The answers to the practice questions below are as given to the IB examiners. The following notes may help you to interpret these and make full use of the guidance given.

- There are no half marks awarded. Each mark is shown by the number in the square brackets [1].
- Points worth single marks are separated from each other by a semicolon (;).
- Alternative possible answers are separated from each other by a dash (/).
- Any answer given in **bold** or underlined *must* be present to score the mark.
- Information in brackets ( ) is not needed to score the mark. You may also find «» brackets to indicate that the information inside the brackets is not needed to score the mark.
- Notes given in italics are to guide the examiner on what to accept/reject in their marking.
- *OWTTE* means 'or words to that effect', so alternative wording that conveys the same meaning can be equally rewarded.

You may notice occasional differences between the calculations or wordings given in the answers and those in the worked solutions. This is because the answers give the final solution with the minimum of working, and the worked solutions provide the extra reasoning and working needed to understand how the answers are attained.

S1.1 Introduction to the particulate nature of matter

Exercises

Q1 C

Q2 B

Q3 C

Q4 **(a)** Fertilizers are soluble.

(b) Large energy requirement and so is expensive.

Q5 The solvent evaporates. The oil has changed chemically by reacting with oxygen in the air.

Q6 Both coins are made from a homogeneous mixture with the same composition: Cu 89%, Al 5%, Zn 5%, Sn 1%.

Q7 X has moved more quickly, therefore the particles of X have smaller mass.

Q8 The average kinetic energy increases with temperature. The particles will move faster and the rate of diffusion increases.

Q9 D

Q10 **(a)** Bubbles will be present throughout the volume of the liquid. A brown gas will be visible above the brown liquid.

(b) As the two states are at the same temperature, the particles have the same average kinetic energy and are moving at the same average speeds. The inter-particle distances in the gas are significantly larger than those in the liquid.

Q11 D

Q12 C

Q13 B

Q14 **(a)** Ice, as it has a melting point of 0 °C.

(b) They are vibrating with increasing frequency.

(c) They are moving freely in a fixed volume as it is the liquid state.

(d) The heat energy is used to separate the particles and not increase their movement.

(e) 52.5 °C

(f) Temperature predicted = 127.5 °C. Water boils at 100 °C, which would be reached after around 980 s, and so the temperature does not follow the predicted pattern. The temperature is likely to be 100 °C.

Q15 D

Q16 B

Q17 C

Q18 D

Q19 Steam will condense on the skin, releasing energy as it forms liquid at the same temperature (portion e–d on S1.1 Figure 5). This is additional to the energy released when both the boiling water and the condensed steam cool on the surface of the skin.

Challenge yourself

1 The molecules of ethanol can partly fit between the spaces between the water molecules, and vice versa.

2 A mixture may appear to be homogeneous with the naked eye but on closer inspection with a microscope may be heterogeneous. The classification depends on the scale on which the components are analysed.

3 In 1827 Robert Brown dropped pollen grains into water and examined them under a microscope. The pollen grains moved around erratically in the water. This so-called 'Brownian motion' was explained in 1905 by Albert Einstein. He realized that the pollen was being jostled by something even smaller than the pollen grains: water molecules. Einstein didn't just base this theory on his observations – he used complex mathematics to show that a particle model could explain Brownian motion.

4 In cold climates, temperature may approach or go below the boiling point of butane so it stays liquid even when released from the pressure it is under when stored in its canister. This makes it ineffective as a fuel.

5 $\Delta E = mc\Delta T$

gradient of graph $= \dfrac{\Delta T}{\Delta E} = \dfrac{1}{mc}$

The gradient is determined by the mass and the specific heat capacity of the substance as solid, liquid and gas.

Practice questions

1 D

2 A

3 B

4 B

5 D

6 A

7 C

8 B

9 B

10 B

11 A

12 **(a)** Smoke: heterogeneous – the solid particles are suspended in the air. [1]

(b) Sugar and water: homogeneous – the sugar dissolves in the water. (If it is a saturated solution with excess sugar that cannot dissolve, the overall mixture is then heterogeneous.) [1]

(c) Steel: homogeneous – it is an alloy of iron and carbon (+ other components) with the same properties throughout. [1]

13 (At certain conditions of low temperature and low humidity), snow changes directly to water vapour by sublimation, without going through the liquid phase. [1]

14 **(a)** Any two from:

forces between gas particles are smaller than between liquid particles

speed of gas particles greater than speed of liquid particles

movement of gas particles is less restricted than that for liquid particles

separation of particles (generally) much greater in a gas than in a liquid [2]

(b) the gas particles do not have the same speed [1]

the speed of individual particles changes with each collision [1]

(c) **(i)** increased kinetic energy of the particles [1]

(ii) heat energy used to separate the particles [1]

no increase in kinetic energy (as indicated by temperature) [1]

15 **(a)** Any value between −15 °C and +10 °C. (The experimental value = −7.2 °C.) [1]

(b) boiling points increase [1]

(c) gas [1]

(d) **(i)** liquid [1]

the molecules are close together but have an irregular arrangement [1]

(ii) increased movement of molecules [1]

with increased separation between particles [1]

(Note the molecules do not break into individual atoms.)

(iii) The bromine molecule break into atoms or ions, which then combine with atoms or ions of other elements. [1]

(iv) Temperature is a measure of the average kinetic energy: increase in temperature results in increase in movement / rate of diffusion. [1]

S1.2 The nuclear atom

Exercises

Q1 Tellurium has a greater proportion of heavier isotopes (those with more neutrons).

Q2

	Species	**No. of protons**	**No. of neutrons**	**No. of electrons**
(a)	${}^{7}Li$	3	4	3
(b)	${}^{1}H$	1	0	1
(c)	${}^{14}C$	6	8	6
(d)	${}^{19}F^{-}$	9	10	10
(e)	${}^{56}Fe^{3+}$	26	30	23

Q3

	Species	**No. of protons**	**No. of neutrons**	**No. of electrons**
	${}^{40}Ca^{2+}$	20	20	18
(a)	${}^{40}_{18}Ar$	18	22	18
(b)	${}^{39}_{19}K^{+}$	19	20	18
(c)	${}^{35}_{17}Cl^{-}$	17	18	18

Q4 D

Q5 Examples are density (related to mass) and, as chlorine is a gas, rate of diffusion.

Q6 C

Q7 C

Q8 B

Q9 A

Q10 C

Q11 No individual atom with this mass exists. Silicon has a number of isotopes and this is an average value. The claim is true.

Q12 D

Q13 B

Q14

$$\text{RAM} = \frac{(50 \times 4.31) + (52 \times 83.76) + (53 \times 9.55) + (54 \times 2.38)}{100}$$

$$= 52.0569 \approx 52.06$$

Q15 $\text{RAM} = \frac{(20 \times x) + (22 \times y)}{100} = 20.18$ and $x + y = 100$;

$y = 100 - x$

$(20 \times x) + (22 \times 100) - (22 \times x) = 2018$

$2x = 2200 - 2018 = 182$

$x = 91$; abundance ${}^{20}Ne = 91\%$

Q16 Let the abundance of ${}^{25}Mg$ be x. Consider 100 atoms.

$$24.31 = \frac{(78.90 \times 24) + (x \times 25) + ([100 - 78.90 - x] \times 26)}{100}$$

$2431 = 1893.6 + 25x + 548.6 - 26x$

$2431 = 2442.2 - x$

$x = 11.20$

${}^{25}Mg$ is 11.20% and ${}^{26}Mg$ is 9.90%

Challenge yourself

1 Potash, soda, magnesia and barytes are compounds of group 1 and 2 elements. These compounds were later broken down into their component elements by electrolysis. (See Reactivity R3.2.)

2 High energy alpha particles can get so close to the nucleus that the strong nuclear force, not included in the Rutherford model, begins to act on them.

3 ${}^{0}_{+1}e + {}^{0}_{-1}e \rightarrow 2\gamma$

Note the left-hand subscripts which represent the charges are balanced.

The negligible masses of the position and electron are converted into energy in the form of two photons.

4 The nucleons are attracted to each other by the short-range strong nuclear force, which is stronger than the electrostatic repulsive forces.

5 $V_{\text{nucleus}} = \frac{4\pi}{3} \times R^3 = \frac{4\pi}{3} \times (1.2 \times 10^{-15} \times A^{1/3})^3$

$= 7.24 \times 10^{-45} \times A\ \text{m}^3$

the mass of the nucleus is A u, i.e. $A \times 1.67 \times 10^{-27}$ kg

$$\text{density } (\rho) = \frac{\text{mass}}{\text{volume}} = \frac{A \times 1.67 \times 10^{-27}\ \text{kg}}{7.24 \times 10^{-45} \times A\ \text{m}^3}$$

$$\approx 2.3 \times 10^{17}\ \text{kg m}^{-3}$$

The density is independent of mass number and so all nuclei have the same density.

This is to be expected as all nuclei are composed of protons and neutrons packed closely together and so have similar densities.

The density of the nucleus is extremely high compared to the bulk density of an element. For example, the density of mercury is 1.35×10^4 kg m^{-3} (so the nucleus is denser by 13 orders of magnitude).

6 mass of H atom = mass of proton (H^+) plus mass of electron

$= 1.672622 \times 10^{-27}$ kg + 9.109383×10^{-31} kg

$= 1.67353 \times 10^{-27}$ kg

The mass of a hydrogen atom is the same as the mass of a hydrogen ion to 3 significant figures (1.67) and the 4th significant figure only differs by 0.001. The mass of the atom is 1.673 and the ion 1.674.

7 The difference in mass between the molecules is very small. The molecules can only be separated over large distances. A large area of ground is needed.

8 Technetium, Tc, has no stable isotopes. Note that its relative mass is an integer and given in parentheses.

Practice questions

1 A

2 C

3 B

4 D

5 no. of protons = 25 [1]

no. of electrons = 25 − 2 = 23 [1]

no. of neutrons = 55 − 25 = 30 [1]

6 $\text{RAM} = \frac{(54 \times 5.95) + (56 \times 91.88) + (57 \times 2.17)}{100}$ [1]

$= 55.90$ [1]

7 **(a)** Cobalt has a greater proportion of heavier isotopes (*OWTTE*) *or* cobalt has greater number of neutrons. [1]

(b) 27 protons and 25 electrons [1]

8 **(a)** different number of neutrons [1]

(b) Same chemical properties because they have the same number of electrons; it is electrons that are responsible for chemical properties. [1]

(c) $\text{RAM} = \frac{(28 \times 92.18) + (29 \times 4.70) + (30 \times 3.12)}{100}$ [1]

$= 28.11$ [1]

S1.3 Electron configurations

Exercises

Q1 C

Q2 D

Q3 **(a)** C

(b) B

Q4 The transition corresponds to a larger energy change in the atom. The emitted photon has more energy and corresponds to a lower wavelength.

Q5 4s < 4p < 4d < 4f

Q6

Sublevel	4s	4p	4d	4f
No. of orbitals	1	3	5	7

Q7 $1s^22s^22p^63s^23p^64s^2$

Q8 $1s^22s^22p_x^22p_y^22p_z^23s^23p_x^13p_y^13p_z^1$, so three unpaired electrons

Q9 Number of orbitals = $n^2 = 4^2 = 16$

$= (1 \times s) + (3 \times p) + (5 \times d) + (7 \times f) = 16$

Q10 C

Q11 C

Q12 **(a)** V is $1s^22s^22p^63s^23p^63d^34s^2$

(b) K is $1s^22s^22p^63s^23p^64s^1$

(c) Se is $1s^22s^22p^63s^23p^63d^{10}4s^24p^4$

(d) Sr is $1s^22s^22p^63s^23p^63d^{10}4s^24p^65s^2$

Q13 D

Q14 B

Q15 B

Q16 D

Q17 **(a)** O^{2-} is $1s^22s^22p^6$

(b) Cl^- is $1s^22s^22p^63s^23p^6$

(c) Ti^{3+} is $1s^22s^22p^63s^23p^63d^1$

(d) Cu^{2+} is $1s^22s^22p^63s^23p^63d^9$

Q18

		3d					4s
(a)	Ti^{2+}	↑	↑				
(b)	Fe^{2+}	↑↓	↑	↑	↑	↑	
(c)	Ni^{2+}	↑↓	↑↓	↑↓	↑	↑	
(d)	Zn^{2+}	↑↓	↑↓	↑↓	↑↓	↑↓	

Q19 **(a)** Ne is $1s^22s^22p^6$

(b) Negatively charged ions would be F^-, O^{2-} or N^{3-}; positively charged ions would be Na^+, Mg^{2+} or Al^{3+}.

Q20 **(a)** [Ar] $3d^8$

(b) [Xe] $5d^{10}6s^2$

(c) [Ne] $3s^23p^6$

(d) [Ne]

Q21 **(a)** Cl is $1s^22s^22p^63s^23p^5$

(b) Nb is $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^35s^2$

(c) Ge is $1s^22s^22p^63s^23p^63d^{10}4s^24p^2$

(d) Sb is $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}5s^25p^3$

Q22 **(a)** Si

(b) Mn

(c) Sr

(d) Sc

Q23 11

Q24 20

Q25 [Kr] $4d^{10}$

Q26 $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}4f^{14}5s^25p^65d^{10}6s^26p^65f^4$

Challenge yourself

1 D

$$E_{\text{photon}} = \frac{hc}{\lambda} = \frac{6.63 \times 10^{-34} \times 3.0 \times 10^{8}}{588 \times 10^{-9}} = 3.38 \times 10^{-19}\ \text{J}$$

$\Delta E_{\text{atom}} = 3.38 \times 10^{-19}$ J

This corresponds to IV → II:
$-2.42 \times 10^{-19} - (-5.80 \times 10^{-19}) = 3.38 \times 10^{-19}$ J

2 The Schrödinger model:

does not have well defined orbits for the electrons

does not treat the electron as a localized particle but gives a probability wave description

predicts the relative intensities of various spectral lines.

3 A

4 $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}5s^25p^64f^{14}5d^{10}6s^26p^66d^27s^2$
or [Rn] $6d^27s^2$

5 **(a)** [Rn] $7s^25f^{14}6d^7$

(b) The first g block element would be
[Rn] $7s^25f^{14}6d^{10}7p^68s^28g^1$

$Z = 86 + 2 + 14 + 10 + 6 + 2 + 1 = 121$

6 **(a)** There would be two types of p orbital and two types of d orbitals.

(b) 4 groups in the p and d block.

Practice questions

1 D

2 A

3 B

4 B

5 C

6 D

7 **(a)** the electron configuration of argon, which is
$1s^22s^22p^63s^23p^6$ [1]

(b) $x = 1$ and $y = 5$ [1]

(c)

↑		↑	↑	↑	↑	↑
4s				3d		

[1]

Accept all six arrows pointing down rather than up.

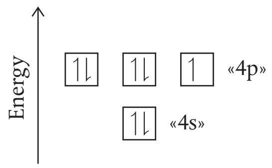
8 **(a)** IV < I < II < III *or* ultraviolet radiation < yellow light < red light < infrared radiation [1]

(b) A continuous spectrum has all colours / wavelengths / frequencies whereas a line spectrum has only lines of sharp / discrete / specific colours / wavelengths / frequencies. [1]

9 **(a)** $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^5$

Accept 3d before 4s.

(b)



Accept double-headed arrows.

S1.4 Counting particles by mass: The mole

Exercises

Q1 Note: these calculations have used $L = 6 \times 10^{23}$

(a) 7.2×10^{22}

(b) 3×10^{24}

(c) 1.2×10^{23}

Q2 0.53 mol H

Q3 0.250 mol

Q4 As the value for A_r is closer to 35 than 37, the ^{35}Cl must be the most abundant isotope. A 50 : 50 mixture would have A_r of 36.

Q5 **(a)** 262.87

(b) 176.14

(c) 164.10

(d) 248.22

Q6 189.1 g

Q7 1.5 mol

Q8 0.0074 mol Cl^-

Q9 1.83×10^{24} C atoms

Q10 171 g (integer values used)

Q11 10.0 g H_2O

Q12 2.0 mol N_2 > 3.0 mol NH_3 > 25.0 mol H_2 > 1.0 mol N_2H_4

Q13 **(a)** CH

(b) CH_2O

(c) $C_{12}H_{22}O_{11}$

(d) C_4H_9

(e) C_4H_7

(f) CH_2O

Q14 $Na_2S_2O_3$

Q15 $CoSO_4.7H_2O$

Q16 $C_{17}H_{25}N$

Q17 NH_3

Q18 6.94 Li

Q19 Atomic masses: S = 32, Se = 79, Te = 128, so the highest % by mass of Cd is in CdS.

Q20 empirical formula CH; molecular formula C_6H_6

Q21 empirical formula H_2PO_3; molecular formula $H_4P_2O_6$

Q22 $C_{10}H_{16}N_5P_3O_{13}$ for both empirical and molecular formulas

Q23 C_3H_8O

Q24 Let change in mass on using the chalk = y g $CaCO_3$

$$\text{mass C in chalk} = \frac{12}{M(CaCO_3)} \times y$$

$$\text{number of C atoms} = 6.02 \times 10^{23} \times \frac{12}{M(CaCO_3)} \times y$$

Q25 2.81 g

Q26 4.93 g

Q27 0.0100 mol

Q28 0.40 mol dm^{-3}

Q29 3.1 cm^3

Q30 0.106 mol dm^{-3} Na_2SO_4 and 0.115 mol dm^{-3} $Pb(NO_3)_2$; assume no side reactions, all $PbSO_4$ precipitates

Q31 Ar

Q32 B

Q33 45 dm^3

Q34 10 dm^3, assuming the H atoms are present as a gas

Challenge yourself

1 Group 1: H has one electron in outer energy level, forms H^+ ion.

Group 17: H needs one electron to fill its outer energy level, can form H^- ion, is a gaseous element (H_2) not a metal.

Not in a group: H is anomalous, and is not a group 1 metal or a group 17 halogen.

2 $FeCl_3.6H_2O$, $CuSO_4.5H_2O$, $Co(NO_3)_2.6H_2O$

3 % N in fertilizer = 18%

P_2O_5 = 44% P by mass, so %P in fertilizer = 51% × 44% = 22%

K_2O = 83% K by mass, so %K in fertilizer = 20% × 83% = 17%

4 As NaOH dissolves, the separated ions Na^+ and OH^- ions become hydrated, i.e. surrounded by H_2O molecules. This involves breaking the hydrogen bonds between the H_2O molecules in pure water and allows closer packing, which reduces the volume.

5 The relationship between concentration and absorbance becomes non-linear at higher concentrations, so interpretation of data to deduce the unknown concentration will be less accurate in this region. It is therefore important that the absorbance of the unknown sample falls within the lower linear part of the graph where it will give a more accurate reading of the unknown concentration.

6 The key reaction in the expansion of the bag is the release of nitrogen gas from the decomposition of sodium azide:

$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$

The reactions to remove the dangerously reactive sodium metal can take several forms, resulting in silicates such as K_2SiO_3 and Na_2SiO_3. For example:

$10Na(s) + 2KNO_3(s) \rightarrow K_2O(s) + 5Na_2O(s) + N_2(g)$

$K_2O(s) + Na_2O(s) + 2SiO_2(s) \rightarrow K_2SiO_3(s) + Na_2SiO_3(s)$

S1.4 Practice questions

1 D

2 B

3 A

4 B

5 D

6 D

7 C

8 D

9 C

10 D

11 A

12 **(a)** 0.115 mol H_2O [1]

(b) 0.0574 mol K_2CO_3 [1]

(c) $K_2CO_3.2H_2O$ [1]

(d) Heat to constant mass; when further heating does not lead to further decrease in mass. [1]

13 % O = 100 − (62.02 + 10.43) = 27.55% [1]

	C	**H**	**O**
% by mass	62.02	10.43	27.55
divide % by mass by A_r	$\frac{62.02}{12.01} = 5.164$	$\frac{10.43}{1.01} = 10.33$	$\frac{27.55}{16.00} = 1.722$
divide by smallest	3.00	6.00	1

[1]

empirical formula = C_3H_6O [1]

14 **(a)** % O = 100 − 43.6 = 56.4% [1]

	P	**O**
% by mass	43.6%	56.4%
divide % by mass by A_r	$\frac{43.6}{30.97} = 1.4078$	$\frac{56.4}{16.00} = 3.525$
divide by smallest	1	2.50

[1]

empirical formula = P_2O_5 [1]

(b) molecular formula = (empirical formula) × n

$285\ g\ mol^{-1} = [(30.97 \times 2) + (16.00 \times 5)] \times n = 141.94 \times n$

$n = \frac{285}{142} = 2$

molecular formula = P_4O_{10} [1]

15 **(a)** $M_r(NaHCO_3) = 22.99 + 1.01 + 12.01 + (16.00 \times 3)$ $= 84.01\ g\ mol^{-1}$ [1]

$n(NaHCO_3)$ = conc. × volume

$= 0.500\ \text{mol dm}^{-3} \times \dfrac{250.00}{1000.00}\ \text{dm}^3$

$= 0.125\ \text{mol}$ [1]

mass($NaHCO_3$) = 0.125 mol × 84.01 g mol^{-1} = 10.5 g [1]

Measure a mass of 10.5 g $NaHCO_3$ using a balance and transfer to a 250 cm^3 volumetric flask. [1]

Add distilled water to the flask, with frequent shaking, up to the mark. [1]

(b) To prepare dilutions of stock solution:

Using a 10 cm^3 pipette transfer 10.00 cm^3 of the stock solution from (**a**) into a 100 cm^3 volumetric flask. Make up to the mark with distilled water → 0.0500 mol dm^{-3} solution. [2]

Using a clean 10 cm^3 pipette transfer 10.00 cm^3 of the 0.0500 mol dm^{-3} solution into a 100 cm^3 volumetric flask. Make up to the mark with distilled water → 0.00500 mol dm^{-3} solution. [2]

S1.5 Ideal gases

Exercises

Q1 A

Q2 C

Q3 A

Q4 D

Q5 D

Q6 3 and 4

Q7 B

Q8 **(a)** Particles are in constant random motion and collide with each other and with the walls of the container in perfectly elastic collisions. The kinetic energy of the particles increases with temperature. There are no intermolecular forces and the volume of the particles is negligible relative to the volume of the gas.

