Answers to practice exam-style questions

THEME S

S1.1 Introduction to the particulate nature of matter

Paper 1 style questions

1	B	**7**	D	**13**	A
2	B	**8**	A	**14**	A
3	C	**9**	B	**15**	D
4	C	**10**	A	**16**	B
5	A	**11**	B		
6	D	**12**	C		

Paper 2 style questions

1

a simple distillation [1]

b filtration [1]

c sublimation [1]

d filtration, then evaporate to dryness [1]

2

a An element is a pure substance that cannot be broken down / decomposed / split up into two or more simpler substances by chemical processes/methods/means [1]
OR
a substance made of one type of atom [1]

b release of light and heat energy OR
a white solid is formed that is different in appearance from reactants (magnesium and sulfur) OR
provide colours (magnesium – silvery/grey and sulfur – yellow) [1]

c Any three from: [3]

- The magnesium and sulfur are chemically bonded in the white solid, but not in the flash mixture.
- The magnesium sulfide is white but the appearance of the flash mixture is yellow and silver.
- The white mixture will not react further if heated, but the flash mixture will react.
- The elements in the mixture could be separated by physical methods, but the elements in the white compound cannot be.
- The magnesium sulfide will have a fixed melting point, but the flash mixture will not.

3

a CO_2 (s) → CO_2 (g) [1]
no new chemical substance formed/change of state/both states of matter contain CO_2 molecules [1]

b molecule [1]
regular/repeating arrangement of molecules/particles (in a lattice) [1]

c weak intermolecular forces (operating between molecules) [1]

d 195 K [1]
25 °C [1]

e endothermic [1]
small amount of heat energy needed to separate molecules/overcome intermolecular forces [1]

f heat energy needed to separate molecules/overcome intermolecular forces during [1]
rapid expansion and evaporation results in rapid cooling [1]
so that the temperature of the carbon dioxide falls below its freezing point [1]

g large average separation distances between gaseous carbon dioxide molecules (between collisions) due to very weak intermolecular forces that the molecules now have and hence are able to become a gas [1]
much shorter distances between closely packed liquid carbon dioxide molecules due to the particles at this temperature not having sufficient energy to overcome the intermolecular forces [1]

h Any two from: [2]

- Dry ice has a fixed shape; liquid carbon dioxide takes the shape of the container.
- Dry ice has a high density; liquid carbon dioxide has a lower density.
- Dry ice cannot flow; liquid carbon dioxide flows

i CO_2 (g) + $Ca(OH)_2$ (aq) → $CaCO_3$ (s) + H_2O (l) [1]

S1.2 The nuclear atom

Paper 1 style questions

1	D	**5**	A	**9**	C
2	C	**6**	B	**10**	D
3	A	**7**	B	**11**	C
4	C	**8**	B	**12**	C

Paper 2 style questions

1

a atom of same element/same number of protons but with different mass number [1]
b 23 protons; 20 electrons; 26 neutrons [1]
c 50 further from A_r value of 50.94; [1]
d carbon-12/${}^{12}C$ [1]
e density/rate of effusion or diffusion [1]

2

a 1.00015 [1]
b One proton (nucleus) [1]
One neutron (nucleus) [1]
One electron (main energy level, (first) shell, (1s) orbital [1]
c Boiling point. Melting point, diffusion rate, solvation ability, density [1]
i $N_2 + 3D_2 \rightarrow 2ND_3$ [1]
ii $2K + 2D_2O \rightarrow 2KOD + D_2$ [1]
d $14 + (3 \times 2) = 20$ [1]

3

a Mass spectrometer [1]
b average mass of an atom / isotopes [1]
relative to 1/12 of the mass of a ${}^{12}C$ atom [1]
c $\frac{(107\times100)+(109\times96)}{100+96}$ [1]
$= 108$ [1]
d The number of protons in the nucleus of an atom/nuclide [1]
e ${}^{107}{}_{47}Ag$ [1]
f Chemical properties depend on the number of (outer or valence) electrons AND isotopes have the same electron configuration/arrangement / same number of electrons [1]
g ${}^{15}_{7}N$
correct symbol, N [1]
correct mass number = 15 and atomic number = 7 [1]

S1.3 Electron configurations

Paper 1 style questions

1	D	**8**	C	**15**	C
2	D	**9**	C	**16**	D
3	B	**10**	C	**17**	C
4	A	**11**	C	**18**	B
5	C	**12**	D	**19**	B
6	D	**13**	C	**20**	D
7	B	**14**	A		

Paper 2 style questions

1

a Both arrows pointing from right to left, [1] for each arrow [2]

b Any transition from higher quantum level to $n = 2$. A specific transition must be given, such as $n = 3$ to $n = 2$ [1]

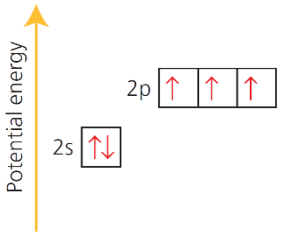
c <u>discrete</u> or <u>quantized</u> energy levels; the energy gap or difference between quantum shells narrows or becomes smaller; [1]
distance between adjacent quantum shells (main energy levels) becomes smaller as energy increases; the energy levels converge – the evidence is the convergence limit at higher frequencies [1]

d From energy level 5 to level 3: 1.55×10^{-19} J [1]
From energy level 6 to level 3: 1.82×10^{-19} J [1]
Energy difference is 2.70×10^{-20} J [1]

2

a Full: $1s^2\ 2s^2\ 2p^3$ [1]
Condensed: [He] $2s^2\ 2p^3$ [1]

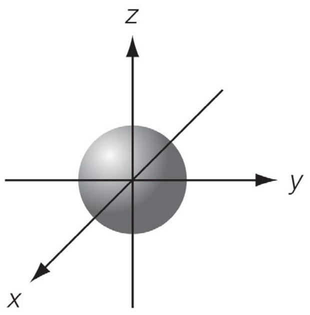
b



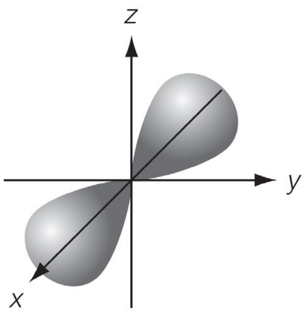
[1]

c first ionization energy of nitrogen greater than oxygen [1]
stronger inter-electronic repulsion between pair of electrons in one of the p orbitals in oxygen require less energy to remove [1]

d symmetrical shape of spherical s orbital [1]
dumbbell-shaped p orbital with two lobes of electron density along x-axis [1]



s orbital



p_x orbital

3

a Minimum amount of energy required to remove one mole of electrons from one mole of gaseous unipositive ions [1]
(M^+) to form one mole of gaseous dispositive ions (M^{2+}) [1]
$Ba^+ (g) \rightarrow Ba^{2+} (g) + e^-$ [1]

b A decrease down group 2 [1]
Valence electrons are progressively further away from the nucleus [1]
less energy is needed to remove them [1]

c Magnesium [1]

4

a The first ionization energy is the minimum energy required to remove one mole of electrons [1] from one mole of gaseous atoms to form one mole of unipositive gaseous ions (under standard thermodynamic conditions. [1]

b In a group 2 element (C) the electron (for the first ionization energy) is removed from a spin pair in an s sublevel [1]. For D, a group 13 element, the electron (for the first ionization energy) is removed from a 3 p sublevel further away from the nucleus. [1] The 3 p sublevel also experiences more shielding. Hence there is a decrease in first ionization energy when moving from C to D. [1]
The increase in nuclear charge accounts for the increase from E (np^2 to F np^3). In addition C cannot be in Group 15 (the other group after which there is a slight decrease) because there is a steady rise for the next 3 elements (D to E to F) indicating the filling of a p-sublevel. [1]

c As you move down group 2 the first ionization energies decrease. As you descend the group the nuclear charge increases due to the presence of additional protons [1], but the shielding effect progressively increases due to the presence of an extra electron shell as you move from one period to the next [1]. The effect of the extra protons is compensated for by the effect of the extra electrons. The only factor left is the additional distance between the outer electron and the nucleus. That lowers the first ionization energy. [1]

d C is in period 3 since in period 2 the group 2 element would have the higher ionization energy while in period 4 the group 2 element would have a lower ionization energy. [1] This can be accounted for by the progressively increasing average distance the valence electron is from the nucleus. [1]
C cannot be in the first period (He) as p-orbitals are being filled, it cannot be in the fourth period (Ca) as ($x + 8$) would not bring you back to Group 2 again. [1]

e There is a slight decrease from F to G due to the presence of a spin pair in the np sub-shell for G. [1]
The resulting electron–electron repulsion is greater than the effect of the increase in nuclear charge and reduces the ionization energy of G. [1]

5

a the minimum amount of energy needed to remove one mole of valence electrons from a mole of (neutral) atoms [1]
in the gaseous state [1]
$Mg(g) \rightarrow Mg^+(g) + e^-$ [1]

b electron removed (from aluminium) is in a p sublevel while that from magnesium is in an s sublevel [1]
the p sublevel is of higher energy / further from nucleus (than s sublevel) / experiences greater shielding from nucleus; [1]

c third electron more difficult to remove because it is taken from a positively charged ion; [1]
and from a (main) energy level closer to nucleus / from second energy level/$n = 2$ while the first electron comes from the third energy level/$n = 3$ [1]

d $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^7\ 4s^2$ [1]
4s [1]

S1.4 Models of the particulate nature of matter

Paper 1 style questions

1	D	8	A	15	C
2	C	9	C	16	A
3	D	10	A	17	D
4	B	11	B	18	A
5	B	12	D	19	C
6	A	13	D	20	D
7	D	14	B		

Paper 2 style questions

1

a) simplest integer ratio of <u>atoms of each element</u> in compound [1]

b) % oxygen = 39.5% [1]

Na 28.4/22.99 Cr 32.1/52.00 O 39.5/16.00 [1]

=1.24 = 0.617 = 2.47

(2:1:4) so empirical formula = Na_2CrO_4 [1]

2

a) amount of iodine atoms = 3.00 g/126.90 g mol^{-1} = 0.0236 mol [1]

b) amount of tin atoms = 0.70 g/118.71 g mol^{-1} = 0.0059 mol [1]

c) x = 0.0236/0.0059 = 4.0 [1]

d) Total % uncertainty = (0.333% + 2.86%) = 3.19% = 3% [1]

absolute uncertainty = (3/100) × 4.0 = 0.12

$x = 4.0 \pm 0.1$ (2 s.f.) [1]

e) SnI_2 = 372.51 g mol^{-1}. Sn = 118.71 g mol^{-1}

(118.71/372.51) × 100 = 31.87% [1]

3

amount of NH_3 = amount of HCl = $0.200 \times \frac{15.0}{1000} = 3.00 \times 10^{-3}$ mol [1]

amount of NH_2CONH_2 = ½ × mol of NH_3 = 1.50×10^{-3} mol [1]

amount of nitrogen = 2 × amount of fertilizer = $(2 \times 1.50 \times 10^{-3}) = 3.00 \times 10^{-3}$ mol [1]

mass of nitrogen = $(3.00 \times 10^{-3}$ mol × 14.01 g mol$^{-1})$ = 0.0420 g [1]

4

To make 0.100 mol dm^{-3} solution, pipette 25.00 cm^3 of the stock solution into a 250 cm^3 volumetric flask and make up to the mark with distilled/deionised water [1]

To make the 0.0100 mol dm^{-3} solution, pipette 25.00 cm^3 of the 0.100 mol dm^{-3} solution prepared earlier into a 250 cm^3 volumetric flask and make up to the mark with distilled/deionised water [1]

Mention of both pipette and volumetric flask required for 3rd mark. [1]

5

a) Avogadro's law: the number of particles in a gas is directly proportional to volume at the same temperature and pressure [1]

b) $C_2H_4\ (g) + 3O_2\ (g) \rightarrow 2CO_2\ (g) + 2H_2O\ (l)$ [1]

c) Gas volumes behave like amounts (in moles).
The ethene produces an amount of carbon dioxide twice that of ethene.
Hence volume of $CO_2 = 200\ cm^3$ [1]

S1.5 Ideal gases

Paper 1 style questions

1	A	8	C	15	D
2	A	9	B	16	D
3	D	10	D	17	C
4	A	11	B	18	A
5	A	12	C	19	D
6	D	13	A	20	D
7	D	14	D		

Paper 2 style questions

1

a) $\frac{V_1}{T_1} = \frac{V_2}{T_2}$; $\frac{955\text{ cm}^3}{278\text{ K}} = \frac{V_2}{303\text{ K}}$; [1]

V2 = 1041 cm3 [1]

The balloon will burst [1]

b) The increase in temperature increases the average kinetic energy of the nitrogen molecules and hence increases their average speed. [1]

They collide more frequently and energetically with each other and the walls of the balloon [1]

2

a) $pV = nRT$; $n = pV/RT$ [1]
$n = (100\text{ kPa} \times 10^3 \times 4.60)/(8.31\text{ J mol}^{-1}\text{ K}^{-1} \times 303\text{K}) = 183\text{ mol}$ [1]
Molar mass of NaO_2 = 54.99 g [1]
Mass of $NaO_2 = (2 \times 183\text{ mol} \times 54.99\text{ g}) = 2.01 \times 10^4\text{ g}$ [1]

b) So the submarine can remain submerged; removal of carbon dioxide which becomes toxic if its concentration is allowed to increase [1]

