of organic compounds that have the same general formula and similar chemical properties because they contain the same functional group.

e the carbon–carbon double bond

Exercise 18.2

3 a alkanes

b alcohols

c carboxylic acids

d alkenes

4 a butane

b methanoic acid

c butanol

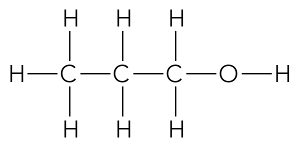
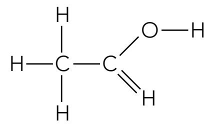
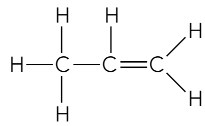
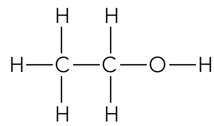
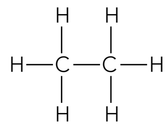
5 a C_nH_{2n+2}

b $C_nH_{2n+1}OH$

c $C_nH_{2n+1}COOH$

Chapter 18 continued

6

Name of compound	Homologous series	Molecular formula	Displayed formula
propanol	**an alcohol**	**C_3H_7OH**	
ethanoic acid	**a carboxylic acid**	CH_3COOH	
propene	**an alkene**	C_3H_6	
ethanol	an alcohol	**C_2H_5OH**	
ethane	an alkane	C_2H_6	

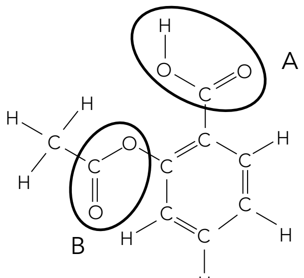
Exercise 18.3

7

Compound	Molecular formula	Structural formula	Empirical formula
butane	C_4H_{10}	**$CH_3CH_2CH_2CH_3$**	C_2H_5
propene	C_3H_6	$CH_3CH{=}CH_2$	**CH_2**
ethanol	C_2H_5OH	**CH_3CH_2OH**	**C_2H_6O**
propanoic acid	C_2H_5COOH	**CH_3CH_2COOH**	**$C_3H_6O_2$**
propyl ethanoate	$CH_3COOC_2H_5$	**$CH_3COOCH_2CH_3$**	**C_2H_4O**

8 **a** because it is an alkane with just carbon–carbon single bonds

b and c



Aspirin

d the hydroxyl group (–OH); the functional group present in alcohols

Chapter 18 continued

9 **a** alkenes
$CH_3CH_2CH_2CH{=}CH_2$
(or $CH_3CH_2CH{=}CHCH_3$)

b (structural) isomerism

10 **a** butane, $CH_3CH_2CH_2CH_3$

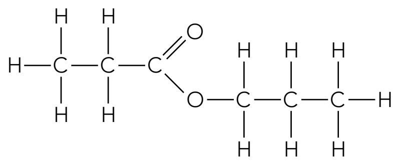
b butan-2-ol, $CH_3CH_2CH(OH)CH_3$

c but-1-ene, $CH_3CH_2CH{=}CH_2$

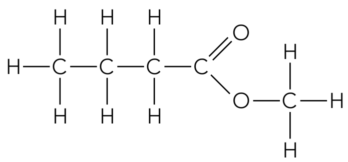
d 1,1-dibromoethane, CH_3CHBr_2

e ethyl ethanoate, $CH_3COOCH_2CH_3$

11 **a** propyl propanoate,
$CH_3CH_2COOCH_2CH_2CH_3$



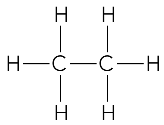
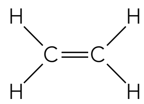
b methyl butanoate, $CH_3CH_2CH_2COOCH_3$



Chapter 19

Exercise 19.1

1 a

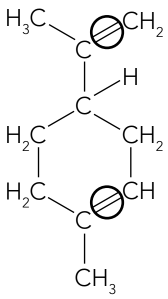
Name of hydrocarbon	ethane	ethene
Molecular formula of hydrocarbon	C_2H_6	C_2H_4
Saturated / unsaturated	**saturated**	**unsaturated**
Displayed formula of hydrocarbon		
Relative molecular mass of hydrocarbon	**30**	**28**

b i Ethane is saturated because all the C—C bonds are single covalent bonds.

ii Ethene is unsaturated because it contains one C=C bond.

c Alkenes C=C

2 a



b $C_{10}H_{16}$

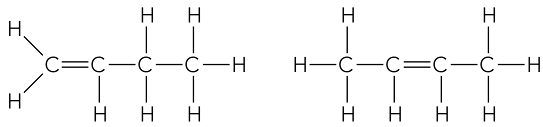
c $10 \times 12 + 16 \times 1 = 136$

d The colour changes from orange–brown to colourless (not clear).

