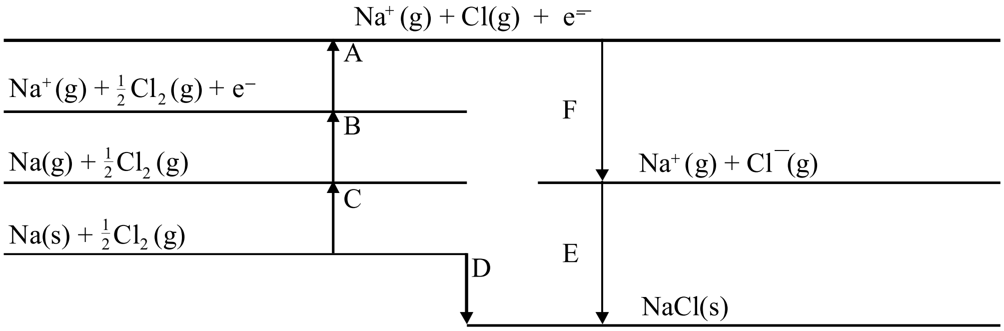
1. The Born-Haber cycle for the formation of sodium chloride from sodium and chlorine may be represented by a series of stages labelled A to F as shown.



(a) (i) Write the letters A to F next to the corresponding definition in the table below

definition	letter	ΔH/kJ mol^{-1}
1st ionisation energy of sodium		+494
1st electron affinity of chlorine		–364
the enthalpy of atomisation of sodium		+109
the enthalpy of atomisation of chlorine		+121
the lattice enthalpy of sodium chloride		–770
the enthalpy of formation of sodium chloride		

(3)

(ii) Calculate the enthalpy of formation of sodium chloride from the data given.

(2)

(b) The lattice enthalpies can be calculated from theory as well as determined experimentally.

	Experimental ΔH/kJ mol^{-1}	Theoretical ΔH/kJ mol^{-1}
Sodium chloride	–770	–766
Silver iodide	–889	–778

Why is the experimental value of the lattice enthalpy of silver iodide (–889kJmol^{-1}) so different from the value calculated theoretically?

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(2)

(c) Explain the trend in first ionisation energies of the elements of Group 1 in the Periodic Table.

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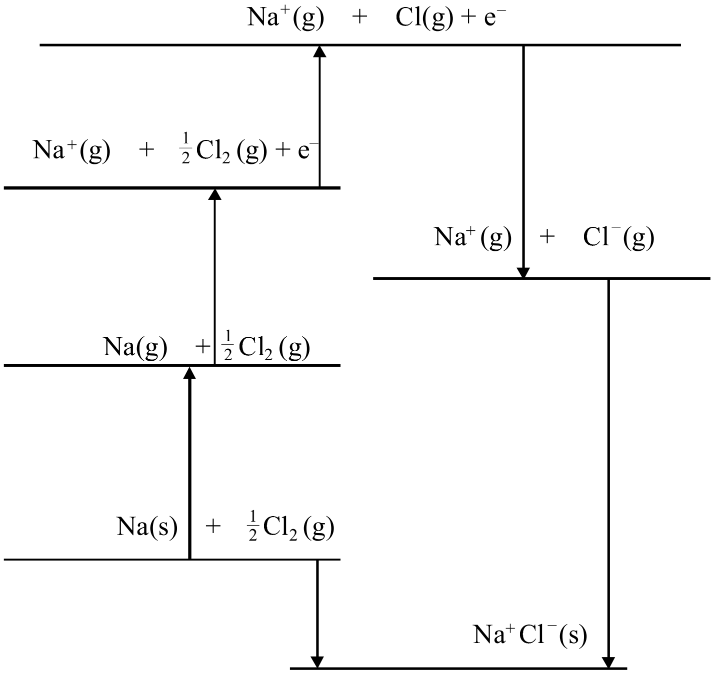
(3)
(Total 10 marks)

2. (a) The Born-Haber cycle for the formation of sodium chloride is shown below.



Use the data below to calculate the lattice enthalpy of sodium chloride.

Enthalpy change	Value of the enthalpy change /kJ mol^{-1}
Enthalpy of atomisation of sodium	+109
1st ionisation energy of sodium	+494
Enthalpy of formation of sodium chloride	–411
Enthalpy of atomisation of chlorine	+121
Electron affinity of chlorine	–364

(2)

(b) Sodium chloride and magnesium oxide have very similar crystal lattices. Suggest why the lattice enthalpy of magnesium oxide is very much larger than that of sodium chloride.

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(2)

(c) The lattice enthalpy of silver iodide can be calculated but the experimental value does not match the calculated value as well as those for sodium chloride match each other.

Explain why the calculated and experimental values for silver iodide are different.

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(2)
(Total 6 marks)

3. (a) State Hess's Law.

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(2)

(b) Define the term **standard enthalpy change of combustion**.

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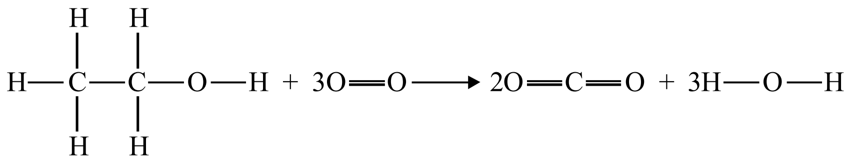
(3)

(c) The equation for the combustion of ethanol in air is

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

and the structural representation of this is:



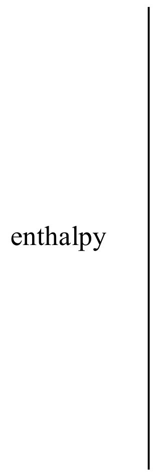
$$\mathrm{H{-}\overset{H}{\underset{H}{C}}{-}\overset{H}{\underset{H}{C}}{-}O{-}H + 3O{=}O \longrightarrow 2O{=}C{=}O + 3H{-}O{-}H}$$

(i) Calculate the enthalpy change for this reaction using the average bond enthalpy values given below.

Bond	Average bond enthalpy/kJ mol^{-1}	Bond	Average bond enthalpy/kJ mol^{-1}
C—H	+412	C—C	+348
C—O	+360	O—H	+463
O—O	+496	C═O	+743

(3)

(ii) Draw and label an enthalpy level diagram to represent this reaction.



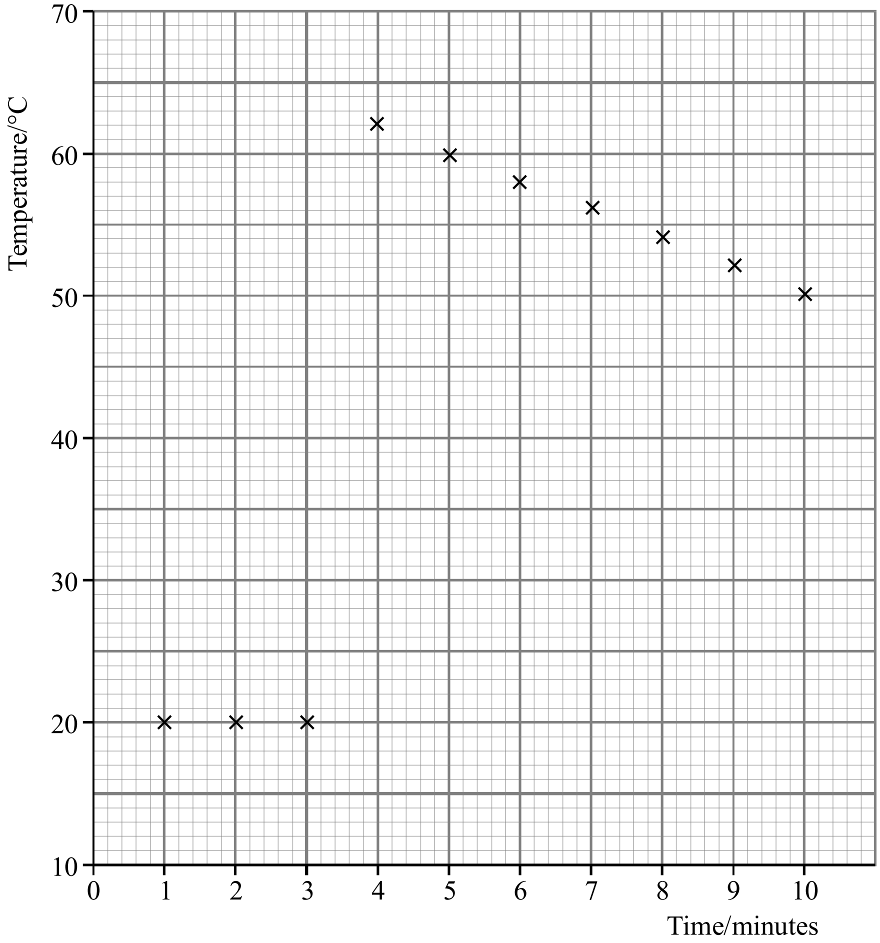
(2)
(Total 10 marks)

4. In an experiment to find the enthalpy change when copper is displaced from a solution of copper ions excess zinc was added to 50.0 cm^3 of 1.00 mol dm^{-3} aqueous copper(ii) sulphate in a plastic cup.

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

The temperature of the solution in the cup was measured every minute for 10 minutes with the zinc being added after 3.5 minutes.

The temperature readings are shown on the graph below.



(a) Suggest two reasons why a series of temperature readings is taken rather than simply initial and final readings.

First reason...

....

Second reason...

....

(2)

(b) Use the graph to calculate the maximum temperature change, ΔT, as the reaction takes place.

ΔT = ...°C

(2)

(c) Calculate the enthalpy change for the reaction using the formula below, giving your answer to an appropriate number of significant figures.

$$\Delta \mathbf{H} = -4.18 \times \Delta \mathbf{T} \ \mathbf{kJ\ mol^{-1}}$$

(2)
(Total 6 marks)

5. (a) (i) Define the term **standard enthalpy of formation**, $\Delta H_f^{\ominus}$.

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(3)

(ii) The following table shows some values of standard enthalpy of formation.

Name	Formula	$\Delta H_f^{\ominus}$/kJ mol^{-1}
ethene	$C_2H_4(g)$	+52.3
hydrogen bromide	$HBr(g)$	–36.2
bromoethane	$C_2H_5Br(g)$	–60.4

Use the data in the table above to calculate the standard enthalpy change for the following reaction.

$$C_2H_4(g) + HBr(g) \rightarrow C_2H_5Br(g)$$

(2)

(iii) State the significance of the sign of the value obtained in part (a)(ii) above.

....

....

(1)

(b) Enthalpy changes can also be calculated using average bond enthalpy data.

Bond	Average bond enthalpy/kJ mol^{-1}
C == C	+612
C — C	+348
C — H	+412
C — Br	+276
H— Br	+366

Use the data in the table above to recalculate the enthalpy change for the reaction in part (a)(ii).

$$C_2H_4(g) + HBr(g) \rightarrow C_2H_5Br(g)$$

(3)

(c) Suggest why the value obtained in part (b) above is likely to be less accurate than that obtained in part (a)(ii).

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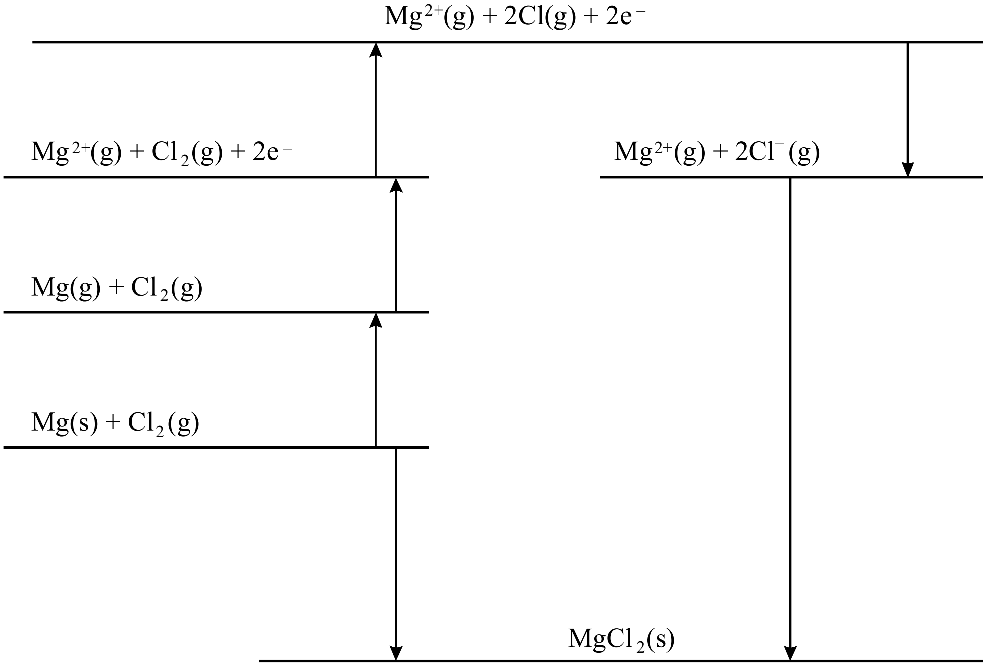
(2)
(Total 11 marks)

6. The formation of magnesium chloride from magnesium and chlorine may be represented by the following Born-Haber cycle:



(a) Define the terms:

Lattice enthalpy.

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(3)

Enthalpy of atomisation.

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(2)

(b) (i) Identify **on the diagram** the chance representing the enthalpy of atomisation of magnesium.

(1)

(ii) Use the data below to calculate the first electron affinity of chlorine.

Enthalpy change	Value of the enthalpy

	change / kJ mol^{-1}
Enthalpy of atomisation of magnesium	+150
1st Ionisation energy of magnesium	+736
2nd Ionisation energy of magnesium	+1450
Enthalpy of formation of magnesium chloride	–642
Enthalpy of atomisation of chlorine	+121
Lattice enthalpy of magnesium chloride	–2493

(2)

(c) Hydrogen gas reacts with sodium metal to form an ionic solid, NaH, which contains sodium cations.

Draw a Born-Haber cycle which could be used to determine the electron affinity of hydrogen.

(3)
(Total 11 marks)

7. The reaction of an acid with a base to give a salt is an exothermic reaction. In an experiment to determine the enthalpy of neutralisation of hydrochloric acid with sodium hydroxide, 50.0cm^3 of 1.00 mol dm^{-3} HCl was mixed with 50.0 cm^3 of 1.10 mol dm^{-3} NaOH. The temperature rise obtained was 6.90 °C.

(a) Define the term **enthalpy of neutralisation**.

....

....

(1)

(b) Assuming that the density of the final solution is 1.00 g cm^{-3} and that its heat capacity is 4.18 J K^{-1} g^{-1}, calculate the heat evolved during the reaction.

(3)

(c) 0.0500 mol of acid was neutralised in this reaction; calculate $\Delta H_{\text{neutralisation}}$ in kJ mol^{-1}.

(2)

(d) Suggest why sodium hydroxide is used in slight excess in the experiment.

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(1)
(Total 7 marks)

8. (a) (i) Define the term **standard enthalpy of combustion**.

