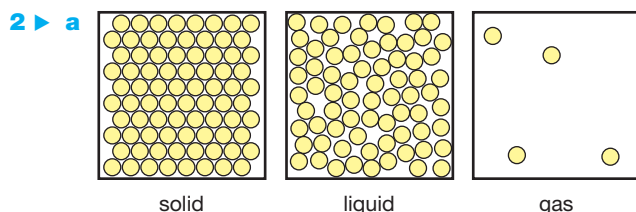


# UNIT 1 ANSWERS

## CHAPTER 1

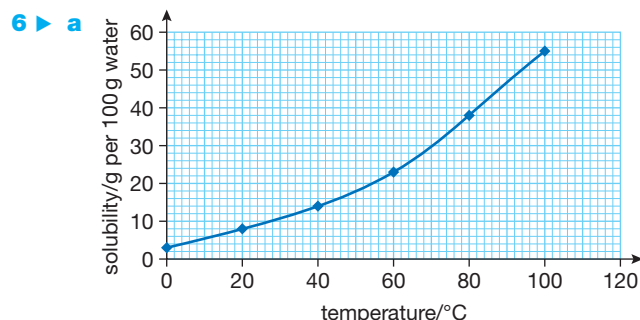
- 1 ► a melting b freezing  
c subliming/sublimation d subliming/sublimation



*Note: Solids should have regularly packed particles touching. Liquids should have most of the particles touching at least some of their neighbours, but with gaps here and there, and no regularity. Gases should have the particles well spaced.*

- b Solids: vibration around a fixed point. Liquids: particles can move around into vacant spaces, but with some difficulty because of the relatively close packing.
- c Evaporation: Some faster moving particles break away from the surface of the liquid. Boiling: Attractive forces are broken throughout the liquid to produce bubbles of vapour.
- 3 ► a i A – gas; B – liquid; C – solid; D – liquid; E – solid  
ii A – gas; B – solid; C – solid; D – liquid; E – solid  
iii A – gas; B – liquid; C – solid; D – gas; E – solid
- b A, because it is a gas.
- c It sublimes and therefore is converted directly from a solid to a gas without going through the liquid stage.
- d D – because it has a lower boiling point the forces of attraction between particles will be weaker therefore it will also evaporate more easily than substance B (the only other substance that is a liquid at 25 °C).
- 4 ► a The ammonia and hydrogen chloride particles have to diffuse through the air in the tube, colliding with air particles all the way.
- b i Its particles will move faster.  
ii It would take slightly longer for the white ring to form, because the gas particles would be moving more slowly at the lower temperature.
- c Ammonia particles are lighter than hydrogen chloride particles and so move faster. The ammonia covers more distance than the hydrogen chloride in the same time.
- d i Ammonium bromide.  
ii The heavier hydrogen bromide particles would move more slowly than the hydrogen chloride particles, and so the ring would form even closer to the hydrobromic acid end than it was to the hydrochloric acid end. The ring will also take slightly longer to form because of the slower moving particles.

- 5 ► Sodium chloride dissolves in water to form a **solution**. The water is called the **solvent** and the sodium chloride is the **solute**. If the solution is heated to 50 °C some of the water **evaporates** until the solution becomes **saturated** and sodium chloride crystals start to form.



- b 94 ± 1 g per 100 g  
The values obtained in this question and in c depend on the line of best fit. In the exam there will always be some tolerance – a range of values will be accepted.
- c From the graph, the solubility at 30 °C is 10 g per 100 g of water.  
 $\frac{40}{100} \times 10 = 4$  g  
Therefore 4 g of sodium chlorate will dissolve.
- d i 53 ± 1 °C  
ii The solubility at 17 °C is 7 ± 1 g per 100 g, therefore 20 – 7 = 13 g must precipitate out of the solution. Answers of 13 ± 1 g are acceptable.

## CHAPTER 2

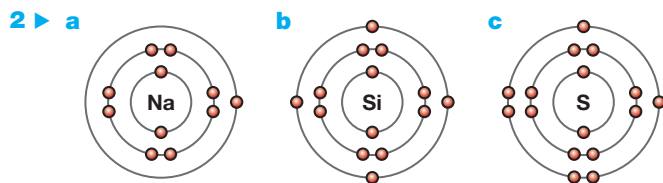
- 1 ►
- | Element  | Compound           | Mixture                   |
|----------|--------------------|---------------------------|
| hydrogen | magnesium oxide    | sea water                 |
| calcium  | copper(II) sulfate | honey                     |
|          |                    | blood                     |
|          |                    | mud                       |
|          |                    | potassium iodide solution |
- 2 ► a mixture b mixture c element  
d element e compound f compound
- 3 ► Substance X is the pure substance – it melts at a fixed temperature. Substance Y is impure – it melts over a range of temperatures.
- 4 ► a crystallisation b (simple) distillation  
c fractional distillation d chromatography  
e filtration
- 5 ► For example: Stir with a large enough volume of cold water to dissolve all the sugar. Filter to leave the diamonds on the filter paper. Wash on the filter paper with more water to remove any last traces of sugar solution. Allow to dry.
- 6 ► a M b R  
c 0.45 ± 0.01 (measure to the centre of the spot and remember to measure from the base line and not from the bottom of the paper)  
d G and T e P

## CHAPTER 3

- 1 ▶ **a** the nucleus **b** electrons  
**c** proton **d** proton and neutron
- 2 ▶ **a** 9  
**b** sum of protons + neutrons in the nucleus  
**c** 9p, 10n, 9e  
**d** The protons and electrons have equal but opposite charges. The atom has no overall charge, therefore there must be equal numbers of protons and electrons.
- 3 ▶ **a** 26p, 30n, 26e  
**b** 41p, 52n, 41e  
**c** 92p, 143n, 92e
- 4 ▶ **a** Atoms with the same atomic number but different mass numbers. They have the same number of protons, but different numbers of neutrons.  
**b**  $^{35}\text{Cl}$ : 17p, 18n, 17e;  $^{37}\text{Cl}$ : 17p, 20n, 17e
- 5 ▶  $\frac{6 \times 7 + 93 \times 7}{100} = 6.93$
- 6 ▶  $\frac{24 \times 78.99 + 25 \times 10.00 + 26 \times 11.01}{100} = 24.32$
- 7 ▶  $\frac{204 \times 1.4 + 206 \times 24.1 + 207 \times 22.1 + 208 \times 52.4}{100} = 207.241$
- 8 ▶ **a** 77 protons, 114 neutrons, 77 electrons  
**b** Iridium-193 has 2 more neutrons in the nucleus.  
**c** More iridium-193 because the relative atomic mass is closer to 193 than 191.
- 9 ▶ This statement is true – it only applies to 1 element, hydrogen ( $^1\text{H}$ ).

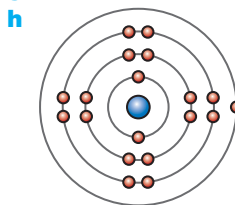
## CHAPTER 4

- 1 ▶ **a** **i** strontium **ii** chlorine **iii** nitrogen  
**iv** caesium **v** neon  
**b** metals: caesium, molybdenum, nickel, strontium, tin  
 non-metals: chlorine, neon, nitrogen



- 3 ▶ **a** 2, 7 **b** 2, 8, 3 **c** 2, 8, 8, 2
- 4 ▶ **a** 5 **b** 7 **c** 4 **d** 8
- 5 ▶ **a** A, F **b** A  
**c** C and D because 5 shells are occupied  
**d** C because there are 7 electrons in the outer shell (energy level)  
**e** B, D  
**f** Calcium – it has 20 electrons and therefore must have 20 protons in the nucleus. The atomic number is therefore 20. Calcium is the element with atomic number 20.

g 82, lead



- 6 ▶ Palladium is a metal and so is likely to have any of the following properties:
- good conductor of electricity
  - forms a basic oxide
  - is shiny when polished or freshly cut
  - is malleable
  - is ductile
  - is a good conductor of heat
- The first two are mentioned specifically on the syllabus.*
- 7 ▶ They have a full outer shell (energy level) and so they have no tendency to form compounds by losing/gaining electrons or sharing electrons.
- 8 ▶ Argon and potassium OR iodine and tellurium.  
 The elements would then be in a different group in the Periodic Table. They would not have the same number of electrons in the outer shell as other members of the group and would react in a completely different way. For example, potassium would be in Group 0 with the noble gases, and argon, which is very unreactive, would be in Group 1, with the highly reactive alkali metals.

## CHAPTER 5

- 1 ▶ **a**  $\text{Fe} + 2\text{HCl} \rightarrow \text{FeCl}_2 + \text{H}_2$   
**b**  $\text{Zn} + \text{H}_2\text{SO}_4 \rightarrow \text{ZnSO}_4 + \text{H}_2$   
**c**  $\text{Ca} + 2\text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2 + \text{H}_2$   
**d**  $2\text{Al} + \text{Cr}_2\text{O}_3 \rightarrow \text{Al}_2\text{O}_3 + 2\text{Cr}$   
**e**  $\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2$   
**f**  $2\text{NaHCO}_3 + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + 2\text{CO}_2 + 2\text{H}_2\text{O}$   
**g**  $2\text{C}_8\text{H}_{18} + 25\text{O}_2 \rightarrow 16\text{CO}_2 + 18\text{H}_2\text{O}$   
**h**  $\text{Fe}_3\text{O}_4 + 4\text{H}_2 \rightarrow 3\text{Fe} + 4\text{H}_2\text{O}$   
**i**  $\text{Pb} + 2\text{AgNO}_3 \rightarrow \text{Pb(NO}_3)_2 + 2\text{Ag}$   
**j**  $2\text{AgNO}_3 + \text{MgCl}_2 \rightarrow \text{Mg(NO}_3)_2 + 2\text{AgCl}$   
**k**  $\text{C}_3\text{H}_8 + 5\text{O}_2 \rightarrow 3\text{CO}_2 + 4\text{H}_2\text{O}$
- 2 ▶ **a** 44 **b** 60 **c** 142 **d** 132  
**e** 286; (The common mistake would be not to multiply the whole water molecule by 10. So the mass of the  $10\text{H}_2\text{O}$  is 180. Students will commonly and wrongly come up with 36 for this by multiplying the  $\text{H}_2$  by 10 but not the O as well. Work out the mass of the whole  $\text{H}_2\text{O}$  first and then multiply it by the number in front. That way you won't make this mistake.)  
**f** 392
- 3 ▶ **a** 13.9% **b** 35%  
**c** 21.2%; (Be careful of the cases where there are two nitrogen atoms in the fertiliser (all except  $\text{KNO}_3$ ). The masses of the nitrogen in those cases will be 28 and not 14.)

- 4 ► In each case, work out the  $M_r$  by adding up the relative atomic masses ( $A_r$  values), and then attach the unit “g” to give the mass of 1 mole.

- a 27 g  
 b 331 g  
 c  $4.30 \times 16 = 68.8$  g  
 d  $0.70 \times 62 = 43.4$  g  
 e  $0.015 \times 85 = 1.275$  g  
 f  $0.24 \times 286 = 68.64$  g *Don't forget the water of crystallisation*

Strictly speaking the answers to d), e) and f) shouldn't be quoted to more than 2 significant figures, because the number of moles is only quoted to that precision.)

- 5 ► In each case, work out the mass of 1 mole as above, and then work out how many moles you've got in the stated mass. You can use the equation:

$$\text{number of moles} = \frac{\text{mass}}{\text{mass of 1 mole}}$$

- a  $\frac{20}{40} = 0.5$  mol  
 b  $\frac{3.20}{160} = 0.0200$  mol  
 c  $\frac{2000}{79.5} = 25.2$  mol; *Don't forget to convert kg to g!*  
 d  $\frac{50}{249.5} = 0.2$  mol  
 e  $\frac{1\,000\,000}{56} = 17\,900$  mol (or 17 857, although this is precise to more significant figures than the  $A_r$ ).  
 f  $\frac{0.032}{64} = 5.0 \times 10^{-4}$  mol (0.0005 mol)

- 6 ► a  $4 \times 58.5 = 234$  g

- b  $\frac{37}{74} = 0.5$  mol  
 c  $\frac{1000}{40} = 25$  mol  
 d  $0.125 \times 79.5 = 9.94$  g (9.9375 g)  
 e  $\frac{4}{0.1} = 40$  g  
 f  $\frac{1}{0.004} = 250$  g

- 7 ► a

	P	H
Combining mass	9.39 g	0.61 g
No. of moles of atoms	$\frac{9.39}{31} = 0.30$	$\frac{0.61}{1} = 0.61$
Ratio of moles	1	2

Empirical formula is  $\text{PH}_2$

- b

	K	N	O
Combining mass	5.85 g	2.10 g	4.80 g
No. of moles of atoms	$\frac{5.85}{39} = 0.15$	$\frac{2.10}{14} = 0.15$	$\frac{4.80}{16} = 0.3$
Ratio of moles	1	1	2

Empirical formula is  $\text{KNO}_2$

- c

	Na	S	O
Combining mass	3.22 g	4.48 g	3.36 g
No. of moles of atoms	$\frac{3.22}{23} = 0.14$	$\frac{4.48}{32} = 0.14$	$\frac{3.36}{16} = 0.21$
Ratio of moles (divide by smallest number)	1 simplifies to 2	1	1.5
		2	3

Empirical formula is  $\text{Na}_2\text{S}_2\text{O}_3$

- d

	Carbon	Hydrogen	Bromine
Given %	22.0	4.6	73.4
No. of moles of atoms	$\frac{22.0}{12} = 1.833$	$\frac{4.6}{1} = 4.6$	$\frac{73.4}{80} = 0.9175$
Ratio of moles (divide by smallest number)	2	5	1

Empirical formula is  $\text{C}_2\text{H}_5\text{Br}$

- 8 ► a The mass of oxygen is  $2.84 - 1.24 = 1.60$  g

	P	O
Combining mass	1.24 g	1.60 g
No. of moles of atoms	$\frac{1.24}{31} = 0.04$	$\frac{1.60}{16} = 0.10$
Ratio of moles	1	2.5

Whole numbers must be used, therefore multiply by 2 to get the empirical formula  $\text{P}_2\text{O}_5$ .

- b  $\text{P}_2\text{O}_5$  has an  $M_r$  of 142

$$\frac{284}{142} = 2$$

Therefore there must be two lots of the empirical formula in a molecule.

Molecular formula is  $\text{P}_4\text{O}_{10}$

- 9 ► a

	Carbon	Hydrogen	Oxygen
Given %	66.7	11.1	22.2
Combining mass in 100 g	66.7 g	11.1 g	22.2 g
No. of moles of atoms	$\frac{66.7}{12} = 5.558$	$\frac{11.1}{1} = 11.1$	$\frac{22.2}{16} = 1.3875$
Ratio of moles (divide by smallest number)	4	8	1

Empirical formula is  $\text{C}_4\text{H}_8\text{O}$

- b** The mass of the empirical formula is  $4 \times 12 + 8 \times 1 + 16 = 72$ . Since this is equal to the relative formula mass, the molecular formula is the same as the empirical formula, that is  $C_4H_8O$ .

- 10 ►** You know the mass of anhydrous sodium sulfate (1.42 g). You can work out the mass of water of crystallisation ( $3.22 - 1.42 \text{ g} = 1.8 \text{ g}$ ).

You can work out the mass of 1 mole of sodium sulfate,  $\text{Na}_2\text{SO}_4 = 142 \text{ g}$ ; and the mass of 1 mole of water = 18 g.

$$\text{Number of moles of sodium sulfate} = \frac{1.42}{142} = 0.01 \text{ mol.}$$

$$\text{Number of moles of water} = \frac{1.8}{18} = 0.1 \text{ mol.}$$

So for every 1 mole of sodium sulfate, there are 10 moles of water.

The value of  $n$  is 10.

- 11 ►** mass of anhydrous calcium sulfate =  $44.14 - 37.34 = 6.80 \text{ g}$   
mass of water of crystallisation =  $45.94 - 44.14 = 1.80 \text{ g}$   
mass of 1 mole of calcium sulfate,  $\text{CaSO}_4 = 136 \text{ g}$

$$\text{number of moles of sodium sulfate} = \frac{6.80}{136} = 0.05 \text{ mol}$$

mass of 1 mole of water = 18 g

$$\text{number of moles of water} = \frac{1.8}{18} = 0.1 \text{ mol}$$

$$\text{number of moles of water} = 0.1 = 2$$

$$\text{number of moles of calcium sulfate} = 0.05$$

the value of  $n = 2$

- 12 ► a** 0.36 mol

- b** From the chemical equation, the number 2 in front of the HCl indicates that 2 mol HCl react with 1 mol  $\text{CaCO}_3$ , therefore 0.4 mol  $\text{CaCO}_3$  react with  $2 \times 0.4 = 0.8 \text{ mol HCl}$ .

- c** 6 mol HCl react to form 3 mol  $\text{H}_2\text{S}$

Therefore the number of moles of  $\text{H}_2\text{S}$  is half the number of moles of HCl.

$$0.4 \text{ mol HCl form } 0.2 \text{ mol H}_2\text{S}$$

- d** 3 mol CO form 2 mol Fe

The number of moles of Fe is  $\frac{2}{3}$  the number of moles of CO.

$$0.9 \times \frac{2}{3} = 0.6 \text{ mol iron}$$

- e**  $0.8 \times \frac{3}{2} = 1.2 \text{ mol hydrogen}$

- 13 ► a** number of moles of iron =  $\frac{10}{56} = 0.179 \text{ mol}$

- b** From the chemical equation, the number of moles of bromine that reacted with this number of moles of iron is  $\frac{3}{2} \times 0.179 = 0.277 \text{ mol}$ .

- c** From the chemical equation, the number of moles of  $\text{FeBr}_3$  will be the same as the number of moles of iron = 0.179 mol.

- d** mass of 1 mol  $\text{FeBr}_3 = 56 + 3 \times 80 = 296$

$$\text{mass of FeBr}_3 = 0.179 \times 296 = 53 \text{ g}$$

(A common mistake here is to multiply the number of moles of  $\text{FeBr}_3$  by the mass of  $2\text{FeBr}_3$ . The 2 has

already been used when you worked out that 2 mol Fe formed 2 mol  $\text{FeBr}_3$  – do not use it again. The equation for working out the mass is

mass = number of moles  $\times$  mass of 1 mole).

- 14 ► a**  $\frac{1.0}{190} = 0.0053 \text{ mol}$

- b** 0.0053 mol

- c**  $0.0053 \times 48 = 0.25 \text{ g}$

- d** number of moles of NaCl =  $0.0053 \times 4 = 0.0212 \text{ mol}$

$$\text{mass of NaCl} = 0.0212 \times 58.5 = 1.2 \text{ g}$$

- e** You can carry out a moles calculation as above:

$$1 \text{ tonne} = 1\,000\,000 \text{ g}$$

$$\text{moles of TiCl}_4 = \frac{1\,000\,000}{190} = 5300 \text{ mol}$$

$$\text{moles of Ti} = 5300 \text{ mol}$$

$$\text{mass of Ti} = 5300 \times 48 = 254\,400 \text{ g}$$

Alternatively, you can reason that, if 1 g  $\text{TiCl}_4$  forms 0.25 g Ti, 1 tonne  $\text{TiCl}_4$  will form 0.25 tonne Ti.

- 15 ►** mass of 1 mole of  $\text{AlCl}_3 = 27 + 3 \times 35.5 = 133.5 \text{ g}$

$$\text{moles of Aluminum chloride} = \frac{2.67}{133.5} = 0.0200 \text{ mol}$$

$$\text{moles of AgCl} = 3 \times 0.0200 = 0.0600 \text{ mol}$$

$$\text{mass of 1 mole of AgCl} = 108 + 35.5 = 143.5 \text{ g}$$

$$\text{mass of AgCl} = 0.0600 \times 143.5 = 8.61 \text{ g}$$

- 16 ► a** mass of 1 mol  $\text{Cr}_2\text{O}_3 = 2 \times 52 + 3 \times 16 = 152 \text{ g}$

$$\text{number of moles of Cr}_2\text{O}_3 = \frac{50}{152} = 0.33 \text{ mol}$$

$$\text{number of moles of Al} = 2 \times 0.33 = 0.66 \text{ mol}$$

$$\text{mass of Al} = 0.66 \times 27 = 17.8 \text{ g}$$

- b** number of moles of Cr =  $2 \times 0.33 = 0.66 \text{ mol}$

$$\text{mass of Cr} = 0.66 \times 52 = 34.3 \text{ g}$$

- c** 5 kg is 5000 g

$$\text{number of moles of Cr}_2\text{O}_3 = \frac{5000}{152} = 33 \text{ mol}$$

$$\text{number of moles of Cr} = 2 \times 33 = 66 \text{ mol}$$

$$\text{mass of Cr} = 66 \times 52 = 3430 \text{ g or } 3.43 \text{ kg}$$

Alternatively, we can reason that 5 kg is 100 times as much as 50 g. If we start with 100 times as much  $\text{Cr}_2\text{O}_3$ , we will make 100 times as much Cr.

