WORKBOOK

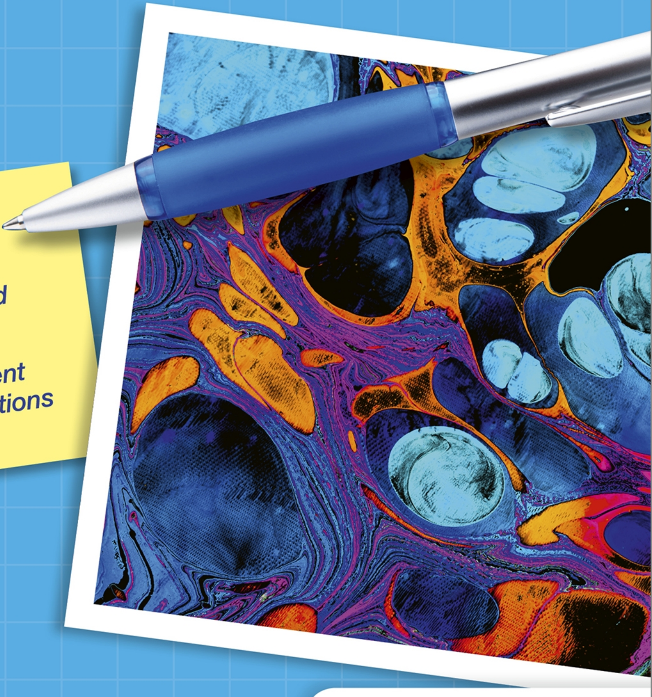
AQA A-LEVEL

Chemistry 2

YEAR 2 TOPICS

- Physical Chemistry
- Inorganic Chemistry
- Organic Chemistry

- ✓ Build confidence with practice questions
- ✓ Practise key maths and practical skills
- ✓ Prepare for assessment with exam-style questions



Nora Henry

Alyn McFarland

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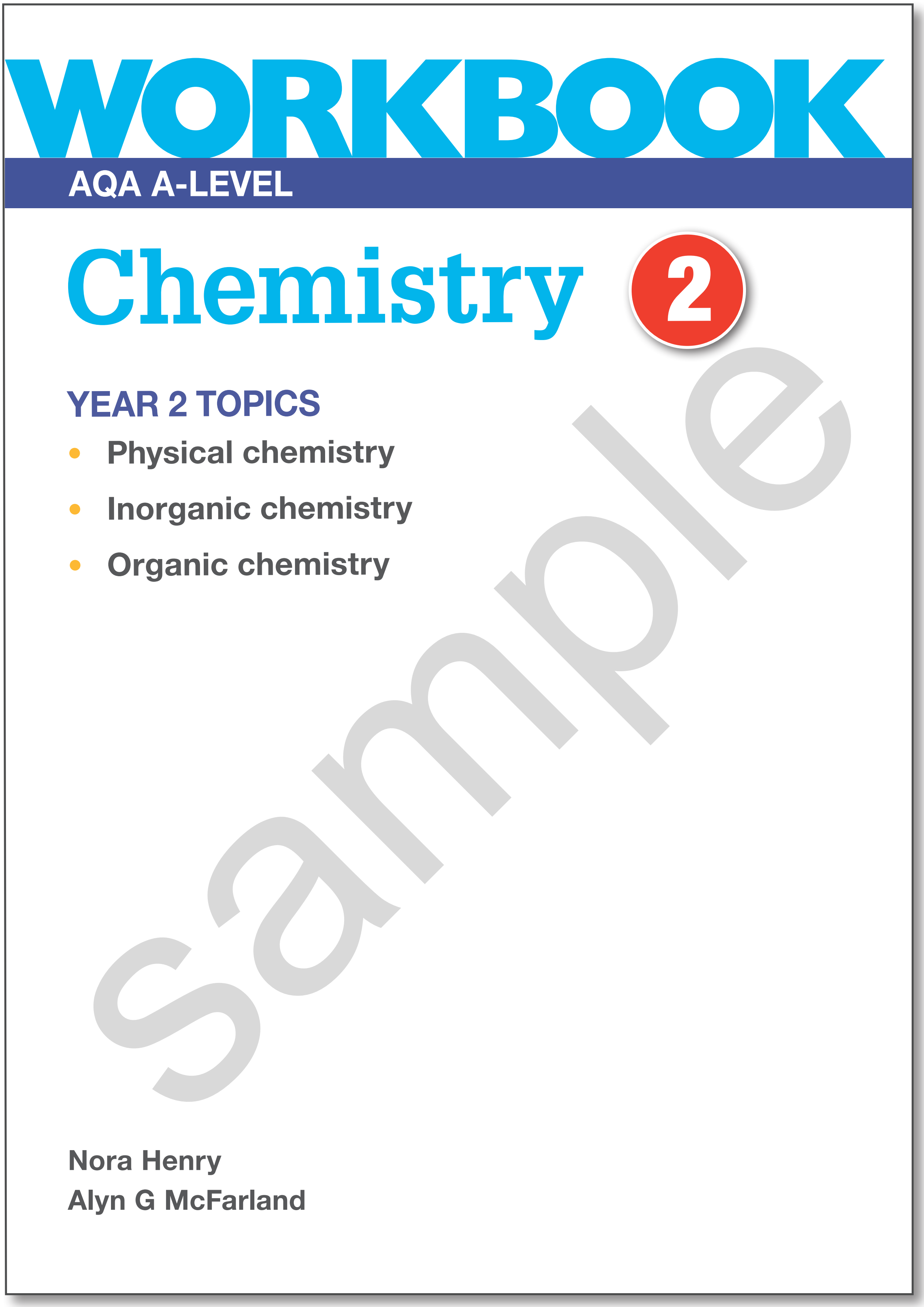