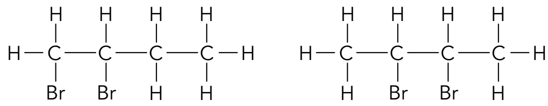
e A thermometer, B (water-cooled) condenser / Liebig condenser, C measuring cylinder

f burning in an insufficient (limited) supply of air (or oxygen)

3 a

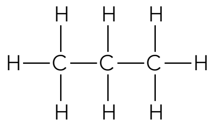


b

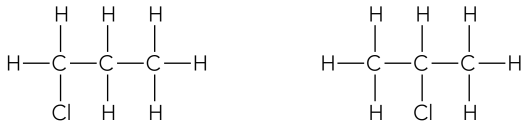


c $C_4H_8 + Br_2 \rightarrow C_4H_8Br_2$

4 a



b

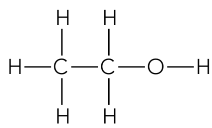


1-chloropropane **2-chloropropane**

c A reaction that is started by light (including ultraviolet light) / the light energy provides the activation energy for the reaction.

Exercise 19.2

5 a



b C_3H_7OH or C_3H_8O

c 3×12 (3 carbons) + 7×1 (7 hydrogens) + $16 + 1$ (for OH group)

$= 36 + 7 + 17 = 60$

6 a retinol, cholesterol and geraniol

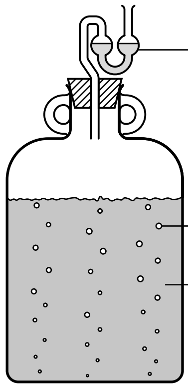
b alkenes

7 a The water has to be in gaseous form and therefore the temperature has to be above the boiling point of water.

b If a high pressure is applied to the chemical system at equilibrium, the equilibrium will shift to lower the pressure and does this by making fewer gas molecules which means more ethanol.

Chapter 19 continued

8 a



b Ethanol is toxic to the yeast and as more is produced it eventually kills the yeast.

c Ethanol is a product of anaerobic respiration – in the presence of oxygen a different reaction takes place / in the presence of oxygen the ethanol may be oxidised.

d Advantage using ethene: can be made using continuous flow method so do not need to stop and start / can make 100% ethanol.

Advantage using sugar: cheap – the materials are easily available and doesn't use fossil fuels.

9 a high temperature and a catalyst

b $C_9H_{20} \rightarrow C_5H_{12} + C_4H_8$

c steam (H_2O)

10 a $C_4H_9OH + \mathbf{6}O_2 \rightarrow \mathbf{4}CO_2 + \mathbf{5}H_2O$

b butan-1-ol + propanoic acid → butyl propanoate + water

c $CH_3CH_2COOCH_2CH_2CH_2CH_3$

Exercise 19.3

11 a $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$

b carbon monoxide and water

c burns to give out lots of heat energy; easy to light; easily available, e.g. from renewable sources; burns cleanly with no solid waste; is a liquid at room temperature and therefore easy to transport

d Ethene is formed from the cracking of fossil fuels, and therefore the production of ethanol by this method makes it a non-renewable fuel. Fossil fuels are non-renewable because they take millions of years to form, so once they are used up there is no more. If ethene is made from glucose, all the sources are from plants and are therefore renewable because more can be planted and grown.

12 a Measurements made are: weighing the masses of ethanol and the burner before and after burning; weighing the water and copper calorimeter to make sure the same amount of water is used each time; temperature measurements.

b Errors are due to heat loss and incomplete combustion.

c $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$

d $x = \frac{13.8}{2.3} \times 2.7 = 16.2$ g of water

13 a A: ethanol; B: $CH_3CH_2CH_2CH_2CH_2OH$; C: 88

b The independent variable is either the number of carbons or the relative molecular mass.

The dependent variable is the boiling point.

To estimate the boiling points of propan-1-ol and hexan-1-ol:

Plot the points; draw line of best fit whether it be a curve or straight line.

