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Premium IB tuition — Physics, Chemistry, Maths

IB DIPLOMA CHEMISTRY · FIRST ASSESSMENT 2025

Reactivity 1
What Drives Chemical Reactions?

Paper 1 Style Practice · Paper 1A and Paper 1B

Medium to hard exam-style questions on Reactivity 1.1 – 1.4, with full worked solutions

Standard Level & Higher Level

Timing and marks

Higher Level: answer every question. Section A: 32 marks; Section B: 28 marks. Total **60 marks**. Suggested time **about 95 minutes** (45 min for Section A, 50 min for Section B).

Standard Level: omit the questions marked HL. Section A: 22 marks; Section B: 22 marks. Total **44 marks**. Suggested time **about 75 minutes**.

These timings use the same rate as the real examination (SL Paper 1: 55 marks in 1 h 30; HL Paper 1: 75 marks in 2 h). Reactivity 1 contains a large amount of additional higher level material (R1.2.3–1.2.5 and all of R1.4), so the HL paper is longer.

Inside this booklet

1. Section A — Paper 1A style: multiple-choice questions
2. Section B — Paper 1B style: data-based questions on experimental work
3. Answer key and worked solutions — Section A
4. Markscheme and worked solutions — Section B

Instructions

- Work under examination conditions: set a timer, and do not look at the solutions until you have finished.
- You need a **calculator** and a copy of the **IB Chemistry data booklet** (first assessment 2025). Where a question gives data, use those values. c(water) = 4.18 J g^{-1} K^{-1}; R = 8.31 J K^{-1} mol^{-1}.
- **Section A:** each question has one correct answer and is worth 1 mark. No marks are deducted for wrong answers, so answer every question. Record your answers in the grid at the end of Section A.
- **Section B:** write your answers in the spaces provided. Show your working in calculations. The marks for each part are shown in brackets.
- Questions tagged HL test Additional Higher Level material and are for HL students only.

Section A — Paper 1A style (multiple choice)

1. Which statement is correct?

A Temperature is a measure of the average kinetic energy of the particles in a substance.
B Heat is a measure of the total kinetic energy of the particles in a substance.
C Heat always flows from a substance with more particles to one with fewer particles.
D A temperature rise of 10 °C is equal to a temperature rise of 283 K.

2. Which row is correct for an exothermic reaction carried out in aqueous solution?

	Temperature of the solution	Sign of ΔH	Relative stability
A	decreases	negative	products more stable
B	increases	positive	reactants more stable
C	increases	negative	products more stable
D	decreases	positive	reactants more stable

3. How much heat is needed to raise the temperature of 50.0 g of water from 20.0 °C to 45.0 °C?

A 0.0523 kJ **B** 5.23 kJ **C** 9.41 kJ **D** 5225 kJ

4. The same amount of heat is supplied to 10.0 g blocks of four metals, all initially at 20 °C. Which block reaches the highest temperature?

A aluminium, c = 0.897 **B** iron, c = 0.449 **C** copper, c = 0.385 **D** gold, c = 0.129

specific heat capacities in J g^{-1} K^{-1}

5. 25.0 cm^3 of 2.00 mol dm^{-3} HCl(aq) was mixed with 25.0 cm^3 of 2.00 mol dm^{-3} NaOH(aq). The temperature rose by 13.0 K. Which expression gives the enthalpy of neutralization in kJ mol^{-1}?

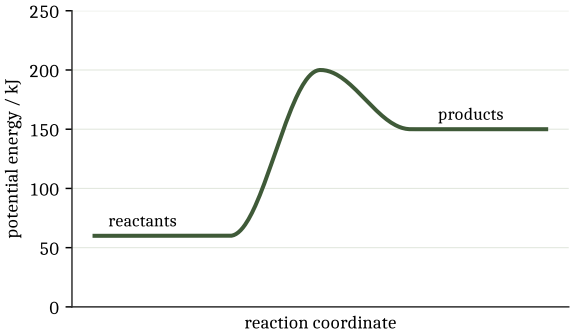
A −(50.0 × 4.18 × 13.0) / (1000 × 0.0500)
B −(25.0 × 4.18 × 13.0) / (1000 × 0.0500)
C −(50.0 × 4.18 × 13.0) / (1000 × 0.100)
D +(50.0 × 4.18 × 13.0) / (1000 × 0.0500)

6. A student used a spirit burner and a copper calorimeter to measure the enthalpy of combustion of ethanol. Which would make the experimental value **less exothermic** than the data booklet value?