(b) At low temperature, the particles have lower kinetic energy, which favours the formation of intermolecular forces and reduces gas pressure.

Q9 The assumptions are most valid when the gas particles are widely separated and the intermolecular forces weakest. These occur at conditions of high temperature and low pressure.

Q10 D

Q11 C

Q12 D

Q13 B

Q14 D

Q15 **(a)** It is is a linear relationship.

(b) The equation of the best fit line is:
$P = 0.3338T + 95.625$

When $P = 0$, $0.3338T + 95.625 = 0$

$T = \dfrac{-95.625}{0.3338} = -286.5\ °\text{C}$

(c) The temperature is lower than expected. The literature value is absolute zero = −273 °C

(d) Proportional; a straight line which passes through/close to the origin.

(e) In the ideal gas model, as the temeprature decreases, the particles decrease their movement until at 0 K there is no movement and the particles are attracted together in a minimal volume. With the experimental results, the volume of the gas will not be precisely zero as the molecules have some volume.

Q16 For a given volume, Curve 1 has the higher pressure. It therefore has the higher temperature.

Q17 **(a)** $P_1V_1 = P_2V_2$

$V_2 = \dfrac{P_1V_1}{P_2} = \dfrac{2.02 \times 10^6 \times 0.0400}{1.01 \times 10^5} = 0.800\ \text{m}^3$

volume of balloon is 0.800 − 0.0400 = 0.760 m^3

(b) The atoms will decrease their speed.

Q18 B

Q19 C

Q20 90 kPa

Q21 16 °C

Q22 3.0 dm^3

Q23 2.8 dm^3

Q24 $M = 131\ \text{g mol}^{-1}$ so gas is Xe

Q25 $90.4\ \text{g mol}^{-1}$

Q26 helium (a greater mass in the same volume; density is mass/volume)

Q27 At higher altitude the external air pressure is less. As the air in the tyre expands on heating (due to friction with the road surface), the internal pressure increases.

Q28 NH_3 shows greater deviation than CH_4 due to stronger intermolecular attractions, especially at low temperature.

Q29 **(a)** II as $\frac{PV}{RT} < 1$ at high pressure

The pressure is reduced by intermolecular forces.

(b) III as $\frac{PV}{RT} > 1$ as the pressure increases

The pressure is greater than an ideal gas due to the space occupied by the molecules.

Challenge yourself

1 If the molecular collisions are not perfectly elastic, then the kinetic energy of the molecules would be gradually lost. The molecules would slow down and the pressure would gradually reduce to zero.

2 **(a)** $PV = \text{constant}: P_1V_1 = P_0V_2$

(b) $\frac{P}{T} = \text{constant}: \frac{P_0}{T_1} = \frac{P_2}{T_2}$

(c) From **(b)**, $P_0 = \frac{P_2T_1}{T_2}$

Substitution in **(a)**, $\frac{P_1V_1}{V_2} = \frac{P_2T_1}{T_2}, \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

3 The atmospheric pressure decreases with altitude. When the pressure inside the balloon is larger than the external pressure the volume of the balloon increases.

4 **(a)** $PV = nRT$; 18.0 g = 1 mol; $n = 1$

$$V = \frac{RT}{P}$$

$$= \frac{8.31 \times 500}{1.00 \times 10^5} = 0.04155\ \text{m}^3 = 41\,550\ \text{cm}^3$$

An approximate value for the volume of the particles can be obtained by considering 18.0 g of water, which has a volume of 18 cm^3

% volume occupied by particles

$$= \frac{18}{41550} \times 100 = 0.04\%$$

(b) $V = \frac{RT}{P}$

$$= \frac{8.31 \times 400}{5.00 \times 10^7} = 6.65 \times 10^{-5}\ \text{m}^3 = 66.5\ \text{cm}^3$$

% volume occupied by particles = 27%

In Challenges 5 and 6, we are considering 1 mol of gas, so $n = 1$.

5 $\left(P + \frac{a}{V^2}\right)(V - b) = RT$

At high pressure, $\frac{a}{V^2}$ can be neglected.

$PV - Pb = RT$

$\frac{PV}{RT} = 1 + \frac{Pb}{RT}$

Note the deviation, which depends on b, increases as the pressure is increased and the temperature is decreased, in agreement with S1.5 Figure 12.

6 $\left(P + \frac{a}{V^2}\right)(V - b) = RT$

At low pressure, the parameter b can be neglected.

$\left(P + \frac{a}{V^2}\right)V = RT$

$PV + \frac{a}{V} = RT$

$\frac{PV}{RT} = 1 - \frac{a}{VRT}$

Note the deviation, which depends on a, increases as the temperature is decreased, in agreement with S1.5 Figure 12.

Practice questions

1	C
2	C
3	D
4	A
5	B
6	C
7	C
8	A
9	C
10	D

11 **(a)** two from:

weak intermolecular / dispersion forces

small atomic volume

smaller number of electrons [2]

(b) $n = \frac{pV}{RT}$

$= \frac{2.0 \times 10^{6} \times 3.0 \times 10^{-2}}{8.31 \times 300}$ [1]

$= 24.1$ (mol) [1]

(c) $N = 24.1 \times 6.02 \times 10_{23} = 1.45 \times 10^{25}$ [1]

volume of 1 atom $= \frac{3.00 \times 10^{-2}\ \mathrm{m}^3}{1.45 \times 10^{25}}$

$= 2.07 \times 10^{-27}\ \mathrm{m}^3$ [1]

Accept value from 1×10^{-27} to 6×10^{-27} m

(d) assuming the atoms each occupy a cube;
distance $= \sqrt[3]{(2.07 \times 10^{-27})} = 1.27 \times 10^{-9}$ m

assuming the atoms each occupy a sphere;
distance $= 7.9 \times 10^{-10}$ m [1]

12 **(a)** temperature: 4
mass: 3
pressure: 3 [1]

(b) 0.0650 kg = 65.0 g

$n = \frac{65.0}{65.02} = 1.00$ (mol) [1]

No penalty for using whole number atomic masses.

(c) $n(N_2) = \frac{3}{2} \times 1.00 = 1.50$ (mol)

$T = 25.00 + 273.15 = 298.15$ K *or* $25.00 + 273 = 298$ K

$P = 1.08 \times 1.01 \times 10^5$ Pa *or* $1.08 \times 1.01 \times 10^2$ kPa *or* 1.09×10^5 Pa *or* 1.09×10^2 kPa

$R = 8.31\ \mathrm{J\,K^{-1}\,mol^{-1}}$ (from the data booklet)

Use $PV = nRT$ (from the data booklet)

$V = \frac{nRT}{P} = \frac{1.50 \times 8.31 \times 298.15}{1.08 \times 1.01 \times 10^5} = 0.034.1\ \mathrm{m}^3$

$= 34.1\ \mathrm{dm}^3$ [4]

Award [4] for correct final answer.

Award (3 max) for 0.0341 dm³ or 22.7 dm³.

Award (3 max) for 34.4 dm³.

Award (2 max) for 22.9 dm³.

Award (2 max) for 0.0227 dm³.

*Award **(2 max)** for 0.034 dm³.*

13 **(a)** Gas behaves ideally; as $p \propto T$ (in K) «at constant V» [1]

$P = 0$ gives $0.400T = -109.2$

$T = \frac{-109.2}{0.400} = -273\ °\mathrm{C}$ [1]

(b) From a point on the graph, find the temperature in K and the pressure, e.g. $T = 0\ °\mathrm{C} = 273$ K, $P = 109.2$ kPa

$PV = nRT$

$n = \frac{PV}{RT} = \frac{109.2 \times 10^3 \times 1.0 \times 10^{-3}}{8.31 \times 273} = 0.048135$ mol [1]

$M = \frac{m}{n} = \frac{0.193}{0.048135} = 4.01\ \mathrm{g\,mol^{-1}}$

The gas is helium. [1]

14 molar volume at STP = 22.4 dm³ mol⁻¹ → $22.4\ \mathrm{dm^3\,mol^{-1}}$

density $= 1.94\ \mathrm{g\,dm^{-3}}$

molar mass = molar volume × density

$= 22.4 \times 1.94 = 43.456\ \mathrm{g\,mol^{-1}}$ [1]

or

$PV = nRT$; $n = \frac{PV}{RT} = \frac{1 \times 10^5 \times 1 \times 10^{-3}}{8.31 \times 273} = 0.044$ mol,

which have mass of 1.94 g

1 mol has mass of 44.01 g [1]

the alkane is C_3H_8 [1]

15 **(a)** The reaction stops after 25 m³ of O_2 is added. [1]

(b) $n(\mathrm{NO}) = \frac{PV}{RT} = \frac{1 \times 10^5 \times 50 \times 10^{-6}}{8.31 \times 273} = 0.0022$ mol [1]

$n(O_2) = \frac{PV}{RT} = \frac{1 \times 10^5 \times 25 \times 10^{-6}}{8.31 \times 273} = 0.0011$ mol [1]

(c) volume = 75 cm³ − 25 cm³ = 50 cm³ [1]
$n(\mathrm{X}) = 0.0022$

(d) $2NO(g) + O_2(g) \rightarrow 2X(g)$ therefore $X = NO_2$ [1]
$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$ [1]

S2.1 The ionic model

Exercises

Q1 D

Q2 D

Q3 C

Q4 A

Q5
- **(a)** KBr
- **(b)** ZnO
- **(c)** Na_2SO_4
- **(d)** $CuBr_2$
- **(e)** $Cr_2(SO_4)_3$
- **(f)** AlH_3

Q6
- **(a)** tin(II) phosphate
- **(b)** titanium(IV) sulfate
- **(c)** manganese(II) hydrogencarbonate
- **(d)** barium sulfate
- **(e)** mercury(I) sulfide
- **(f)** vanadium(III) oxide
- **(g)** chromium(III) oxide
- **(h)** lead (IV) oxide

Q7 X_3Y_2

Q8 Mg: Z = 12: electron configuration [Ne] $3s^2$

Br: Z = 35: electron configuration [Ar] $3d^{10}4s^24p^5$

The magnesium atom loses its two electrons from the 3s orbital to form Mg^{2+}, with electron configuration [Ne].

Two bromine atoms each gain one electron into their 4p sub-shell to form Br^-, with electron configuration [Kr].

The ions attract each other by electrostatic forces and form a lattice with the formula $MgBr_2$.

Q9 Ionic bonding between NH_4^+ ions and NO_3^- ions. Covalent bonding within NH_4^+ ions and NO_3^- ions.

Q10 B

Q11 B

Q12
- **(a)** D
- **(b)** A
- **(c)** C

Q13 Test the melting point: ionic solids have high melting points.

Test the solubility: ionic compounds usually dissolve in water but not in hexane.

Test the conductivity: ionic compounds in aqueous solution are good conductors.

Challenge yourself

1 Our calculation does not include the shielding effect of electrons in a lower sublevel. The two 3s electrons in aluminium, for example, partially shield the 3p electron so it (the 3p electron) experiences a smaller effective nuclear charge then the +3 calculated using the simple model. The reduced effective nuclear charge makes it easier to remove the 3p electron from aluminium than to remove one of the 3s electrons in magnesium.

2 For example, the production of chlorine during the electrolysis of sodium chloride provides indirect evidence for the existence of ions. The negative chloride ion is attracted to the positive electrode and is discharged with the production of gaseous chlorine.

3 The ionic radius of Mg^{2+} is smaller than Na^+. The smaller radius would also lead to increased attraction between the ions.

Our analysis has also assumed that the compounds have similar structures. This may not be so, in which case, the structures may also affect the melting point.

4 Aluminium fluoride (1564 K) has a higher melting point than magnesium fluoride (1534 K), due to the increased charge of the aluminium cation.

The melting point of aluminium oxide (2345 K) is less than magnesium oxide (3125 K), which is not expected from the ionic model. This suggests that aluminium oxide has some covalent character.

5 Some things to consider:

- Model shows cubic structure very clearly.
- Model shows ionic bonds with grey connections as directional. Ionic bonds are due to electrostatic attraction of ions. They are not physical connections.
- Model shows ions as separated. Ions in real structure fill most of the volume, as illustrated in other models shown earlier in the chapter.

Practice questions

1 C

2 C

3 B

4 D

5 B

6 C

7 D

8 A

9 Electrostatic attraction increases with charge and decreases with distance. [1]

Mg^{2+} has smaller radius than Ba^{2+}, O^{2-} has smaller radius than S^{2-} [1]

so ionic radius of MgO is smaller than BaS. [1]

10 **(a)** electrostatic attraction [1]

between Fe^{3+} and O^{2-} ions [1]

(b) Fe^{3+}: $1s^22s^22p^63s^23p^63d^54s^2$ [1]

O^{2-}: $1s^22s^22p^6$ [1]

(c) iron(III) oxide [1]

(d) Applied force moves layers so ions of the same charge are forced to be closer [1]

repulsion between layers makes the compound brittle. [1]

(e) $Fe_3O_4 - Fe_2O_3 = FeO$, so the additional ion of iron is Fe^{2+} [1]

11 **(a)** Structure: lattice of ions [1]

Bonding: electrostatic attraction between K^+ ions and S^{2-} ions [1]

(b) K^+: $1s^22s^22p^63s^23p^6$ [1]

S^{2-}: $1s^22s^22p^63s^23p^6$ [1]

(c) S^{2-} is larger than K^+ [1]

Both have same electron configuration but K has larger nuclear charge so attracts electrons more strongly. [1]

(d) Solid: doesn't conduct because no mobile ions [1]

Molten: conducts because of free moving / mobile ions in molten state [1]

12 **(a)** +5 [1]

Ionization energies increase in a linear pattern and then there is a large jump when 6th electron removed / removal of 6th electron is unfavourable. [1]

(b) X_2O_5 [1]

Ion would have large charge density and so would polarise the anion; compound would show covalent character. [1]

(c) X^{2+} and X^{3+} [1]

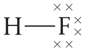
S2.2 The covalent model

Exercises

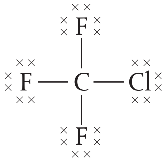
Q1 C

Q2 A

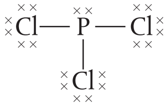
Q3 **(a)**



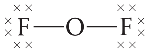
(b)



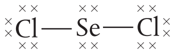
(c)



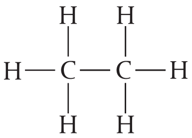
(d)



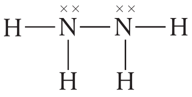
(e)



(f)



(g)



Q4 **(a)** 16

(b) 24

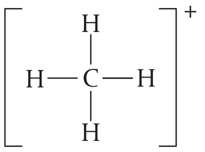
(c) 32

(d) 8

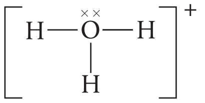
(e) 20

(f) 26

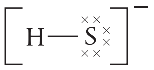
Q5 **(a)**



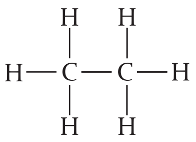
(b)



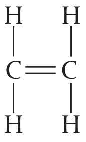
(c)



Q6 **(a)**



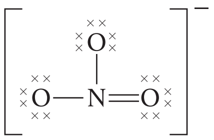
(b)



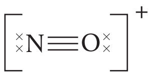
(c) H—C=C—H

Q7 C

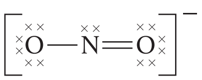
Q8 **(a)**



(b)



(c)



Q9 C

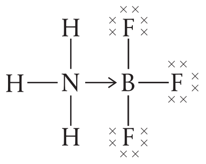
Q10 A

Q11 D

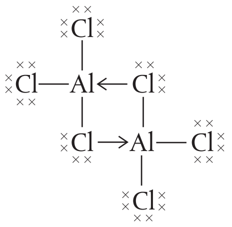
Q12 A

Q13 B

Q14



Q15



Q16 **(a)** Bent / V-shaped, 105° / less than 109.5°

(b) Tetrahedral, 109.5°

(c) Linear, 180°

(d) Trigonal pyramidal, 107° / less than 109.5°

(e) Triangular planar, 120°

(f) Trigonal pyramidal, 107° / less than 109.5°

(g) Bent / V-shaped, 105° / less than 109.5°

Q17 **(a)** Linear, 180°

(b) Bent / V-shaped, less than 120° / in the range 100–119°

(c) Bent / V-shaped, less than 109.5° / in the range 100–108°

(d) Trigonal pyramidal, less than 109.5° / in the range 100–108°

Q18 **(a)** 4

(b) 3 or 4

(c) 2 or 5

(d) 4

(e) 3

Q19 D

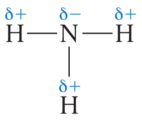
Q20 **(a)** $\overset{\delta+}{\mathrm{H}}-\overset{\delta-}{\mathrm{Br}}$

(b) $\overset{\delta-}{\mathrm{O}}=\overset{\delta+}{\mathrm{C}}=\overset{\delta-}{\mathrm{O}}$

(c) $\overset{\delta+}{\mathrm{Cl}}-\overset{\delta-}{\mathrm{F}}$

(d) $\mathrm{O}=\mathrm{O}$

(e)



$\overset{\delta+}{\mathrm{H}}-\overset{\delta-}{\mathrm{N}}(\overset{\delta+}{\mathrm{H}})-\overset{\delta+}{\mathrm{H}}$

Q21 **(a)** C 2.6, H 2.2, difference = 0.4
C 2.6, Cl 3.2, difference = 0.6, more polar

(b) Si 1.9, Li 1.0, difference = 0.9
Si 1.9, Cl 3.2, difference = 1.3, more polar

(c) N 3.0, Cl 3.2, difference = 0.2
N 3.0, Mg 1.3, difference = 1.7, more polar

Q22 **(a)** polar
(b) non-polar
(c) polar
(d) non-polar
(e) non-polar
(f) polar
(g) non-polar
(h) non-polar

Q23 *cis* isomer has a net dipole moment

Q24 **(a)** Linear, 180°, non-polar
(b) Bent/V-shaped, less than 109.5°/in range 100–108°, polar
(c) Tetrahedral, 109.5°, polar
(d) Triangular planar, 120°, nonpolar

Q25 Similarities: composed of carbon atoms only, both are covalent network structures.

Differences: each carbon is bonded to four other carbon atoms in diamond and three other carbon atoms in graphite; geometry around carbon atoms is tetrahedral in diamond and trigonal planar in graphite; diamond is a continuous C–C structure whereas graphite is distinct layers of carbon atoms (layers are held together by weak intermolecular forces); graphite has delocalized electrons but diamond does not.

Q26 Similarities: strong, high melting points, insoluble in water, non-conductors of electricity, good thermal conductors.

Differences: diamond is stronger and more lustrous; silicon can be doped to be an electrical conductor.

Q27 Graphite and graphene have delocalized electrons that are mobile and so conduct electrical charge. In diamond all electrons are held in covalent bonds and so none are mobile.

Q28 A metal
B covalent network
C polar molecular
D non-polar molecular
E ionic compound

Q29 A

Q30 **(a)** London dispersion forces
(b) London dispersion, dipole–dipole, hydrogen bonding
(c) London dispersion forces
(d) London dispersion, dipole–dipole
(e) London dispersion, dipole–dipole, dipole–induced dipole

Q31 **(a)** C_2H_6
(b) H_2S
(c) Cl_2
(d) HCl

Q32 C

Q33 R_F spot 1 = $\frac{9.0}{12.0}$ = 0.75; R_F spot 2 = $\frac{6.0}{12.0}$ = 0.50

Q34 **(a)** R_F spot X = $\frac{24}{80}$ = 0.30; R_F spot Y = $\frac{72}{80}$ = 0.90

(b) Compound Y is more soluble in the non-polar solvent (mobile phase) so must be less polar than compound X.

Challenge yourself

1 F_2 has a lower bond enthalpy than expected from its atomic radius due to repulsion. The bond length is so short that the lone pairs in the two atoms repel each other, weakening the bond.

2 The definition of electronegativity relates to the attraction of an atom to electrons in a covalent bond. As group 18 elements tend not to bond with other atoms, we do not assign them a Pauling electronegativity value.

3 When bonded to a more electronegative element. Only F is more electronegative than O. For example, in OF_2.

4 The high thermal conductivity of diamond is because of its strong covalent bonds. When heated the bonds becoming vibrationally excited, and, as they are all connected, thermal energy can be readily transferred through the network from one bond to the next. Silicon is similarly a good thermal conductor (but not a good electrical conductor), which is why computer chips need to be cooled to prevent overheating.

5 Diamonds are kinetically stable with respect to graphite, as the conversion has a very high activation energy (see Reactivity 2.2.4). So the reaction generally occurs too slowly to be observed.

6 Run each solution out from separate burettes, and see whether the stream of liquid is deflected in the presence of a charged rod. Only the polar solution will show deflection.

Test solubility with ionic and covalent solvents. The polar solvent will be a better solvent for polar / ionic solutes; the non-polar solvent better for covalent / non-polar solutes.

Practice questions

1 C

2 A

3 A

4 A

5 B

6 C

7 D (please note that there is an error in the question in the first print run, the answer for D should be NH_3).

8 B

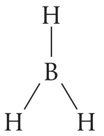
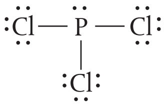
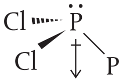
9 B

10 hydrogen bonding in butan-1-ol stronger; [1]

than dipole–dipole attractions in butanal. [1]

Accept converse argument. Do not penalize 'dipole–dipole bonding' instead of 'dipole–dipole attractions'.

11

BH_3	PCl_3
(a) *Lewis formula:*  *Allow x's, dots or lines to represent electrons.*	**(a)** *Lewis formula:*  *Allow x's, dots or lines to represent electrons.* *Lone pairs must be included for mark.*
(b) *Shape:* Triangular / trigonal planar; *Bond angle:* 120°	**(b)** *Shape:* Triangular / trigonal pyramidal; *Bond angle:* Less than 109.5° / in range 100–108°
(c) *Polarity:* non-polar **and** *Explanation:* no net dipole moment / B and H very similar electronegativity values / «small» bond dipole but molecule symmetrical / «small» bond dipoles cancel out / *OWTTE*;	**(c)** *Polarity:* polar **and** *Explanation:* net dipole moment / polar P–Cl bonds **and** asymmetrical molecule / bond dipoles do not cancel /  / *OWTTE*;

[8]

Do not allow ECF in this question from incorrect Lewis formula.