- d** 5 tonnes is 5 000 000 g

$$\text{number of moles of Cr}_2\text{O}_3 = \frac{5\,000\,000}{152} = 33\,000 \text{ mol}$$

$$\text{number of moles of Cr} = 2 \times 33\,000 = 66\,000 \text{ mol}$$

$$\text{mass of Cr} = 66\,000 \times 52 = 3\,430\,000 \text{ g or } 3.43 \text{ tonnes}$$

Alternatively, we can reason that, if 5 kg of  $\text{Cr}_2\text{O}_3$  produces 3.43 kg Cr, then 5 tonnes of  $\text{Cr}_2\text{O}_3$  will produce 3.43 tonnes of Cr.

- 17 ► a** mass of 1 mol  $\text{CuO} = 63.5 + 16 = 79.5 \text{ g}$

$$\text{number of moles of CuO} = \frac{4.00}{79.5} = 0.0503 \text{ mol}$$

$$\text{number of moles of CuSO}_4 = 0.0503 \text{ mol}$$

$$\text{number of moles of CuSO}_4 \cdot 5\text{H}_2\text{O} = 0.0503 \text{ mol}$$

$$\text{mass of 1 mol CuSO}_4 \cdot 5\text{H}_2\text{O} = 63.5 + 32 + 4 \times 16 + 5 \times (16 + 2) = 249.5 \text{ g}$$

$$\text{mass of CuSO}_4 \cdot 5\text{H}_2\text{O} = 0.0503 \times 249.5 = 12.55 \text{ g}$$

- b** the theoretical yield is 12.55 g  
the actual yield is 11.25 g  
 $\frac{11.25}{12.55} \times 100 = 89.64\%$
- 18 ► a** mass of 1 mol ethanol =  $2 \times 12 + 6 \times 1 + 16 = 46$  g  
number of moles of ethanol =  $\frac{20.0}{46} = 0.435$  mol  
number of moles of ethyl ethanoate = 0.435 mol  
mass of 1 mol ethyl ethanoate =  $4 \times 12 + 2 \times 16 + 8 \times 1 = 88$  g  
mass of ethyl ethanoate =  $0.435 \times 88 = 38.3$  g  
**b** theoretical yield = 38.3 g  
actual yield = 30.0 g  
percentage yield =  $\frac{30.0}{38.3} \times 100 = 78.3\%$
- 19 ► a** 0.5 mol HCl would react with 0.25 mol  $\text{Na}_2\text{CO}_3$ .  
There is more than 0.25 mol  $\text{Na}_2\text{CO}_3$ , therefore  $\text{Na}_2\text{CO}_3$  is in excess.  
**b** 0.02 mol  $\text{O}_2$  would react with  $\frac{0.02}{5} = 0.004$  mol  $\text{C}_3\text{H}_8$ .  
There is more than 0.004 mol  $\text{C}_3\text{H}_8$ , therefore  $\text{C}_3\text{H}_8$  is in excess.  
**c** 28 g of CO is  $\frac{28}{28} = 1$  mol  
1 mol CO would react with  $\frac{1}{3} = 0.33$  mol  $\text{Fe}_2\text{O}_3$ .  
There is more than 0.33 mol  $\text{Fe}_2\text{O}_3$ , therefore  $\text{Fe}_2\text{O}_3$  is in excess.  
**d** 16 g  $\text{O}_2$  is  $\frac{16}{32} = 0.5$  mol  
16 g  $\text{SO}_2$  is  $\frac{16}{64} = 0.25$  mol  
0.25 mol  $\text{SO}_2$  would react with 0.125 mol  $\text{O}_2$   
There is more than 0.125 mol  $\text{O}_2$ , therefore  $\text{O}_2$  is in excess.
- 20 ► a** 1.0 g of  $\text{CaCO}_3$  is  $\frac{1.0}{100} = 0.010$  mol  
0.010 mol  $\text{CaCO}_3$  would react with  $2 \times 0.010 = 0.020$  mol HCl  
There is less than 0.020 mol HCl, therefore there is not enough HCl to react with all the  $\text{CaCO}_3$ . Therefore  $\text{CaCO}_3$  is in excess.  
**b** To calculate the number of moles of  $\text{CO}_2$  you must use the number of moles of HCl because not all the  $\text{CaCO}_3$  reacted.  
moles of HCl = 0.015 mol  
moles of  $\text{CO}_2 = 0.5 \times 0.015 = 0.0075$  mol  
mass of  $\text{CO}_2 = 0.0075 \times 44 = 0.33$  g  
(If you got the answer 0.44 g you used the number of moles of  $\text{CaCO}_3$ .  $\text{CaCO}_3$  was in excess; therefore not all of it will react.)
- 2 ► a**  $2.0 \times 24 = 48 \text{ dm}^3$  (48 000  $\text{cm}^3$ )  
**b**  $0.10 \times 24 = 2.4 \text{ dm}^3$  (2400  $\text{cm}^3$ )  
**c**  $1 \times 10^{-3} \times 24 = 0.024 \text{ dm}^3$  (24  $\text{cm}^3$ )
- 3 ► a** 200  $\text{cm}^3$  of chlorine is  $\frac{200}{24\,000} = 0.00833$  mol  
(be careful with units here – if the volume is in  $\text{cm}^3$  you must use 24 000 as the molar volume)  
mass of 1 mol of  $\text{Cl}_2 = 2 \times 35.5 = 71$  g  
mass of 0.00833 mol  $\text{Cl}_2 = 0.00833 \times 71 = 0.592$  g  
**b** mass of 1 mol of  $\text{O}_2 = 2 \times 16 = 32$  g  
number of moles of  $\text{O}_2 = 0. \frac{16}{32} = 0.005$  mol  
0.005 mol  $\text{O}_2$  has a volume of  $0.005 \times 24\,000 = 120 \text{ cm}^3$  (0.12  $\text{dm}^3$ )  
**c** 1  $\text{dm}^3$  of the gas is  $\frac{1}{24} = 0.0417$  mol  
the mass of 0.0417 mol is 1.42 g  
mass of 1 mol of the gas =  $\frac{1.42}{0.0417} = 34.1$  g
- 4 ►**  $\frac{0.240}{24} = 0.0100$  mol  
From the chemical equation, the number of moles of  $\text{H}_2$  is the same as the number of moles of Mg: 0.0100 mol.  
volume of hydrogen =  $0.0100 \times 24 = 0.24 \text{ dm}^3$  (240  $\text{cm}^3$ )
- 5 ►**  $\frac{1}{24} = 0.0417$  mol  $\text{O}_2$   
From the chemical equation, the number of moles of  $\text{KNO}_3$  is twice the number of moles of  $\text{O}_2$   
 $2 \times 0.0417 = 0.0833$  mol  
mass of 1 mol  $\text{KNO}_3 = 39 + 14 + 3 \times 16 = 101$  g  
mass of 0.0417 mol  $\text{KNO}_3 = 0.0833 \times 101 = 8.42$  g
- 6 ►** mass of 1 mol  $\text{MnO}_2 = 55 + 2 \times 16 = 87$  g  
number of moles of  $\text{MnO}_2 = \frac{2.00}{87} = 0.0230$  mol  
from the chemical equation: 1 mol  $\text{MnO}_2$  produces 1 mol  $\text{Cl}_2$   
number of moles of  $\text{Cl}_2$  produced = 0.0230 mol  
volume of  $\text{Cl}_2$  produced =  $0.0230 \times 24\,000 = 552 \text{ cm}^3$  (0.552  $\text{dm}^3$ )
- 7 ►** mass of 1 mol  $\text{H}_2\text{SO}_4 = 98$  g  
number of moles of  $\text{H}_2\text{SO}_4 = \frac{4.90}{98} \text{ mol} = 0.0500$  mol  
this number of moles is in 1  $\text{dm}^3$  of solution  
concentration = 0.0500 mol/ $\text{dm}^3$  (0.0500 to show that the answer is precise to 3 significant figures)
- 8 ►** mass of 1 mol KOH =  $39 + 16 + 1 = 56$  g  
mass of 0.200 mol =  $0.200 \times 56 = 11.2$  g  
concentration = 11.2 g/ $\text{dm}^3$
- 9 ►** In each of these questions the number of moles is given by:  
number of moles = volume in  $\text{dm}^3 \times$  concentration in mol/ $\text{dm}^3$   
to convert a volume in  $\text{dm}^3$  to  $\text{cm}^3$  divide by 1000  
**a**  $\frac{25.0}{1000} \times 0.100 = 0.00250$  mol  
**b**  $\frac{200}{1000} \times 0.200 = 0.0400$  mol

## CHAPTER 6

- 1 ► a**  $\frac{2.4}{24} = 0.10$  mol  
**b**  $\frac{480}{24} = 20$  mol  
**c**  $\frac{100}{24\,000} = 0.00417$  mol  
**d**  $\frac{1500}{24\,000} = 0.0625$  mol

- 9 ►** In each of these questions the number of moles is given by:  
number of moles = volume in  $\text{dm}^3 \times$  concentration in mol/ $\text{dm}^3$   
to convert a volume in  $\text{dm}^3$  to  $\text{cm}^3$  divide by 1000  
**a**  $\frac{25.0}{1000} \times 0.100 = 0.00250$  mol  
**b**  $\frac{200}{1000} \times 0.200 = 0.0400$  mol



$$\text{c } \frac{75.0}{1000} \times 0.150 = 0.01125 \text{ mol}$$

$$\text{d } \frac{22.4}{1000} \times 0.280 = 0.00627 \text{ mol}$$

$$10 \blacktriangleright \text{concentration (mol/dm}^3\text{)} = \frac{\text{number of moles (mol)}}{\text{volume (dm}^3\text{)}}$$

$$\text{a } \text{concentration} = \frac{0.100}{2} = 0.0500 \text{ mol/dm}^3$$

$$\text{b } \text{volume in dm}^3 = \frac{25.0}{1000} = 0.025 \text{ dm}^3$$

$$\text{concentration} = \frac{0.0200}{0.0250} = 0.800 \text{ mol/dm}^3$$

$$\text{c } \text{volume in dm}^3 = \frac{27.8}{1000} = 0.0278 \text{ dm}^3$$

$$\text{concentration} = \frac{0.00150}{0.0278} = 0.0540 \text{ mol/dm}^3$$

$$11 \blacktriangleright \text{volume (dm}^3\text{)} = \frac{\text{number of moles (mol)}}{\text{concentration (mol/dm}^3\text{)}}$$

$$\text{a } \frac{0.500}{0.100} = 5.00 \text{ dm}^3 \text{ (5.00 indicates that the answer is to 3 significant figures)}$$

$$\text{b } \frac{0.00500}{0.0200} = 0.250 \text{ dm}^3 \text{ (250 cm}^3\text{)}$$

$$\text{c } \frac{0.0200}{0.500} = 0.0400 \text{ dm}^3 \text{ (40.0 cm}^3\text{)}$$

$$12 \blacktriangleright \text{no. of moles of copper(II) sulfate} = \frac{20}{1000} \times 0.100 = 0.00200 \text{ mol}$$

equation shows that 1 mol  $\text{CuSO}_4$  produces 1 mol  $\text{BaSO}_4$

no. of moles  $\text{BaSO}_4$  formed = 0.00200 mol

mass of 1 mol  $\text{BaSO}_4$  =  $137 + 32 + 4 \times 16 = 233 \text{ g}$

mass of 0.00200 mol  $\text{BaSO}_4$  =  $0.00200 \times 233 = 0.466 \text{ g}$

$$13 \blacktriangleright 25.0 \text{ cm}^3 \text{ of } 2.00 \text{ mol/dm}^3 \text{ HCl contains } \frac{25.0}{1000} \times 2.00 \text{ mol} = 0.0500 \text{ mol}$$

From the chemical equation, the number of moles of  $\text{CaCO}_3$  that reacts is half the number of moles of HCl, that is  $0.5 \times 0.0500 = 0.0250 \text{ mol}$ .

mass of 1 mol  $\text{CaCO}_3$  = 100 g

mass of 0.0250 mol  $\text{CaCO}_3$  =  $0.0250 \times 100 = 2.50 \text{ g}$

number of moles of  $\text{CO}_2$  = 0.0250 mol

volume of  $\text{CO}_2$  =  $0.0250 \times 24\,000 = 600 \text{ cm}^3$

$$14 \blacktriangleright \text{number of moles of hydrogen peroxide} = \frac{30.0}{1000} \times 0.0200 = 0.000600 \text{ mol (} 6.00 \times 10^{-4}\text{)}$$

from the chemical equation, the number of moles of  $\text{O}_2$  produced is half the number of moles of  $\text{H}_2\text{O}_2$

number of moles of  $\text{O}_2$  =  $0.000300 \text{ mol (} 3.00 \times 10^{-4}\text{)}$

volume of oxygen =  $0.000300 \times 24\,000 = 7.20 \text{ cm}^3$

$$15 \blacktriangleright \text{a } \text{no. of moles of NaOH solution} = \frac{25.0}{1000} \times 0.400 \text{ mol} = 0.0100 \text{ mol}$$

the equation shows that you need half as many moles of sulfuric acid = 0.00500 mol

$$\text{volume (dm}^3\text{)} = \frac{\text{number of moles (mol)}}{\text{concentration (mol/dm}^3\text{)}}$$

$$\text{volume} = \frac{0.00500}{0.200} = 0.0250 \text{ dm}^3 \text{ or } 25.0 \text{ cm}^3$$

$$\text{b } \text{mass of 1 mol } \text{CaCO}_3 = 100 \text{ g}$$

$$\text{number of moles of } \text{CaCO}_3 = \frac{10.0}{100} = 0.100 \text{ mol}$$

from the chemical equation 0.100 mol  $\text{CaCO}_3$  reacts with  $2 \times 0.100 = 0.200 \text{ mol HCl}$

$$\text{volume (dm}^3\text{)} = \frac{\text{number of moles (mol)}}{\text{concentration (mol/dm}^3\text{)}}$$

$$\text{volume} = \frac{0.200}{2.00} = 0.100 \text{ dm}^3 \text{ or } 100 \text{ cm}^3$$

16  $\blacktriangleright$  In each of these questions the number of moles is given by  
number of moles = volume in  $\text{dm}^3 \times$  concentration in  $\text{mol/dm}^3$   
to convert a volume in  $\text{dm}^3$  to  $\text{cm}^3$  divide by 1000

$$\text{a } \text{no. of moles of NaOH solution} = \frac{25.0}{1000} \times 0.100 \text{ mol} = 0.00250 \text{ mol}$$

the equation shows that you need the same number of moles of nitric acid = 0.00250 mol

$$\text{concentration (mol/dm}^3\text{)} = \frac{\text{number of moles (mol)}}{\text{volume (dm}^3\text{)}}$$

$$20.0 \text{ cm}^3 \text{ is } \frac{20.0}{1000} = 0.0200 \text{ dm}^3$$

$$\text{concentration} = \frac{0.00250}{0.0200} = 0.125 \text{ mol/dm}^3$$

$$\text{b } \text{no. of moles of nitric acid} = \frac{30.0}{1000} \times 0.100 \text{ mol} = 0.00300 \text{ mol}$$

the equation shows that you need half the number of moles of sodium carbonate =  $0.5 \times 0.00300 = 0.00150 \text{ mol}$

$$\text{concentration (mol/dm}^3\text{)} = \frac{\text{number of moles (mol)}}{\text{volume (dm}^3\text{)}}$$

$$25.0 \text{ cm}^3 \text{ is } \frac{25.0}{1000} = 0.0250 \text{ dm}^3$$

$$\text{concentration} = \frac{0.00150}{0.0250} = 0.0600 \text{ mol/dm}^3$$

$$\text{c } \text{no. of moles of potassium carbonate solution} = \frac{25.0}{1000} \times 0.250 \text{ mol} = 0.00625 \text{ mol}$$

the equation shows that you need twice the number of moles of ethanoic acid =  $2 \times 0.00625 = 0.0125 \text{ mol}$

$$\text{concentration (mol/dm}^3\text{)} = \frac{\text{number of moles (mol)}}{\text{volume (dm}^3\text{)}}$$

$$12.5 \text{ cm}^3 \text{ is } \frac{12.5}{1000} = 0.0125 \text{ dm}^3$$

$$\text{concentration} = \frac{0.0125}{0.0125} = 1.00 \text{ mol/dm}^3$$

$$17 \blacktriangleright \text{a } \text{no. of moles of hydrochloric acid} = \frac{18.8}{1000} \times 0.0400 \text{ mol} = 0.000752 \text{ mol (} 7.52 \times 10^{-4}\text{)}$$

the equation shows that you need half the number of moles of calcium hydroxide

number of moles of calcium hydroxide =  $0.5 \times 0.000752 = 0.000376 \text{ mol (} 3.76 \times 10^{-4}\text{)}$

$$\text{concentration (mol/dm}^3\text{)} = \frac{\text{number of moles (mol)}}{\text{volume (dm}^3\text{)}}$$

$$25.0 \text{ cm}^3 \text{ is } \frac{25.0}{1000} = 0.0250 \text{ dm}^3$$

$$\text{concentration} = \frac{0.000376}{0.0250} = 0.0150 \text{ mol/dm}^3$$

**b** mass of 1 mol  $\text{Ca}(\text{OH})_2 = 40 + 2 \times (16+1) = 74 \text{ g}$

there are 0.0150 mol in  $1 \text{ dm}^3$

the mass of 0.0150 mol is  $0.0150 \times 74 = 1.11 \text{ g}$

therefore the concentration is  $1.11 \text{ g/dm}^3$

**18 ▶ a** 0.1 mol  $\text{HNO}_3$  reacts with 0.05 mol  $\text{Na}_2\text{CO}_3$ . There is more than this present, therefore  $\text{Na}_2\text{CO}_3$  is in excess.

**b** no. of moles of  $\text{Na}_2\text{CO}_3 = \frac{20.0}{1000} \times 0.100 \text{ mol} = 0.00200 \text{ mol}$

0.0200 mol  $\text{HNO}_3$  reacts with 0.0100 mol  $\text{Na}_2\text{CO}_3$ . There is less than this present, therefore  $\text{HNO}_3$  is in excess.

**c** no. of moles of  $\text{Na}_2\text{CO}_3 = \frac{25.0}{1000} \times 0.300 \text{ mol} = 0.00750 \text{ mol}$

no of moles of  $\text{HNO}_3 = \frac{20.0}{1000} \times 0.400 \text{ mol} = 0.00800 \text{ mol}$

0.00800 mol  $\text{HNO}_3$  reacts with 0.00400 mol  $\text{Na}_2\text{CO}_3$ . There is more than this present, therefore  $\text{Na}_2\text{CO}_3$  is in excess.

## CHAPTER 7

**1 ▶ a i** An atom or group of atoms which carries an electrical charge.

**ii** Attractions between positively and negatively charged ions holding them together.

**b** Correct electronic structures for:

**i** Na 2,8,1 and Cl 2,8,7

**ii** Li 2,1 and O 2,6

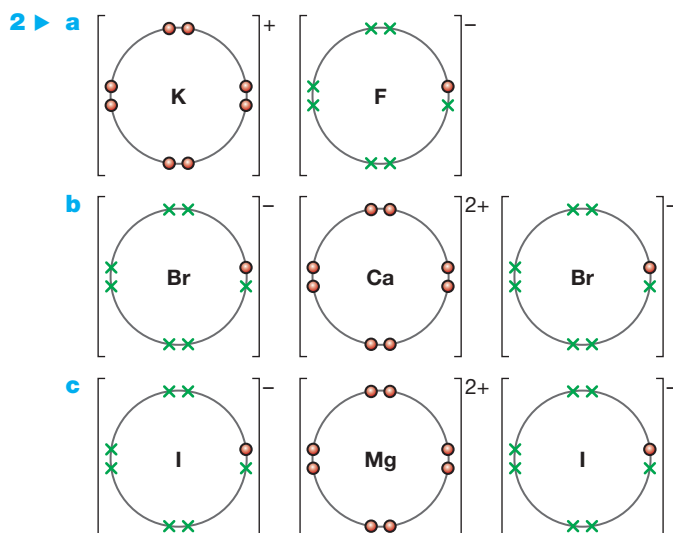
**iii** Mg 2,8,2 and F 2,7.

