AQA A-LEVEL CHEMISTRY · 7405

# CHEMISTRY



## EXAM QUESTION PACK

### 3.1.4 ENERGETICS

| SCORE | PERCENTAGE |
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## Q1.

This question is about energetics.

(a) Write an equation, including state symbols, for the reaction with an enthalpy change equal to the enthalpy of formation for iron(III) oxide.

_____

**(1)**

(b) **Table 1** contains some standard enthalpy of formation data.

**Table 1**

| | $CO(g)$ | $Fe_2O_3(s)$ |
|---|---|---|
| **$\Delta_f H^\ominus$ / kJ mol$^{-1}$** | −111 | −822 |

$Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(s) + 3CO_2(g)$ $\Delta H = -19$ kJ mol$^{-1}$

Use these data and the equation for the reaction of iron(III) oxide with carbon monoxide to calculate a value for the standard enthalpy of formation for carbon dioxide.

Show your working.

$\Delta_f H^\ominus$ _____ kJ mol$^{-1}$

**(3)**

(c) Some enthalpy data are given in **Table 2**.

**Table 2**

| **Process** | **$\Delta H$ / kJ mol$^{-1}$** |
|---|---|
| $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$ | −92 |
| $N_2(g) \longrightarrow 2N(g)$ | +944 |
| $H_2(g) \longrightarrow 2H(g)$ | +436 |



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Use the data from **Table 2** to calculate the bond enthalpy for N−H in ammonia.

N−H bond enthalpy _____ kJ $mol^{-1}$

**(3)**

(d) Give one reason why the bond enthalpy that you calculated in part (c) is different from the mean bond enthalpy quoted in a data book (388 kJ $mol^{-1}$).

_____

_____

_____

**(1)**

**(Total 8 marks)**

## Q2.

This question is about enthalpy changes.

(a) Write an equation, including state symbols, to show the reaction taking place when the standard enthalpy of combustion for ethanol is measured.

_____

**(2)**

(b) State the name given to the enthalpy change represented by the following chemical equation.

Explain why this enthalpy change would be difficult to determine directly.

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

Enthalpy change _____

Explanation _____

_____

_____



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**(2)**

(c) Standard enthalpies of combustion for carbon and carbon monoxide are −393 kJ $mol^{-1}$ and −283 kJ $mol^{-1}$, respectively.

Use these data to calculate the enthalpy change for the reaction in part **(b)**.

Enthalpy change = _____ kJ $mol^{-1}$

**(2)**

(d) Use the following data to calculate a value for the Xe–F bond enthalpy in $XeF_4$

$$Xe(g) + 2F_2(g) \rightarrow XeF_4(g) \quad \Delta H = -252 \text{ kJ mol}^{-1}$$
$$F_2(g) \rightarrow 2F(g) \quad \Delta H = +158 \text{ kJ mol}^{-1}$$

Xe–F bond enthalpy = _____ kJ $mol^{-1}$

**(3)**

(e) Suggest a reason why the value calculated in part **(d)** differs from the mean Xe–F bond enthalpy quoted in a data source.

_____

_____



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_____

**(1)**

**(Total 10 marks)**

## Q3.

Standard enthalpy of combustion data can be used to calculate enthalpies of formation.

(a) State the meaning of the term standard enthalpy of combustion.

_____

_____

_____

_____

**(3)**

(b) The equation corresponding to the enthalpy of formation of propan-1-ol is shown.

$$3C(s) + 4H_2(g) + \frac{1}{2}O_2(g) \longrightarrow CH_3CH_2CH_2OH(l)$$

**Table 1** contains some standard enthalpy of combustion data.

**Table 1**

| | **C(s)** | **$H_2(g)$** | **$CH_3CH_2CH_2OH(l)$** |
|---|---|---|---|
| $\Delta H_c^{\ominus}$ / kJ mol$^{-1}$ | –394 | –286 | –2010 |

Use data from **Table 1** to calculate a value for the standard enthalpy of formation of propan-1-ol. Show your working.

_____

_____

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**(3)**

(c) An equation for the complete combustion of gaseous propan-1-ol is shown.



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$$CH_3CH_2CH_2OH(g) + 4\tfrac{1}{2}O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(g) \quad \Delta H = -1893 \text{ kJ mol}^{-1}$$

**Table 2** shows some bond enthalpy data.

**Table 2**

| | C–H | C–O | O–H | C=O | O=O |
|---|---|---|---|---|---|
| Bond enthalpy / kJ mol$^{-1}$ | 412 | 360 | 463 | 805 | 496 |

Use data from **Table 2** and the enthalpy change for this reaction to calculate a value for the bond enthalpy of a C–C bond in propan-1-ol.

_____

_____

_____

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_____

**(3)**

**(Total 9 marks)**

## Q4.

(a) Propanone can be formed when glucose comes into contact with bacteria in the absence of air.

(i) Balance the following equation for this reaction of glucose to form propanone, carbon dioxide and water.

.....$C_6H_{12}O_6$ $\longrightarrow$ .....$CH_3COCH_3$ + .....$CO_2$ + .....$H_2O$

**(1)**

(ii) Deduce the role of the bacteria in this reaction.

_____

**(1)**

(b) Propanone is also formed by the oxidation of propan−2−ol.

(i) Write an equation for this reaction using [O] to represent the oxidising agent.

_____

**(1)**

(ii) State the class of alcohols to which propan−2−ol belongs.



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**(1)**

(c) A student determined a value for the enthalpy change when a sample of propanone was burned. The heat produced was used to warm some water in a copper calorimeter.
The student found that the temperature of 150 g of water increased by 8.0 °C when $4.50 \times 10^{-3}$ mol of pure propanone was burned in air.

Use the student's results to calculate a value, in kJ $mol^{-1}$, for the enthalpy change when one mole of propanone is burned.
(The specific heat capacity of water is 4.18 J $K^{-1}$ $g^{-1}$)

**(3)**

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

**(3)**

(e) Use the mean bond enthalpy data in the table and the equation given below the table to calculate a value for the standard enthalpy change when gaseous propanone is burned.

| | C−H | C−C | C−O | O−H | C=O | O=O |
|---|---|---|---|---|---|---|



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| | | | | | | |
|---|---|---|---|---|---|---|
| **Mean bond enthalpy / kJ mol$^{-1}$** | 412 | 348 | 360 | 463 | 805 | 496 |

$$CH_3COCH_3(g) + 4O_2(g) \longrightarrow 3CO_2(g) + 3H_2O(g)$$

_____

_____

_____

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**(3)**

(f) Suggest **two** reasons why the value obtained by the student in part (c) is different from the value calculated in part (e).

Reason 1 _____

_____

_____

Reason 2 _____

_____

_____

**(2)**

**(Total 15 marks)**

## Q5.

Ethanol is an important industrial compound.

(a) Ethanol can be produced by the hydration of ethene.
The equation for the equilibrium that is established is

$$H_2C{=}CH_2(g) + H_2O(g) \rightleftharpoons CH_3CH_2OH(g) \quad \Delta H = -42 \text{ kJ mol}^{-1}$$

The operating conditions for the process are a temperature of 300 °C and a pressure of 7 MPa.
Under these conditions, the conversion of ethene into ethanol is 5%.

(i) Identify the catalyst used in this process.
Deduce how an overall yield of 95% is achieved in this process without changing the operating conditions.

_____



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_____

_____

_____

**(2)**

(ii) Use your knowledge of equilibrium reactions to explain why a manufacturer might consider using an excess of steam in this process, under the same operating conditions.

_____

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_____

**(3)**

(iii) At pressures higher than 7 MPa, some of the ethene reacts to form a solid with a relative molecular mass greater than 5000.

Deduce the identity of this solid.

Give **one** other reason for **not** operating this process at pressures higher than 7 MPa.
Do **not** include safety reasons.

_____

_____

_____

**(2)**

(b) Write an equation for the reaction that has an enthalpy change that is the standard enthalpy of formation of ethanol.

_____

**(2)**

(c) When ethanol is used as a fuel, it undergoes combustion.

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

_____

_____

_____



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_____

_____

_____

**(3)**

(ii) Consider these bond enthalpy data.

| | C–H | C–C | C–O | O=O | C=O | O–H |
|---|---|---|---|---|---|---|
| **Bond enthalpy / kJ mol$^{-1}$** | 412 | 348 | 360 | 496 | 805 | 463 |

Use these data and the equation to calculate a value for the enthalpy of combustion of gaseous ethanol.

$$CH_3CH_2OH(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(g)$$

_____

_____

_____

_____

_____

_____

_____

**(3)**

(d) Gaseous ethanol can be used to convert hot copper(II) oxide into copper.

(i) Deduce the role of ethanol in this reaction.

_____

**(1)**

(ii) Draw the structure of the organic compound with $M_r$ = 60 that is produced in this reaction.

**(1)**

**(Total 17 marks)**

## Q6.

A student calculated that a value for the enthalpy change of neutralisation is –51.2 kJ mol$^{-1}$.