Allow ***(1 max)*** *for stating that* PCl_3 *is polar and* BH_3 *is non-polar without giving a reason or if explanations are incorrect.*

Allow polar bonds cancel for BH_3 *and polar bonds do not cancel for* PCl_3*.*

Do not allow asymmetric molecule as reason for PCl_3 *or symmetric molecule for* BH_3 *as reason alone.*

12 **(a)** *Award* ***(2 max)*** *for three of the following features:*

Bonding

Graphite ***and*** *C_{60} fullerene*: covalent bonds **and** van der Waals' / London / dispersion forces

Diamond: covalent bonds (and van der Waals' / London / dispersion forces)

Delocalized electrons

Graphite ***and*** *C_{60} fullerene*: delocalized electrons

Diamond: no delocalized electrons

Structure

Diamond: network / giant structure / macromolecular / three-dimensional structure **and** *Graphite*: layered structure / two-dimensional structure / planar

C_{60} fullerene: consists of molecules / spheres made of atoms arranged in hexagons / pentagons

Bond angles

Graphite: 120° **and** *Diamond*: 109°

C_{60} fullerene: bond angles between 109° and 120°

Allow Graphite: sp^2 ***and*** *Diamond: sp^3.*

Allow C_{60} fullerene: sp^2 and sp^3.

Number of atoms each carbon is bonded to Graphite ***and*** *C_{60} fullerene*: each C atom attached to 3 others

Diamond: each C atom attached to 4 atoms / tetrahedral arrangement of C (atoms) [6 max]

(b) **(i)** network / giant structure / macromolecular;

each Si bonded covalently to 4 oxygen atoms **and** each O atom bonded covalently to 2 Si atoms / single covalent bonds [2]

Award ***(1 max)*** *for answers such as network-covalent, giant-covalent or macromolecular-covalent.*

Both M1 and M2 can be scored by a suitable diagram.

(ii) *Silicon dioxide*: strong / covalent bonds in network / giant structure / macromolecule

Carbon dioxide: weak / van der Waals' / dispersion / London forces between molecules [2]

(c) triple (covalent) bond;

one electron pair donated by oxygen to carbon atom / dative (covalent) / coordination bond; [2]

*Award (**1 max**) for representation of C≡O.*

*Award **[2]** if CO shown with dative covalent bond.*

13 Methoxymethane is very weakly polar / weak dipole–dipole forces exist between methoxymethane molecules; [1]

Accept London dispersion forces / weak van der Waals' forces

Ethanol contains a hydrogen atom bonded directly to an electronegative oxygen atom / hydrogen bonding can occur between two ethanol molecules / intermolecular hydrogen bonding in ethanol; [1]

the forces of attraction between molecules are stronger in ethanol than in methoxymethane / hydrogen bonding stronger than van der Waals' / dipole–dipole attractions. [1]

Award ***(2 max)*** *if covalent bonds breaking during boiling is mentioned in the answer.*

Penalise only once if no reference given to intermolecular nature of hydrogen bonding or London dispersion / dipole−dipole / van der Waals' forces.

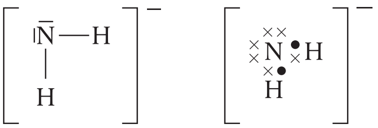
14 **(a)**

;

trigonal pyramidal;

less than 109° / in the range of 100−108°; [3]

(b)



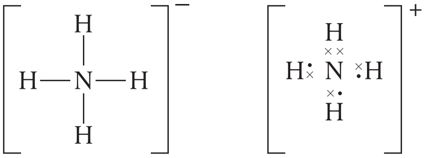
;

Must include negative charge for the mark.

bent / V-shaped;

less than 109° / in the range of 100−106°; [3]

(c)



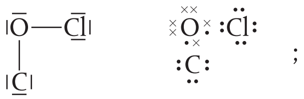
;

tetrahedral;

109.5°; [3]

Penalise once only if electron pairs are missed off outer atoms

15 **(a)** **(i)**

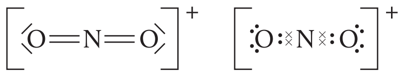


;

bent / v-shaped;

less than 109° / in the range of 100–106°; [3]

(ii)



;

linear;

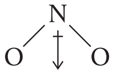
180°; [3]

Allow dots, crosses or lines in Lewis formula.

*Penalize missing charge, missing bracket once only in **(i)** and **(ii)**.*

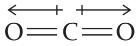
*Lone pairs required for BOTH **(i)** and **(ii)**.*

(b) NO_2:



*Award **[1]** for correct representation of the bent shape and **[1]** for showing the net dipole moment, or explaining it in words (unsymmetrical distribution of charge).*

CO_2:



*Award **[1]** for correct representation of the linear shape and for showing the two equal but opposite dipoles or explaining it in words (symmetrical distribution of charge).* [3]

For both species, allow either arrow or arrow with bar for representation of dipole moment.

Allow correct partial charges instead of the representation of the vector dipole moment.

Ignore incorrect bonds.

Lone pairs not needed.

(c) Structure:

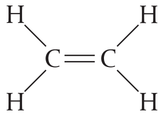
network / giant lattice / macromolecular / repeating tetrahedral units; [1]

Bonding:

(single) covalent (bonds). [1]

It is not necessary to identify which part refers to structure and bonding specifically.

16 **(a)** $H-C\equiv C-H$; [1]



[1]

(b) ethyne bond is shorter because of the greater number of shared electrons so stronger electrostatic attraction to (bonded) nuclei; [1]

(c) London (dispersion) forces; [1]

17 **(a)** measure of how far a component spot has travelled in comparison to the solvent front / mobile phase; [1]

Accept formula

$$R_f = \frac{\text{distance travelled by spot}}{\text{distance travelled by solvent front}}$$

(b) $(R_f = \frac{32}{96} =)$ 0.33; [1]

(c) R_f value compared with literature values «that used the same stationary phase and solvent / mobile phase». [1]

(d) component will have a different degree of solubility in a different solvent; [1]

R_f value is dependent on solubility of component in the mobile phase «and stationary phase»; [1]

(e) thin layer chromatography / TLC;

Accept GLC or HPLC [1]

S2.3 The metallic model

Exercises

Q1 B

Q2 **(a)** Delocalized electrons are able to move freely so can move across a metallic structure in response to an applied voltage.

(b) Delocalized electrons and closely packed cations enable efficient transfer of thermal energy.

(c) The metallic bond is non-directional due to the random movement of delocalised electrons. This allows the metallic bond to remain intact when the lattice of cations undergoes conformation changes under pressure.

Q3 **(a)** malleability, thermal conductivity, high melting point, chemically unreactive due to oxide layer

(b) light, strong, forms alloys

(c) thermal conductivity, high melting point, non-corrosive, chemically unreactive due to oxide layer

(d) light, strong, non-corrosive, chemically unreactive due to oxide layer

4 Beryllium has stronger electrostatic attraction in its metallic bonding as it forms a structure of smaller cations with greater magnitudes of charge (Be^{2+}) and delocalizes two electrons for each cation formed. Lithium has larger cations with only one positive charge (Li^{+}) and delocalizes only one electron for each cation formed.

Q5 C

Q6 The strength of metallic bonding decreases down a group as the size of the cation increases, reducing the attraction between the delocalized electrons and the positively charged nuclei. Weaker bonds require less energy to break, so melting points are lower.

Challenge yourself

1 The additional delocalization of d-electrons will lead to a stronger metallic bond in transition metals (due to higher magnitudes of charge on the lattice of cations). This will lead to higher melting points and densities. The large number of delocalized electrons will lead to increased electrical conductivity in comparison to Group 1 and 2 metals.

Practice questions

1 C

2 D

3 A

4 B

5 **(a)** the electrostatic attraction; [1]

between lattice of cations AND delocalized electrons; [1]

mobile electrons are able to carry charge across a metallic structure. [1]

(b) calcium has smaller ionic radius; [1]

greater attraction between delocalized electrons and «nuclei of» cations; [1]

or

strontium has a larger ionic radius; [1]

weaker attraction between delocalized electrons and «nuclei of» cations. [1]

6 metallic bond / electrostatic attraction / movement of delocalized electrons is non-directional; [1]

metallic bond remains intact when structure undergoes conformation changes; [1]

closely packed cations AND delocalized electrons; [1]

allow for efficient transfer of thermal energy «across metallic structure». [1]

S2.4 From models to materials

Exercises

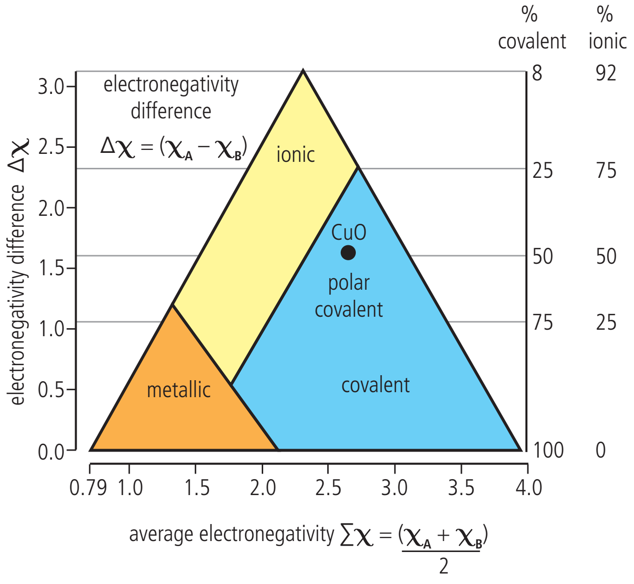
Q1

	Substance	$\chi_{average}$	$\Delta\chi$	% ionic character	Bonding
(a)	Cl_2O	3.3	0.2	6	(Polar) covalent
(b)	$PbCl_2$	2.5	1.4	44	Polar covalent
(c)	Al_2O_3	2.5	1.8	56	Ionic
(d)	HBr	2.6	0.8	25	Polar covalent
(e)	NaBr	1.95	2.1	66	Ionic

The % ionic character is taken from the bonding triangle (S2.4 Figure 2).

Q2

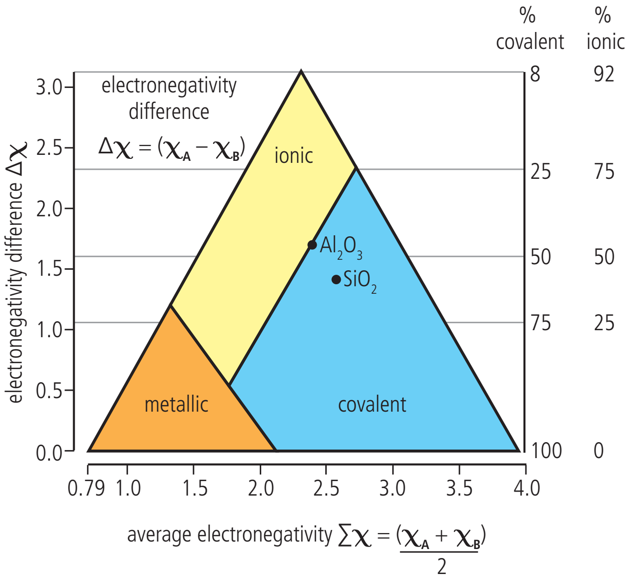
Substance	$\chi_{average}$	$\Delta\chi$	% ionic character	Bonding
CuO	2.65	1.5	47	Polar covalent



Q3

Substance	$\chi_{average}$	$\Delta\chi$	Bonding
SiO_2	2.65	1.5	Polar covalent
Al_2O_3	2.5	1.8	Polar covalent/ionic

Applying these data, we can position these compounds in the bonding triangle. SiO_2 is positioned lower and to the right of Al_2O_3 and therefore has the greater covalent character.



Q4 A

Q5 B

Q6 Alloying element(s) disrupts regular / repeating (metal) lattice. It is difficult for one layer to slide over another / atoms smaller than the metal cations can fit into the (holes of) metal lattice, disrupting bonding. Can make the metal harder / stronger / more corrosion resistant / brittle.

Q7 An alloy is a mixture of one or more metals with other elements. Composites may not contain a metal. Alloys are generally homogeneous mixtures whereas a composite is always a heterogeneous mixture with the components present as separate phases.

Q8 B

Q9 C

Q10 A

Q11 *Recyclable* means a substance can be processed by chemical means into new products.

Biodegradable means a substance will be broken down in the natural environment into harmless end products.

Reusable means a substance can be reused without being changed physically or chemically.

Reusable plastics potentially have a low impact on the environment, as they can remain in circulation rather than accumulating in land-fill or oceans, and there are no side-products of processing. This depends though on individual responsibility for reusing rather than disposing.

Biodegradable plastics also reduce the environmental impact of plastics as they break down into harmless products. This depends though on time and on the conditions during the breakdown process. In some cases, biodegradable plastics may not break down fully.

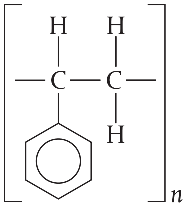
Recyclable plastics enable parts or all of the substance to be reused rather than discarded, which reduces their environmental impact. But the sorting, processing and manufacturing stages in the recycling process have side-products, energy costs and waste products that may themselves have an environmental impact.

The best way to reduce the amount of plastic waste is to reduce consumption. For example, limiting the use of single-use plastics and packaging, and encouraging more reusable plastics. Individuals also must be mindful of responsible disposal of plastics according to their chemical nature and ability to be recycled, biodegraded or composted.

Challenge yourself

1 Alloys are a mixture of metals in which the components are held together by metallic bonding. These bonds are able to form between atoms of the same metals and atoms of different metals due to their non-directional nature. In order to separate a metal from an alloy, the bonds between the different metals have to be broken. This can be done by heating until the metal with the lower melting point forms a liquid and can be separated by filtration from the remaining solid. This is not an economically efficient process. Alloys can also be separated into their component metals by chemical means, such as selectively dissolving with acids.

2



Practice questions

1 A

2 A

3 A

4 **(a)** Y: [1]

bottom left position is metallic bonding. [1]

(b) X (ionic compound), Y (metal) [1]

generally have high melting points. [1]

(c) X; ionic compounds are brittle. [1]

When the regular structure of the lattice is displaced, the ionic bonds can no longer hold the oppositely charged ions in position. [1]

(d) Z (covalent substance) [1]

has low boiling point due to relatively weak intermolecular forces. [1]

5

Substance	χ_{average}	$\Delta\chi$	**Position in triangle and type of bonding**	**Predicted properties**	
Sn	2.0	0.0	metallic	high mp and bp, electrical and thermal conductor	[2]
P_4O_{10}	2.8	1.2	polar covalent	soluble in water, high mp	[2]
Cd_3Mg	1.5	0.4	metallic	high mp and bp, electrical and thermal conductor	[2]
MgO	2.35	2.1	ionic	high mp and bp, conductor of electricity when molten or in aqueous solution	[2]
NCl_3	3.1	0.2	covalent	low mp and bp, non-conductor of electricity, not soluble in water	[2]

6 Polymerization reactions of polypropylene are addition reactions with only one product, the polymer. So all the atoms of reactant are found in the product and the atom economy is 100%. In this sense, the reactions are “green chemistry”, as there is no waste or by-product. However, factors such as energy usage, catalysts and yield less than 100% may all have environmental impact. [2]

7 The alloy is stronger than the pure metal. [1]

Adding atoms of a different size disrupts the regular metal lattice so that it is difficult for one layer to slide over another. [1]

Alloying can make the metal harder, stronger and more resistant to corrosion. [1]

8 **(a)** average electronegativity = $\frac{3.0 + 2.0}{2} = 2.5$

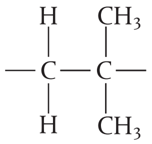
electronegativity difference = 3.0 − 2.0 = 1.0 [1]

bonding in $SbBr_3$ is polar covalent [1]

(b) bonding in $LaBr_3$ is ionic bonding *or* electrostatic forces between ions [1]

slight movement brings ions of same charge adjacent to each other, causing the crystal to break *or* slight movement results in repulsion between layers, causing the crystal to break [1]

9 **(a)**



or

$-CH_2C(CH_3)_2-$ [1]

Continuation bonds needed for mark.

No penalty if brackets present or ‘n’ appears after the bracket / formula.

(b) same mass of product as reactant, thus 100% [1]

accept “less than 100%” only if a reason is given, e.g. the catalyst is not converted into the product, or other reasonable answer.

(c) due to stability of plastics / strong covalent bonds

or

low volatility preventing good mixing with oxygen gas

or

lack of / insufficient oxygen

or

plastics are often parts of devices with non-combustible components, which mechanically prevent the combustion of plastic components

or

PVC already partly oxidized (because some C−H bonds are replaced with C−Cl bonds), so it cannot produce enough heat for complete combustion

or

many industrial / household materials contain additives that reduce their flammability / act as flame retardants [1]

10 **(a)** metallic bonding [1]

(b) In MgB_2, difference in EN = 0.7, average EN = 1.65 [1]

bonding is on the boundary between ionic and metallic

or

not just one type of bonding [1]

S3.1 The periodic table: Classification of elements

Exercises

Q1 D

Q2 C

Q3 B

Q4

	Element	Period	Group
(a)	helium	1	18
(b)	chlorine	3	17
(c)	barium	6	2
(d)	francium	7	1

Q5 **(a)** Periods are rows and groups are columns.

(b) $1s^22s^22p^63s^23p^3$

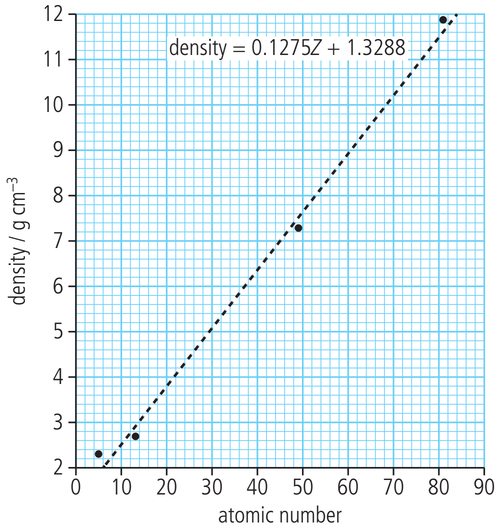
The valence energy level is the third principal energy level, so the element is in period 3. It has the $3p^3$ configuration, so it is in the third group of the p block, which is Group 15.

Q6 Element 51 is antimony (Sb), which is in Group 15. Its valence electrons are $5s^25p^3$, and so it has 5 valence electrons.

Q7 C

Q8 A

Q9



Predicted density 5.3 g cm^{-3}. This agrees with the measured value of 5.1 g cm^{-3}.

Q10 C

Q11 A

Q12 B

Q13 C

Q14 D

Q15 D

Q16 D

Q17 B

Q18 **(a)** Half the distance between the nuclei of neighbouring atoms of the same element bonded together by a covalent bond.

(b) **(i)** The noble gases do not form bonds so the distance between neighbouring atoms is not defined.

(ii) The atomic radii decrease across a period. The nuclear charge increases as the number of protons increases. Electrons are added to the same main energy level. The electrostatic attraction between the outer electrons and the nucleus increases.

Q19 Si^{4+} has two occupied energy levels (isoelectronic with Ne) and Si^{4-} has three (isoelectronic with Ar) and so Si^{4-} is larger.

Q20 **(a)** The electron in the outer electron energy level (level 4) is removed to form K^+. The net attractive force increases as the electrons in the third energy level experience a greater effective nuclear charge.

(b) P^{3-} has one more principal energy level than Si^{4+}. Positive ions are smaller than their parent ions while negative ions are larger than their parent ions.

(c) The ions have the same electron configuration: both have two complete shells; the two extra protons in Na^+ attract the electrons more strongly.

Q21 $Cl^- > Cl > Cl^+$

Q22 D

Q23 D

Q24 B

Q25 Sodium floats on the surface; it melts into a sphere; there is fizzing / effervescence / bubbles; sound is produced; solution gets hot; white smoke is produced.

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

Q26 Down the group, the reactivities of the alkali metals increase but those of the halogens decrease.

Q27 They all have valence electron configuration ns^1. They react by losing this electron. The electron is easier to remove as the atomic radii increase due to decreased attraction between the electron and the nucleus.

Q28 They all have valence electron configuration $ns^{2n}p^5$. They react by gaining an electron. An electron is easier to add as the atomic radii decrease due to increased attraction between the electron and the nucleus.

Q29 A

Q30 D

Q31 C

Q32 **(a)** $MgO(s), SiO_2(s), P_4O_{10}(s), SO_2(g)$

(b) MgO: giant structure with ionic bonding; strong attraction between oppositely charged ions.

SiO_2 (quartz): giant structure covalent bonding; strong covalent bonds throughout structure.

$P_4O_{10}(s)$ molecular, covalent bonding; weak van der Waals' forces between molecules; P_4O_{10} is larger molecule and so has stronger intermolecular bonding.

SO_2: molecular, covalent bonding; weak van der Waals' forces between molecules; SO_2 is smaller molecule and so has weaker intermolecular bonding.

(c) MgO: alkaline solution

$MgO(s) + H_2O(l) \rightarrow Mg(OH)_2(aq)$

SiO_2 (quartz): oxide is insoluble

P_4O_{10}: acidic solution

$P_4O_{10}(s) + 6H_2O(l) \rightarrow 4H_3PO_4(aq)$

SO_2: acidic solution

(d) **(i)** $Al_2O_3(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2O(l)$

(ii) $Al_2O_3(s) + 2NaOH(aq) + 3H_2O(l) \rightarrow 2NaAl(OH)_4(aq)$

Q33 The oxides of Na and Mg are basic; the oxide of Al is amphoteric; the oxides of Si to Cl are acidic. Ar forms no oxide.

$Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$

$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$

Q34 **(a)** Natural rain contains dissolved carbon dioxide, which reacts with water to form carbonic acid:

$CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$

(b) sulfuric acid

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

$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$

$H_2O(l) + SO_3(g) \rightarrow H_2SO_4(aq)$

(c) nitric acid

production reduced by use of lean burn engines, catalytic converters, recirculation of exhaust gases

Q35 C

Q36 C

Q37 **(a)** N = −3, H = +1
(b) Sn = +2, Cl = −1
(c) H = +1, O = −2
(d) N = +5, O = −2
(e) Pb = +4, O = −2
(f) P = +5, O = −2
(g) Cl = +7, O = −2
(h) H = +1, O = −2, Cl = +1

Q38 **(a)** nitric(V) acid
(b) nitric(III) acid
(c) phosphoric(V) acid
(d) lead(IV) oxide
(e) lead(II) sulfate(VI)

Q39 **(a)** Cu_2O
(b) Fe_2O_3
(c) SnO_2
(d) $NaBrO_3$

Q40 **(a)** I is oxidized (−1 to 0), Cl is reduced (0 to −1)
(b) O is oxidized (−2 to 0), F is reduced (0 to −1)
(c) P is oxidized (0 to +5), F is reduced (0 to −1)
(d) Na is oxidized (0 to +1), H is reduced (0 to −1)

Challenge yourself

1 ytterbium, yttrium, terbium, erbium

2 $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 4f^7 6s^2$ *or* $[Xe]4f^7 6s^2$

3 Two liquids, ten gases

4 'Metalloid' refers to the properties of certain elements in relation to the periodic table. 'Semiconductor' refers to the physical properties of materials (including alloys and compounds). There is a partial overlap between the two sets.

5 Coal and oil are fossilized decayed plants or animals; the plants and animals would have contained amino acids. The amino acids methionine and cysteine contain sulfur. Coal and oil with a higher percentage of sulfur are considered 'dirty' because of the sulfur dioxide pollution that is produced on combustion. Sulfur dioxide results in acid rain.

6 $CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(s) + H_2O(l) + CO_2(g)$; $CaSO_4(s)$ has low solubility

$CaCO_3(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2O(l) + CO_2(g)$

Practice questions

1 C

2 A

3 B

4 B

5 A

6 B

7 D

8 A

9 A

10 B

11 A

12 **(a)** the amount of energy required to remove one (mole of) electron(s) [1]

from (one mole of) an atom(s) in the gaseous state [1]

(b) greater positive charge on nucleus / greater number of protons / greater core charge [1]

greater attraction by Mg nucleus for electrons (in the same shell) / smaller atomic radius [1]

13 $Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$ [1]

$SO_3(l) + H_2O(l) \rightarrow H_2SO_4(aq)$ [1]

State symbols are not needed.

Na_2O is basic and SO_3 is acidic [1]

14 **(a)** solution becomes yellow / orange / brown / darker [1]

chlorine is more reactive than iodine (and displaces it from solution) / *OWTTE* [1]

Allow correct equation $Cl_2(g) + 2KI(aq) \rightarrow 2KCl(aq) + I_2(s)$ *for second mark or stating that iodine /* I_2 *is formed.*

(b) no colour change / nothing happens as fluorine is more reactive than chlorine / *OWTTE* [1]

15 **(a)** atomic number / Z [1]

Accept nuclear charge / number of protons.

(b) Across period 3: increasing number of protons / atomic number / Z / nuclear charge [1]

(atomic) radius / size decreases / same shell / energy level / similar shielding / screening (from inner electrons) [1]

No mark for shielding / screening or shielding / screening increases.

Noble gases: do not form bonds (easily) / have a full / stable octet / shell / energy level / cannot attract more electrons [1]

Do not accept 'inert' or 'unreactive' without reference to limited ability / inability to form bonds or attract electrons.

16 **(a)** Na: 11 p, 11 / 2.8.1 e^- and Na^+: 11 p, 10 / 2.8 e^-

or

Na^+ has two electron shells / energy levels, Na has three / *OWTTE* [1]

Na^+ has greater net positive charge / same number of protons pulling smaller number of electrons [1]

(b) Si^{4+}: 10 e^- in two (filled) energy levels / electron arrangement 2.8 / *OWTTE* [1]

P^{3-}: 18 e^- in three (filled) energy levels / electron arrangement 2.8.8, thus larger / *OWTTE* [1]

or

Si^{4+} has two energy levels whereas P^{3-} has three / P^{3-} has one more (filled) energy level [1]

Si^{4+} has 10 e^- whereas P^{3-} has 18 e^- / Si^{4+} has fewer electrons / P^{3+} has more electrons [1]

17 **(a)** ions are in fixed positions in the solid state / there are no moveable ions / *OWTTE* [1]

Do not accept answer that refers to atoms or molecules.

(b) $2O^{2-} \rightarrow O_2 + 4e^- / O^{2-} \rightarrow \frac{1}{2}O_2 + 2e^-$ [1]

Accept e instead of e^-.

(c) **(i)** basic [1]

allow alkaline

(ii) $Na_2O + H_2O \rightarrow 2NaOH / Na_2O + H_2O \rightarrow 2Na^+ + 2OH^-$ [1]

do not accept $\leftrightharpoons$

18 **(a)** repeating pattern of (physical and chemical) properties [1]

(b) 2.8.8 [1]

Two of: the outer energy level / shell is full; the increased charge on the nucleus; great(est) attraction for electrons [2]

(c) 17 p in Cl nucleus attract the outer shell electrons more than 11 p in Na nucleus / greater nuclear charge attracts outer shell electrons more [1]

Allow converse for Na. Do not accept 'has larger nucleus'.

(d) S^{2-} has one proton less / smaller nuclear charge so outer level held less strongly / *OWTTE* [1]

Allow converse for chloride. Do not accept 'has larger nucleus'.

(e) The radii of the metal atoms increase (from Li → Cs) (so the forces of attraction are less between them) / *OWTTE* [1]

The forces of attraction between halogen molecules are van der Waals' forces [1]

these forces increase with increasing mass / number of electrons. [1]

S3.2 Functional groups: Classification of organic compounds

Exercises

Q1 **(a)** $CH_3C(CH_3)(OH)CH_3$

(b) $CH_3CH(CH_3)CH_2CH_3$

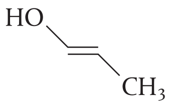
(c) $CH_3CHCHCH_3$

(d) $CH_3C(CH_3)_2CH_2CH_3$

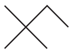
Q2 **(a)**

(b)

(c)



(d)



Q3 **(a)** $CH_3CH(CH_3)CH(CH_3)CH_3$

(b) $CH_3CHClCHBrCH_3$

(c) $CH_3CH_2C(CH_3)CHCH_3$

(d) $CH_3CHCClCHClCH_3$

Q4 **(a)** class: carboxylic acid; functional group: carboxyl (acid)

(b) class: alkene; functional group: alkenyl

(c) class: alcohol; functional group: hydroxyl

(d) class: ester; functional group: carboxyl (ester)

(e) class: aldehyde; functional group: carbonyl (aldehyde)

(f) class: amine; functional group: amino

Q5 B

Q6 D

Q7 D

Q8 A

Q9 B

Q10 Order of increasing volatility: $CH_3CH_2COOH < CH_3CH_2CH_2Br < CH_3CH_2CH_3$

Reasoning: strength of intermolecular forces goes in order: alkane (London dispersion forces) < halogenoalkane (dipole–dipole) < carboxylic acid (H-bonding); volatility is reverse to this order.

Q11 The solubility of the alcohols in water will decrease as the chain length increases. The alcohols contain a highly polar –OH group and a non-polar carbon chain. As the carbon chain increases in length the alcohol will become less polar overall and therefore less soluble in water which is a polar solvent.

Q12 **(a)** class: alkane; name: 2-methylbutane

(b) class: carboxylic acid; name: 2-methylpropanoic acid

(c) class: ketone; name: butanone

(d) class: ester; name: propyl ethanoate

(e) class: alcohol; name: propan-1-ol

(f) class: aldehyde; name: pentanal

Q13 **(a)** $HCOOH$

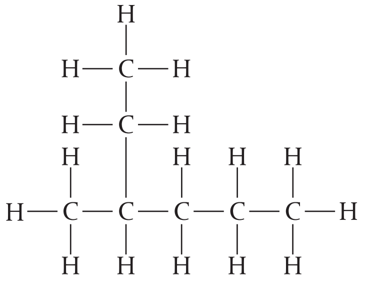
(b) $CH_3CH_2CCCH_2CH_3$

(c) $CH_3CH_2CH_2CH_2CHO$

(d) $CH_3COCH_2CH_3$

(e) $CH_3CH_2OCH_2CH_2CH_3$

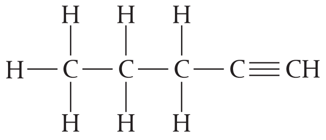
(f) $CH_3CHBrCH_2OH$**Q14**

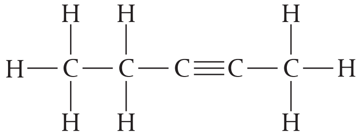


The longest chain contains six carbons. The correct name is 3-methylhexane.

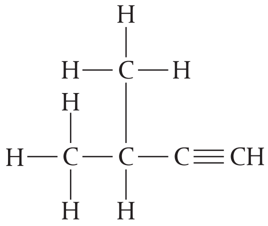
Q15 D

Q16 **(a)**



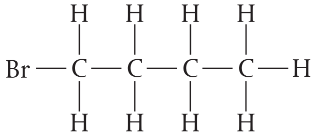


(b)

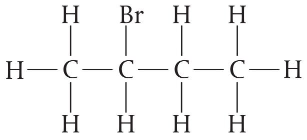


Q17 D

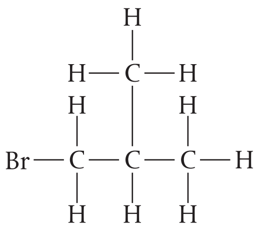
Q18 1-bromobutane, primary



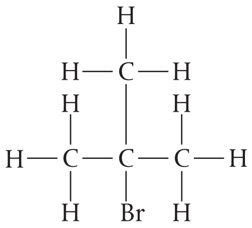
2-bromobutane, secondary



1-bromo-2-methylpropane, primary



2-bromo-2-methylpropane, tertiary



Q19 **(a)** $CH_3OCH_2CH_3$ (methoxyethane) and either $CH_3CH_2CH_2OH$ (propan-1-ol) or $CH_3CH(OH)CH_3$ (propan-2-ol)

(b) CH_3COCH_3 (propanone) and CH_3CH_2CHO (propanal)

(c) $CH_3CH_2CH_2C(O)OH$ (butanoic acid) and either $CH_3C(O)OCH_2CH_3$ (ethyl-ethanoate) or $HC(O)OCH_2CH_2CH_3$ (propyl methanoate) or $CH_3CH_2C(O)OCH_3$ (methyl propanoate)

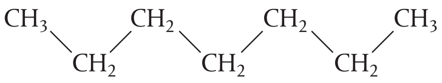
Challenge yourself

1 Using the experimental data, the ratio of the carbon–carbon bond length to the radius of a carbon nucleus can be calculated;

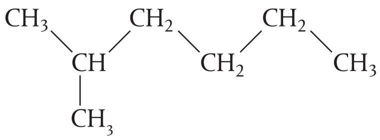
$$\text{Ratio} = \frac{154 \times 10^{-12}\,\text{m}}{2.7 \times 10^{-15}\,\text{m}} = 57\,000$$

Therefore, the plastic stick used in a model would need to be 57 000 times longer than the radius of the plastic ball used to represent a carbon nucleus.

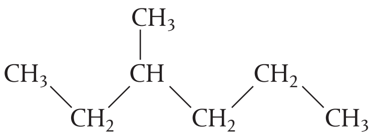
$57\,000 \times 0.5\,\text{cm} = 28\,500\,\text{cm} = 285\,\text{m}$



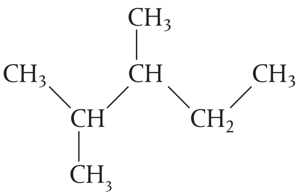
heptane



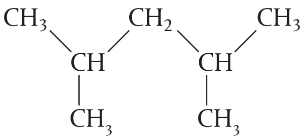
2-methylhexane



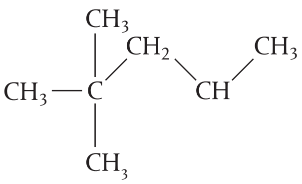
3-methylhexane



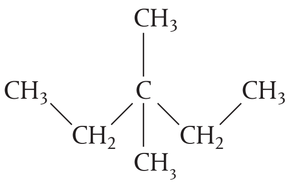
2,3-dimethylpentane



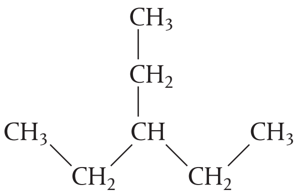
2,4-dimethylpentane



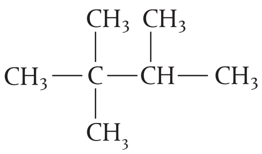
2,2-dimethylpentane



3,3-dimethylpentane



3-ethylpentane



2,2,3-trimethylbutane

Practice questions

1 B

2 D

3 A

4 D

5 A 1-bromobutane [1]

B 2-bromobutane [1]

C 2-bromo-2-methylpropane [1]

D 1-bromo-2-methylpropane [1]

6 **(a)** The compound contains the $-COO-$ group, so it is an ester. [1]

(b) ethyl butanoate [1]

7 A: 1-chlorobutane; it is a primary compound

B: 2-chloro-2-methylpropane; it is a tertiary compound

1 mark for correct name for A.

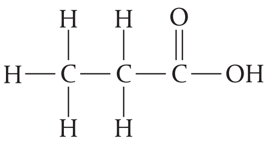
1 mark for correct name for B.

1 mark for primary for A **and** tertiary for B. [3]

8 **(a)** ethyl methanoate and methyl ethanoate are both esters [1]

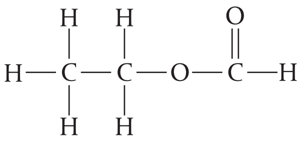
(b) the three compounds are structural isomers [1]

(c) propanoic acid



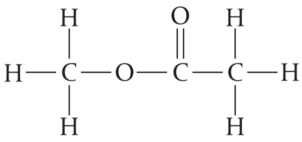
[1]

ethyl methanoate



[1]

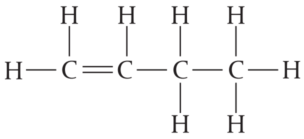
methyl ethanoate



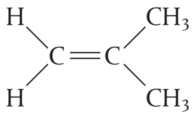
[1]

9 **(a)** but-1-ene and 2-methylpropene are both alkenes

(b)

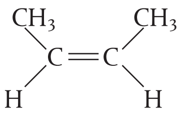


but-1-ene



2-methylpropene

(c) but-2-ene



10 **(a)** $CH_3{-}CH_2{-}CH_2{-}CH_2I$ primary; $CH_3{-}CH_2{-}CHI{-}CH_3$ secondary; $CH_3{-}C(CH_3)I{-}CH_3$ tertiary

(b) Secondary: 2-iodopropane,
Tertiary: 2-iodo-2-methylpropane

(c) All three isomers contain the same polar C-I bond but the strength of the London Dispersion Forces will be weakest for the isomer with the most branching.

The tertiary isomer, 2-iodo-2-methylpropane, has the most branching and will therefore have the weakest intermolecular forces and lowest boiling point.

11 **(a)** Methoxypropane is an ether.

(b) Butanol is a structural isomer of methoxypropane:

$CH_3\text{-}CH_2\text{-}CH_2\text{-}CH_2\text{-}OH$

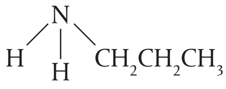
(c) Butanol is an alcohol and has hydrogen bonding as the dominant intermolecular force.

Methoxypropane is an ether so has dipole-dipole as the dominant intermolecular force.

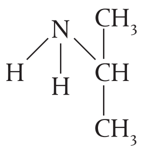
Hydrogen bonding is the stronger intermolecular force so butanol will have a higher boiling point than methoxypropane.

12 **(a)**

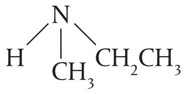
Primary:



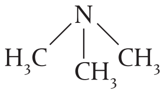
or



Secondary:



Tertiary:



(b) The primary and secondary isomers of C_4H_9N are all able to hydrogen bond. (The tertiary isomer does not have a hydrogen bonded to the nitrogen atom so it cannot hydrogen bond.) The isomer with the least branching will have the strongest London Dispersion Forces so the straight chain, primary amine (propan-1-amine) has the strongest overall intermolecular forces and the highest boiling point.

R1.1 Measuring enthalpy changes

Exercises

Q1 Heat is transferred from the object with a high temperature T_H to the object at temperature T_C.

The two objects will have the same final temperature. The final temperature is closer to T_H than T_C.

Q2 $\Delta T = 310 - 300 = 10\ K$

$\Delta T = (310 - 273) - (300 - 273) = 10\ °C$

Q3 B

Q4 B

Q5 B

Q6 C

Q7 A

Q8 B

Q9 A

Q10 A

Q11 D

Q12 C

Q13 A

Q14 A

Q15 C

Q16 C

Q17 **(a)** B
(b) C

Q18 B

Q19 $q = mc\Delta T$, so $\Delta T = \frac{q}{mc}$

$$\frac{q}{mc} = \frac{100}{100 \times 0.138} = 7.2\ °C$$

$T = 25.0 + 7.2 = 32.2\ °C$

Q20 **(a)** $\Delta T = 36.50 - 25.85 = 10.65\ °\text{C (or K)}$

$q = mc\Delta T$

$q = [m(H_2O) \times c(H_2O) \times \Delta T(H_2O)] + [m(Cu) \times c(Cu) \times \Delta T(Cu)]$

$= (200.00\ \text{g} \times 4.18\ \text{g}^{-1}\ \text{K}^{-1} \times 10.65\ \text{K}) + (120.00\ \text{g} \times 0.385\ \text{g}^{-1}\ \text{K}^{-1} \times 10.65\ \text{K})$

$= 8903.4 + 492.0\ \text{J}$

$= 9395.4\ \text{J}$

$$n(C_6H_{12}O_6) = \frac{1.10\ \text{g}}{180.18\ \text{g mol}^{-1}}$$

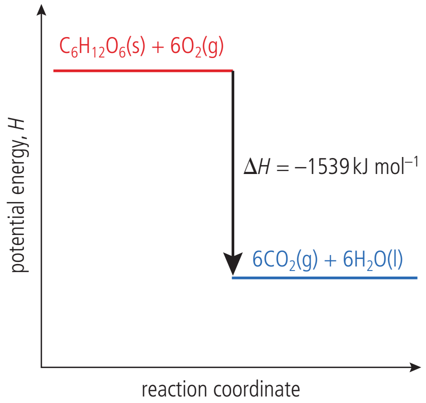
$= 6.11 \times 10^{-3}\ \text{mol}$

$$\Delta H_c = \frac{9395.4\ \text{J}}{6.11 \times 10^{-3}\ \text{mol}}$$

$= -1539 \times 10^3\ \text{J mol}^{-1}$

$= -1539\ \text{kJ mol}^{-1}$

(b)



Q21 $q = mc\Delta T$

$= 150.00\ \text{g} \times 4.18\ \text{J g}^{-1}\ \text{K}^{-1} \times (31.5 - 25.0)\ \text{K}$

$= 4075.5\ \text{J}$

$$n(P) = \frac{0.0500\ \text{g}}{30.97\ \text{g mol}^{-1}}$$

$= 1.614 \times 10^{-3}\ \text{mol}$

$$\Delta H_c = \frac{4075.5\ \text{J}}{1.614 \times 10^{-3}\ \text{mol}}$$

$= 2525 \times 10^3\ \text{J mol}^{-1}$

$\approx 2500\ \text{kJ mol}^{-1}$

The precision of the answer is limited by the precision of measurement of the temperature difference. The value is lower than the literature value owing to heat losses and incomplete combustion.

Q22 **(a)** $m(\text{propan-1-ol}) = 45.65 - 45.05 = 0.60\ \text{(g)}$

$n(\text{propan-1-ol}) = 0.60 / 60.11 = 0.00998\ \text{mol}$

(b) $\Delta T = 41.5 - 23.7 = 17.8\ \text{(K)}$

$q = mc\Delta T$

$= 210 \times 4.18 \times 17.8$

$= 15.62\ \text{kJ}$

(c) $$\text{enthalpy change (kJ mol}^{-1}) = \frac{15.62\ \text{kJ}}{0.00998\ \text{mol}}$$

$= -1565\ \text{kJ mol}^{-1}$

(d) Not all heat produced transferred to water as heat is lost to surroundings. Incomplete combustion, with carbon and carbon monoxide being formed instead of carbon dioxide also leads to a less exothermic reaction.

Challenge yourself

1 When heat energy is added to water some of the energy is needed to break hydrogen bonds and so less energy is available to increase the kinetic energy. The presence of hydrogen bonds reduces the temperature increase and so increases the specific heat capacity.

2 $Q = m \times c \times \Delta T$

temperature change (water and aluminium)
$= 77.5 - 24.5 = 53.0 \pm 0.2\ \text{K}$

energy lost by brass = energy gained by water and calorimeter

energy gained by water and calorimeter
$= (200.00 \times 4.18 \times 53.0) + (80.00 \times 0.900 \times 53.0)\ \text{J}$

$= 44\,308 + 3\,816 = 48\,124\ \text{J}$

energy lost by brass = 48 124 J

$212.10 \times 0.400 \times (T_{\text{Bunsen}} - 77.5) = 48\,124\ \text{J}$

$$T_{\text{Bunsen}} - 77.5 = \frac{48\,124}{212.10 \times 0.400}$$

$= 567.2\ °\text{C}$

temperature of brass in Bunsen flame = 567.2 + 77.5 °C

= 644.7 » 645 °C

Practice questions

1 A

2 D

3 B

4 C

5 D

6 **(a)** **(i)** $m(\text{methanol}) = 80.557 - 80.034 = 0.523\ (\text{g})$ [1]

$$n(\text{methanol}) = \frac{0.523\ \text{g}}{32.05\ \text{g mol}^{-1}} = 0.0163\ (\text{mol})$$ [1]

Award **[2]** *for correct final answer.*

(ii) $\Delta T = 26.4 - 21.5 = 4.9\ (\text{K})$ [1]

$q = mc\Delta T = 20.000 \times 4.18 \times 4.9\ (\text{J})$
or $20.000 \times 4.18 \times 4.9 \times 10^{-3}\ (\text{kJ})$ [1]

$= 410\ \text{J}$ *or* $0.41\ \text{kJ}$ [1]

Award **[3]** *for correct final answer.*

(iii) $$\Delta H^{\ominus}_{c} = -\frac{410\ (\text{J})}{0.0163\ (\text{mol})} \text{ or } -\frac{0.41\ (\text{kJ})}{0.0163\ (\text{mol})}$$ [1]

$= -25\,103\ \text{J mol}^{-1}$ *or* $-25\ \text{kJ mol}^{-1}$ [1]

Award **[2]** *for correct final answer. Award* **[1]** *for (+)25 (kJ mol^{-1}).*

(b) not all heat produced transferred to water / heat lost to surroundings/environment / *OWTTE* / incomplete combustion (of methanol) / water forms as $H_2O(l)$ instead of $H_2O(g)$ [*Do not allow just 'heat is lost'*] [2]

7 **(a)** all heat is transferred to water/copper sulfate solution / no heat loss;

specific heat capacity of zinc is zero/negligible / no heat is absorbed by the zinc;

density of water/solution = 1.0 / density of solution = density of water;

heat capacity of cup is zero / no heat is absorbed by the cup;

specific heat capacity of solution = specific heat capacity of water;

temperature uniform throughout solution;

Award **[1]** *each for any two. Accept 'energy' instead of 'heat'.* [2]

(b) **(i)** $T_{\text{final}} = 73.0\ (^{\circ}\text{C})$ [1]

Allow in the range 72 to 74 (°C).