- From the line, see what temperature corresponds with three carbons or a relative molecular mass of 60.
- Extend line on for six carbons or relative molecular mass of 102.

c The value for hexan-1-ol, because the value for propan-1-ol lies on the line already given. The hexan-1-ol is an estimate by extrapolation of what the value would be.

Chapter 19 continued

Exercise 19.4

14 **a** A **b** C and E **c** D
d B and D **e** A **f** C_2H_4O

15 **a** H_2O

b 1 kg = 1000 g: M_r(ethanol) = 46; M_r (CH_3COOH) = 60
Number of moles of ethanol = $\frac{920}{46} \times 10^3$ = 2×10^4 (20 000)

From the equation the number of moles of ethanoic acid = number of moles of ethanol = 20 000. Therefore, mass of ethanoic acid formed = 20 000 × 60 = 1 200 000 or 1.2×10^6 g or 1200 kg.

c ethanoic acid + sodium carbonate → sodium ethanoate + carbon dioxide + water

16 **a** If the condenser is returned to the usual slanting position, the reaction mixture will not be returned to the flask and some of the ethanol will not be oxidised to ethanoic acid. Some of the ethanol will distill over and not be available to react.

b $C_2H_5OH + \mathbf{2}[O] \rightarrow CH_3COOH + \mathbf{H_2O}$

c The group knew that they needed **ethanoic** acid as well as ethanol to prepare ethyl ethanoate. So, they took the ethanol and divided it into two portions. One portion was oxidised to **ethanoic** acid by adding it to **acidified** potassium **manganate**(VII). The mixture was then refluxed for an hour and the **ethanoic** acid produced was distilled off. The ethanoic acid was then mixed with the **ethanol** and some concentrated **sulfuric** acid added. The concentrated **sulfuric** acid acted as a **catalyst** for the reaction. After the mixture was refluxed together, the ester **ethyl** ethanoate was distilled off.

d $CH_3COOH(l) + C_2H_5OH(l) \rightleftharpoons CH_3COOC_2H_5(l) + \mathbf{H_2O}(l)$

Chapter 20

Exercise 20.1

1 A: refinery gas / heating and cooking

B: gasoline (petrol) / fuel for cars

C: naphtha / for making chemicals (chemical feedstock)

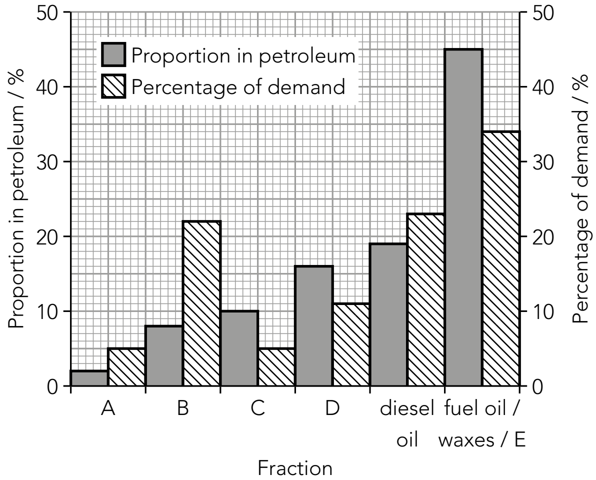
D: kerosene (paraffin) / fuel for aircraft (jet fuel)

E: bitumen / road surfacing

2 **a** boiling point

b Chain length decreases and viscosity decreases as you ascend the tower.

3 **a**

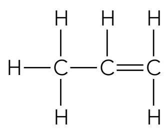


b B (gasoline), diesel oil; E (bitumen), fuel oil and waxes

c B (gasoline) + diesel oil = 22 + 23 = 45%

4 **a** $C_{15}H_{32} \rightarrow C_{12}H_{26} + C_3H_6$

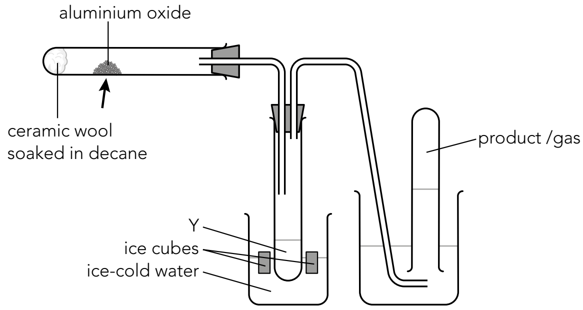
b



5 **a** $C_{10}H_{22}(l) \rightarrow C_8H_{18}(l) + \mathbf{C_2H_4(g)}$

decane octane **ethene**

b



c It is a catalyst.

d See diagram.

e Y = octane.