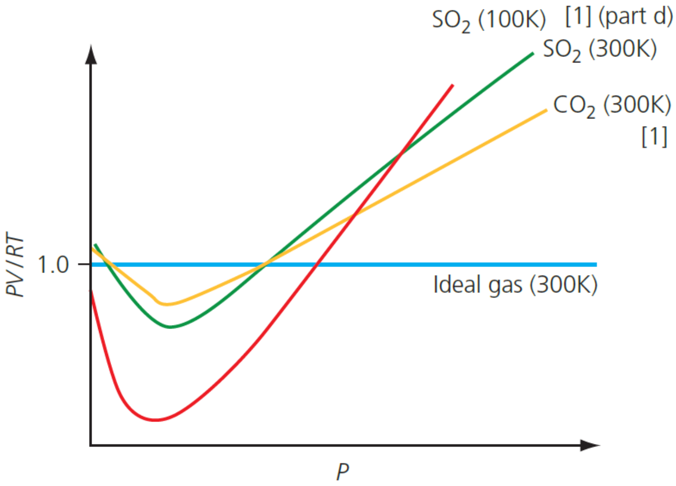
3

a) The particles have no forces (attractive or repulsive) acting between them; [1]
These particles do not take up any significant space, meaning their atomic/molecular volume is completely ignored (negligible) [1]

b) Real gases deviate most from ideal behaviour at high pressure and low temperature. [1]
At high pressure the volume of the container is decreased and the particles are closer together [1]
The volume occupied by the gas particles becomes significant compared to the volume of the container; the value of pV is larger than it should be if behaving ideally [1]
At low temperature the gas particles have less kinetic energy and the particles will tend to attract each other during collisions [1]
Intermolecular forces of attraction between the gas particles become significant/stronger [1]
The impact of a gas particle with the wall of the container is decreased; the value of pV is smaller than it should be if behaving ideally [1]

c) SO_2 deviates more from ideal gas behaviour than CO_2. SO_2 is a polar molecule (V-shape); CO_2 is a non-polar molecule (linear) [1]
Hence there are strong dipole-dipole forces between SO_2 molecules but only weaker London dispersion forces between CO_2 molecules [1]
Or
The molecular size of SO_2 is greater than CO_2 and it has a larger number of electrons. [1]
The larger, more diffuse electron cloud forms larger temporary dipoles and hence there are stronger London dispersion forces between SO_2 molecules than between CO_2 molecules [1]

d) At the lower temperature of 100 K, the SO_2 molecules have less kinetic energy and they move slower (on average) Hence, the intermolecular forces of attractions between SO_2 molecules become more significant [1]



4

a) $n(Ar) = 5.19\ g/39.95\ g\ mol^{-1} = 0.1299\ mol$ [1]
Using $pV = nRT$, $p(0.01 \times 0.3) = 0.1299 \times 8.31 \times 293$
$p = 105428$ Pa or 105.4 kPa [1]

b) Using $pV = nRT$,
$105428\ (V) = 0.13 \times 8.31 \times 323$
$V = 0.003309\ m^3$, [1]
height of gas = 0.003309/0.01 = 0.3309 m;
distance moved = (0.3309 m – 0.300 m) = 0.0309 m [1]

5 $M_r = mRT/pV = (0.425 \times 8.31 \times 398.5)/(52.0 \times 10^{-6} \times 101000) = 267.97 = 268$ (3sf) [1]

$(0.002/0.425 + 0.3/52.0 + 0.5/398.5 + 1000/101000) \times 268 = 5.797$ (4 sig. fig.) = 6 (same number of dp as the answer) [1]

M_r of $AlCl_3$: (268 ± 6) no units which indicates that $AlCl_3$ ($M_r = 133$) exists as a dimer (Al_2Cl_6) in the gas phase [1]

S2.1 The ionic model

Paper 1 style questions

1	A
2	B
3	C
4	D
5	D
6	A
7	A
8	B
9	D
10	C
11	A
12	B
13	D
14	D
15	D
16	B
17	A
18	B
19	B
20	B

Paper 2 style questions

1

a) It is an empirical formula because it shows the smallest integer ratio between the cations and anions (the actual number of ions present will vary with sample size [1]

b) $2Ag\ (s) + F_2\ (g) \rightarrow 2AgF\ (s)$ [1]

c) Synthesis/redox/reduction and oxidation/electron transfer [1]

d) Ionic/ionic bonds/ionic bonding [1]
Electrostatic attraction between cations and anions/oppositely charged ions [1]

e) Any four from: [4]

- High melting point
- High boiling point.
- Solid under standard conditions
- Brittle
- Electrically conducting when molten or in aqueous solution
- Non-conducting as a solid.

f) Minimum of 8 ions shown in a three-dimensional arrangement placed correctly. [1]
The charges are Ag^+ and F^- and must be shown and be in a 1:1 ratio (do not allow multiple charges on ions [1]

g) The lithium ion is small and relatively polarizing due to its charge density [1]
and the large bromide ion's valence electrons are polarizable [1]
There is thus a certain amount of covalent character in the bonding [1]

2

a)

i. Caesium fluoride has low volatility which is a consequence of having a relatively high melting and boiling point. This is due to the strong attractive electrostatic forces of attraction operating between oppositely charged ions in the lattice. [1]
Large amounts of thermal energy are required to break down the lattice [1]

ii. Caesium fluoride is a non-conductor in the solid state: the ions are not free to move when a voltage is applied.
It becomes an electrical conductor when melted or when an aqueous solution is formed. The electrical conductivity is due to the presence of mobile lithium and fluoride ions [2]

iii. Caesium fluoride is soluble in water (a polar solvent) because both ions are hydrated by the water molecules. The polar water molecules hydrate the caesium and fluoride ions via ion-dipole forces.

The attraction between water molecules and ions is energetically favourable to break up the lattice and release hydrated ions [2]

b) Enthalpy change when one mole of a solid ionic compound is separated into gaseous ions (at infinite distance) under standard conditions [1]

$CsF (s) \rightarrow Cs^+ (g) + F^- (g)$ [1]

3

a) Mg^{2+} and SO_4^{2-} [1]

Electrostatic forces of attraction between oppositely charged ions/anions and cations. (Do not accept ionic bonding/bonds) [1]

b) Sulfur and oxygen [1]

c) Aluminium [1]

+3/tripositive, Al^{3+} [1]

d) $Al_2(SO_4)_3$ [1]

4

a) $MgCl_2$ has a giant ionic structure while $AlCl_3$ has a simple/polymeric molecular structure.[1]

More energy is required to overcome the electrostatic forces of attraction between Mg^{2+} and Cl^- compared to instantaneous dipole-induced dipole attraction between $AlCl_3$ molecules. Hence $MgCl_2$ should have a higher melting point. [1]

b) Molten/fused liquid [1]

or in aqueous solution/dissolved in water [1]

Mobile ions present (that move in response to a potential difference) [1]

5

a) PbF_2, $PbCl_2$ and $PbBr_2$ have giant ionic structure with strong ionic bonding which involves electrostatic forces of attraction between oppositely charged ions. [1]

PbF_2, $PbCl_2$ and $PbBr_2$ have the same cation and the anions have the same charge. Since the ionic radius increases from fluoride to bromide, the interionic distance increases from PbF_2 to $PbBr_2$. Hence the strength of ionic bonding, lattice enthalpy and melting point decreases from PbF_2 to $PbBr_2$. [1]

b) Due to its high charge and small size, the lead(IV) ion, Pb^{4+}, has a high charge density and hence high polarizing power, which causes it to distort the electron cloud of chloride ions to such an extent that $PbCl_4$ is predominantly a polar covalent compound. [1]

Hence, $PbCl_4$, having a simple covalent/molecular structure, consists of discrete molecules held together by weak London (dispersion) forces, which requires a small amount of thermal energy to overcome. Thus, it has a low melting point. [1]

S2.2 The covalent model

Paper 1 style questions

1	A	**8**	D	**15**	A
2	D	**9**	A	**16**	D
3	D	**10**	A	**17**	C
4	B	**11**	D	**18**	D
5	B	**12**	B	**19**	C
6	B	**13**	A	**20**	A
7	C	**14**	B		

Paper 2 style questions

1

a) 1 mark for information in each cell below: [4]

	Electron domain geometry	Molecular geometry
Nitrogen	Tetrahedral	Trigonal pyramidal
Carbon	Trigonal planar	Trigonal planar

b) sp^2 [1]

c) sigma bonds: 9; pi bonds: 2 and lone pairs: 6 [3]

d) Ethanediamide has a greater molar mass/larger size and so has more electrons/greater electron density. [1]
Thus it has stronger/more extensive London (dispersion) forces / forms more hydrogen bonds per molecule (lone pairs on four atoms and also four hydrogen atoms). [1]

e) It can form hydrogen bonds with water molecules [1]

2

a) An increase for both melting and boiling points [1]

b) This is due to the increase in the number of polarizable electrons available for formation of temporary dipoles. [1]
The extent of the London (dispersion) forces attraction increases and hence the thermal energy required to separate the atoms increases [1]

c) Differences: 23; 221; 28 and 258 [1]
The perfluorocarbon molecules behave most like the noble gas of nearest molar mass [1]

d) The perfluorocarbons have lower boiling points than the corresponding hydrocarbons because of the high value of the electronegativity of fluorine, and hence much lower polarisability of the molecules, leading to weaker dispersion forces. [1]
The hydrocarbons have no strongly electronegative atoms so high polarisability, leading to stronger London (dispersion) forces and hence higher boiling points. Hydrocarbons are also larger, elongated molecules so they have a greater area of close contact. [2]

3

a) graphite; diamond (in this order) [2]

b) cubic boron nitride will be harder; hexagonal boron nitride will be softer [1]
there are weak intermolecular forces (London (dispersion forces) between the layers in hexagonal boron nitride; [1]
cubic boron nitride is held together by strong covalent bonds in a rigid three-dimensional lattice [1]

c) Graphene will be conducting since it has free / delocalised pi/ mobile electrons [1]
Carbon-60 will be a poor conductor, (although there is delocalization within each cage) it consists of separate/simple/discrete molecules/spheres/particles, (so no delocalization of electrons from one sphere to the next) or pi electrons are trapped within each molecule / sphere [1]

4 The C–H bond in $CHCl_3$ is relatively polarised due to the electron-withdrawing effect of the electronegative chlorine atoms. $CHCl_3$ has a relatively large molecular dipole moment with δ^+ H and δ^- Cl. [1]

CCl_4 has no dipole moment. Although each C-Cl bond is polar, the tetrahedral shape of the molecule leads to overall (vectorial) cancellation of these bond dipoles. [1]
The polar water molecules interact more strongly with the polar $CHCl_3$ molecules than with the non-polar CCl_4 molecules so the polar $CHCl_3$ is more soluble. (In addition, there is evidence that the hydrogen atoms in $CHCl_3$ can participate in hydrogen bonding.) [1]

5

a) Si_nH_{2n+2} and Si_6H_{14} [2]

b) There is an increase in the boiling and melting points of the silanes across the table from left to right [1]

c) trisilane Si_3H_8 or tetrasilane Si_4H_{10} [1]

d) The silanes are simple molecular substances [1]
London (dispersion) forces operate between the non-polar molecules. [1]
They increase with increase in molecule size/number of electrons/molar mass due to the increase in the size of the temporary or induced dipole [1]

e) $SiH_4\ (l) + 2O_2\ (g) \rightarrow SiO_2\ (s) + 2H_2O\ (l)$ [1]

f) Tetrahedral, 109.5 degrees/109 degrees/109 degrees 28’ [2]

g) The silicon atom is larger than the carbon atom/the carbon atom is smaller than the silicon atom. The silicon atom has an additional electron shell (main energy level) than the carbon atom/the carbon atom has one less electron shell (main energy level) than the silicon atom [1]
Weaker attraction for the silicon nucleus and shared electron pair in the silicon-hydrogen bond due to larger size of silicon atom with more diffuse orbitals; stronger attraction for the shared pair in the carbon-hydrogen bond [1]

h) In methane the difference is 0.4 and silane difference is 0.3/methane has a difference of 0.1 more than silane. Greater shielding in silicon/less shielding in carbon or less effective orbital overlap in silicon/more effective orbital overlap in carbon [1]
Hydrogen is more electronegative than silicon but less electronegative than carbon, or in methane, hydrogen will have a δ^+ charge but in silane hydrogen will have a δ^- charge, or the C-H bonds in methane are more polar (than the Si-H bonds in silane) [1]

i) Three-dimensional covalent network structure involving tetrahedrally bonded silicon/each silicon atom joined to four others by single (sigma) covalent bonds (except at surface) [1]

Shiny/lustrous surface, hard, brittle, poor electrical conductor/semi-conductor, poor thermal conductor, high melting point, high boiling point [1]

6

a) There is a stationary phase and a mobile phase. [1]
The components of the mixture dissolve, absorb, or partition to differing extents between the two phases. [1]

b) i. A [1]
ii. B [1]
iii. No, the two 'oranges' are different substances [1]
since they have different R_f values [1]
iv. R_f value = $\frac{9}{22}$ = 0.41 [1]

c) i. R_f value of compound 1 = $\frac{1.6}{12.6}$ = 0.13 [1]
R_f value of compound 2 = $\frac{9.1}{12.6}$ = 0.72 [1]
Compound 1 has a greater affinity for the thin layer than compound 2. Since the thin layer is composed of silica gel, compound 1 is more polar than compound 2. [1]
ii. TLC is faster than paper chromatography. The thin layer may be made from a variety of different solids. Hence, a range of mixtures can be separated by using different stationary and mobile phases [1]

7 Any two of: [1 each]

- planar/flat X-ray diffraction studies
- carbon to carbon bond lengths all equal
- carbon to carbons bonds intermediate in length between C–C and C=C
- all C–C–C bond angles equal