$\Delta T = 73.0 - 24.8 = 48.2\ (^{\circ}\text{C})$ [1]

Allow in the range 47 to 49 (°C). Award **[2]** *for correct final answer. Allow ECF if* T_{final} *or* T_{initial} *correct.*

(ii) temperature decreases at uniform rate (when above room temperature) / *OWTTE* [1]

(iii) 10.1 (kJ) [1]

Allow in the range 9.9 to 10.2 (kJ).

(c) Complete colour change shows all the copper has reacted

so $$n(\text{Zn}) = n(\text{CuSO}_4) = \frac{1.00 \times 50.0}{1000} = 0.0500\ (\text{mol})$$ [1]

(d) $-201\ \text{kJ mol}^{-1}$ [1]

Allow in the range −197 to −206 (kJ mol^{-1}). Value must be negative to award mark.

8 **(a)** $\Delta T = 30.3 - 24.5 = 5.8\ \text{K}$ [1]

$q = m(H_2O) \times c(H_2O) \times \Delta T(H_2O)$

$= 100.0\ \text{g} \times 4.18\ \text{J g}^{-1}\ \text{K}^{-1} \times 5.8\ \text{K}$

$= 2424.4\ \text{J}$ [1]

KOH is the limiting reagent.

$$n(\text{KOH}) = \frac{50 \times 0.950}{1000} = 0.0475\ \text{mol}$$ [1]

$$\Delta H = -\frac{2424.4}{0.0475} = -68.64 \times 10^{3}\ \text{J mol}^{-1}$$

$= -51.04\ \text{kJ mol}^{-1}$ [1]

(b) Assumptions: no heat loss, $c(\text{solution}) = c(\text{water})$, $m(\text{solution}) = m(H_2O)$, $\text{density}(H_2O) = 1.00$ [1]

9 $\Delta H_{\text{reaction}} = -\Delta H_{\text{water}}$

$q = mc\Delta T$

$\Delta H_{\text{water}} = 100.00\ \text{g} \times 4.18\ \text{J g}^{-1}\ \text{K}^{-1} \times (21.79 - 25.55)\ \text{K}$
$= -1571.68\ \text{J}$ [1]

$$n(\text{NH}_4\text{Cl}) = \frac{5.35\ \text{g}}{53.50\ \text{g mol}^{-1}} = 0.100\ \text{mol}$$ [1]

$$\Delta H = \frac{-\Delta H_{\text{water}}}{n(\text{NH}_4\text{Cl})} = \frac{-(-1571.68\ \text{J})}{0.100\ \text{mol}} = +15\,716.8\ \text{J mol}^{-1}$$

$= +15.7\ \text{kJ mol}^{-1}$ [1]

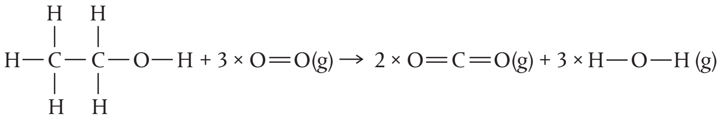
R1.2 Energy cycles in reactions

Exercises

Q1	B
Q2	A
Q3	1 × C–C + 6 × C–H
Q4	B
Q5	B
Q6	B
Q7	D

Q8 $-482\ \mathrm{kJ\ mol^{-1}}$

$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$



$$\mathrm{H{-}\overset{H}{\underset{H}{C}}{-}\overset{H}{\underset{H}{C}}{-}O{-}H} + 3 \times \mathrm{O{=}O(g)} \rightarrow 2 \times \mathrm{O{=}C{=}O(g)} + 3 \times \mathrm{H{-}O{-}H\ (g)}$$

Bonds broken	**$\Delta H^\ominus$ / kJ mol^{-1}**	**Bonds formed**	**$\Delta H^\ominus$ / kJ mol^{-1}**
C–C	+346	4 C=O	4 × (−804)
3 × O=O	3 × (+498)	6 H–O	6 × (−463)
O–H	+463		
C–O	+358		
5 × C–H	5 × (+414)		
Total	+4731		−5994

$\Delta H^\ominus = +4731 - 5994\ \mathrm{kJ\ mol^{-1}} = -1263\ \mathrm{kJ\ mol^{-1}}$

The calculated value is less exothermic than the enthalpy of combustion in Table 14 of the data booklet ($-1367\ \mathrm{kJ\ mol^{-1}}$) because the bond enthalpy calculation assumes all species are in the gaseous state: water and ethanol are liquids.

Q9 B

Q10 B

Q11 $\Delta H^\ominus = -394 - (-283)\ \mathrm{kJ} = -111\ \mathrm{kJ\ mol^{-1}}$

Q12 $\Delta H^\ominus = -180.5 + (+66.4) = -114.1\ \mathrm{kJ\ mol^{-1}}$

Q13 $\Delta H^\ominus = 2 \times (-33.2) + (+9.16) = -57.24\ \mathrm{kJ\ mol^{-1}}$

Challenge yourself

1 **(a)** The bond enthalpies generally decrease down the halogen group but the bond enthalpy in F_2 is weaker than Cl_2 and Br_2. This is due to repulsion between the non-bonding pairs in the two fluorine atoms.

(b)

Average of E(H–H) and E(F–F) / kJ mol^{-1}	**E(H–F) / kJ mol^{-1}**	**$E(H{-}F) - \left(\frac{E(H{-}H) + E(F{-}F)}{2}\right)$ / kJ mol^{-1}**
$\frac{436 + 159}{2} = +297.5$	+567	+269.5

The difference shows that the H–F bond is very polar and F is very electronegative.

2 The difference in the values is largely due to the assumption that H_2O is gaseous in the bond enthalpy calculation.

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$
$\Delta H^\ominus = -891\ \mathrm{kJ\ mol^{-1}}$

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$
$\Delta H^\ominus = -808\ \mathrm{kJ\ mol^{-1}}$

The difference is due to formation of H_2O in different states:

$2H_2O(l) \rightarrow 2H_2O(g)$ $\Delta H^\ominus = 891 - 808 = +83\ \mathrm{kJ\ mol^{-1}}$
$H_2O(l) \rightarrow H_2O(g)$ $\Delta H^\ominus = +41.5\ \mathrm{kJ\ mol^{-1}}$

There are (on average) 2 hydrogen bonds between each water molecule so the strength of a hydrogen bond is approximately $20\ \mathrm{kJ\ mol^{-1}}$.

This assumes that all other molecular interactions, such as dipole–dipole and London forces, are negligible, which is an additional approximation.

3 **(a)** $E(O{-}O) = +144\ \mathrm{kJ\ mol^{-1}}$, $E(O{=}O) = +498\ \mathrm{kJ\ mol^{-1}}$. The average value $= 321\ \mathrm{kJ\ mol^{-1}}$.

This suggests that the oxygen–oxygen bond in ozone is intermediate between a single bond and double bond, as discussed in Structure 2.2.

(b) The O=O bond in O_2 is stronger than the bond in O_3. Less energy / longer wavelength radiation is required to dissociate O_3 than O_2.

4 The compounds must be heated for reaction to occur. It is difficult to distinguish any energy changes due to the reaction from the heat input.

Practice questions

1 C

2 B

3 C

4 D

5 B

6 A

7 A

8 $NH_3(g) \rightarrow N(g) + 3H(g)$ [1]

$\Delta H^{\ominus} = -\frac{1}{2}x + \frac{1}{2}y + \frac{1}{2}z = \frac{1}{2}(y + z - x)$ [1]

Bond energy $= \frac{\Delta H^{\ominus}}{3} = \frac{1}{6}(y + z - x)$ [1]

9 bonds broken: 4 × N−H, 1 × N−N, 1 × O=O = +2220 (kJ mol^{-1}) [1]

bonds formed: 1 × N≡N, 4 × O−H = −2797 (kJ mol^{-1}) [1]

enthalpy change = −2797 + 2220 = −577 kJ mol^{-1} [1]

Award **[3]** *for correct final answer.*

10 ∑bonds broken = 4 × 414 «kJ» + 2 × 463 «kJ» / 2582 «kJ [1]

∑bonds formed = 1077 «kJ» + 3 × 436 «kJ» / 2385 «kJ» [1]

ΔH «= ∑bonds broken − ∑bonds formed =(2582 kJ − 2385 kJ)» = «+»197 «kJ» [1]

Award **[3]** *for final correct answer.*
Award **[2 Max]** *for final answer of −197 kJ*

R1.3 Energy from fuels

Exercises

Q1 C

Q2 A

Q3 C

Q4 8.78 g

Q5 The compound contains C and H and possibly O

$n(CO_2) = \frac{0.88}{44.01} = 0.0200 = n(C)$

0.60 g contains 0.0200 × 12.01 = 0.24 g of C

$n(H_2O) = \frac{0.36}{18.02} = 0.0200$; $n(H) = 0.0400$

0.60 g contains 0.0200 × 2 × 1.01 = 0.0404 g of H

Mass of O = 0.60 − (0.24 + 0.04) = 0.32 g; $n(O) = \frac{0.32}{16.00} = 0.02$

Empirical formula = CH_2O

Q6 $C_xH_y + (x + \frac{y}{4})O_2(g) \rightarrow xCO_2(g) + \frac{y}{2}H_2O(l)$

Q7 $C_xH_y + (\frac{x}{2} + \frac{y}{4})O_2(g) \rightarrow xCO(g) + \frac{y}{2}H_2O(l)$

Q8 $C_6H_{14} + 8O_2(g) \rightarrow 3CO(g) + 3CO_2(g) + 7H_2O(l)$

Q9 **(a)** $2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$

(b) 1 mol of C_8H_{18} reacts with $12\frac{1}{2}$ mol of O_2

114.26 g of C_8H_{18} reacts with 400 g of O_2

1.00 g of C_8H_{18} reacts with 3.50 g of O_2

(c) mass of air $= \frac{100}{20} \times 3.50$ g =17.5 g

(d) CO affects oxygen uptake in the blood. It is absorbed by the lungs and binds to haemoglobin in red blood cells more effectively than oxygen and is bound permanently. This prevents oxygen from being transported around the body.

(e) catalytic converters

Q10 **(a)** Carbon has a higher enthalpy of combustion than hydrogen (per mol) but the atomic mass of hydrogen is much less than the atomic mass of carbon. As the percentage carbon content decreases and the percentage hydrogen content increases, the specific energy increases.

(b) Carbon dioxide is a greenhouse gas and causes global warming.

(c)

Fuel	**Specific energy / kJ g^{-1}**	**Carbon content by mass / %**	**Carbon dioxide produced / g**
Coal	32	94	108
Oil	42	83	72
Natural gas	55	75	50
Wood	15	70	171
Hydrogen	142	0	0

Hydrogen is the best fuel and wood is the worst fuel on this basis. But, the carbon dioxide in the wood was taken out of the atmosphere when the wood was growing and more wood can be grown. Taking this into account, coal is the worst fuel.

Q11 **(a)** Empirical formula: $C_{135}H_{96}O_9NS$ or $C_{136}H_{97}O_9NS$. (It typically also contains trace amounts of silicon, sodium, calcium, aluminium, nickel, copper, zinc, arsenic, lead and mercury.)

(b) Some elements burn to give acidic oxides that dissolve / react with water to give acid rain.

$S(s) + O_2(g) \rightarrow SO_2(g)$; $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$

$H_2O(aq) + SO_2(g) \rightarrow H_2SO_3(aq)$;

$H_2O(l) + SO_3(g) \rightarrow H_2SO_4(aq)$

$2N(g) + O_2(g) \rightarrow 2NO(g)$;

$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$;

$2NO_2(g) + H_2O(l) \rightarrow HNO_3(aq) + HNO_2(aq)$

Q12 **(a)** $\frac{\text{useful output energy}}{\text{total input energy}} = \frac{\text{useful electrical energy}}{\text{total input heat energy}} = 0.38$

heat energy produced by combustion per second $= \frac{5.00 \times 10^5 \text{ kJ s}^{-1}}{0.38}$

mass per second $= \frac{5.00 \times 10^5 \text{ kJ s}^{-1}}{0.38 \times 33.0 \text{ kJ g}^{-1}}$

$= 39\,872$ g s^{-1} » 39.9 kg s^{-1}

(b) We have the balanced equation:

$4CH(s) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(g)$

no. of moles of CO_2 = no. of moles of CH

no. of moles of CH burnt per second

$= \frac{39872}{12.01 + 1.01} = 3062$ mol

mass of CO_2 per second $= 3062 \times 44.01$

$= 134\,775$ g » 135 kg

Q13 Carbon-containing fuels are non-renewable.

They are needed as chemical feedstocks.

Their combustion adds carbon dioxide to the atmosphere which contributes to global warming.

Q14 **(a)** High specific energy / energy density. As a liquid it is convenient to handle and deliver. Easy to vaporize, which assists combustion.

(b) It was formed by the partial decomposition of marine plants millions of years ago.

Q15 It is more efficient.

It produces more thermal energy per unit of mass / has a higher specific energy / energy density.

It produces less CO_2 per unit of output energy.

Q16 **(a)** Fossil fuels and biomass are derived from the sun through photosynthesis.

(b) Advantages: renewable and has little environmental impact.

Disadvantages: photosynthesis is not very efficient, so relatively little of the available solar energy is trapped. Land used to grow plants for biomass could be used to grow plants for food.

Q17 **(a)** $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$

(b) chlorophyll

(c) Process: fermentation

Equation: $C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$

Conditions: acidity / absence of oxygen / below 40 °C. Yeast provides enzyme.

Q18 **(a)** methane

(b) particulates (soot), hydrocarbons, carbon monoxide

(c) Fossil fuels are being phased out due to carbon dioxide production and limited resources. Biomass is a renewable source.

Q19 **(a)** $H_2(g)$ is oxidized at the anode:
$2H_2(g) \rightarrow 4H^+(aq) + 4e^-$

$O_2(g)$ is reduced at the cathode:
$4H^+(aq) + O_2(g) + 4e^- \rightarrow 2H_2O(l)$

(b) Less waste heat produced and more chemical energy converted to useful energy. They can use a renewable energy source to form the hydrogen, which is more efficient.

Challenge yourself

1 The compound contains C, H and N.

$n(CO_2) = \frac{0.456}{44.01} = 0.010361 = n(C)$

0.20 g contains 0.010361×12.01 g of C

= 0.124439 g of C

$n(H_2O) = \frac{0.248}{18.02} = 0.013762$

$n(H) = 2 \times 0.013762 = 0.027525$

0.20 g contains 0.027525×1.01 g of H

= 0.0278 g of H

mass of N = 0.20 − 0.124439 − 0.0278 = 0.047761 g

$n(N) = \frac{0.047761}{14.01} = 0.003409$

	C	H	N
Mol	0.010361	0.027525	0.003409
Ratio	3	8	1

Empirical formula = C_3H_8N

2 combination of nitrogen and oxygen at high temperatures: $N_2(g) + O_2(g) \rightarrow 2NO(g)$

acid rain and (photochemical) smog

3 **(a)** % mass of C in coal:

$$\frac{12.01}{13.02} \times 100 = 92.24\%$$

% mass of C in wood:

$$\frac{5 \times 12.01}{5 \times 12.01 + 9 \times 1.01 + 4 \times 16.00} \times 100 = 45.10\%$$

(b) Carbon footprint of coal:

mass of coal needed:

$$\frac{1.33 \times 10^6\,\text{kJ}}{31\,\text{kJ g} -1} \times \frac{100}{65} = 66\,005\,\text{g} = 66.0\,\text{kg}$$

mass of carbon dioxide produced:

$$66.0\,\text{kg} \times \frac{92.24\%}{100} \times \frac{44.01}{12.01} = 223.1\,\text{kg}$$

Carbon footprint of wood:

mass of wood needed:

$$\frac{1.33 \times 10^6\,\text{kJ}}{22\,\text{kJ g} -1} \times \frac{100}{70} = 86\,364\,\text{g} = 86.4\,\text{kg}$$

mass of carbon dioxide produced:

$$86.4\,\text{kg} \times \frac{45.10\%}{100} \times \frac{44.01}{12.01} = 142.7\,\text{kg}$$

(c) Carbon footprint for wood doesn't account for the carbon dioxide taken in while the wood was growing or that more wood can be grown relatively quickly, taking carbon dioxide out of the atmosphere,

4 **(a)** Distant from localized areas of pollution; data present an accurate measure of global levels of CO_2.

(b) The annual variation is due to CO_2 uptake by growing plants. Over a year, the uptake is highest in the northern hemisphere springtime. Considerably more of the northern hemisphere is land (for plants) compared with the southern hemisphere.

(c) Photosynthesis: $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$

CO_2 dissolves in water: $CO_2 + H_2O \rightarrow H_2CO_3(aq)$

(d) Decreased level of photosynthesis: less CO_2 taken in by plants.

5 **(a)** 1%

(b) Wavelength of radiation not absorbed by chlorophyll. Some radiation is reflected or heats the surface of the Earth. Plants do not cover all the Earth's surface.

6 **(a)** $2C_{16}H_{23}O_{11} + 19H_2O + O_2 \rightarrow xH_2 + yCO + zCO_2$

$2C_{16}H_{23}O_{11}$		$19H_2O$	O_2	xH_2	yCO	zCO_2	
$-\frac{1}{16}$					+2	+4	C
	+1	+1		0			H
	−2	−2	0		−2	−2	O

(b) C is oxidized

O and H are reduced

(c) Balancing the H atoms:

$2C_{16}H_{23}O_{11} + 19H_2O + O_2 \rightarrow 42H_2 + yCO + zCO_2$

$x = 42$

(d) Total change of oxidation number of H = −84

Total change of oxidation number of O = −4

Total increase in oxidation number of C = +88

Balancing the change in oxidation numbers for C: $+88 - 2 = 2y + 4z$

Balancing the C atoms: $y + z = 32$

Solving the equations: $2z = 86 - 64 = 22$

$z = 11$ and $y = 21$

$2C_{16}H_{23}O_{11} + 19H_2O + O_2 \rightarrow 42H_2 + 21CO + 11CO_2$

(e) $2H_2 + CO \rightarrow CH_3OH$

$42H_2 + 21CO \rightarrow 21CH_3OH$

2 molecules of wood produce 21 molecules of methanol

1 molecule of wood produces 10.5 molecules of methanol

7 The entropy decrease is smaller for the reaction which produces one mole of gaseous water, which leads to a larger efficiency.

Practice questions

1 **(a)** long wavelength / infrared / IR radiation from Earth's surface

(some of this radiation) is absorbed (by gas) [1]

Do not accept 'trapped' or 'blocked'.

Do not award mark for 'IR from Sun'.

causes (increased) vibration in bonds

re-radiates heat back to the Earth [1]

Accept 're-transmits'

Do not accept 'reflects/bounces'

(b) no change in polarity/dipole (moment) when molecule vibrates; [1]

(c) two from:

CO: prevents haemoglobin from transporting oxygen

Particulate / Soot: respiratory problems, lung diseases, carcinogenic

Nitrogen oxides: photochemical smog / acid deposition/rain [2]

2 **(a)** two from:

wide availability

produce energy at appropriate rate

ease of transportation

current technology is based on fossil fuels

high energy density / specific energy [2]

(b) hydrogen [1]

3 **(a)** coal 92.2%, petrol 84.1% and natural gas 74.8% [3]

(b) two from:

they have higher specific energy

liquid or gaseous state make them more convenient to use / easier to transport

produce less pollution / smaller carbon foot print [2]

(c) two from:

hydrogen has a very high specific energy / energy density

it is clean burning, producing only H_2O when it is burned

can be generated by electrolysis, electricity can come from renewable sources [2]

4 **(a)**

Formula	M /g mol⁻¹	DH_c / kJ mol⁻¹	Specific enthalpy / kJ g⁻¹
C_3H_8	44.11	−2219	$\frac{2219}{44.11} = 50.31$
C_4H_{10}	58.14	−2878	$\frac{2878}{58.14} = 49.50$

[2]

(b) $pV = nRT$

$\rho = \frac{m}{V}$; $M = \frac{m}{n}$

$\rho = \frac{PM}{RT}$

C_3H_8

$\rho = \frac{1.00 \times 10^5 \times 44.11}{8.13 \times 273} = 1987 \text{ g m}^{-3}$

$= 1.987 \times 10^{-3} \text{ g cm}^{-3}$ heat produced per cm³

$= 50.31 \text{ kJ g}^{-1} \times 1.987 \times 10^{-3} \text{ g cm}^{-3}$

$= 0.09999 \text{ kJ g}^{-1}$

» 0.10 kJ g^{-1}

C_4H_{10}

$\rho = \frac{1.00 \times 10^5 \times 58.14}{8.13 \times 273} = 2620 \text{ g m}^{-3}$

$= 2.620 \times 10^{-3} \text{ g cm}^{-3}$

heat produced per cm³

$= 49.50 \text{ kJ g}^{-1} \times 2.620 \times 10^{-3} \text{ g cm}^{-3}$

$= 0.1297 \text{ kJ g}^{-1}$

» 0.13 kJ g^{-1} [3]

5 **(a)** photosynthesis [1]

(b) $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$ [1]

(c) fermentation [1]

$C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$ [1]

6 **(a)**

Fuel	Formula	Standard enthalpy of combustion ΔH / kJ mol⁻¹	Relative molecular mass	Specific energy / kJ g⁻¹
methane	CH_4	−891	16.05	55.5
methanol	CH_3OH	−726	32.05	22.7

[2]

(b) CH_4:

4 × H is +4, so ox. no. (C) = −4

CH_3OH:

4 × H is +4 and O is −2, so ox. no. (C) = −2 [1]

(c) the more oxidized the C the lower the specific energy [1]

(d) for $C_6H_{12}O_6$, specific energy $= \frac{2803}{180.18} = 15.56 \text{ kJ g}^{-1}$

for $C_6H_{12}O_6$: 12 × H is +12, 6 × O is −12, so ox. no. C = 0

the result supports the hypothesis [3]

7 **(a)** Cathode: $2H_2O + O_2 + 4e^- \rightarrow 4OH^-$

Anode: $2H_2 + 4OH^- \rightarrow 4H_2O + 4e^-$

Award ***[1]*** *if both equations given but at wrong electrodes.* [2]

8 **(a)** $M_r(C_8H_{18}) = 114.26$ and $\Delta Hc^\ominus = -5470$ «kJ mol^{-1}»;

«specific energy = $\dfrac{5470\text{ kJ}}{0.11426\text{ kg}}$ =» 4.79×10^4 / 47873 / 47900 «kJ kg^{-1}»;

Award ***[2]*** *for correct final answer.*

Accept "48 x 10^3 *«kJ kg*$^{-1}$*»"* ***OR*** *"47.9 x* 10^3 *«kJ kg*$^{-1}$*»".*

(b) Wood is less useful because it requires «about three times» more mass for same energy. [1]

Accept "octane is more useful because it has higher specific energy".