Diagrams (similar to those in the chapter) showing transfer of electrons, and the charges and electronic structures of the ions formed (or words to the same effect).

In (i), show 1 electron transferred from Na to Cl leaving  $\text{Na}^+ [2,8]^+$  and  $\text{Cl}^- [2,8,8]$

In (ii), show 2 lithium atoms each giving 1 electron to O leaving  $2 \times \text{Li}^+ (2)^+$  and  $\text{O}^{2-} [2,8]^{2-}$

In (iii), show 1 Mg giving an electron each to 2 fluorines leaving  $\text{Mg}^{2+} [2,8]^{2+}$  and  $2 \times \text{F}^- [2,8]^-$



**3 ▶**

	<b>a formula</b>	<b>b name</b>
<b>i</b>	magnesium	$\text{Mg}^{2+}$
<b>ii</b>	strontium	$\text{Sr}^{2+}$
<b>iii</b>	potassium	$\text{K}^+$
<b>iv</b>	oxygen	$\text{O}^{2-}$
<b>v</b>	sulfur	$\text{S}^{2-}$
<b>vi</b>	caesium	$\text{Cs}^+$
<b>vii</b>	chlorine	$\text{Cl}^-$
<b>viii</b>	iodine	$\text{I}^-$
<b>ix</b>	aluminium	$\text{Al}^{3+}$
<b>x</b>	calcium	$\text{Ca}^{2+}$
<b>xi</b>	nitrogen	$\text{N}^{3-}$
		nitride

**4 ▶** PbO NaBr

$\text{MgSO}_4$   $\text{ZnCl}_2$

$\text{K}_2\text{CO}_3$   $(\text{NH}_4)_2\text{S}$

$\text{Ca}(\text{NO}_3)_2$   $\text{Fe}(\text{OH})_3$

$\text{FeSO}_4$   $\text{CuCO}_3$

$\text{Al}_2(\text{SO}_4)_3$   $\text{Ca}(\text{OH})_2$

$\text{CoCl}_2$   $\text{CaO}$

$\text{AgNO}_3$   $\text{FeF}_3$

$\text{NH}_4\text{NO}_3$   $\text{RbI}$

$\text{Na}_2\text{SO}_4$   $\text{Cr}_2\text{O}_3$

**5 ▶ a** The electrostatic forces of attraction between oppositely charged ions are strong and require a lot of energy to break.

**b** The ions are held tightly in place in the giant lattice structure and are not free to move.

**c** The ions are free to move (it is important to use the word **ions** here; any mention of electrons will score 0 in an exam).

**6 ▶** Potassium chloride will have a lower melting point than calcium oxide. The charges on the ions in KCl ( $\text{K}^+$  and  $\text{Cl}^-$ ) are lower than in CaO ( $\text{Ca}^{2+}$  and  $\text{O}^{2-}$ ). There are weaker electrostatic forces of attraction between oppositely charged ions in KCl than in CaO; these forces require less energy to break than the forces in CaO.

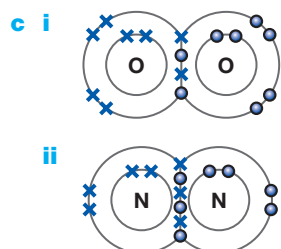
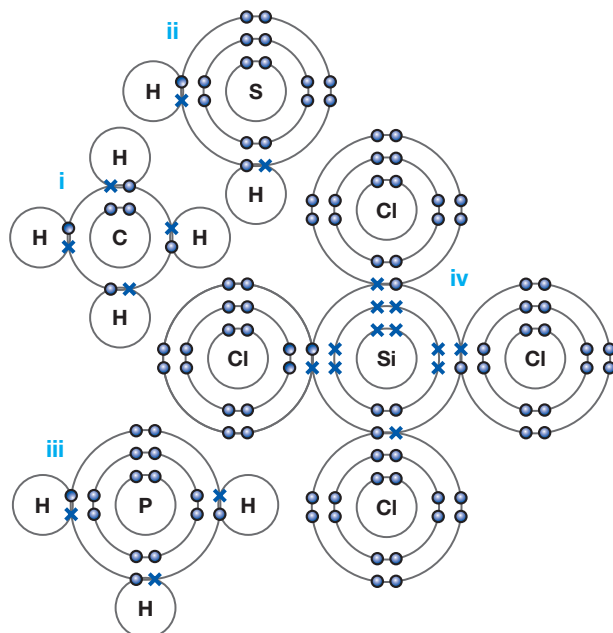
## CHAPTER 8

**1 ▶**

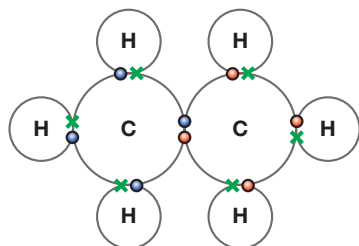
<b>a</b>	$\text{MgO}$	ionic
<b>b</b>	$\text{CH}_3\text{Br}$	covalent
<b>c</b>	$\text{H}_2\text{O}_2$	covalent
<b>d</b>	$\text{FeCl}_2$	ionic
<b>e</b>	$\text{NaF}$	ionic
<b>f</b>	$\text{HCN}$	covalent

**2 ▶ a** A pair of electrons that is shared between two atoms. The atoms are held together because the nucleus of each is attracted to the shared pair.

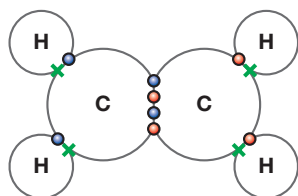
- b** (It doesn't matter whether students use dots or crosses or just different colours- or what positions (N, S, E, W) the hydrogens occupy in the  $H_2S$  or  $PH_3$ .)



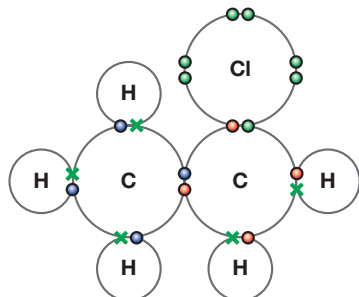
**3 ▶ a**



**b**

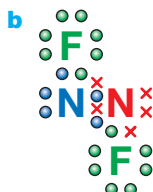


**c**



(It doesn't matter what variations of colours or dots and crosses are used. In the chloroethane case, the Cl could equally well have been drawn together with either of the other two hydrogen atoms on the right-hand carbon atom.)

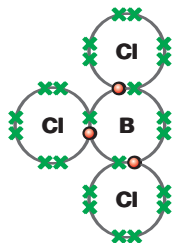
- 4 ▶** Carbon dioxide has a simple molecular structure; diamond has a giant covalent structure. When carbon dioxide sublimates, only the weak intermolecular forces of attraction must be broken – but when diamond sublimates, the strong covalent bonds must be broken. A lot more energy is required to break the strong covalent bonds in diamond than the weak intermolecular forces in carbon dioxide.
- 5 ▶ a** Simple molecular because it is a liquid at room temperature. Only weak intermolecular forces of attraction must be broken to melt solid hexane. Compounds with giant structures have high melting points and boiling points and will be solids at room temperature.
- b** Pentane has a lower boiling point. The intermolecular forces of attraction are weaker in pentane because the relative formula mass is lower. Weaker intermolecular forces require less energy to break.
- c** It will not conduct electricity because there are no ions present and all the electrons are held tightly in atoms or covalent bonds.
- 6 ▶ a** To break apart diamond, strong covalent bonds must be broken, which requires a large amount of energy. Much less energy is required to break the weak forces of attraction between the layers in graphite.
- b**  $C_{60}$  fullerene has a molecular structure but graphite has a giant structure. To melt  $C_{60}$  fullerene, only weak intermolecular forces must be broken, but to melt graphite strong covalent bonds must be broken. Much less energy is required to break the weak intermolecular forces in  $C_{60}$  fullerene than the strong covalent bonds in graphite.
- c** Each C atom in graphite only forms 3 bonds so there is one electron left over on each on each atom. These delocalised electrons are free to move throughout the layers.
- d** All the outer shell electrons in diamond are held tightly in covalent bonds and unable to move around.
- 7 ▶** (Weakest intermolecular forces of attraction) hydrogen, phosphorus trifluoride, ammonia, ethanol, water, ethanamide (strongest intermolecular forces of attraction). Higher intermolecular attractions produce higher boiling points – more energy has to be supplied to overcome stronger forces of attraction between molecules.
- 8 ▶ a** Nitrogen usually forms 3 bonds because it has 5 electrons in its outer shell. Each N forms 1 bond to F, therefore there must be a double bond between the two N atoms. Formation of a double bond results in each N having 8 electrons in its outer shell.





(It doesn't matter what variations of colours or dots and crosses are used. The F atoms could also be drawn in different positions.)

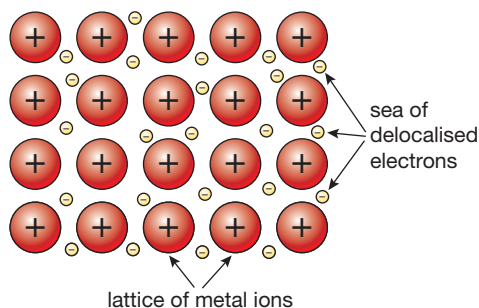
9 ► a



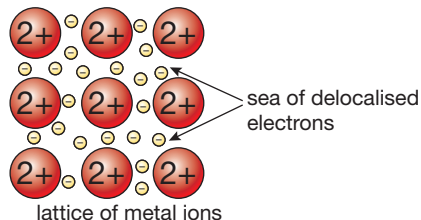
b The B atom does not have 8 electrons in its outer shell.

## CHAPTER 9

1 ► a



b



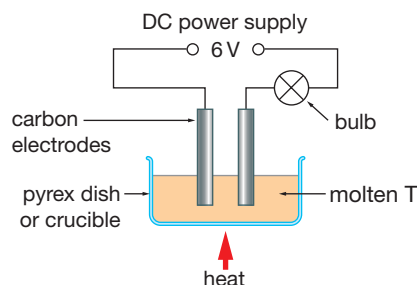
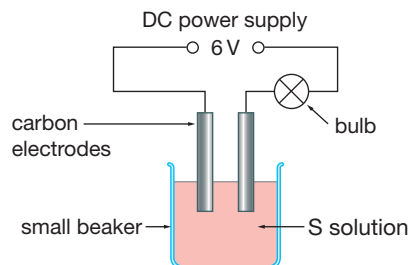
(The diagrams should show at least 9 ions but can be of any size beyond that. Electrons could be shown as  $e^-$ .)

- 2 ► a Mg 2, 8, 2. The two outer electrons will be lost from each Mg atom to form the sea of delocalised electrons. There will be a regular arrangement of  $Mg^{2+}$  ions.
- b Na will form  $Na^+$  ions, Mg will form  $Mg^{2+}$  ions and Al will form  $Al^{3+}$  ions. There will be stronger electrostatic attraction between the metal ions and the delocalised electrons when the charge on the ion is higher.  $Al^{3+}$  has the strongest attraction between the metal ions and the delocalised electrons, therefore most energy has to be supplied to overcome the forces of attraction.
- c Delocalised electrons are free to move.
- d The layers of metal ions are able to slide over each other without changing the bonding.
- 3 ► a giant covalent  
b molecular  
c molecular  
d giant ionic  
e giant metallic  
f molecular  
g giant metallic

## CHAPTER 10

- 1 ►
- |   | cathode  | anode    |
|---|----------|----------|
| a | lead     | bromine  |
| b | zinc     | chlorine |
| c | hydrogen | iodine   |
| d | sodium   | iodine   |
| e | copper   | chlorine |
| f | hydrogen | chlorine |
| g | hydrogen | oxygen   |
| h | hydrogen | oxygen   |
- 2 ► a  $Mg^{2+} + 2e^- \rightarrow Mg$   
b  $Al^{3+} + 3e^- \rightarrow Al$   
c  $2Br^- \rightarrow Br_2 + 2e^-$   
d  $2O_2^- \rightarrow O_2 + 4e^-$   
e  $2Cl^- \rightarrow Cl_2 + 2e^-$   
f  $Ni^{2+} + 2e^- \rightarrow Ni$   
g  $4OH^- \rightarrow O_2 + 2H_2O + 4e^-$   
h  $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$   
i  $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$
- 3 ► a Ions weren't free to move.  
b anode  
c iodine:  $2I(l) \rightarrow I_2(g) + 2e^-$   
d  $K(l) + e^- \rightarrow K(l)$   
e anode: bromine  
cathode: sodium  
f anode:  $2Br^- \rightarrow Br_2 + 2e^-$   
cathode:  $Na^+ + e^- \rightarrow Na$
- 4 ► a i  $Pb^{2+}(l) + 2e^- \rightarrow Pb(l)$   
ii  $2Br^-(l) \rightarrow Br_2(g) + 2e^-$   
iii oxidised: bromide ions. Reduced: lead(II) ions.  
b i  $2H^+(aq) + 2e^- \rightarrow H_2(g)$   
ii  $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$   
iii oxidised: chloride ions. Reduced: hydrogen ions.  
c i  $2H^+(aq) + 2e^- \rightarrow H_2(g)$   
ii  $2Br^-(aq) \rightarrow Br_2(aq \text{ or } l) + 2e^-$   
iii oxidised: bromide ions. Reduced: hydrogen ions.  
d i  $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$   
ii  $4OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e^-$   
iii oxidised: hydroxide ions. Reduced: copper(II) ions.  
e i  $2H^+(aq) + 2e^- \rightarrow H_2(g)$   
ii  $4OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e^-$   
iii oxidised: hydroxide ions. Reduced: hydrogen ions.  
f i  $Mg^{2+}(l) + 2e^- \rightarrow Mg(s \text{ or } l)$   
ii  $2I(l) \rightarrow I_2(g) + 2e^-$   
iii oxidised: iodide ions. Reduced: magnesium ions.  
g i  $2H^+(aq) + 2e^- \rightarrow H_2(g)$   
ii  $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$   
iii oxidised: chloride ions. Reduced: hydrogen ions.

- 5 ► The melting point of S is too high to reach using a Bunsen, and so you would have to test a solution in water. On the other hand, T would melt easily, and won't dissolve. Heat it until it melts.

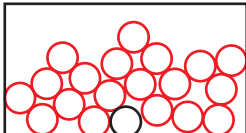


(You could also do the electrolysis of the solution using the more complicated apparatus in the chapter, but since there is no need to collect anything, there isn't much point.)

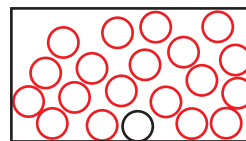
If the substances are electrolytes, the bulbs will light up, and there will be signs of activity around the electrodes (gases given off, solids deposited, etc).

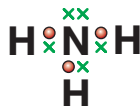
- 6 ► a  $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s})$
- b i The  $\text{Cu}^{2+}(\text{aq})$  ions are responsible for the blue colour and these are removed from the solution, therefore the blue colour fades.  
water dissociates  
 $\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$   
Oxygen is produced at the anode. The half equation for this is:  
 $4\text{OH}^- \rightarrow 2\text{H}_2\text{O} + \text{O}_2 + 4\text{e}^-$   
Because the  $\text{OH}^-$  is removed from solution, more water must dissociate to replace them. This produces an excess of  $\text{H}^+$  ions in solution, so the solution is acidic.
- ii There is no change in colour because the concentration of  $\text{Cu}^{2+}(\text{aq})$  remains the same; when 1  $\text{Cu}^{2+}(\text{aq})$  ion is removed at the cathode another one replaces it at the anode.  
There is no change in acidity because the  $\text{OH}^-$  ions are not removed from the solution to form oxygen.

## END OF UNIT 1 QUESTIONS

- 1 ► a  $-255^\circ\text{C}$  (1) The temperature is between the melting point and boiling point.
- b  Particles randomly arranged (1) and mostly touching each other. (1)

The diagram below would only score 1 mark – for the random arrangement



- c number of protons: 1 (1)  
number of neutrons: 2 (1)  
number of electrons: 1 (1)
- d i 2, 5 (1)
- ii   
three covalent bonds shown, each with a dot and a cross (1)  
the remaining 2 electrons on N (1)
- e  $\text{NH}_4\text{SO}_4$  is incorrect (1) because the ammonium ion is  $\text{NH}_4^+$  and the sulfate ion is  $\text{SO}_4^{2-}$
- f fractional distillation (1)
- 2 ► a The elements in the Periodic Table are arranged in order of **atomic number** (1). The vertical columns are called **groups** (1) and contain elements which have the same number of **electrons** (1) in their **outer shell** (1).
- b H (1) and Ar (1)
- c i The second property is not a property of metals.  
Metal oxides are basic (1)  
and would react with acids rather than with alkalis such as sodium hydroxide. (1)
- ii The grey circles represent silicon atoms. (1)  
A silicon atom has 4 electrons in its outer shell and will form 4 covalent bonds. (1)  
An oxygen atom has 6 electrons in its outer shell and will only form 2 covalent bonds. (1)  
(Note: although some of the 'grey atoms' only appear to form 2 bonds this is because only part of the giant structure is shown. None of the 'red atoms' form 4 bonds.)
- iii  $\text{SiO}_2$  would be a solid at room temperature because it has a giant structure. (1)  
A lot of energy is required to break all the strong covalent bonds. (1)
- 3 ► a idea of electron transfer (1)  
electron transfer from Sr to Br (1)  
the correct number of electrons transferred (1)
- b high melting point (1)  
strong electrostatic forces of attraction between oppositely-charged ions (1)  
require a lot of energy to break (1)
- c The relative atomic mass of an element is the weighted average mass of the isotopes of the element. (1)  
It is measured on a scale on which a carbon-12 atom has a mass of exactly 12. (1)
- d  $\frac{50.69 \times 79 + 49.31 \times 81}{100}$  (1)  
79.99 (1)
- e  $\text{Sr}(\text{NO}_3)_2$  (1)

- 4 ► a Diamond has a giant covalent structure (1), where all the atoms are held together by strong covalent bonds – a lot of energy is required to break these bonds. (1)
- b i Graphite has the higher melting point. (1)  
Graphite has a giant structure but C<sub>60</sub> fullerene has a simple molecular structure. (1)  
To melt graphite, all the strong covalent bonds between carbon atoms must be broken – this requires a lot of energy. (1)  
To melt C<sub>60</sub> fullerene, only the weak intermolecular forces of attraction must be broken – this requires less energy. (1)
- ii Graphite has delocalised electrons. (1)  
Delocalised electrons are free to move (throughout the structure). (1)
- iii C<sub>60</sub> has a molecular structure rather than a giant structure. (1)  
Electrons cannot move between molecules. (1)



The correct diagram is above. (There is a double bond between the O atoms.)

- b K<sub>2</sub>O is the correct formula (1)



The correct diagram is above. (Carbon forms 4 bonds and oxygen forms 2 bonds.)

- 6 ► a mass of lead = 24.16 – 17.95 = 6.21 g (1)
- b mass of oxygen = 24.80 – 24.16 g = 0.64 g (1)
- c
- |   | Pb       | O           |
|---|----------|-------------|
| combining masses                                      | 6.21 g   | 0.64 g      |
| no of moles of atoms                                  | 6.21/207 | 0.64/16 (1) |
|   | = 0.03   | = 0.04      |
| ratio of moles  | 3        | : 4 (1)     |
| empirical formula: Pb <sub>3</sub> O <sub>4</sub> (1) |          |             |

- 7 ► a number of moles of copper =  $\frac{2.00}{63.5}$  = 0.0315 mol (1)

From the chemical equation, the same number of moles of copper(II) nitrate will be produced.

number of moles of Cu(NO<sub>3</sub>)<sub>2</sub> is 0.0315 mol (1)

M<sub>r</sub> of Cu(NO<sub>3</sub>)<sub>2</sub> = 63.5 + 2 × (14 + 3 × 16) = 187.5

mass of Cu(NO<sub>3</sub>)<sub>2</sub> = 187.5 × 0.0315 = 5.91 g (1)

- b The student would use crystallisation. (1)

Heat the solution in an evaporating basin to drive off some of the water until the solution becomes saturated. (1)

Allow to cool then filter off the crystals that form. (1)

- c percentage yield =  $\frac{\text{actual yield}}{\text{theoretical yield}} \times 100$   
 $= \frac{5.23}{7.61} \times 100$  (1)  
 = 68.7% (1)

- 8 ► a M<sub>r</sub> of FeS<sub>2</sub> = 56 + 2 × 32 = 120  
480 kg is 480 000 g  
number of moles of FeS<sub>2</sub> =  $\frac{480\,000}{120}$  = 4000 mol (1)

From the chemical equation, the number of moles of Fe<sub>2</sub>O<sub>3</sub> is half the number of moles of FeS<sub>2</sub>.

number of moles of Fe<sub>2</sub>O<sub>3</sub> = 2000 mol (1)

Mr of Fe<sub>2</sub>O<sub>3</sub> = 2 × 56 + 3 × 16 = 160

mass of Fe<sub>2</sub>O<sub>3</sub> = 2000 × 160 = 320 000 g or 320 kg (1)

- b From the chemical equation 4000 mol FeS<sub>2</sub> produces 8000 mol SO<sub>2</sub>. (1)

volume = 8000 × 24 = 192 000 dm<sup>3</sup> (1)

- c i 2SO<sub>2</sub> + O<sub>2</sub> → 2SO<sub>3</sub>  
all formulae correct (1)  
correct balancing (1)

- ii From the equation 2 mol SO<sub>2</sub> react with 1 mol O<sub>2</sub>. So the number of moles of O<sub>2</sub> will be half as many. Therefore the number of moles of O<sub>2</sub> will be 4000 mol. (1)

volume = 4000 × 24 = 96 000 dm<sup>3</sup> (1)

For the second mark we could just reason that since half as many moles are required, then the volume will also be half.