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(3)

(ii) The values for the standard enthalpy of combustion of graphite and carbon monoxide are given below:

	$\Delta H_c^{\ominus}$/kJ mol^{-1}
C (graphite)	–394
CO(g)	–283

Use these data to find the standard enthalpy change of formation of carbon monoxide using a Hess's law cycle.

$$C(graphite) + \frac{1}{2}O_2(g) \rightarrow CO(g)$$

(3)

(iii) Suggest why it is not possible to find the enthalpy of formation of carbon monoxide directly.

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(1)

(iv) Draw an enthalpy level diagram below for the formation of carbon monoxide from graphite.

(1)

(b) Natural gas consists of methane, CH_4. When methane burns completely in oxygen the reaction occurs as shown in the equation

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ $\Delta H_c = -890 \text{ kJ mol}^{-1}$

Methane does not burn unless lit.

Use this information to explain the difference between thermodynamic and kinetic stability.

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(4)
(Total 12 marks)

9. A student was required to determine the enthalpy change for the reaction between iron and copper sulphate solution.

The student produced the following account of their experiment.

A piece of iron, mass about 3 g, was placed in a glass beaker. 50 cm^3 of 0.5 mol dm^{-3} aqueous copper sulphate solution was measured using a measuring cylinder and added to the beaker. The temperature of the mixture was measured immediately before the addition and every minute afterwards until no further change took place.

$$Fe + CuSO_4 \rightarrow FeSO_4 + Cu$$

Timing	before addition	1 min	2 mins	3 mins	4 mins	5 mins
Temperature/°C	22	27	29	26	24	22

(a) Suggest **two** improvements you would make to this experiment. Give a reason for each of the improvements suggested.

Improvement 1 ....

....

Reason 1 ....

....

Improvement 2 ....

....

Reason 2 ....

....

(4)

(b) In an improved version of **the same experiment** a maximum temperature rise of 15.2 °C occurred when reacting excess iron with 50.0 cm^3 of 0.500 mol dm^{-3} aqueous copper sulphate solution.

(i) Using this data and taking the specific heat capacity of all aqueous solutions as 4.18 Jg^{-1} deg^{-1} calculate the heat change.

(1)

(ii) Calculate the number of moles of copper sulphate used.

(1)

(iii) Calculate the enthalpy change of this reaction in kJ mol^{-1}.

(2)
(Total 8 marks)

10. (a) Define the term **standard enthalpy of combustion**, making clear the meaning of **standard** in this context.

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(3)

(b) Use the enthalpies of combustion given below to find the enthalpy change for the reaction:

$$2C(graphite) + 2H_2(g) + O_2(g) \rightarrow CH_3COOH(l)$$

	$\Delta H_{combustion}$/kJ mol^{-1}
C(graphite)	–394
H_2(g)	–286
CH_3COOH(l)	–874

(3)

(c) With reference to ethanoic acid, CH_3COOH, what is the enthalpy change obtained in (b) called?

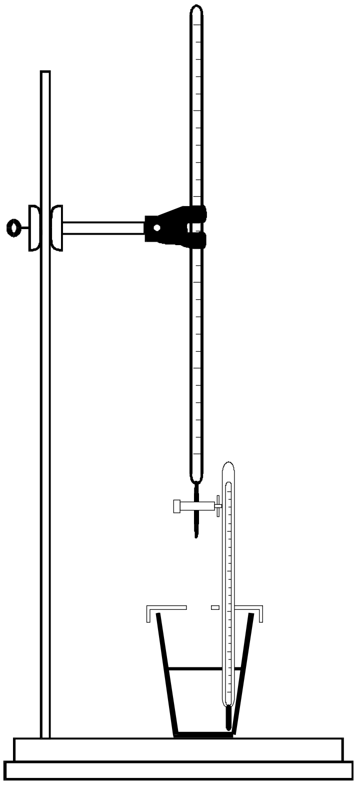
………………………………………………………………………………………..

(1)

(d) Draw an enthalpy level diagram to represent the enthalpy change for the combustion of graphite. Show both the enthalpy levels of the reactants and products and an energy profile which represents the activation energy for the reaction.

(3)
(Total 10 marks)

11. (a) This question is about finding the formula of copper hydroxide. The method is as follows:



20.0 cm^3 of an aqueous solution of a copper salt of concentration 1.00 mol dm^{-3} was placed in a polystyrene cup and its temperature measured using a thermometer graduated in 0.1 °C intervals.

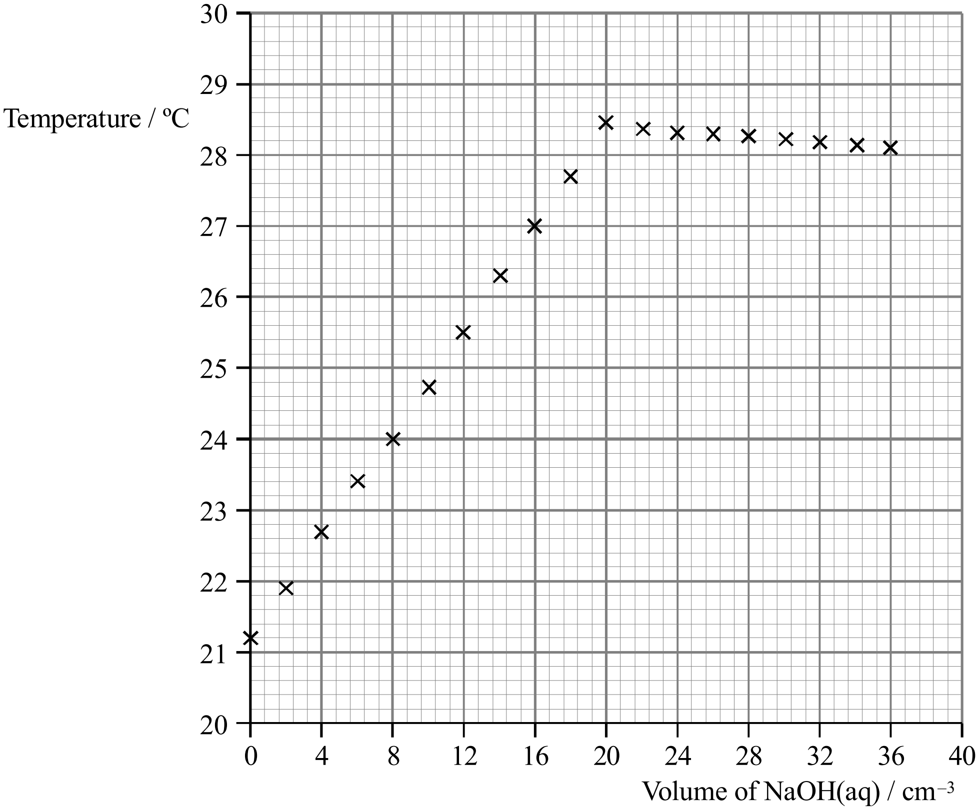
A burette was filled with aqueous sodium hydroxide, of concentration 2.00 mol dm^{-3}.

2.00 cm^3 of sodium hydroxide solution was run into the solution of the copper salt and the temperature was measured immediately.

As soon as possible a further 2.00 cm^3 of sodium hydroxide solution was run in and the temperature measured again.

This process of adding 2.00 cm^3 portions of sodium hydroxide solution and measuring the temperature was continued until a total of 36.0 cm^3 of the sodium hydroxide solution had been added.

The temperature readings are shown in the graph below.



(i) Explain why the temperature reaches a maximum and then falls slightly on addition of further sodium hydroxide solution.

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(2)

(ii) From the graph, what volume of the aqueous sodium hydroxide was required for complete reaction?

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(1)

(iii) Calculate the amount (number of moles) of sodium hydroxide in this volume of solution.

(1)

(iv) Calculate the amount (number of moles) of copper ions that have reacted.

(1)

(v) Write the ratio of moles of copper ions to hydroxide ions reacting.

(1)

(vi) Write the formula of the copper hydroxide that is produced.

(1)

(b) The data call be used to find the enthalpy change for the reaction between sodium hydroxide and the copper salt.

(i) Use the graph to find the temperature rise that occurs for complete reaction.

……………………………………………………………………………………..

(1)

(ii) Find the heat change, q, that occurs in the polystyrene cup for complete reaction. Use the formula

$$q = 168 \times \Delta T \text{ joules}$$

(1)

(iii) Use your results from (a)(iv) and (b)(ii) above, to find the molar enthalpy change, ΔH, for the reaction. Give the correct sign and units to the answer.

(3)

(c) Identify one potential source of error in this experiment, and say what you would do to reduce its effect.

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(2)
(Total 14 marks)

12. An excess of zinc powder was added to 20.0 cm^3 of a solution of copper(II) sulphate of concentration 0.500 mol dm^{-3}. The temperature increased by 26.3 °C.

(a) How many moles of copper(II) sulphate were used in this experiment?

(1)

(b) Calculate the enthalpy change, ΔH, in kJ mol^{-1} for this reaction given that:

energy change =	specific heat capacity	× mass of solution	× temperature change
/J	/J g^{-1} K^{-1}	/g	/K

Assume that the mass of solution is 20.0 g and the specific heat capacity of the solution is 4.18 J $g^{-1}K^{-1}$.

(2)
(Total 5 marks)

13. Urea, which is used as a fertillser in much of mainland Europe, Asia and Africa, is manufactured by the reaction of ammonia and carbon dioxide.

$$2NH_3(g) + CO_2(g) \rightarrow NH_2CONH_2(s) + H_2O(l)$$

(a) Define the term **standard enthalpy of formation**, $\Delta H_f^{\ominus}$, of urea.

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(3)

(b) Calculate the enthalpy change, $\Delta H^{\ominus}$ for the reaction above, given the following standard enthalpies of formation.

Substance	$\Delta H_f^{\ominus}$ / kJ mol^{-1}
$NH_3(g)$	–46.2
$CO_2(g)$	–393.5
$NH_2CONH_2(s)$	–632.2
$H_2O(l)$	–285.8

(3)
(Total 6 marks)

14. (a) Halogenoalkanes react with many nucleophiles.

Define the term **nucleophile**.

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....

....

(2)

(b) (i) Identify the reagent and conditions necessary for the conversion of iodoethane to ethylamine, $C_2H_5NH_2$.

Reagent:..

Conditions:..

....

(3)

(ii) State why the rate of reaction would be slower if bromobutane were used in place of iodoethane, with all other conditions remaining the same.

....

....

(1)

(c) Iodoethane reacts with water to form ethanol and hydrogen iodide.

$C_2H_5I + H_2O \rightarrow C_2H_5OH + HI \qquad \Delta H_f^{\ominus} = +36\ \text{kJ mol}^{-1}$

Use some or all of the data below to calculate the C—I bond enthalpy.

Bond	Bond enthalpy / kJ mol^{-1}	Bond	Bond enthalpy / kJ mol^{-1}
C—H	413	H—I	298
C—C	347	C—O	358
H—O	464		

(3)

(d) Ethanol was heated under reflux with an excess of a mixture of potassium dichromate(VI) and dilute sulphuric acid. Draw the full structural formnula of the organic product.

(1)
(Total 10 marks)

15. (a) The equation below shows the reaction which occurs when ammonia is dissolved in water.

$$NH_3(g) + H_2O(1) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

(i) Explain why water is classified as an acid in this reaction.

....

....

(1)

(ii) The ammonia is acting as a weak base in this reaction.

What is the difference between a weak base and a strong base?

....

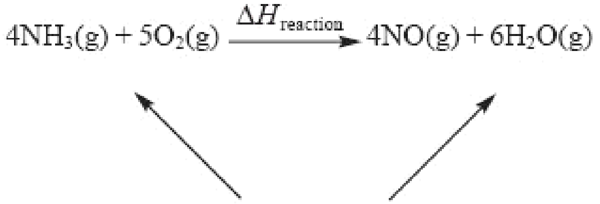
....

....

(1)

(b) Ammonia reacts with oxygen to form the gases nitrogen(II) oxide and steam.

(i) Complete the Hess cycle below so that $\Delta H_{reaction}$ can be calculated using standard enthalpy changes of formation. Include state symbols.



....

(2)

(ii) Calculate $\Delta H_{\text{reaction}}$ for this reaction using the following data.

ΔH_f [$NH_3(g)$] = – 46.1 kJ mol^{-1}

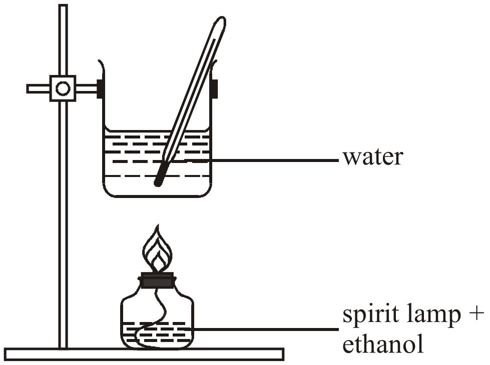
ΔH_f [$NO(g)$] = + 90.2 kJ mol^{-1}

ΔH_f [$H_2O(g)$] = – 241.8 kJ mol^{-1}

Include a sign and units in your answer and give your answer to **three** significant figures.

(3)
(Total 7 marks)

16. The apparatus shown in the diagram below may be used to find the enthalpy of combustion of alcohols.



Using the apparatus, a student recorded the results included in the table below.

Alcohol = ethanol, C_2H_5OH
Molar Mass (C_2H_5OH) = 46.0 g mol^{-1}
Volume of water in beaker = 200 cm^3
$\therefore$ mass of water in beaker = 200 g
Weighings
Spirit lamp + ethanol before combustion = 198.76 g
Spirit lamp + ethanol after combustion = 197.68 g
Temperatures
Water before heating = 19.5 °C
Water after heating = 38.1 °C
Specific heat capacity of water = 4.18 J g^{-1} °C^{-1}

(a) What assumption is the student making about water to be able to state that its mass is numerically equal to its volume?

....

(1)

(b) Calculate the heat gained by the water. Give your answer in kJ.

(2)

(c) Calculate the amount (number of moles) of ethanol used.

(2)

(d) Using your values from (b) and (c), calculate the enthalpy of combustion of ethanol. Give your answer to a number of significant figures consistent with the readings in the table. Include a sign and units in your answer.

(3)

(e) The student's evaluation of the experiment is given below.

> My calculated value of the enthalpy of combustion was numerically much less than the data book value. The reasons for my low value include:
>
> 1 heat losses to the surrounding air;
> 2 when I re-checked the mass of the spirit lamp and ethanol after combustion, I noticed that it had lost mass even when it was not being used;
> 3 a black solid which formed on the base of the beaker.

(i) Explain why the spirit lamp and ethanol **lost mass** even when not in use.

....

....

(1)

(ii) Suggest the identity of the black solid. Explain why its formation will lead to a low value for the enthalpy of combustion.

Identity ....

Explanation ....

....

(2)
(Total 11 marks)

17. This question is about a self-heating can of coffee.

The bottom of the can has a compartment containing copper(II) nitrate solution. When a button on the bottom of the can is pressed, magnesium powder is released into the compartment where it reacts with the copper(II) nitrate solution.

(a) (i) Write an ionic equation for the reaction between magnesium powder and copper(II) ions. Include state symbols, but omit any spectator ions.

(2)

(ii) Show how the standard enthalpy change for this reaction could be calculated from the standard enthalpies of formation of copper(II) ions and magnesium ions. You should include a Hess cycle in your answer.

(3)

(b) The can contains 150 g of a solution of coffee in water.

The temperature of the solution needs to increase by 60 °C to produce a hot drink.