8

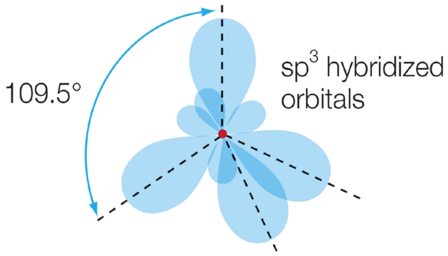
a) The Valence Shell Electron Pair Repulsion (VSEPR) is a theory that describes the prediction of the molecules' shape taking into account the number of electron domains which surround a central atom. [1]
Electrons around the central atom will experience repulsion with neighbouring pairs of bonding and non-bonding electrons [1]
The molecule or ion will spontaneously arrange itself in order to minimise this repulsion and reach the state with lowest energy [1]

b) electron domain geometry: tetrahedral [1]
molecular shape: trigonal pyramidal [1]
O–Cl–O bond angle: any value between 106° and 108° [1]

c) In A, the formal charge of Cl is 0 while O is -1 [1]
in B, the formal charge of Cl is -1 while O is 0 [1]
Structure A more likely as more electronegative oxygen carries the negative charge. [1]

9

a)



[1]

b) Three of the sp^3 hybridized orbitals of the carbon overlap head-on with the s orbitals of each of the three hydrogen atoms to form sigma bond [1]

The fourth sp^3 hybridized orbital of the carbon overlaps head-on with an sp^2 hybridised orbital of the carbon to also form a sigma bond [1]

S2.3 The metallic model

Paper 1 style questions

1	A
2	B
3	C
4	B
5	D
6	A
7	C
8	A
9	A
10	C
11	C
12	C
13	C
14	C
15	B
16	C

Paper 2 style questions

1

a) Solid metals consist of a regular lattice of cations [1]
formed by the loss of one or more delocalized valence electrons [1]
which form a delocalized cloud, gas or sea of electrons throughout the lattice. [1]
The bonding consists of electrostatic attraction of the cations for these delocalized electrons. [1]

b) The electrostatic attraction between cations and delocalized valence electrons is relatively strong [1]
and requires considerable heat /energy to overcome it/break them. [1]
The delocalized electrons in the solid [1]
and the delocalized valence electrons and cations in the liquid [1]
are free to move under an applied potential difference. [1]
Heat is conducted in a metal by the increased kinetic energy of the randomly moving electrons [1]
and by increased vibrations of the regularly close packed cations. [1]

2

Sodium, magnesium and aluminium are metallic and hence are good electrical conductors due to the presence of delocalized valence electrons. [1]

From sodium, magnesium to aluminium, the atomic and ionic size decrease and the number of valence electrons available for delocalization increases from one to three. Thus, the metallic bonding strength increases and hence the melting point increases since more heat energy is required to break or overcome the metallic bonding. In addition, with more valence electrons for delocalization, the electrical conductivity also increases. [1]

Silicon has a giant covalent structure with strong silicon-silicon single covalent bonds. A large amount of heat energy is required to break all these bonds and hence it has a high melting point. It is a semiconductor with a low electrical conductivity intermediate between that of an insulator and a conductor. There are no delocalized valence electrons. [1]

Phosphorus (P4), sulfur (S8), chlorine (Cl2) and argon (Ar) are simple molecular covalent substances and hence are held together in the solid state by London (dispersion) forces. A small amount of heat is required to break these intermolecular forces and hence their melting points are low. [1]

S8 has 128 electrons, P4 has 60 electrons, Cl2 has 34 electrons and Ar has only 18 electrons (ignoring any changes in molecular shapes). The strength or extent of intermolecular forces increases with molecular size (number of electrons available for polarization). Thus, sulfur (S8) has the highest melting point followed by phosphorus (P4), chlorine (Cl2) and then argon (Ar) with the lowest melting point. [1]

All of these substances are insulators and non-conductors when solids or liquids due to the absence of mobile charge carriers. They consist of molecules and no ions or delocalized electrons are present. [1]

3

a) positive ions with a charge +2/dipositive cations [1]
(attract) delocalized (valence) electrons (or sea of, or free, or mobile electrons) [1]

b) strontium and barium have similar ionic radii so they can fit into each other's lattices as an alloy cools down to form a solid solution [1]
all of the barium and strontium cations are attracted to the sea of delocalized electrons [1]

c) Correct reference to [position in group and therefore] size of cations/proximity/distance of electrons from the cations for first mark using any of the following arguments (or reverse for barium): [1]

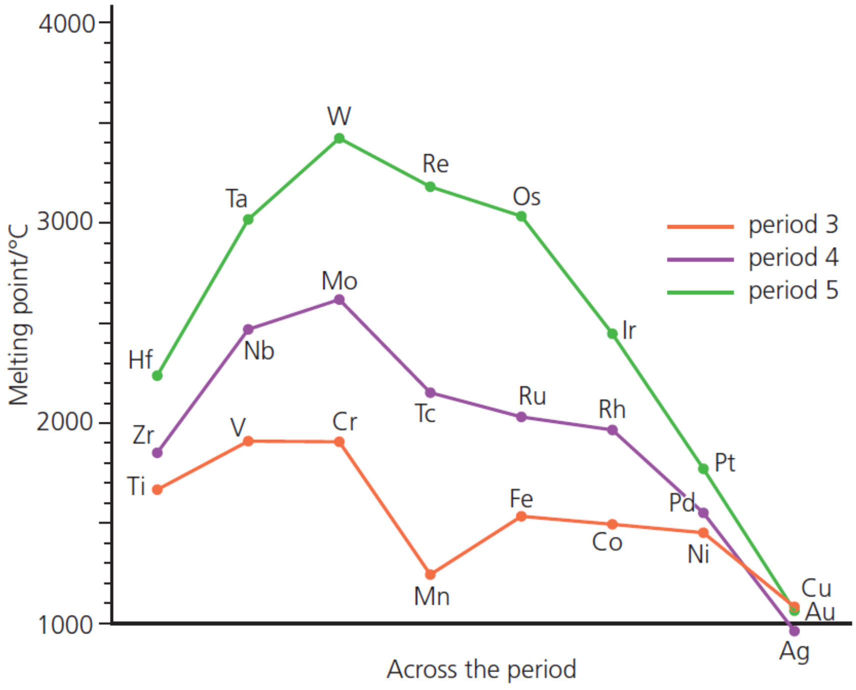
- strontium has delocalized/valence electrons closer to cations/positive ions/atoms/nucleus
- cations/positive ions/atoms are smaller/higher charge density
- cation/positive ion/atom has fewer (electron) shells/main energy levels

Relative strength of metallic bonding for second mark using any of the following arguments (or reverse for barium): [1]

- (strontium) has stronger attraction between the cations/positive ions/ and the delocalized valence electrons
- stronger metallic bonding

d) Chromium has a higher melting/boiling point than strontium [1]
Chromium has stronger metallic bonds as both the 3d and 4s electrons can be used in metallic bonding, hence more energy is required to overcome the stronger metallic bonds. Strontium uses only 5s electrons for metallic bonding [1]
or
Chromium has a greater density than calcium [1]
Chromium has a greater atomic mass but its atomic radius is smaller. Hence atomic volume is smaller. Since density = mass/volume, the density of chromium is greater than calcium [1]

4



a) [1] for correct plotting; [1] for correct labels and axes; [1] for key or labels of metals

b) Generally, down the groups, the melting point of metals increases except for the elements in the last group / except for the metals copper, Cu, gold, Au and silver, Ag. [1]

Across all periods, the melting point increases for the first three metals (elements) listed before it shows a decreasing trend for the next few metals. [1]

However, for periods 4 and 5, the 4th element has a lower melting point than expected. [1]

c) It is supported by the third row melting point data, as demonstrated by tungsten (W), which has the maximum number of d electrons in its period of the periodic table. [1]

It is not supported by first and second row melting point data, as demonstrated by manganese. [1]

S2.4 From models to materials

Paper 1 style questions

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

Paper 2 style questions

1

a) The measure of the ability of an atom in a covalent bond [1]
to attract a shared pair/bonding pair of electrons [1]

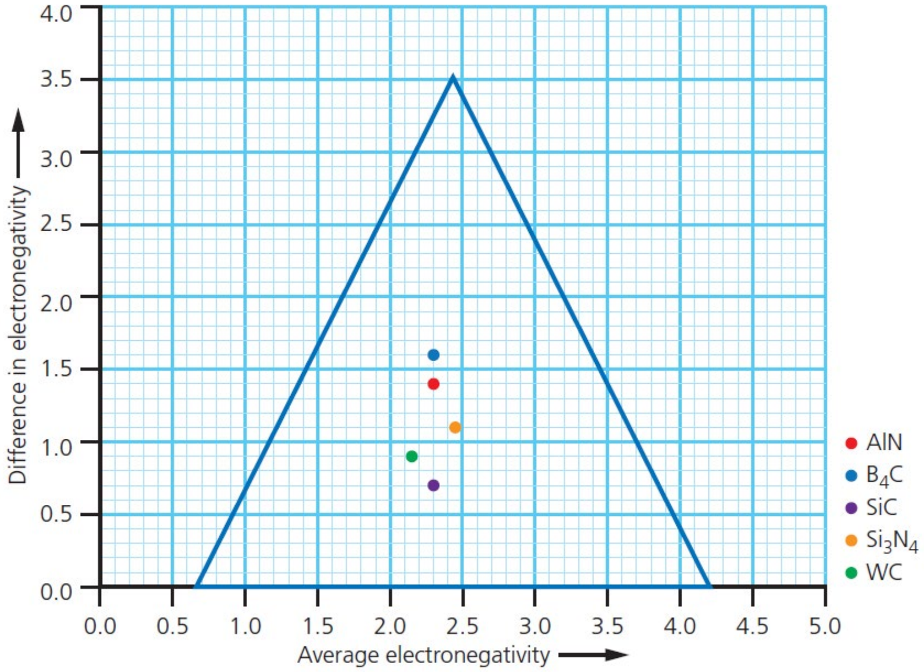
b) These elements are metals [1]
They lose their valence electrons which become delocalized and form a current when a potential difference is applied [1]

i. The more electronegative atom will gain electrons. High values of electronegativity tend to correlate with low values of electron affinities [1]

ii. The less electronegative atom will lose electrons. Low values of electronegativity tend to correlate with low values of ionization energy [1]

2

a) [1]



AlN is polar covalent (close to ionic border) [1]
B_4C is polar covalent [1]
SiC is polar covalent [1]
Si_3N_4 is polar covalent [1]
WC is polar covalent. [1]

b) SiC is more metallic [1]

c) It does not make correct prediction for elements that have multiple oxidation states, for example, $SnCl_2$ and $SnCl_4$ [1]

3

In the pure metal all the atoms/cations are the same size (radius/volume); suitable diagram of pure metal structure [1]

Atoms/cations arranged into a regular, repeating arrangement/lattice [1]

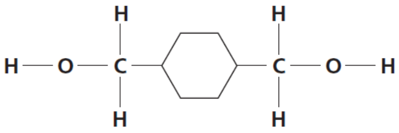
In an alloy the atoms/cations are of different sizes/suitable diagram of alloy structure [1]

Atomic lattice disrupted [1]

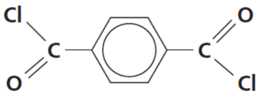
Atoms/cations in pure metals structure can slide across each other when force is applied to the metal when it is bent (accept an explanation based on dislocations) [1]

In an alloy the sliding of rows of metal atoms is prevented by different sized metals atoms/cations [1]

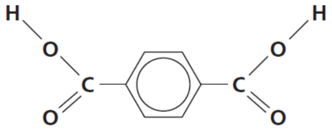
4 a) [2]



and either



or



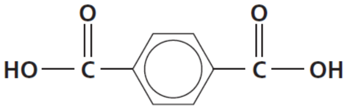
b) A small molecule, such as H_2O or HCl, is eliminated when the monomers join. [1]

c) Ester linkage/ester bond/-CO-O- [1]

5

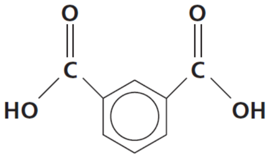
a) They contain benzene rings [1]

b)



and

[2]



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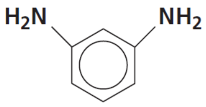


[2]

c) Hydrogen bonding/hydrogen bonds [1]

d) Nomex has a lower tensile strength than Kevlar because the chains cannot be aligned as closely [1]

for maximising hydrogen bonding interactions between adjacent chains [1]

e) The amide groups become protonated (at the nitrogens) due to the sulfuric acid acting as a proton donor; disrupting the hydrogen bonding [1]

The sulfuric acid molecules then solvate the protonated chains [1]

S3.1 The periodic table: classification

Paper 1 style questions

1	C	11	D	21	C
2	A	12	B	22	C
3	C	13	C	23	C
4	B	14	A	24	B
5	C	15	C	25	A
6	D	16	D	26	D
7	D	17	D	27	C
8	C	18	B	28	D
9	A	19	D	29	A
10	C	20	B	30	D

Paper 2 style questions

1 There is a decrease in ionic radii from Na^+ to Al^{3+}: all the ions have the electron arrangement of 2, 8 or $1s^2\ 2s^2\ 2p^6$ (that is they, are isoelectronic) [1]
There is a progressive increase in the nuclear charge due to the additional protons: the sodium ion has eleven protons, the magnesium ion has twelve electrons and the aluminium ion has thirteen protons. [1]
All of the electrons experience a higher effective nuclear charge and the electrons in the two shells (main energy levels) are pulled progressively closer towards the nucleus. [1]
There is a decrease ionic radii from P^{3-}, S^{2-} to Cl^-: all the ions have the electron arrangement of 2, 8, 8 or $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ (that is, they are isoelectronic) [1]
There is a progressive increase in the nuclear charge due to the additional protons: the phosphide ion has fifteen protons, the sulfide ion has sixteen electrons and the chloride ion has seventeen protons [1]
All of the electrons experience a higher effective nuclear charge and the electrons in the three shells (main energy levels) are pulled progressively closer towards the nucleus [1]