$2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$

f Test: add bromine water and mix.

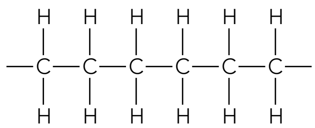
Results: the bromine is decolourised or the colour of the bromine water goes from orange to colourless.

g The other product would have to contain only one carbon and be an alkene, which has to have at least two carbons.

Exercise 20.2

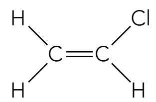
6 **a** Poly(ethene) is an **addition** polymer formed from many **ethene** molecules. In this reaction, the starting molecules can be described as **monomers**. The process is known as **polymerisation**.

b



7 a–v / b–ii / c–iv / d–i / e–iii

8



9 **a** not broken down by natural means / not digested by biological organisms (bacteria / fungi)

b calcium chloride, water, carbon dioxide

Chapter 20 continued

10 a aquatic life (e.g. fish, turtles, birds) can get tangled with plastic bags / ropes, etc. and be strangled by them / larger creatures can take bags, etc. into their stomachs / fragments of larger plastic objects and microbeads can be ingested and interfere with feeding and digestion of food

b any exfoliating / mildly abrasive products such as cosmetic facial scrubs, shower gels and toothpastes

Exercise 20.3

11

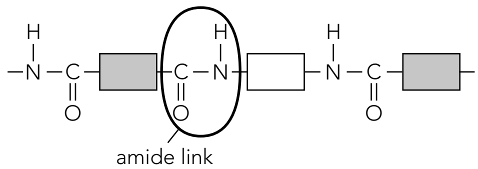
	Addition polymerisation	Condensation polymerisation
Monomers used	usually many molecules of a single **unsaturated** monomer, usually contains a carbon–carbon **double** bond	molecules of **two** different monomers usually used; monomers contain a reactive **functional** group at each end of the molecule
Reaction taking place	an **addition** reaction	a **condensation** reaction with loss of a small molecule (usually **water**) each time a monomer joins the chain
Nature of product	only **one** product formed – the polymer	two products: the **polymer** plus another, small, molecule
	non-biodegradable	can be biodegradable
	resistant to acids	PET can be **hydrolysed** back to monomers by acids or alkalis

12 a A, C and D.

b They have only one functional group in the molecule.

13 a polyamide

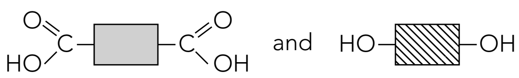
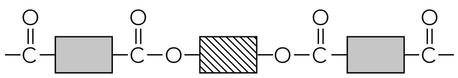
b



c water / H_2O

d nylon

14

Monomers used	
Structure of the polymer formed (show just three monomers joined)	
Other product formed	Name: **water** Formula: H_2O
Type of polymer formed	~~polyamide~~ (**polyester**) ~~polysaccharide~~

Chapter 20 continued

15 **a** They are both condensation polymers / both polyamides.

b Synthetic polyamides are made from a maximum of two monomers. A protein chain is made from up to 20 different (amino acid) monomers with each protein having a unique amino acid sequence.

c amino acids

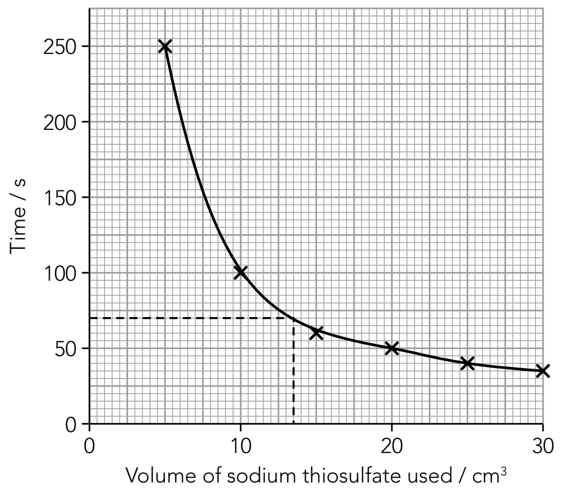
Chapter 21

Exercise 21.1

1 Measuring cylinder, because the large volume of solution needs to be added rapidly; volumetric pipettes are designed to only dispense small volumes at a time, but are more accurate.