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Nora Henry
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Contents

1 Physical chemistry

2 Inorganic chemistry

3 Organic chemistry

About this book

1. **This workbook** will help you to prepare for AQA A-level Chemistry Topics 3.1.8–3.1.12, 3.2.4–3.2.6 and 3.3.7–3.3.16.

2. **These topics** could be assessed in:
   - A-level Paper 1, which lasts for 2 hours. Paper 1 is worth 35% of the A-level. There is a mixture of short- and long-answer questions, worth 105 marks.
   - A-level Paper 2, which lasts for 2 hours. Paper 2 is worth 35% of the A-level. There is a mixture of short- and long-answer questions, worth 105 marks.
   - A-level Paper 3, which lasts for 2 hours and covers any content and any practical skills. Paper 3 is worth 30% of the A-level. Questions on practical techniques and data analysis are worth 40 marks. There are 20 marks for questions testing across the specification and 30 marks for multiple-choice questions, worth a total of 90 marks.

3. **For each topic** in this workbook there are:
   - stimulus materials, including key terms and concepts
   - short-answer questions that build up to short- and long-answer questions and multiple-choice questions in exam-style
   - spaces for you to write or plan your answers
   - questions that test your mathematical skills

4. **Answering the questions** will help you to build your skills and meet the assessment objectives AO1 (knowledge and understanding), AO2 (application) and AO3 (analysis). Mathematical skills make up a minimum of 20% of the total marks across the A-level.

5. **Worked answers** are included for some practice questions to help you understand how to gain the most marks.

6. **Marks available** are indicated for all questions so that you can gauge the level of detail required in your answers.

7. **Timings** are given for the exam-style questions to make your practice as realistic as possible.

8. Icons next to a question will help you to identify:

   ⚗ where the practical elements of the course are covered

   🖩 where your calculation skills are tested

   ⚙ where questions draw on synoptic knowledge, i.e. content from more than one topic

   60 how long these questions should take you

9. **You still need** to read your textbook and refer to your revision guides and lesson notes.

10. **Answers** are available at: **www.hoddereducation.co.uk/workbookanswers.**

1 Physical chemistry

3.1.8 Thermodynamics

Born–Haber cycles

Lattice enthalpy may be enthalpy of lattice formation (exothermic) or enthalpy of lattice dissociation (endothermic).

A **Born–Haber cycle** links the lattice enthalpy and enthalpy of formation to enthalpy changes of atomisation (bond enthalpy), ionisation and electron affinity.

The **enthalpy of solution** of an ionic compound may be calculated from lattice enthalpy and hydration enthalpy values of the ions.

The **perfect ionic model** gives a theoretical value of the lattice enthalpy.

Practice questions ?

1 Which one of the following changes is exothermic? **1 mark**

A $Na(s) \rightarrow Na(g)$ ○

B $Na(g) \rightarrow Na^+(g) + e^-$ ○

C $O(g) + e^- \rightarrow O^-(g)$ ○

D $O^-(g) + e^- \rightarrow O^{2-}(g)$ ○

2 The following changes occur in the Born–Haber cycle for sodium fluoride.

Enthalpy change	ΔH/kJ mol^{-1}
$Na(s) \rightarrow Na(g)$	+108
$Na(g) \rightarrow Na^+(g) + e^-$	+500
$\frac{1}{2}F_2(g) \rightarrow F(g)$	+79
$F(g) + e^- \rightarrow F^-(g)$	−348
$Na(s) + \frac{1}{2}F_2(g) \rightarrow NaF(s)$	−569

Calculate the lattice enthalpy (dissociation) for sodium fluoride.