I. heat lost to the surroundings II. incomplete combustion of the ethanol III. a thermometer that reads 0.5 °C too high throughout

A I and II only **B** I and III only **C** II and III only **D** I, II and III

7. What is the activation energy of the **reverse** reaction shown in the potential energy profile?



A 140 kJ **B** 90 kJ **C** 50 kJ **D** 200 kJ

8. Which equation represents the standard enthalpy change of combustion of propan-1-ol?

A $2C_3H_7OH(l) + 9O_2(g) \rightarrow 6CO_2(g) + 8H_2O(l)$
B $C_3H_7OH(l) + 4½O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$
C $C_3H_7OH(l) + 4½O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$
D $C_3H_7OH(l) + 3O_2(g) \rightarrow 3CO(g) + 4H_2O(l)$

9. For which process is ΔH equal to four times the average C–H bond enthalpy?

A $CH_4(g) \rightarrow C(g) + 4H(g)$ **B** $CH_4(g) \rightarrow C(s) + 2H_2(g)$
C $C(g) + 4H(g) \rightarrow CH_4(g)$ **D** $CH_4(g) \rightarrow C(g) + 2H_2(g)$

10. Use the bond enthalpies C=C 614, H–H 436, C–C 346 and C–H 414 kJ mol^{-1}. What is ΔH, in kJ mol^{-1}, for $C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g)$?

A +124 **B** +704 **C** −124 **D** −1174

11. Why do enthalpy changes calculated from average bond enthalpies often differ from experimental values?

I. The bond enthalpies are averages taken over many different compounds.
II. Bond enthalpies apply only to substances in the gaseous state.
III. Bond enthalpies are measured only at high temperatures.

A I only **B** I and III only **C** II and III only **D** I and II only

12. The enthalpy change for A → B is ΔH_1, and for C → B is ΔH_2. What is the enthalpy change for A → C?

A $\Delta H_1 + \Delta H_2$ **B** $\Delta H_1 - \Delta H_2$ **C** $\Delta H_2 - \Delta H_1$ **D** $-\Delta H_1 - \Delta H_2$

13. $C(s) + O_2(g) \rightarrow CO_2(g)$, $\Delta H = -394$ kJ; $2CO(g) + O_2(g) \rightarrow 2CO_2(g)$, $\Delta H = -566$ kJ. What is ΔH for $C(s) + ½O_2(g) \rightarrow CO(g)$?

A +172 kJ **B** −677 kJ **C** −960 kJ **D** −111 kJ

14. Why is the F–F bond enthalpy (159 kJ mol^{-1}) lower than the Cl–Cl bond enthalpy (242 kJ mol^{-1})?

A Fluorine has a higher electronegativity than chlorine.

B The lone pairs on the two small fluorine atoms are close together and repel strongly.

C The F–F bond is longer than the Cl–Cl bond.

D Fluorine molecules have weaker London (dispersion) forces.

15. Which equation represents the incomplete combustion of butane to carbon monoxide?

A $2C_4H_{10} + 13O_2 \rightarrow 8CO + 10H_2O$

B $C_4H_{10} + 4O_2 \rightarrow 4CO + 5H_2O$

C $2C_4H_{10} + 9O_2 \rightarrow 8CO + 10H_2O$

D $2C_4H_{10} + 5O_2 \rightarrow 8CO + 10H_2$

16. Which fuel releases the most energy per mole of carbon dioxide produced?

Fuel	ΔH_c / kJ mol^{-1}
A methane, CH_4	−891
B propane, C_3H_8	−2219
C octane, C_8H_{18}	−5470
D ethanol, C_2H_5OH	−1367

17. The enthalpy of combustion of propane (M = 44.11 g mol^{-1}) is −2219 kJ mol^{-1}. What is its specific energy?

A 0.0199 kJ g^{-1} **B** 2219 kJ g^{-1} **C** 9.79 × 10^4 kJ g^{-1} **D** 50.3 kJ g^{-1}

18. Which is a disadvantage of biofuels compared with fossil fuels?

A They are non-renewable.