(i) Calculate the energy change needed to produce a temperature increase of 60 °C in the coffee, using the relationship

Energy change = 4.2 × mass of solution × temperature change.

Remember to include a unit in your answer.

(2)

(ii) The standard enthalpy change for this reaction is -530 kJ mol^{-1}.

Calculate the number of moles of reactants needed to produce the energy change in (i).

(1)

(iii) A solution of copper(II) nitrate of concentration 8.0 mol dm^{-3} is used.

Use your answer to (ii) to calculate the volume, in cm^3, of copper(II) nitrate solution needed.

Your answer should be given to two significant figures.

(1)

(c) Suggest TWO reasons why the temperature of the coffee may **not** increase by as much as 60 °C.

....

....

....

....

(2)
(Total 11 marks)

18. The reaction between chlorine and methane, in the presence of ultraviolet light, involves the formation of free radicals and includes the following steps:

A $Cl_2 \rightarrow 2Cl^\bullet$ $\Delta H^\ominus = +242 \text{ kJ mol}^{-1}$

B $CH_4 + Cl^\bullet \rightarrow HCl + CH_3^\bullet$ $\Delta H^\ominus = +4 \text{ kJ mol}^{-1}$

C $Cl_2 + CH_3^\bullet \rightarrow CH_3Cl + Cl^\bullet$ $\Delta H^\ominus = -97 \text{ kJ mol}^{-1}$

D $Cl^\bullet + Cl^\bullet \rightarrow Cl_2$

E $CH_3^\bullet + CH_3^\bullet \rightarrow CH_3CH_3$

F $Cl^\bullet + CH_3^\bullet \rightarrow CH_3Cl$ $\Delta H^\ominus = -339 \text{ kJ mol}^{-1}$

(a) (i) What is meant by a **free radical**? ....

....

....

(1)

(ii) Draw a 'dot-and-cross' diagram, showing outer shell electrons only, for a chlorine free radical.

(1)

(iii) What type of bond breaking occurs in step **A**?

....

(1)

(b) Which of the steps, **A** to **F**, are chain propagation steps?

....

(1)

(c) (i) Write the equation for the overall reaction between one mole of chlorine and one mole of methane molecules.

(1)

(ii) Calculate the standard enthalpy change, $\Delta H^\ominus$, for this reaction.

(2)

(d) (i) What is the value of $\Delta H^\ominus$ for step **D**? ....

(1)

(ii) Would you expect step **E** to be exothermic or endothermic? Justify your answer.

....

....

....

(1)

(e) The overall reaction was repeated using bromine gas instead of chlorine gas.

Would you expect step **A** for bromine to be more or less endothermic than step **A** for chlorine? Justify your answer.

....

....

....

(2)

(Total 11 marks)

19. In two similar, separate experiments the enthalpy changes for the reactions of sodium hydrogencarbonate and sodium carbonate with excess dilute hydrochloric acid were determined.

(a) The first experiment was to find the enthalpy change, ΔH_1, for the reaction

$$NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + CO_2(g) + H_2O(l)$$

Measurement	Reading
Mass of solid sodium hydrogencarbonate added to hydrochloric acid.	5.00 g
Volume of hydrochloric acid	50.0 cm^3
Temperature of hydrochloric acid before addition of solid sodium hydrogencarbonate	22.0 °C
Final temperature of solution	15.5 °C
Molar mass of sodium hydrogencarbonate	84.0 g mol^{-1}
Specific heat capacity of solution	4.18 J g^{-1} $°C^{-1}$

(i) Calculate the amount (moles) of sodium hydrogencarbonate used.

(1)

(ii) Calculate the heat absorbed in the reaction in kJ.
[Assume that 1 cm^3 of solution has a mass of 1 g]

(2)

(iii) Calculate the value of ΔH_1 in kJ mol^{-1}. Include a sign in your answer expressing it to a number of significant figures suggested by the data in the table.

(2)

(b) In the second experiment the enthalpy change for the reaction between sodium carbonate and dilute hydrochloric acid was measured.

$$Na_2CO_3(s) + 2HCl\,(aq) \rightarrow 2NaCI(aq) + CO_2(g) + H_2O(l)$$

The molar enthalpy change, ΔH_2, was calculated to be –35.6 kJ mol^{-1}

(i) Give TWO ways in which the temperature change differs when equal molar amounts of sodium hydrogencarbonate and sodium carbonate react separately with the same volume of hydrochloric acid.

....

....

....

(2)

(ii) Give ONE assumption that has been made in calculating the values of ΔH_1, and ΔH_2 from experimental results.

....

....

(1)
(Total 8 marks)

20. In the manufacture of beer, brewers often add small amounts of salts of Group 2 elements to the water used. These salts influence the chemical reactions during the brewing process. Two such salts are calcium sulphate and magnesium sulphate.

(a) A flame test can be used to confirm that a sample of a salt contains calcium ions.

(i) Describe how you would carry out a flame test.

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(3)

(ii) A positive test results in a brick-red flame colour. Describe the changes that occur in calcium ions to produce a colour.

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(2)

(iii) Impurities in the salt may lead to other colours being observed in the flame. What metal ion is likely to be present if a yellow flame is seen?

....

(1)

(b) Magnesium sulphate can be used in its anhydrous form, $MgSO_4(s)$, or in its hydrated form, $MgSO_4.7H_2O(s)$.

An experiment was carried out to find the enthalpy change when hydrated magnesium sulphate dissolved completely in water.

$$MgSO_4.7H_2O(s) \xrightarrow{\text{excess water}} MgSO_4(aq) + 7H_2O(l)$$

12.3 g of hydrated magnesium sulphate was added to 100 g of water in a simple calorimeter and the temperature was found to fall by 1.1 °C.

(i) Calculate the energy change, in joules, that occurred in the experiment, using the relationship

Energy change (J) = 4.18 × mass of water × temperature change

(2)

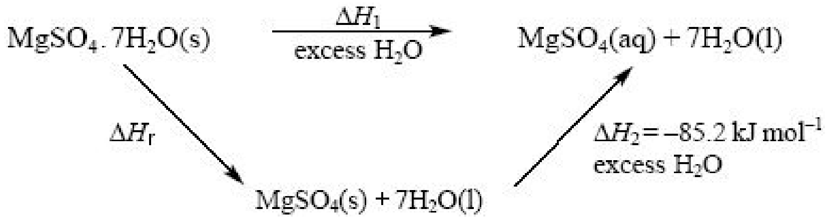
(ii) Calculate the number of moles of hydrated magnesium sulphate used in the experiment. Use the Periodic Table as a source of data.

(2)

(iii) Use your answers to (i) and (ii) to calculate the enthalpy change for the reaction. Include a sign and units in your final answer, which should be given to 2 significant figures.

(2)

(c) The enthalpy change as hydrated magnesium sulphate is converted to anhydrous magnesium sulphate is very difficult to measure. The Hess Cycle below can be used to find this enthalpy change, ΔH_r.



(i) Use the cycle to write an expression for ΔH_r using ΔH_1 and ΔH_2.

(1)

(ii) Use your expression in (c)(i) and your answer from (b)(iii) to calculate ΔH_r.

Include a sign and units in your final answer, which should be given to 2 significant figures.

(2)
(Total 15 marks)

21. Phosphine, PH_3, is a hydride of the Group 5 element, phosphorus.

(a) (i) Draw a ‘dot-and-cross’ diagram of a phosphine molecule. You should include only outer shell electrons.

(1)

(ii) Draw the shape you would expect for the phosphine molecule, suggesting a value for the HPH bond angle.

HPH bond angle ....

(2)

(iii) Explain the shape of the phosphine molecule you have given in your answer in (ii).

Justify your value for the HPH bond angle.

....

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....

(2)

(b) (i) Write a balanced equation, including state symbols, for the atomisation of phosphine gas.

....

(1)

(ii) Use your answer to (i) and the data below to calculate the standard enthalpy change of atomisation of phosphine at 298 K. Include a sign and units in your answer.

$\Delta H^{\ominus}_{f}[PH_3(g)] = +\ \ 5.4\ kJ\ mol^{-1}$

$\Delta H^{\ominus}_{at}[\frac{1}{2}H_2(g)] = +218.0\ kJ\ mol^{-1}$

$\Delta H^{\ominus}_{at}[P(s)] = +314.6\ kJ\ mol^{-1}$

(3)

(iii) Calculate a value for the bond energy of the bond between phosphorus and hydrogen, using your answer to (ii).

(1)
(Total 10 marks)

22. Methane, CH_4, is used as a domestic and industrial fuel and as a reagent in the petrochemical industry.

(a) Define the term **standard enthalpy of combustion**.

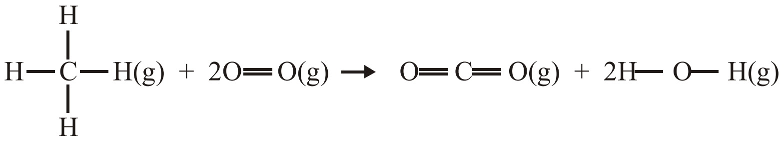
....

....

....

(3)

(b) Methane burns in oxygen according to the equation:



$$\mathrm{H{-}\underset{\underset{H}{|}}{\overset{\overset{H}{|}}{C}}{-}H(g)\ +\ 2O{=}O(g) \longrightarrow O{=}C{=}O(g)\ +\ 2H{-}O{-}H(g)}$$

Use the average bond enthalpy data shown below to calculate the enthalpy change of this reaction.

Bond	Bond enthalpy/kJ $\mathrm{mol^{-1}}$
C—H	+435
O==O	+498
C==O	+805
H—O	+464

(3)

(e) Methane is the feedstock in the manufacture of hydrogen according to the equation:

$$CH_4(g) + 2H_2O(g) \rightarrow CO_2(g) + 4H_2(g)$$

Given the enthalpy of formation data below, draw a **labelled** Hess's law cycle and use it to calculate the enthalpy change of this reaction.

Substance	Enthalpy of formation/kJ mol^{-1}
$CH_4(g)$	–75
$CO_2(g)$	–394
$H_2O(g)$	–242

(4)
(Total 10 marks)

23. In an experiment to find the enthalpy of neutralisation of a monobasic acid, HX, with an alkali, the following procedure was followed:

Step 1 25.0 cm^3 of 1.00 mol dm^{-3} dilute aqueous acid, HX, was measured into a polystyrene cup.

Step II A 0-100 °C thermometer was placed in the acid. The temperature of the acid was immediately read and recorded.

Step III 5.00 cm^3 portions of aqueous sodium hydroxide were added to the acid from a burette. After each addition, the temperature of the solution was read and recorded. The thermometer was removed and rinsed with water between each addition. A total of 50.0 cm^3 of aqueous sodium hydroxide was added.

(a) Suggest ONE change that could be made at **Step II** and ONE change that could be made at **Step III** to improve the accuracy of the experiment.

Step II ....

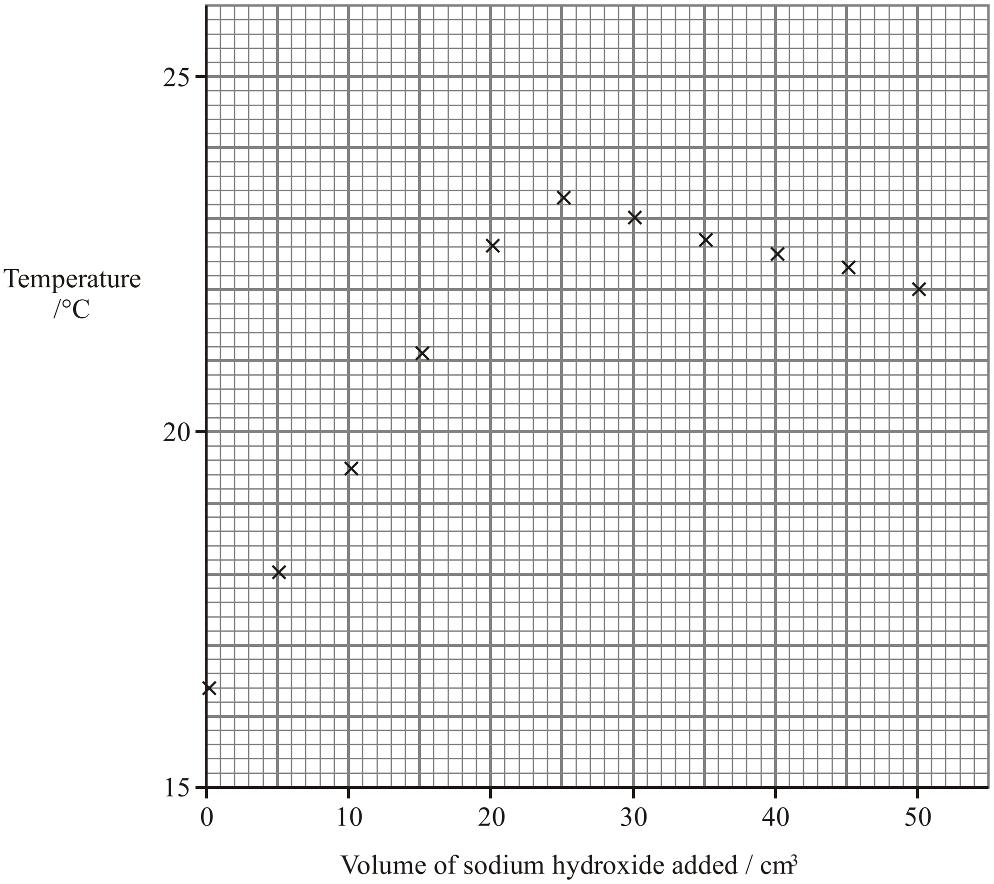
....

Step III ....

....

(2)

(b) The readings of temperature and volume are plotted on the grid. Draw two separate straight lines of best fit, extending the two lines so that they intersect.



(2)

(c) From the graph, read off the maximum temperature rise, ΔT, and the volume of aqueous sodium hydroxide added at neutralisation, V_N.

ΔT = ... °C V_N = ... cm^3

(2)

(d) (i) Use the formula below to calculate the heat evolved in the neutralisation.

$$\text{Heat evolved} = \frac{(V_N + 25) \times \Delta T \times 4.18}{1000}\ \text{kJ}$$

(1)

(ii) Given that the amount (moles) of acid neutralised was 0.025 mol, calculate the enthalpy of neutralisation, ΔH_{neut}, in units of kJ mol^{-1}.

ΔH_{neut} = ... kJ mol^{-1}

(2)
(Total 9 marks)

24. A reaction of ammonium dichromate(VI) is shown by the following equation.

$$(NH_4)_2Cr_2O_7(s) \xrightarrow{\text{heat}} N_2(g) + 4H_2O(g) + Cr_2O_3(s)$$

(a) What type of reaction is this?

....

(1)

(b) The enthalpy change for this reaction can be calculated from standard enthalpy changes of formation.

(i) State fully what is meant by the **standard enthalpy change of formation**, $\Delta H_f^\ominus$, of a compound.