- 9 ► a A saturated solution is a solution which contains as much dissolved solid as possible at a particular temperature. (1)

- b number of moles =  $\frac{32.8}{1000} \times 0.100$  = 3.28 × 10<sup>-3</sup> mol or 0.00328 mol (1)

- c From the chemical equation, the number of moles of Sr(OH)<sub>2</sub> is half the number of moles of HCl.

number of moles of Sr(OH)<sub>2</sub> = 1.64 × 10<sup>-3</sup> or 0.00164 mol (1)

volume of Sr(OH)<sub>2</sub> in dm<sup>3</sup> is  $\frac{25}{1000}$  = 0.025 dm<sup>3</sup> (1)

concentration of Sr(OH)<sub>2</sub> =  $\frac{0.00164}{0.025}$  = 0.0656 mol/dm<sup>3</sup> (1)

- d M<sub>r</sub> of Sr(OH)<sub>2</sub> = 88 + 2 × (16 + 1) = 122

mass of Sr(OH)<sub>2</sub> in 1 dm<sup>3</sup> of solution is 122 × 0.0656 = 8.00 g (1)

1 dm<sup>3</sup> is 1000 cm<sup>3</sup>, therefore mass of Sr(OH)<sub>2</sub> present in 100 cm<sup>3</sup> or 100 g of water is one tenth of this, that is 0.800 g (1)

- e (The three points in bold are essential and then 1 mark each for 2 further points.)

**an empty evaporating basin is weighed (1)**

some of the strontium hydroxide solution is poured into the evaporating basin (1)

care must be taken to not pour in any solid (1)

**weigh the evaporating basin and strontium hydroxide solution (1)**

the evaporating basin is heated over a tripod and gauze using a Bunsen burner (1)

until all the water is driven off / heat to constant mass (1)

**weigh the evaporating basin and solid strontium hydroxide (1)**

- 10 ► a Number of moles of sodium hydroxide =  $\frac{25.0}{1000} \times 0.100$  = 2.50 × 10<sup>-3</sup> mol or 0.0025 mol (1)

From the chemical equation, the number of moles of sulfuric acid is half of this.

Number of moles of sulfuric acid =  $1.25 \times 10^{-3}$  mol or 0.00125 mol (1)

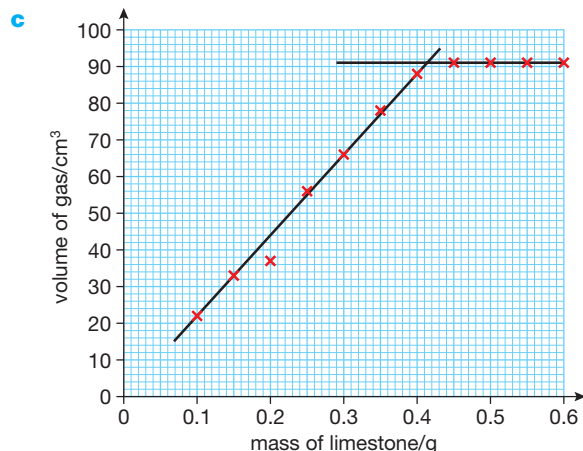
Volume of sulfuric acid in  $\text{dm}^3$  =  $0.0200 \text{ dm}^3$  (1)

Concentration of sulfuric acid is  $\frac{0.00125}{0.0200} = 0.0625 \text{ mol/dm}^3$  (1)

- b i** lattice of positive ions (1)  
sea of delocalised electrons (1)  
electrostatic attraction between the positive ions and delocalised electrons (1)
- ii** the layers of positive ions/atoms (1)  
can slide over each other (1)
- iii** number of moles =  $\frac{100}{1000} \times 0.0625 = 0.00625 \text{ mol}$   
or  $6.25 \times 10^{-3} \text{ mol}$  (1)
- iv**  $0.100 \text{ g of Mg is } \frac{0.100}{24} = 0.00417 \text{ mol}$  or  $4.17 \times 10^{-3} \text{ mol}$  (1)  
Therefore the sulfuric acid is in excess because only 0.00417 mol would be required to react with all the Mg and there is more present than this. (1)
- v** The number of moles of magnesium must be used to work out the number of moles of  $\text{H}_2$  because the sulfuric acid is in excess.  
From the chemical equation, the number of moles of  $\text{H}_2$  is the same as the number of moles of Mg.  
The number of moles of  $\text{H}_2$  is 0.00417 mol or  $4.17 \times 10^{-3} \text{ mol}$ . (1)  
volume of hydrogen =  $0.00417 \times 24000 = 100 \text{ cm}^3$  (1)
- vi** the percentage yield =  $\frac{94}{100} \times 100 = 94\%$  (1)
- c i** hydrogen gas (1)
- ii**  $4\text{OH}^- \rightarrow \text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^-$   
**OR**  $2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$   
for formulae for all species correct (1)  
for correctly balanced (1)

**11 ► a** measuring cylinder/pipette/burette (1)

**b** volume is  $33 \text{ cm}^3$  (1)



plotting points correctly (2)

line of best fit (1)

- d i** Student 3's result (1)  
The volume of gas recorded is too low and does not lie on the line of best fit. (1)

**ii** Some gas could have escaped from the apparatus / there was a leak in the apparatus / they could have used less than 0.15 g limestone / they used a different concentration of hydrochloric acid that had too low a concentration / they used too low a volume of hydrochloric acid. (1)

**e** All the hydrochloric acid has been used up. (1)  
There is no more hydrochloric acid to react with any additional limestone, so no more gas can be produced. (1)

**f** 0.42 g (1)  
(This is the point at which the two lines cross because this is the point at which adding any more  $\text{CaCO}_3$  does not produce any more gas.)

**g i**  $\text{CaCO}_3 (\text{s}) + 2\text{HCl} (\text{aq}) \rightarrow \text{CaCl}_2 (\text{aq}) + \text{H}_2\text{O} (\text{l}) + \text{CO}_2 (\text{g})$

**ii** Number of moles of  $\text{CO}_2 = \frac{91}{24000} = 3.79 \times 10^{-3} \text{ mol}$  or 0.00379 mol. (1)

From the chemical equation, the number of moles of  $\text{CaCO}_3$  is the same as the number of moles of  $\text{CO}_2$ .

Number of moles of  $\text{CaCO}_3$  is  $3.79 \times 10^{-3} \text{ mol}$  or 0.00379 mol. (1)

Mr of  $\text{CaCO}_3$  is  $40 + 12 + 3 \times 16 = 100$

mass of  $\text{CaCO}_3 = 0.00379 \times 100 = 0.379 \text{ g}$  (1)

**iii**  $\frac{0.379}{0.42} \times 100$  (1)  
= 90 % (1)

You have worked out above the mass of limestone that reacts with the hydrochloric acid is 0.42 g. The volume of  $\text{CO}_2$  produced by this mass of limestone is  $91 \text{ cm}^3$ .  $91 \text{ cm}^3$  of  $\text{CO}_2$  is produced by 0.379 g of  $\text{CaCO}_3$ , therefore this is the mass of  $\text{CaCO}_3$  present in 0.42 g of limestone.

## UNIT 2 ANSWERS

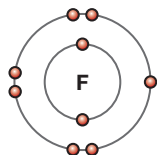
### CHAPTER 11

- 1 ►** They have the same number of electrons in the outer shell (1 electron) and all react in the same way.
- 2 ► a** A = lithium
- b** B = potassium; C = hydrogen; D = potassium hydroxide
- c**  $2\text{K}(\text{s}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{KOH}(\text{aq}) + \text{H}_2(\text{g})$
- d** The paper goes blue/purple.
- e** Lots of heat evolved. Melting point of potassium is low.
- f** E is sodium and F is sodium oxide  
sodium + oxygen → sodium oxide  
 $4\text{Na}(\text{s}) + \text{O}_2(\text{g}) \rightarrow 2\text{Na}_2\text{O}(\text{s})$
- 3 ► a** FALSE      **b** TRUE      **c** TRUE
- d** FALSE, the formula of lithium chloride is  $\text{LiCl}$
- 4 ► a** More dense. Density increases down the group.
- b** 1

- c Edexcelium hydroxide and hydrogen.
- d More reactive than francium. Reactivity increases down the group.  
Reactivity increases as the size of the atom increases.  
aAn edexcelium atom is larger than a francium atom.  
The outer electron would be more easily lost from edexcelium – it is less strongly attracted by the nucleus.
- e  $2\text{Ed} + 2\text{H}_2\text{O} \rightarrow 2\text{EdOH} + \text{H}_2$
- f Alkaline, because  $\text{OH}^-$  ions are formed.
- g  $\text{Ed}_2\text{O}$

## CHAPTER 12

1 ► a

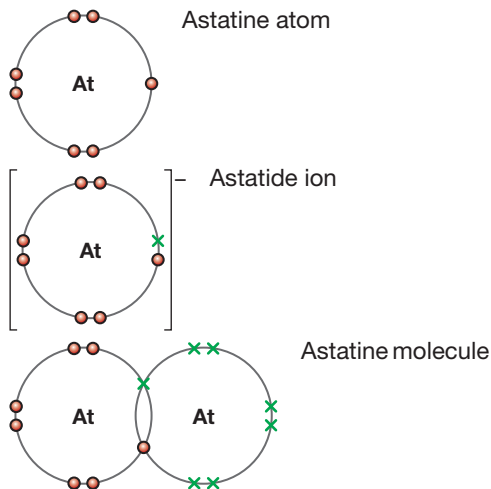


- b A fluorine atom has 7 electrons in its outer shell – the group number indicates the number of outer shell electrons.

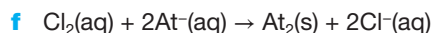


- d There are weaker intermolecular forces of attraction between fluorine molecules than between chlorine molecules. Weaker forces require less energy to break. The intermolecular forces are weaker because fluorine has a lower relative molecular mass.
- e No reaction because chlorine is less reactive than fluorine and therefore cannot displace it.

2 ► a

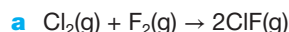


- b Astatine would be a solid – melting point increases down the group and iodine is a solid.
- c Reactivity decreases down the group so astatine would be less reactive than iodine.  
Astatine is a larger atom than iodine and so has a lower attraction for electrons.
- d pH 1–2
- e Caesium astatide would be a colourless/white solid because it is an ionic compound. It would probably be soluble in water because many ionic compounds are soluble in water and most of the alkali metal halides are soluble in water (e.g. sodium chloride, potassium iodide etc.).

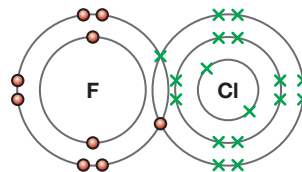


The chlorine is reduced in this reaction because it gains electrons. The  $\text{At}^-$  is oxidised in the reaction because it loses electrons. Oxidation and reduction both occur, therefore it is a redox reaction.

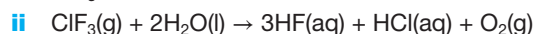
3 ► a



b



The boiling point increases as the relative molecular mass increases and the strength of the intermolecular forces increases.



## CHAPTER 13

1 ► a 78.1%    b 21.0%    c 0.04%    d 0.9%

2 ► a  $95 - 80 = 15 \text{ cm}^3$  of oxygen

$$\frac{15}{95} \times 100 = 15.8\%$$

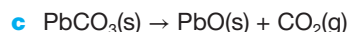
- b The answer is less than expected. The apparatus was not left long enough; there was not enough time for all the oxygen in the air to react. The experiment could be improved by leaving the apparatus set up for longer.

3 ► a

The student should follow a procedure such as:

- Weigh a boiling tube.
- Pour 1 g of the metal carbonate into the boiling tube and weigh the boiling tube again.
- Connect a delivery tube to the boiling tube. Connect the other end of the delivery tube to a gas syringe or put in to a measuring cylinder filled with water.
- Heat the boiling tube.
- Allow the apparatus to cool.
- Record the volume of gas collected.
- Re-weigh the boiling tube and contents.

b Carbon dioxide.



d Thermal decomposition.

e Sodium carbonate is more thermally stable than the other carbonates; it does not decompose when heated.

f The  $M_r$  of  $\text{PbCO}_3$  is 267, whereas the  $M_r$  of  $\text{CuCO}_3$  is 123.5.

1 g of  $\text{CuCO}_3$  is more moles than 1 g of  $\text{PbCO}_3$  – more than twice as many moles.

When more moles of solid are heated, more moles of gas will be produced, so a greater volume of gas will be produced.

g The mass goes down because  $\text{CO}_2$  is given off. In the first experiment the mass goes down by 0.15 g, which must be the mass of  $89 \text{ cm}^3$  of gas.



From the first experiment:

1 cm<sup>3</sup> of gas has a mass of  $\frac{0.16}{89} = 0.00180$  g

100 cm<sup>3</sup> has a mass of  $100 \times 0.00180 = 0.180$  g

- h** The student did not heat the carbonate for long enough so not all the carbonate decomposes.  
Some of the gas escaped/leaked from the apparatus.

- 4 ► a** Carbon dioxide is one of the gases responsible for the greenhouse effect. These gases absorb IR (infrared) radiation that has been emitted from the Earth's surface. They then re-radiate it in all directions, warming the atmosphere. Heating of the atmosphere could lead to climate change.

## CHAPTER 14

- 1 ► a** sodium, aluminum, iron, copper  
**b i** magnesium oxide, copper  
**ii**  $\text{Mg(s)} + \text{CuO(s)} \rightarrow \text{MgO(s)} + \text{Cu(s)}$  (Include state symbols in all equations for preference.)  
**iii** Copper(II) oxide (CuO) has been reduced to copper (Cu) because it has lost oxygen - reduction is the loss of oxygen.  
**iv** Copper(II) oxide (CuO) is the oxidizing agent because it oxidizes the magnesium (gives oxygen to it) and is, in the process, reduced.  
**c i** Zinc is higher in the reactivity series because it takes the oxygen from the cobalt(II) oxide  
**ii** A reducing agent is a substance which reduces something else. Zinc removes oxygen from the cobalt(II) oxide. Removal of oxygen is reduction.  
**iii** Zinc because it gains oxygen - oxidation is gain of oxygen.  
**d** aluminium, manganese, chromium (Statement 1: Al is above Cr. Statement 2: Mn is below Al. Statement 3: Mn is above Cr. Putting this together gives the final list.)
- 2 ► a** oxidised; gain of oxygen  
**b** reduced; loss of oxygen  
**c** oxidised; loss of electrons  
**d** reduced; gain of electrons
- 3 ►** Magnesium is above lead because it removes the oxygen from the lead(II) oxide.
- 4 ► a** Either: grey iron filings become coated with brown solid. Or: solution fades from blue to colourless (very pale green).  
**b** iron (Fe) has been oxidised (to Fe<sup>2+</sup>) because it has lost electrons - oxidation is loss of electrons.  
**c**  $\text{Fe(s)} + \text{CuSO}_4\text{(aq)} \rightarrow \text{FeSO}_4\text{(aq)} + \text{Cu(s)}$
- 5 ► a** nickel, copper, silver  
**b i** Either: colour of solution changes from blue to green. Or: nickel becomes coated with brown solid.  
**ii**  $\text{Ni(s)} + \text{CuSO}_4\text{(aq)} \rightarrow \text{NiSO}_4\text{(aq)} + \text{Cu(s)}$   
**iii**  $\text{Ni(s)} + \text{Cu}^{2+}\text{(aq)} \rightarrow \text{Ni}^{2+}\text{(aq)} + \text{Cu(s)}$   
 Nickel has been oxidised by loss of electrons.

- 6 ► a** X is between iron and hydrogen in the reactivity series. It displaces hydrogen from dilute hydrochloric acid, and copper from copper(II) sulfate, and so must be above hydrogen and copper. It won't displace iron from iron(II) sulfate, and so must be below iron.

- b i** yes: X nitrate and silver  
**ii** no reaction  
**iii** no reaction (at least, not in the short term)  
**iv** yes: X chloride and copper  
**v** yes: X sulfate and hydrogen

- 7 ► a** hydrogen  
**b** aluminium chloride  
**c**  $2\text{Al(s)} + 6\text{HCl(aq)} \rightarrow 2\text{AlCl}_3\text{(aq)} + 3\text{H}_2\text{(g)}$   
**d** Aluminium is covered by a very thin, but very strong, layer of aluminium oxide which prevents the acid getting at the aluminium underneath. On heating, the acid reacts with the oxide and removes it. The aluminium then shows its true reactivity, and produces a vigorous reaction.

- 8 ►** Drop a very small piece into cold water. If it reacts, judge its reactivity relative to K, Na, Ca or Mg.

If it doesn't react, add a small piece to dilute hydrochloric acid and warm if necessary. Rapid reaction in the cold would place it as 'similar to magnesium'. A few bubbles of hydrogen in the cold, but more on heating would place it as 'similar to iron or zinc'.

If there is no reaction, then it is 'below hydrogen'.

- 9 ► a** The iron/steel must be exposed to oxygen (air) and water.  
**b** Painting prevents the oxygen and water from coming into contact with the iron/steel.  
**c** It is iron/steel coated with zinc.  
**d** It would take much longer before the car went rusty. Zinc is more reactive than iron and so corrodes in preference to the iron. The zinc prevents iron rusting even when the coating is scratched. When paint is scratched the steel underneath will rust - this will not happen when the steel is galvanised, therefore the life of the car is extended.

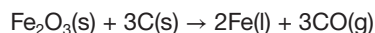
**10 ► a**

	Pearsonium chloride solution	Mollium chloride solution	Rosium chloride solution	Amelium chloride solution
pearsonium	X	X	X	X
mollium	✓	X	✓	✓
rosium	✓	X	X	X
amelium	✓	X	✓	X

- b** mollium > amelium > rosium > pearsonium  
**c** mollium + pearsonium chloride → mollium chloride + pearsonium

## CHAPTER 15

- 1 ►** Iron is below carbon in the reactivity series and so can be extracted by heating with carbon.

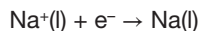


Aluminium is above carbon in the reactivity series and so heating aluminium oxide with carbon will not produce a reaction. Aluminium must be extracted by electrolysis.



- 2 ▶ a Sodium is high in the reactivity series, and so will need to be produced by electrolysis. It can be manufactured by electrolysis molten sodium chloride using suitable electrodes (industrially, carbon anode, steel cathode), keeping the sodium and chlorine apart to prevent them from reacting back to give sodium chloride again.