2 Results would be more precise but not more accurate, as accuracy depends on human reaction time.

3



4 **a** 13.5 cm^3 (13–14 cm^3 depending on the line drawn on the graph)

b 16.5 cm^3 (or 30 minus the value given in **4a**)

5 $\frac{20}{40} \times 0.5 = 0.25\,mol/dm^3$

6 **a** Additional apparatus needed: thermometer and Bunsen burner

- Use fixed volume of thiosulfate.
- Measure temperature of solution.
- Add acid and time.
- Retake temperature.
- Repeat, but heat thiosulfate solution to raise temperature by 5 °C.
- Further repeats increasing temperature by 5 °C each time.
- Plot a graph of time against final temperature of mixture.
- To make a fair test, change only one variable (temperature) at a time while keeping all other conditions (solution concentration, etc.) the same.

b Toxic sulfur dioxide gas will be released from solution if the mixture is heated too strongly (the solubility of gases decreases with increased temperature).

Exercise 21.2

7 **a**

Elements	Mixtures	Compounds
aluminium	brass	carbon dioxide
sodium	dilute nitric acid	copper sulfate
zinc	seawater	methane

b Elements. Mixtures are easily separated by physical means, and compounds can be separated chemically into their constituent elements.

c electrolysis

8 **a** simple distillation

b Any insoluble solid in a liquid, with appropriate filtrate and residue from example given.

c Dissolve in a solvent (water), filter and heat to crystallise.

9 Heat the solution to evaporate water (crystallisation point) until crystals are seen at the surface or until a sample taken with a glass rod forms crystals. Allow to cool and filter the crystals formed from the remaining solution.

10 Simple distillation: the liquid is evaporated by boiling and condensed leaving the solid behind. Example: production of drinking water from seawater. Fractional distillation: more than one liquid is present / the mixture is boiled and the condensate is collected at different temperatures / the collecting vessel is changed when the boiling point of a component is exceeded. Example: separating different fractions from petroleum.

11 **a** Impure water would have a higher boiling point.

b An impure metal would have a lower melting point.

12 Weigh the sample of powdered brass / add hydrochloric acid in excess / wait until reaction stops (no more bubbling) / filter to obtain copper / wash and dry copper / weigh the copper and calculate the mass of zinc by subtraction.

Chapter 21 continued

Exercise 21.3

13 A: solvent front: the level that the solvent reaches during chromatography run

B: origin / base line: the level where the samples are placed

14 between 0 cm and 1 cm / must be below the origin, e.g. 0.5 cm

15 A pencil was used rather than an ink pen because soluble components in the ink could have risen up the paper with the sample and interfered with the result.

16 No. There are no spots in common / no spots at the same height.

17 The component at 5 cm, as it has moved furthest.

18 Mixture X: one component was insoluble and remained on the baseline (as shown).

Mixture Y: two of the components moved together in the solvent used (or one component was colourless and so did not show on the chromatogram).

19 **a** By comparing the spots with those produced by known substances (standards).

b By running the sample using a different solvent to try to separate them (or by using a locating agents to detect any colourless component).

20 **a** spot moves 2 – 1 cm = 1 cm, solvent moves 7 – 1 cm = 6 cm; R_f = 1/6 = 0.17 (2 d.p.)

b By using a locating agent which reacts with the substance in the spot making it coloured / spray the locating agent on the chromatogram / hang to allow the reaction to take place and to dry.

Chapter 22

Exercise 22.1

1

Cation (metal ion)	Colour of flame
lithium, Li^+	**red**
sodium, Na^+	**yellow**
potassium, K^+	**lilac**
calcium, Ca^{2+}	**orange–red**
barium, Ba^{2+}	**light green**
copper(II), Cu^{2+}	**blue–green**

2 The procedure for the flame test is:

- Dip a clean (nichrome) wire probe into hydrochloric acid.
- Dip the wire into the compound to be tested.
- Hold the wire near the edge of the roaring blue flame of a Bunsen burner.
- Note the colour of the flame produced.

Results: the antacids containing sodium carbonate will give a yellow flame; this colour tends to mask the colour produced by the presence of other metals / the antacid containing just calcium carbonate gives an orange red flame.

3 Solid carbonates would not react directly with NaOH(aq) / first acidify each solid antacid with a small amount of dilute hydrochloric acid to form the metal chlorides / then add excess aqueous sodium hydroxide separately to each sample. The antacid containing calcium ions will give a white precipitate (of calcium hydroxide).