Worked example

The enthalpy of lattice dissociation is for the following change:

$$NaF(s) \rightarrow Na^+(g) + F^-(g)$$

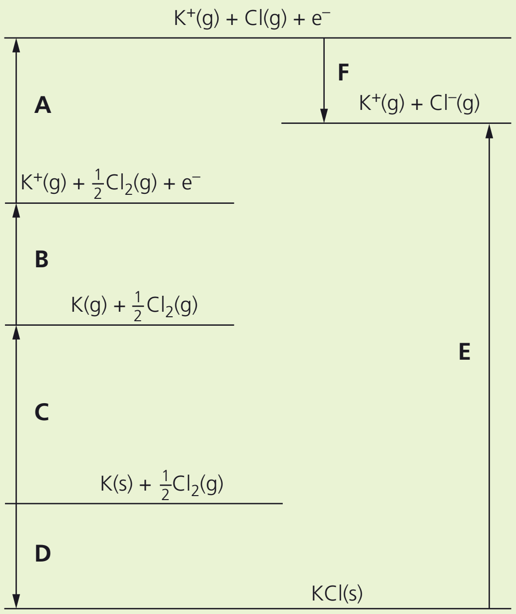
So, reversing the formation, and atomising and ionising the elements, will give you the same change.

$$\Delta_{LE}H = -\Delta_f H + \Delta_a H + \Delta_{IE1} H + \Delta_a H + \Delta_{EA1} H \checkmark$$

$$\Delta_{LE}H = -(-569) + 108 + 500 + 79 + (-348) = +908\,(\text{kJ mol}^{-1}) \checkmark$$

Watch for data supplied for $F_2 \rightarrow 2F$, as this is the bond dissociation enthalpy or 2× enthalpy of atomisation of fluorine, and you need to take into account the moles of F required. Also, in this case, you only need 1× first electron affinity as there is 1 mol of F^- in NaF. Be sure you can identify the various changes around a Born–Haber cycle and the number of moles to which they refer.

3 The Born–Haber cycle for potassium chloride is shown below. The enthalpy changes are not drawn to scale.



a Which letter represents the enthalpy of atomisation of potassium? **1 mark**

....

b Which letter represents the enthalpy of lattice dissociation of potassium chloride? **1 mark**

....

c Which letter represents the first ionisation energy of potassium? **1 mark**

....

d Which letter represents the first electron affinity of chlorine? **1 mark**

....

e The following values are given:

Enthalpy change	ΔH/kJ mol^{-1}
A	+121
B	+420
C	+89
D	−437
E	+703

Calculate a value for F. **3 marks**

....

....

....

....

4 The table below gives data for the Born–Haber cycle for magnesium oxide.

Enthalpy change	**ΔH/kJ mol^{-1}**
Enthalpy of atomisation of magnesium	+150
First ionisation energy of magnesium	+736
Second ionisation energy of magnesium	+1450
Bond dissociation enthalpy of oxygen	+496
First electron affinity of oxygen	–142
Second electron affinity of oxygen	+844
Enthalpy of formation of magnesium oxide	–602

a Write equations, including state symbols, which represent the following changes in the Born–Haber cycle.

i first ionisation energy of magnesium **1 mark**

....

ii second ionisation energy of magnesium **1 mark**

....

iii enthalpy of formation of magnesium oxide **1 mark**

....

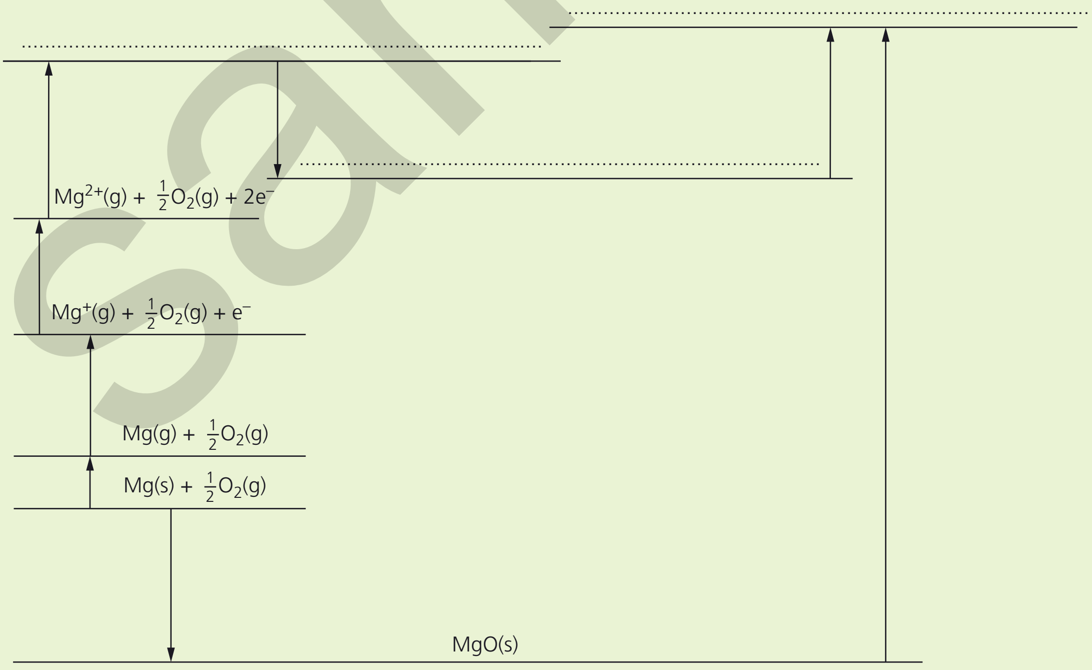
iv second electron affinity of oxygen **1 mark**