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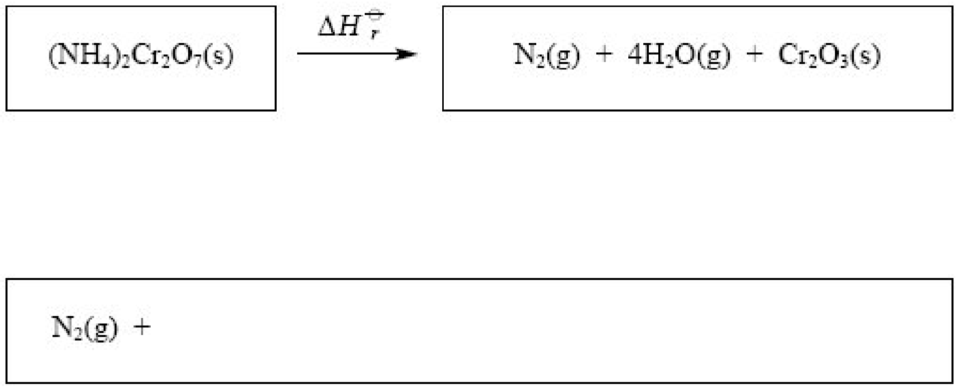
....

....

(3)

(ii) Complete the Hess cycle for the reaction so that you can calculate the enthalpy change of the reaction from standard enthalpy changes of formation.



(3)

(iii) What is the value of $\Delta H_f^\ominus[N_2(g)]$? ....

(1)

(iv) Calculate $\Delta H^\ominus_r$ for the reaction using the following data. Remember to include a sign and units in your answer.

$\Delta H_f^\ominus[(NH_4)_2Cr_2O_7(s)] = -1810 \text{ kJ mol}^{-1}$

$\Delta H_f^\ominus[H_2O(g)] = -242 \text{ kJ mol}^{-1}$

$\Delta H_f^\ominus[Cr_2O_3(s)] = -1140 \text{ kJ mol}^{-1}$

(3)

(c) In this reaction, water vapour is formed which condenses to liquid water on cooling. Is this reaction $H_2O(g) \rightarrow H_2O(l)$ exothermic or endothermic?

Justify your answer.

....

....

....

....

(2)
(Total 13 marks)

25. Chlorine can be converted to the gas chlorine(I) oxide, Cl_2O.

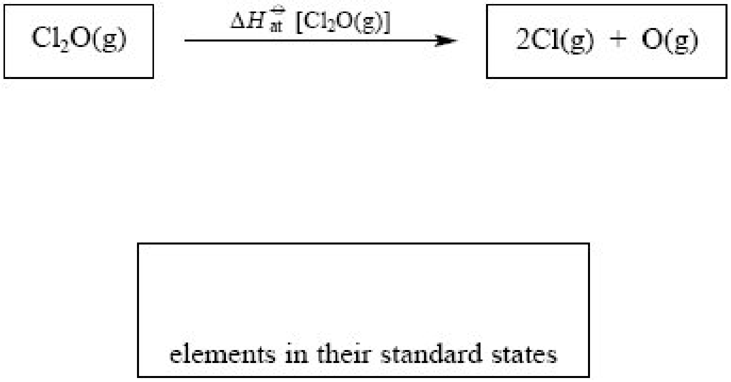
The standard molar enthalpy change of formation of chlorine(I) oxide and the standard molar enthalpy changes of atomisation of chlorine and oxygen are given below:

$\Delta H_f^\ominus [Cl_2O(g)] = +80.3 \text{ kJ mol}^{-1}$

$\Delta H_{at}^\ominus [\tfrac{1}{2}Cl_2(g)] = +121.7 \text{ kJ mol}^{-1}$

$\Delta H_{at}^\ominus [\tfrac{1}{2}O_2(g)] = +249.2 \text{ kJ mol}^{-1}$

A partially completed Hess cycle involving chlorine(I) oxide is shown below:



(i) Insert the appropriate formulae, showing the correct quantities of each element, into the box above. Include state symbols in your answer.

(1)

(ii) Insert arrows between the boxes, writing the correct numerical data alongside the appropriate arrows.

(1)

(iii) Use the cycle to calculate $\Delta H_{at}^{\ominus}$ [$Cl_2O(g)$].

(1)

(iv) Calculate the Cl—O bond energy in chlorine(I) oxide.

(1)
(Total 14 marks)

26. (a) Enthalpy changes can be calculated using average bond enthalpy data.

(i) The enthalpy change to convert methane into gaseous atoms is shown below.

$$CH_4(g) \rightarrow C(g) + 4H(g) \qquad \Delta H = +1664 \text{ kJ mol}^{-1}$$

Calculate the average bond enthalpy of a C—H bond in methane.

(1)

(ii) Use the data in the table below and your answer to (a)(i) to calculate the enthalpy change for

$$2C(g) + 2H_2(g) + Br_2(g) \rightarrow CH_2BrCH_2Br(g)$$

Bond	Average bond enthalpy / kJ mol^{-1}	Bond	Average bond enthalpy / kJ mol^{-1}
C—C	+348	H—H	+436
Br—Br	+193	C—Br	+276

(3)

(b) The standard enthalpy of formation of 1,2-dibromoethane, CH_2BrCH_2Br, is –37.8 kJ mol^{-1}.

Suggest the main reason for the difference between this value and your calculated value in (a)(ii).

....

....

....

(1)
(Total 5 marks)

27. In an experiment to find the enthalpy change for the reaction

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

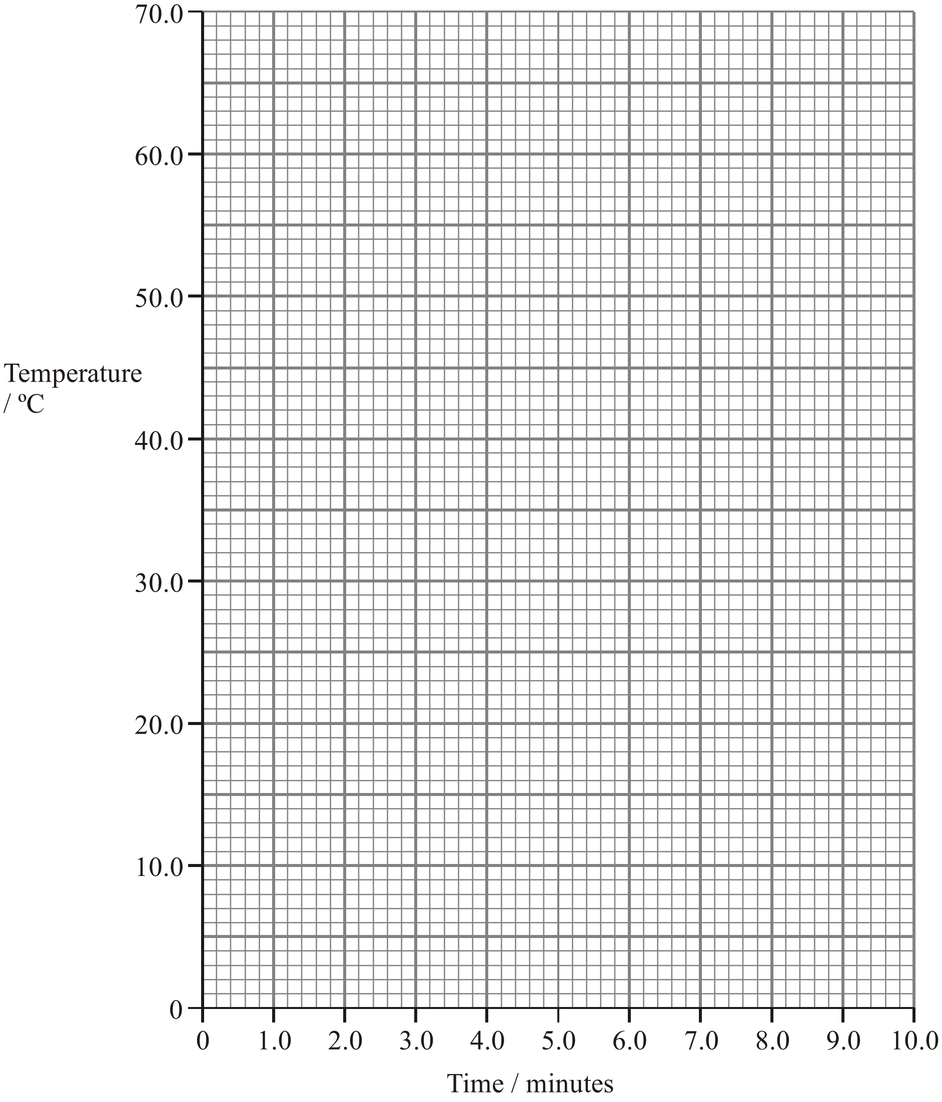
a student was given the following list of instructions:

- weigh out 5.0 g of zinc powder into a weighing bottle
- use a measuring cylinder to transfer 50 cm^3 of 1.0 mol dm^{-3} aqueous copper(II) sulphate into a polystyrene cup, firmly held in a 250 cm^3 beaker
- stir the solution with the thermometer and record the temperature to the nearest 0.5 °C
- continue to stir the solution, recording its temperature every minute
- at exactly 3.5 minutes, add the zinc powder to the aqueous copper(II) sulphate, stirring continuously
- record the temperature of the solution every minute from 4.0 to 9.0 minutes.

The temperature readings obtained are shown in the table below.

Time/min	0	1.0	2.0	3.0	4.0	5.0	6.0	7.0	8.0	9.0
Temperature/°C	20.0	20.0	20.0	20.0	63.0	60.5	59.0	57.0	55.5	53.0

(a) (i) Plot a graph of temperature against time on the grid below.



(2)

(ii) Use the graph to calculate the maximum temperature change, ΔT. Show clearly on the graph how you obtained your answer.

ΔT = ...°C

(2)

(iii) Give ONE reason why a series of temperature readings is obtained instead of just the starting and maximum temperatures.

....

....

(1)

(b) (i) Calculate the heat change, in joules.

The specific heat capacity of the solution is 4.18 J g^{-1} $°C^{-1}$.

(1)

(ii) What assumption have you made about the solution in your calculation in (i)?

....

....

(1)

(iii) Calculate the amount (moles) of copper (II) sulphate, $CuSO_4$, in 50 cm^3 of a 1.0 mol dm^{-3} solution.

(1)

(iv) Calculate the enthalpy change for this reaction in kJ mol^{-1}.

(2)

(c) Suggest TWO improvements that could be made to the experimental procedure.

Give a reason for each.

Improvement 1 ....

....

....

Reason ....

....

....

Improvement 2 ....

....

....

Reason ....

....

....

(4)
(Total 14 marks)

28. The values of the lattice energies of potassium iodide and calcium iodide experimentally determined from Born-Haber cycles and theoretically calculated from an ionic model are shown below.

	Experimental lattice energy /kJ mol^{-1}	Theoretical lattice energy /kJ mol^{-1}
Potassium iodide, KI(s)	– 651	– 636
Calcium iodide, CaI$_2$(s)	–2074	–1905

(i) Explain why the experimental lattice energy of potassium iodide is less exothermic than the experimental lattice energy of calcium iodide.

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(3)

(ii) Explain why the experimental and theoretical values of the lattice energy are almost the same for potassium iodide, but are significantly different for calcium iodide.

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(3)

(Total 6 marks)

29. When solutions of potassium carbonate and calcium chloride are mixed together, the following reaction takes place

$$CaCl_2(aq) + K_2CO_3(aq) \rightarrow CaCO_3(s) + 2KCl(aq)$$

(a) Re-write the above equation as an ionic equation. Include state symbols, but omit any spectator ions.

(2)

(b) An experiment was carried out to measure the enthalpy change for this reaction. 50 cm^3 of a 1.00 mol dm^{-3} solution of potassium carbonate was added to 50 cm^3 of a 1.00 mol dm^{-3} solution of calcium chloride. The temperature fell by 1.5 °C.

(i) Calculate the energy taken in from the surroundings using the relationship

energy =	mass of solution	×	specific heat capacity of solution	×	temperature change
/J	/g		/J g^{-1} °C^{-1}		/°C

You may assume that

- 1.0 cm^3 of solution has a mass of 1.0 g.
- The specific heat capacity of the solution is 4.2 J g^{-1} °C^{-1}.

Energy taken in = ... J

(1)

(ii) How many moles of calcium chloride are used in this experiment?

(1)

(iii) Calculate the enthalpy change for the reaction, giving your answer to two significant figures. Include a sign and units in your answer.

(2)

(iv) Which measurement is likely to have caused the major source of error in this experiment? Explain your answer.

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(1)

(v) What apparatus should be used to contain the reaction mixture during this experiment?

....

(1)

(c) If the experiment in (b) was repeated, but using only 25 cm^3 of each solution, predict what the fall in temperature would be.

....

(1)

(Total 9 marks)

30. (a) State Hess's Law.

....

....

....

....

(2)

(b) Methane burns in oxygen.

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

(i) Calculate the enthalpy change for this reaction, using the bond enthalpies given below.

	Bond enthalpy / kJ mol^{-1}
C – H	+435
O = O	+498
C = O	+805
H – O	+464

(3)

(ii) State the name of this enthalpy change.

....

(1)

(iii) The value of this enthalpy change, under standard conditions, is –890 kJ mol^{-1}. State the meaning of **standard conditions.**

....

....

....

(2)

(iv) Suggest, with a reason, why the enthalpy change calculated in (i) is different from the standard value quoted in (iii).

....

....

....

....

(2)

(c) Although the reaction between methane and oxygen is exothermic, it does not occur unless the mixture is ignited.

Use these facts to explain the difference between thermodynamic and kinetic stability.

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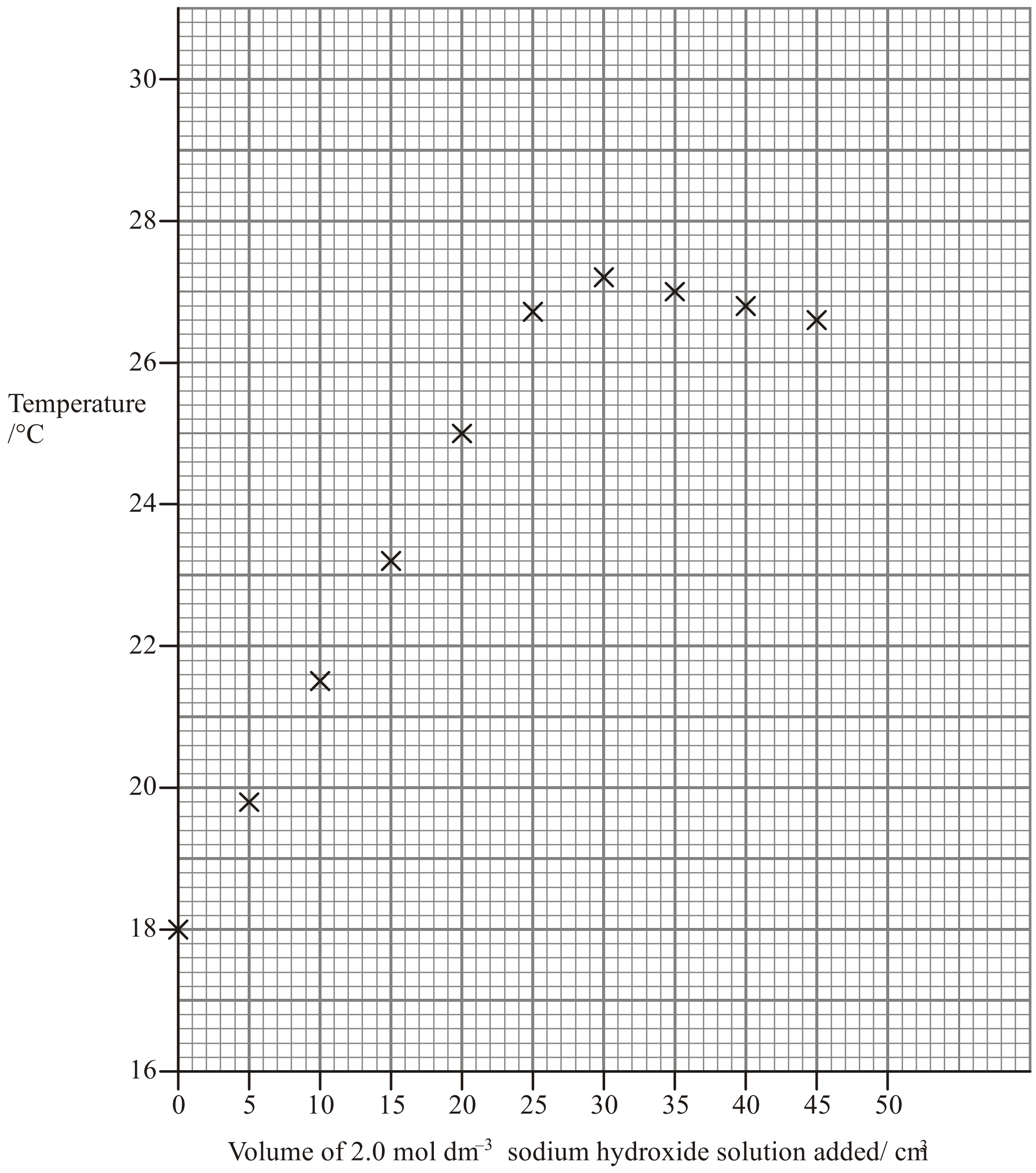
(4)
(Total 14 marks)

31. A student carried out an experiment to find the concentration of a solution of nitric acid and also its enthalpy of neutralisation.