Sodium would be released at the cathode:



- b It needed the discovery of electricity.
- c Any other metals high in the reactivity series – e.g. potassium, lithium, calcium, magnesium.
- 3 ▶ Because lead is below iron in the reactivity series it is also below carbon. Lead can be extracted from its ore by heating with carbon.
- 4 ▶ a Chromium(III) oxide contains the  $\text{Cr}^{3+}$  ion. The oxide ion has the formula  $\text{O}^{2-}$ , therefore the formula of chromium(III) oxide is  $\text{Cr}_2\text{O}_3$ .  

$$\text{Cr}_2\text{O}_3 + 2\text{Al} \rightarrow 2\text{Cr} + \text{Al}_2\text{O}_3$$
- b Al is the reducing agent because it takes the O away from the  $\text{Cr}_2\text{O}_3$  – it reduces it.
- c Chromium is less reactive than aluminium because the aluminium is able to take the oxygen away from the chromium oxide.
- 5 ▶ a  $2\text{FeTiO}_3 + 7\text{Cl}_2 + 6\text{C} \rightarrow 2\text{TiCl}_4 + 2\text{FeCl}_3 + 6\text{CO}$
- b  $\text{TiCl}_4 + 2\text{Mg} \rightarrow \text{MgCl}_2 + \text{Ti}$
- c Magnesium is more reactive than titanium as it is able to take the chlorine away from the  $\text{TiCl}_4$ ; the magnesium is able to reduce the  $\text{TiCl}_4$ .
- d Magnesium is more reactive than carbon, so it could be extracted from its ore by electrolysis or by heating with a more reactive metal such as sodium.
- e Titanium is more expensive than magnesium. Although it is expensive to extract magnesium using electrolysis, the magnesium is used in the extraction of titanium, therefore titanium must be more expensive.

- 6 ▶ a i It has a low density. (Note: saying that it has a **low density** is usually better than just saying it is **light**)
- ii Alloys are stronger than the pure metals. Alloys contain some atoms of different sizes; this makes it more difficult for the layers of positive ions to slide over each other.
- iii Advantages:  
 An aluminium body is lighter than a steel body and so the car will use less fuel. An aluminium body will not corrode.  
 Disadvantage:  
 Aluminium is more expensive than steel.
- b i a good conductor of electricity and ductile (can be drawn out into wires)
- ii does not react with water and malleable (can be hammered into shape)

- c High-carbon steel is brittle and not very malleable so it would be difficult to make it into the correct shapes for car bodies. It would also not bend/deform very well during a crash – the fact that mild steel deforms upon impact (because it is malleable) is important for absorbing some of the force upon impact (from a Physics point of view, the change in momentum occurs over a longer period of time).

## CHAPTER 16

	methyl orange	phenolphthalein	litmus
a	yellow	pink	blue
b	red	colourless	red

2 ▶

solution	pH	strongly acidic	weakly acidic	neutral	weakly alkaline	strongly alkaline
potassium iodide	7			✓		
propanoic acid	4.2		✓			
sodium carbonate	9.5				✓	
potassium hydroxide	13					✓
iron(III) chloride	2.4	✓				
nitric acid	1.3	✓				

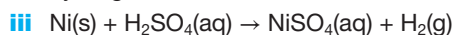
- 3 ▶ a  $\text{Na}_2\text{O} + 2\text{HNO}_3 \rightarrow 2\text{NaNO}_3 + \text{H}_2\text{O}$
- b  $2\text{KOH} + \text{H}_2\text{SO}_4 \rightarrow \text{K}_2\text{SO}_4 + 2\text{H}_2\text{O}$
- 4 ▶ Neutralisation – reaction a) involves an acid reacting with a base (sodium oxide) and reaction b) involves an acid reacting with an alkali (KOH).
- 5 ▶ Measure out  $25.0\text{cm}^3$  of potassium hydroxide solution using a **pipette**. Transfer the potassium hydroxide solution to a **conical flask**. Add a few drops of an **indicator**. Put the sulfuric acid into the **burette**. Add the acid to the alkali until the **indicator changes colour**.

## CHAPTER 17

- 1 ▶ a copper(II) oxide, copper(II) hydroxide, copper(II) carbonate
- b  $\text{CuO}(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{CuSO}_4(\text{aq}) + \text{H}_2\text{O}(\text{l})$   
 $\text{Cu}(\text{OH})_2(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{CuSO}_4(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$   
 $\text{CuCO}_3(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{CuSO}_4(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
- 2 ▶ a A = copper; B = copper(II) oxide; C = copper(II) sulfate; D = copper(II) carbonate; E = carbon dioxide; F = copper(II) nitrate
- b i  $\text{CuO}(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{CuSO}_4(\text{aq}) + \text{H}_2\text{O}(\text{l})$
- ii  $\text{CuCO}_3(\text{s}) + 2\text{HNO}_3(\text{aq}) \rightarrow \text{Cu}(\text{NO}_3)_2(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
- 3 ▶ a i Description could include: Silvery metal. Colourless acid. Slow bubbles of gas on warming. Formation of green solution. Possible disappearance of nickel if very small quantities were used.

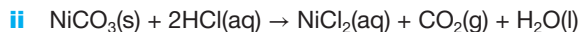
Pops with lighted splint held to mouth of tube.

ii hydrogen



- b i** Description could include: Green solid. Colourless acid. Bubbles of gas produced in the cold. Formation of green solution. Green powder disappears if excess acid is used.

Testing for gas: Bubble through lime water to give a white precipitate (milky, cloudy).



Some teachers may prefer to write this as:



**4 ► soluble**

sodium chloride

zinc nitrate

iron(III) sulfate

potassium sulfate

aluminium nitrate

ammonium chloride

magnesium nitrate

sodium phosphate

potassium dichromate(VI)

**insoluble**

lead(II) sulfate

calcium carbonate

lead(II) chloride

copper(II) carbonate

silver chloride

barium sulfate

calcium sulfate (almost insoluble)

nickel(II) carbonate

chromium(III) hydroxide

- 5 ► a** Dilute sulfuric acid in a beaker/flask. Heat gently. Add copper(II) oxide a little at a time until no more reacts. Filter into an evaporating basin. Evaporate gently until a sample will crystallise on cooling. Leave to crystallise. Separate and dry crystals.



- 6 ► a** using a pipette: Need to be able to measure the volume of sodium carbonate solution accurately so that exactly the same volume can be used later without the indicator.

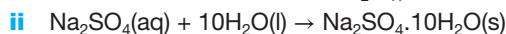
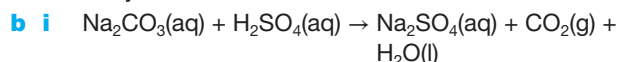
few drops of methyl orange were added: An indicator to show when the solution becomes 'neutral'. (In fact the colour change happens around pH 4. The pH changes very quickly around the end point for the indicator, and the difference in volume of acid added to take the pH from the desired pH 7 to the actual pH 4 for the indicator will be a fraction of a drop.)

until the solution became orange: Exactly the right volume of acid has been added (the 'neutral' colour for methyl orange).

without the methyl orange: Otherwise the crystals would be contaminated by methyl orange.

evaporated until... Shows when the solution is sufficiently concentrated to crystallise on cooling. If a sample will crystallise, so will the bulk of the solution.

left to cool... Crystals form on cooling because the solubility of the sodium sulfate is lower in the cold.



- 7 ► a** A solution of any soluble silver salt (silver nitrate is always used) plus a solution of any soluble chloride including dilute hydrochloric acid  
 $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl(s)}$

- b** A solution of any soluble calcium salt (calcium chloride or nitrate) plus a solution of a soluble carbonate (sodium, potassium or ammonium carbonate).  
 $\text{Ca}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{CaCO}_3(\text{s})$

- c** A solution of a soluble lead(II) salt (most commonly, lead(II) nitrate) plus a solution of any soluble sulfate including dilute sulfuric acid  
 $\text{Pb}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{PbSO}_4(\text{s})$

- d** A solution of a soluble lead(II) salt (most commonly, lead(II) nitrate) plus a solution of any soluble chloride including dilute hydrochloric acid.  
 $\text{Pb}^{2+}(\text{aq}) + 2\text{Cl}^-(\text{aq}) \rightarrow \text{PbCl}_2(\text{s})$

- 8 ►** Mix solutions of barium chloride or barium nitrate and sodium carbonate, potassium carbonate or ammonium carbonate. Filter, wash and dry the precipitate.  
 $\text{Ba}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{BaCO}_3(\text{s})$

- 9 ► a** A: dilute sulfuric acid + solid zinc (or zinc oxide, hydroxide or carbonate)  
 $\text{H}_2\text{SO}_4(\text{aq}) + \text{Zn(s)} \rightarrow \text{ZnSO}_4(\text{aq}) + \text{H}_2(\text{g})$   
 or  $\text{H}_2\text{SO}_4(\text{aq}) + \text{ZnO(s)} \rightarrow \text{ZnSO}_4(\text{aq}) + \text{H}_2\text{O(l)}$   
 or  $\text{H}_2\text{SO}_4(\text{aq}) + \text{Zn(OH)}_2(\text{s}) \rightarrow \text{ZnSO}_4(\text{aq}) + 2\text{H}_2\text{O(l)}$   
 or  $\text{H}_2\text{SO}_4(\text{aq}) + \text{ZnCO}_3(\text{s}) \rightarrow \text{ZnSO}_4(\text{aq}) + \text{H}_2\text{O(l)} + \text{CO}_2(\text{g})$

- b** C: solutions of a soluble barium salt + soluble sulfate (including  $\text{H}_2\text{SO}_4$ )  
 $\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})$

There is no point in writing full equations for precipitation reactions like this. The ionic equation is always easier.

- c** B: dilute nitric acid + potassium hydroxide or carbonate solution  
 $\text{HNO}_3(\text{aq}) + \text{KOH(aq)} \rightarrow \text{KNO}_3(\text{aq}) + \text{H}_2\text{O(l)}$   
 or  $2\text{HNO}_3(\text{aq}) + \text{K}_2\text{CO}_3(\text{aq}) \rightarrow 2\text{KNO}_3(\text{aq}) + \text{H}_2\text{O(l)} + \text{CO}_2(\text{g})$

- d** A: dilute nitric acid + solid copper(II) oxide or hydroxide or carbonate.  
 $2\text{HNO}_3(\text{aq}) + \text{CuO(s)} \rightarrow \text{Cu(NO}_3)_2(\text{aq}) + \text{H}_2\text{O(l)}$   
 or  $2\text{HNO}_3(\text{aq}) + \text{Cu(OH)}_2(\text{s}) \rightarrow \text{Cu(NO}_3)_2(\text{aq}) + 2\text{H}_2\text{O(l)}$   
 or  $2\text{HNO}_3(\text{aq}) + \text{CuCO}_3(\text{s}) \rightarrow \text{Cu(NO}_3)_2(\text{aq}) + \text{H}_2\text{O(l)} + \text{CO}_2(\text{g})$

(Unlike the other simple acids, nitric acid also reacts with copper metal, but it doesn't produce hydrogen, and is not on the Edexcel International GCSE syllabus. Any equation using copper to produce hydrogen should be disallowed.)

- e** C: solutions of a soluble lead(II) salt (normally the nitrate) + a soluble chromate(VI). (Since you don't know anything specifically about chromates, choose one of sodium, potassium or ammonium chromate (VI) – these are bound to be soluble because all sodium, potassium and ammonium salts are soluble.)  
 $\text{Pb}^{2+}(\text{aq}) + \text{CrO}_4^{2-}(\text{aq}) \rightarrow \text{PbCrO}_4(\text{s})$

- 10 ► a** acid:  $\text{H}_3\text{O}^+$  base:  $\text{CO}_3^{2-}$

- b** acid:  $\text{H}_2\text{SO}_4$  base:  $\text{MgO}$

- c** acid:  $\text{HNO}_3$  base:  $\text{NH}_3$

- d** acid:  $\text{H}_2\text{SO}_4$       base:  $\text{H}_2\text{O}$   
**e** acid:  $\text{HCl}$       base:  $\text{NH}_3$   
**f** acid:  $\text{CH}_3\text{COOH}$       base:  $\text{NH}_3$

## CHAPTER 18

- 1 ▶ a** chlorine      **b** ammonia      **c** carbon dioxide  
**d** hydrogen      **e** oxygen
- 2 ▶ a** Clean a nichrome or platinum wire by dipping it into concentrated hydrochloric acid and then into a flame until no colour shows. Moisten the wire with concentrated hydrochloric acid, dip it into the solid, and then back into flame. Lithium ions give a red flame colour.  
**b** Warm the solid very gently with sodium hydroxide solution. Test any gases given off with a piece of damp red litmus paper. If it turns blue, ammonia is being given off from an ammonium compound.  
**c** Make a solution in pure water. Add dilute hydrochloric acid + barium chloride solution. A white precipitate shows the sulfate ions.  
**d** Add dilute hydrochloric acid or dilute nitric acid. (Not sulfuric acid which forms an insoluble layer of calcium sulfate around the calcium carbonate, which stops the reaction.) Look for a colourless odourless gas turning limewater milky.  
**e** Make a solution in pure water. Add dilute nitric acid + silver nitrate solution. A yellow precipitate shows the iodide ions.
- 3 ▶ a** A = iron(III) chloride; B = iron(III) hydroxide; C = silver chloride  
**b**  $\text{B: Fe}^{3+}(\text{aq}) + 3\text{OH}^{-}(\text{aq}) \rightarrow \text{Fe}(\text{OH})_3(\text{s})$  or:  $\text{FeCl}_3(\text{aq}) + 3\text{NaOH}(\text{aq}) \rightarrow \text{Fe}(\text{OH})_3(\text{s}) + 3\text{NaCl}(\text{aq})$   
 $\text{C: Ag}^{+}(\text{aq}) + \text{Cl}^{-}(\text{aq}) \rightarrow \text{AgCl}(\text{s})$  or  $3\text{AgNO}_3(\text{aq}) + \text{FeCl}_3(\text{aq}) \rightarrow 3\text{AgCl}(\text{s}) + \text{Fe}(\text{NO}_3)_3(\text{aq})$
- 4 ▶ a** D = iron(II) sulfate; E = iron(II) hydroxide; F = barium sulfate  
**b**  $\text{E: Fe}^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Fe}(\text{OH})_2(\text{s})$   
or:  $\text{FeSO}_4(\text{aq}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{Fe}(\text{OH})_2(\text{s}) + \text{Na}_2\text{SO}_4(\text{aq})$   
 $\text{F: Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})$   
or  $\text{BaCl}_2(\text{aq}) + \text{MgSO}_4(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + \text{MgCl}_2(\text{aq})$
- 5 ▶ a** G = potassium carbonate; H = potassium nitrate; I = carbon dioxide (*Note: g could also be potassium hydrogen carbonate, but, because hydrogen carbonates haven't been covered in the course, a student is unlikely to come up with it.*)  
**b**  $\text{CO}_3^{2-}(\text{s}) + 2\text{H}^{+}(\text{aq}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$  or  $\text{K}_2\text{CO}_3(\text{s}) + 2\text{HNO}_3(\text{aq}) \rightarrow 2\text{KNO}_3(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
- 6 ▶ a** J = ammonium bromide; K = ammonia; L = silver bromide  
**b**  $\text{Ag}^{+}(\text{aq}) + \text{Br}^{-}(\text{aq}) \rightarrow \text{AgBr}(\text{s})$
- 7 ▶ a** Dissolve the mixture in the minimum possible amount of hot water, and then allow it to cool again. The small amount of the more soluble potassium carbonate will stay in solution, but crystals of potassium nitrate will be formed on cooling. Filter these off, and wash them on the filter paper with a small amount of very cold water. Allow them to dry.

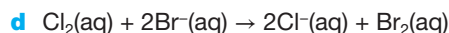
- b** Take a sample and add any named dilute acid. If the crystals are free of potassium carbonate, there won't be any fizzing.

- 8 ▶ a** Add a few drops of the liquid to some anhydrous copper(II) sulfate. If the colour changes from white to blue then water is present.  
**b** The boiling point of the liquid must be measured. If the liquid boils at  $100^{\circ}\text{C}$  it is pure water. The freezing/melting point can also be measured – pure water freezes/melts at  $0^{\circ}\text{C}$ .  
**c** The test-tube becoming warm indicates that an exothermic reaction has occurred. An exothermic reaction with sodium hydroxide solution could mean that the colourless liquid is an acid – neutralisation reactions are exothermic. The fact that there is no other visual reaction indicates that there is no metal present that has an insoluble hydroxide.  
The second test indicates that the solution does not contain carbonate ions – carbonate ions would produce fizzing with dilute hydrochloric acid.  
The third test indicates that no sulfate ions are present – the solution cannot be sulfuric acid.  
The last test indicates the presence of chloride ions. With the results of the other tests, this suggests that the solution is dilute hydrochloric acid.  
The student could confirm this by adding an indicator, to see whether the liquid is acidic.

## END OF UNIT 2 QUESTIONS

- 1 ▶ a** They all have 1 electron in the outer shell. (1)  
**b** Potassium oxide. (1)  
**c** The piece of lithium floats; hydrogen gas is formed; the final solution is alkaline.  
**d i** it forms a 1+ ion in compounds; it reacts with water and air  
**ii** The density increases from lithium to sodium but decreases from sodium to potassium (1)  
The trend is not clear for the whole group (1)
- 2 ▶ a** Because there will be no reaction (1)  
between a halogen and its halide ion / chlorine cannot reaction with chloride ions / bromine cannot reaction with bromide ions / iodine cannot reaction with iodide ions (1)
- b**
- |                   | potassium chloride solution | potassium bromide solution | potassium iodide solution |
|-------------------|-----------------------------|----------------------------|---------------------------|
| chlorine solution |                             | orange solution formed     | brown solution formed     |
| bromine solution  | no reaction                 |                            | brown solution formed     |
| iodine solution   | no reaction                 | no reaction                |                           |
- c** chlorine + potassium bromide → potassium chloride + bromine (1)





$\text{Cl}^-(\text{aq})$  (1)

$\text{Br}_2(\text{aq})$  (1)

balancing (1)

- e chlorine gains electrons to form chloride ions (1)  
bromide ions lose electrons to form bromine (1)  
gain of electrons is reduction and loss of electrons is oxidation (1)



ii test: damp litmus paper/universal indicator paper (1)  
positive result: the paper is bleached (goes white) (1)

iii test: add dilute nitric acid (1) followed by silver nitrate solution (1)  
positive result: white precipitate (1)

3 ► a nitrogen (1)

b i The oxygen in the air is used up as it reacts with the iron. (1)

Water is pushed in by the atmosphere to take its place/the pressure in the measuring cylinder is reduced. (1)

ii All the oxygen has reacted and so there is nothing to react with the iron.

iii Hydrated iron(III) oxide. (1)

iv There is no oxygen present (1); Oxygen is required for the splint to burn. (1)

v The total amount of oxygen present in the original sample of air is  $94 - 75 = 19 \text{ cm}^3$  of oxygen. (1)  
 $9 \text{ cm}^3$  of oxygen is used up after one day, therefore there is  $19 - 9 = 10 \text{ cm}^3$  of oxygen still present after one day. (1)

This is in a total volume of  $85 \text{ cm}^3$  (1) (reading on the measuring cylinder).

$$\text{Percentage oxygen} = \frac{10}{85} \times 100 = 11.8 \%. (1)$$

vi These will produce hydrogen when they react with the water present (1)

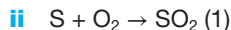
Hydrogen gas will take up space in the measuring cylinder, therefore the volume change will not show the volume of oxygen used up. (1)

vii Suggest one change that the student could make to the *apparatus* so he does not have to wait so long for the results. (1)

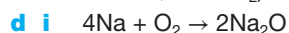
Use more iron filings/more finely-divided (higher surface area) iron filings/warm the apparatus. (1)

c Sulfur reacts with oxygen. A teacher burns a piece of sulfur in a gas jar of oxygen.

i blue



iii the litmus solution will be red (1)  
No-metal oxides are acidic and dissolve in water (if soluble, like  $\text{SO}_2$ ) to form acidic solutions. (1)



ii hydroxide ion (1)

all formulae (1)

balancing (1)

4 ► a Put dilute sulfuric acid into a beaker and heat it on a tripod and gauze using a Bunsen burner. (1)

Add excess copper(II) oxide (1)

Filter off the excess copper(II) oxide (1)

b Test: add hydrochloric acid followed by barium chloride solution (1)

Positive result: white precipitate (1)

c copper(II) hydroxide (1)

d i In a new line: Any 2: Concentration of copper(II) sulfate (1)

Volume of copper(II) sulfate solution (1)

How finely divided the metal is (1) (*if the metal is very finely divided it will react more quickly and the heat will be given out more quickly, so there is less chance for heat to be lost to the surroundings. This will give a more accurate value for the temperature change*)

ii Zinc > nickel > silver (1)

iii The temperature rise would be zero (1)

Copper will not react with copper(II) sulfate / copper ions (1)

iv The mass of 1 mole of zinc (65 g) is very similar to the mass of 1 mole of nickel (59 g) (1)

Because the student used the same mass of each metal, the number of moles of nickel and zinc is very similar. (1)