B Growing the crops uses land that could be used to grow food.

C They release carbon that was fixed millions of years ago.

D They cannot be made from waste materials.

19. Which gases absorb infrared radiation emitted by the Earth's surface?

I. CO_2 II. CH_4 III. N_2

A I and II only **B** I and III only **C** II and III only **D** I, II and III

20. What is the half-equation at the cathode of a hydrogen–oxygen fuel cell with an alkaline electrolyte?

A $H_2 + 2OH^- \rightarrow 2H_2O + 2e^-$

B $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$

C $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$

D $4OH^- \rightarrow O_2 + 2H_2O + 4e^-$

21. Which row describes the anode of a methanol fuel cell with an acidic electrolyte?

	Process	Half-equation
A	reduction	$CH_3OH + H_2O + 6e^- \rightarrow CO_2 + 6H^+$
B	reduction	$1½O_2 + 6H^+ + 6e^- \rightarrow 3H_2O$
C	oxidation	$CH_3OH \rightarrow CO_2 + 4H^+ + 4e^-$
D	oxidation	$CH_3OH + H_2O \rightarrow CO_2 + 6H^+ + 6e^-$

22. For $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$, $\Delta H = -184$ kJ mol^{-1}. E(H–H) = 436 and E(Cl–Cl) = 242 kJ mol^{-1}. What is the H–Cl bond enthalpy, in kJ mol^{-1}?

A 247 **B** 431 **C** 339 **D** 862

23. HL Which substance has a standard enthalpy change of formation of zero at 298 K?

A Hg(l) **B** $O_3(g)$ **C** C(diamond) **D** $Br_2(g)$

24. HL ΔH_f / kJ mol^{-1}: $Fe_2O_3(s)$ −824, CO(g) −111, $CO_2(g)$ −394. What is ΔH for $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$?

A +25 kJ **B** −358 kJ **C** −25 kJ **D** −1515 kJ

25. HL ΔH_c / kJ mol^{-1}: $C_2H_4(g)$ −1411, $C_2H_5OH(l)$ −1367. What is ΔH for $C_2H_4(g) + H_2O(l) \rightarrow C_2H_5OH(l)$?

A +44 kJ **B** −2778 kJ **C** −1411 kJ **D** −44 kJ

26. HL Data / kJ mol^{-1}: ΔH_f(LiF) −616; ΔH_{atom}(Li) +159; IE_1(Li) +520; E(F–F) +158; EA(F) −328. What is the lattice enthalpy of LiF, in kJ mol^{-1}?

A +1046 **B** −186 **C** +1125 **D** +1702

27. HL Which compound has the largest lattice enthalpy?

A NaF **B** MgO **C** CaS **D** $MgCl_2$

28. HL Which reaction has the largest increase in entropy?

A $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
B $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
C $2NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g)$
D $H_2(g) + I_2(g) \rightarrow 2HI(g)$

29. HL A reaction has $\Delta H = -50.0$ kJ mol^{-1} and $\Delta S = -100$ J K^{-1} mol^{-1}. What is ΔG at 600 K?

A −110 kJ mol^{-1} **B** −49.9 kJ mol^{-1} **C** +59 950 kJ mol^{-1} **D** +10.0 kJ mol^{-1}

30. HL A reaction has $\Delta H = +180$ kJ mol^{-1} and $\Delta S = +160$ J K^{-1} mol^{-1}. Under which conditions is it spontaneous?

A above 1125 K **B** below 1125 K
C above 1.13 K **D** at all temperatures

31. HL A reaction has $\Delta G^{\ominus} = -11.4$ kJ mol^{-1} at 298 K. What is the value of K?

A 0.0100 **B** 100 **C** 4.60 **D** 1.00

32. HL Which row is correct for a reaction mixture at equilibrium?

	ΔG	$\Delta G^{\ominus}$	Q
A	0	0	1
B	$-RT \ln K$	0	K
C	0	$-RT \ln K$	K
D	$-RT \ln Q$	0	1

Answer grid — Section A

Q		Q		Q		Q	
1		9		17		25	
2		10		18		26	
3		11		19		27	
4		12		20		28	
5		13		21		29	
6		14		22		30	
7		15		23		31	
8		16		24		32	