The design of a possible hand-warmer using hydrochloric acid and sodium hydroxide was discussed. It was proposed that 500 cm$^3$ of hydrochloric acid should be used in a flexible, sealed plastic container with a breakable tube of solid sodium hydroxide also in the container. On breaking the tube, the sodium hydroxide would be released, react with the



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acid and produce heat.
A 40 °C temperature rise was thought to be suitable.

(a) Calculate the heat energy, in J, required to raise the temperature of the reaction mixture by 40 °C. Assume that the reaction mixture has a density of 1.00 g $cm^{-3}$ and a specific heat capacity of 4.18 J $K^{-1}$ $g^{-1}$.
Assume that all of the heat energy given out is used to heat the reaction mixture.

_____

_____

_____

_____

**(2)**

(b) Use your answer from part (a) and the value for the enthalpy change of neutralisation of –51.2 kJ $mol^{-1}$ to calculate the minimum amount, in moles, and hence the minimum mass of sodium hydroxide required in the breakable tube.
(If you could not complete the calculation in part (a) assume that the heat energy required was 77 400 J. This is **not** the correct answer).

Show your working.

Moles of NaOH _____

_____

Mass of NaOH _____

_____

**(3)**

(c) Use the amount, in moles, of sodium hydroxide from part (b) to calculate the minimum concentration, in mol $dm^{-3}$, of hydrochloric acid required in the 500 $cm^3$ of solution used in the sealed container.

_____

_____

_____

**(1)**

(d) Suggest **one** possible risk to a person who uses a hand-warmer containing sodium hydroxide and hydrochloric acid.

_____

_____

**(1)**

(e) A commercial hand-warmer uses powdered iron sealed in a plastic container.
A valve allows air to enter the container, and oxygen in the air reacts slowly with the iron to form solid iron(III) oxide. The heat released warms the container.



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(i) Write an equation for this reaction between iron and oxygen to form iron(lll) oxide.

_____

**(1)**

(ii) One version of an iron-oxygen hand-warmer advertises that it is designed to stay warm for up to four hours.
Other than by increasing the amount of iron in the container, state **one** change to the iron in the hand-warmer that would increase this time.
Explain why this change to the iron might **not** be an advantage.

Change to the iron _____

_____

Explanation _____

_____

_____

**(3)**

(f) Another type of hand-warmer uses sodium thiosulfate. Sodium thiosulfate is very soluble in water at 80 °C but is much less soluble at room temperature.
When a hot, concentrated solution of sodium thiosulfate is cooled it does not immediately crystallise. The sodium thiosulfate stays dissolved as a stable 'super-saturated' solution until crystallisation is triggered.
Heat energy is then released when the sodium thiosulfate crystallises.

(i) This type of hand-warmer is re-usable.
Suggest **one** environmental advantage that a sodium thiosulfate hand-warmer has over the other two types.

_____

**(1)**

(ii) Describe the **two** steps that you would take to make the sodium thiosulfate hand-warmer ready for re-use.

Step 1 _____

_____

Step 2 _____

_____

**(2)**

**(Total 14 marks)**

## Q7.

(a) Iron is extracted from iron(III) oxide using carbon at a high temperature.

(i) State the type of reaction that iron(III) oxide undergoes in this extraction.

_____



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**(1)**

(ii) Write a half-equation for the reaction of the iron(III) ions in this extraction.

_____

**(1)**

(b) At a high temperature, carbon undergoes combustion when it reacts with oxygen.

(i) Suggest why it is **not** possible to measure the enthalpy change directly for the following combustion reaction.

$$C(s,graphite) + \frac{1}{2}O_2(g) \longrightarrow CO(g)$$

_____

_____

**(1)**

(ii) State Hess's Law.

_____

_____

_____

**(1)**

(iii) State the meaning of the term *standard enthalpy of combustion.*

_____

_____

_____

_____

_____

**(3)**

(c) Use the standard enthalpies of formation in the table below and the equation to calculate a value for the standard enthalpy change for the extraction of iron using carbon monoxide.

| | $Fe_2O_3(s)$ | $CO(g)$ | $Fe(l)$ | $CO_2(g)$ |
|---|---|---|---|---|
| $\Delta H_f$/ kJ mol$^{-1}$ | - 822 | - 111 | +14 | - 394 |

$$Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(l) + 3CO_2(g)$$



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_____

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_____

**(3)**

(d) (i) Write an equation for the reaction that represents the standard enthalpy of formation of carbon dioxide.

_____

**(1)**

(ii) State why the value quoted in part (c) for the standard enthalpy of formation of $CO_2(g)$ is the same as the value for the standard enthalpy of combustion of carbon.

_____

_____

**(1)**

**(Total 12 marks)**

## Q8.

This question is about the extraction of titanium from titanium(IV) oxide by a two-stage process.
The first stage in the process produces titanium(IV) chloride. In the second stage, titanium(IV) chloride is converted into titanium.
The enthalpy change for the second stage can be determined using Hess's Law.

(a) Give **one** reason why titanium is **not** extracted directly from titanium(IV) oxide using carbon.

_____

_____

**(1)**

(b) Give the meaning of the term *enthalpy change*.

_____

_____

_____

**(1)**

(c) State Hess's Law.



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**(1)**

(d) Define the term *standard enthalpy of formation*.

**(3)**

(e) The following standard enthalpy of formation data refer to the second stage in the extraction of titanium.

| | $TiCl_4(g)$ | Na(l) | NaCl(s) | Ti(s) |
|---|---|---|---|---|
| $\Delta H_f^{\ominus}$/ kJ mol$^{-1}$ | –720 | +3 | –411 | 0 |

(i) State why the value for the standard enthalpy of formation of Na(l) is **not** zero.

**(1)**

(ii) Use data from the table to calculate a value for the standard enthalpy change of the following reaction.

$$TiCl_4(g) + 4Na(l) \longrightarrow 4NaCl(s) + Ti(s)$$



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_____

**(3)**

(iii) State the role of sodium in this reaction.

_____

**(1)**

**(Total 11 marks)**

## Q9.

Comparison of lattice enthalpies from Born-Haber cycles with lattice enthalpies from calculations based on a perfect ionic model are used to provide information about bonding in crystals.

(a) Define the terms *enthalpy of atomisation* and *lattice dissociation enthalpy*.

Enthalpy of atomisation _____

_____

_____

Lattice dissociation enthalpy _____

_____

_____

**(4)**

(b) Use the following data to calculate a value for the lattice dissociation enthalpy of sodium chloride.

| | $\Delta H^{\ominus}$/kJ mol$^{-1}$ |
|---|---|
| $Na(s) \longrightarrow Na(g)$ | +109 |
| $Na(g) \longrightarrow Na^+(g) + e^-$ | +494 |
| $Cl_2(g) \longrightarrow 2Cl(g)$ | +242 |
| $Cl(g) + e^- \longrightarrow Cl^-(g)$ | −364 |
| $Na(s) + \frac{1}{2}Cl_2(g) \longrightarrow NaCl(s)$ | −411 |

_____

_____

_____

_____

_____

_____

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_____

**(3)**

(c) Consider the following lattice dissociation enthalpy ($\Delta H_L^\circ$) data.

| | NaBr | AgBr |
|---|---|---|
| $\Delta H_L^\circ$(experimental)/kJ mol$^{-1}$ | +733 | +890 |
| $\Delta H_L^\circ$(theoretical)/kJ mol$^{-1}$ | +732 | +758 |

The values of $\Delta H_L^\circ$ (experimental) have been determined from Born–Haber cycles.

The values of $\Delta H_L^\circ$ (theoretical) have been determined by calculation using a perfect ionic model.

(i) Explain the meaning of the term *perfect ionic model*.

_____

_____

_____

_____

**(2)**

(ii) State what you can deduce about the bonding in NaBr from the data in the table.

_____

_____

**(1)**

(iii) State what you can deduce about the bonding in AgBr from the data in the table.

_____

_____

**(1)**

**(Total 11 marks)**

## Q10.

A scientist used mass spectrometry to analyse a sample of the air near a fertiliser factory. The sample of air included traces of a gas which was shown by its molecular ion to have a precise $M_r$ = 44.00105

(a) State the meaning of the term *molecular ion*.

_____

_____

**(1)**



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(b) (i) Use the following data to show that the trace gas was dinitrogen oxide ($N_2O$).

Show your working.

| Atom | Precise relative atomic mass |
|---|---|
| $^{12}C$ | 12.00000 |
| $^{14}N$ | 14.00307 |
| $^{16}O$ | 15.99491 |

_____

_____

**(1)**

(ii) Propane is used as a fuel in the fertiliser factory. State why both propane and its combustion product, carbon dioxide, might have been identified as the trace gas if the scientist had used relative molecular masses calculated to one decimal place.