Because silver does not react it does not matter how much the student uses (1)

v Nickel gets oxidized. (1)

Ni loses electrons to form  $\text{Ni}^{2+}$  and oxidation is loss of electrons. (1)

e Heat the solution to drive off the water and form (small) crystals of hydrated copper(II) sulfate. (1)

Keep heating the solid. (1)

f i Dip a piece of (nichrome) wire into concentrated hydrochloric acid and then into the solid. (1)

Put the solid into the blue/roaring/non-luminous flame of a Bunsen burner. (1)

ii blue-green (1)

5 ► a An **ore** is a sample of rock that contains enough of a mineral for it to be worthwhile to extract the metal. (1)

b

	thermal decomposition	neutralisation	redox
reaction 1			✓
reaction 3	✓		



products (1)

balancing (1)

d carbon is the reducing agent (1)

it reduces the  $\text{MoO}_3$  by removing the oxygen (1)

e Molybdenum is below carbon in the reactivity series (1)

Carbon must be more reactive than molybdenum to be able to take the oxygen away from molybdenum oxide. (1)



**f** Because aluminium has to be extracted from its ore using electrolysis. (1)

Electrolysis is very expensive because it uses a large amount of electricity, therefore aluminium is more expensive than carbon. (1)

**g** Add sodium hydroxide solution and warm. (1)

If ammonium ions are present a colourless gas will be given off that turns moist red litmus paper blue. (1)

**6 ► a** Mix together solutions (1) of lead(II) nitrate and sodium sulfate (1). Filter the mixture (1) Wash the residue with distilled/pure water (1); Leave the solid in a warm oven to dry. (1)

**b** Because lead(II) iodide is yellow and insoluble in water. (1)

The precipitate could just be lead(II) iodide. (1)



Dot and cross in the covalent bond. (1)

All other electrons. (1)

**b** A proton/hydrogen ion/ $\text{H}^+$  is transferred from the HCl to the water. (1)

HCl is an acid because it donates a proton/hydrogen ion/ $\text{H}^+$  (1)

$\text{H}_2\text{O}$  is a base because it accepts a proton/hydrogen ion/ $\text{H}^+$  (1)

**c**

Reagent	Solid or solution?		Salt	Other product
magnesium	solid	hydrochloric acid	magnesium chloride	hydrogen
copper(II) oxide	solid	hydrochloric acid	copper(II) chloride	water
sodium hydroxide	solution	hydrochloric acid	sodium chloride	water
silver nitrate	solution	hydrochloric acid	silver chloride	nitric acid

**d i** a white precipitate is formed (1)

**ii** hydrochloric acid contains chloride ions (1)

Will always give a positive test for chloride ions. (1)

**e**  $\text{CuCO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{CuCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$   
 reactants (1)  
 products (1)  
 balance (1)

**8 ► a** iron(II) (1) chloride (1)

**b** B = iron(II) hydroxide (1)  $\text{Fe}(\text{OH})_2$  (1)

**c** D = iron(III) hydroxide (1)  $\text{Fe}(\text{OH})_3$  (1)

**d** redox (1)

The iron(II) is oxidised to iron(III) (1)

**9 ► a**  $\text{Na}_2\text{SO}_4$  (1)

**b** Sodium hydroxide solution/sulfuric acid is measured out using a pipette and put into conical flask/beaker. (1)  
 Sulfuric acid/sodium hydroxide placed in burette. (1)

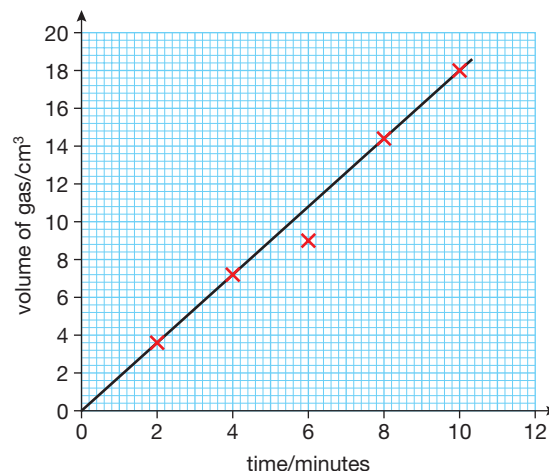
Named indicator added e.g. methyl orange/phenolphthalein but not universal indicator. (1)

Acid/alkali added to alkali/acid until indicator changes colour (1)

Procedure repeated without indicator using same volumes. (1)

Sodium carbonate/sodium hydrogencarbonate could also be used instead of sodium hydroxide.

**c i**



all point plotted correctly (2) – 1 for each mistake  
 line of best fit (1)

**ii** the point at 6 minutes (1);  $10.8 \pm 0.2 \text{ cm}^3$  (1)

**iii** as the time increases the volume of gas increases (1)  
 directly proportional relationship (1)  
 just stating that the volume is directly proportional to the time scores (2)

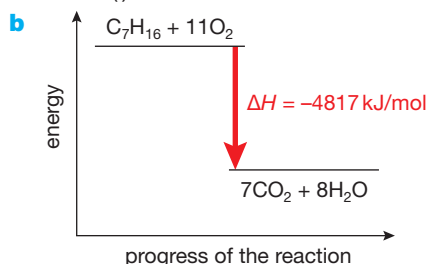
**iv** test: burning splint (1)  
 positive result: squeaky pop (1)

## UNIT 3 ANSWERS

### CHAPTER 19

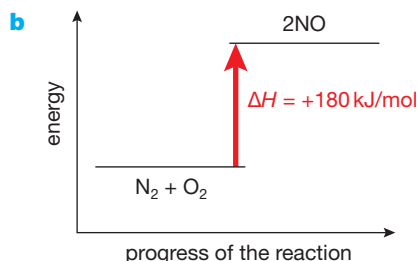
**1 ► a** A reaction in which heat energy is given out to the surroundings.

Correctly balanced equations for any two exothermic reactions e.g. any combustion reactions (metals, hydrogen, hydrocarbons, etc, in oxygen), neutralisation reactions involving oxides or hydroxides and acids, magnesium and acids. For example,  $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l})$  or  $\text{NaOH}(\text{aq}) + \text{HCl}(\text{aq}) \rightarrow \text{NaCl}(\text{aq}) + \text{H}_2\text{O}(\text{l})$

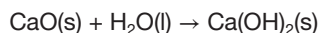


- c Energy is needed to break the bonds in heptane and in oxygen. Energy is released when new bonds are made to produce the carbon dioxide and water. More energy is released when the new bonds are made than was used to break the old ones.

- 2 ► a A reaction in which heat energy is absorbed from the surroundings.



- 3 ► a exothermic    b exothermic    c endothermic  
d exothermic    e endothermic    f exothermic
- 4 ► The cans have two chambers. The outer one contains the beverages or food to be heated. The inner chamber contains calcium oxide and water. The calcium oxide and water are kept separate by a seal. Pushing a button on the bottom of the can breaks the seal and allows the calcium oxide and water to come into contact. You get the reaction below which is exothermic.



- 5 ► a Bonds broken:  $4 \times \text{C-H} = 4 \times (+413) = +1652 \text{ kJ}$   
 $1 \times \text{Br-Br} = +193 \text{ kJ}$   
 total = +1845 kJ  
 Bonds made:  $3 \times \text{C-H} = 3 \times (-413) = -1239 \text{ kJ}$   
 $1 \times \text{C-Br} = -290 \text{ kJ}$   
 $1 \times \text{H-Br} = -366 \text{ kJ}$   
 total = -1895 kJ  
 Overall change = +1845 - 1895 = -50 kJ (exothermic)
- b Bonds broken:  $1 \times \text{H-H} = +436 \text{ kJ}$   
 $1 \times \text{Cl-Cl} = +243 \text{ kJ}$   
 total = +679 kJ  
 Bonds made:  $2 \times \text{H-Cl} = 2 \times (-432) = -864 \text{ kJ}$   
 Overall change = +679 - 864 = -185 kJ (exothermic)
- c Bonds broken:  $2 \times \text{H-H} = 2 \times (+436) = +872 \text{ kJ}$   
 $1 \times \text{O=O} = +498 \text{ kJ}$   
 total = +1370 kJ  
 Bonds made:  $4 \times \text{O-H} = 4 \times (-464) = -1856 \text{ kJ}$   
 Overall change = +1370 - 1856 = -486 kJ (exothermic)
- d Bonds broken:  $1 \times \text{N}\equiv\text{N} = +944 \text{ kJ}$   
 $3 \times \text{H-H} = 3 \times (+436) = +1308 \text{ kJ}$   
 total = +2252 kJ  
 Bonds made:  $6 \times \text{N-H} = 6 \times (-388) = -2328 \text{ kJ}$   
 Overall change = +2252 - 2328 = -76 kJ (exothermic)
- 6 ► a Her first two results weren't reliable. There was too much difference between them.  
 b Any two from: Wear a lab coat; avoid touching the hot copper can; tie long hair back.  
 c Heat energy change =  $Q = mc\Delta T$   
 Mass of water being heated =  $m = 100 \text{ g}$

c is the specific heat capacity of the water (it is the water that is being heated)

$$c = 4.18 \text{ J/g}^\circ\text{C}$$

$$\text{Temperature change of water} = \Delta T = 55.0 - 19.0 = 36.0^\circ\text{C}$$

$$\text{Heat evolved} = Q = mc\Delta T = 100 \times 4.18 \times 36.0 \text{ J} = 15048 \text{ J} = 15.0 \text{ kJ to 3 significant figures}$$

- d Mass of hexane burnt =  $35.62 - 35.23 \text{ g} = 0.39 \text{ g}$

$$\text{Heat evolved per gram} = \frac{15.048}{0.39} = 38.6 \text{ kJ/g to 3 significant figures}$$

The combustion reaction is exothermic, therefore heat energy released per gram = -38.6 kJ/g

- e Number of moles of hexane,  $\text{C}_6\text{H}_{14}$  burnt =  
 $n = \frac{\text{mass (m)}}{\text{relative molecular mass (Mr)}} = \frac{0.39}{86} = 0.00453 \text{ mol}$

The molar enthalpy change of combustion of hexane ( $\Delta H$ )

$$= \frac{\text{heat energy change (Q)}}{\text{number of moles of ethanol burned (n)}}$$

$$= \frac{15.048}{0.00453} = 3320 \text{ kJ/mol to 3 significant figures}$$

3 significant figures

The combustion reaction is exothermic, therefore  $\Delta H = -3320 \text{ kJ/mol}$

- f Any two from: Misreading one of the weighings of the spirit burner so that it looked as if less hexane had been burnt than was really the case; misreading the thermometer to give a final temperature higher than it should have been; adding less than  $100 \text{ cm}^3$  of water to the flask, so that the temperature went up more than it should because the heat was going into a smaller volume of water.
- g Any two from: Heat loss to the surroundings; heat is lost to warm up the copper calorimeter or the thermometer; incomplete combustion of the fuel.

- 7 ► a Heat energy change =  $Q = mc\Delta T$   
 Mass of solution being heated =  $m = 50 \text{ g}$ , the mass of the lithium chloride is relatively small and it is ignored in the calculation.  
 c is the specific heat capacity of the diluted solution of lithium chloride, which we assume to be the same as the heat capacity of water  $c = 4.18 \text{ J/g}^\circ\text{C}$   
 Temperature change of water =  $\Delta T = 33.5 - 17.0 = 16.5^\circ\text{C}$   
 Heat evolved =  $Q = mc\Delta T = 50 \times 4.18 \times 16.5 \text{ J} = 3448.5 \text{ J} = 3.45 \text{ kJ to 3 significant figures}$
- b Number of moles of lithium chloride dissolved =  
 $n = \frac{\text{mass (m)}}{\text{relative formula mass (Mr)}} = \frac{5.15}{42.5} = 0.121 \text{ mol}$
- c The enthalpy change of solution  

$$\Delta H = \frac{\text{heat energy change (Q)}}{\text{number of moles of lithium chloride dissolved (n)}}$$

$$= \frac{3.45}{0.121} = 28.5 \text{ kJ/mol to 3 significant figures}$$

The dissolving of lithium chloride is exothermic, therefore  $\Delta H = -28.5 \text{ kJ/mol}$

- d** Any two from: Heat loss to the surroundings; heat is lost to warm up the solution container or the thermometer; incomplete transfer of the salt from the weighing boat into the water; some salt crystal does not dissolve completely in water.

**8 ► a** The number of moles of NaOH( $n$ ) =  
 concentration( $C$ )  $\times$  volume( $V$ ) =  $0.50 \times \frac{200}{1000} = 0.10 \text{ mol}$

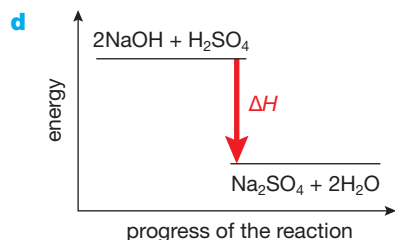
- b** Heat given out by the reaction =  $Q = mc\Delta T = 300 \times 4.18 \times 4.5 = 5643 \text{ J} = 5.64 \text{ kJ}$  to 3 significant figures  
 Here we assume the density of the reaction mixture is the same as water and the specific heat capacity of the mixture is the same as heat capacity of water.

- c** The enthalpy change of neutralisation

$$\Delta H = \frac{\text{heat energy change (Q)}}{\text{number of moles of NaOH reacted (n)}}$$

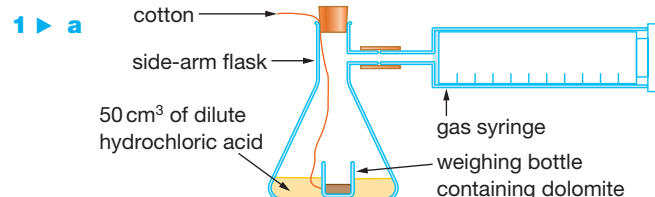
$$= \frac{5.65}{0.10} = 56.4 \text{ kJ/mol to 3 significant figures}$$

The neutralisation reaction is exothermic, therefore  $\Delta H = -56.4 \text{ kJ/mol}$ .



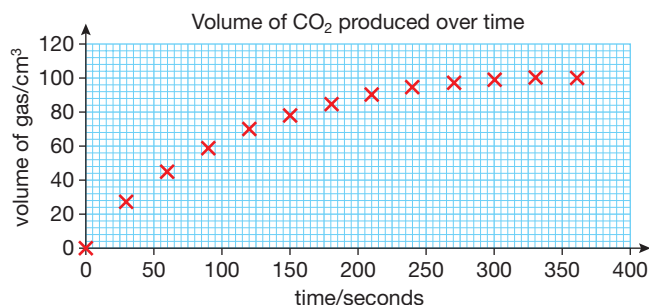
- e** As double the amount of reactants react, the amount of heat energy released from the reaction doubles. The same amount of solution gets heated up so the temperature rise should double.

## CHAPTER 20



Collection of gas over water into an inverted measuring cylinder is an acceptable alternative. Pieces of dolomite are put in a weighing bottle into the conical flask before the reaction starts. This is to make sure that the reaction can be started without losing any carbon dioxide.

- b** Graph should be completely smooth with the axes properly labelled.



- c** At the very beginning of the reaction. The reaction can only happen when acid particles collide with the solid dolomite. Numbers of acid particles per unit volume are greatest at the beginning of the reaction before any get used up. Therefore the greatest number of collisions per second and the fastest reaction is at the beginning.

- d** 70 seconds (read this off the graph and allow some tolerance depending on the size of graph paper available.)

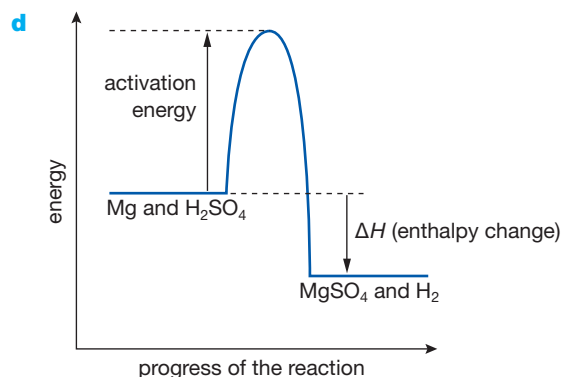
- e** Volume produced within the first 80 seconds = 55 cm  
 (The average rate =  $\frac{55}{80} = 0.688 \text{ cm}^3/\text{s}$  to 3 significant figures)

- f** **i** There would be a lower initial rate; same volume of gas.  
**ii** There would be a lower initial rate; half the volume of gas (50 cm<sup>3</sup>).  
**iii** The initial rate would be the same; half the volume of gas (50 cm<sup>3</sup>). (The initial rate depends on the original concentration of the acid that is still the same.)  
**iv** The initial rate would be faster; same volume of gas.

- 2 ► a** Time taken for the reaction would increase. A reaction happens when acid particles collide with the magnesium. The concentration of acid is less, so there is less number of acid particles per unit volume. There will be fewer collisions per second, and therefore a slower reaction.

- b** Time taken for the reaction would decrease. The acid particles are moving faster and so collide with the magnesium more often. Reaction only happens if the energy of the collision equals or exceeds activation energy. At higher temperatures more acid particles have energy greater than or equal to the activation energy, so a greater proportion of the collisions are successful.

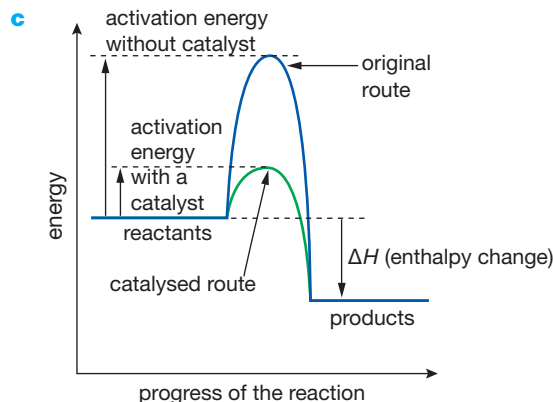
- c** Answers could include: Acid will be used up quickly immediately around the magnesium – stirring brings fresh acid into contact with it; bubbles of hydrogen form around the magnesium, preventing acid from reaching it – stirring helps to dislodge the bubbles; bubbles of hydrogen lift the magnesium to the surface (sometimes above the surface) of the acid, lowering contact between acid and magnesium – stirring helps to prevent this.



- 3 ► a** Reactions only happen if collisions have energies equalling or exceeding activation energy. Catalysts provide an alternative route for the reaction with a lower activation energy. More particles now have energy greater than or equal to the activation energy, so there will be more successful collisions per unit time.

- b** To find out whether it speeds the reaction up: You could do this most simply by having two test-tubes with equal volumes of the same hydrogen peroxide solution side-by-side. Add some copper(II) oxide to one and look for the faster production of bubbles.

To show that it is unchanged: Use a known mass of copper(II) oxide. When the reaction stops, filter through previously weighed filter paper, allow to dry, and re-weigh. Show that the mass of copper(II) oxide is unchanged. (If it had changed, and you hadn't lost any during the separation process, it must have reacted in some way.)



## CHAPTER 21

- 1 ▶ a** The reversible symbol ' $\rightleftharpoons$ ' shows that the reaction can go both ways. The reactants can form products and the products can react to form the reactants.
- b** White solid decomposes to form colourless gases.
- c** Ammonium chloride, ammonia and hydrogen chloride
- 2 ▶ a** Right
- b** Left
- c** Left
- d** No change
- e** No change
- 3 ▶ a** 'Dynamic' means that the reactions are still happening and the rate of the forward reaction is equal to the rate of the reverse reaction. 'Equilibrium' means that the concentrations of the reactants and the products remain constant.
- b** The proportion of sulfur trioxide will increase. An increase in pressure shifts the position of equilibrium to the right because this side has fewer moles of gas molecules.
- c** The position of equilibrium shifts to the right. A decrease in temperature favours the forward exothermic reaction.
- 4 ▶ a** Less carbon monoxide and hydrogen would be produced. (A high pressure favours the reverse reaction, as the left hand side of the reaction has fewer number of moles of gas molecules.)
- b** A high temperature favours the endothermic forward reaction and shifts the position of equilibrium to the right hand side of the reaction.

- c** Using a catalyst increases the rate of the reaction, as it provides an alternative pathway for the reaction with lower activation energy. More particles will have energy greater than or equal to the activation energy so there will be more successful collisions per unit time. More products will be made quickly and relatively low temperature and pressure can be used in the presence of a catalyst. This saves energy and money. Adding a catalyst has no effect on the position of equilibrium, as the rate of the forward reaction and the rate of the reverse reaction are increased by an equal amount.