- The solutions of nitric acid and sodium hydroxide were allowed to reach the same temperature.
- 50.0 cm^3 of the nitric acid was pipetted into a polystyrene cup.
- A burette was filled with a solution of 2.0 mol dm^3 sodium hydroxide, NaOH.
- The initial temperature of the acid was recorded.
- The sodium hydroxide was added to the acid in 5.0 cm^3 portions.
- After each addition, the mixture was stirred and the maximum temperature recorded.
- This was repeated until 45 cm^3 of the sodium hydroxide solution had been added.

The student plotted the results, as shown below.



(a) Complete the graph by drawing two intersecting straight **lines** of best fit.

(1)

(b) The point where the lines cross represents the neutralisation of the nitric acid by the sodium hydroxide solution.

Use the graph to find:

(i) the volume of 2.0 mol dm^3 sodium hydroxide solution, NaOH, that reacts exactly with the 50 cm^3 of the nitric acid.

....

(1)

(ii) the maximum temperature **change,** ΔT, in the reaction.

....

(1)

(c) The information in (b)(i) and the equation below can be used to calculate the concentration of the nitric acid.

$$NaOH(aq) + HNO_3(aq) \rightarrow NaNO_3(aq) + H_2O(l)$$

(i) Calculate the amount (moles) of sodium hydroxide used to neutralise the 50 cm^3 of nitric acid.

(1)

(ii) Write the amount (moles) of nitric acid in 50.0 cm^3 of the solution.

....

(1)

(iii) Hence calculate the concentration of nitric acid, HNO_3, in mol dm^3.

(2)

(d) (i) Use the data from (b) to calculate the heat change for this reaction.

The density of the mixture produced at neutralisation is 1.0g cm^{-3} and the specific heat capacity of the mixture is 4.2 J g^{-1} $°C^{-1}$.

Heat change = mass × specific heat capacity × ΔT

(2)

(ii) Use your answer from (d)(i) and (c)(iii) to calculate the enthalpy of neutralisation per mole of nitric acid, HNO_3. Include a sign and units with your answer.

(3)

(e) The enthalpy of neutralisation found by this method may be **less exothermic** than the data book value because of heat loss.

Suggest ONE way to reduce the error due to heat loss.

....

....

(1)
(Total 13 marks)

32. (a) Calculate the number of atoms in 3.50 g of lithium.

Use the Periodic Table as a source of data.

[The Avogadro constant, $L = 6.02 \times 10^{23}$ mol^{-1}]

(2)

(b) The equation for the reaction of lithium with hydrochloric acid is shown below.

$$2Li(s) + 2HCl(aq) \rightarrow 2LiCl(aq) + H_2(g)$$

(i) Rewrite this equation as an ionic equation, omitting the spectator ions.

(1)

(ii) Draw a ‘dot and cross’ diagram of lithium chloride showing **all** the electrons. Indicate charges clearly on your diagram.

(2)

(iii) The value of the standard enthalpy change for the reaction, $\Delta H^{\ominus}$, is –557 kJ mol^{-1}. State TWO of the reaction conditions necessary for this enthalpy change to be **standard**.

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(2)
(Total 7 marks)

33. (a) Define the term **standard enthalpy of formation**.

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....

(3)

(b) The dissociation of phosphorus pentachloride is a reversible reaction.

$$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$$

(i) Use the values of enthalpy of formation given to calculate ΔH for the forward reaction.

	ΔH_f / kJ mol^{-1}
$PCl_5(g)$	– 399
$PCl_3(g)$	– 306

(1)

(ii) Explain, with reasons, the effect that raising the temperature would have on the composition of the equilibrium mixture.

....

....

....

(2)

(iii) Other than by changing the temperature, suggest how the amount of PCl_5 present at equilibrium could be increased. Give a reason for your answer.

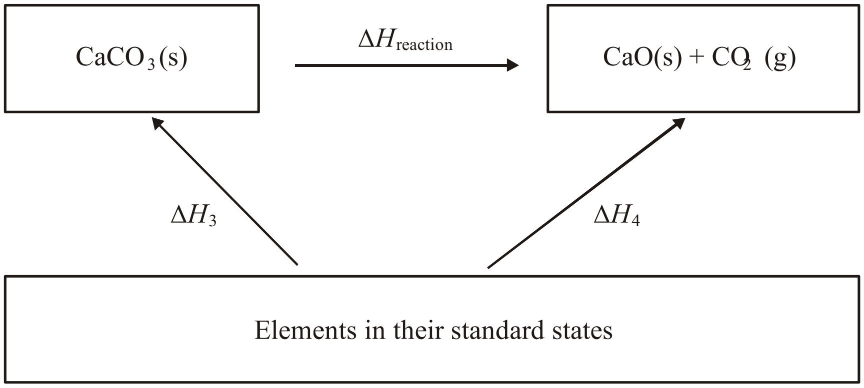
....

....

....

(2)
(Total 8 marks)

34. The enthalpy change for the thermal decomposition of calcium carbonate cannot be measured directly, but can be found by carrying out two reactions as shown in the Hess cycle below.



(a) Suggest ONE reason why it is difficult to measure $\Delta H_{reaction}$ directly by experiment.

....

....

....

....

(1)

(b) In an experiment to find ΔH_1 a student added 2.00 g of finely powdered calcium carbonate to 20.0 cm^3 of 2.50 mol dm^{-3} hydrochloric acid solution (an excess) in a polystyrene container. The temperature rose from 20.5 °C to 23.0 °C.

(i) Why is the calcium carbonate used in this experiment finely powdered, rather than in lumps? Explain why this is important for an accurate result.

....

....

....

(2)

(ii) Calculate the energy change using the relationship below.

$$\underset{/\mathrm{J}}{\text{Energy change}} = \underset{/\mathrm{J\,g^{-1}\,K^{-1}}}{4.2} \times \underset{/\mathrm{g}}{\text{mass of solution}} \times \underset{/\mathrm{K}}{\text{temperature change}}$$

Assume that the mass of the solution is 20 g.

(1)

(iii) Calculate the enthalpy change, ΔH_1. Include a sign and units in your answer.
[The molar mass of $CaCO_3$ is 100 g mol^{-1}]

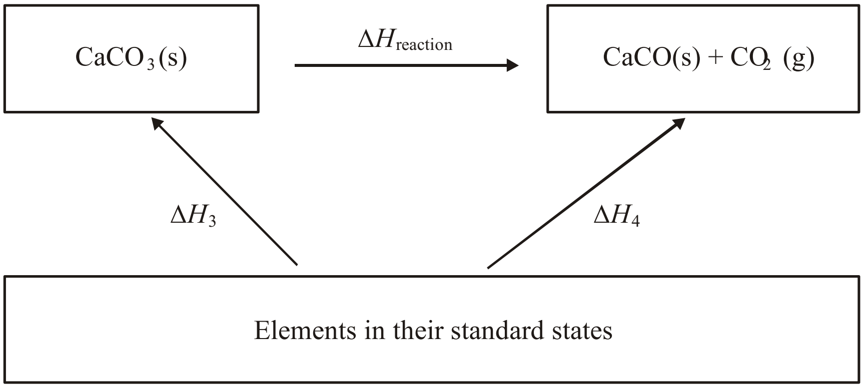
(3)

(iv) In another experiment, the value of ΔH_2 was found to be –181 kJ mol^{-1}.

Use this result and your answer to (iii) to calculate the value of $\Delta H_{reaction}$.

(2)

(c) The student checked the experimental results using information from the *Book of data* in another Hess cycle.

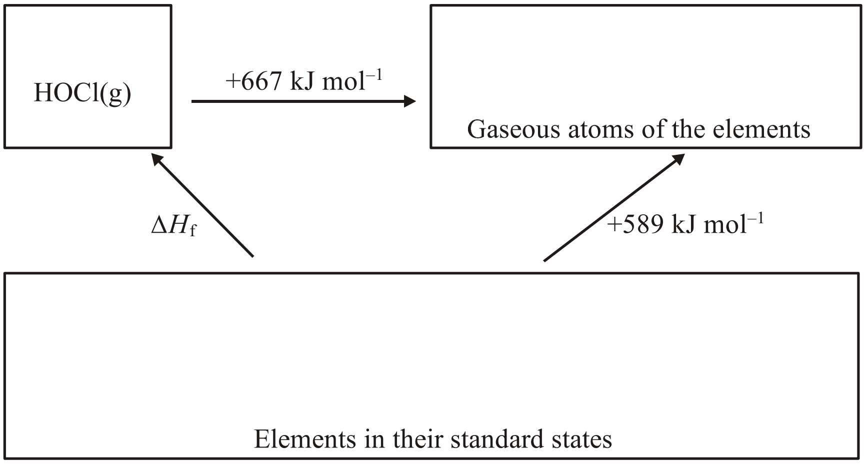


Name the enthalpy change represented by ΔH_3.

....

(1)
(Total 10 marks)

35. The Hess cycle below can be used to estimate the enthalpy change of formation, ΔH_f, of the unstable gaseous compound with the formula HOCl(g).



(a) (i) Insert formulae, with state symbols, into the appropriate boxes, to show the correct quantities of each element.

(1)

(ii) Use the cycle to calculate a value for the enthalpy change of formation, ΔH_f [HOCl(g)].

(1)

(iii) Assuming that the H—O bond energy is +464 kJ mol^{-1}, calculate a value for the O—Cl bond energy.

(1)

(b) (i) Draw a ‘dot and cross’ diagram for the HOCl molecule showing outer electrons only.

(2)

(ii) Predict the HOCl bond angle. Justify your answer.

Angle ....

Justification ....

....

....

....

....

(2)

(c) HOCl(g) can be made from chlorine(I) oxide by the reversible reaction

$$Cl_2O(g) + H_2O(g) \rightleftharpoons 2HOCl(g)$$

What effect, if any, would an increase in pressure have on the proportion of HOCl(g) at equilibrium? Justify your answer.

....

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....

(2)
(Total 9 marks)

36. (a) Define the term **standard enthalpy of formation**.

....

....

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....

....

(3)

(b) In the Haber process, ammonia is manufactured from nitrogen and hydrogen as shown in the equation.

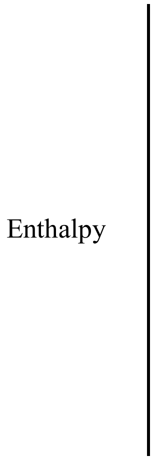
$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

(i) Use the bond enthalpies below to calculate the standard enthalpy of formation of ammonia.

Bond	**Bond enthalpy / kJ mol^{-1}**
N≡N in N^2	+945
H–H in H^2	+436
N–H in NH^3	+391

(4)

(ii) Draw a labelled enthalpy level diagram for the formation of ammonia in the Haber process.



(2)

(iii) State the temperature used in the Haber process and explain in terms of the rate of reaction and position of equilibrium, why this temperature is chosen.

Temperature ....

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....

....

(3)

(iv) Identify the catalyst used in the Haber process and state what effect, if any, it has on the equilibrium yield of ammonia.

Catalyst ....

Effect on yield ....

(2)

(v) Explain why it is necessary to use a catalyst in this process.

....

....

....

....

(1)

(c) The pressure used in the Haber process is 250 atmospheres.

(i) State and explain an advantage of increasing the pressure to 1000 atmospheres.

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(2)

(ii) Suggest a disadvantage of using a pressure of 1000 atmospheres.

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(1)
(Total 18 marks)

37. The enthalpy change for the reaction of anhydrous aluminium chloride, $AlCl_3$, with water can be found as follows:

- Add about 100 cm^3 of distilled water to a weighed polystyrene cup.
- Measure the steady temperature of the water.
- Add anhydrous aluminium chloride to the polystyrene cup, with stirring.
- Measure the highest temperature reached.
- Re-weigh the polystyrene cup and contents

Data
Mass of anhydrous aluminium chloride = 4.00 g
Mass of solution = 104 g
Initial temperature = 17.5 °C
Highest temperature reached = 43.5 °C
Specific heat capacity of the solution = 4.09 J g^{-1} $°C^{-1}$

(i) Calculate the heat change in this experiment.

(2)

(ii) Assuming that 100 cm^3 of water is a large excess, calculate the enthalpy change, in kJ mol^{-1}, when one mole of aluminium chloride reacts. Include a sign and unit in your answer.

(3)
(Total 5 marks)

38. (a) (i) Write the equation which represents the change occurring when the standard enthalpy of atomisation of bromine is measured.

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(2)

(ii) Define **lattice energy**.

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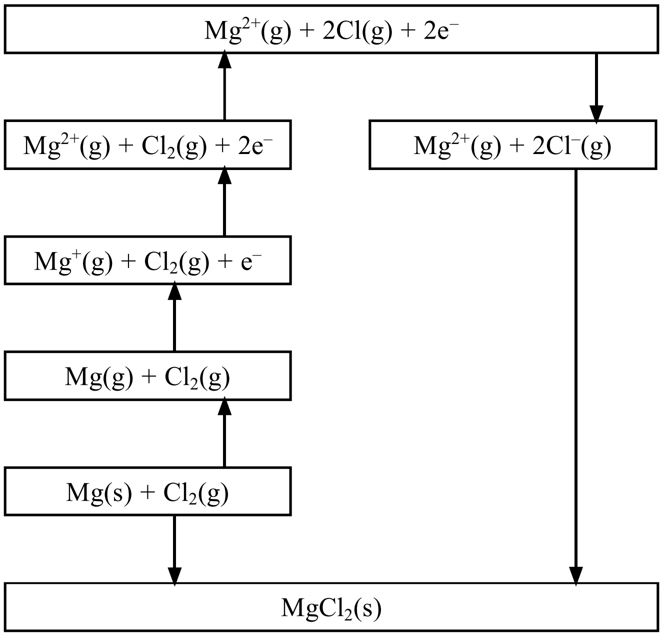
(3)

(b) A Born-Haber cycle for the formation of magnesium chloride is shown below.