_____

_____

**(1)**

(iii) State why the precise relative atomic mass for the $^{12}C$ isotope is exactly 12.00000

_____

**(1)**

(c) Dinitrogen oxide is formed when ammonia is oxidised according to the following equation.

$$2NH_3(g) + 2O_2(g) \rightarrow N_2O(g) + 3H_2O(l)$$

(i) Use the standard enthalpies of formation in the table below to calculate a value for the standard enthalpy change of this reaction.

| | $NH_3(g)$ | $O_2(g)$ | $N_2O(g)$ | $H_2O(l)$ |
|---|---|---|---|---|
| $\Delta H_f^o$/ kJ mol$^{-1}$ | –46 | 0 | +82 | –286 |

_____

_____

_____



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**(3)**

(ii) State **one** condition necessary for enthalpies of formation to be quoted as standard values at a specified temperature of 298 K.

_____

**(1)**

**(Total 8 marks)**

## Q11.

Hess's Law is used to calculate the enthalpy change in reactions for which it is difficult to determine a value experimentally.

(a) State the meaning of the term *enthalpy change*.

_____

**(1)**

(b) State Hess's Law.

_____

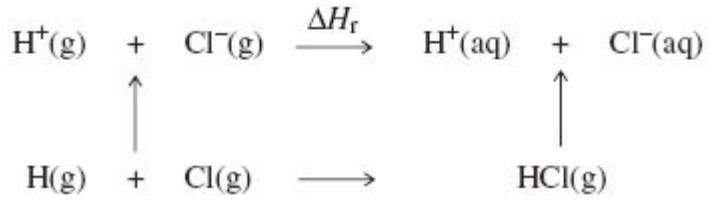
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**(1)**

(c) Consider the following table of data and the scheme of reactions.

| Reaction | Enthalpy change / kJ mol$^{-1}$ |
|---|---|
| $HCl(g) \rightarrow H^+(aq) + Cl^-(aq)$ | –75 |
| $H(g) + Cl(g) \rightarrow HCl(g)$ | –432 |
| $H(g) + Cl(g) \rightarrow H^+(g) + Cl^-(g)$ | +963 |




Use the data in the table, the scheme of reactions and Hess's Law to calculate a value for $\Delta H_r$

_____

_____

_____

_____



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_____

**(3)**

**(Total 5 marks)**

## Q12.

Barium can be extracted from barium oxide (BaO) in a process using aluminium.
A mixture of powdered barium oxide and powdered aluminium is heated strongly.
The equation for this extraction process is shown below.

$$3BaO(s) + 2Al(s) \rightarrow 3Ba(s) + Al_2O_3(s)$$

Some standard enthalpies of formation are given in the table below.

| Substance | BaO(s) | $Al_2O_3(s)$ |
|---|---|---|
| $\Delta H_f^{\⦵}$/kJ mol$^{-1}$ | –558 | –1669 |

(a) (i) State what is meant by the term *standard enthalpy of formation*.

_____

_____

_____

_____

_____

**(3)**

(ii) State why the standard enthalpy of formation of barium and that of aluminium are both zero.

_____

**(1)**

(iii) Use the data to calculate the standard enthalpy change for the reaction shown by the equation above.

_____

_____

_____

_____

_____

**(3)**

(b) (i) Suggest the major reason why this method of extracting barium is expensive.

_____

**(1)**



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(ii) Using barium oxide and aluminium powders increases the surface area of the reactants. Suggest **one** reason why this increases the rate of reaction.

_____

**(1)**

(c) (i) Write an equation for the reaction of barium with water.

_____

**(1)**

(ii) A solution containing barium ions can be used to test for the presence of sulfate ions in an aqueous solution of sodium sulfate.

Write the **simplest ionic** equation for the reaction which occurs and state what is observed.

Simplest ionic equation

_____

Observation _____

**(2)**

(iii) State how barium sulfate can be used in medicine. Explain why this use is possible, given that solutions containing barium ions are poisonous.

Use _____

Explanation _____

_____

**(2)**

**(Total 14 marks)**

## Q13.

The combustion of hydrocarbons is an important source of energy.

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

_____

_____

_____

**(3)**

(b) (i) Write an equation for the complete combustion of ethane, $C_2H_6$.

_____

(ii) Use the standard enthalpies of formation given below to calculate the standard enthalpy of combustion of ethane.

| Formula and state of compound | $C_2H_6(g)$ | $CO_2(g)$ | $H_2O(l)$ |
|---|---|---|---|



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| | | | |
|---|---|---|---|
| Standard enthalpy of formation (at 298 K)/kJ $mol^{-1}$ | –85 | –394 | –286 |

**(4)**

(c) A container and its contents of total heat capacity 120 J $K^{-1}$ were heated using a methane burner. Calculate the maximum theoretical temperature rise when 0.10 g of methane was completely burned. The standard enthalpy of combustion of methane is –890 kJ $mol^{-1}$.

**(4)**

**(Total 11 marks)**

## Q14.

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

**(3)**

(b) Use the data in the table to calculate the standard enthalpy of formation of liquid methylbenzene, $C_7H_8$

| Substance | C(s) | $H_2(g)$ | $C_7H_8(l)$ |
|---|---|---|---|
| Standard enthalpy of combustion, $\Delta H_c^{\ominus}$ /kJ $mol^{-1}$ | –394 | –286 | –3909 |

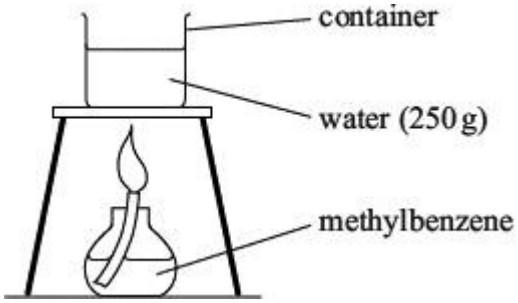
$7C(s) + 4H_2(g) \rightarrow C_7H_8(l)$

**(3)**

(c) An experiment was carried out to determine a value for the enthalpy of combustion of liquid methylbenzene using the apparatus shown in the diagram.



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Burning 2.5 g of methylbenzene caused the temperature of 250 g of water to rise by 60°C. Use this information to calculate a value for the enthalpy of combustion of methylbenzene, $C_7H_8$

(The specific heat capacity of water is 4.18 J $K^{-1}$ $g^{-1}$. Ignore the heat capacity of the container.)

**(4)**

(d) A 25.0 $cm^3$ sample of 2.00 mol $dm^{-3}$ hydrochloric acid was mixed with 50.0 $cm^3$ of a 1.00 mol $dm^{-3}$ solution of sodium hydroxide. Both solutions were initially at 18.0 °C.

After mixing, the temperature of the final solution was 26.5°C.

Use this information to calculate a value for the standard enthalpy change for the following reaction.

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

In your calculation, assume that the density of the final solution is 1.00 g $cm^{-3}$ and that its specific heat capacity is the same as that of water. (Ignore the heat capacity of the container.)

**(4)**

(e) Give **one** reason why your answer to part (d) has a much smaller experimental error than your answer to part (c).

**(1)**

**(Total 15 marks)**



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## Q15.

(a) Explain the meaning of the terms *mean bond enthalpy* and *standard enthalpy of formation*.

*Mean bond enthalpy* _____

_____

_____

*Standard enthalpy of formation* _____

_____

_____

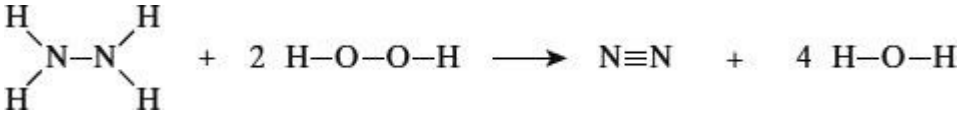
_____

**(5)**

(b) Some mean bond enthalpies are given below.

| Bond | N–H | N–N | N≡N | H–O | O–O |
|---|---|---|---|---|---|
| Mean bond enthalpy/kJ mol$^{-1}$ | 388 | 163 | 944 | 463 | 146 |

Use these data to calculate the enthalpy change for the following gas-phase reaction between hydrazine, $N_2H_4$, and hydrogen peroxide, $H_2O_2$



_____

_____

_____

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_____

**(3)**

(c) Some standard enthalpies of formation are given below.

| | $N_2H_4(g)$ | $H_2O_2(g)$ | $H_2O(g)$ |
|---|---|---|---|
| $\Delta H_f^\circ$ /kJ mol$^{-1}$ | +75 | –133 | –242 |

These data can be used to calculate the enthalpy change for the reaction in part (b).

$$N_2H_4(g) + 2H_2O_2(g) \rightarrow N_2(g) + 4H_2O(g)$$

(i) State the value of $\Delta H_f^\ominus$ for $N_2(g)$.



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(ii) Use the $\Delta H_f^{\ominus}$ values from the table to calculate the enthalpy change for this reaction.

**(4)**

(d) Explain why the value obtained in part (b) is different from that obtained in part (c) (ii).