....

v enthalpy of lattice dissociation of magnesium oxide **1 mark**

....

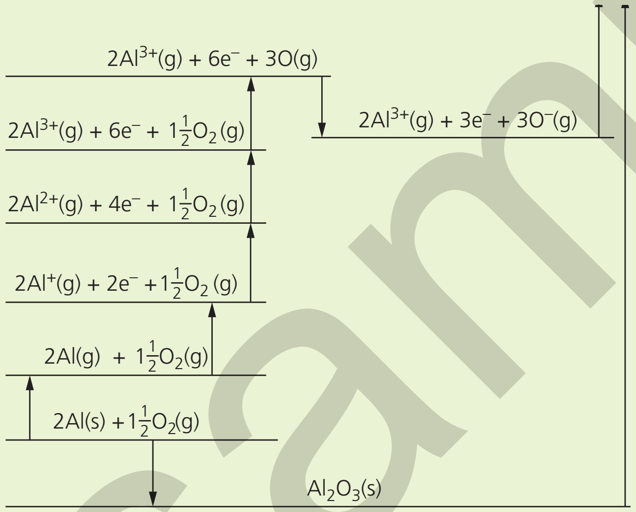
b The Born–Haber cycle below is for magnesium oxide.



i Complete the Born–Haber cycle by showing what is present at the three empty levels. **3 marks** ii Using the constructed Born–Haber cycle, calculate the enthalpy of lattice dissociation of magnesium oxide. **3 marks**

5 The table and diagram below give the enthalpy changes and the Born–Haber cycle for aluminium oxide.

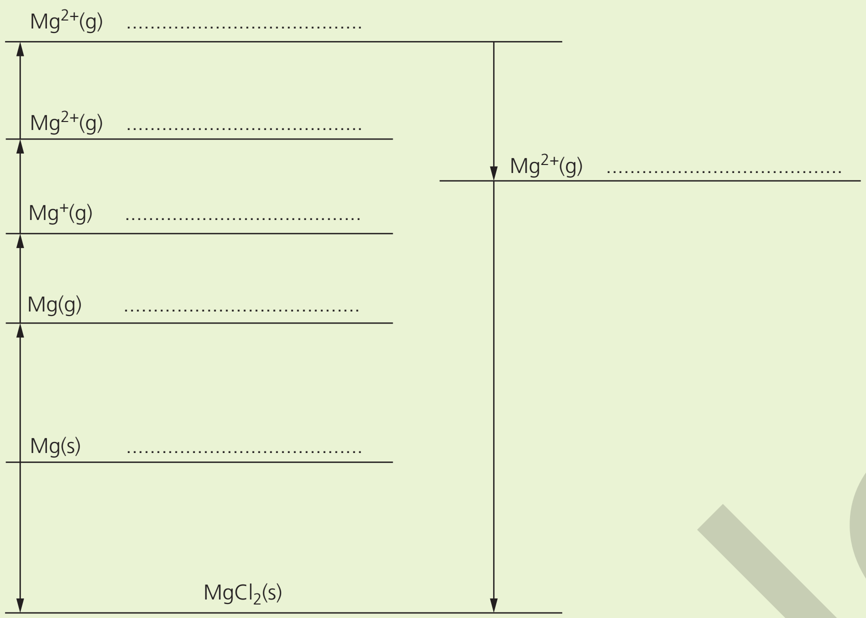
	Enthalpy change	**ΔH/kJ mol^{-1}**
A	Enthalpy of atomisation of oxygen	+248
B	First electron affinity of oxygen	–142
C	Second electron affinity of oxygen	+844
D	First ionisation enthalpy of aluminium	+580
E	Second ionisation enthalpy of aluminium	+1800
F	Third ionisation enthalpy of aluminium	To be calculated
G	Enthalpy of atomisation of aluminium	+324
H	Enthalpy of lattice dissociation of aluminium oxide	+15 180
I	Enthalpy of formation of aluminium oxide	–1676



a Label the enthalpy changes on the Born–Haber cycle with the letters A to I from the table above. **4 marks**

b Calculate a value for the third ionisation enthalpy of aluminium. **4 marks**

6 The diagram below shows the Born–Haber cycle for magnesium chloride.



a Complete the Born–Haber cycle, giving the species present at each level. Include state symbols. **6 marks**

b The table below gives some enthalpy values associated with magnesium chloride.