4 **a** Use a flame test: barium gives a light green flame.

b Barium carbonate is insoluble in water but would dissolve in the acid environment of the stomach and cause poisoning.

5 **a** Ca^{2+} and CO_3^{2-}

b The fizzing with HCl shows a gas is formed, and this is CO_2 which is given by carbonates.

c calcium carbonate

d All carbonates (apart from those of Group I and ammonium) are insoluble. Therefore calcium carbonate is insoluble and after filtration will always be in the residue.

6 **a** strong pungent smell of ammonia and the red litmus is turned blue

b nitric acid

c white precipitate

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

e ammonium chloride, or NH_4Cl

f $NH_4Cl + NaOH \rightarrow NaCl + H_2O + NH_3$

Exercise 22.2

7 **a** chromium Cr^{3+}

b copper Cu^{2+}

c iron(II) Fe^{2+}

8 **a** When Fe^{2+} is exposed to air it will form Fe^{3+} by oxidation.

b iron(III) Fe^{3+}

c copper Cu^{2+} / blue–green flame

9 Drops of copper sulfate are added to the ammonia solution, so the ammonia solution is always in excess and the pale blue precipitate never appears.

Chapter 22 continued

10 a

Test	Observations
1 A sample of the solid mixture was dissolved in distilled water.The solution was acidified with dilute HCl(aq) and a solution of $BaCl_2$ added.	A white precipitate was formed.
2 A sample of the solid was placed in a test tube. NaOH(aq) was added and the mixture warmed. A piece of moist red litmus paper was held at the mouth of the tube.	The solid dissolved and pungent fumes were given off. The litmus paper turned **blue**, indicating the presence of **ammonium** ions.
3 A sample of the solid was dissolved in distilled water to give a **colourless** solution. NaOH(aq) was added dropwise until in excess.	A **white** precipitate was formed which was **soluble** in excess alkali.
4 A further sample of the solid was dissolved in distilled water. Concentrated ammonia solution (NH_3(aq)) was added dropwise until in excess.	A **white** precipitate was formed. On addition of excess alkali, the precipitate was **soluble**.

b ammonium sulfate ($(NH_4)_2SO_4$) and zinc sulfate ($ZnSO_4$)

c The precipitate in both tests is zinc hydroxide ($Zn(OH)_2$).

Exercise 22.3

11 a carbon dioxide

b Test: bubble the gas through colourless limewater.
Result: the limewater turns cloudy/milky.

c i Test: Use a glowing splint
Observations: splint relights

ii Test: Use a lighted splint
Observations: ‘pops’ with a lighted splint

12 a Ammonia: turns damp red litmus paper blue.

Chlorine: bleaches damp litmus paper white.

b It would not show which was which / both gases would give identical results.

13 a A paper dipped in a solution of acidified potassium manganate(VII) is held in the gas. The paper turns from purple to colourless.

b It is a reducing agent.

c damages buildings / damaging to tree growth / damaging to aquatic life in lakes and streams / causes breathing problems

Chapter 22 continued

Exercise 22.4

14 burette: to measure the volume of the solution added
volumetric pipette: to measure the fixed volume of the solution in the flask
conical flask: to contain and swirl the solutions being reacted

15 thymolphthalein, it is an indicator to show the endpoint

16 a

	Titration			
	1	2	3	4
Final burette reading / cm^3	**25.9**	**48.6**	**32.4**	**28.5**
First burette reading / cm^3	**0.0**	**23.3**	**6.9**	**3.1**
Volume of solution A / cm^3	**25.9**	**25.3**	**25.5**	**25.4**
Best titration result (✓)		✓	✓	✓

b 25.9 cm^3 because it is the rough or pilot value

c 25.4 cm^3 (omitting rejected value)

d to ensure adequate mixing

e Any remaining solution in the flask could interfere with the next titration and give false results.

17 a The water does not affect the result as the quantity of acid and alkali are separately measured. Water is only the solvent.

b A smaller volume would be needed to neutralise the same volume of alkali since it contains more acid per unit volume (a higher concentration).

Exercise 22.5

18 The cobalt chloride paper would change from blue to pale pink (test for the presence of water).

19 a white precipitate (sulfate ions)

20 water D (lowest pH value)

21 water C (most sodium ions and most chloride ions)

22 Water D. Impurities increase the boiling point of water. D is purest (lowest total solids) and so has the lowest boiling point.

23 Carbonates can decompose on heating / some solids splash out during boiling.

24 To decide whether the water might pose any hazards to health / for quality control.