- 5 ▶ a** A high pressure favours the reaction producing fewer moles of gas, shifting the position of equilibrium to the right hand side.
- b** Very high pressures are expensive to produce for two reasons: all the reaction vessels and pipework have to be built much more strongly, and it takes a lot more energy to drive the compressors producing the very high pressures. The extra ammonia produced isn't worth the extra cost, so a compromise of 200 atmospheres pressure is used.
- c** At a low temperature. Lower temperature favours the forward exothermic reaction, shifting the position of equilibrium to the right hand side to produce more ammonia.
- d** At low temperatures the reaction is extremely slow even in the presence of a catalyst.  $450^{\circ}\text{C}$  is chosen because it gives a compromise between the rate and the yield of the reaction.

## END OF UNIT 3 QUESTIONS



correct number of outer shell electrons for both H and Br (1)

sharing of 1 pair of electrons between the two atoms (1)

**b**

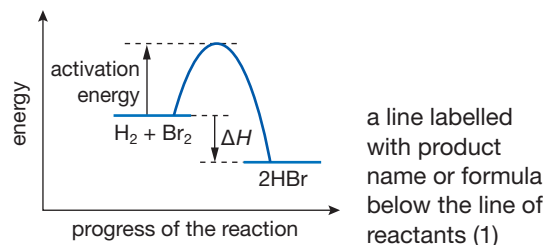
bond	bond energy / kJ/mol
H-H	436
Br-Br	196
H-Br	368

Bonds broken:  $1 \times \text{H-H} = +436 \text{ kJ}$   
 $1 \times \text{Br-Br} = +196 \text{ kJ}$   
 total = +632 kJ

Bonds made:  $2 \times \text{H-Br} = -736 \text{ kJ}$

Overall change =  $+632 - 736 = -104 \text{ kJ}$  (exothermic)

- c i and ii**



use of single-headed arrows to represent activation energy (1) and enthalpy change (1), as shown

2 ► a i 19.3 (1), 16.6 (1), 2.7 (1)

ii A

b i  $Q = mc\Delta T = 100.0 \times 4.2 \times (23.2 - 15.9) = 3.066 \text{ kJ}$   
calculation of temperature change (1)

using the correct mass (100 g, award mark if 105 g is used) (1)

correct answer for  $Q$  (1)

ii  $n = \text{mass of calcium chloride} / M_r = \frac{5}{(40 + 35.5 \times 2)}$   
 $= \frac{5}{111} = 0.0450 \text{ mol}$  to 3 significant figures.

calculation of  $M_r$  (1)

correct answer to 2–4 significant figures (1)

iii  $\Delta H = \frac{3.066}{0.0450} = 68.1 \text{ kJ/mol}$  to 3 significant figures.

The molar enthalpy change is  $-68.1 \text{ kJ/mol}$  as the dissolving is exothermic.

dividing  $Q$  by  $n$  (1)

correct answer (must have a negative sign) (1)

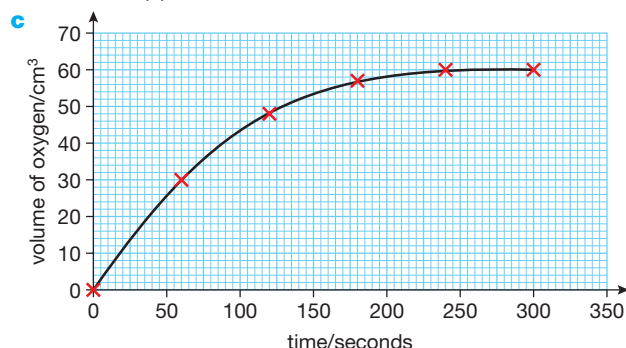
c heat loss to the surrounding air through beaker (1)  
some magnesium chloride is left on the weighing boat/  
did not dissolve in water completely (1)

3 ► a  $2\text{H}_2\text{O}_2(\text{aq}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g})$

correct formulae of reactants and products (1)

correct balancing (1)

b To prevent the loss of oxygen at the beginning of the reaction. (1)



correct labelling of axes with units (1)

correct points plotted (2)

a smooth curve of best fit, going through all of the points (1)

d i 125 seconds

ii  $43 \text{ cm}^3$

iii  $\frac{53}{150} = 0.353 \text{ cm}^3/\text{sec}$

correct numerical answer (1)

correct unit (1)

e Reaction has stopped (1) because all the hydrogen peroxide is used up (1).

f i B (1)

ii Weigh a sample of manganese(IV) oxide and add to hydrogen peroxide (1). Oxygen is produced at a faster rate with manganese(IV) oxide than without (1). Filter the reaction mixture and dry the solid (1).

Re-weigh the solid and the mass should be the same as before if it acts as a catalyst (1).

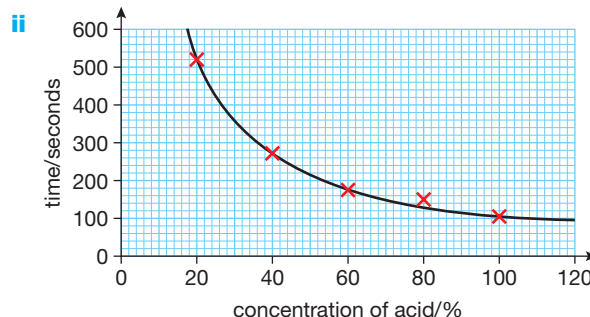
g Shallower curve than the original (1). but the end volume remains the same (1).

h Shallower curve than the original (1) and only  $30 \text{ cm}^3$  (half the volume) of gas is produced (1).

4 ► a the mass (1). and the surface area / the size (1).

b To let the gas escape but keep the liquid inside the conical flask. (1)

c i B



correct labelling of axes with units (1)

correct points plotted (2)

a smooth curve of best fit, going through all but one point (1)

iii 150 seconds

iv C

d i The mass of  $\text{CO}_2$  given off is directly (1) proportional (1) to the concentration of the acid.

ii Increasing the concentration of the acid increases the rate of reaction. (1) This is because there are more acid particles within a fixed volume (1) so the frequency of successful collision between the acid and the marble chip increases (1).

5 ► a i The gas particles have more kinetic energy (1) so more particles have energy greater than or equal to the activation energy (1). There are more successful collisions per unit time (1).

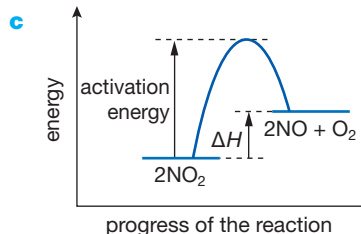
ii The rate increases (1) as gas particles are closer together and there are more frequent collisions (or more successful collisions per unit time) (1).

iii Reactions happen on the surface of the catalyst (1) and gauzes have greater surface area (1).

b The use of a catalyst increases the rate of the reaction without itself being used up (1). It saves money which would be spent on increasing temperature or pressure (1).

6 ► a reversible reaction (1), enthalpy change (1)

b The forward reaction is endothermic. (1)





- d i** Increasing the temperature favours the forward endothermic reaction. (1)  
The position of equilibrium shifts to the right. (1)  
The reaction mixture becomes more colourless. (1)
- ii** Increasing the pressure favours the left hand side of the reaction as it has fewer moles of moles of gases. (1)  
The position of equilibrium shifts to the left. (1)  
The reaction mixtures becomes more brown. (1)
- e** No change in the position of equilibrium (1) The use of a catalyst increases the rate of both the forward and reverse reaction equally (1).

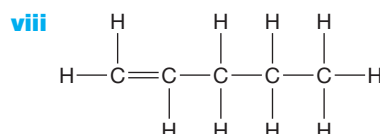
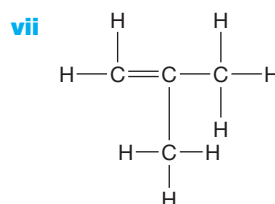
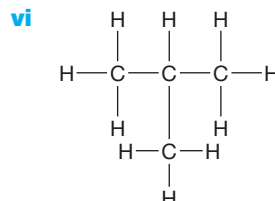
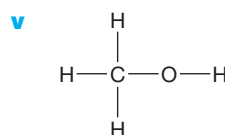
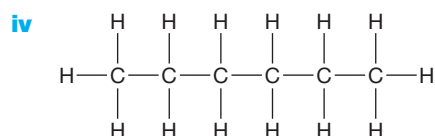
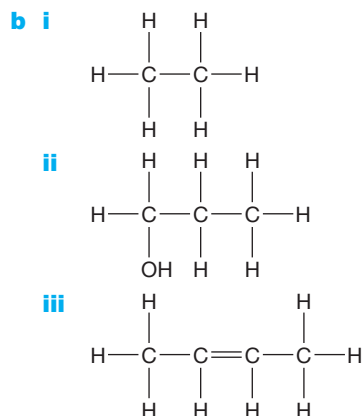
**7 ► a** A and D

- b i** Increasing the pressure favours the right hand side of the reaction which has fewer moles of moles of gas molecules. (1)  
The position of equilibrium shifts to the right. (1)  
The yield of  $\text{NH}_3$  increases. (1)
- ii** As temperature increases, the percentage of  $\text{NH}_3$  at equilibrium decreases. (1)
- iii**  $\Delta H$  for the forward reaction is negative. (1)  
Increasing temperature favours the reverse endothermic reaction. (1)
- iv** Increasing temperature increases the rate of reaction. (1)  
More particles will have energy greater than or equal to the activation energy. (1)  
The frequency of successful collisions increases. (1)

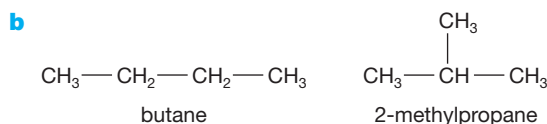
## UNIT 4 ANSWERS

### CHAPTER 22

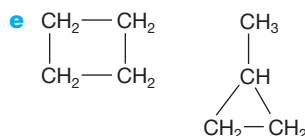
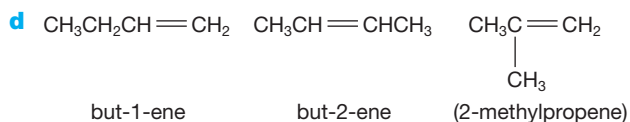
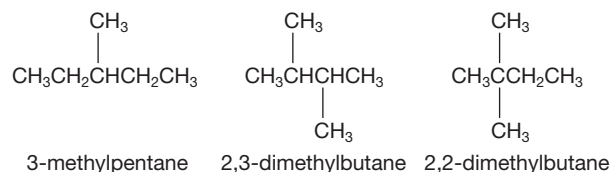
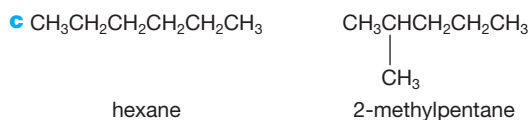
- 1 ► a i** methane  
**ii** propane  
**iii** hexane  
**iv** propene  
**v** ethene  
**vi** but-1-ene  
**vii** propan-1-ol  
**viii** butan-2-ol



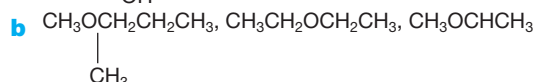
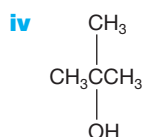
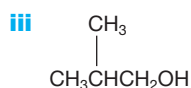
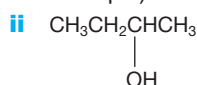
- 2 ► a** The existence of molecules with the same molecular formula but different structural formulae.



(In these and subsequent formulae, if you aren't asked specifically for displayed formulae, these quicker forms are acceptable.)



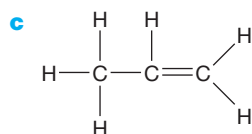
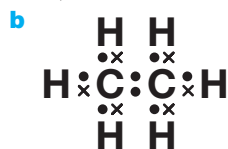
- 3 ► a i (first answer is shown in the textbook as an example)



These compounds are known as ethers.

- 4 ► a addition  
 b combustion  
 c addition  
 d substitution  
 e combustion  
 f substitution

- 5 ► a C, it contains element other than carbon and hydrogen.



- d  $\text{C}_n\text{H}_{2n+2}$   
 e Same functional group / similar chemical properties; Shows a gradation in physical properties; Each member differs from the next by a  $-\text{CH}_2-$ .  
 f  $\text{C}_5\text{H}_{12} + 8\text{O}_2 \rightarrow 5\text{CO}_2 + 6\text{H}_2\text{O}$

## CHAPTER 23

- 1 ► a Carbon and hydrogen.  
 b The crude oil (mixture of hydrocarbons) is heated until it boils. The vapour passes into a fractionating column. The temperature is higher at the bottom of the column than at the top. Different fractions condense and are drawn off at different heights in the column. The hydrocarbons with the highest boiling points (longer chains) condense towards the bottom of the column. The smaller hydrocarbon molecules travel further up the column until they condense and are drawn off.  
 c Gasoline – petrol for car; Diesel – fuel for lorries or buses  
 d Any two from: refinery gas, kerosene, fuel oil or bitumen.  
 e The average size of the molecules in gasoline is smaller than in diesel. Diesel is darker in colour and more viscous than gasoline.

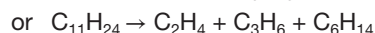
- f Any between  $n = 5 - 10$  for  $\text{C}_n\text{H}_{2n+2}$



- ii Carbon monoxide is poisonous as it reduces the blood's ability to carry oxygen around the body.

- 2 ► a The sulfur (or sulfur compound) burns to make sulfur dioxide. The sulfur dioxide reacts with water and oxygen in the atmosphere to produce sulfuric acid that falls as acid rain.  
 b The spark in the engine causes nitrogen to react with oxygen to give various oxides of nitrogen.  
 c Sulfur dioxide reacts with water and oxygen in the atmosphere to produce sulfuric acid. The sulfuric acid in acid rain can react with calcium carbonate and corrode the buildings.

- 3 ► a Crude oil produces too many larger hydrocarbons / not enough of the more useful smaller ones. Smaller alkanes can be used as fuel for cars. Cracking also produces alkenes that can be used to make polymers.  
 b Heat the vaporised fraction in the presence of a silicon dioxide or aluminium oxide catalyst at  $600-700^\circ\text{C}$ .  
 c  $\text{C}_{11}\text{H}_{24} \rightarrow 2\text{C}_2\text{H}_4 + \text{C}_7\text{H}_{16}$   
 d Any other valid cracking equation starting with  $\text{C}_{11}\text{H}_{24}$ . For example:



or lots of other variants. In each case, at least one hydrocarbon should be an alkane ( $\text{C}_n\text{H}_{2n+2}$ ), and at least one an alkene ( $\text{C}_n\text{H}_{2n}$ ).

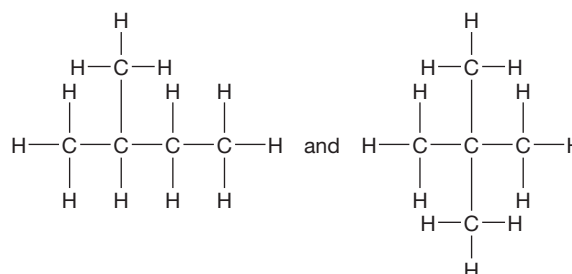
- 4 ► This is entirely open to your imagination and ability to think both logically and laterally. It is impossible to suggest 'right' answers.

## CHAPTER 24

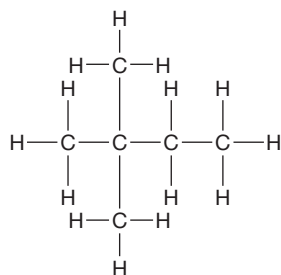
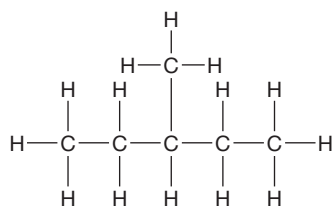
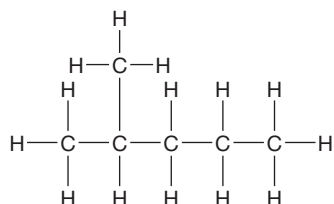
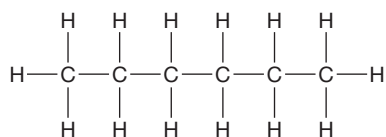
- 1 ► a Contains only C-C single bonds and has no double or triple bonds.  
 b i  $\text{C}_{11}\text{H}_{24}$   
 ii Liquid  
 iii  $\text{C}_{11}\text{H}_{24}(\text{l}) + 17\text{O}_2(\text{g}) \rightarrow 11\text{CO}_2(\text{g}) + 12\text{H}_2\text{O}(\text{l})$   
 iv  $2\text{C}_{11}\text{H}_{24} + 23\text{O}_2 \rightarrow 22\text{CO} + 24\text{H}_2\text{O}$   
 Carbon monoxide is poisonous, as it reduces the ability of the blood to carry oxygen around the body.

- 2 ► a  $\text{C}_5\text{H}_{12}$   
 b  $\text{C}_4\text{H}_{10}$   
 c  $\text{C}_6\text{H}_{14}$

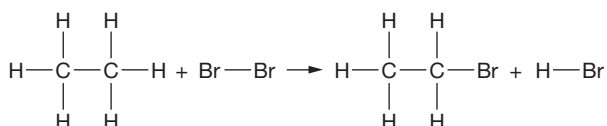
- 3 ► a



**b** Any 2 of the following:



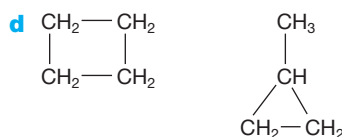
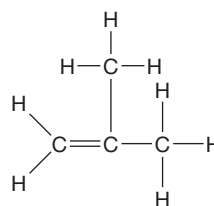
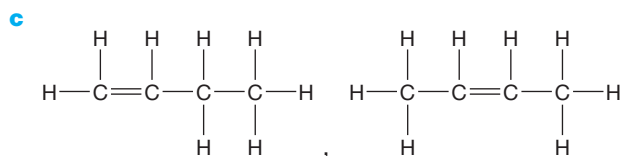
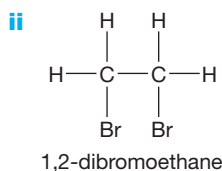
- 4 ▶ a** UV light  
**b** Substitution  
**c**



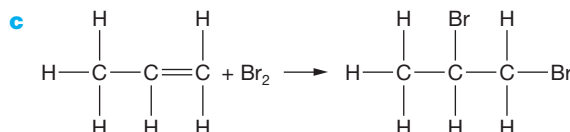
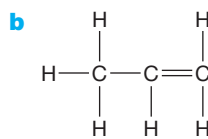
- 5 ▶ a**  $\text{C}_4\text{H}_{10} + \text{Cl}_2 \rightarrow \text{C}_4\text{H}_9\text{Cl} + \text{HCl}$   
**b**  $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{Cl}$        $\text{CH}_3\text{CH}(\text{Cl})\text{CH}_2\text{CH}_3$   
 1-chlorobutane                  2-chlorobutane

## CHAPTER 25

- 1 ▶ a i** Any two from: same general formula; same functional group or similar chemical properties; shows a gradation in physical properties; each member differs from the next by a  $-\text{CH}_2-$ .  
**ii** Compounds containing one or more carbon carbon double or triple bond.  
**b i** Starting: orange; Finishing: colourless



**2 ▶ a** a carbon-carbon double bond



- d i**  $\text{CH}_3\text{CH}_3 + \text{Br}_2(\text{g}) \rightarrow \text{CH}_3\text{CH}_2\text{Br} + \text{HBr}(\text{g})$   
**ii** In this (substitution) reaction, one of the hydrogens has been replaced by a bromine during the reaction. In the previous (addition) reaction, nothing was lost when the two molecules combined together.

- 3 ▶ a** 1,2-dichloropropane  $\text{CH}_2\text{ClCHClCH}_3$   
**b** ethane  $\text{CH}_3\text{CH}_3$

**4 ▶ a** Cracking.

Conditions: 1) The presence of a silicon dioxide (silica) or aluminium oxide (alumina) catalyst 2) heat the vaporized alkanes at high temperatures of 600 – 700 °C.