9 **(a)** specific energy = $\dfrac{\text{energy released}}{\text{molecular mass}}$

$= \dfrac{5470\text{ kJ mol }-1}{114.26\text{ g mol }-1} = 47.9$ «kJ g^{-1}»

energy density = specific energy × density

= 47.9 kJ g^{-1} × 0.703 g cm^{-3}» = 33.7 «kJ cm^{-3} »

Do ***not*** *accept "−47.9 «kJ g*$^{-1}$*»".*

Do ***not*** *accept "−33.7 «kJ cm*$^{-3}$*»" unless "−47.9 «kJ g*$^{-1}$*»" already penalized.* [2]

(b) energy is lost «to the surroundings» as heat/sound/friction

or

energy is lost to the surroundings «as heat/sound/friction»

or

incomplete combustion

Do ***not*** *accept just "energy is lost".* [1]

10 **(a)** Methane is the major component of natural gas. It has the formula CH_4. [1]

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$ [1]

(b) Natural gas is the cleanest of the fossil fuels to burn as it has a high H:C ratio. [1]

The combustion of natural gas produces minimal amounts of carbon monoxide, hydrocarbons and particulates. [1]

It does contribute to global warming but does not contribute to acid rain, unlike coal and oil. [1]

(c) Natural gas is the fossil fuel in the shortest supply and it is unevenly distributed around the world. Oil is expected to last a little longer. Coal, which is distributed more evenly around the world, is expected to last longer still. [1]

R2.1 How much? The amount of chemical change

Exercises

Q1 **(a)** $CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$

(b) $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

(c) $H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

(d) $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

(e) $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

Q2 **(a)** $2K + 2H_2O \rightarrow 2KOH + H_2$

(b) $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

(c) $Cl_2 + 2KI \rightarrow 2KCl + I_2$

(d) $4CrO_3 \rightarrow 2Cr_2O_3 + 3O_2$

(e) $Fe_2O_3 + 3C \rightarrow 3CO + 2Fe$

Q3 **(a)** $2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$

(b) $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$

(c) $3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$

(d) $6H_2O_2 + 2N_2H_4 \rightarrow 2N_2 + 10H_2O + O_2$

(e) $4C_2H_7N + 15O_2 \rightarrow 8CO_2 + 14H_2O + 2N_2$

Q4 **(a)** 2.50 mol

(b) 5.63 mol

(c) 665.4 g

Q5 **(a)** $2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$

(b) 1.59 g

Q6 4.355 kg

Q7 **(a)** $CaCO_3 \rightarrow CaO + CO_2$

(b) 92.8%

(c) $CaCO_3$ is the only source of CO_2, all the $CaCO_3$ undergoes complete decomposition, all CO_2 released is captured, heating does not cause any change in the mass of the other minerals present.

Q8 **(a)** 2.40 mol

(b) 0.01101 mol

(c) 44 mol

Q9 **(a)** 35.64 dm^3

(b) 5.7 dm^3

Q10 0.6524 dm^3

Q11 0.138 mol Br_2 and 0.1561 mol Cl_2, so more molecules of Cl_2

Q12 0.113 dm^3

Q13 0.28 dm^3

Q14 0.178 mol dm^{-3}

Q15 0.0220 mol dm^{-3}, 0.0802% HCl

Q16 0.106 mol dm^{-3} Na_2SO_4 and 0.115 mol dm^{-3} $Pb(NO_3)_2$; assume no side reactions, all $PbSO_4$ precipitates, $PbSO_4$ is completely dry and none is lost in the drying process.

Q17 C

Q18 B

Q19 **(a)** 10 cm^3

(b) 5 cm^3

Q20 **(a)** 85.2 g

(b) 1.25 g H_2

Q21 5.22 g $C_2H_4Cl_2$

Q22 253 g theoretical $CaSO_3$; 78.2%

Q23 3.16 g of ester

Q24 107 g of C_6H_6 needed

Q25 72.0%

Q26 63.2%

Q27 Method 2

Challenge yourself

1 52% NH_3 by mass; assuming no side reactions occur and gases behave as ideal gases.

2 Many reactions with 'useless' by-products could have high stoichiometric yield under optimum conditions but have a low atom economy.

For example, methanoic acid production:

$$2NaCOOH + H_2SO_4 \rightarrow 2HCOOH + Na_2SO_4$$

For 100% yield, the atom economy is

$$\frac{2 \times 46.03}{(2 \times 68.01) + 98.09} \times 100\,\% = 39.32\%$$

3 1217 tonnes (There is an error in the formula of tri-ammonium phosphate in the first print run of the book. The correct formula for tri-ammonium is $(NH_4)_3PO_4$. Using the corrected formula gives an answer of 1242 tonnes.)

4 320 tonnes / 3.20×10^5 kg

5 0.225 mol dm^{-3} (or round to 0.23 mol dm^{-3})

6 $[Na_2CO_3] = \dfrac{YP}{X}$ mol dm^{-3}

$[NaHCO_3] = \dfrac{Y(Q - P)}{X}$ mol dm^{-3}

7 100 mol O_2; 3.2 kg

Practice questions

1 D

2 D

3 B

4 C

5 C

6 D

7 D

8 D

9 D

10 A

11 **(a)** $n(N_2) = \dfrac{80.0}{22.7} = 3.52$ mol; [1]

$n(NaN_3) = \dfrac{2}{3} \times 3.52 = 2.35$ mol; [1]

$m(NaN_3) = 2.35 \times (22.99 + 3 \times 14.01) = 153$ g [1]

(b) $n(Na) = n(NaN_3) = 2.35$ mol; [1]

$m(Na) = 2.35 \times 22.99 = 54.0$ g [1]

12 NH_3/ammonia in excess AND by 10 dm^3 [1]

25.0 dm^3 N_2 produced [1]

13 **(a)** $n(HCl) = 0.200$ mol $dm^{-3} \times 0.02720$ $dm^3 = 0.00544$ *or* 5.44×10^{-3} (mol) [1]

(b) $n(HCl)$ excess is $0.100\,mol\,dm^{-3} \times 0.02380\,dm^3 = 0.00238$ or $2.38 \times 10^{-3}\,mol$ [1]

Penalise not dividing by 1000 once only in (a) and (b).

(c) $n(HCl)$ reacted $= 0.00544 - 0.00238 = 0.00306$ or 3.06×10^{-3} (mol) [1]

(d) $2HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$ *or*

$2H^+(aq) + CaCO_3(s) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$

[2]

*Award **[1]** for correct reactants and products.*

*Award **[1]** if the equation is correctly balanced.*

*Award **(1 max)** for the following equations:*

$2HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + H_2CO_3(aq)$

$2H^+(aq) + CaCO(s) \rightarrow Ca^{2+}(aq) + HCO(aq)$

Ignore state symbols.

(e) $n(CaCO_3) = \frac{1}{2}n(HCl) = \frac{1}{2} \times 0.00306 = 0.00153$ *or* $1.53 \times 10^{-3}\,mol$ [2]

*Award **[2]** for correct final answer.*

(f) $M_r(CaCO_3) = 40.08 + 12.01 + 3 \times 16.00 = 100.09$ or $100.1\,g\,mol^{-1}$;

Accept 100.

$m(CaCO_3) = n \times M = 0.00153\,mol \times 100.09\,g\,mol^{-1} = 0.153\,g$;

$\%CaCO_3 = \frac{0.153}{0.188} \times 100 = 81.4\%$ *or* 81.5% [3]

Ignore state symbols.

Accept answers in the range 79.8% to 81.5%.

*Award **[3]** for correct final answer.*

(g) Only $CaCO_3$ reacts with acid / impurities are inert / non-basic / impurities do not react with the acid / nothing else in the eggshell reacts with acid / no other carbonates. [1]

Do not accept 'all calcium carbonate reacts with acid'.

14 **(a)** M(desired product) $= 3(12.01\,g\,mol^{-1}) + 6(1.01\,g\,mol^{-1}) + 6(14.01\,g\,mol^{-1}) = 126.15\,g\,mol^{-1}$ AND M(all reactants) $= 24(1.01\,g\,mol^{-1}) + 12(14.01\,g\,mol^{-1}) + 6(12.01\,g\,mol^{-1}) + 6(16.00\,g\,mol^{-1}) = 360.42\,g\,mol^{-1}$ [1]

atom economy $= \frac{126.15\,g\,mol^{-1}}{360.42\,g\,mol^{-1}} \times 100\% = 35.0\%$ [1]

(b) $n((NH_2)_2CO) = \frac{14.0}{60.07} = 0.233\,mol$; [1]

$n(C_3H_6N_6) = \frac{0.233 \times 1}{6} = 0.0388\,mol$ AND

«theoretical yield» $m(C_3H_6N_6) = 0.0388 \times 126.15 = 4.89\,g$ [1]

(c) % yield $= \frac{4.12}{4.89} \times 100 = 84.3\%$ [1]

Using alternative answer % yield $= \frac{4.12}{5.00} \times 100 = 82.4\%$

*Allow ECF from part **(b)**.*

R2.2 How fast? The rate of chemical change

Exercises

Q1 Reaction gives off a gas, so change in volume of gas or mass of flask and contents could be measured.

Reaction involves purple MnO_4^- ions being reduced to colourless Mn^{2+} ions, so colorimetry could be used.

Reaction involves a change in the concentration of ions (23 on the reactants side and 2 on the products sides), so conductivity could be used.

Q2 C

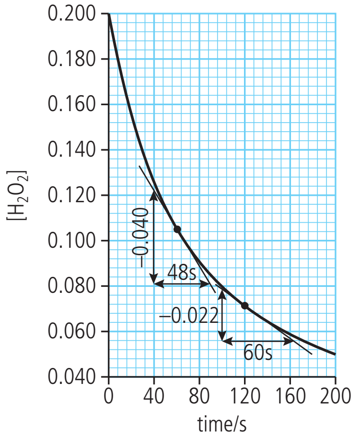
Q3 **(a)** Measure the decrease in the mass of flask and contents.

Measure the increase in pH of the reaction mixture.

Measure the increase in volume of gas collected.

(b) The rate of the reaction decreases with time because the concentration of the acid decreases.

Q4



At 60 s, rate $= 8.3 \times 10^{-4}\,mol\,dm^{-3}\,s^{-1}$

At 120 s, rate $= 3.7 \times 10^{-4}\,mol\,dm^{-3}\,s^{-1}$

Q5 D

Q6 A

Q7 The reaction requiring the simultaneous collision of two particles is faster. The simultaneous collision of three particles is statistically less likely.

Q8 B

Q9 B

Q10 B

Q11 The ashes must contain a catalyst that speeds up the reaction between sugar and oxygen. (Deduced from the fact that all other factors that affect reaction rate can be ruled out.)

Q12 **(a)** $2CO(g) + 2NO(g) \rightarrow 2CO_2(g) + N_2(g)$

(b Carbon monoxide and nitrogen monoxide are toxic gases.

(c) The increased surface area of the catalyst in contact with exhaust gases will increase catalyst efficiency.

(d) Catalytic activity involves the catalyst interacting with the gases, and the reaction occurring on its surface. As temperature increases, the increased kinetic energy of the gases increases the frequency with which they bind to the catalyst.

(e) Catalytic converters reduce pollution from cars but do not remove it completely. As in **(d)**, they are not effective when the engine first starts from cold, when an estimated 80% of pollution occurs. Other pollutants in car exhausts are not removed by the catalyst, e.g. ozone, sulfur oxides and particulates. Also the catalytic converter increases the output of CO_2, a serious pollutant because of its greenhouse gas properties.

Challenge yourself

1 Collecting a gas over warm water will cause its temperature, and therefore its volume, to increase.

2 A homogeneous catalyst is likely to be more difficult to recover as it is in the same phase as the reactants and products. Possible approaches might include a solvent extraction or distillation.

Practice questions

1 B

2 A

3 A

4 C

5 B

6 C

7 C

8 A

9 C

10 D

11 B

12 D

13 A

14 B

15 **(a)** $ZnCO_3(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + CO_2(g) + H_2O(l)$ [2]

Award **[1]** *for correct formulas and balancing;* **[1]** *for correct state symbols.*

(b) CO_2 is produced and escapes in an open system. [1]

(c) the rate of reaction is greatest at the start and decreases with time;

the concentration of acid decreases as it reacts and so collisions between reactants become less frequent;

the rate approaches zero as the limiting reactant is used up. [3]

(d) draw a tangent to curve A and curve B at $t = 0$;

measure the gradient of each tangent;

tangent on curve A has higher gradient than tangent on curve B [3]

(e) A at a higher temperature than B;

catalyst used in A but not in B;

the $ZnCO_3(s)$ in A was in smaller pieces than in B [3]

Accept $ZnCO_3$ *had larger surface area in A (or converse argument).*

(f) «Higher temperature» increases the frequency of collisions;

«and» increase proportion of particles with kinetic energy greater than activation energy / more particles with KE > E_a;

«A catalyst» provides an alternate reaction route with a lower E_a / decreases activation energy;

«therefore» proportion of particles with kinetic energy greater than activation energy increases / more particles have KE > E_a;

«Smaller particle size/greater surface area» leads to higher frequency of collisions between reactants;

«therefore» higher frequency of successful collisions. [6]

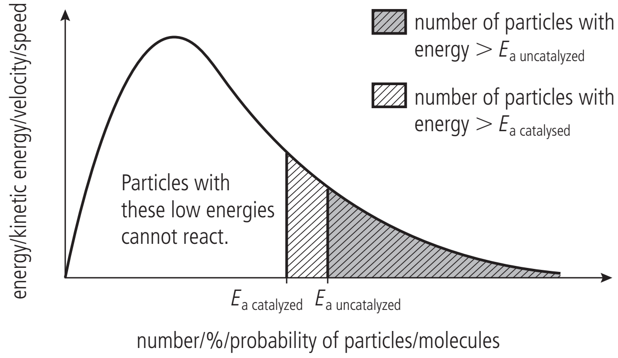
(g) The change in mass would be much less as $H_2(g)$ has a much lower molar mass than $CO_2(g)$;

Not satisfactory as the change in mass as the reaction proceeds is too small to measure accurately / very small changes in mass would have large experimental uncertainty / *OWTTE*. [2]

16 **(a)** T2 is greater than T1. [1]

(b) Higher temperature leads to higher average kinetic energy of particles;
graph shifts to the right «and area under curve remains constant». [2]

(c)



[4]

(d) **(i) and (ii)** «Catalysts» do not change energy/potential energy of reactants and products;

«so» enthalpy change/difference between PEreactants and PEproducts «at constant pressure» does not change;

«catalysts» decrease Ea for forward and backward reaction equally;

reaction occurs faster but yield remains constant. **[4]**

R2.3 How far? The extent of chemical change

Exercises

Q1 A

Q2 C

Q3 B

Q4 **(a)** $K = \dfrac{[NO_2]^2}{[NO]^2\,[O_2]}$

(b) $K = \dfrac{[NO_2]^4\,[H_2O]^6}{[NH_3]^4\,[O_2]^7}$

(c) $K = \dfrac{[CH_3OH]\,[Cl^-]}{[CH_3Cl]\,[OH^-]}$

Q5 **(a)** $N_2O_4(g) \rightleftharpoons 2NO_2(g)$

(b) $CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3H_2(g)$

Q6 **(a)** $3F_2(g) + Cl_2(g) \rightleftharpoons 2ClF_3(g)$

$K = \dfrac{[ClF_3]^2}{[F_2]^3[Cl_2]}$

(b) $2NO(g) \rightleftharpoons N_2(g) + O_2(g)$

$K = \dfrac{[N_2][O_2]}{[NO]^2}$

(c) $CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3H_2(g)$

$K = \dfrac{[CO][H_2]^3}{[CH_4][H_2O]}$

Q7 **(a)** Mostly reactants

(b) Mostly reactants

(c) Mostly products

Q8 2.00

Q9 3.60×10^{-3}

Q10 B

Q11 D

Q12 C

Q13 **(a)** Shift to the left

(b) Shift to the right

(c) No shift in equilibrium

Q14 **(a)** Shift to the left

(b) Shift to the right

(c) This is equivalent to an increase in pressure, so shifts to the left

(d) Shift to the right

(e) Shift to the right

Q15 **(a)** amount of CO will decrease

(b) amount of CO will decrease

(c) amount of CO will increase

(d) no change in amount of CO

Q16 C

Q17 B

Q18 The Haber process is exothermic in the forward direction. Therefore, increasing the temperature will decrease the value of *K*. This represents a decrease in the reaction yield.

Challenge yourself

1 Earth receives energy from the Sun and disperses energy, largely as heat. But exchange of matter is minimal. The only exceptions to Earth being a closed system are matter received from space, such as asteroids and space dust, and matter lost to space, such as spacecraft.

2 The different values of *K* indicate different stabilities of the hydrogen halides. The bonding is the strongest in HCl and the weakest in HI.

This is largely because of the size of the atoms. As I has a larger atomic radius that Cl, the bonding pair in HI is further from the nucleus than the bonding pair in HCl. This means that the bonding pair in HI experiences a weaker pull. The HI bond breaks more easily and so the dissociation reaction to form H_2 and I_2 is favoured.

3 **(a)** $K' = K^2 = 77\,300$

(b) $K' = \dfrac{1}{\sqrt{K}} = 0.0600$

(c) $K' = \dfrac{1}{K^3} = 4.65 \times 10{-}8$

4 The value for *K* at 298 K for the reaction $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$ is extremely low, so the equilibrium mixture lies to the left with almost no production of NO. But at higher temperatures, such as in vehicle exhaust fumes, the reaction shifts to the right and a higher concentration of NO is produced. This gas is easily oxidised in the air, producing the brown gas NO_2 which is responsible for the brownish haze: $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$.

5 The atom economy of the Haber process is 100% as there is only one product. In other words, there is no waste. But, due to the reversible nature of this reaction, this does not mean that all reactants are converted into product, so the stoichiometric yield is less than 100%.

It is the goal of these industries to maximize yield and efficiency by choosing the optimum conditions, taking equilibrium and kinetic considerations into account.

Practice questions

1 C

2 A

3 D

4 A

5 C

6 D

7 A

8 D

9 **(a)** $K = \dfrac{[SO_3]^2}{[O_2][SO_2]^2}$ [1]

Ignore state symbols.

Square brackets required.

(b) yield (of SO_3) increases / equilibrium moves to right / more SO_3 formed;

3 gaseous molecules → 2 gaseous molecules / decrease in volume of gaseous molecules / fewer gaseous molecules on right-hand side [2]

Do not allow ECF.

(c) yield (of SO_3) decreases;

forward reaction is exothermic / reverse/ backwards reaction is endothermic / equilibrium shifts to absorb (some of) the heat [2]

Do not accept exothermic reaction or Le Châtelier's principle.

Do not allow ECF.

(d) rates of both forward and reverse reactions increase equally; no effect on position of equilibrium; no effect on value of K [3]

10 **(a)** exothermic

Accept either of the following for the second mark.

increasing temperature favours endothermic/ reverse reaction; as yield decreases with increasing temperature [2 max]

(b) yield increases / equilibrium moves to the right / more ammonia;

increase in pressure favours the reaction which has fewer moles of gaseous products [2]

(c) $«K» = \frac{[NH_3]^2}{[N_2][H_2]^3}$ [1]

11 **(a)** reactants and products in same phase/state;

rate of forward reaction = rate of reverse reaction;

concentrations of reactants and products remain constant / macroscopic properties remain constant [2 max]

Do not accept concentrations are equal.