2

a) $8s^1$ (or [Rn] $8s^1$) [1]
b) metal – it is in group 1 and has one valence electron [1]
that can readily be lost to form a cation [1]
c) Z has the greatest number of electron shells [1]
the increase in nuclear charge is more than outweighed by the increase in shielding and distance of the valence electrons from the nucleus [1]
d) The melting point of Z would be expected to be the lowest in group 1. [1]
This is due to the expected weakest metallic bonding present in the solid. [1]
The weak metallic bonding is due to cations of low charge density with the greatest distance between the cations and delocalized valence electrons [1]
e) +1 [1]
f) basic [1]

3 X: Na [1]
Oxide of Y: SO_3 [1]
$H_2SO_4 + Na_2O \rightarrow Na_2SO_4 + 2H_2O$ [1]

4 Add (warm dilute) acid to the sample and check if powder reacts to form a soluble product.
If it undergoes acid/base reaction, it is magnesium oxide MgO. [1]
$MgO\ (s) + 2HNO_3\ (aq) \rightarrow Mg(NO_3)_2\ (aq) + H_2O\ (l)$.
Add hot, concentrated potassium hydroxide solution to the sample and check if powder reacts to form a soluble product. [1]
If it undergoes acid/base reaction, it is silicon dioxide, SiO_2.
$SiO_2\ (s) + 2KOH\ (aq) \rightarrow K_2SiO_3\ (aq) + H_2O\ (l)$ [1]

5

II Reduction or redox [1]

III ligand exchange/ligand replacement/ligand substitution [1]

Ti^{4+} in $[TiCl_6]^{2-}$ has empty 3d orbitals and hence d-d transitions are not possible [1]

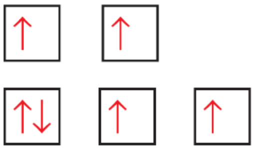
Ti^{3+} in $[TiCl_6]^{3-}$ has a partially filled 3d orbital, hence d-d transitions are possible [1]

Violet $[Ti(H_2O)_6]^{3+}$ absorbs in the yellow region while orange $[TiCl_6]^{3-}$ absorbs in the blue region. [1]

A blue colour has a shorter wavelength than yellow and thus indicates a larger energy gap between the two sets of d orbitals. [1]
Titanium can show variable oxidation states as the energy difference between the 3d and 4s orbitals of titanium is relatively small. [1]
With sufficient energy, the 4s and inner 3d electrons can be removed. [1]

6

a) A transition metal is a metal that forms one or more cations in which the metal has a partially filled d sublevel [1]
b) +3 [1]
c) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 4d^6$ [1]
d) 4 [1]
e) Octahedral [1]



[1]

S3.2 Functional groups: classification of compounds

Paper 1 style questions

1	B	12	B
2	A	13	A
3	D	14	B
4	A	15	A
5	B	16	B
6	C	17	D
7	D	18	D
8	B	19	C
9	B	20	A
10	D	21	D
11	A	22	D

Paper 2 style questions

1

a) Pentane has more electrons than propane and hence the instantaneous dipole-induced dipole interactions / London (dispersion) forces of attraction between the pentane molecules are stronger than that of propane. [1]
Thus, more energy is required to break the stronger London (dispersion) forces of attraction between pentane than propane, leading to the higher boiling point of pentane [1]

b) Pentane has a larger surface area of contact between its molecules than 2,2-dimethylpropane (which has the same number of electrons.) [1]
Therefore the instantaneous dipole-induced dipole interactions or London (dispersion) forces of attraction between the pentane molecules are more extensive and requires more energy to be broken/overcome [1]

c) Structural isomerism; chain = isomerism [2]

2

a) $C_9H_{10}O_4$ [1]

b) Hydroxyl, ester, alkenyl, carbonyl [2]

c) Tertiary alcohol [1]
the carbon atom bonded to the –OH group has three other carbon atoms bonded to it [1]

3

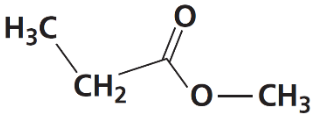
Radio waves [1]
Infrared spectroscopy [1]
Functional groups between the two molecules are identical/the same [1]
4 peaks for propan-1-ol and 3 peaks for propan-2-ol; [1]
Number of hydrogen atoms in each chemical environment is 3:2:2:1 for propan-1-ol and 6:1:1 in propan-2-ol; [1]

4

a) The number of different chemical environments/hydrogen environments/proton environments [1]

The number of (identical) hydrogens/protons on the neighbouring carbon atom(s) [1]

b)



[1]

Any two of the following: [2]

- 3 types of hydrogen/proton environments
- 2 × CH_3 (peak area 3) and 1 x CH_2 (peak area 2)
- one of the CH_3 groups is next to oxygen (at 4.1 ppm), the proton is the most de-shielded/downfield
- one of the CH_3 groups is next to carbon (at 0.9 ppm), proton most shielded/upfield
- the CH_2 group next to >C=O at 2 ppm, signal is in between the two extremes
- one of the CH_3 groups that has no neighbouring protons produces a singlet
- one of the CH_3 next to CH_2 (2+1) produces a triplet
- the CH_2 next to CH_3/C_2H_5 (3+1) produces a quartet

Award 1 mark maximum if structure $CH_3COOCH_2CH_3$ given and the explanation involves only integration or/and splitting patterns.

5

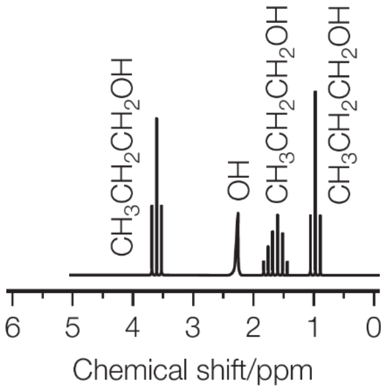
a) 3350 cm-1 due to O—H bond [1]
2900 cm-1 due to C—H bond; [1]
award [1] if both bonds given but not linked to correct absorption.

b) (comparison of) fingerprint regions/with library, database or known spectra [1]

c) same molecular ion/parent ion/M^+/due to $[C_4H_9OH]^+$/$[C_4H_{10}O]^+$ [1]

d) i. ($m/z = 57$) $[C_4H_9]^+$ /$[(CH_3)_3C]^+$ [1]
ii. (m/z = 45) $[C_2H_5O]^+$ /$[CH_3CHOH]^+$ /$[C_2H_4OH]^+$ [1]
Penalize missing + charge once only in (c) and (d).
i. (four peaks because) four different hydrogen environments/four different types of proton/*OWTTE* [1]
6:2:1:1 [1]
ii. $(CH_3)_3COH$/more detailed formula [1]

6



[1]

It has four hydrogen environments and there will be splitting of the methyl group and both sets of methylene (-CH_2-) protons. [1]

In the high resolution ^1H NMR spectrum the methyl group (**CH_3**CH_2CH_2OH) is a triplet at approximately 1 ppm, [1]

and the alcohol or hydroxyl group (CH_3CH_2C**H_2OH**) is a broad singlet at approximately 2.3 ppm. [1]

Of the two methylene groups, the one next to the highly electronegative oxygen atom of the –OH group is most downfield at approximately 3.6 ppm. This methylene group (CH_3CH_2**CH_2**OH) is split into a triplet. The methylene group adjacent to the methyl group appears as a sextet [1]

7

a) i. award [1] each for any two of: [2]

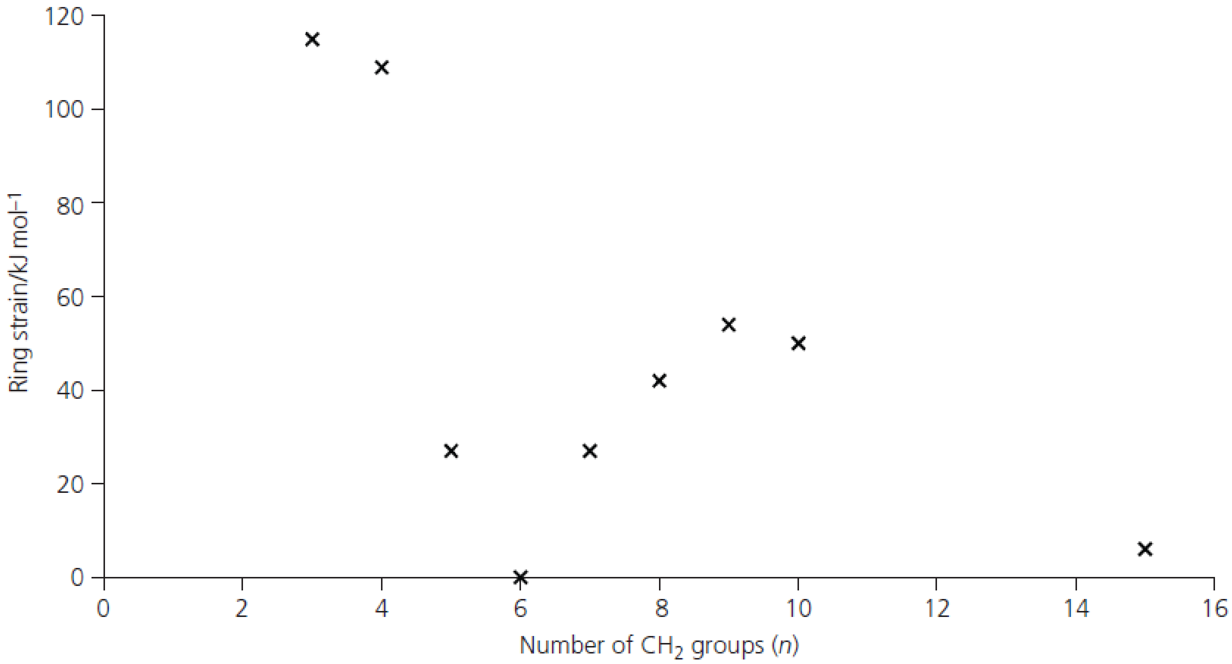
- a group of organic molecules with the same general formula
- neighbouring members differing by –CH2-
- similar chemical properties
- gradation/trend in physical properties
- same functional group

ii. a compound/molecule containing carbon and hydrogen only [1]

iii. containing only single (C–C) bonds/no multiple (carbon–carbon) bonds [1]

b) C_nH_{2n} [1]

c) [1 mark for correct plotting of line graph or bar chart]
[1 mark for correct axes and labels]



d) When the number of -CH$_2$- groups increases from 3 to 6, ring strain decreases from 115 to 0 kJ mol^{-1}. [1]
From $n = 6$ to $n = 9$, ring strain increases again from 0 to 54 kJ mol^{-1} [1]
It then decreases again when $n = 9$ to $n = 15$ from 54 kJ mol^{-1} to 6 kJ mol^{-1} [1]

e) The most stable cycloalkane is cyclohexane [1]
while the least stable is cyclopropane [1]

f) Given that the interior angle of a hexagon is 120°, the bond angle deformation of cyclohexane should be 120° as well, but it is 109.5° [1]
Cyclohexane is non-planar and its bond angles must be the preferred tetrahedral bond angle of 109.5°. A planar structure would have bond angles of 120°, like benzene [1]

THEME R

R1.1 Measuring enthalpy changes

Paper 1 style questions

1	A	**7**	B	**13**	C
2	D	**8**	D	**14**	A
3	C	**9**	C	**15**	D
4	A	**10**	D	**16**	B
5	C	**11**	A		
6	B	**12**	C		

Paper 2 style questions

1 **a** The enthalpy change when the amounts of reactants in a specified stochiometric equation [1] react together under standard conditions of 298 K and 100 kPa [1] with all compounds in their standard states and solutions at a concentration of 1.00 mol dm^{-3}. [1]

b Experiment:

Measure volume of dilute acid using a measuring cylinder. [1]

Place in a polystyrene beaker with lid. [1]

Take temperature of the acid and ensure constant reading. [1]

Place mixture on a stirrer / mention of stirring during reaction [1]

Add small volume (for example 5 cm^3) of alkali and record temperature. [1]

Repeat additions of alkali should be made and recording of the temperature until the mixture has started to cool. [1]

Processing data:

Plot graph of temperature vs volume of acid. [1]

Extrapolate cooling curve back to time of mixing. [1]

Calculate heat energy change using $Q = mc\Delta T$. m is combined mass of solutions at neutralisation point. [1]

Find moles of either reactant used with concentration × volume.[1]

Calculate ΔH using $-Q$/moles of reactant. Q in kJ. [1]

Max 9 marks

2 **a** It is exothermic because heat is released to the surroundings and the temperature rises. [1]

b To make any heat loss as small as possible, so that all the heat will be given out very rapidly. [1]

c Heat released = mass × specific heat capacity × temperature increase

Amount of LiOH or HCl used = 0.500 dm^3 × 0.050 mol dm^{-3} = 0.025 mol[1]

Heat released = 100 g × 4.18 J g^{-1} °K^{-1} × 3.5 °C = 1463 J = 1.463 kJ [1]

ΔH = (1.463 kJ ÷ 0.025 mol) = −58.5 kJ mol^{-1} [1] correct value [1] sign

d Significant heat loss to the surroundings. [1] Insulate the reaction vessel and use a lid OR Draw a temperature versus time graph and extrapolate to calculate the rise in temperature.[1]

e The temperature change would be the same, namely, 3.5 °C. [1] This occurs because the amount of lithium hydroxide reacted would be the same. The excess hydrochloric acid would not react; lithium hydroxide is the limiting reagent. [1]