**(1)**

**(Total 13 marks)**

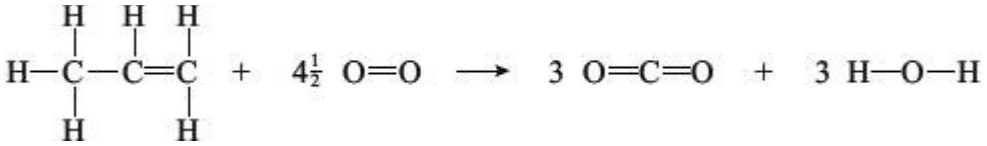
## Q16.

(a) Define the term *standard enthalpy of combustion*, $\Delta H_c$⦵

**(3)**

(b) Use the mean bond enthalpy data from the table and the equation given below to calculate a value for the standard enthalpy of combustion of propene. All substances are in the gaseous state.

| Bond | C == C | C—C | C—H | O == O | O == C | O—H |
|---|---|---|---|---|---|---|
| Mean bond enthalpy/ kJ mol$^{-1}$ | 612 | 348 | 412 | 496 | 743 | 463 |





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**(3)**

(c) State why the standard enthalpy of formation, $\Delta H_f^\circ$, of oxygen is zero.

_____

**(1)**

(d) Use the data from the table below to calculate a more accurate value for the standard enthalpy of combustion of propene.

| Compound | $C_3H_6(g)$ | $CO_2(g)$ | $H_2O(g)$ |
|---|---|---|---|
| Standard enthalpy of formation, $\Delta H_f^\circ$/ kJ mol$^{-1}$ | +20 | –394 | –242 |

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**(3)**

(e) Explain why your answer to part (b) is a less accurate value than your answer to part (d).

_____

_____

**(2)**

**(Total 12 marks)**

## Q17.

When 0.10 g of propane was burned the quantity of heat evolved was 5.0 kJ. The enthalpy of combustion of propane in kJ mol$^{-1}$ is

**A** −800

**B** −1500

**C** −2200

**D** −2900

**(Total 1 mark)**

## Q18.

(a) The table below contains some mean bond enthalpy data.

| Bond | H–O | O–O | O=O |
|---|---|---|---|



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| | | | |
|---|---|---|---|
| Mean bond enthalpy/kJ mol$^{-1}$ | 463 | 146 | 496 |

The bonding in hydrogen peroxide, $H_2O_2$, can be represented by H–O–O–H. Use these data to calculate the enthalpy change for the following reaction.

$$H_2O_2(g) \rightarrow H_2O_2(g) + \tfrac{1}{2} O_2(g)$$

**(3)**

(b) The standard enthalpy of formation, $\Delta H_f^{\ominus}$ for methane, is –74.9 kJ mol$^{-1}$. Write an equation, including state symbols, for the reaction to which this enthalpy change applies.

**(2)**

(c) The enthalpy changes for the formation of atomic hydrogen and atomic carbon from their respective elements in their standard states are as follows.

$$\tfrac{1}{2} H_2(g) \rightarrow H(g) \qquad \Delta H^{\ominus} = +218 \text{ kJ mol}^{-1}$$

$$C(s) \rightarrow C(g) \qquad \Delta H^{\ominus} = +715 \text{ kJ mol}^{-1}$$

(i) By reference to its structure, suggest why a large amount of heat energy is required to produce free carbon atoms from solid carbon.

(ii) Parts (b) and (c) give enthalpy data for the formation of $CH_4(g)$, H(g) and C(g). Use these data and Hess's Law to calculate the value of the enthalpy change for the following reaction.

$$CH_4(g) \rightarrow C(g) + 4H(g)$$

(iii) Use your answer from part (c)(ii) to calculate a value for the mean bond



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enthalpy of a C–H bond in methane.

_____

**(5)**

**(Total 10 marks)**

## Q19.

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

_____

_____

_____

_____

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_____

**(3)**

(b) State Hess's Law and use it, together with the data given in the table below, to calculate the standard enthalpy change for the following reaction.

$MgO(s) + 2HCl(g) \rightarrow MgCl_2(s) + H_2O(l)$

| | MgO | HCl(g) | $MgCl_2$ | $H_2O$ |
|---|---|---|---|---|
| $\Delta H_f^{\ominus}$/kJ mol$^{-1}$ | –602 | –92 | –642 | –286 |

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**(4)**

(c) In an experiment, an excess of solid magnesium oxide was added to 50 cm$^3$ of 3.0 mol dm$^{-3}$ hydrochloric acid. The initial temperature of the solution was 21 °C. After reaction, the temperature had risen to 53 °C. (The specific heat capacity of water is 4.2 J K$^{-1}$ g$^{-1}$)



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Use this information to calculate the enthalpy change for the reaction of one mole of magnesium oxide with hydrochloric acid. For your calculation you should assume that all the heat from the reaction is used to raise the temperature of 50 g of water.

**(8)**

**(Total 15 marks)**

## Q20.

The table below contains some standard enthalpy of formation data.

| Substance | C(s) | $N_2(g)$ | $H_2O(g)$ | $CO_2(g)$ | $NH_4NO_3(s)$ |
|---|---|---|---|---|---|
| $\Delta H_f^{\ominus}$ / kJ mol$^{-1}$ | 0 | 0 | –242 | –394 | –365 |

(a) Why are the values of the standard enthalpy of formation for carbon and nitrogen zero?

**(1)**

(b) State Hess's Law.



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_____

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**(2)**

(c) Use $\Delta H_f^{\ominus}$ data from the table to calculate a value for the enthalpy change for the following reaction.

$$NH_4NO_3(s) + \frac{1}{2}C(s) \rightarrow N_2(g) + 2H_2O(g) + \frac{1}{2}CO_2(g)$$

_____

_____

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_____

**(3)**

**(Total 6 marks)**

## Q21.

(a) What is the meaning of the term *enthalpy change*?

_____

_____

_____

**(2)**

(b) (i) Define the term *standard enthalpy of formation* of a compound.

_____

_____

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_____

(ii) Write an equation, including state symbols, for the formation from its elements of solid sodium sulphate, $Na_2SO_4$

_____

**(5)**

(c) State Hess's Law.

_____

_____

**(1)**

(d) Some standard enthalpy changes are difficult to measure directly but can be



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determined from standard enthalpies of combustion.
Maleic acid, $C_4H_4O_4$, reacts with oxygen to form carbon dioxide and water as shown by the following equation.

$$C_4H_4O_4(s) + 3O_2(g) \rightarrow 4CO_2(g) + 2H_2O(l)$$

Use the standard enthalpy of combustion data given below to calculate a value for the standard enthalpy change for the following reaction.

$$4C(s) + 2H_2(g) + 2O_2(g) \rightarrow C_4H_4O_4(s)$$

| | $C_4H_4O_4(s)$ | C(s) | $H_2(g)$ |
|---|---|---|---|
| $\Delta H_c^{\ominus}$ / kJ mol$^{-1}$ | –1356 | –393.5 | –285.8 |

**(3)**
**(Total 11 marks)**

## Q22.

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

**(3)**

(b) State Hess's law.

**(1)**

(c) Propanone, $CH_3COCH_3$, burns in oxygen as shown by the equation



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$CH_3COCH_3(l) + 4\ O_2(g) \rightarrow 3H_2O(l) + 3CO_2(g)$

Use the data given below to calculate the standard enthalpy of combustion of propanone.

| | $CO_2(g)$ | $H_2O(l)$ | $CH_3COCH_3(l)$ |
|---|---|---|---|
| $\Delta H_f^{\ominus}$/kJ mol$^{-1}$ | –394 | –286 | –248 |

_____

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**(3)**

**(Total 7 marks)**

## Q23.

Using the data below, which is the correct value for the standard enthalpy of formation for $TiCl_4(l)$?

$C(s) + TiO_2(s) + 2Cl_2(g) \rightarrow TiCl_4(l) + CO_2(g)$ $\quad \Delta H^{\ominus} = -232$ kJ mol$^{-1}$

$Ti(s) + O_2(g) \rightarrow TiO_2(s)$ $\quad \Delta H_f^{\ominus} = -912$ kJ mol$^{-1}$

$C(s) + O_2(g) \rightarrow CO_2(g)$ $\quad \Delta H_f^{\ominus} = -394$ kJ mol$^{-1}$

**A** −1538 kJ mol$^{-1}$

**B** −1094 kJ mol$^{-1}$

**C** −750 kJ mol$^{-1}$

**D** +286 kJ mol$^{-1}$

**(Total 1 mark)**

## Q24.

When ethanamide ($CH_3CONH_2$) burns in oxygen the carbon is converted into carbon dioxide, the hydrogen is converted into water and the nitrogen forms nitrogen gas.

| Substance | ethanamide | carbon dioxide | water |
|---|---|---|---|
| Enthalpy of formation ($\Delta H_f^{\ominus}$) / kJ mol$^{-1}$ | −320 | −394 | −286 |

Using the data above, which one of the following is a correct value for the enthalpy of combustion of ethanamide?