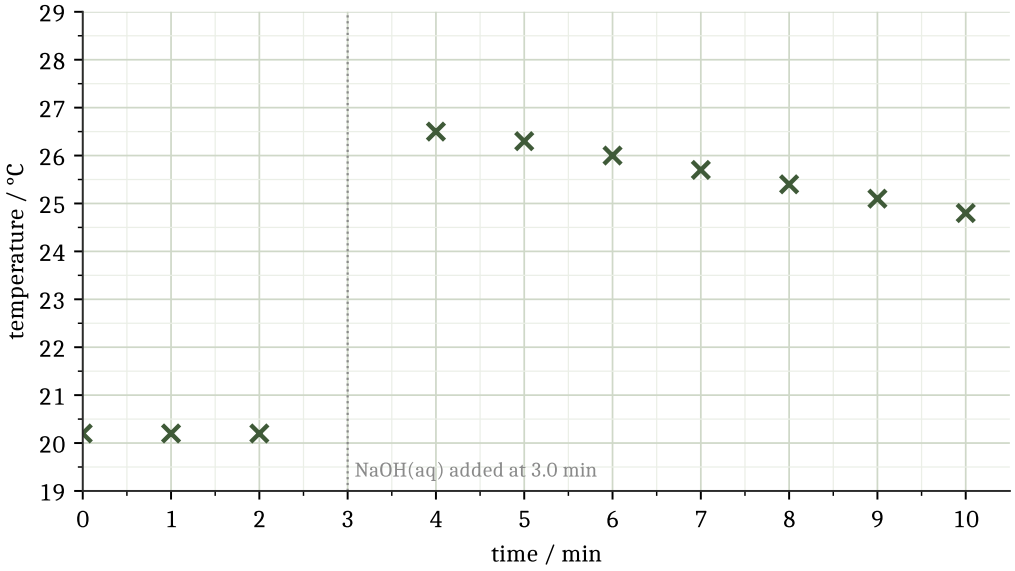
Section A total: SL 22 marks (Questions 1–22); HL 32 marks.

Section B — Paper 1B style (data-based questions on experimental work)

Question 1 Enthalpy of neutralization with a cooling correction (8 marks)

A student placed 50.0 cm^3 of 1.00 mol dm^{-3} hydrochloric acid in an expanded polystyrene cup and recorded its temperature every minute. At exactly 3.0 minutes, 50.0 cm^3 of 1.10 mol dm^{-3} sodium hydroxide solution, at the same initial temperature, was added and the mixture was stirred. The temperature was then recorded every minute until 10.0 minutes. The results are plotted below.

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$



(a) Explain why the temperature was recorded for three minutes before the sodium hydroxide was added. **[1]**

(b) By drawing suitable lines on the graph, determine the temperature change that would have occurred if no heat had been lost to the surroundings. **[2]**

(c) Show that the sodium hydroxide is in excess. **[1]**

(d) Calculate the enthalpy change of neutralization, in kJ mol^{-1}. State the assumption you make about the solution. **[2]**

(e) The literature value is −57.2 kJ mol^{-1}. Suggest two reasons why the student's value differs from it, other than heat loss. **[2]**

Question 2 Enthalpies of combustion of a series of alcohols (8 marks)

A student burned four alcohols, one at a time, in a spirit burner under a copper calorimeter containing 150.0 g of water. The student calculated the enthalpy of combustion of three of them.

Alcohol	M / g mol^{-1}	mass burned / g	ΔT of water / K	experimental ΔH_c / kJ mol^{-1}	data booklet ΔH_c / kJ mol^{-1}
methanol, CH_3OH	32.05	0.350	8.1	−465	−726
ethanol, C_2H_5OH	46.08	0.400	12.5	−903	−1367
propan-1-ol, C_3H_7OH	60.11	0.420	14.1	?	−2021
butan-1-ol, C_4H_9OH	74.14	0.450	16.2	−1673	−2676

(a) Calculate the experimental enthalpy of combustion of propan-1-ol. **[3]**

(b) Calculate the percentage error in your answer to (a). **[1]**

(c) The data booklet values increase by a nearly constant amount from one alcohol to the next. Explain this in terms of the bonds broken and formed. **[2]**

(d) All four experimental values are about 35–40 % less exothermic than the data booklet values. Suggest two specific modifications to the apparatus or method, each linked to the error it reduces. **[2]**

Question 3 Using Hess's law to find an enthalpy of decomposition (6 marks)

Potassium hydrogencarbonate decomposes on heating: $2KHCO_3(s) \rightarrow K_2CO_3(s) + H_2O(l) + CO_2(g)$. To find ΔH for this reaction, a student carried out two experiments, each with 30.0 cm^3 of 2.0 mol dm^{-3} HCl(aq) (an excess) in a polystyrene cup.