3 **a** (23.70 − 23.03) °C = 0.67 °C / K [1]

b 1.28×10^{-3} mol [1]

c $\Delta H = -\frac{10.114 \times 0.67}{1.28 \times 10^{-3}} = -5300$ kJ mol^{-1} [1] for magnitude [1] correct sign

d $\frac{-5.6 \times 10^{3} - -5.3 \times 10^{3}}{-5.6 \times 10^{3}} \times 100 = 5.4\%$ [1]

e $\Delta H_c^\ominus$ sucrose > $\Delta H_c^\ominus$ TNT suggests size of enthalpy change not important, or this alone cannot produce a powerful explosive. [1]

Rate of reaction for TNT is greater suggesting this is an important factor. [1]

Moles of gas generated by sucrose is 11 more than started with. For 1 mol of TNT, 7.5 mol of gas are generated. This suggests producing a lot of gas is not important. [1]

R1.2 Energy cycles in reactions

Paper 1 style questions

1	C	**7**	A	**13**	D
2	A	**8**	B	**14**	C
3	B	**9**	C	**15**	B
4	C	**10**	C	**16**	B
5	D	**11**	D	**17**	C
6	D	**12**	C		

Paper 2 style questions

1 **a** $1N(NO_2)_3(g) + 2\ CH_3OH(l) \rightarrow 2N_2(g) + 2CO_2(g) + 4H_2O(l)$ [1]

b Products from the reaction are non-toxic / normal components of the atmosphere/ no chlorine produced [1]

c $N(NO_2)_3 \rightarrow 2N_2(g) + 3O_2(g)$

$\Delta H = [(6 \times 305) + (3 \times 158)] - [(2 \times 945) + (3 \times 498)]$ [1]

$= 2304 - 3384$ [1]

$= -1080\ \text{kJ mol}^{-1}$ [1]

2 **a** $\Delta H = (614 + 436) - (346 + 414 \times 2)$ [1] for correct multiples [1] for correct signs

$\Delta H = -124$ [1]

Exothermic [1]

b Heat energy required $= \frac{1000 \times 4.18 \times 80}{1000} = 334.4$ kJ [1] for expression [1] for correct value

$n = \frac{-Q}{\Delta H} = \frac{-334}{-4000} = 0.084$ mol [1]

3 **a** Energy needed to break (1 mol) of a bond in the gaseous state averaged over similar compounds. [1]

b $\Delta H_c^{\ominus}$, $C_2H_5OH(s) + 3O_2(g) \rightarrow 2CO_2\rightarrow(g) + 3H_2O(l)$ [1]

$\Delta H_c^{\ominus} = \sum \Delta H_f^{\ominus}(\text{products}) - \sum \Delta H_f^{\ominus}(\text{reactants})$

$= (2 \times -394) + (3 \times -286) - -278$ [1]

$= -1368\ \text{kJ mol}^{-1}$[1]

c Heat loss to the environment [1]

d $\Delta H_f^{\ominus}$, $3C(s) + 3H_2(g) + O_2(g) \rightarrow C_2H_5COOH(l)$ [1]

$\Delta_f H^{\ominus} = \sum \Delta H_c^{\ominus}(\text{reactants}) - \sum \Delta H_c^{\ominus}(\text{products})$ [1]

$= (3 \times -394) + (3 \times -286) - -1527$ [1]

$= -513\ \text{kJ mol}^{-1}$[1]

4 **a** 0.0250 mol [1]

b $Q = 50 \times 4.18 \times 9.7 = 2027$ J [1]

$\Delta H_1 = \frac{-Q}{n} = \frac{-2027}{0.025 \times 1000} = -81\ \text{kJ mol}^{-1}$ [1]

c i $\Delta H = \Delta H_1 - \Delta H_2$

$= -81 - 18$

$= -99\ \text{kJ mol}^{-1}$ [1]

ii $\frac{103-99}{103} \times 100 = 3.9\%$ [1]

d Any 2 points from: [2]

MgSO$_4$ is not completely anhydrous.
MgSO$_4$ is impure.
Heat lost to surroundings
Heat capacity of solution is taken to be that of water
Density of solution taken to be 1 g cm^{-3}.
Uncertainty in the thermometer is high because of small temperature change.
All the solid did not dissolve.
Conditions are not standard, the literature value is at standard conditions.

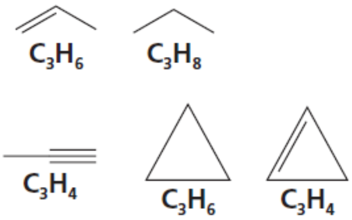
R1.3 Energy from fuels

Paper 1 style questions

1	B	5	B	9	D
2	D	6	D	10	B
3	C	7	A	11	D
4	C	8	C	12	A

Paper 2 style questions

1 **a**



1 mark for 3 correct structures, 2 marks for 4 correct structures [2]

b Reacted with lime / CaO / sodalime / $Ca(OH)_2$ / KOH / NaOH / alkali neutralized with a base [1]

c **i** $\frac{891\ \text{kJ mol}^{-1}}{16.05\ \text{g mol}^{-1}} = 55.5\ \text{kJ g}^{-1}$ [1]

ii Shale gas will have a lower specific energy because it is not pure methane. It is 10% carbon dioxide, CO_2, and this will not be combusted. [1]

d Shale gas has the lowest percentage of carbon OR Shale gas has a high specific energy and low carbon content OR The molecules present have a shorter carbon chain/ shorter hydrocarbon [1]

2 **a** Advantages of ethanol [max 2]

- lower greenhouse emissions
- reduced crude oil use
- renewable energy source.

Disadvantages of ethanol [max 2]

- lower specific energy
- potential damage to engines
- high energy cost of distillation step
- land taken which could be used for growing food.

b $C_4H_9OH + 6O_2 \rightarrow 4CO_2 + 5H_2O$ [1]

c Although butan-1-ol releases ×3 more energy per mole of fuel burnt, the energy released per dm^3 for the two alcohols is much closer. [1]

1 mole of methanol has a mass of 32.1 g, whereas 1 mole of butan-1-ol has a mass of 72.1 g. This means that the energy density of the two alcohols is:

methanol $\frac{726\ \text{kJ mol}^{-1}}{32.1\ \text{g mol}^{-1}} = 22.6\ \text{kJ g}^{-1}$

butan-1-ol $\frac{2676\ \text{kJ mol}^{-1}}{72.1\ \text{g mol}^{-1}} = 37.1\ \text{kJ g}^{-1}$

The density of the two alcohols is similar, and so when burning equal volumes of each alcohol, butan-1-ol will not release 3 times the amount of energy.

d Energy in kJ mol^{-1} = 1367 kJ mol^{-1}

Energy per gram = $\frac{1367 \text{ kJ mol}-1}{46.08 \text{ g mol}^{-1}}$ = 29.60 kJ g^{-1} [1]

Energy per cm^3 = 29.60 kJ g^{-1} × 0.780 g cm^{-3} = 23.1 kJ cm^{-3} [1]

23.1 kJ cm^{-3} × 1000 = 23 100 kJ dm^{-3} [1]

3 **a** Anode: $CH_3OH + H_2O \rightarrow 6H^+ + 6e^- + CO_2$ [1]

Cathode: $\frac{3}{2}O_2 + 6H^+ + 6e^- \rightarrow 3H_2O$ [1]

b Anode: $CH_3CH(OH)CH_3 + 5H_2O \rightarrow 3CO_2 + 18H^+ + 18e^-$ [1]

Cathode: $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$ (accept multiples) [1]

Overall: $CH_3CH(OH)CH_3 + \frac{9}{2}O_2 + \rightarrow 3CO_2 + 4H_2O$ (accept multiples) [1]

c Propan-2-ol is a liquid but hydrogen is a gas, so propan-2-ol is easier to transport and store. Propan-2-ol also has a higher energy density. [1]

4 **a** alkanes [1]

b **i** Mass burnt = 10.8 × 10700 = 115 560 kg

= 116 tonnes (3 s.f.) [1]

ii amount of kerosene = $\frac{116 \times 10^6 \text{ g}}{198.44 \text{ g mol}^{-1}}$ = 5.85×10^5 mol [1]

mass of CO_2 formed = 5.85×10^5 mol × 14 × 44.01 g mol^{-1} = 3.60×10^8 g (360 tonnes) [1]

R1.4 Entropy and spontaneity (HL only)

Paper 1 style questions

1	B	**7**	B	**13**	D
2	A	**8**	A	**14**	C
3	D	**9**	C	**15**	B
4	B	**10**	C	**16**	A
5	B	**11**	C	**17**	D
6	B	**12**	A		

Paper 2 style questions

1 **a** The entropy increases due to an increase in the number of moles of gaseous particles, resulting in more disorderliness in the system and more ways of distributing the energy (quanta)/arranging the particles. [1]

b $\Delta G = \Delta H - T\Delta S$; $\Delta G = +117 - (298 \times \frac{175}{1000}) = +64.9\ \text{kJ mol}^{-1}$ [1]

Since $\Delta G^{\ominus}$ is positive, the decomposition is not spontaneous at room temperature. [1]

2 **a** The enthalpy change when one mole of a compound is formed from [1]

its elements (in their standard states), under standard conditions (100 kPa). [1]

b **i** $SiCl_4(l) + 2H_2O(l) \rightarrow SiO_2(s) + 4HCl(g)$

[1] for formulas and balancing; [1] for state symbols

ii $\Delta H^{\ominus} = \sum \Delta H_f^{\ominus}$ (products) – $\sum \Delta H_f^{\ominus}$ (reactants) = $[-910 + 4(-92.0)] - [-640 + 2(-286)]$ [1]

$-66.0\ \text{kJ mol}^{-1}$ [1]

c Positive entropy change. 4 moles of gas (hydrogen chloride) are produced. [1]

There is an increase in the disorder of the system as there are more ways of arranging the particles and quanta of energy in the system. [1]

d $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus}$ and since $\Delta H^{\ominus}$ is negative and $\Delta S^{\ominus}$ is positive [1], $\Delta G^{\ominus}$ is always negative.[1] So the process is spontaneous at all temperatures.

3 **a** $\Delta S^{\ominus} = 88.3 - 77.8 = +10.5\ \text{J K}^{-1}\ \text{mol}^{-1}$; [1]

Minimum temp when $\Delta G = 0$, thus $4.2 - T \times 0.0105 = 0$; [1]

$T = 400$ K [1]

b $\Delta G^{\ominus} = -8.31\ \text{J K}^{-1}\ \text{mol}^{-1}\ (298\ \text{K}) \ln (2.97 \times 10^{-13})$

$= 7.14 \times 10^{4}\ \text{J mol}^{-1} = +71.4\ \text{kJ mol}^{-1}$ [1]

c The sign of $\Delta G^{\ominus}$ is positive, so the reaction is not spontaneous under standard conditions. The value shows that the position of equilibrium lies significantly to the left, and there will be, to a very good approximation, no reaction. [1]

4 **a** Δn of gas = (2 mol − 1 mol) = +1 mol; ΔS is positive; OR As temperature increases to 1300 K, ΔG becomes negative, so ΔS must be positive.

When the temperature is low, ΔG is positive and the reaction is non-spontaneous. Since ΔS is positive, only a positive value of ΔH could allow ΔG to be positive. [1]

b $+78\ \text{kJ mol}^{-1} = \Delta\text{H} - 378\ \text{K} \times \Delta S = A$ $-58\ \text{kJ mol}^{-1} = \Delta\text{H} - 1300\ \text{K} \times \Delta S = B$

Equation A – equation B:

$136\ \text{kJ mol}^{-1} = 922\ \text{K} \times \Delta S$ [1]

$\Delta S = +0.148\ \text{kJ K}^{-1}\ \text{mol}^{-1}$ ($+148\ \text{J K}^{-1}\ \text{mol}^{-1}$) [1]

Putting ΔS into A

$+78\ \text{kJ mol}^{-1} = \Delta H - 378\ \text{K} \times 0.148\ \text{kJ K}^{-1}\ \text{mol}^{-1}$ [1]

$\Delta H = +134\ \text{kJ mol}^{-1}$ [1]

R2.1 How much? The amount of chemical change

Paper 1 style questions

1	D	9	B	17	B
2	A	10	C	18	D
3	C	11	D	19	C
4	B	12	B	20	D
5	D	13	A	21	C
6	C	14	B	22	D
7	B	15	C		
8	C	16	D		

Paper 2 style questions

1 **a** $LiH(s) + H_2O(l) \rightarrow LiOH(aq) + H_2(g)$ [1]

Amount of LiH(s) = (1.59 g/7.95 g mol^{-1}) = 0.200 mol [1]

Amount of $H_2(g)$ = 0.200 mol

Volume of $H_2(g)$ = 0.200 mol × 22.7 dm^3 mol^{-1} = 4.54 dm^3 [1]

b $LiOH(aq) + HCl(aq) \rightarrow LiCl(aq) + H_2O(l)$ [1]

Amount of LiOH(aq) = 0.200 mol

Amount of HCl needed to neutralize = 0.200 mol [1]

Molar concentration of HCl(aq) = (72.92 g dm^{-3}/36.46 g mol^{-1}) = 2.000 mol dm^{-3} [1]

Amount (mol) = vol (dm^3) × molarity (mol dm^{-3})

0.200 mol = vol (dm^3) × 2.00 (mol dm^{-3}); vol (dm^3) = 0.100 dm^3 = 100 cm^3 [1]

2 **a** $2Cu_2O(s) + Cu_2S(s) \rightarrow 6Cu(s) + SO_2(g)$ [2]

1 mark for balanced equation; 1 mark for state symbols

b $n(Cu_2O)$ = (12.0 × 10^3 g/143.10 g mol^{-1}) = 83.9 mol [1]

$n(Cu_2S)$ = (4.00 × 10^3 g ÷159.16 g mol^{-1}) = 25.1 mol [1]

Penalize failure to convert from kg to g once only.