(b) $K = \frac{[HI]^2}{[H_2][I_2]}$ [1]

(c) no change to position of equilibrium [1]

(d) the reaction is exothermic / heat is given out / ΔH is negative [1]

(e) $K' = \frac{1}{K} = \frac{1}{160} = 6.25 \times 10^{-3}$ [1]

(f) no effect (on the value of the equilibrium constant);

as it speeds up forward and reverse reaction / concentrations of reactants and products do not change / position of equilibrium does not change / no change in yield [2]

R3.1 Proton transfer reactions

Exercises

Q1 **(a)** HSO_3^-

(b) $CH_3NH_3^+$

(c) C_2H_5COOH

(d) HNO_3

(e) HF

(f) H_2CO_4

Q2 **(a)** $H_2PO_4^-$

(b) CH_3COO^-

(c) HSO_3^-

(d) SO_4^{2-}

(e) O_2^-

(f) Br^-

Q3 **(a)** CH_3COOH (acid) / CH_3COO^- (base)

BH_3 (base) / NH_4^+ (acid)

(b) CO_3^{2-} (base) / HCO_3^- (acid)

H_3O^+ (acid) / H_2O (base)

(c) NO_2^- (base)/NHO_2 (acid)

Q4 $HPO_4^{2-}(aq) + H_2O(l) \rightleftharpoons PO_4^{3-}(aq) + H_3O^+(aq)$ (acid behaviour)

$HPO_4^{2-}(aq) + H_2O(l) \rightleftharpoons H_2PO_4^-(aq) + OH^-(aq)$ (base behaviour)

Q5 **(a)** Dissolved carbon dioxide forms the weak acid carbonic acid, which ionises partially in solution to release H^+ ions.

$H_2O(l) + CO_2(g) \rightleftharpoons H_2CO_3(aq)$

$H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq)$

(b) $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$

$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$

(c) Lime, CaO, is basic and so can help to neutralise acidity.

$CaO(s) + H_2SO_4(aq) \rightarrow CaSO_4(s) + H_2O(l)$

Q6 Burning coal releases non-metal oxides, such as SO_2 and nitrogen oxides, NO_x, which contribute to acid rain. Burning coal also contributes to climate change from the release of greenhouse gases, and to respiratory diseases from release of particulates and some heavy metals. Coal mining also has large impacts on the environment and health.

Q7 pH increases by 1 unit

Q8 pH = 4.72

Q9 pH = 2.0

Q10 pH = 4.49, acidic

Q11 $[H^+] = 5.0 \times 10^{-9}$ mol dm^{-3}

Q12 The pH scale converts $[H^+]$ values with a wide range of concentrations, and most often with negative exponents, into a simpler numeric scale. pH numbers are usually positive and have no exponents. The pH scale is a smaller range of numbers than $[H^+]$ values in mol dm^{-3}.

Q13 temperature < 298 K

Q14 $[H^+] = 1.3 \times 10^{-9}$ mol dm^{-3}

Q15 $[H^+] = 1.5 \times 10^{-7}$, pH = 6.8

Q16 $[H^+] = 1.0 \times 10^{-9}$ mol dm^{-3},
$[OH^-] = 1.0 \times 10^{-5}$ mol dm^{-3}

Q17
- **(a)** $[OH^-] = 2.9 \times 10^{-6}$ mol dm^{-3}; basic
- **(b)** $[H^+] = 1.0 \times 10^{-12}$ mol dm^{-3}; basic
- **(c)** $[H^+] = 1.0 \times 10^{-4}$ mol dm^{-3}; acidic
- **(d)** $[OH^-] = 1.2 \times 10^{-10}$ mol dm^{-3}; acidic

Q18 $[OH^-] = 2.2 \times 10^{-7}$, basic

Q19
- **(a)** pH = 6.9
- **(b)** pH = 2
- **(c)** pH = 4.8

Q20 NH_3 acts as a weak Brønsted–Lowry base because it ionizes only partially in solution to accept a proton.

$NH_3(g) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

The equilibrium position of this reaction lies to the left, in favour of reactants.

Q21 B

Q22 A

Q23
- **(a)** H_2CO_3
- **(b)** HCOOH

Q24
- **(a)** $H_2SO_4(aq) + CuO(s) \rightarrow CuSO_4(aq) + H_2O(l)$
- **(b)** $HNO_3(aq) + NaHCO_3(s) \rightarrow NaNO_3(aq) + H_2O(l) + CO_2(g)$
- **(c)** $H_3PO_4(aq) + 3KOH(aq) \rightarrow K_3PO_4(aq) + 3H_2O(l)$
- **(d)** $HCl(aq) + CH_3NH_2(l) \rightarrow CH_3NH_3^+Cl^-$

Q25 B

Q26
- **(a)** nitric acid + sodium carbonate / sodium hydrogencarbonate / sodium hydroxide

  $2HNO_3(aq) + Na_2CO_3(aq) \rightarrow 2NaNO_3(aq) + H_2O(l) + CO_2(g)$
- **(b)** hydrochloric acid + ammonia solution

  $HCl(aq) + NH_4OH(aq) \rightarrow NH_4Cl(aq) + H_2O(l)$
- **(c)** copper(II) oxide + sulfuric acid

  $H_2SO_4(aq) + CuO(s) \rightarrow CuSO_4(aq) + H_2O(l)$
- **(d)** methanoic acid + potassium hydroxide

  $HCOOH(aq) + KOH(aq) \rightarrow KCOOH(aq) + H_2O(l)$

Q27 pH 7.0; equivalence point

Q28 $HNO_3(aq) + KOH(aq) \rightarrow KNO_3(aq) + H_2O(l)$

pH = 2.0

Q29 0.0083 mol dm^{-3}

Challenge yourself

1 The product NH_4Cl forms smoke which is a solid, from gaseous reactants and so can be classified as sublimation or deposition.

2 $2OH^- \rightarrow O^{2-} + H_2O$

This shows the transfer of a proton from one hydroxide ion to another. So OH^- is both accepting H^+ to form H_2O (basic behaviour) and losing H^+ to form O^{2-} (acidic behaviour).

The two equations showing amphiprotic behaviour are:

$OH^- + H^+ \rightarrow H_2O$ and $OH^- \rightarrow O^{2-} + H^+$

3 $Al_2(SO_4)_3(s) + 6H_2O(l) \rightarrow 2Al(OH)_3(s) + 3H_2SO_4(aq)$

Aluminium sulfate dissolves to form aluminium hydroxide, which is insoluble and so is responsible for the cloudy appearance, and a dilute solution of sulfuric acid, which causes the low pH.

A more detailed explanation is that $Al_2(SO_4)_3(s)$ ionizes in water to form Al^{3+} and SO_4^{2-} ions. As Al^{3+} is small with a high charge density, it hydrolyzes water that is bound to it in the hydrated complex, releasing H^+ and lowering the pH.

$$[Al(H_2O)_6]^{3+}(aq) \rightleftharpoons [Al(H_2O)_5OH]^{2+}(aq) + H^+(aq)$$

Aluminium sulfate is the salt of a strong acid and a weak base, and forms an acidic solution.

4 Vinegar is ethanoic acid, CH_3COOH. When added to soil this will lower the pH and cause hydrangea flowers to turn blue, as explained in the text.

5 In weak acids and weak bases, the majority of the acid or base is not ionized at the start of the reaction, so energy is used to break bonds in ionizing the reactants. This endothermic process reduces the net energy released when water forms. For example, the neutralization reaction between NaOH and CH_3COOH is less exothermic than the neutralization reaction between NaOH and HCl.

Practice questions

1 C

2 D

3 B

4 A

5 A

6 C

7 A

8 D

9 D

10 B

11 B

12 **(a)** **(i)** use a pH meter, use a suitable indicator

(ii)

HCOOH(aq)	5	weak acid
KCl(aq)	7	neutral salt
$HNO_3(aq)$	1	strong acid
Na(OH)(aq)	13	strong base
$NH_3(aq)$	10	weak base

(b) **(i)** $2NaHCO_3(s) + 2HCl(aq) \rightarrow Na_2CO_3(aq) + CO_2(g) + H_2O(l)$

effervescence

solid being used up

increase in temperature/exothermic reaction [5]

(ii) HCO_3^- can act as a Brønsted–Lowry acid and base by giving and accepting protons:

as acid: $HCO_3^- \rightarrow H^+ + CO_3^{2-}$

as base: $HCO_3^- \rightarrow H^+ + H_2CO_3$ [4]

(c) Conductivity can be quantified, so more objective. [1]

The method described in (a) is qualitative only. [1]

13 **(a)** *Acid*: proton/H^+ donor and Base: proton/H^+ acceptor

Do not accept OH^- for base.

Weak base: (base/electrolyte) partially dissociated/ionized (in solution/water) **and** *Strong base*: (base/electrolyte assumed to be almost) completely/100% dissociated/ionized (in solution/water) / *OWTTE*

NH_3 / $CH_3CH_2NH_2$ [3]

Allow either name or formula or other suitable example.

(b) sulfurous acid/H_2SO_3

corrodes marble / limestone buildings / statues / leaching in soils / harms / kills plants

or

nitrous acid / HNO_2

corrodes marble / limestone buildings / statues / leaching in soils / harms / kills plants

or

carbonic acid / H_2CO_3

corrodes marble / limestone buildings / statues / acidification of lakes [2]

Do not allow oxides (e.g. CO_2).

Do not accept just corrodes or damages.

14 **(a)** $(K_w) = [H^+][OH^-]$ / $(K_w) = [H_3O^+][OH^-]$

Do not award mark if [ ] omitted or other brackets are used. [1]

(b) $[H^+]$ increases, $[OH^-]$ decreases but still some present (K_w constant) / $[OH^-]$ cannot go to zero as equilibrium present / $[OH^-] = \frac{K_w}{[H^+]}$, thus $[OH^-]$ cannot be zero *OWTTE*; [1]

(c) (changing *T* disturbs equilibrium) endothermic reaction / forward reaction favoured / equilibrium shifts to the right;

to use up (some of the) heat supplied;

K_w increases (as both $[H^+]$ and $[OH^-]$ increase); [3]

(d) (as $[H^+]$ increases) pH decreases / pH < 7;

No mark for more acidic.

inverse relationship between pH and $[H^+]$ /

$\text{pH} = -\log[H^+]$ / $\text{pH} = \log_{10}\frac{1}{[H^+]}$; [2]

Accept $[H_3O^+]$ in place of $[H^+]$.

R3.2 Electron transfer reactions

Exercises

Q1 B

Q2 C

Q3 C

Q4 **(a)** CH_4: C = −4, H = +1

(b) $CuSO_4$: Cu = +2, S = +6, O = −2

(c) C_2H_4: H = +1, O = −2

(d) CO: C = +2, O = −2

(e) $K_2Cr_2O_7$: K = +1, Cr = +6, O = −2

(f) K_2CrO_4: K = +1, Cr = +6, O = −2

(g) H_2O_2: H = +1, O = −1

(h) C_4H_{10}: C = −2.5, H = +1

Q5 **(a)** $Sn^{2+}(aq) + 2Fe^{3+}(aq) \rightarrow Sn^{4+}(aq) + 2Fe^{2+}(aq)$

+2 +3 +4 +2

oxidation: Sn +2 → +4, so $Sn^{2+}(aq)$ is oxidized

reduction: Fe +3 → +2, so $Fe^{3+}(aq)$ is reduced

(b) $2FeCl_2(aq) + Cl_2(aq) \rightarrow 2FeCl_3(aq)$

+2 −1 0 +3 −1

oxidation: Fe +2 → +3, so $2FeCl_2(aq)$ is oxidized

reduction: Cl 0 → −1, so $Cl_2(aq)$ is reduced

(c) $2H^+(aq) + S_2O_3^{2-}(aq) \rightarrow S(s) + SO_2(g) + H_2O(l)$

+1 +2 −2 0 +4 −2 +1 −2

oxidation: S +2 → +4, so $S_2O_3^{2-}(aq)$ is oxidized

reduction: S +2 → 0, so $S_2O_3^{2-}(aq)$ is also reduced

(d) $2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$

+1 −1 +1 −2 0

oxidation: O −1 → 0, so $H_2O_2(aq)$ is oxidized

reduction: O −1 → −2, so $H_2O_2(aq)$ is also reduced

(e) $I_2(aq) + SO_3^{2-}(aq) + H_2O(l) \rightarrow 2I^-(aq) +$

0 +4 −2 +1 −2 −1

$SO_4^{2-}(aq) + 2H^+(aq)$

+6 −2 +1

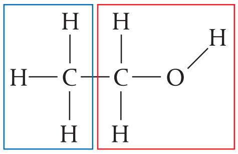
oxidation: S +4 → +6, so $SO_3^{2-}(aq)$ is oxidized

reduction: I 0 → −1, so $I_2(aq)$ is reduced

Q6 **(a)** assign H = +1, O = −2

Therefore, in ethanol,
$(6 \times +1) + -2 + (2 \times C) = 0$, so C = −2

This is an average value as the two carbon atoms are not equivalent.



We can divide the molecule into two regions, each with a zero charge.

For the carbon atom in the red region (CH_2OH),
$C + (3 \times +1) + -2 = 0$, so C = −1

For the carbon atom in the blue region (CH_3),
$C + (3 \times +1) = 0$, so C = −3

The average oxidation state $= \frac{-3 + -1}{2} = -2$ as expected.

(b) In CO_2, the oxidation state of C is +4

Ethanol is oxidized:

- the elements have combined with oxygen
- hydrogen has been removed
- the oxidation state of carbon has increased from −2 to +4
- the electron density of the carbon atom has reduced but no electron loss has led to production of positive ions.

Q7 **(a)** reducing agent = $H_2(g)$;
oxidizing agent = $Cl_2(g)$

(b) reducing agent = $Al(s)$;
oxidizing agent = $PbCl_2(s)$

(c) reducing agent = $KI(aq)$;
oxidizing agent = $Cl_2(aq)$

(d) reducing agent = $CH_4(g)$;
oxidizing agent = $O_2(g)$

Q8 A

Q9 C

Q10 **(a)** $Ca(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2(g)$
0 +1 +2 0
oxidation: $Ca(s) \rightarrow Ca^{2+}(aq) + 2e^-$
reduction: $2H^+(aq) + 2e^- \rightarrow H_2(g)$

(b) $2Fe^{2+}(aq) + Cl_2(aq) \rightarrow 2Fe^{3+}(aq) + 2Cl^-(aq)$
+2 0 +3 −1
oxidation: $2Fe^{2+}(aq) \rightarrow 2Fe^{3+}(aq) + 2e^-$
reduction: $Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$

(c) $Sn^{2+}(aq) + 2Fe^{3+}(aq) \rightarrow Sn^{4+}(aq) + 2Fe^{2+}(aq)$
+2 +3 +4 +2
oxidation: $Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2e^-$
reduction: $2Fe^{3+}(aq) + 2e^- \rightarrow 2Fe^{2+}(aq)$

(d) $Cl_2(aq) + 2Br^-(aq) \rightarrow 2Cl^-(aq) + Br_2(aq)$
0 −1 −1 0
oxidation: $2Br^-(aq) \rightarrow Br_2(aq) + 2e^-$
reduction: $Cl_2(aq) + 2e^- \rightarrow 2Cl^-(aq)$

Q11 **(a)** $2I^-(aq) + HSO_4^-(aq) + 3H^+(aq) \rightarrow I_2(aq) + SO_2(g) + 2H_2O(l)$

(b) $I_2(aq) + 5OCl^-(aq) + H_2O(l) \rightarrow 2IO_3^-(aq) + 5Cl^-(aq) + 2H^+(aq)$

(c) $2MnO_4^-(aq) + 5H_2SO_3(aq) \rightarrow 2Mn^{2+}(aq) + 5SO_4^{2-}(aq) + 3H_2O(l) + 4H^+(aq)$

Q12 **(a)** solution changes from purple to colourless

(b) $C_2O_4^{2-}(aq) \rightarrow 2CO_2(g) + 2e^-$

(c) $MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$

(d) $2MnO_4^-(aq) + 16H^+(aq) + 5C_2O_4^{2-}(aq) \rightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 10CO_2(g)$

(e) 6.16×10^{-3}

(f) 6.16×10^{-3}

(g) 24.7%

Q13 **(a)** 0.117%

(b) solution changes from orange to green

Q14 D

Q15 B

Q16 **(a)** $CuCl_2(aq) + Ag(s)$; no reaction
Cu is a more reactive metal than Ag

(b) $3Fe(NO_3)_2(aq) + 2Al(s) \rightarrow 2Al(NO_3)_3(aq) + 3Fe(s)$
Al is a more reactive metal than Fe, so is able to reduce Fe^{2+}

(c) $2NaI(aq) + Br_2(aq) \rightarrow 2NaBr(aq) + I_2(aq)$
Br is a more reactive non-metal than I, so is able to oxidize I^-

(d) $KCl(aq) + I_2(aq)$; no reaction
Cl is a more reactive non-metal than I

Q17 **(a)** $W > X > Y > Z$

(b) **(i)** no reaction
(ii) no reaction

Q18 C

Q19 B

Q20 **(a)** $2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$
$2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$

(b) $Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$
$Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

(c) $Zn(s) + 2CH_3CO_2H(aq) \rightarrow Zn(CH_3CO_2)_2 + H_2(g)$
$Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$

Q21 **(a)** $Zn(s) + SO_4^{2-}(conc.) + 4H^+(aq) \rightarrow Zn^{2+}(aq) + SO_2(g) + 2H_2O(l)$

(b) $NO_3^-(conc.) + 4Zn(s) + 10H^+(aq) \rightarrow NH_4^+(aq) + 4Zn^{2+}(aq) + 3H_2O(l)$

(c) $2NO_3^-(aq) + 3Cu(s) + 8H^+(aq) \rightarrow 2NO(aq) + 3Cu^{2+}(aq) + 4H_2O(l)$

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

Q22 A

Q23 A

Q24 **(a)** Zn is a more reactive metal than Fe, so is oxidized

Zn / Zn^{2+} is anode; $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

Fe / Fe^{2+} is cathode; $Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$

(b) Mg is a more reactive metal than Fe, so is oxidized

Mg / Mg^{2+} is the anode; $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$

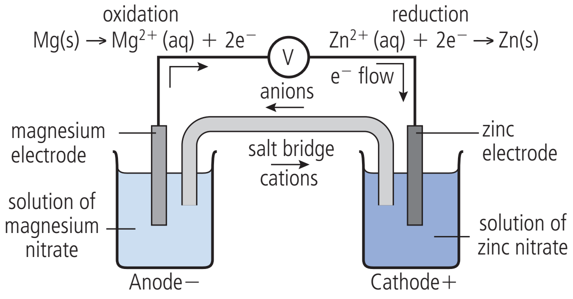
Fe / Fe^{2+} is the cathode; $Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$

(c) Mg is a more reactive metal than Cu, so is oxidized

Mg / Mg^{2+} is the anode; $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$

Cu / Cu^{2+} is the cathode; $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

Q25 **(a)** Mg is a more reactive metal than Zn so will be oxidized in the voltaic cell.



(b) $Mg(s) \mid Mg^{2+}(aq) \mid\mid Zn^{2+}(aq) \mid Zn(s)$

Q26 Fe is more reactive than Cu and so will reduce Cu^{2+} ions in solution.

The iron spatula would slowly dissolve as it is oxidized to Fe^{2+} ions. Copper metal (red/brown) would precipitate as Cu^{2+} ions are reduced. The blue colour of the solution would fade, as Cu^{2+} ions are removed.

Q27 **(a)** Oxidation number increases from 0 to +2. Pb(s) is oxidized.

(b) $Pb(s) + SO_4^{2-}(aq) \rightarrow PbSO_4(s) + 2e^-$

(c) $PbSO_4$ is insoluble: the Pb^{2+} ions do not disperse into solution.

(d) Advantage: delivers large amounts of energy over short periods; rechargeable.

Disadvantage: heavy mass; lead and sulfuric acid could cause pollution.

Q28 **(a)** $CH_3OH(aq) + H_2O(l) \rightarrow CO_2(g) + 6H^+(aq) + 6e^-$

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

(c) Advantage: does not need to be recharged, less hazardous if broken/exposed to the environment, weighs less for same energy output

Disadvantage: more expensive, needs constant supply of fuel, methanol fuel cells produce carbon dioxide, does not produce high current

Q29 **(a)** $H_2(g) + 2OH^-(aq) \rightarrow 2H_2O(l) + 2e^-$

$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$

(b) Fuel cells (and an electric motor) are more efficient as less waste heat produced and more chemical energy converted to useful electrical energy. They may also use a renewable energy source, which is more efficient.

(c) All convert chemical energy **directly** into electrical energy.

Q30 D

Q31 B

Q32 D

Q33 **(a)** at anode: $2Br^-(l) \rightarrow Br_2(l) + 2e^-$

at cathode: $2K^+(l) + 2e^- \rightarrow 2K(l)$

(b) at anode: $2F^-(l) \rightarrow F_2(g) + 2e^-$

at cathode: $Mg^{2+}(l) + 2e^- \rightarrow Mg(l)$

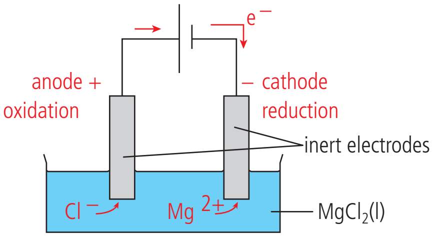
(c) at anode: $S^{2-}(l) \rightarrow S(l) + 2e^-$

at cathode: $Zn^{2+}(l) + 2e^- \rightarrow Zn(l)$

(d) at anode: $2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$

at cathode: $Ca^{2+}(l) + 2e^- \rightarrow Ca(l)$

Q34 **(a)**



(b) anode: $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

cathode: $Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$

overall: $Mg^{2+}(aq) + 2Cl^-(aq) \rightarrow Mg(s) + Cl_2(g)$

Q35 C

Q36 C

Q37 B

Q38 $C_4H_9OH(l) + 6O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)$

$C_5H_{11}OH(l) + 7.5O_2(g) \rightarrow 5CO_2(g) + 6H_2O(g)$

Q39 **(a)** butanone; orange → green

(b) methanal; orange → green

(c) no reaction; no colour change

Q40 B

Q41 C

Q42 **(a)** Use $LiAlH_4$ in dry ether and heat. The acid is reduced first to the aldehyde and then to the alcohol.

$$CH_3CH_2COOH \xrightarrow{+H} CH_3CH_2CH_2OH$$

(b) Ethanal is heated with $NaBH_4(aq)$.

$$CH_3CHO \xrightarrow{+H} CH_3CH_2OH$$

Q43 $CH_3CHO \rightarrow CH_3CH_2OH$

Average oxidation state of C in ethanal $= \frac{(-4+2)}{2}$

$= \frac{-2}{2} = -1$

Average oxidation state of C in ethanol $= \frac{(-6+2)}{2}$

$= \frac{-4}{2} = -2$

The carbon atoms decrease their average oxidation state by 1, so 2 electrons are needed:

$CH_3CHO + 2e^- \rightarrow CH_3CH_2OH$

Add $2H^+$ to balance the charges:

$CH_3CHO + 2H^+ + 2e^- \rightarrow CH_3CH_3OH$

Q44 Hydrogenation (of vegetable oils) with nickel catalyst.