Experiment	Reaction	solid added	ΔT / K
1	$KHCO_3(s) + HCl(aq) \rightarrow KCl(aq) + H_2O(l) + CO_2(g)$	3.00 g $KHCO_3$ (M = 100.12)	−5.5
2	$K_2CO_3(s) + 2HCl(aq) \rightarrow 2KCl(aq) + H_2O(l) + CO_2(g)$	2.76 g K_2CO_3 (M = 138.21)	+5.1

(a) Calculate ΔH_1 for experiment 1, in kJ mol^{-1}. Assume the mass of the solution is 30.0 g. **[2]**

(b) ΔH_2 for experiment 2 is −32.0 kJ mol^{-1}. Use your answer to (a) and Hess's law to calculate ΔH for the decomposition of potassium hydrogencarbonate. **[2]**

(c) Explain why ΔH for the decomposition cannot be measured directly in a calorimeter. **[1]**

(d) State why the hydrochloric acid must be in excess in both experiments. **[1]**

HL Question 4 Gibbs energy from equilibrium data (6 marks)

The equilibrium constant, K, for $N_2O_4(g) \rightleftharpoons 2NO_2(g)$ was determined at four temperatures by measuring the intensity of the brown colour of NO_2.

Temperature / K	298	323	348	373
K	0.147	0.877	4.05	15.3

(a) Calculate $\Delta G^{\ominus}$, in kJ mol^{-1}, at 298 K and at 373 K. **[2]**

(b) Assuming $\Delta H^{\ominus}$ and $\Delta S^{\ominus}$ are constant over this range, use your two values to determine $\Delta S^{\ominus}$ and $\Delta H^{\ominus}$. **[2]**

(c) Determine the temperature at which K = 1. **[1]**

(d) Explain the sign of $\Delta S^{\ominus}$. **[1]**

Section B total: SL 22 marks (Questions 1–3); HL 28 marks.

End of questions

Answer key — Section A

Q	Ans	Q	Ans	Q	Ans	Q	Ans
1	A	**9**	A	**17**	D	**25**	D
2	C	**10**	C	**18**	B	**26**	A
3	B	**11**	D	**19**	A	**27**	B
4	D	**12**	B	**20**	C	**28**	C
5	A	**13**	D	**21**	D	**29**	D
6	A	**14**	B	**22**	B	**30**	A
7	C	**15**	C	**23**	A	**31**	B
8	B	**16**	A	**24**	C	**32**	C

Worked solutions — Section A

1. A. Heat is energy transferred because of a temperature difference; it flows from higher to lower temperature, not from more to fewer particles. A change of 10 °C is a change of 10 K.

2. C. Heat passes from the system to the water, so the temperature rises; the products have lower enthalpy, so ΔH is negative and the products are more stable.

3. B. $Q = 50.0 \times 4.18 \times 25.0 = 5225$ J = 5.23 kJ. C uses 45.0 instead of ΔT; D forgets to convert J to kJ.

4. D. $\Delta T = Q/mc$, and m and Q are the same, so the smallest c gives the largest rise.

5. A. The whole 50.0 g of solution warms. $n(H_2O) = 2.00 \times 0.0250 = 0.0500$ mol. B uses only one solution's volume; C uses the total moles of both reactants; D has the wrong sign.

6. A. I reduces the heat reaching the water; II releases less energy per mole. III is a constant offset that cancels when ΔT is calculated.

7. C. The reverse reaction climbs from the products (150 kJ) to the peak (200 kJ): 50 kJ. A is E_a(forward); B is ΔH.

8. B. One mole of fuel, complete combustion, water as a liquid (standard state). A is for 2 mol; C forms steam; D is incomplete.