Copper(I) sulfide, Cu_2S, is the limiting reagent

since 25.1 mol × 2 = 50.2 mol of Cu_2O needed so Cu_2O is excess reactant [1]

c $n(Cu) = 6 \times n(Cu_2S)$ = 6 × 25.1 mol = 151 mol; [1]

$m(Cu)$ = 151 mol × 63.55 g mol^{-1} = 9600 g = 9.6 kg [1]

If Cu_2O given in **a**, allow 3 × $n(Cu_2O)$ and 3 × $n(Cu_2O)$ × 63.55.

3 one mark for each correct row in the table [3]

	$CS_2(g)$	+	$3O_2(g)$	→	$CO_2(g)$	+	$2SO_2(g)$
Initial volume/ cm^3	20		100		–		–
Change in volume / cm^3	−20		−60		+20		+40
Final volume / cm^3	–		40		20		40

Total volume of gases at the end of reaction = (40 + 20 + 40) cm^3 = 100 cm^3 [1]

Both CO_2 and SO_2 are acidic gases which will react with NaOH(aq)

Hence the % of final volume dissolved in alkali (60/100 × 100%) = 60% [1]

4 **a** Cumene process = $\frac{94.12}{152.21}$ × 100% = 61.8% [1]

Catalytic process = $\frac{94.12}{122.14}$ × 100% = 77.0% [1]

The catalytic process is more sustainable as it has a higher atom economy. [1]

b any 3 from: [3]

- the number of steps in the synthesis
- the type of waste produced in the synthesis
- the solvents required in the synthesis
- the amount of energy needed in the synthesis
- the toxicity and reactivity of chemicals.

5 **a** $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$ [1]

b Equal volumes of all gases contain equal numbers of particles (at the same temperature and pressure) [1]

c 666 (2/3) cm^3 of hydrogen and 333 (1/3) cm^3 of oxygen; 2:1 ratio. [1]

d It increases up to 600 cm^3 hydrogen to a maximum [1]

and then decreases to zero when the volume of hydrogen is 0 cm^3. [1]

e There is no oxygen for it to react with and hence no water formed. [1]

f 4.04 g of hydrogen reacts with 32.00 g of oxygen to form 36.04 g of water. [1]

The mass ratio is 4.04:32.00:36.04 = 1:8:9 (integers) [1]

6 **a** Amount of copper atoms = 19.76 g/63.55 g mol^{-1} = 0.312 mol [1]

Amount of oxygen atoms = 4.98 g/16.00 g mol^{-1} = 0.311 mol [1]

CuO [1]

+2 [1]

b Some of the oxide is not reduced completely due to insufficient heating / time [1]

Ensure the air hole of Bunsen burner is fully opened / use a smaller mass to ensure complete decomposition [1]

7 **a** Amount of butan-1-ol = (6.25 g/74.12 g mol^{-1}) = 0.084 3 mol

Amount of butyl ethanoate = (6.57 g/116.16 g mol^{-1}) = 0.056 56 mol [1]

Percentage yield = (0.056 56 mol/0.084 3 mol) × 100 = 67.1% [1]

b Atom economy = 100 × (116.16/134.17) [1]

= 86.6% [1]

c Amount of ethanoic acid = (5.45 g/60.05 g mol^{-1}) = 0.090 8 mol

Amount of butyl ethanoate = (9.806 g/116.16 g mol^{-1}) = 0.084 4 mol [1]

Percentage yield = (0.084 4 mol/0.0908 mol) × 100 = 93.1% [1]

d The atom economy is)116.17/253.16) × 100 [1]

= 45.9% [1]

e Link between yield and an explanation required: (high percentage) yield shows a high % conversion of reactants into products [1]

Link between atom economy and an explanation required: (low) atom economy shows a lot of waste (product) or (low) atom economy shows not much desired product [1]

f Less waste (products) or a higher atom economy [1]

Less toxic reactants or less toxic (waste) products, or less corrosive reactants, or less corrosive (waste) products, or less harmful reactants, or less harmful (waste) products, or less hazardous reactants, or less hazardous (waste) products. [1]

R2.2 How fast?

The rate of chemical change

Paper 1 style questions

1	D
2	D
3	A
4	B
5	C
6	B
7	D
8	B
9	B
10	D
11	A
12	D
13	C
14	B
15	B
16	A
17	C
18	B
19	B
20	C
21	B

Paper 2 style questions

1 a i it is decreasing; [1]
less frequent (effective) collisions / fewer (effective) collisions per second or (unit) time between the ions; [1]

ii reactant(s) used up/reaction is complete [1]
Do not accept reaction reaches equilibrium.

b i It would increase. [1]
Accept a quantitative answer such as 'doubles'.
more frequent collisions; [1]
collisions or colliding species/ions have more energy (OWTTE); [1]
more molecules with kinetic energy greater than activation energy [1]

ii rate would be lower [1]
smaller surface area [1]

2 a i iodine – oxidizing agent [1]
H^+/H_2SO_4 – catalyst [1]

ii $NaHCO_3$ added prior to the titration to quench reaction by reacting with the catalyst, H_2SO_4
so that concentration of iodine present at different time intervals can then be determined [1]

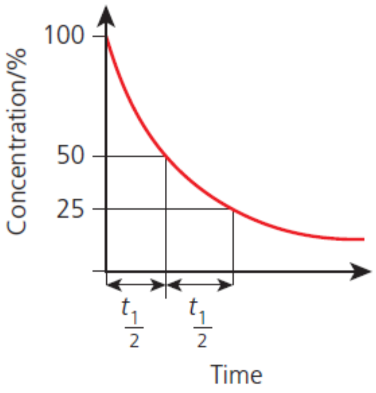
b i zero order with respect to iodine [1]
First order with respect to sulfuric acid [1]

ii rate [1]

iii Magnitude of gradient would increase. [1]

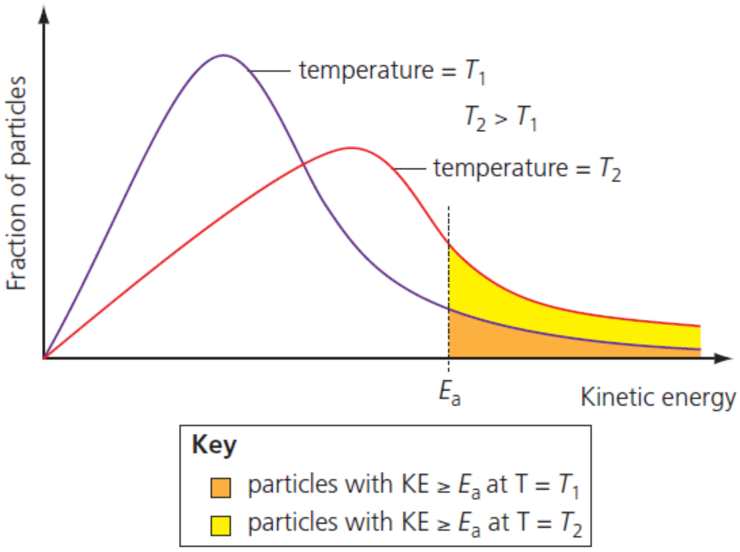
iv Volume of thiosulfate required = $20.0\ cm^3$
Amount of I_2 reacted = $0.01 \times 1.0 \times 10^{-3} = 1.0 \times 10^{-5}$ mol [1]
Amount of thiosulfate required = $1.0 \times 10^{-5} \times 2 = 2.0 \times 10^{-5}$ mol
Concentration of thiosulfate used = $1.00 \times 10^{-3}\ mol\ dm^{-3}$ [1]

v



[2] (one mark for graph (and axes labels) and one mark for showing constant half-life)

vi

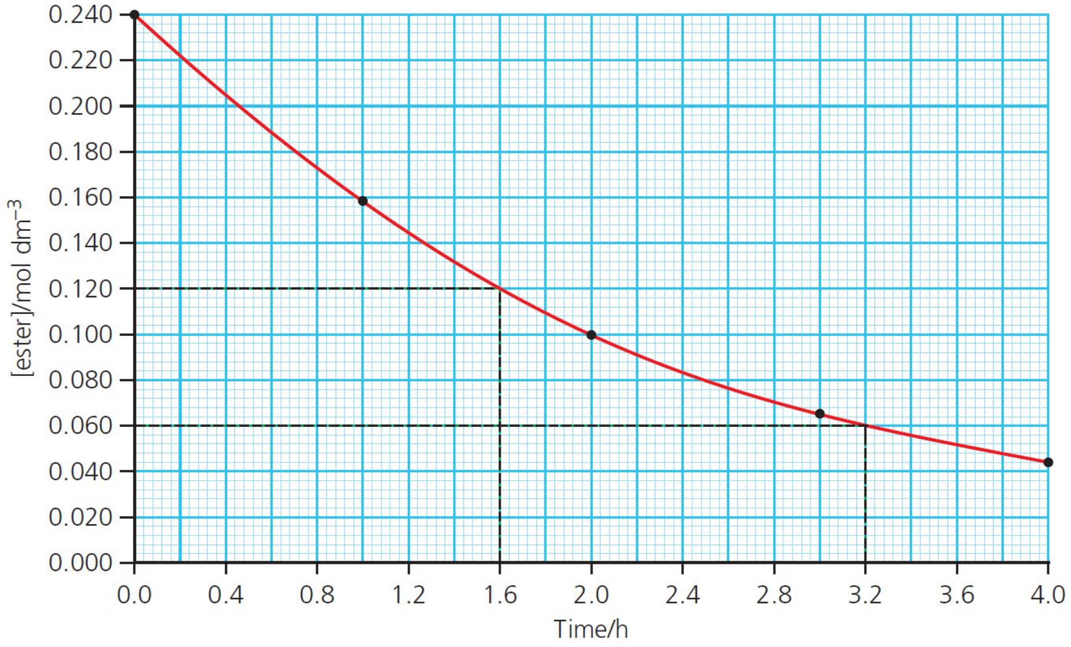


curves starting from origin, correctly labelled with temperature [1]
labelled axes with single E_a marked out. (2 E_a lose this mark) [1]
At higher temperatures, the average kinetic energy of molecules increases. [1]
The number of reactant particles with kinetic energy ≥ E_a will increase and the frequency of effective collisions increases. Hence rate of reaction will increase. [1]

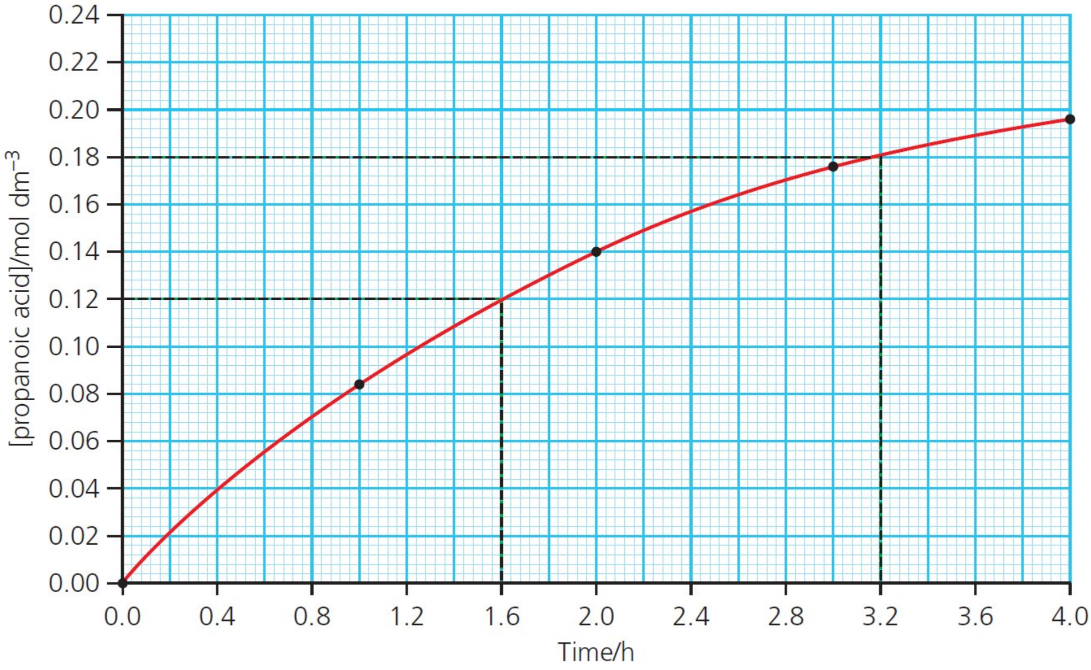
3 a i The cold water is a quenching agent / meant to quench the reaction. [1]
It is necessary to slow down the reaction significantly (by lowering concentration and temperature of the reaction) so that the reaction is considered to have stopped at that instant. [1]

Time / h	**Concentration of propanoic acid / mol dm^{-3}**	**Concentration of ester / mol dm^{-3}**
0	0.000	**0.240**
1	0.084	**0.156**
2	0.140	**0.100**
3	0.178	**0.062**
4	0.195	**0.045**

EITHER [Reactant]–time graph



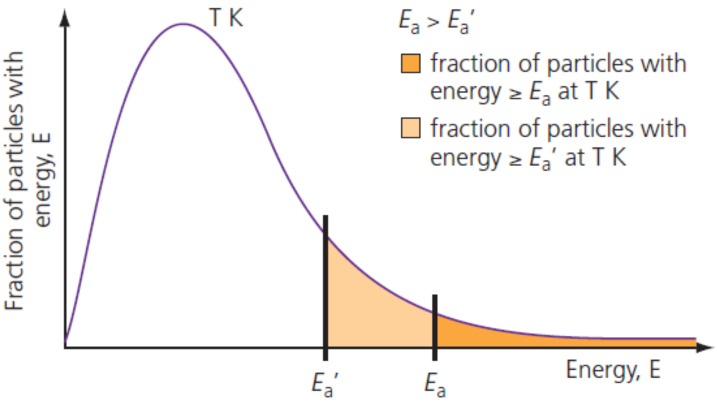
OR [Product]–time graph



In either case:

labelled axes with units [1]
points plotted [1]
smooth curve [1]
appropriate lines shown to determine at least two half-life values [1]
half-life is approximately 1.65 h, or 99 min or 5940 s [1]
The half life is constant, hence it is first order with respect to the ester. [1]

iii The catalyst speeds up the rate of reaction [1]
by providing an alternative reaction pathway with lower activation energy. [1]