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**A** $-1823\ kJ\ mol^{-1}$

**B** $-1183\ kJ\ mol^{-1}$

**C** $-1000\ kJ\ mol^{-1}$

**D** $-360\ kJ\ mo1^{-1}$

**(Total 1 mark)**

## Q25.

In which one of the following reactions is the standard enthalpy change equal to the standard enthalpy of formation of lithium fluoride?

**A** $Li(g) + F(g) \rightarrow LiF(s)$

**B** $Li^+(g) + F^-(g) \rightarrow LiF(s)$

**C** $Li^+(aq) + F^-(g) \rightarrow LiF(s)$

**D** $Li(s) + \frac{1}{2}F_2(g) \rightarrow LiF(s)$

**(Total 1 mark)**



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## Q1.

This question is about enthalpy changes.

(a) When ethanoic acid reacts with sodium hydroxide, the enthalpy change, $\Delta H$, is –56.1 kJ $mol^{-1}$

$$CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$$

Calculate the temperature rise when 25 $cm^3$ of 2.0 mol $dm^{-3}$ aqueous ethanoic acid react with 25 $cm^3$ of 2.0 mol $dm^{-3}$ aqueous sodium hydroxide.

Assume that both solutions have the same initial temperature, have a density of 1.0 g $cm^{-3}$ and a specific heat capacity of 4.18 J $K^{-1}$ $g^{-1}$

Temperature rise _____ °C

**(4)**

(b) A student recorded the temperature of aqueous ethanoic acid in a polystyrene cup for three minutes.

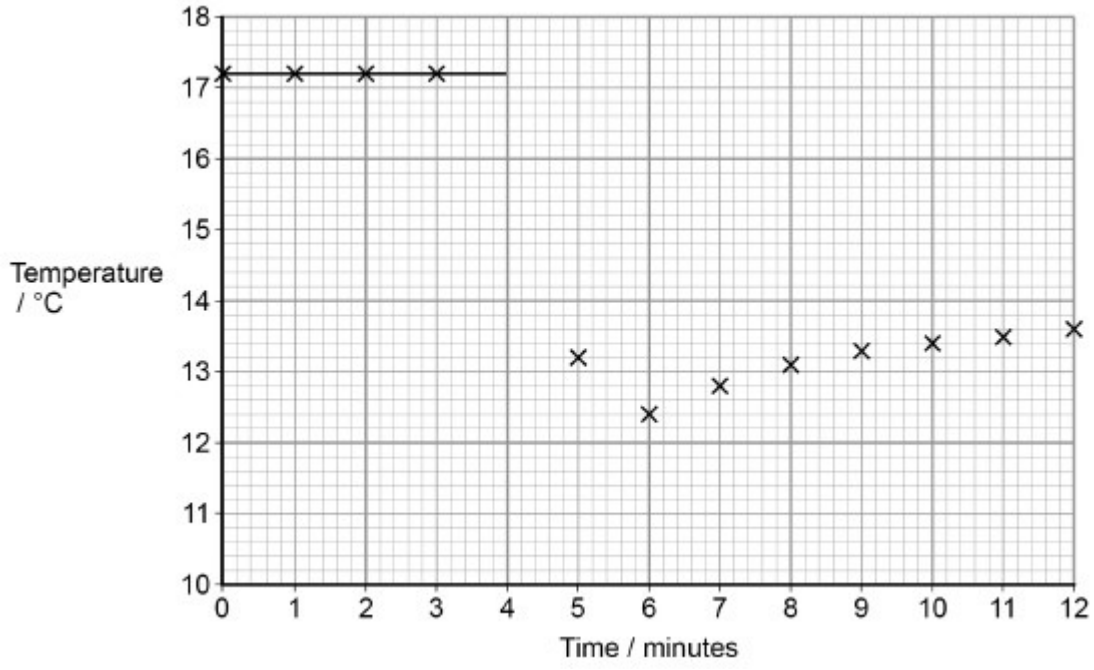
At the fourth minute, the student added sodium hydrogencarbonate.

The student stirred the mixture and carried on recording the temperature every minute for several minutes.

The student's measurements are shown in the graph.

A best-fit line showing the temperature before mixing has been drawn.

Draw an appropriate best-fit line on the graph and use it to find the temperature change at the time of mixing.




Temperature change at time of mixing _____ °C

**(2)**



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**(Total 6 marks)**

## Q2.

The oxidation of propan-1-ol can form propanal and propanoic acid.
The boiling points of these compounds are shown in the table.

| Compound | Boiling point / °C |
|---|---|
| propan-1-ol | 97 |
| propanal | 49 |
| propanoic acid | 141 |

In a preparation of propanal, propan-1-ol is added dropwise to the oxidising agent and the aldehyde is separated from the reaction mixture by distillation.

(a) Explain, with reference to intermolecular forces, why distillation allows propanal to be separated from the other organic compounds in this reaction mixture.

**(3)**

(b) Give **two** ways of maximising the yield of propanal obtained by distillation of the reaction mixture.

1. _____



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_____

_____

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2. _____

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_____

**(2)**

(c) Describe how you would carry out a simple test-tube reaction to confirm that the sample of propanal obtained by distillation does **not** contain any propanoic acid.

_____

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**(2)**

(d) A student carried out an experiment to determine the enthalpy of combustion of ethanol.
Combustion of 457 mg of ethanol increased the temperature of 150 g of water from 25.1 °C to 40.2 °C

Calculate a value, in kJ $mol^{-1}$, for the enthalpy of combustion of ethanol in this experiment.
Give your answer to the appropriate number of significant figures.

(The specific heat capacity of water is 4.18 J $K^{-1}$ $g^{-1}$)

Enthalpy of combustion _____ kJ $mol^{-1}$



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**(3)**

(e) A mixture of isomeric alkenes is produced when pentan-2-ol is dehydrated in the presence of hot concentrated sulfuric acid. Pent-1-ene is one of the isomers produced.

Name and outline a mechanism for the reaction producing pent-1-ene.

Name of mechanism _____

Mechanism

**(4)**

(f) A pair of stereoisomers is also formed in the reaction in **part (e)**.

Name the less polar stereoisomer formed.
Explain how this type of stereoisomerism arises.

Name _____

Explanation _____

_____

_____

_____

**(2)**

**(Total 16 marks)**

## Q3.

A student carried out an experiment to find the temperature rise for a reaction between hydrochloric acid and sodium hydroxide solution.

- The student used a measuring cylinder to place 50 $cm^3$ of 0.400 mol $dm^{-3}$ hydrochloric acid into a glass beaker.
- The student recorded the temperature at one-minute intervals for three minutes.
- At the fourth minute the student added 50 $cm^3$ of 0.400 mol $dm^{-3}$ sodium hydroxide solution and stirred to mix the solutions, but did not record the temperature.
- The student recorded the temperature at one-minute intervals for a further eight minutes.

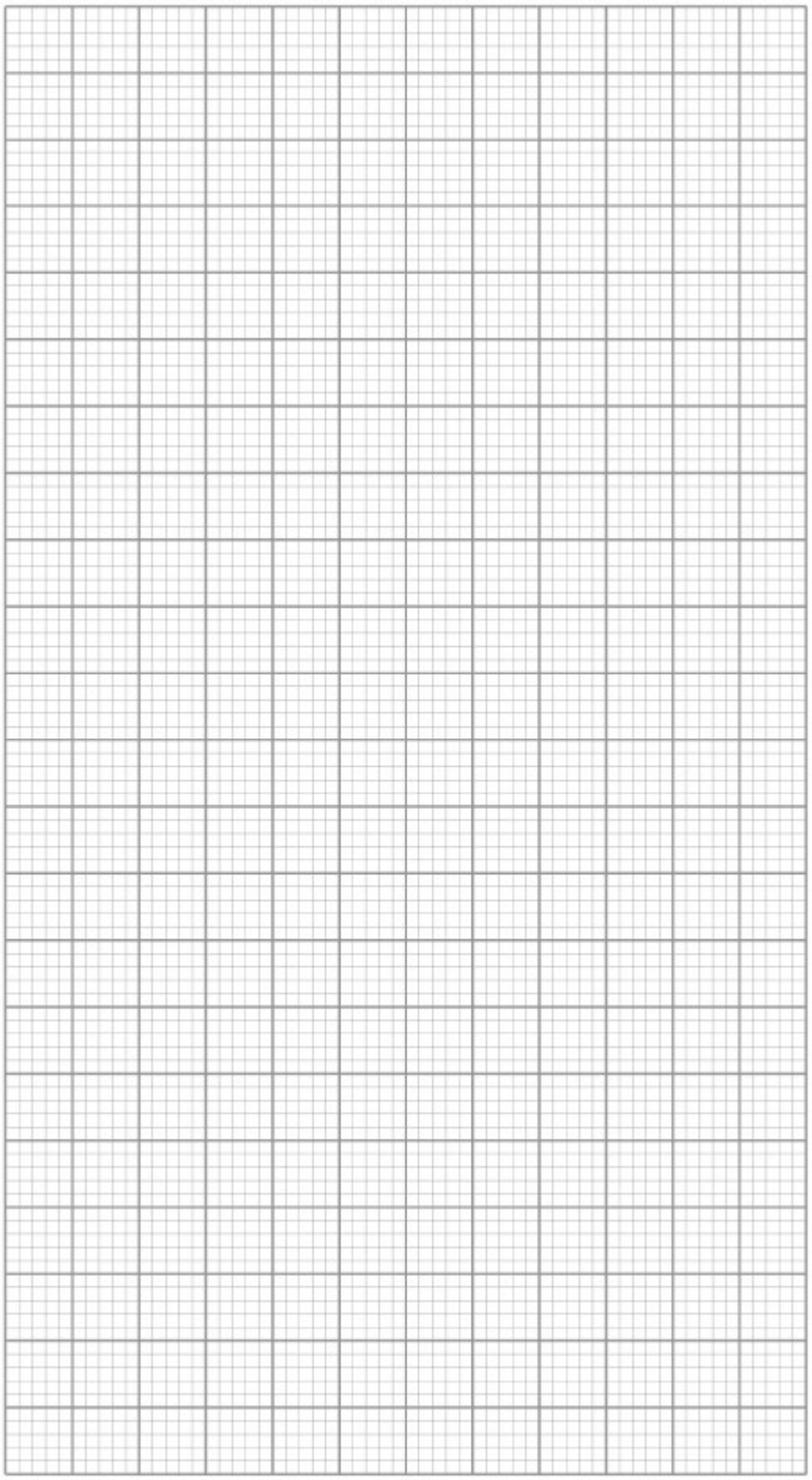
The results are shown in the table.