	Value / kJ mol^{-1}
Enthalpy of atomisation of magnesium	+150
1st ionisation energy of magnesium	+736
2nd ionisation energy of magnesium	+1450
Enthalpy of atomisation of chlorine	+122
Enthalpy of formation of magnesium chloride	–642
Lattice energy of magnesium chloride	–2526

The theoretically calculated value for the lattice energy of magnesium chloride is –2326 kJ mol^{-1}.

Explain, in terms of the bonding in magnesium chloride, why the experimentally determined value of –2526 kJ mol^{-1} is significantly different from the theoretical value.

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(2)

(c) The table shows values for the lattice energies of the metal chlorides of some Group 2 metals.

Group 2 metal chloride	$MgCl_2$	$CaCl_2$	$SrCl_2$	$BaCl_2$
Lattice energy/ kJ mol^{-1}	–2526	–2237	–2112	–2018

Explain why these lattice energies become less exothermic from $MgCl_2$ to $BaCl_2$.

(3)
(Total 10 marks)

39. An experiment was carried out to find the enthalpy change for the reaction of zinc powder with copper(II) sulphate solution.

$$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$$

50cm^3 of copper(II) sulphate solution, of concentration 1.0 mol dm^{-3}, was put into a polystyrene cup and the temperature of the solution measured. After one minute, 5.0 g of zinc powder was added, the mixture stirred with a thermometer and the temperature measured every 30 s.

(a) (i) What is meant by a spectator ion?

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(1)

(ii) Give the formula of the spectator ion in this reaction.

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(1)

(iii) Write the equation for this reaction, omitting the spectator ion.

(1)

(b) How would you measure the 50 cm^3 of copper(II) sulphate solution?

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(1)

(c) Give TWO reasons why it is better to use a polystyrene cup, rather than a metal container, to obtain more accurate results.

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(2)

(d) Calculate the number of moles of each of the reactants and hence deduce which reactant

is completely used up.
Use the Periodic Table as a source of data.

Moles of zinc powder

Moles of copper(II) sulphate

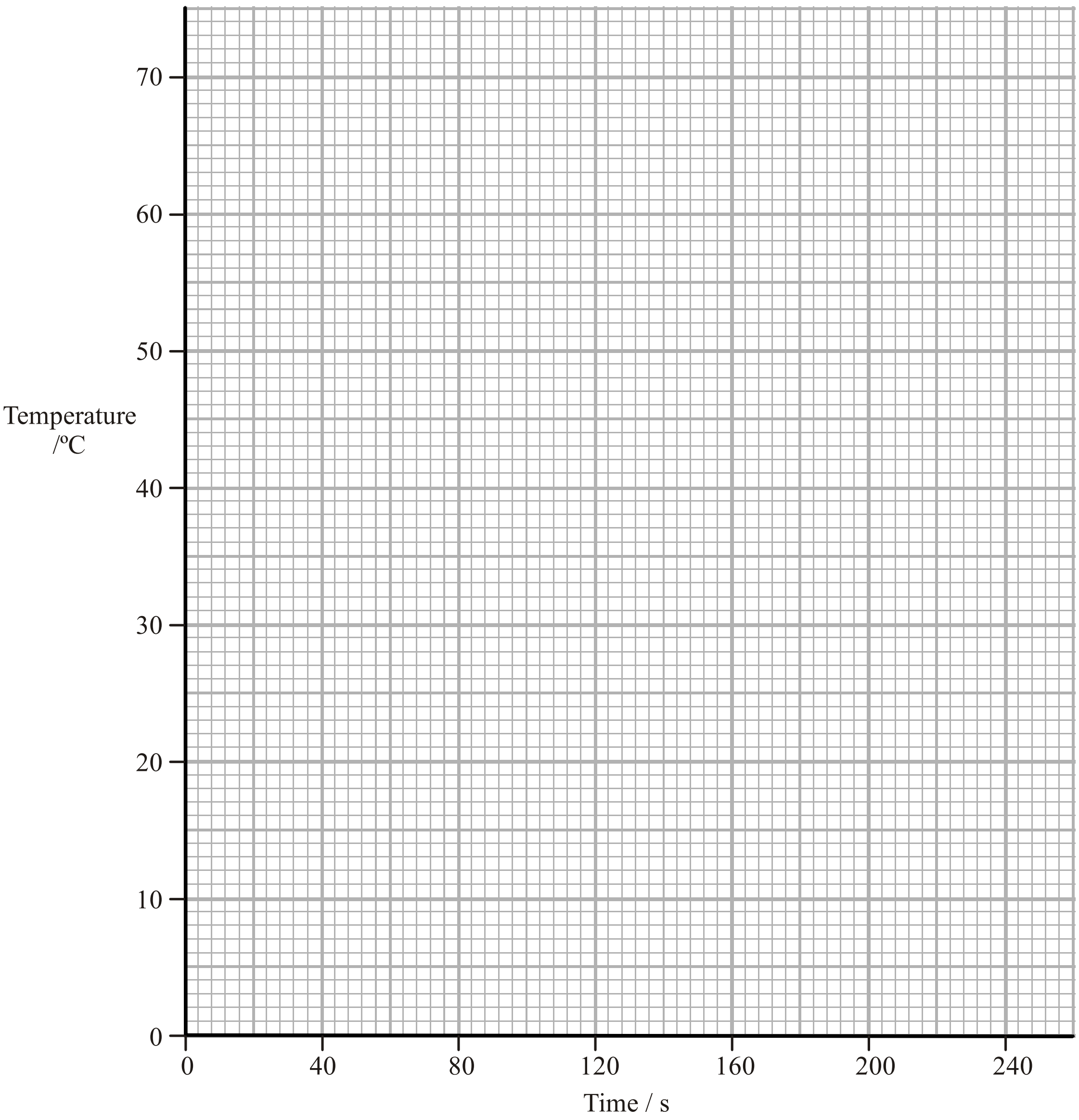
Reactant used up ....

(3)

(e) The following results were obtained.

Time /s	0	60	90	120	150	180	210
Temperature /°C	22	22	60	65	63	61	59

(i) On the graph paper below, plot the results of this experiment.



(2)

(ii) Explain the shape of your graph

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(2)

(iii) The maximum recorded temperature in this experiment was 65°C. Use your graph to estimate a more accurate maximum temperature.

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(1)

(f) (i) Calculate the energy change in this experiment using your answer to (e)(iii) and the relationship

energy change =	mass of solution	×	specific heat capacity of solution	×	temperature **rise**
/J	/g		/J °C^{-1} g^{-1}		/°C

You may assume that

- 1.0 cm^3 of solution has a mass of 1.0 g
- The specific heat capacity of the solution is 4.2 J °C^{-1} g^{-1}

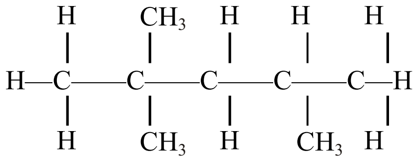
(1)

(ii) Use your answers to (d) and (f)(i) to calculate ΔH for this reaction. Include a sign and units in your answer.

(3)
(Total 18 marks)

40. This question is about some of the chemicals used in car engines and their reactions.

(a) Compound **X**, shown below, is one component of petrol.



$$\begin{array}{ccccccccccc} & & H & & CH_3 & & H & & H & & H \\ & & | & & | & & | & & | & & | \\ H & - & C & - & C & - & C & - & C & - & C & - & H \\ & & | & & | & & | & & | & & | \\ & & H & & CH_3 & & H & & CH_3 & & H \end{array}$$

(i) Name **X**.

....

(1)

(ii) Give the **empirical** formula of **X**.

....

(1)

(iii) **X** can be made by cracking decane, $C_{10}H_{22}$.

Assuming only one other product forms in a cracking reaction, deduce the **molecular** formula of this other product.

(1)

(iv) What is the sign of the enthalpy change for the reaction in which decane is cracked? Give a reason for your answer.

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(1)

(v) If the air supply in a car engine is poor, there is not enough air for carbon dioxide to be produced.

Use this information to suggest ONE possible equation for the combustion of **X** in this engine. Use the molecular formula of **X** in your equation.

(2)

(b) When air enters a car engine, as well as the fuel burning, nitrogen and oxygen can react to form nitrogen(II) oxide.

$$N2(g) + O_2(g) \rightleftharpoons 2NO(g) \qquad \Delta H = +\ 180\ \text{kJ mol}^{-1}$$

(i) What, if any, is the effect on the percentage of nitrogen(II) oxide in an equilibrium mixture of these three gases if the pressure and temperature are increased? Explain your answers.

Increase in pressure

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....

Increase in temperature

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(2)

(ii) In a car exhaust pipe, nitrogen(II) oxide passes over a catalytic converter. The following reaction occurs.

$2NO(g) + 2CO(g) \rightarrow N_2(g) + 2CO_2(g) \qquad \Delta H = -746\ \text{kJ mol}^{-1}$

Explain why this reaction speeds up when the car engine has been running for a few minutes.

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(1)

(iii) A textbook says "The catalytic converter converts the gases coming out of the engine into less harmful ones".

State, with a reason, which of the four gases in the equation in (ii) you consider to be **least** harmful.

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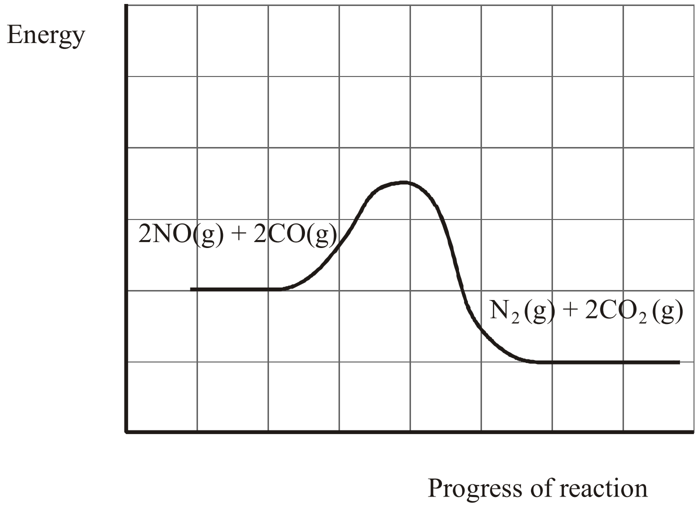
(1)

(iv) The diagram below shows the reaction profile for the change which occurs in the catalytic converter.



On the diagram, show the activation energy, E_A.

Add a line showing the reaction profile if no catalyst is present.

(2)
(Total 12 marks)

41. (a) Define the term **standard enthalpy of combustion**.

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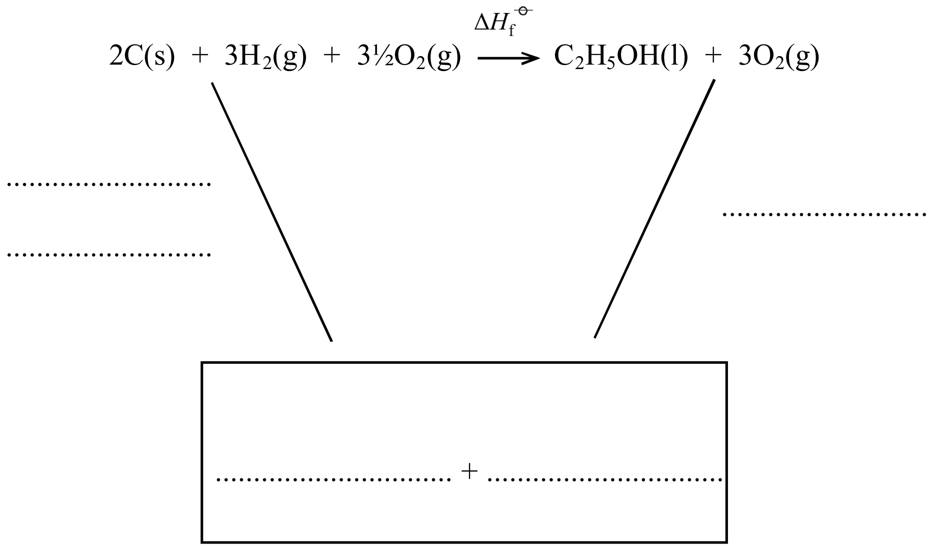
(3)

(b) The following standard enthalpies of combustion are needed to calculate the standard enthalpy of formation of ethanol, C_2H_5OH.

Substance	**Standard enthalpy of combustion /kJ mol^{-1}**
carbon, C (s, graphite)	–394
hydrogen, H_2(g)	–286
ethanol, C_2H_5OH (l)	–1371

(i) Complete the Hess's Law cycle by filling in the box and labelling the arrows with the enthalpy changes.



$$2C(s) + 3H_2(g) + 3\tfrac{1}{2}O_2(g) \xrightarrow{\Delta H_f^\ominus} C_2H_5OH(l) + 3O_2(g)$$

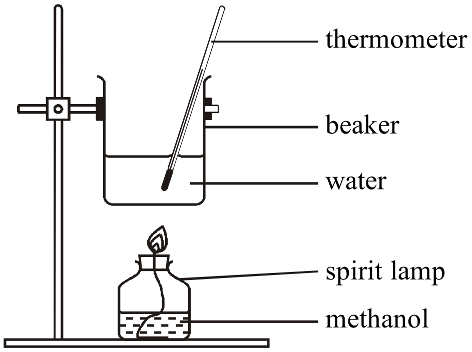
(3)

(ii) Use your Hess's Law cycle to calculate the standard enthalpy of formation of ethanol.

(2)
(Total 8 marks)

42. The apparatus used and the recordings made by a student, carrying out an experiment to determine the enthalpy of combustion of methanol, are shown below.

Diagram



Results

Molar mass (methanol) = 32 g mol^{-1}
Volume of water in beaker = 50 cm^3
Mass of water in beaker = 50 g
Weighings
    Spirit lamp + methanol before combustion = 163.78 g
    Spirit lamp + methanol after combustion = 163.44 g
Temperatures
    Water before heating = 22.0 °C
    Water after heating = 43.5 °C
Specific heat capacity of water = 4.18 J g^{-1} $°C^{-1}$

Observations

- When the spirit lamp was being weighed its mass was continually falling.
- A black substance formed on the bottom of the beaker as the methanol burned.

(a) (i) Calculate the amount (moles) of methanol, CH_3OH, burned.

(2)

(ii) Calculate the heat gained by the water. Give your answer in kJ.

(2)

(iii) Use your values from (i) and (ii) to calculate the enthalpy of combustion of methanol in kJ mol^{-1}. Include a sign with your answer.

ΔH = ... kJ mol^{-1}

(2)

(b) (i) The thermometer used in the experiment can be read to an accuracy of±0.5 °C. Calculate the percentage error in the temperature change.

(1)

(ii) Calculate the maximum temperature change that could have occurred during the experiment.

(1)

(c) (i) Give a reason why the mass of the spirit lamp fell as it was being weighed.

....

....