**b** Test: shake with bromine water

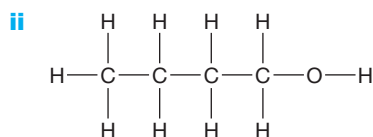
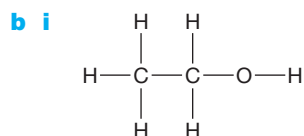
Result with propane: no colour change (solution remains orange)

Result with propene: solution changes colour from orange to colourless

- 5 ▶ a**  $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{OH}$  or  $\text{CH}_3\text{CH}_2\text{CH}(\text{OH})\text{CH}_3$   
 butan-1-ol                  butan-2-ol  
**b**  $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{Cl}$  or  $\text{CH}_3\text{CH}_2\text{CHClCH}_3$   
 1-chlorobutane                  2-chlorobutane  
**c**  $\text{CH}_3\text{CH}_2\text{CHBrCH}_2\text{Br}$   
 1,2-dibromobutane

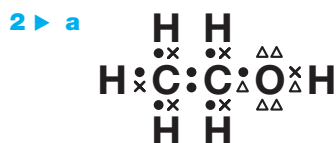
## CHAPTER 26

- 1 ▶ a i**  $\text{CH}_3\text{OH}$   
**ii**  $\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$



**c i** Butan-1-ol

**ii** Ethanol



**b** Any two from: 300 °C; 60–70 atmosphere pressure; phosphoric acid catalyst

**c i** Sugar cane

**ii** Yeast provides the enzyme (biological catalyst) for the reaction; 30–40 °C is the optimum temperature for the reaction. If the temperature is too low, the rate of the reaction is too slow. If the temperature is too high, the enzyme will be denatured; In the presence of air (aerobic conditions), enzymes in the yeast produce carbon dioxide and water instead of ethanol.

**iii** Countries like Brazil have plenty of land and warm climate to grow sugar cane. There is also a lack of crude oil resource in these countries, so it is too expensive to hydrate ethene to produce ethanol.

**d**  $\text{C}_2\text{H}_5\text{OH} + 3\text{O}_2 \rightarrow 2\text{CO}_2 + 3\text{H}_2\text{O}$

**3 ► a**  $\text{C}_6\text{H}_{12}\text{O}_6 \rightarrow 2\text{C}_2\text{H}_5\text{OH} + 2\text{CO}_2$

**b i** potassium dichromate(VI) in dilute sulfuric acid

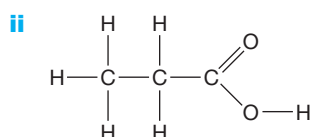
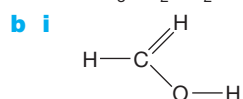
**ii** orange to green

**iii**  $\text{CH}_3\text{CH}_2\text{OH} + 2[\text{O}] \rightarrow \text{CH}_3\text{COOH} + \text{H}_2\text{O}$

**c**  $\text{C}_2\text{H}_5\text{OH} \rightarrow \text{C}_2\text{H}_4 + \text{H}_2\text{O}$

## CHAPTER 27

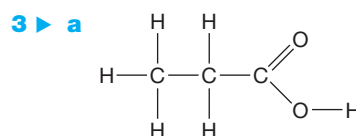
**1 ► a i**  $\text{CH}_3\text{COOH}$   
**ii**  $\text{CH}_3\text{CH}_2\text{CH}_2\text{COOH}$



**c i** Butanoic acid

**ii** Ethanoic acid

**2 ► c**



**b** Orange to green.

**c** Propan-1-ol is very flammable and should not be heated with an open flame. The vapour can easily catch fire when heated with a Bunsen burner.

**d i** bubbles/fizzes/effervescence; Mg decreases in size  
 $\text{Mg}(\text{aq}) + 2\text{CH}_3\text{COOH}(\text{aq}) \rightarrow (\text{CH}_3\text{COO})_2\text{Mg}(\text{aq}) + \text{H}_2(\text{g})$

**ii** bubbles/fizzes/effervescence

$\text{Na}_2\text{CO}_3(\text{aq}) + 2\text{CH}_3\text{COOH}(\text{aq}) \rightarrow 2\text{CH}_3\text{COONa}(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$

**4 ► a**  $\text{CH}_3\text{COOH} + \text{NaHCO}_3 \rightarrow \text{CH}_3\text{COONa} + \text{CO}_2 + \text{H}_2\text{O}$

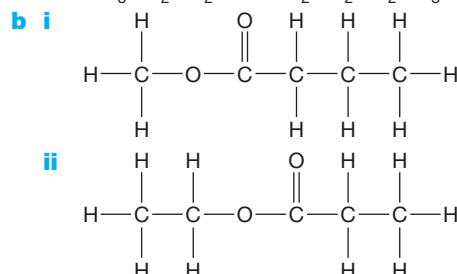
**b** Neutralisation.

**c** Bubbles will form.

## CHAPTER 28

**1 ► a i**  $\text{CH}_3\text{COOCH}_2\text{CH}_3$

**ii**  $\text{CH}_3\text{CH}_2\text{CH}_2\text{COOCH}_2\text{CH}_2\text{CH}_2\text{CH}_3$



**c i** Methyl propanoate.

**ii** Ethyl methanoate.

**2 ► a**  $\text{CH}_3\text{COOCH}_2\text{CH}_3$        $\text{CH}_3\text{COOCH}_2\text{CH}_2\text{CH}_3$   
 $\text{CH}_3\text{CH}_2\text{COOCH}_2\text{CH}_3$        $\text{CH}_3\text{CH}_2\text{COOCH}_2\text{CH}_2\text{CH}_3$

**b** • Add 1 cm<sup>3</sup> of ethanoic acid and 1 cm<sup>3</sup> of ethanol into a boiling tube, then add a few drops of concentrated sulfuric acid.

• Warm at about 80 °C.

• Allow the contents of the tube to cool. When cool, pour the mixture into a beaker containing sodium carbonate solution.

• A layer of ester will separate and float on top of the water.

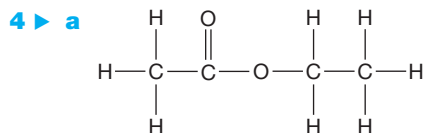
• Repeat with other combinations of acid and alcohol.

**c** Esters can be detected by smelling the product: gently waft the odour towards nose. For example, propyl ethanoate smells like pear.

**d** Food flavourings and perfumes.

<b>3 ► a</b>	<b>Alcohol</b>	<b>Carboxylic acid</b>
	Methanol	Propanoic acid
	$\text{CH}_3\text{OH}$	$\text{CH}_3\text{CH}_2\text{COOH}$
<b>b</b>	Propan-1-ol	Methanoic acid
	$\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$	$\text{HCOOH}$

- c** Butan-1-ol      Ethanoic acid  
 $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{OH}$        $\text{CH}_3\text{COOH}$



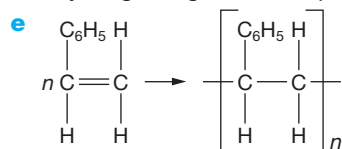
- b** Ethanoic acid.  
**c** By heating with potassium dichromate(VI) in dilute sulfuric acid.  
**d**  $\text{CH}_3\text{COOH(l)} + \text{CH}_3\text{CH}_2\text{OH(l)} \rightleftharpoons \text{CH}_3\text{COOCH}_2\text{CH}_3\text{(l)} + \text{H}_2\text{O(l)}$   
**e** Heat at about  $80^\circ\text{C}$  with a few drops of concentrated sulfuric acid as a catalyst.

- 5 ▶** Test 1: Take out  $3\text{cm}^3$  of each of the three samples and put them into three separate test-tubes. Add a few pieces of Mg or a spatula tip worth of  $\text{Na}_2\text{CO}_3$  solid into each of the test-tubes. The one that bubbles contains ethanoic acid.

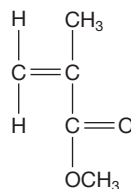
Test 2: To the remaining two samples, take out another  $3\text{cm}^3$  of each and put them into two separate boiling tubes. Add a few drops of potassium dichromate(VI) in dilute sulfuric acid. Heat the boiling tubes in a warm water bath. The one that changes colour from orange to green contains ethanol.

## CHAPTER 29

- 1 ▶ a** Unsaturated: containing one or more carbon carbon double or triple bond.  
 Hydrocarbon: compounds containing carbon and hydrogen only.  
**b**  $\text{C}_2\text{H}_4 + 3\text{O}_2 \rightarrow 2\text{CO}_2 + 2\text{H}_2\text{O}$   
**c** Ethene can be used to make addition polymer polyethene, which is used for making plastic bags, milk bottles etc.
- 2 ▶ a** Joining up of lots of little molecules (the monomers) to make one big one (a polymer).
- b**
- 
- c** (The "continuation" bonds at each end are an important part of this structure. Marks will be lost in an exam if they are omitted.)
- 
- d** Joining two or more molecules together without anything being lost in the process.

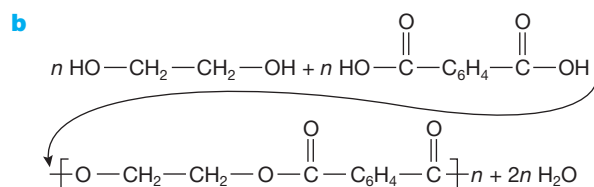
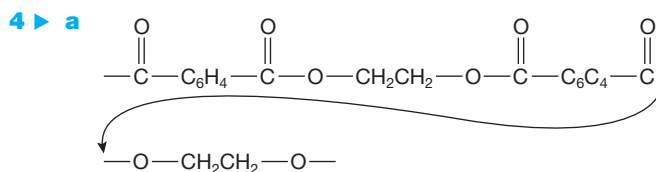
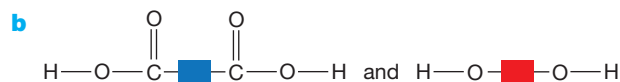


- f** Drawing the molecule to show its relationship with the structure of the polymer as drawn in the question:

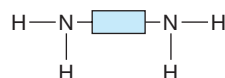


- g** Either: buried in landfill sites because there would be no emission of poisonous carbon monoxide or no carbon dioxide to contribute to global warming or no sulfur dioxide to contribute to acid rain. Or: incineration because no space is filled up for landfills or energy can be generated to provide heat for office buildings.

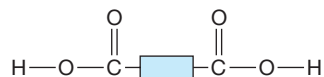
- 3 ▶ a** Condensation polymer



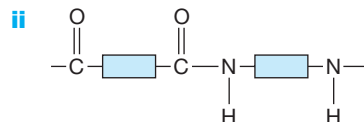
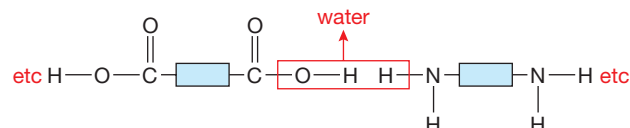
- 5 ▶ a** Drawing 1,6-diaminohexane as:



and hexanedioic acid as:




- i** In a condensation reaction, when two molecules join together a small molecule is lost in the process. When the two molecules above join together, a molecule of water is lost every time they come together. In a condensation polymerisation reaction, this happens repeatedly.

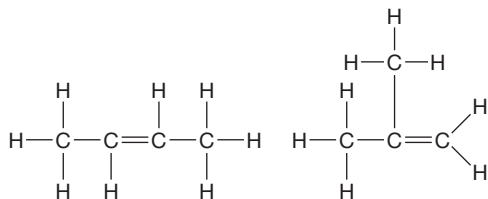




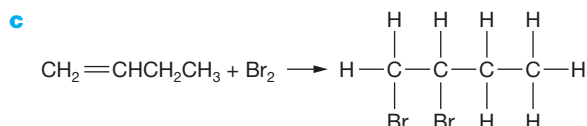
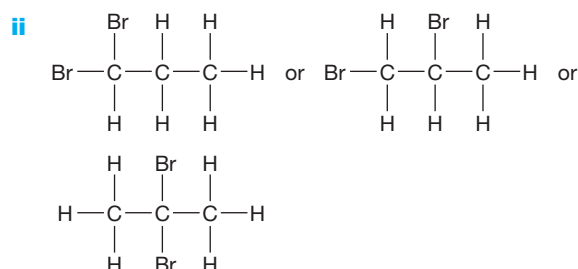
- b i** There will be four more  $\text{CH}_2$  groups in the “box” in the part of the diagram which comes from the dioic acid.
- ii** The bonding between the two monomers will be identical, as will the size of the diamine.

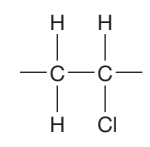
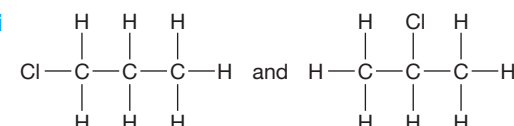
## END OF UNIT 4 QUESTIONS

- 1 ► a** D (1)
- b** Used as liquefied petroleum gas / domestic heating or cooking. (1)
- c** Gasoline. (1)
- d** Crude oil is separate into fractions by fractional distillation. (1)  
It is heated until it boils and the vapour enters a fractionating column (1), which is cooler at the top and hotter at the bottom (1). The fractions with a lower boiling point condense nearer to the top of the column and can be tapped off / the fractions with a higher boiling point condense nearer to the bottom of the column and can be tapped off (1).
- e** As the number of carbon atoms increase, the boiling point of a hydrocarbon increases. (1)  
This is because the intermolecular forces increases so it takes more energy to break them during boiling. (1)
- f**  $2\text{C}_{15}\text{H}_{32} + 31\text{O}_2 \rightarrow 30\text{CO} + 32\text{H}_2\text{O}$   
correct formulae of products (1)  
correct balancing of the equation (1)
- g i**  (1)
- ii** C
- h i** 600–700 °C (1); silica / alumina as a catalyst, accept ‘aluminium oxide / silicon dioxide / aluminosilicate / zeolite or the correct formulae. (1)
- ii** To make more of the smaller-chain fractions, for example, gasoline to use as petrol. (1)  
To produce more alkenes that can be used for making polymers (plastics). (1)
- iii**  $\text{C}_9\text{H}_{20}$  (1)
- iv** The compound contains one or more carbon – carbon double or triple bond. (1)
- v**



- 2 ► a** E contains an element that is not carbon or hydrogen. (1)
- b i** 1,3-dibromopropane 1 mark for dibromopropane



- d i** B and D
- ii**  $\text{C}_n\text{H}_{2n+2}$
- iii** Same functional group / same or similar chemical properties. (1)  
Shows a trend in physical properties. (1)  
Or Each member differs from the next by a  $-\text{CH}_2-$  unit. (1)
- e i** Poly(chloroethene) or polyvinylchloride (1)
- ii**
- 
- iii** They are inert. (1)  
Or non-biodegradable. (1)  
Or if they are burned, they release toxic gases including carbon monoxide or hydrogen chloride. (1)
- iv** No colour change / bromine water remains orange. (1)  
Polymer of compound A is saturated / contains no  $\text{C}=\text{C}$  bond. (1)
- f i** UV light (1)
- ii** Substitution. (1)
- iii**
- 

## 3 ► a

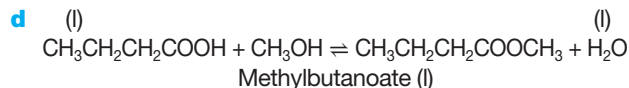
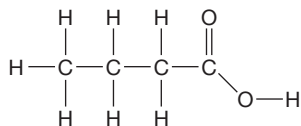
	C	H	O
Mass (in 100 g)	55	9	36
$A_r$	/12	/1	/16
Number of moles	4.58	9	2.25
/smallest number of moles	/2.25	/2.25	/2.25
Ratio	2	4	1
The empirical formula is $\text{C}_2\text{H}_4\text{O}$ .			

**b**  $2 \times 12 + 4 \times 1 + 1 \times 16 = 44$

$$\frac{88}{44} = 2$$

The molecular formula is  $C_4H_8O_2$ .

**c** butanoic acid (1)



**4 ▶ a** Test 'Add bromine water'

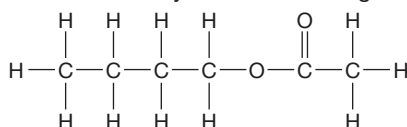
Result with compound 1 Solution changes colour from orange to colourless

Result with compound 2 Solution stays orange

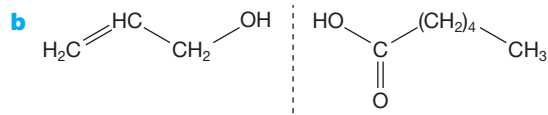
**b** 1 mark for the correct structure; 1 mark for the continuation bonds

**c** Poly(but-2-ene) (1)

**5 ▶ a i** 1 mark for the ester linkage; 1 mark for the correct structure of butyl and ethanoate group



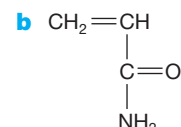
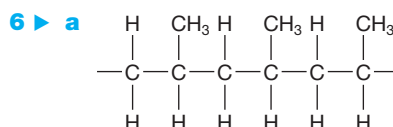
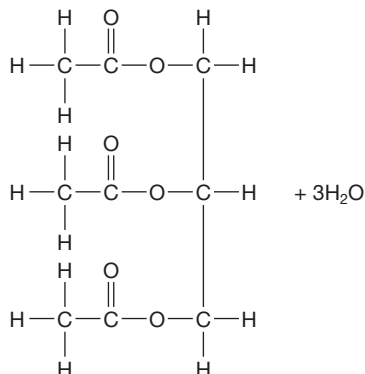
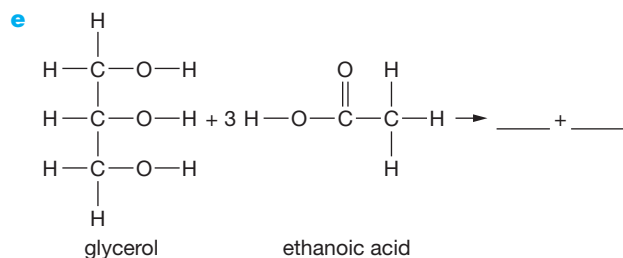
**ii** Potassium dichromate. (VI); Dilute sulfuric acid. (1); Heat. (1)



**c i** The reaction is reversible. (1)

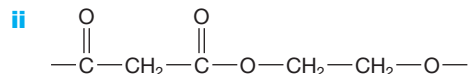
**ii** Concentrate sulfuric acid. (1)

**d** No effect on the position of equilibrium. (1); Catalyst increase the rate of the forward reaction and the rate of the reverse reaction equally. (1)



**c i** In condensation polymerisation, two different types of monomers are used (1)

A small molecule such as water or hydrogen chloride is produced in condensation polymerisation as well as the polymer (1)



**d i** A polyester that is biodegradable (1)

**ii** C

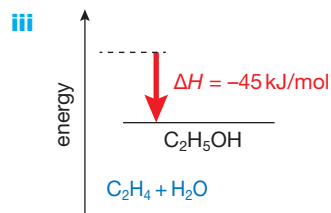
**7 ▶ a i** Phosphoric acid (1); 300 °C or 60–70 atm pressure. (1)

**ii** Bond breaking  $4 \times \text{C}-\text{H} + 1 \times \text{C}=\text{C} + 2 \times \text{O}-\text{H} = 4 \times 412 + 612 + 2 \times 463 = 3186 \text{ kJ/mol}$ . (1)

Bond forming  $5 \times \text{C}-\text{H} + 1 \times \text{C}-\text{C} + 1 \times \text{C}-\text{O} + 1 \times \text{O}-\text{H} =$

$5 \times 412 + 1 \times 348 + 1 \times 360 + 1 \times 463 = 3231 \text{ kJ/mol}$ . (1)

Overall  $3186 - 3231 = -45 \text{ kJ/mol}$ . (1)



**b i** Fermentation (1)

**ii** To provide the enzyme (zymase) which speeds up the reaction. (1)

**iii** In the presence of oxygen, yeast would produce carbon dioxide and water rather than ethanol. (1)

**iv** Warm climate or plenty of land to grow sugar canes (1); Lack of oil resources to produce ethane. (1)

**c i** Dilute sulfuric acid (1); Heating. (1)

**ii**  $\text{CH}_3\text{CH}_2\text{OH} + 2[\text{O}] \rightarrow \text{CH}_3\text{COOH} + \text{H}_2\text{O}$

**iii** Orange (1) to green. (1)

**iv**  $\text{Mg(s)} + 2\text{CH}_3\text{COOH(aq)} \rightarrow (\text{CH}_3\text{COO})_2\text{Mg(aq)} + \text{H}_2\text{(g)}$

1 mark for the correct formulae of products; 1 mark for correct balancing.

**d i** Water is formed in the reaction. (1)

**ii** Pour the reaction mixture into water (or sodium carbonate solution) (1); The alcohol and the acid will dissolve but the ester will float on the surface of water. (1)