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| Time / min | 0 | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | 12 |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| Temperature / °C | 19.8 | 19.8 | 19.8 | 19.8 | | 21.4 | 21.7 | 21.6 | 21.5 | 21.4 | 21.3 | 21.2 | 21.1 |

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



Use your graph to find the temperature rise, $\Delta T$, at the fourth minute.



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Show your working on the graph by drawing suitable lines of best fit.

$\Delta T$ _____ °C

**(5)**

(b) The uncertainty in each of the temperature readings from the thermometer used in this experiment was ±0.1° C

Calculate the percentage uncertainty in the value for the temperature rise.

Percentage uncertainty _____

**(1)**

(c) Suggest a change to the experiment that would minimise heat loss.

_____

_____

**(1)**

(d) Suggest and explain another change to the experiment that would decrease the percentage uncertainty in the use of the same thermometer.

_____

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_____

**(2)**

(e) A second student completed an experiment to determine the enthalpy of neutralisation for the reaction between ethanedioic acid solution (HOOCCOOH) and potassium hydroxide solution.

The student added 25 $cm^3$ of 0.80 mol $dm^{-3}$ ethanedioic acid solution to 75 $cm^3$ of 0.60 mol $dm^{-3}$ potassium hydroxide solution.
The temperature increased by 3.2 °C

Give an equation for the reaction between ethanedioic acid solution and potassium hydroxide solution.



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Calculate the enthalpy change ($\Delta H$) per mole of water formed in this reaction.
Assume that the specific heat capacity of the reaction mixture is 4.2 J $K^{-1}$ $g^{-1}$
Assume that the density of the reaction mixture is 1.00 g $cm^{-3}$
Equation _____

$\Delta H$ _____ kJ $mol^{-1}$

**(5)**

(f) In a similar experiment to that in **part (e)**, the enthalpy of neutralisation for the reaction between sulfuric acid and potassium hydroxide solution was found to be –57.0 kJ $mol^{-1}$ per mole of water formed.

Suggest an explanation for the difference between this value and your answer to **part (e)**.

(If you were unable to obtain an answer to **part (e)** you should assume a value of –28.5 kJ $mol^{-1}$. This is **not** the correct answer.)

_____

_____

_____

_____

_____

_____

**(2)**

**(Total 16 marks)**

## Q4.

A student planned and carried out an experiment to determine the enthalpy of reaction when magnesium metal displaces zinc from aqueous zinc sulfate.

$$Mg(s) + Zn^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Zn(s)$$

The student used this method:



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- A measuring cylinder was used to transfer 50 $cm^3$ of a 1.00 mol $dm^{-3}$ aqueous solution of zinc sulfate into a glass beaker.
- A thermometer was placed in the beaker.
- 2.08 g of magnesium metal powder were added to the beaker.
- The mixture was stirred and the maximum temperature recorded.

The student recorded a starting temperature of 23.9 °C and a maximum temperature of 61.2 °C.

(a) Show by calculation which reactant was in excess.

Use the data to calculate the experimental value for enthalpy of reaction in kJ $mol^{-1}$(Assume that the specific heat capacity of the solution is 4.18 J $K^{-1}g^{-1}$and the density of the solution is 1.00 g $cm^{-3}$).

Reactant in excess _____

Enthalpy of reaction _____ kJ $mol^{-1}$

**(6)**

(b) Another student used the same method and obtained a value for the enthalpy of reaction of −142 kJ $mol^{-1}$

A data book value for the enthalpy of reaction is −310 kJ $mol^{-1}$

Suggest the most likely reason for the large difference between the student's experimental value and the data book value.

_____

_____



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_____

**(1)**

(c) Suggest how the students' method, and the analysis of the results, could be improved in order to determine a more accurate value for the enthalpy of reaction.

Justify your suggestions.

Do not refer to the precision of the measuring equipment. Do not change the amounts or the concentration of the chemicals.

_____

_____

_____

_____

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_____

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_____

**(6)**

**(Total 13 marks)**

## Q5.

Anhydrous magnesium chloride, $MgCl_2$, can absorb water to form the hydrated salt $MgCl_2.4H_2O$

$$MgCl_2(s) + 4H_2O(l) \longrightarrow MgCl_2.4H_2O(s)$$

(a) Suggest **one** reason why the enthalpy change for this reaction cannot be determined directly by calorimetry.

_____

_____

**(1)**

(b) Some enthalpies of solution are shown in Table 1.

**Table 1**

| Salt | Enthalpy of solution / kJ mol$^{-1}$ |
|---|---|
| $MgCl_2(s)$ | −155 |
| $MgCl_2.4H_2O(s)$ | −39 |

Calculate the enthalpy change for the absorption of water by $MgCl_2(s)$ to form $MgCl_2.4H_2O(s)$.



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Enthalpy change _____ kJ mol$^{-1}$

**(2)**

(c) Describe how you would carry out an experiment to determine the enthalpy of solution of anhydrous magnesium chloride.

You should use about 0.8 g of anhydrous magnesium chloride.

Explain how your results could be used to calculate the enthalpy of solution.

_____

_____

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_____

**(6)**

(d) Anhydrous magnesium chloride can be formed by direct reaction between its elements.

$$Mg(s) + Cl_2(g) \longrightarrow MgCl_2(s)$$

The free-energy change, $\Delta G$, for this reaction varies with temperature as shown in **Table 2**.

**Table 2**

| T / K | $\Delta G$ / kJ mol$^{-1}$ |
|---|---|
| 298 | −592.5 |
| 288 | −594.2 |
| 273 | −596.7 |
| 260 | −598.8 |
| 240 | −602.2 |