9. A. Bond enthalpy refers to breaking bonds in gaseous molecules to give gaseous atoms. C is bond formation (−4E).

10. C. Broken: C=C + H–H = 1050; formed: C–C + 2C–H = 346 + 828 = 1174; $\Delta H = 1050 - 1174 = -124$. The other four C–H bonds are unchanged.

11. D. Statement III is false: bond enthalpies are quoted for 298 K.

12. B. A → B → C: ΔH_1 + (reverse of C → B) = $\Delta H_1 - \Delta H_2$.

13. D. Use the first equation and reverse half the second: −394 + 283 = −111 kJ.

14. B. The F–F bond is short (142 pm, shorter than Cl–Cl), so the lone pairs on the two atoms repel strongly and weaken the bond.

15. C. 8 C → 8CO; 20 H → $10H_2O$; O atoms: 8 + 10 = 18 = $9O_2$. A is complete combustion's O_2 with CO products; D makes H_2.

16. A. Energy per mol CO_2: methane 891, propane 740, octane 684, ethanol 684 kJ. Methane has the highest H : C ratio.

17. D. 2219 / 44.11 = 50.3 kJ g^{-1}.

18. B. Biofuels are renewable (A false); their carbon was fixed recently (C describes fossil fuels); biogas is made from waste (D false).

19. A. CO_2 and CH_4 have vibrations that change the dipole moment. N_2 has none.

20. C. Oxygen is reduced at the cathode; in alkali, OH^- ions form. A is the alkaline anode; B is the acidic cathode.

21. D. The fuel is oxidized at the anode; carbon goes from −2 to +4, releasing 6 electrons, and H_2O supplies the extra oxygen atom.

22. B. −184 = (436 + 242) − 2E(H–Cl), so 2E = 862 and E = 431. D forgets that two H–Cl bonds form.

23. A. Mercury is a liquid at 298 K, so Hg(l) is its standard state. O_3, diamond and Br_2(g) are not standard states.

24. C. [3(−394)] − [−824 + 3(−111)] = −1182 + 1157 = −25 kJ. Elements count as zero.

25. D. Reactants − products: (−1411 + 0) − (−1367) = −44 kJ. Water does not burn.

26. A. $\Delta H_{lat} = -\Delta H_f + \Delta H_{atom}(Li) + IE + \frac{1}{2}E(F–F) + EA = 616 + 159 + 520 + 79 - 328 = +1046$. C uses the full F–F bond enthalpy; B has the wrong sign for ΔH_f; D has the wrong sign for EA.

27. B. 2+ and 2− ions with small radii. CaS also has 2+/2− ions but both are larger; $MgCl_2$ contains 1− ions.

28. C. 0 mol gas → 2 mol gas. B gives 1 mol gas; A decreases; D is about zero.

29. D. $\Delta G = -50.0 - 600(-0.100) = -50.0 + 60.0 = +10.0$ kJ mol^{-1}. C uses ΔS in J without converting.

30. A. ΔH and ΔS both positive: spontaneous when $T\Delta S > \Delta H$, i.e. above T = 180 / 0.160 = 1125 K.

31. B. ln K = 11 400 / (8.31 × 298) = 4.60; $K = e^{4.60} = 100$. C is ln K, not K.

32. C. At equilibrium $\Delta G = 0$ and Q = K. $\Delta G^{\ominus}$ is a constant equal to $-RT \ln K$; it is zero only if K = 1.

Markscheme — Section B

Question 1 (8 marks)

(a) To establish a steady (constant) initial temperature of the acid, which is needed to calculate the temperature change and to draw the line before mixing **[1]**.

(b) A line of best fit through the cooling points extrapolated back to 3.0 min, giving about 26.9 °C, and a horizontal line through the initial points at 20.2 °C **[1]**; ΔT = 26.9 − 20.2 = **6.7 K** (accept 6.5–6.9 K) **[1]**.

(c) n(HCl) = 1.00 × 0.0500 = 0.0500 mol; n(NaOH) = 1.10 × 0.0500 = 0.0550 mol; the reaction is 1 : 1, so NaOH is in excess and HCl is limiting **[1]**.