Enthalpy change	$\Delta H/\mathrm{kJ\,mol^{-1}}$
Enthalpy of atomisation of magnesium	+150
First electron affinity of chlorine	−364
First ionisation energy of magnesium	+736
Second ionisation energy of magnesium	+1450
Enthalpy of formation of magnesium chloride	−642
Lattice enthalpy (formation) of magnesium chloride	−2492
Enthalpy of hydration of magnesium ions	−1920
Enthalpy of hydration of chloride ions	−364

i Use the data in the table to calculate a value for the enthalpy of atomisation of chlorine. **3 marks**

....

....

....

ii Calculate the enthalpy of solution of magnesium chloride. **3 marks**

....

....

....

iii State and explain whether you would expect a calcium ion to have a higher or lower hydration enthalpy than a magnesium ion. **2 marks**

7 The enthalpy values in the table below are for silver chloride.

Enthalpy change	ΔH/kJ mol^{-1}
$Ag(s) \rightarrow Ag(g)$	+289
$Ag(g) \rightarrow Ag^+(g) + e^-$	+732
$Cl_2(g) \rightarrow 2Cl(g)$	+242
$Cl(g) + e^- \rightarrow Cl^-(g)$	−364
$AgCl(s) \rightarrow Ag^+(g) + Cl^-(g)$	+905

a Use the data in the table to calculate a value for the enthalpy of formation of silver chloride. **2 marks**

b A value for the enthalpy of lattice dissociation of silver chloride calculated using the perfect ionic model is +883 kJ mol^{-1}.

i Explain what assumptions are made in the perfect ionic model. **2 marks**

ii Explain what may be deduced by comparing the experimental value of the enthalpy of lattice formation of silver chloride given in the table with the theoretical value calculated using the perfect ionic model. **2 marks**

Gibbs free-energy change, ΔG, and entropy change, ΔS

Entropy (S) is a measure of the disorder of a system. **ΔS** is the change in entropy.

S and ΔS are both measured in J K^{-1} mol^{-1}.

Gibbs free energy change is $\Delta G = \Delta H - T\Delta S$, where a reaction is feasible if $\Delta G = 0$ or is less than zero.

Calculations involving ΔG use temperature (T) in kelvin (K), entropy change (ΔS) in kJ K^{-1} mol^{-1} and enthalpy change (ΔH) in kJ mol^{-1}.

A graph of ΔG against T has an intercept on the ΔG axis of ΔH, and the intercept on the horizontal axis shows the temperature at which a reaction becomes feasible. The gradient of the line is $-\Delta S$.

Practice questions

8 Which one of the following reactions shows a decrease in entropy? **1 mark**

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

B $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$ ○

C $Na_2CO_3(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$ ○

D $NaCl(s) + H_2SO_4(l) \rightarrow NaHSO_4(s) + HCl(g)$ ○

9 Which one of the following reactions would be feasible at any temperature? **1 mark**

A an endothermic reaction with a decrease in entropy ○

B an endothermic reaction with an increase in entropy ○

C an exothermic reaction with a decrease in entropy ○

D an exothermic reaction with an increase in entropy ○

10 Strontium carbonate decomposes on heating, as shown by the following equation.

$SrCO_3(s) \rightarrow SrO(s) + CO_2(g)$

The table below shows the enthalpy of formation and standard entropy values of the reactants and products.

Substance	$SrCO_3(s)$	$SrO(s)$	$CO_2(g)$
$\Delta_f H$/kJ mol^{-1}	–1218	–590	–394
S/J K^{-1} mol^{-1}	97.1	54.4	214

Calculate the enthalpy change and entropy change for the reaction, and the temperature at which this thermal decomposition becomes feasible.

Worked example

$\Delta H = (-590) + (-394) - (-1218) = +234\,\text{kJ mol}^{-1}$ ✓

$\Delta S = 54.4 + 214 - 97.1 = +171.3\,\text{J K}^{-1}\,\text{mol}^{-1}$ ✓

$\Delta G = \Delta H - T\Delta S$

$\Delta G = 0$ when reaction becomes feasible

$\Delta H - T\Delta S = +234 - T(0.1713) = 0$

$T = \dfrac{234}{0.1713} = 1366\,\text{K}$ ✓

Remember that when using enthalpy of formation data:

$\Delta H = \Sigma\Delta_f H(\text{products}) - \Sigma\Delta_f H(\text{reactants})$

Also, that

$\Delta S = \Sigma S(\text{products}) - \Sigma S(\text{reactants})$

$\Delta G = 0$ as the reaction becomes feasible.