(1)

(ii) Suggest the identity of the black substance that forms on the beaker. State the effect on the value of the enthalpy of combustion obtained.

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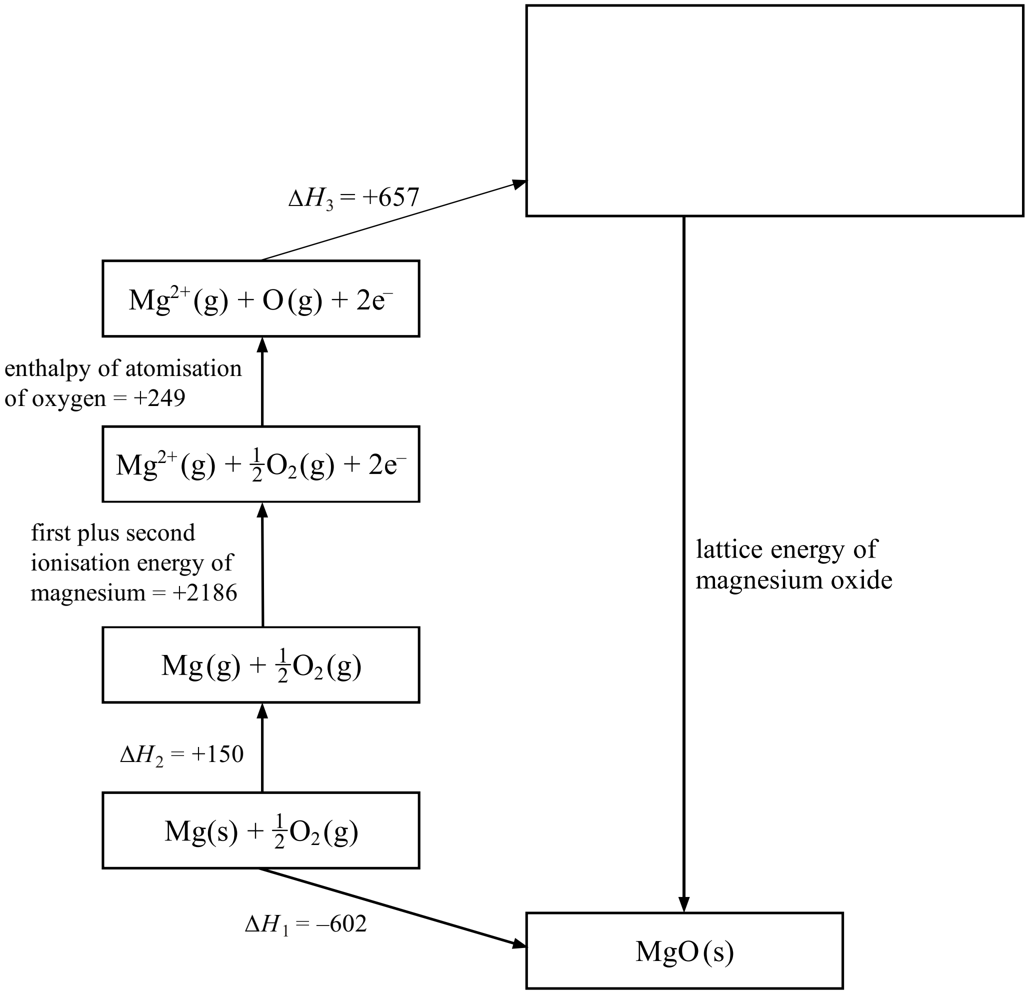
(2)

(Total 11 marks)

43. (a) An incomplete Born-Haber cycle for the formation of magnesium oxide, MgO, from its constituent elements is shown below. All numerical values are in kJ mol^{-1}.



(i) Complete the empty box on the cycle by writing in the formulae of the missing species. State symbols are required.

(1)

(ii) Identify each of the following enthalpy changes by name:

ΔH_1 ....

ΔH_2 ....

ΔH_3 ....

(3)

(iii) Use the Born-Haber cycle to calculate the lattice energy of magnesium oxide.

(2)

(b) Magnesium iodide is another compound of magnesium. The radius of the magnesium ion is 0.072 nm, whereas the radius of the iodide ion is much larger and is 0.215 nm.

(i) Describe the effect that the magnesium ion has on an iodide ion next to it in the magnesium iodide lattice.

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(1)

(ii) What TWO quantities must be known about the ions in a compound in order to calculate a theoretical lattice energy?

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(2)

(iii) Suggest how the value of the theoretical lattice energy would compare with the experimental value from a Born-Haber Cycle for magnesium iodide.

Give a reason for your answer.

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(2)
(Total 11 marks)

44. Calcium hydroxide decomposes on strong heating to form calcium oxide and water.

$$Ca(OH)_2(s) \rightarrow CaO(s) + H_2O(l)$$

Two samples of **calcium hydroxide** were taken, each weighing exactly 1.00 g.

The first sample was cautiously added to 25.0 cm^3 of dilute hydrochloric acid contained in a glass beaker. The temperature rise was measured and found to be 16.5 °C.

The other sample was heated for some time. It was then allowed to cool and then added to another 25.0 cm3 portion of hydrochloric acid as before. In this case the temperature rose by 25.5 °C.

In both cases, the acid used was an excess.

(a) (i) Calculate the energy produced by the reaction of each solid with the acid.

Use the relationship

Energy produced / J = mass of solution / g × 4.2 / J °C^{-1} g^{-1} × temperature rise /°C

You may assume that 1.0 cm^3 of solution has a mass of 1.0 g. Ignore the mass of the solid.

For the solid calcium hydroxide

For the solid calcium oxide

(1)

(ii) How many moles of calcium hydroxide were used in each experiment?
[Molar mass of $Ca(OH)_2$ = 74.0 g mol^{-1}]

(1)

(iii) Using your answers to (a)(i) and (ii), calculate the enthalpy changes for each reaction.

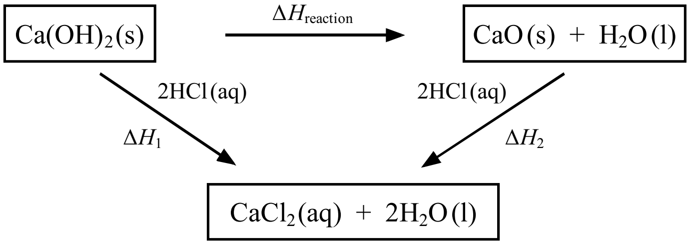
Give your answers to **two significant figures**. Include the sign and units for each answer.

For the solid calcium hydroxide, ΔH_1

For the solid calcium oxide, ΔH_2

(2)

(b) A Hess cycle for all these reactions is shown below.



(i) Use this Hess cycle and your answers in (a)(iii) to calculate $\Delta H_{\text{reaction}}$. Include a sign and units.

(2)

(ii) Apart from the approximations involved in using the equation given in (a)(i), give TWO other potential sources of error which are likely to affect the accuracy of the results.

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(2)

(iii) Suggest why $\Delta H_{\text{reaction}}$ is difficult to determine directly by experiment.

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(1)
(Total 9 marks)

45. This question is about ammonia, NH3, which is produced as shown in the following equation.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

(a) Use oxidation numbers to explain why this is a redox reaction.

....

....

....

(2)

(b) (i) Use the average (mean) bond enthalpy data to calculate a value for the enthalpy change for this reaction. You are reminded to show **all** your working.

Bond	Average bond enthalpy / kJ mol^{-1}
N≡N	944
H—H	436
N—H	388

(3)

(ii) The actual standard enthalpy change for this reaction is –92 kJ mol^{-1}. Explain why the value you calculated in (b)(i) is not the same as this.

....

....

(1)

(iii) At room temperature, a mixture of nitrogen and hydrogen is thermodynamically unstable with respect to ammonia, but is kinetically stable.

Use the data in (b)(i) and (ii) to help you explain why this mixture is

thermodynamically unstable

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kinetically stable

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(3)

(c) The manufacturer of ammonia would like to achieve a high rate of reaction and a high equilibrium yield of product.

(i) State and explain, in terms of collision theory, TWO ways to increase the rate of the reaction. An increase in pressure does **not** alter the rate in this process.

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(6)

(ii) State and explain TWO ways to increase the equilibrium yield of ammonia.

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(4)
(Total 19 marks)

46. The enthalpy change for the reaction between aqueous sodium hydroxide solution and aqueous hydrochloric acid was determined by the following method:

- Aqueous hydrochloric acid was titrated against 25.0 cm^3 of 1.50 mol dm^{-3} aqueous sodium hydroxide solution using a suitable indicator. The mean (or average) titre was 22.75 cm^3.
- 25.0 cm^3 of the sodium hydroxide solution was carefully measured into a polystyrene cup and 22.75 cm^3 of the hydrochloric acid was transferred to a clean dry beaker. Both solutions were allowed to stand for five minutes before their temperatures were noted.
- The hydrochloric acid was then added to the sodium hydroxide solution, the mixture stirred thoroughly and the highest temperature noted.
- The experiment was repeated three times giving an average temperature change of +10.5°C.

(a) (i) Calculate the heat produced in the reaction, in joules.

Use the approximations that the density of the final solution is 1.00 g cm^{-3} and its specific heat capacity is 4.18 J g^{-1} K^{-1}.

(2)

(ii) Calculate the enthalpy change for the reaction, in kJ mol^{-1}.

(3)

(b) State ONE assumption made when calculating this enthalpy change, other than those stated in (a)(i).

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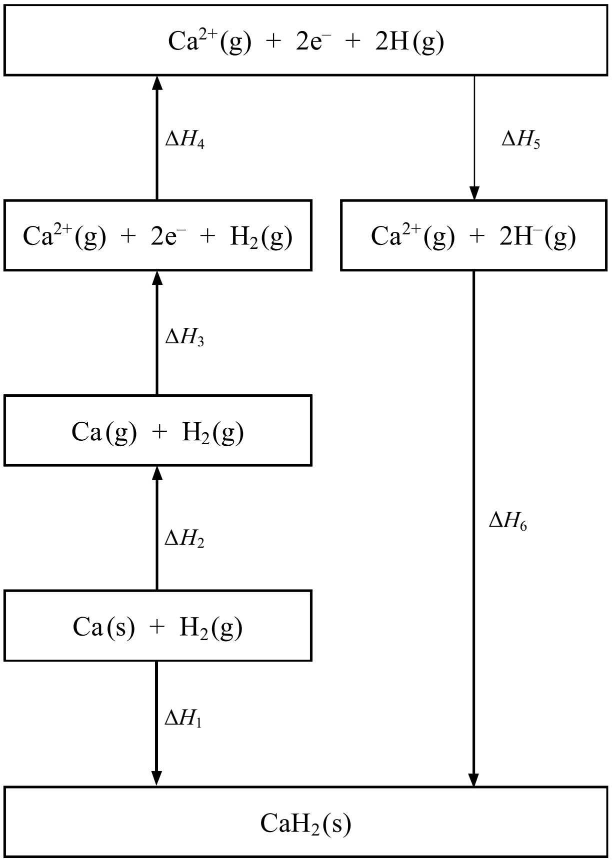
(1)

(Total 6 marks)

47. The Born-Haber cycle below represents the enthalpy changes when calcium hydride, CaH_2, is formed from its elements.



(a) Write down in terms of **one** of the symbols ΔH_1 to ΔH_6

(i) the lattice energy of calcium hydride ....
(1)

(ii) the first electron affinity of hydrogen ....
(1)

(b) Use the data below to calculate the standard enthalpy of formation of calcium hydride, $CaH_2(s)$.

	value /kJ mol^{-1}
enthalpy of atomisation of calcium	+178
first plus second ionisation energies of calcium	+1735
enthalpy of atomisation of hydrogen	+218
first electron affinity of hydrogen	–73
lattice energy of calcium hydride	–2389

Calculation:

(2)

(c) Explain why the lattice energy of magnesium hydride, $MgH_2(s)$, is more exothermic than the lattice energy of calcium hydride, $CaH_2(s)$.

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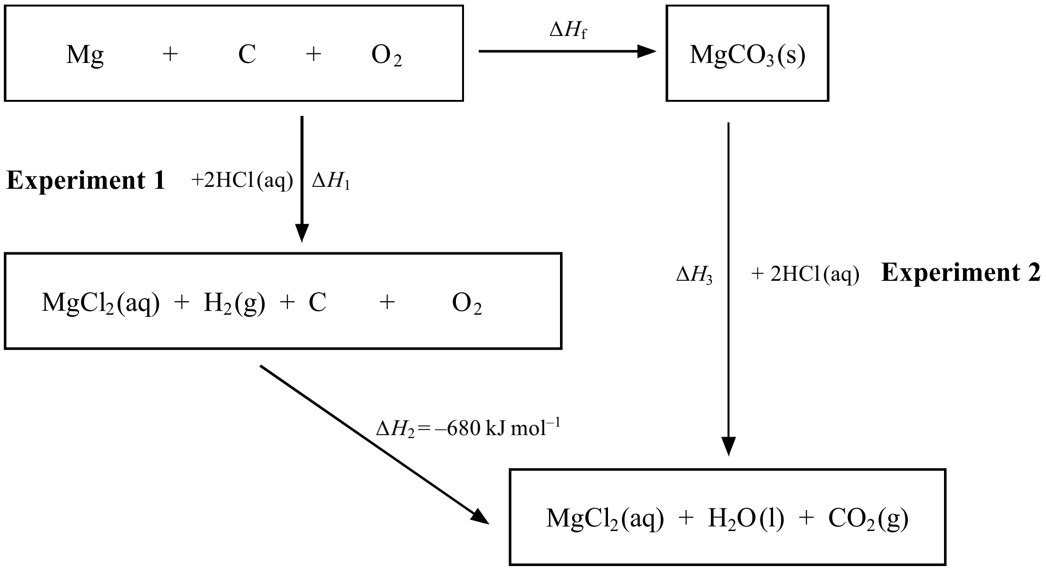
(3)
(Total 7 marks)

48. Two experiments were carried out in order to calculate the enthalpy change of formation of magnesium carbonate, $MgCO_3$.

A Hess cycle for these reactions is shown below.



(a) Complete the Hess cycle above for the formation of magnesium carbonate from its elements by balancing the equations and adding state symbols.

(2)

(b) In **Experiment 1** the temperature of 100 cm^3 of hydrochloric acid was measured. After one minute, 0.100 g of magnesium was added to the excess acid and the temperature measured every minute. The following results were obtained:

Time / min	0	1	2	3	4	5	6
Temp / °C	21.0	21.0	25.3	25.1	24.9	24.8	24.7

(i) How many moles of magnesium were used in this experiment?

Use the Periodic Table as a source of data.

....

(1)

(ii) The initial concentration of the hydrochloric acid was 2.00 mol dm^{-3}.

Calculate the number of moles of hydrochloric acid at the start and hence the number remaining at the end of the experiment.

(3)

(iii) Plot the graph of temperature against time.

(2)

(iv) Calculate the energy change in this experiment assuming the temperature rise is 4.5 °C. Use the expression

Energy change (J) = 4.2 × mass of solution × temperature change

[Assume that 1 cm^3 of solution has a mass of 1 g]

(1)

(v) Use your answer to (iv) to calculate ΔH_1 for one mole of magnesium reacting with hydrochloric acid. Include a sign and units in your answer.

(2)

(vi) Suggest why a temperature rise of 4.5 °C was used in the calculation in (iv).