(d) Q = 100.0 × 4.18 × 6.7 = 2801 J; n(H_2O) = 0.0500 mol; ΔH = −2.80 / 0.0500 = **−56.0 kJ mol^{-1}** (follow through from (b); negative sign required) **[1]**. Assumption: the solution has the density (1.00 g cm^{-3}, so 100 g) and specific heat capacity (4.18 J g^{-1} K^{-1}) of water **[1]**.

(e) Any two **[2]**: heat absorbed by the cup and thermometer is ignored; the specific heat capacity or density of the solution is not exactly that of water; the solutions were not initially at exactly the same temperature; uncertainty in reading the thermometer and in extrapolating the line; the concentrations of the solutions were not exactly as stated. *Do not accept "heat loss" or "human error".*

Question 2 (8 marks)

(a) Q = 150.0 × 4.18 × 14.1 = 8841 J (8.84 kJ) **[1]**; n = 0.420 / 60.11 = 6.99 × 10^{-3} mol **[1]**; ΔH_c = −8.84 / 6.99 × 10^{-3} = **−1.27 × 10^3 kJ mol^{-1}** (−1265; negative sign required) **[1]**.

(b) |−1265 − (−2021)| / 2021 × 100 = **37.4 %** **[1]**.

(c) Each successive alcohol has one more CH_2 group **[1]**; so on combustion one extra C–C and two extra C–H bonds (plus 1½ O=O) are broken, and one more CO_2 and H_2O are formed (two C=O and two O–H bonds), adding an approximately constant extra amount of energy released **[1]**.

(d) Any two, each linked to its error **[2]**: use a draught shield / lid / insulate the calorimeter to reduce heat loss to the surroundings; place the flame closer to the calorimeter base to reduce heat loss; include the heat absorbed by the copper calorimeter (m c ΔT for copper); ensure a good supply of air (oxygen) to reduce incomplete combustion; keep the cap on the burner and reweigh immediately to reduce mass lost by evaporation.

Question 3 (6 marks)

(a) Q = 30.0 × 4.18 × 5.5 = 690 J; n($KHCO_3$) = 3.00 / 100.12 = 0.0300 mol **[1]**; ΔH_1 = +0.690 / 0.0300 = **+23.0 kJ mol^{-1}** (positive, because the temperature fell) **[1]**.

(b) Both routes give 2KCl(aq) + 2H_2O(l) + 2CO_2(g): $2\Delta H_1 = \Delta H + \Delta H_2$ **[1]**; ΔH = 2(+23.0) − (−32.0) = **+78.0 kJ** (per mole of reaction as written; follow through from (a)) **[1]**.

(c) The solid must be heated strongly and continuously, so the heat absorbed by the reaction cannot be separated from (measured against) the heat supplied **[1]**.

(d) So that all of the solid reacts (the solid is the limiting reactant), and the amount in mol used in the calculation is that of the solid **[1]**.

HL **Question 4 (6 marks)**

(a) 298 K: $\Delta G^{\ominus} = -8.31 \times 298 \times \ln 0.147 = +4750$ J = **+4.75 kJ mol^{-1} [1]**; 373 K: $-8.31 \times 373 \times \ln 15.3$ = **−8.46 kJ mol^{-1} [1]**.

(b) gradient = $-\Delta S = (-8.46 - 4.75) / (373 - 298) = -0.176$ kJ K^{-1} mol^{-1}, so $\Delta S^{\ominus}$ = **+176 J K^{-1} mol^{-1} [1]**; $\Delta H^{\ominus} = \Delta G + T\Delta S = 4.75 + 298 \times 0.176$ = **+57.2 kJ mol^{-1} [1]**.

(c) K = 1 when $\Delta G^{\ominus} = 0$: $T = \Delta H/\Delta S = 57.2 / 0.176$ = **325 K** (accept 323–327 K; consistent with K = 0.877 at 323 K) **[1]**.

(d) Positive: one mole of gas forms two moles of gas, so there are more ways of distributing the matter and energy **[1]**.

These practice papers are a free resource, shared for any IB student who finds them useful.

If a particular idea here still isn't clicking, specialist one-to-one support in Physics, Chemistry and Maths from a past A Level examiner is available at iBLaurel — call or message **+44 7404 710097**, or use the contact form at **www.iBLaurel.com**.

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