Note that ΔS has units of J K^{-1} mol^{-1} but, to be used in the $\Delta G = \Delta H - T\Delta S$ equation, it is converted to kJ K^{-1} mol^{-1} by dividing by 1000.

Temperature is measured in kelvin (K).

11 When lead(II) nitrate is heated, it decomposes according to the following equation.

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

The table below gives the absolute entropy values and the enthalpy of formation values for all the substances in the reaction.

Substance	$Pb(NO_3)_2(s)$	$PbO(s)$	$NO_2(g)$	$O_2(g)$
S/J K^{-1} mol^{-1}	217.9	68.7	240	205
$\Delta_f H$/kJ mol^{-1}	−451.9	−217.3	+33.2	0

a Explain why the enthalpy of formation of oxygen is zero. **1 mark**

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b Calculate ΔS for this reaction. **3 marks**

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c Calculate ΔH for this reaction. **3 marks**

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d Show that this reaction is not feasible at 500 K. **4 marks**

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e Calculate the temperature at which this reaction becomes feasible. **3 marks**

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12 The enthalpy of fusion (melting) of magnesium is +9 kJ mol^{-1}. The absolute entropy values of magnesium as a solid and a liquid are 32.7 J K^{-1} mol^{-1} and 42.7 J K^{-1} mol^{-1}.

Calculate the melting point of magnesium in kelvin (K). **4 marks**

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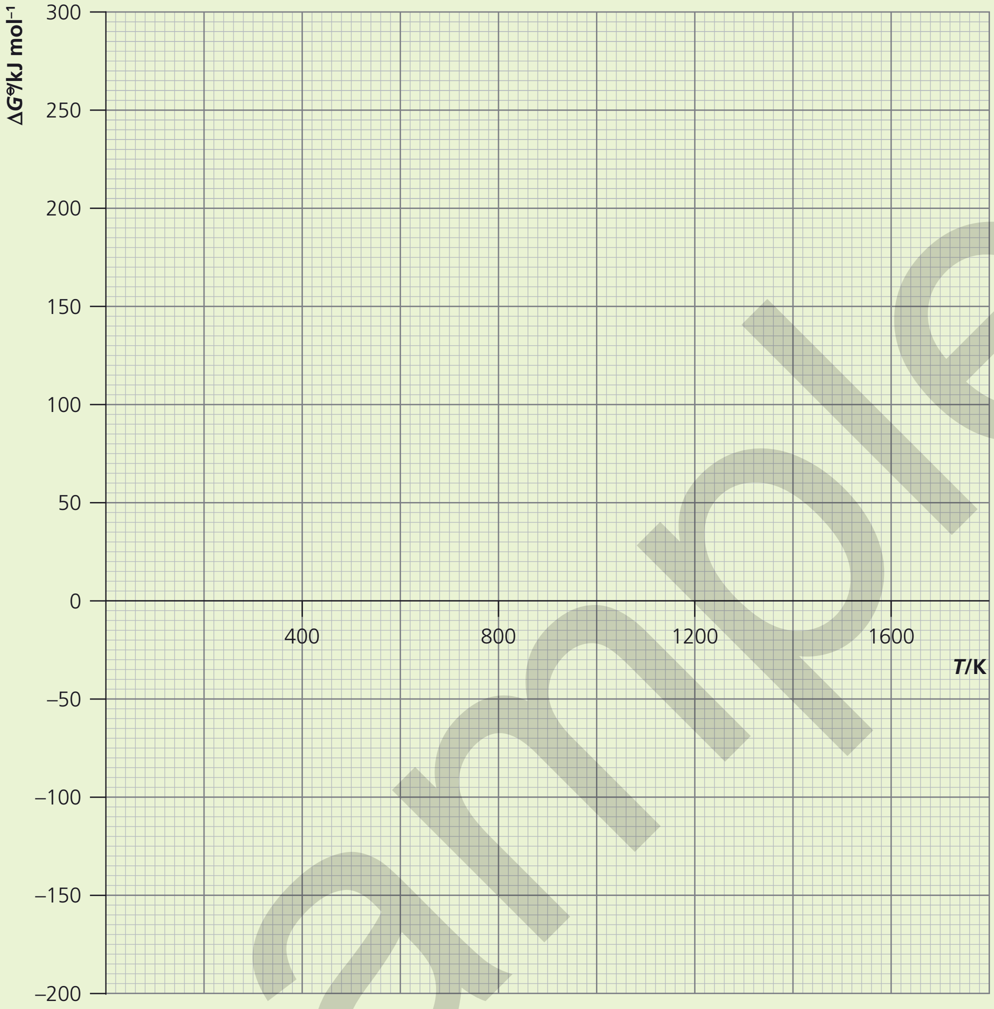
13 When heated strongly, copper(II) sulfate decomposes according to the following equation.