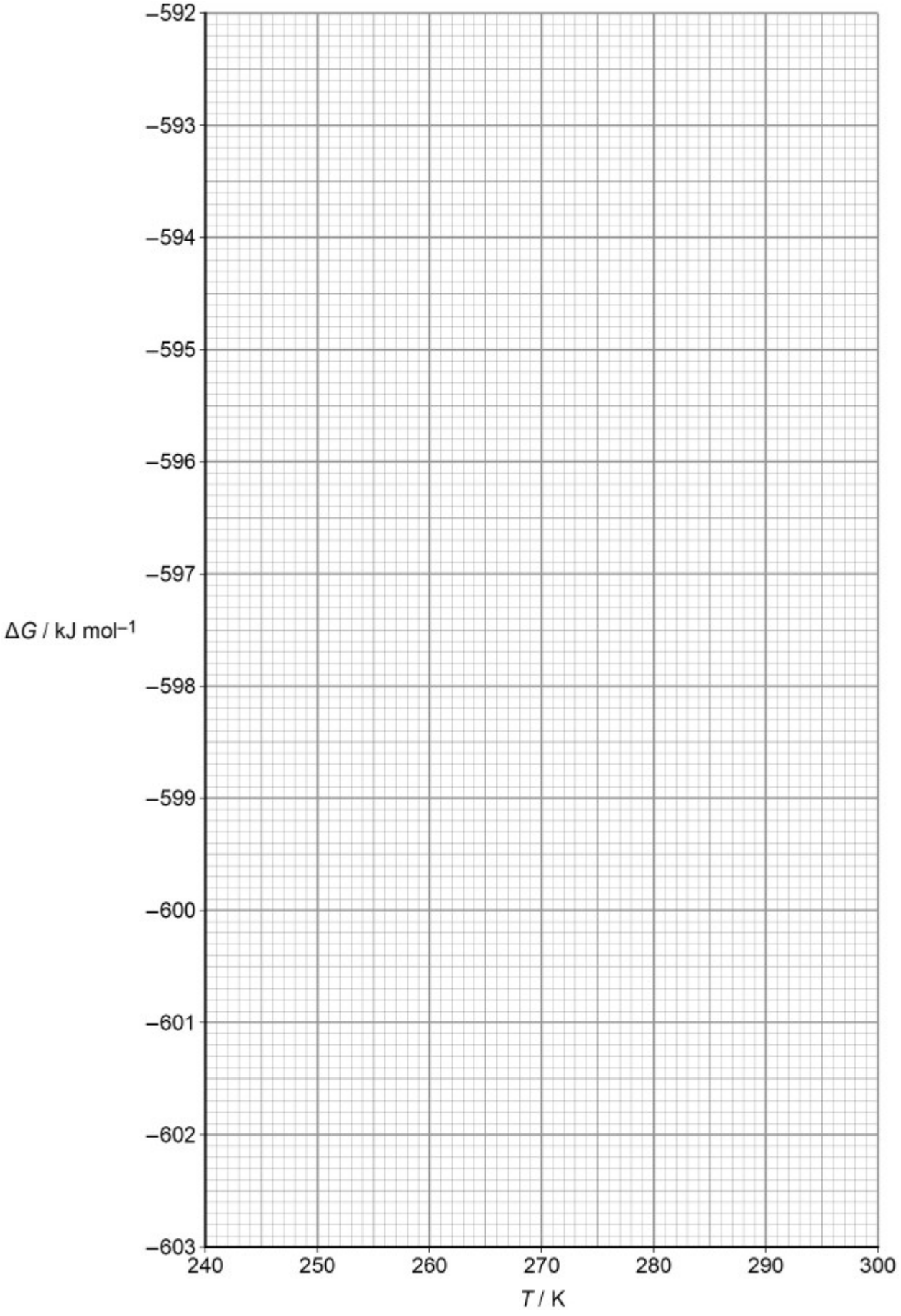


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Use these data to plot a graph of free-energy change against temperature on the grid below.

Calculate the gradient of the line on your graph and hence calculate the entropy change, $\Delta S$, in $J K^{-1} mol^{-1}$, for the formation of anhydrous magnesium chloride from its elements.

Show your working.




ΔS _____ $JK^{-1}mol^{-1}$



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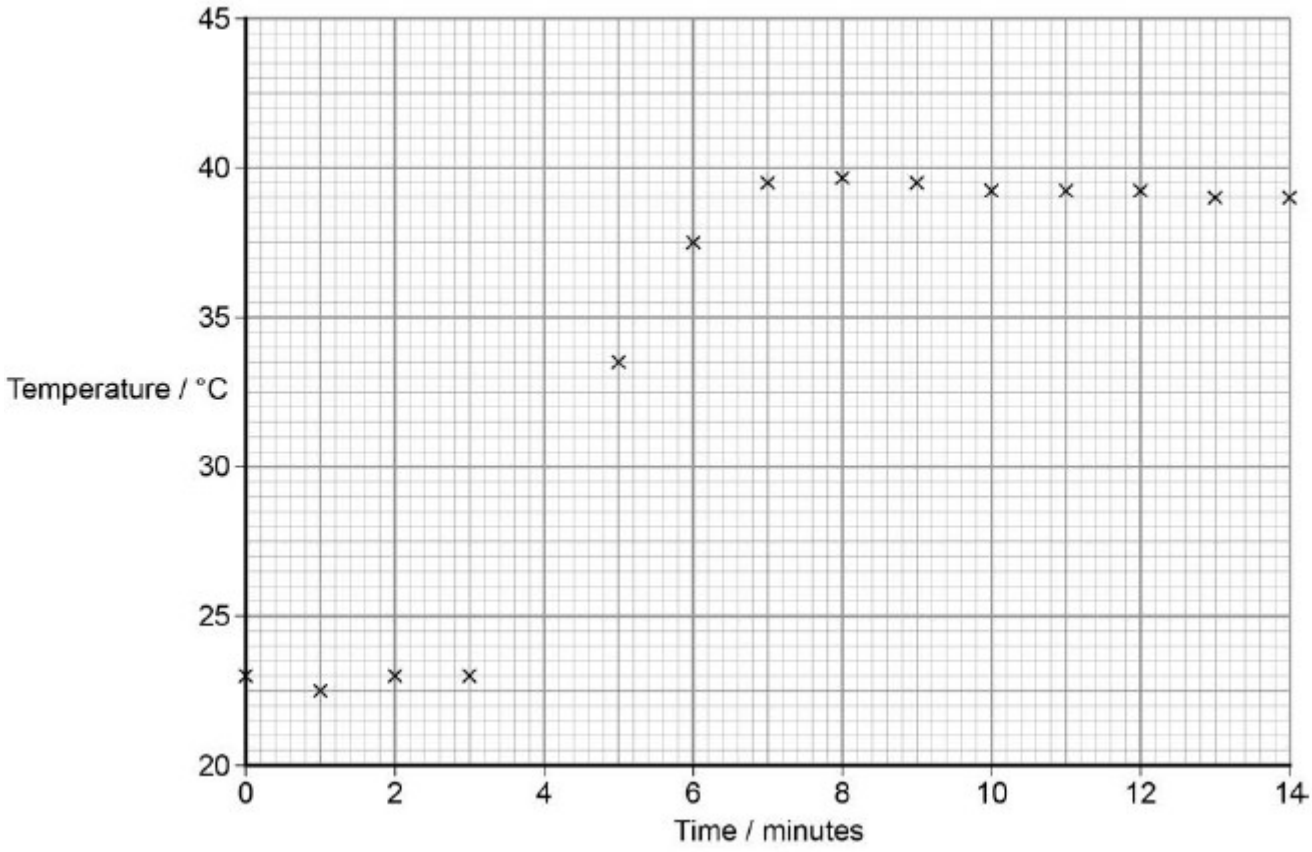
**(5)**
**(Total 14 marks)**

## Q6.

A student carried out a reaction between magnesium ribbon and aqueous trichloroethanoic acid in order to determine the enthalpy change. The equation for the reaction is shown:

$$Mg(s) + 2H^+(aq) \longrightarrow Mg^{2+}(aq) + H_2(g)$$

The student measured the initial temperature of the trichloroethanoic acid and again every minute for 3 minutes before adding the magnesium ribbon at the fourth minute.
The student continued to measure the temperature every minute for a further 10 minutes.
The graph for these measurements is shown below.




The student used 240 mg of magnesium and 10.0 $cm^3$ of aqueous trichloroethanoic acid (an excess).

Use these data and information determined from the graph above to calculate the enthalpy change, in kJ $mol^{-1}$, for this reaction.
Show your working.
Give your answer to an appropriate precision.
(The specific heat capacity of water = 4.18 J $K^{-1}$ $g^{-1}$)



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Enthalpy change = _____ kJ $mol^{-1}$

**(Total 7 marks)**

## Q7.

What is the temperature rise, in K, when 504 J of heat energy are absorbed by 0.110 kg of solid iron?
Specific heat capacity of iron = 0.448 J $K^{-1}$ $g^{-1}$

**A** $9.78 \times 10^{-2}$ ⬭

**B** $1.02 \times 10^{1}$ ⬭

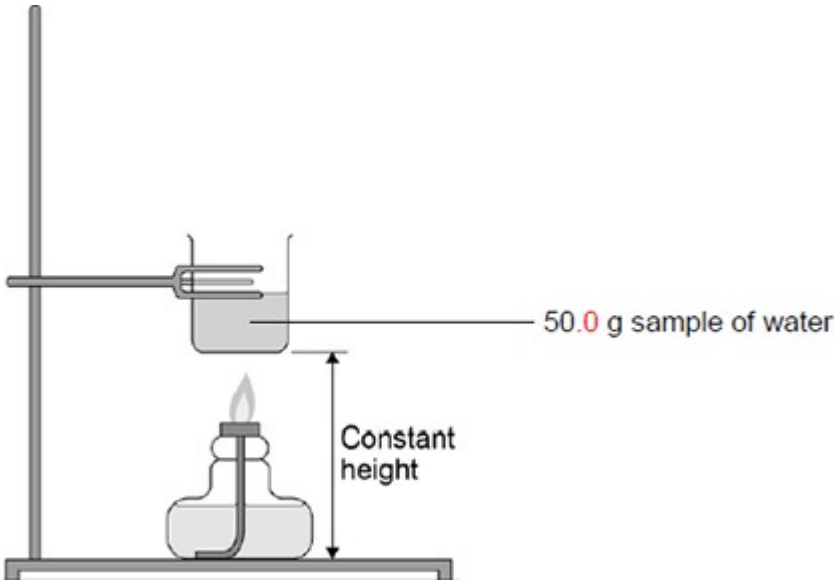
**C** $2.83 \times 10^{2}$ ⬭

**D** $1.02 \times 10^{4}$ ⬭

**(Total 1 mark)**

## Q8.

The figure below shows apparatus used in an experiment to determine the enthalpy of combustion of leaf alcohol.