Oils are liquid as they contain unsaturated fatty acid chains with C=C.

Margarines are solids as they contain saturated fatty acid chains with C–C.

Double bond has restricted rotation which gives molecule kinks which restricts packing of molecules and reduces intermolecular forces.

Saturated compounds can pack more effectively and have stronger intermolecular forces.

Q45 **(a)** $CH_3COCH_3 \rightarrow CH_3CH(OH)CH_3$

(b) $CH_3CO_2H \rightarrow CH_3CH_2OH$

(c) $CH_2CHCl \rightarrow CH_3CH_2Cl$

(d) $C_3H_4 \rightarrow C_3H_8$

Challenge yourself

1 assign H = +1 and O = −2, with C variable

For $C_6H_{12}O_6$, $6 \times C + (12 \times +1) + (6 \times -2) = 0$, so C = 0

$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$

$+4\ -2 \qquad +1\ -2 \qquad 0\ +1\ -2 \qquad 0$

oxidation: O −2 → 0, so $6H_2O(l)$ is oxidized

reduction: C +4 → 0, so $CO_2(g)$ is reduced

2 H_2O_2: H = +1, O = −1

The oxidation state for oxygen is −1, which is halfway between 0 (element) and −2 (usual oxidation state in compounds). The oxygen can be oxidized (to 0) or reduced (to −2). It will more easily be reduced from −1 to −2 as it is a very electronegative element, and so acts mainly as an oxidizing agent.

3 $3I_2(aq) + 6KOH(aq) \rightarrow 5KI(aq) + KIO_3(aq) + 3H_2O(l)$

I changes from 0 (in I_2) to −1 (in KI), which is reduction.

I changes from 0 (in I_2) to +5 (in KIO_3), which is oxidation.

These changes occur simultaneously.

4

Most active								**Least active**
Mg	Al	C	Zn	Fe	Pb	H_2	Cu	Ag

Carbon is more reactive than zinc and iron as it can remove oxygen from their oxides.

Carbon is less reactive than aluminium as electricity is needed for the extraction of aluminium.

Hydrogen is more reactive than copper, as copper metal does not react with dilute acids, but less reactive than lead.

5 At the negative electrode (anode); $H_2(g)$ is oxidized:

$H_2(g) + 2OH^-(aq) \rightarrow 2H_2O(l) + 2e^-$

At the positive electrode (cathode) $O_2(g)$ is reduced:

$2H_2O(l) + O_2(g) + 4e^- \rightarrow 4OH^-$

6 At the anode, oxide ions are oxidized:

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

The oxygen is produced at the carbon anode at the operating temperature of around 800 °C, so the carbon reacts.

7 Aluminium oxide has some covalent character, as aluminium has quite a high electronegativity (1.6).

8 Oxidation state of hydrogen in both compounds = +1

In propane (8 hydrogens) the average oxidation state of C = $\frac{-8}{3}$

In propene (6 hydrogens) the average oxidation state of C = $\frac{-6}{3} = -2$

Combustion of propane involves a bigger change of oxidation state and more oxygen is needed.

9 Consider first the oxidation reaction:
ethanol → ethanal

The average oxidation state of C in ethanol (C_2H_6O)
$= \frac{(-6+2)}{2} = -2$

The average oxidation state of C in ethanal (C_2H_4O)
$= \frac{(-4+2)}{2} = -1$

As both carbon atoms increase their oxidation state by +1 (from −2 to −1), two electrons are released:

$C_2H_5OH \rightarrow CH_3CHO + 2e^-$

Add $2H^+$ to balance the charges:

$C_2H_5OH \rightarrow CH_3CHO + 2e^- + 2H^+$

Reduction: $Cr_2O_7^{2-} + 6e^- + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O$

To deduce the overall reaction, balance the number of electrons in the half reactions, so need 3× the ethanol/ethanal equation:

$Cr_2O_7^{2-} + 6e^- + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O$

$3C_2H_5OH \rightarrow 3CH_3CHO + 6e^- + 6H^+$

add and simplify:

$Cr_2O_7^{2-} + 6e^- + 14H^+ + 3C_2H_5OH \rightarrow 2Cr^{3+} + 7H_2O + 3CH_3CHO + 6e^- + 6H^+$

$Cr_2O_7^{2-} + 8H^+ + 3C_2H_5OH \rightarrow 2Cr^{3+} + 7H_2O + 3CH_3CHO$

10 $CHO \rightarrow CH_2OH$

The average oxidation state of C in CHO = (−1 + 2) = +1

The average oxidation state of C in CH_2OH l = (−3 +2) = −1

As the carbon atom decreases the oxidation state by 2 : 2 electrons are released.

$C_2H_5OH \rightarrow CH_3CHO + 2e^-$

$C_2H_5OH \rightarrow CH_3CHO + 2e^- + 2H^+$

Practice questions

1 A

2 B

3 C

4 D

5 A

6 D

7 A

8 A

9 A

10 **(a)** 1 : 2 [1]

Accept 2 Fe^{3+} : 1 Fe^{2+}

*Do **not** accept 2 : 1 only*

(b) $H_2O_2(aq) + 2H^+(aq) + 2e^- \rightarrow 2H_2O(l)$ [1]

11 **(a)** salt bridge [1]

movement of ions ***OR*** balance charge [1]

*Do **not** accept "to complete circuit" unless ion movement is mentioned for M2.*

(b) *Positive electrode (cathode):* $Ag^+(aq) + e^- \rightarrow Ag(s)$ [1]

Negative electrode (anode): $Mg(s) \rightarrow Mg^{2+}(aq) + 2^{e-}$ [1]

*Award **[1 max]** if correct equations given at wrong electrodes.*

(c) in external wire from left to right [1]

12 **(a)** An electrolytic cell converts electrical energy to chemical energy and a voltaic cell converts chemical energy to electrical energy / an electrolytic cell uses electricity to carry out a (redox) chemical reaction and a voltaic cell uses a (redox) chemical reaction to produce electricity / an electrolytic cell requires a power supply and a voltaic cell does not.

An electrolytic cell involves a non-spontaneous (redox) reaction and a voltaic cell involves a spontaneous (redox) reaction.

In an electrolytic cell, the cathode is negative and the anode is positive and vice versa for a voltaic cell.

A voltaic cell has two separate solutions and an electrolytic cell has one solution / a voltaic cell has salt bridge and an electrolytic cell has no salt bridge.

In an electrolytic cell, oxidation occurs at the positive electrode/anode and in a voltaic cell, oxidation occurs at the negative electrode/anode and vice versa. [2 max]

(b) (solid) ions in a lattice / ions cannot move;

(molten) ions mobile / ions free to move [2]

(c) Reduction occurs at the cathode / negative electrode and oxidation occurs at the anode / positive electrode

Cathode / negative electrode: $Na^+ + e^- \rightarrow Na$

Anode / positive electrode: $2Cl \rightarrow Cl_2 + 2e^-$/

$Cl^- \rightarrow \frac{1}{2}Cl_2 + e^-$

*Award **(1 max)** if the two electrodes are not labelled/ labelled incorrectly for the two half-equations.*

Overall cell reaction:

$Na^+(l) + Cl^-(l) \rightarrow Na(l) + \frac{1}{2}Cl_2(g)$ [5]

*Award **(1)** for correct equation and **(1)** for correct state symbols.*

Allow NaCl(l) instead of $Na^+(l)$ and $Cl^-(l)$.

(d) Al does not corrode / rust; Al is less dense / better conductor / more malleable [1]

Accept Al is lighter (metal compared to Fe).

Accept converse argument.

13 **(a)** cathode: $2H_2O + O_2 + 4e^- \rightarrow 4OH^-$

anode: $2H_2 + 4OH^- \rightarrow 4H_2O + 4e^-$

If both equations given but at wrong electrodes award (1). [2]

(b) cathode: nickel hydroxide / $Ni(OH)_2$

anode: cadmium hydroxide / $Cd(OH)_2$

cell equation:

$Cd + 2H_2O + 2NiO(OH) \rightarrow Cd(OH)_2 + 2Ni(OH)_2$ [3]

14 **(a)** contains no lithium / metal / uses lithium salt in an organic solvent (as electrolyte)

involves movement of lithium ions (between electrodes) [2]

(b) Anode:

$LiC_6 \rightarrow Li^+ + 6C + e^-$ / Li^+ ions dissociate from anode (and migrate to cathode)

Cathode:

$Li^+ + e^- + MnO_2 \rightarrow LiMnO_2$ / $Li^+ + e^- + CoO_2 \rightarrow LiCoO_2$ / $Li^+ + e^- + FePO_4 \rightarrow LiFePO_4$ / $Li^+ + e^- + NiO_2 \rightarrow LiNiO_2$ / Li^+ ions are inserted into metal oxide / phosphate (structure)

*Award **(1)** if electrodes are reversed.* [2]

(c) Similarity:

both convert chemical energy directly into electrical energy / both use spontaneous redox reactions (to produce energy) / both are electrochemical cells / voltaic cells / galvanic cells

Difference:

fuel cells are energy conversion devices and rechargeable batteries are energy storage devices / fuel cells require constant supply of reactants and batteries have stored chemical energy / provide power until stored

chemicals are used up / batteries can be recharged and fuel cells do not need recharging (have a continuous supply of fuel) / fuel cells are more expensive than rechargeable batteries / the reactions in a rechargeable battery are reversible and in a fuel cell are not [2]

15 **(a)** **(i)** melting point of the cryolite solution is much lower than the melting point of alumina/ Al_2O_3 / it lowers the melting point of the mixture / cell operates at lower temperature [1]

Allow lowers melting point or lowers melting point of aluminium oxide.

Do not allow lowers melting point of aluminium.

(ii) positive electrode: $2O^{2-} \rightarrow O_2 + 4e^-$ /

$O^{2-} \rightarrow \frac{1}{2}O_2 + 2e^-$

negative electrode: $Al^{3+} + 3e^- \rightarrow Al$ [2]

*Award **(1)** for correct equations but wrong electrodes.*

Allow e instead of e^-.

(b) use of fossil fuels (to provide energy)

oxidation of the (graphite) positive electrode / anode [2]

16 **(a)** Al is more reactive than Fe / Al is higher than Fe in the reactivity series / it is harder to reduce aluminium ores compared to iron ores / Fe^{3+} is a better oxidizing agent than Al^{3+} / *OWTTE* [2]

(b) **(i)** $Fe_3O_4 + 4CO \rightarrow 3Fe + 4CO_2$

(ii) $Fe_3O_4 + 4H_2 \rightarrow 3Fe + 4H_2O$ [2]

17 **(a)** Positive electrode: graphite/carbon

negative electrode: graphite / carbon (on a steel liner) [2]

(b) much less energy required to recycle than to produce Al from ore / *OWTTE*

less production of CO_2/greenhouse gases (graphite used in the electrolysis is converted into CO_2) / the more that is recycled the less there will be in landfill sites / *OWTTE* [2]

18 **(a)** **(i)** copper: 0 to +2 / increases by 2 / +2 / 2+

Allow zero/nought for 0.

nitrogen: +5 to +4 / decreases by 1 / −1 / 1− [2]

Penalize missing + sign or incorrect notation such as 2+, 2^+ or II, once only.

(ii) nitric acid / HNO_3 / NO_3^- / nitrate [1]

(b) **(i)** 0.100×0.0285

2.85×10^{-3} (mol) [2]

*Award **(2)** for correct final answer.*

(ii) 2.85×10^{-3} (mol) [1]

(iii) $(63.55 \times 2.85 \times 10^{-3}) = 0.181$ g [1]

Allow use of 63.5.

(iv) $\left(\frac{0.181}{0.456} \times 100 =\right)$ 39.7% [1]

(v) $\left(\frac{44.2 - 39.7}{44.2} \times 100 =\right)$ 10% / 10.1% [1]

Allow 11.3%, i.e. percentage obtained in (iv) is used to divide instead of 44.2%.

R3.3 Electron sharing reactions

Exercises

Q1 A

Q2 B

Q3 D

Q4 D

Q5 D

Q6

Q7 Thermolytic fission occurs when a bond breaks due to the addition of heat. Photolytic fission occurs when a bond breaks after the absorption of light energy (a photon).

Q8 The three stages of a radical reaction are initiation, propagation and termination.

Initiation: Br–Br $\xrightarrow{UV}$ Br• + Br•

Propagation: CH_4 + Br• → CH_3• + HBr

Termination: CH_3• + Br• → CH_3Br

Q9 **(a)** The reaction proceeds through a radical mechanism.

Initiation: Br–Br $\xrightarrow{UV}$ Br• + Br•

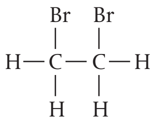
Propagation: C_2H_6 + Br• → C_2H_5• + HBr

Termination: C_2H_5• + Br• → C_2H_5Br

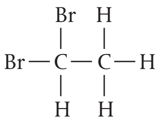
(b) Butane can be formed if two ethyl radicals formed in the propagation steps combine:

C_2H_5• + C_2H_5• → C_4H_{10}

(c)



1,2-dibromoethane

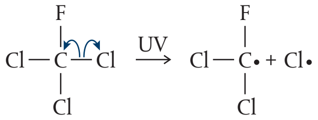


1,1-dibromoethane

Q10 **(a)** The bromine water would be decolourized as radical substitution reactions require UV light to initiate the reaction.

(b) The bromine water would not be decolourized as there is no UV light to initiate the reaction.

Q11 **(a)** The absorption of UV light causes homolytic fission of a C–Cl bond and the formation of CCl_2F• and Cl• radicals



(b) Cl•(g) + O_3(g) → O_2(g) + ClO•(g)

ClO•(g) + O_3(g) → O_2(g) + Cl•(g)

A chlorine radical reacts in the first step but another is formed in the second step, which makes the chlorine radical a catalyst.

Challenge yourself

1 For the C–F bond:

bond energy per molecule = $\dfrac{492 \times 10^3\,\text{J mol}^{-1}}{6.02 \times 10^{23}\,\text{mol}^{-1}}$

= $8.17^{28} \times 10^{-19}$ J

Need to use the equation

$$\lambda = \frac{hc}{E} =$$

$$\frac{6.63 \times 10^{-34}\,\text{J s}^{-1} \times 3.00 \times 10^8\,\text{m s}^{-1}}{8.17^{28} \times 10^{-19}}\text{J} = 2.43 \times 10^{-7}\,\text{m}$$

= 243 nm

For the C–Cl bond:

bond energy per molecule = $\dfrac{324 \times 10^3\,\text{J mol}^{-1}}{6.02 \times 10^{23}\,\text{mol}^{-1}}$

= 5.38206×10^{-19} J

$$\lambda = \frac{hc}{E} = \frac{6.63 \times 10^{-34}\,\text{J s} \times 3.00 \times 10^8\,\text{m s}^{-1}}{5.3821 \times 10^{-19}\,\text{J}}$$

= 3.70×10^{-7} m

= 370 nm

Practice questions

1 A

2 D

3 **(a)** HBr [1]

(b) The reaction requires UV light to break the Br–Br bond and initiate the reaction. [1]

(c) Initiation: Br–Br $\xrightarrow{UV}$ Br• + Br• [1]

Propagation: C_5H_{12} + Br• → C_5H_{11}• + HBr [1]

Termination: C_5H_{11}• + Br• → $C_5H_{11}Br$ [1]

(d) **(i)** The three products are structural isomers. [1]

(ii) $CH_2BrCH_2CH_2CH_2CH_3$, 1-bromopentane [1]

$CH_3CHBrCH_2CH_2CH_3$, 2-bromopentane [1]

$CH_3CH_2CHBrCH_2CH_3$, 3-bromopentane [1]

4 **(a)** Radical substitution [1]

(b) A radical substitution mechanism has three stages:

Initiation: Cl–Cl $\xrightarrow{UV}$ Cl• + Cl• [1]

Propagation: C_3H_8 + Cl• → C_3H_7• + HCl [1]

Termination: C_3H_7• + Cl• → C_3H_7Cl [1]

(c) $CH_2ClCH_2CH_3$, 1-chloropropane [1]

$CH_3CHClCH_3$, 2-chloropropane [1]

(d) Propane is regarded as kinetically stable as most of its reactions have high activation reactions and it won't react at regular temperatures. The products of these reactions have a lower enthalpy than propane so it is thermodynamically unstable compared to these products. [2]

5 **(a)** The reaction occurs through a radical mechanism which has three stages;

Initiation: Cl–Cl $\xrightarrow{UV}$ Cl• + Cl•

Propagation: C_2H_6 + Cl• → C_2H_5• + HCl

Termination: C_2H_5• + Cl• → C_2H_5Cl

(b) Small amounts of butane could be formed in the reaction if two ethyl radicals, formed in the propagation steps, combine in a termination step:

C_2H_5• + C_2H_5• → C_4H_{10} [1]

R3.4 Electron-pair sharing reactions

Exercises

Q1 A

Q2 C

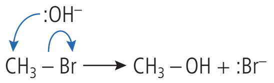
Q3 D

Q4 Nucleophilic substitution reactions occur when a nucleophile replaces a substituent, such as a halogen, in an organic compound.

In the reaction between NaOH and chloroethane, OH^- acts as the nucleophile and it substitutes the chlorine atom in chloroethane. (:Cl^- is the leaving group.)

$CH_3CH_2Cl + OH^- \rightarrow CH_3CH_2OH + Cl^-$

Q5



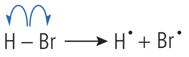
Q6 Section 12 in the data booklet lists the strengths of covalent bonds. From this table the bond strengths for the carbon halogen bonds can be compared.

Bond	C–F	C–Cl	C–Br	C–I
Bond enthalpy / kJ mol^{-1}	492	324	285	228

The bond that is the easiest to break will generate the best leaving group. The weakest bond is C–I so I^- will be the best leaving group.

Q7 B

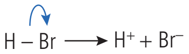
Q8 Homolytic fission:



$H-Br \longrightarrow H^{\bullet} + Br^{\bullet}$

The one-headed arrows (fish hooks) show that the bond breaks with one electron going to each atom, to form two radicals.

Heterolytic fission:



$H-Br \longrightarrow H^{+} + Br^{-}$

The two-headed arrow show that the bond breaks with the pair of electrons going to the more electronegative atom, to form a positive ion and a negative ion.

Q9 B

Q10 D

Q11 D

Q12 **(a)** This is an addition reaction with HBr: an alkene is converted to a halogenoalkane.

$CH_3CH_2CHBrCH_3$, 2-bromobutane

(b) With conc. H_2SO_4 followed by dilution with water, –H and –OH add across the alkene double bond, forming the alcohol.

CH_3CH_2OH, ethanol

(c) This is an addition reaction with bromine: an alkene is converted to an dibromoalkane.

$CH_3CHBrCH_2Br$, 1,2-dibromopropane

Q13 Bromine water readily reacts with alkenes but will only react with alkanes in the presence of UV light. If no UV light sources are present then:

adding bromine water to an alkene results in an instantaneous colour change (brown to colourless)

adding bromine water to an alkane results in no colour change.

Practice questions

1 D

2 A

3 C

4 C

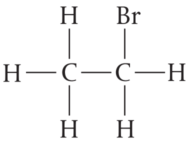
5 B

6 B

7 D

8 **(a)** Electrophiles accept electron pairs (from nucleophiles)
Nucleophiles donate electron pairs (to electrophiles)
1 mark for each

(b)



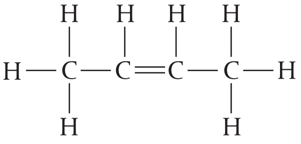
$$\begin{array}{ccccccc} & & H & & Br & & \\ & & | & & | & & \\ H & - & C & - & C & - & H \\ & & | & & | & & \\ & & H & & H & & \end{array}$$

bromoethane
1 mark for structure
1 mark for name

9 **(a)** Step 1: Addition
Step 2: Nucleophilic substitution
1 mark for each

(b) Coordination bond [1]

10 **(a)**



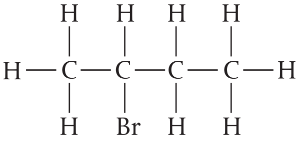
$$\begin{array}{ccccccccccc} & & H & & H & & H & & H & & \\ & & | & & | & & | & & | & & \\ H & - & C & - & C & = & C & - & C & - & H \\ & & | & & & & & & | & & \\ & & H & & & & & & H & & \end{array}$$

1 mark for structure

(b) Concentrated H_2SO_4 and water [1]

(c) The colour of the solution would change from orange to clear as the orange coloured Br_2 reacts and colourless products are formed. [1]

(d)



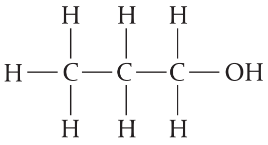
$$\begin{array}{ccccccccccc} & & H & & H & & H & & H & & \\ & & | & & | & & | & & | & & \\ H & - & C & - & C & - & C & - & C & - & H \\ & & | & & | & & | & & | & & \\ & & H & & Br & & H & & H & & \end{array}$$

2-bromobutane
1 mark for structure
1 mark for name

11 **(a)** $\mathrm{C_3H_7I + NaOH \rightarrow C_3H_7OH + NaI}$

$(\mathrm{C_3H_7I + OH^- \rightarrow C_3H_7OH + I^-})$

(b)



$\mathrm{H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-OH}$

propan-1-ol

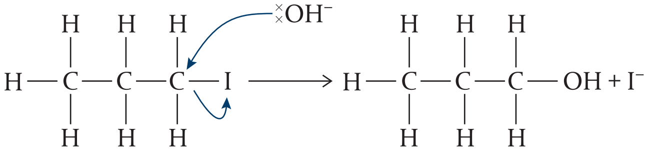
1 mark for structure AND name

(c) Nucleophile: $\mathrm{OH^-}$

Electrophile: $\mathrm{C_3H_7}$

1 mark for each

(d)



1 mark for double headed arrow from electron pair on OH⁻ to C

1 mark for double headed arrow from C–I bond to I

1 mark for correct products

(e) Heterolytic fission [1]