$CuSO_4(s) \rightarrow CuO(s) + SO_3(g)$

a On the axes below, plot a graph of ΔG against T using the following values. **3 marks**

T/K	400	800	1200	1600
ΔG/kJ mol^{-1}	+145	+71	−4	−78



b Calculate the gradient of the line you have drawn. Give your answer to 3 significant figures. **2 marks**

c What is the intercept on the ΔG axis? **1 mark**

d State a value for ΔH for this reaction. Include units in your answer. **2 marks**

e State a value for ΔS for this reaction. Include units in your answer. **2 marks**

f At what temperature does the reaction become feasible? **1 mark**

....

35

Exam-style questions ?

1 a The enthalpy changes shown below relate to magnesium fluoride.

Enthalpy change	ΔH/kJ mol^{-1}
$Mg(s) \rightarrow Mg(g)$	+150
$Mg(g) \rightarrow Mg(g) + e^-$	+736
$Mg^+(g) \rightarrow Mg^{2+}(g) + e^-$	+1450
$F_2(g) \rightarrow 2F(g)$	+158
$F(g) + e^- \rightarrow F^-(g)$	–348
$MgF_2(s) \rightarrow Mg^{2+}(g) + 2F^-(g)$	+2883
$MgF_2(s) \rightarrow Mg^{2+}(aq) + 2F^-(aq)$	–20
$Mg^{2+}(g) \rightarrow Mg^{2+}(aq)$	–1891

i Calculate the enthalpy of formation of magnesium fluoride. **3 marks**

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....

ii Calculate the hydration enthalpy of the fluoride ion. **3 marks**

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....

....

b The theoretical value for the enthalpy of lattice dissociation of magnesium fluoride calculated from the perfect ionic model is +2881 kJ mol^{-1}. The experimental value is given in the table above.

i Explain what you understand by the *perfect ionic model*. **1 mark**

....

....

ii Explain the bonding in magnesium fluoride based on these values. **2 marks**

....

....

....

2 The table below shows some absolute entropy values and enthalpies of formation of some substances in different states.

Substance	S/J K^{-1} mol^{-1}	$\Delta_f H$/kJ mol^{-1}
Al(s)	28	0
H_2O(l)	70	–286
H_2O(g)	189	–242
Al_2O_3(s)	51	–1676
H_2(g)	131	0
O_2(g)	205	0
MgO(s)	27	–601
MgO(l)	48	To be calculated

a Aluminium reacts with water vapour according to the following equation.

$$2Al(s) + 3H_2O(g) \rightarrow Al_2O_3(s) + 3H_2(g)$$

i Calculate the enthalpy change for this reaction. **3 marks**

....

....

....

ii Calculate the entropy change for this reaction. **3 marks**

....

....

....

iii Show that the reaction of aluminium with water vapour is feasible at 500 K. **3 marks**

....

....

....

b The melting point of magnesium oxide is 3125 K. Calculate a value for the enthalpy of formation of MgO(l). **6 marks**

....

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3.2.6 Reactions of ions in aqueous solution

Many metal hydroxides and carbonates are insoluble in water and form **precipitates** when solutions containing these ions are mixed.

The observations, including the colour of the precipitate, can be used to identify the metal ions present in the solution.

Hydroxide ions are present in sodium hydroxide solution and ammonia solution.

Practice questions ?

1 This question relates to required practical 11.

The table below gives details of three different hexaaqua complexes.

Complex	Observations when sodium hydroxide solution is added	Observations when ammonia solution is added	Observations when sodium carbonate solution is added
A	White precipitate that dissolves in excess sodium hydroxide solution	White precipitate that does not dissolve in excess ammonia solution	White precipitate Bubbles of gas
B	Blue precipitate that does not dissolve in excess sodium hydroxide solution	Blue precipitate that dissolves in excess ammonia solution to form a deep-blue solution	Blue–green precipitate
C	Green precipitate that does not dissolve in excess sodium hydroxide solution	Green precipitate that does not dissolve in excess ammonia solution	Green precipitate

a Identify the complexes A, B and C. **3 marks**

A: ....

B: ....

C: ....

b Write an equation for the formation of the white precipitate when A reacts with sodium hydroxide solution. **1 mark**

....

c Give the formula of the complex which is responsible for the deep-blue solution when B reacts with excess ammonia solution. **1 mark**

....

d Write an equation for the formation of the green precipitate when C reacts with sodium carbonate solution. **1 mark**

....

e Name the gas given off when A reacts with sodium carbonate solution. **1 mark**

....