[1]

With a catalyst, as shown by the Maxwell–Boltzmann distribution, there is a higher fraction of molecules with kinetic energy ≥ lowered activation energy and the rate of reaction increases. [1]

b **i** Comparing experiments 1 and 2, when $[OH^-]$ increased by 1.5 times while keeping [ester] constant, initial rate increased by 1.5 times. [1]
Hence, order of reaction with respect to $[OH^-]$ is 1. [1]

ii Comparing experiments 2 and 3,

$$r = \left(\frac{0.06}{0.03} \times \frac{0.02}{0.015} \times 4.05 \times 10^{-3}\right) = 0.0108 \text{ mol dm}^{-3}\text{ s}^{-1} \quad [1]$$

4 **a** **i** The shorter the time, the faster the initial rate of reaction (or rate = 1/time). [1]

ii Initial rate for experiments 1: $1/60 = 0.0167\text{ s}^{-1}$
Initial rate for experiments 2: $1/48 = 0.0208\text{ s}^{-1}$
Initial rate for experiments 3: $1/40 = 0.0250\text{ s}^{-1}$
Initial rate for experiments 4: $1/34 = 0.0294\text{ s}^{-1}$
The relative rate between experiments 1 and 2:0.0208/0.0167 = 1.25 [1]
Using experiments 1 and 2, when the concentration of HSO_3^- is constant and the concentration of HCOH increases 1.25 times, the rate also increases 1.25 times. Hence the order of reaction with respect to H_2CO is one. [1]
Comparing experiments 2 and 3 and using the rate equation for each:

$$\frac{0.0208}{0.025} = \frac{k[0.04]^x k[0.05]^1}{k[0.05]^x[0.06]^1} \quad [1]$$

$$\frac{0.0208}{0.025} = (4/5)^x \times \frac{5}{6};\ x = 0$$

therefore the order of reaction is zero. [1]

iii rate = $k[H_2CO]$ [1]

units of rate constant, $k = \frac{\text{mol dm}^{-3}\text{ s}^{-1}}{\text{mol dm}^{-3}} = \text{s}^{-1}$ [1]

iv Elementary step 1: $HSO_3^- + H_2O \rightarrow SO_3^{2-} + H_3O^+$ [1]
Elementary step 2: $H_2O + SO_3^{2-} + H_2CO \rightarrow CH_2(OH)SO_3^- + OH^-$ [1]

R2.3 How far? The extent of chemical change

Paper 1 style questions

1	C	**6**	B	**11**	A
2	A	**7**	D	**12**	B
3	A	**8**	C	**13**	D
4	C	**9**	B	**14**	B
5	D	**10**	C	**15**	D

Paper 2 style questions

1 **a** 200 °C [1]
600 atm [1]
(Allow the 'highest pressure and the lowest temperature'.)

b **i** Yield increases / equilibrium moves to the right / more ammonia. [1]
4 (gas) molecules → 2 / decrease in volume / fewer molecules on right-hand side. [1]

ii Yield decreases / equilibrium moves to the left/ less ammonia; [1]
exothermic reaction. [1]

c High pressure expensive/greater cost of operating at high pressure/reinforced pipes, etc. needed. [1]
Lower temperature: greater yield but lowers rate. [1]
(*Do not award a mark for simply saying 'compromise'*.)

d $\frac{[NH_3]^2}{[N_2][H_2]^3}$ [1]

e **i** Artificial fertilizers / increasing crop yields; production of high explosives for mining [1 max].

ii The claim is not valid since catalysts do not alter the yield / position of equilibrium [1]
they only increase the rates of the forward and backward reactions equally so equilibrium is reached faster. [1]

2 **a** $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ [1]
Equilibrium concentration of $[SO_3] = 0.020$ mol dm^{-3} [1]
At equilibrium, $[SO_2] = (0.1 – 0.020) = 0.080$ mol dm^{-3} [1]
At equilibrium $[O_2] = (0.050 – 0.010) = 0.040$ mol dm^{-3} [1]
$K = \frac{(0.020)2}{(0.080)2 \times 0.040} = 1.56$[1]

3 **a** Dynamic equilibrium refers to a reversible reaction in which the forward and the backward reactions are taking place at the same rate [1]
and hence, the concentrations of reactants and products remain constant (no net or overall change in concentrations of reactants and products) [1]

b Pressure:
By Le Châtelier's Principle, with a high pressure of 50 MPa, the equilibrium position will shift to the right so as to favour the production of less gaseous molecules to reduce the pressure. [1]
At higher pressure, the reaction proceeds at a faster rate. [1]
Catalyst:
There is no change to the equilibrium position, as the catalyst speeds up the rate of both forward and backward reaction to the same extent. [1]

Hence, the rate of the reaction increases, compared to that under standard conditions, without a catalyst. [1]

c A low temperature favours an increase in yield [1]
but the rate of reaction is low and a long time is needed to reach equilibrium. [1]

4 **a** $n = pV/RT = 4.90$ mol [1]

b

	CO	$2H_2$	CH_3OH
Initial / mol	3.2	3.2	0
Charge / mol	$-x$	$-2x$	$+x$
Equilibrium / mol	$3.20 - x$	$3.20 - 2x$	x

$n = 4.90 = 3.20 - x + 3.20 - 2x + x$
$4.90 = 6.40 - 2x$
$n(CH_3OH) = x = 0.750$ mol [1]
$n(CO) = 3.20 - 0.750 = 2.45$ mol [1]
$n(H_2) = 3.20 - 2(0.750) = 1.70$ mol [1]

c $K = [CH_3OH]/([CO_2] \times [H_2]^2)$
$= (0.750/10.0)/((2.45/10) \times (1.7/10)^2) = 10.6$ [1]

d $\Delta G = -RT\ln K = (8.31\ \text{J mol}^{-1}\ \text{K}^{-1}) \times (273 + 120\ \text{K}) \times \ln 10.6$
$\Delta G = -7710\ \text{J mol}^{-1} = -7.710\ \text{kJ mol}^{-1}$ [1]

R3.1 Proton transfer reactions

Paper 1 style questions

1	B	**9**	B	**17**	D
2	D	**10**	C	**18**	B
3	C	**11**	A	**19**	B
4	A	**12**	B	**20**	C
5	A	**13**	B	**21**	D
6	C	**14**	C	**22**	B
7	B	**15**	A		
8	A	**16**	B		

Paper 2 style questions

1 **a** Acids are proton / H^+ donors. [1]
Bases are proton / H^+ acceptors. [1]

b $H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq)$ [1]

c a species that is formed after an acid loses a proton [1]

d $H^+(aq) + \mathbf{HCO_3}^-(aq) \rightarrow H_2CO_3(aq)$ (proton acceptor) [1]
$OH^- + \mathbf{HCO_3}^-(aq) \rightarrow H_2O(l) + CO_3^{2-}(aq)$ (proton donor) [1]

2 **a** Amount of $Ca(OH)_2$ reacted = 0.250 × (22.30/1000) = 5.575×10^{-3} mol [1]
Amount of lactic acid = 2 × 5.575×10^{-3} mol = 0.01115 mol [1]
Molar mas of lactic acid = 1.00 g / 0.01115 mol = 89.7 [1]

b $[H^+] = 10^{-pH}$; $10^{-2.43} = 3.72 \times 10^{-3}$ mol dm^{-3} [1]
$K_a = [H^+]^2/[HA] = (3.72 \times 10^{-3})^2/(0.100) = 1.38 \times 10^{-4}$ mol dm^{-3} [1]

c **i** A buffer solution is a solution that resists changes in pH when small amounts of acid or base are added to it.
H^+ is removed by reaction with lactate ions: $H^+ + C_3H_5O_3^- \rightarrow C_3H_6O_3$
or $H^+ + A^- \rightarrow HA$ [1]

ii OH^- is removed by reaction with lactic acid:
$OH^- + C_3H_6O_3 \rightarrow C_3H_5O_3^- + H_2O$
or
$HA + OH^- \rightarrow A^- + H_2O$ [1]

d $[A^-] = [HA] = 0.1$ mol dm^{-3}
Amount of lactate ions = 0.1 × 500/1000 = 0.05 mol [1]
Mass of sodium lactate needed = 0.05 × 112.0 = 5.60 g [1]

3 **a** $[H^+(aq)]$ from 0.100 mol dm^{-3} $H_2CO_3 = 10^{-3.68} = 2.09 \times 10^{-4}$ mol dm^{-3} [1]
Since the concentration of $H^+(aq)$ is less than the concentration of H_2CO_3, H_2CO_3 undergoes partial dissociation. Hence, H_2CO_3 is a weak acid. [1]

b Increase in
$[H^+(aq)] = 10^{-8.14} - 10^{-8.25}$
$= (7.24 \times 10^{-9}) - (5.62\times 10^{-9})$ mol dm^{-3}
$= 1.62 \times 10^{-9}$ mol dm^{-3} [1]

Percentage increase in $[H^+(aq)] = \frac{1.62 \times 10^{-9}}{5.62 \times 10^{-9}}$ 100% = 28.8% [1]

c Acid rain arising from dissolved oxides of nitrogen (NO_x) and SO_2 in the atmosphere. [1]

4 **a** volume NaOH = 19.00 cm^3 [1]

amount of sodium hydroxide $= \frac{19.00}{1000} \times 0.50 = 9.50 \times 10^{-3}$ mol [1]

b

$K_a = \frac{[H^+][A^-]}{[HA]}$ [1]

c Amount of weak monoprotic acid = 9.50×10^{-3} mol [1]

initial concentration of weak acid $= 9.50 \times 10^{-3} \div \frac{20}{1000} = 0.475$ mol dm^{-3} [1]

$[H^+] = 10^{-2.10} = 0.00794$ mol dm^{-3}

$[H^+] = [A^-] = 0.00794$ mol dm^{-3} [1]

$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{[0.00794][0.00794]}{[0.475]}$$

$= 1.33 \times 10^{-4}$ mol dm^{-3} [1]

d Cresol red [1]

Its pH transition range lies within the sharp pH change over the equivalence point. [1]

5 **a** Acidic because it is hydrogen ion donor and basic because it is a hydrogen ion acceptor. [1]

Such substances are described as amphiprotic. [1]

b water molecule, H_2O [1]

c pH is the negative logarithm to the base of ten of the hydrogen ion molar concentration. $pH = -\log_{10}[H^+(aq)]$ [1]

7.00 [1]

pOH is the negative logarithm to the base of ten of the hydroxide ion molar concentration. $pOH = -\log_{10}[OH^-(aq)]$ [1]

7.00 [1]

d **i** $K_w = [H^+(aq)] \times [OH^-(aq)]$ [1]

ii Water is more ionized at a higher temperature. [1]

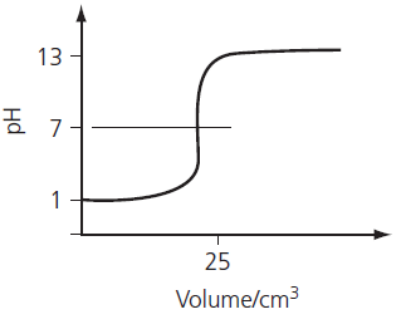
Le Chatelier's principle implies that endothermic reactions are favoured by increases in temperature. [1]

Hence the ionization of water molecules is endothermic. [1]

iii None / unchanged [1]

6

a



i The curve should include the following:
starting pH = 1 [1]
equivalence point: 25.00 cm^3 of KOH [1]
pH at equivalence point = 7 [1]
pH to finish = 12 – 13 [1]

ii $K_a = 10^{-4.76} = 1.74 \times 10^{-5}$ [1]
$K_a = [H^+(aq)]^2 \div [CH_3COOH]$ / $1.74 \times 10^{-5} = \frac{[H^+(aq)]^2}{0.100}$[1]
$[H^+(aq)] = 1.32 \times 10^{-3}$ (mol dm^{-3}); [1]
starting pH = 2.88 [1]
pH at equivalence point: 8–9; [1]
The pH at the equivalence is above 7 because of the equilibrium formed between the ethanoate ions in the salt produced and water molecules. $CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$.This leads to the formation of a small OH^- which reduced the concentration of H^+ and raises the pH. This phenomenon is known as salt hydrolysis . [1]

b i $HIn \rightleftharpoons H^+ + In^-$ [1]
and two colours indicated / or one colour and colourless: [1]
In acid the equilibrium moves to the left or vice versa [1]

ii Phenolphthalein / phenol red / bromothymol blue;
the colour change of indicator occurs within the range of pH at the equivalence point / on vertical part of graph; [2]

R3.2 Electron transfer reactions

Paper 1 style questions

1	D	**11**	C	**21**	A
2	B	**12**	B	**22**	C
3	D	**13**	B	**23**	A
4	C	**14**	C	**24**	A
5	A	**15**	D	**25**	A
6	C	**16**	A	**26**	D
7	B	**17**	B	**27**	C
8	D	**18**	B	**28**	B
9	D	**19**	C		
10	B	**20**	C		

Paper 2 style questions

1 a i magnesium (Mg), iron (Fe), copper (Cu), and gold (Au). [1]

ii the loss or removal of electrons from a chemical species [1]
For example, iron atoms lose electrons during their reaction with copper(II):
ions: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^{-}$ [1]

iii Reduction involves a decrease in oxidation state. [1]