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(1)

(c) 2.2 g of magnesium carbonate was added to 100 cm^3 of the same acid in **Experiment 2**.

The temperature changed from 21.0 °C to 23.5 °C resulting in an energy change of 1.05 kJ.

(i) Calculate the mass of one mole of magnesium carbonate, $MgCO_3$ and hence the number of moles of magnesium carbonate used in this experiment.
Use the Periodic Table as a source of data.

(2)

(ii) Using the method in part (b)(v), calculate ΔH_3.

(1)

(d) Using your answers to (b)(v) and (c)(ii), calculate the enthalpy change of formation, ΔH_f, of magnesium carbonate, $MgCO_3$.
Include a sign and units in your answer.

(2)

(e) Why is it impossible to measure ΔH_f of $MgCO_3(s)$ directly?

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(1)
(Total 18 marks)

49. Calculate the standard enthalpy change of formation of gaseous silicon tetrachloride, $\Delta H^\ominus_f [SiCl_4(g)]$.

Your answer should include a sign and units.

Use the Hess cycle below and the following data at 298 K.

$\Delta H^{o}_{at} [Si(s)]$ = +455.6 kJ mol^{-1}

$\Delta H^{o}_{at} [½Cl_2]$ = +121.7 kJ mol^{-1}

Bond energy, E (Si-Cl) = +407.4 kJ mol^{-1}

$$\begin{tikzcd}
SiCl_4(g) \arrow[rr, "\Delta H^\ominus"] & & Si(g) + 4Cl(g) \\
& Si(s) + 2Cl_2(g) \arrow[ul, "\Delta H^\ominus_f [SiCl_4(g)]"] \arrow[ur] &
\end{tikzcd}$$

(Total 3 marks)

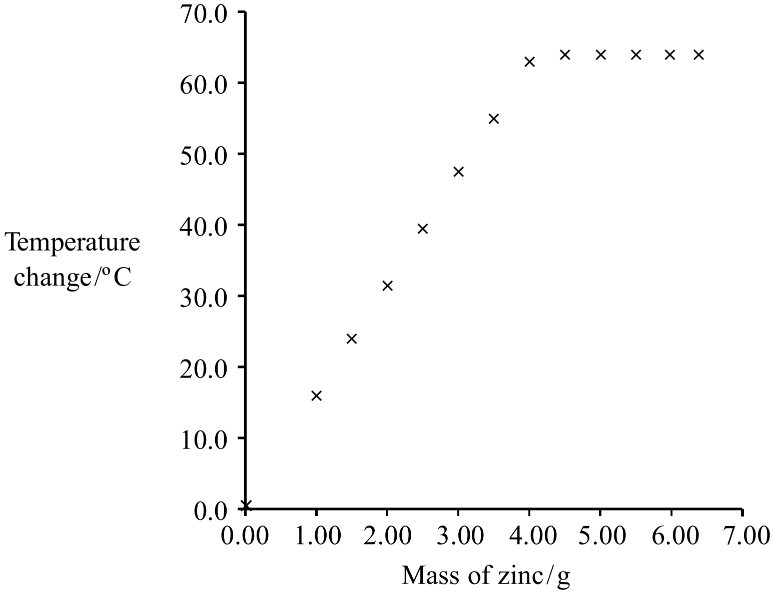
50. An experiment was carried out to measure the enthalpy change for the reaction of zinc with aqueous copper(II) sulphate.

The equation for the reaction is

$$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$$

- A measuring cylinder was used to transfer separate 50 cm3 samples of 1.25 mol dm–3 copper(II) sulphate solution into polystyrene cups.
- Weighed amounts of zinc powder were added to each sample in turn.
- Each mixture was stirred thoroughly and the temperature rise noted with a thermometer accurate to 0.5 °C.

The results of this experiment are summarised on the graph below.



(a) Explain why the graph initially shows a rise in temperature and then levels off.

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(2)

(b) (i) Suggest why the mass of metal is **not** used in the calculation of the heat change.

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(1)

(ii) The graph shows that the maximum temperature change is 63.5 °C. Use this value to calculate the maximum heat change, in joules, in this reaction.

You should assume that the density of the solution is 1.00 g cm^{-3} and its heat capacity is the same as water, 4.18 J g^{-1} $°C^{-1}$.

(1)

(iii) From the heat change calculated in (b)(ii) calculate the enthalpy change, in kJ mol^{-1}, for the reaction. Include the appropriate sign and give your answer to **three** significant figures.

(4)

(c) (i) It is suggested that the precision of the experiment would be improved by using a thermometer accurate to 0.1 °C.

Explain why this suggestion is incorrect.

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....

(1)

(ii) Suggest a simple practical change to the **method** that would make the experiment more accurate.

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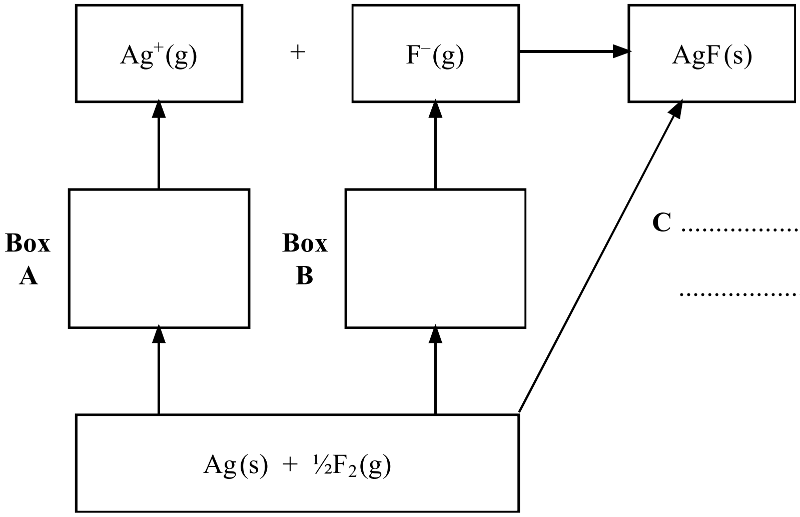
(1)

(Total 10 marks)

51. (a) The following data were collected to use in a Born-Haber cycle for silver fluoride, AgF.

	Value /kJ mol^{-1}
enthalpy of atomisation of silver	+285
first ionisation energy of silver	+731
enthalpy of atomisation of fluorine	+79
enthalpy of formation of silver fluoride	–205
lattice energy of silver fluoride	–958

On the following outline of a Born-Haber cycle, complete boxes **A** and **B** by adding the formula **and** state symbol for the appropriate species. Write the name of the enthalpy change at **C**.



C ....

....

(3)

(b) ΔH_{latt} (theoretical) is the lattice energy calculated assuming the crystal lattice is completely ionic.
ΔH_{latt} (experimental) is the lattice energy determined experimentally using the Born-Haber cycle.

Values for the silver halides are listed below.

Formula of halide	ΔH_{latt} (theoretical) / kJ mol^{-1}	ΔH_{latt} (experimental) / kJ mol^{-1}	ΔH_{latt} (theoretical) minus ΔH_{latt} (experimental) / kJ mol^{-1}
AgF	–920	–958	38
AgCl	–833	–905	72
AgBr	–816	–891	75
AgI	–778	–889	111

(i) Explain why the **theoretical** lattice energies become less exothermic from AgF to AgI.

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(3)

(ii) Explain why the values of the theoretical and experimental lattice energies are different.

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(2)

(iii) Explain why the difference between the theoretical and experimental lattice energies increases from AgF to AgI.

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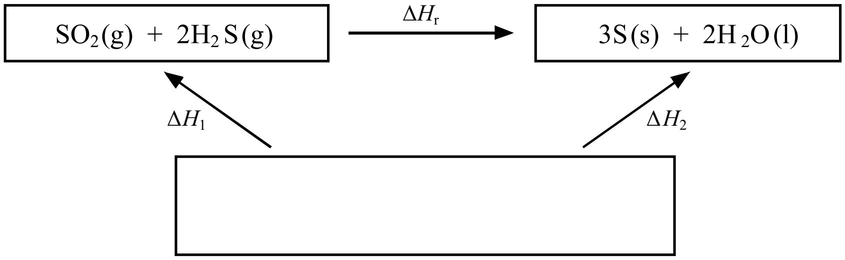
....

(2)

(Total 10 marks)

52. The Hess cycle below can be used to find the enthalpy change, ΔH_r, for the reaction between hydrogen sulphide and sulphur dioxide, using standard enthalpy changes of formation.



(i) Complete the cycle by filling in the empty box.

(2)

(ii) What is meant by the **standard enthalpy change of formation**, $\Delta H_f^\ominus$, of a compound?

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(3)

(iii) Use the cycle and the data below to calculate the enthalpy change of the reaction, ΔH_r.

	$\Delta H_f^\ominus$ / kJ mol^{-1}
SO_2 (g)	–296.8
H_2S (g)	–20.6
H_2O (l)	–285.8

(2)
(Total 7 marks)

53. Which of the equations shown below represents the reaction for which ΔH is the standard enthalpy change of formation, $\Delta H^\ominus_{f\,298}$, for ethanol, C_2H_5OH. Ethanol melts at 156 K and boils at 352 K.

A $2C(g) + 6H(g) + O(g) \rightarrow C_2H_5OH(g)$

B $2C(s) + 3H_2(g) + O_2(g) \rightarrow C_2H_5OH(l)$

C $2C(s) + 3H_2(g) + O(g) \rightarrow C_2H_5OH(g)$

D $2C(s) + 3H_2(g) + \frac{1}{2}O_2(g) \rightarrow C_2H_5OH(l)$

(Total 1 mark)

54. Airbags, used as safety features in cars, contain sodium azide, NaN3. An airbag requires a large volume of gas to be produced in a few milliseconds. The gas is produced in this reaction:

$$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g) \qquad \Delta H \text{ is positive}$$

When the airbag is fully inflated, 50 dm3 of nitrogen gas is produced.

(a) Calculate the number of molecules in 50 dm^3 of nitrogen gas under these conditions.

[The Avogadro constant = 6.02×10^{23} mol^{-1}. The molar volume of nitrogen gas under the conditions in the airbag is 24 dm^3 mol^{-1}].

(2)

(b) Calculate the mass of sodium azide, NaN_3, that would produce 50 dm^3 of nitrogen gas.

(3)

(c) What will happen to the temperature in the airbag when the reaction occurs?

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(1)

(d) The airbag must be strong enough not to burst in an accident. An airbag which has burst in an accident is hazardous if the sodium azide in it has decomposed.

Explain why this is so.

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(2)
(Total 8 marks)

55. A student investigated a reaction which could be used to warm up coffee in self-heating cans.

$$Mg(s) + Cu(NO_3)_2(aq) \rightarrow Mg(NO_3)_2(aq) + Cu(s)$$

In the self-heating cans, the bottom has a compartment containing copper(II) nitrate solution. When a button on the bottom of the can is pressed, the magnesium powder is released into the compartment where it reacts with the copper(II) nitrate solution.

(a) A student investigated the enthalpy change for this reaction by measuring

50.0 cm^3 of 0.300 mol dm^{-3} copper(II) nitrate solution into a 100 cm3 beaker and adding 1g (an excess) of magnesium powder.

The results are shown below.

Temperature of copper(II) nitrate solution at start = 22 °C
Temperature of mixture after reaction = 43 °C

(i) Calculate the energy change which took place. The specific heat capacity of the solution is 4.20 J g^{-1} K^{-1}.

Which is the correct value for the energy change in joules?

(1)

(ii) How many moles of copper(II) nitrate were used in the experiment?

(1)

(iii) Calculate the enthalpy change for the reaction. You should include a sign and units in your answer.

(2)

(iv) Suggest two changes you would make to the equipment used in order to improve the accuracy of the result.

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(2)

(b) The ionic equation for the reaction is shown below:

$Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s) \qquad \Delta H = -532 \text{ kJ mol}^{-1}$

Would the following affect the value of the experimental result?

Explain your answer, stating the effect, if any, on the value of the enthalpy change obtained.

(i) The student used 2 g rather than 1g of magnesium.

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(2)

(ii) The heat losses that occurred from the student's beaker.

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(2)

(c) The temperature in the self-heating can needs to increase by 60 °C to produce a hot drink.

Suggest a change you could make to the mixture in the experiment in (a) to produce a greater temperature rise. You are **not** expected to do a calculation.

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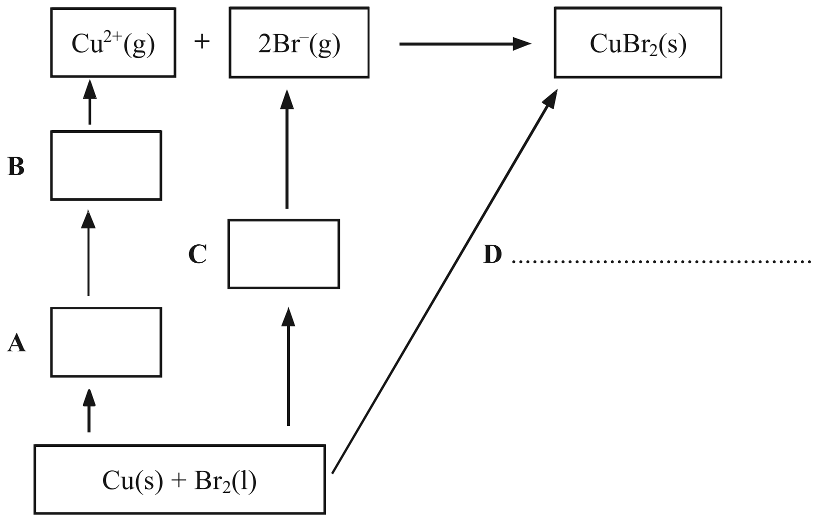
(1)
(Total 11 marks)

56. The following data can be used in a Born-Haber cycle for copper(II) bromide, $CuBr_2$.

Enthalpy change of atomisation of bromine $\Delta H^{\ominus}_{at}[\frac{1}{2}Br_2(l)]$	+111.9 kJ mol^{-1}
Enthalpy change of atomisation of copper, $\Delta H^{\ominus}_{at}[Cu(s)]$	+338.3 kJ mol^{-1}
First ionisation energy of copper, $E_{m1}[Cu(g)]$	+746.0 kJ mol^{-1}
Second ionisation energy of copper, E_{m2} [Cu(g)]	+1958.0 kJ mol^{-1}
Electron affinity of bromine, $E_{aff}[Br(g)]$	–342.6 kJ mol^{-1}
Enthalpy change of formation of $CuBr_2(s)$, $\Delta H^{\ominus}_f$ [$CuBr_2(s)$]	–141.8 kJ mol^{-1}

(a) On the following outline of a Born-Haber cycle complete the boxes **A**, **B**, and **C** by putting in the formula and state symbol for the appropriate species and writing the name of the enthalpy change **D**.



(3)

(b) Use the data to calculate a value for the lattice energy of copper(II) bromide.

Give a sign and units in your answer.

(3)

(c) When the lattice energy of copper(II) bromide is calculated from ionic radii and charges, the result is a value numerically about 10% less than the one obtained from the Born-Haber cycle.

(i) What does this suggest about the nature of the bonding in copper(II) bromide?

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....

(1)

(ii) Draw a diagram to show how the smaller copper ion alters the shape of the larger bromide ion.

(1)

(Total 8 marks)