2 A solution contains the complex $[Fe(H_2O)_6]^{3+}$.

a Give the oxidation state of iron, the coordination number and the shape of this complex. **3 marks**

....

....

b What would be observed if sodium hydroxide solution was added to the solution containing the complex $[Fe(H_2O)_6]^{3+}$ until in excess? **2 marks**

....

....

c Write an equation for the reaction occurring in **b** between the complex $[Fe(H_2O)_6]^{3+}$ and hydroxide ions. **1 mark**

....

d What is observed when sodium carbonate solution is added to a solution of the complex $[Fe(H_2O)_6]^{3+}$? **2 marks**

....

....

e Write an equation for the reaction occurring in **d** between the complex $[Fe(H_2O)_6]^{3+}$ and carbonate ions. **2 marks**

....

f Explain, using an equation, why a solution containing $[Fe(H_2O)_6]^{3+}$ is acidic. **1 mark**

....

3 This question relates to required practical 11.

Three solutions, P, Q and R, were added to known solutions and the observations were recorded in the table below.

Solution	$FeSO_4$(aq)	$Al(NO_3)_3$(aq)	$CuCl_2$(aq)
P	Green precipitate	White precipitate that dissolves in excess of solution P to form a colourless solution	Blue precipitate
Q	White precipitate	No precipitate	No precipitate
R	Green precipitate	White precipitate and bubbles of a colourless gas	Blue–green precipitate

a Identify the solutions P, Q and R. **3 marks**

P: ....

Q: ....

R: ....

b Give the formula of the following:

i green precipitate when P is added to $FeSO_4$(aq) **1 mark**

....

ii white precipitate when Q is added to $FeSO_4$(aq) **1 mark**

....

iii white precipitate and colourless solution when P is added to $Al(NO_3)_3(aq)$ **2 marks**

....

....

iv blue precipitate when P is added to $CuCl_2(aq)$ **1 mark**

....

v colourless gas when R is added to $Al(NO_3)_3(aq)$ **1 mark**

....

vi blue–green precipitate when R is added to $CuCl_2(aq)$ **1 mark**

....

c Write an equation for the production of the white precipitate and colourless gas when R is added to $Al(NO_3)_3(aq)$. **2 marks**

....

25

Exam-style questions ?

1 This question relates to required practical 11.

Which one of the reactions below would produce a brown solid? **1 mark**

A $Fe(NO_3)_3(aq) + NaOH(aq)$ ○

B $CuSO_4(aq) + Na_2CO_3(aq)$ ○

C $FeSO_4(aq) + KOH(aq)$ ○

D $Al_2(SO_4)_3(aq) + Na_2CO_3(aq)$ ○

2 This question relates to required practical 11.

A solution is thought to contain $[Al(H_2O)_6]^{3+}$ ions.

a State the colour of the solution. **1 mark**

....

b A gas is produced when sodium carbonate solution is added to the solution.

i Name the gas produced. **1 mark**

....

ii Write an ionic equation for the reaction that produces the gas. **2 marks**

....

c What would be observed if potassium hydroxide solution was added to the solution slowly until in excess? Write equations for the reactions that occur. **4 marks**

....

....

3 This question relates to required practical 11.

a An inorganic solid, X, is tested, as shown below.

Test	Observations
(1) Add X to a test tube and add 10 cm^3 of deionised water.	Solid dissolves to form a blue solution.
(2) To the solution from test 1, add potassium hydroxide solution dropwise and then until in excess.	Blue precipitate forms which does not dissolve in excess potassium hydroxide solution.
(3) To the solution from test 1, add ammonia solution dropwise and then until in excess.	Blue precipitate forms which dissolves in excess ammonia solution to form a deep-blue solution.
(4) To the solution from test 1, add sodium carbonate solution.	A blue–green precipitate forms.

i Give the formula of the complex formed when solid X dissolves in water. **1 mark**

ii Write an equation for the formation of the blue precipitate in test 2. **1 mark**

iii Give the formula of the complex in the deep-blue solution in test 3. **1 mark**

iv Give the formula of the blue–green precipitate in test 4. **1 mark**

b The tests were repeated with solid Y, which is iron(III) chloride. Complete the table below to give the observations for each of the tests. **7 marks**

Test	Observations
(1) Add Y to a test tube and add 10 cm^3 of deionised water.	
(2) To the solution from test 1, add potassium hydroxide solution dropwise and then until in excess.	
(3) To the solution from test 1, add ammonia solution dropwise and then until in excess.	
(4) To the solution from test 1, add sodium carbonate solution.	

Total 20 marks