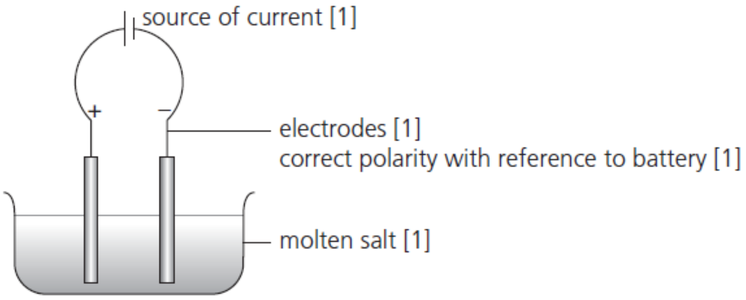
For example: copper(II) ions gain electrons during their reaction with iron atoms:
$Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$ [1]

iv Magnesium (Mg) is the strongest reducing agent. [1]
It reduces Fe^{2+}; iron(Fe) reduces Cu^{2+}; Cu reduces Au^{+}. [1]

v gold ions, Ag^{+} [1]
Every metal present can reduce gold ions to gold atoms. [1]

vi Does not react. [1]
It is too low in the reactivity series and hence is less reactive than magnesium. [1]

b



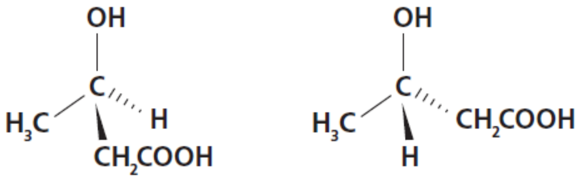
c at the cathode (negative electrode) [1]
electrons are given to the copper(II) cations (positive ions) [1]

electrons are removed from the iodide anions (negative ions) at the anode (positive electrode) [1]
copper atoms formed (at the cathode) / $Cu^{2+} + 2e^- \rightarrow Cu$; [1]
iodine molecules formed (at the anode) / $2I^- \rightarrow I_2 + 2e^-$. [1]

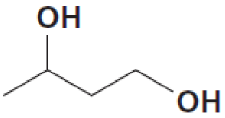
d Electrons flow through the external circuit or wires; [1]
ions (cations and anions) move through the electrolyte. [1]

2 a hydroxyl, carboxyl [2]

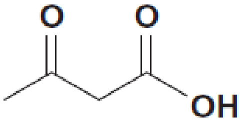
b [2]



c [1]



d i [1]



ii propanone [1]

3 a Oxidation state of chromium is changed from +6 ($Cr_2O_7^{2-}$) to +3 (Cr^{3+}); hence $Cr_2O_7^{2-}$ is reduced. [1]

b $C_2H_5OH + H_2O \rightarrow CH_3CO_2H + 4H^+ + 4e^-$ [1]

c $3C_2H_5OH + 2Cr_2O_7^{2-} + 16H^+ \rightarrow 3CH_3CO_2H + 4Cr^{3+} + 11H_2O$ [1]

d Amount of $Cr_2O_7^{2-}$ reacted $= \frac{37.50}{1000} \times 1.50 = 0.05625$ mol [1]

Amount of ethanol in 25.00 $cm^3 = \frac{3}{2} \times 0.056\,25 = 0.084\,38$ mol [1]

Mass of ethanol in 25.00 cm^3 = 0.084 38 mol × 46.08 g mol^{-1} = 3.888 g [1]

Volume of ethanol in 25.00 cm^3 = (3.888 g / 0.790 g cm^{-3}) = 4.92 cm^3 [1]

e Volume of ethanol in 100.00 cm^3 = (4.92 cm^3 × 4) = 19.7 cm^3 [1]

Concentration of ethanol = 19.7 % [1]

4 a Anode: $CH_3CH(OH)CH_3 + 5H_2O \rightarrow 3CO_2 + 18H^+ + 18e^-$ [1]
Cathode: $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$ [1]

b $2CH_3CH(OH)CH_3 + 9O_2 \rightarrow 6CO_2 + 8H_2O$ [1]

c $1.56\ V = +1.23\ V - E^{\ominus}_{anode}$; $E^{\ominus}_{anode} = -0.33\ V$ [1]

d Propan-2-ol is a liquid but hydrogen is a gas under standard conditions. Propan-2-ol is easier to store and transport. [1]

5 a To conduct electricity [1]

b Cathode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ [1]
Anode: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$ [1]

c Volume of copper layer = $10 \times \frac{0.5}{10} = 0.5\ \text{cm}^3$ [1]

Mass of copper layer = $8.96\ \text{g cm}^{-3} \times 0.5\ \text{cm}^3 = 4.48\ \text{g}$ [1]

Amount of copper atoms in the layer = $\frac{4.48\ \text{g}}{63.55\ \text{mol}^{-1}} = 0.0705\ \text{mol}$

Amount of electrons required = $(2 \times 0.0705\ \text{mol}) = 0.141\ \text{mol}$ [1]

Amount of charge required = $0.141\ \text{mol} \times 9.65 \times 10^4\ \text{C mol}^{-1} = 13\,607\ \text{C}$ [1]

Time required to copper plate = $\frac{13\,607\ \text{C}}{20\ \text{A}} = 680\ \text{s}$. [1]

6 Chlorine is produced at high concentration / oxygen is produced at low concentration. [1]

At low concentrations of Cl^-, $E^\ominus$ (O_2/H_2O) is more positive than $E^\ominus$ (Cl_2/Cl^-); hence H_2O is oxidized to oxygen, O_2. [1]

At high concentration of Cl^-, Cl^- is oxidized in preference over water molecules, H_2O. [1]

7 **a** Oxidation: loss of electrons [1] and reduction: gain of electrons [1].

b Oxidizing agent: substance reduced / removes electrons from another substance / causes another substance to be oxidized [1]

Reducing agent: substance oxidized / gains electrons from another substance / causes another substance to be reduced [1]

c $E^\ominus$ (H_2O_2/H_2O) is very / more positive [1]

hence more likely to gain electrons / undergo reduction [1]

d $E^\ominus = +1.10\ \text{V}$; [1]

the reaction is feasible/spontaneous as $E^\ominus > 0$ [1]

8 **a** **i** The diagram must show two electrodes (labelled as nickel and manganese) must be dipping into aqueous solutions labelled with their ions: $Ni^{2+}(aq)$ and $Mn^{2+}(aq)$. [1]
The two electrodes must be connected via a salt bridge. [1]
and an external circuit consisting of wire and a voltmeter. [1]

ii Hydrogen gas at 100 kPa pressure passed over a platinum electrode. [1]
The temperature of all chemicals is 298 K [1]
and the hydrogen ion (proton) concentration is 1 mol dm^{-3}. [1]

b **i** $Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$ [1]
$Mn(s) \rightarrow Mn^{2+}(aq) + 2e^-$ [1]
Manganese atoms are oxidized. [1]
Manganese(II) ions are the oxidizing agent. [1]

ii Manganese is the anode. [1]
The electron flow is from manganese to nickel. [1]

iii $E^\ominus_{\text{cell}} = +0.92\ \text{V}$ [1]
ΔG is negative. [1]

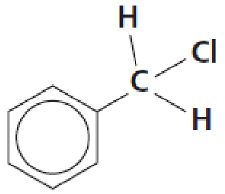
R3.3 Electron sharing reactions

Paper 1 style questions

1	C	5	B	9	C
2	C	6	D	10	D
3	A	7	A	11	A
4	B	8	D		

Paper 2 style questions

1 **a** [1]



b Initiation: $Cl_2 \rightarrow 2Cl\bullet$ [1]
Propagation: $C_6H_5CH_3 + Cl\bullet \rightarrow C_6H_5CH_2\bullet + HCl$ [1]
$C_6H_5CH_2\bullet + Cl_2 \rightarrow C_6H_5\ CH_2Cl + Cl\bullet$ [1]

UV light initiates the reaction by causing homolytic fission of the chlorine molecule. [1]

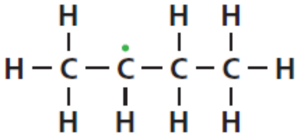
c $C_{14}H_{14}$ diphenylethane [1]

2 **a** radical substitution [1]
hydrogen bromide [1]

b $C_4H_{10} + 2Br_2 \rightarrow C_4H_8Br_2 + 2HBr$ [1]

c 6 [1]

d [1]



3 **a** A species with one or more unpaired electrons. [1]

b

[1]

c homogeneous catalysis [1]
Iron can exist in variable oxidation states (+2/+3). [1]

R3.4 Electron pair sharing reactions

Paper 1 style questions

1	C	**8**	D	**15**	B
2	A	**9**	A	**16**	C
3	C	**10**	A	**17**	C
4	D	**11**	B	**18**	A
5	D	**12**	D	**19**	C
6	B	**13**	D	**20**	A
7	D	**14**	A	**21**	D

Paper 2 style questions

1 **a** [1]

b **i** electrophilic addition. [1]
chloroethane [1]

ii hydroxide ions / sodium hydroxide / potassium hydroxide [1]

2 Ethanol – use of ethanol or other water miscible solvent (added to each halogenoalkane / liquid in separate containers) [1]

Fair test – use of equal volumes/amounts / specified volumes/amounts in each tube or warm the tubes in a water bath / specified temperature / room temperature [1]

Silver nitrate - silver nitrate (solution) /silver ions, $Ag^+(aq)$ to each tube (of halogenoalkane) [1]

Time – find the time taken for a precipitate to form / obscure a dot or cross on test-tube. [1]

Rate (1/time) – expected trend is 2-iodobutane > 2-bromobutane > 2-chlorobutane or 2-iodobutane is the fastest and 2-chlorobutane is the slowest [1]

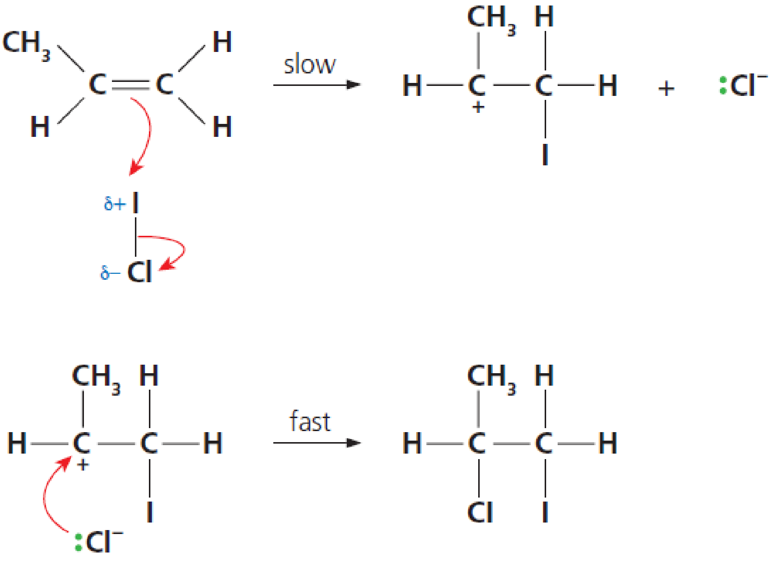
C–X Bond enthalpy linked to rate of hydrolysis. C–I < C–Br < C–Cl [1]

3 **a** The I–Cl bond has a permanent dipole. Since the iodine side (has a delta positive charge and) is more attracted to the electron-rich carbon=carbon double bond.

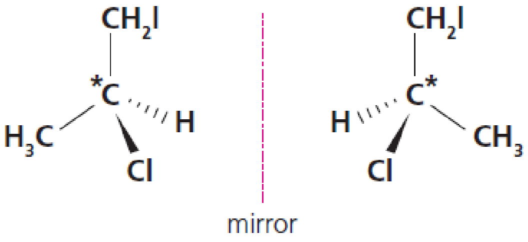
OR

I–Cl is a stronger electrophile because it has a permanent dipole and is more attracted to the >C=C< bond with the pi pair of electrons. [1]

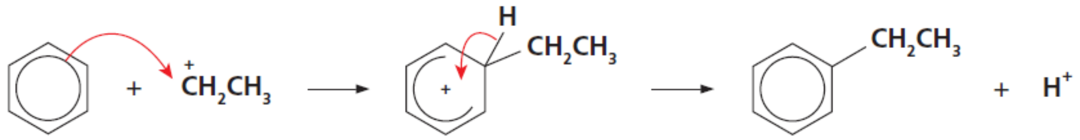
b electrophilic addition [1]
iodine atom acting as the electrophile [1]
two correct steps using curly arrows of the mechanism [2]
Major product shown with Cl on the middle carbon [1]



c [1]



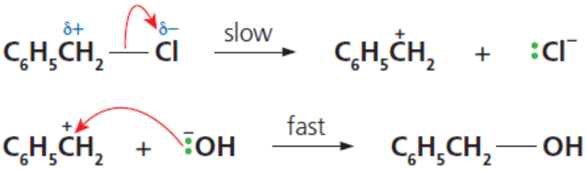
4 **a** Z = chloroethane / bromoethane / iodoethane [1]



curly arrow from benzene ring to carbocation. [1]

correct structure of intermediate and curly arrow from C–H bond back to form ring [1]

b **i** S_N1 [1]



Step 1 showing curly arrow and formation of carbocation intermediate [1]

Step 2: arrow from hydroxide ion to carbocation and formation of product [1]

ii The benzyl cation, $C_6H_5CH_2^+$, intermediate formed is stable. [1]
It is stabilized by delocalisation of the pi electrons from the benzene ring, hence favouring the S_N1 mechanism. [1]

c **i** coordination bond / coordinate bond [1]
ii A Lewis acid is an electron pair acceptor. [1]
Cu^{2+} is the Lewis acid [1]