The alcohol is placed in a spirit burner and weighed. The burner is lit and the alcohol allowed to burn for a few minutes. The flame is extinguished and the burner is re-weighed. The temperature of the water is recorded before and after heating.

The following table shows the results obtained.



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| | |
|---|---|
| Initial mass of spirit burner and alcohol / g | 56.38 |
| Final mass of spirit burner and alcohol / g | 55.84 |
| Initial temperature of water / °C | 20.7 |
| Final temperature of water / °C | 40.8 |

(a) Write an equation for the complete combustion of leaf alcohol ($CH_3CH_2CH=CHCH_2CH_2OH$).

_____

**(1)**

(b) Use the results from the table above to calculate a value for the enthalpy of combustion of leaf alcohol. Give units in your answer.
(The specific heat capacity of water is 4.18 J $K^{-1}$ $g^{-1}$)

Enthalpy of combustion = _____ Units = _____

**(4)**

(c) State how your answer to part (b) is likely to differ from the value quoted in reference sources.
Give **one** reason for your answer.

_____

_____

_____

_____

_____

**(2)**

(d) A 50.0 g sample of water was used in this experiment.

Explain how you could measure out this mass of water without using a balance.



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**(2)**

**(Total 9 marks)**

## Q9.

A 5.00 g sample of potassium chloride was added to 50.0 g of water initially at 20.0 °C. The mixture was stirred and as the potassium chloride dissolved, the temperature of the solution decreased.

(a) Describe the steps you would take to determine an accurate minimum temperature that is **not** influenced by heat from the surroundings.

_____

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**(4)**

(b) The temperature of the water decreased to 14.6 °C.

Calculate a value, in kJ $mol^{-1}$, for the enthalpy of solution of potassium chloride.

You should assume that only the 50.0 g of water changes in temperature and that the specific heat capacity of water is 4.18 J $K^{-1}$ $g^{-1}$.
Give your answer to the appropriate number of significant figures.

Enthalpy of solution = _____ kJ $mol^{-1}$



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**(4)**

(c) The enthalpy of solution of calcium chloride is −82.9 kJ $mol^{-1}$.
The enthalpies of hydration for calcium ions and chloride ions are −1650 and −364 kJ $mol^{-1}$, respectively.

Use these values to calculate a value for the lattice enthalpy of dissociation of calcium chloride.

Lattice enthalpy of dissociation = _____ kJ $mol^{-1}$

**(2)**

(d) Explain why your answer to part **(c)** is different from the lattice enthalpy of dissociation for magnesium chloride.

_____

_____

_____

_____

_____

**(2)**
**(Total 12 marks)**

## Q10.

An engineer was trying to develop a new fuel for a motorboat by blending mixtures of different alcohols in order to find out which mixture released the most energy when used in the engine.

The engineer had a number of alcohols in unlabelled bottles. It was decided to identify the alcohols by determining their enthalpies of combustion and comparing these values with those from a data book.

(a) Outline a simple practical experiment that the engineer could use, including the measurements to be taken, in order to determine the enthalpy of combustion for one of the unknown alcohols. You do **not** need to include details of any calculations.

_____

_____

_____



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_____

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_____

**(5)**

(b) Other than heat loss to the surroundings, identify **two** major sources of error in the experiment. Do **not** refer to the precision of the equipment.

_____

_____

_____

_____

_____

**(2)**

(c) The engineer found that the experimental values for the enthalpies of combustion of butan-1-ol and methylpropan-2-ol were very similar and so these values could not be used to distinguish between the two alcohols.

Identify a reagent that the engineer could use to distinguish between these two alcohols.
Give the observation in each case.

Reagent _____

_____

Butan-1-ol _____

_____

Methylpropan-2-ol _____

_____

**(3)**

(d) The filter in the air intake for the engine in the motorboat may become partially blocked by dust and debris.

Explain with the aid of an equation why combustion of methylpropan-2-ol under these circumstances would be of economic and environmental concern to the engineer.

_____

_____

_____



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**(3)**
**(Total 13 marks)**

## Q11.

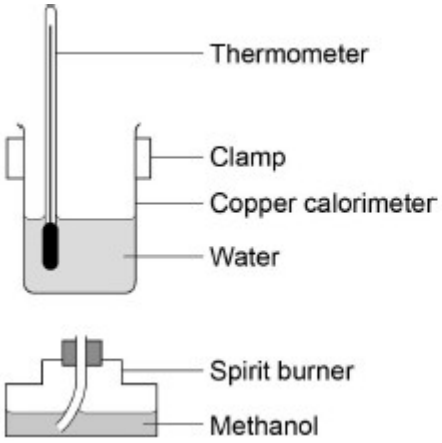
Alcohols such as methanol ($CH_3OH$), ethanol ($CH_3CH_2OH$) and propan-1-ol ($CH_3CH_2CH_2OH$) are good fuels.

(a) A student carried out an experiment to determine the enthalpy of combustion of methanol.

Methanol was placed in a spirit burner and the mass of the spirit burner measured. The student placed 100 g of water in a copper calorimeter and clamped it above the spirit burner. The burner was lit and allowed to burn for a few minutes. The flame was then extinguished and the new mass of the spirit burner found.

The measured temperature rise was 38.0 °C. The specific heat capacity of water is 4.18 J $K^{-1}$ $g^{-1}$.

A diagram of the apparatus is shown alongside a table which shows the measurements the student recorded.




| | |
|---|---|
| Mass of burner containing methanol before experiment | 214.02 g |
| Mass of burner containing methanol after experiment | 212.37 g |

Use the student's data to calculate an experimental value for the enthalpy of combustion of methanol in kJ $mol^{-1}$.



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_____

_____

_____

**(4)**

(b) Suggest **one** reason, other than incomplete combustion or heat transfer to the atmosphere, why the student's value for the enthalpy of combustion of methanol is different from that in a Data Book.

_____

_____

**(1)**

(c) The uncertainty in each of the temperature readings from the thermometer in this experiment was ±0.25 °C. This gave an overall uncertainty in the temperature rise of ±0.5 °C.

Calculate the percentage uncertainty for the use of the thermometer in this experiment.

_____

_____

**(1)**

(d) The student said correctly that using a thermometer with an overall uncertainty for the rise in temperature of ±0.5 °C was adequate for this experiment.

Explain why this thermometer was adequate for this experiment.

_____

_____

_____

**(1)**

(e) The enthalpy of combustion of ethanol is −1371 kJ $mol^{-1}$. The density of ethanol is 0.789 g $cm^{-3}$.

Calculate the heat energy released in kJ when 0.500 $dm^3$ of ethanol is burned. Give your answer to an appropriate number of significant figures.

_____

_____

_____

_____

**(3)**

**(Total 10 marks)**



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## Q12.

(a) Propanone can be formed when glucose comes into contact with bacteria in the absence of air.

(i) Balance the following equation for this reaction of glucose to form propanone, carbon dioxide and water.

.....$C_6H_{12}O_6$ ⟶ .....$CH_3COCH_3$ + .....$CO_2$ + .....$H_2O$

**(1)**

(ii) Deduce the role of the bacteria in this reaction.

_____

**(1)**

(b) Propanone is also formed by the oxidation of propan−2−ol.

(i) Write an equation for this reaction using [O] to represent the oxidising agent.

_____

**(1)**

(ii) State the class of alcohols to which propan−2−ol belongs.

_____

**(1)**

(c) A student determined a value for the enthalpy change when a sample of propanone was burned. The heat produced was used to warm some water in a copper calorimeter.
The student found that the temperature of 150 g of water increased by 8.0 °C when $4.50 \times 10^{-3}$ mol of pure propanone was burned in air.

Use the student's results to calculate a value, in kJ $mol^{-1}$, for the enthalpy change when one mole of propanone is burned.
(The specific heat capacity of water is 4.18 J $K^{-1}$ $g^{-1}$)

_____

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_____

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_____

**(3)**



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

_____

_____

_____

_____

_____

_____

**(3)**

(e) Use the mean bond enthalpy data in the table and the equation given below the table to calculate a value for the standard enthalpy change when gaseous propanone is burned.

| | **C−H** | **C−C** | **C−O** | **O−H** | **C=O** | **O=O** |
|---|---|---|---|---|---|---|
| **Mean bond enthalpy / kJ mol$^{-1}$** | 412 | 348 | 360 | 463 | 805 | 496 |

$$CH_3COCH_3(g) + 4O_2(g) \longrightarrow 3CO_2(g) + 3H_2O(g)$$

_____

_____

_____

_____

_____

_____

_____

**(3)**

(f) Suggest **two** reasons why the value obtained by the student in part (c) is different from the value calculated in part (e).

Reason 1 _____

_____

_____

Reason 2 _____

_____

_____

**(2)**

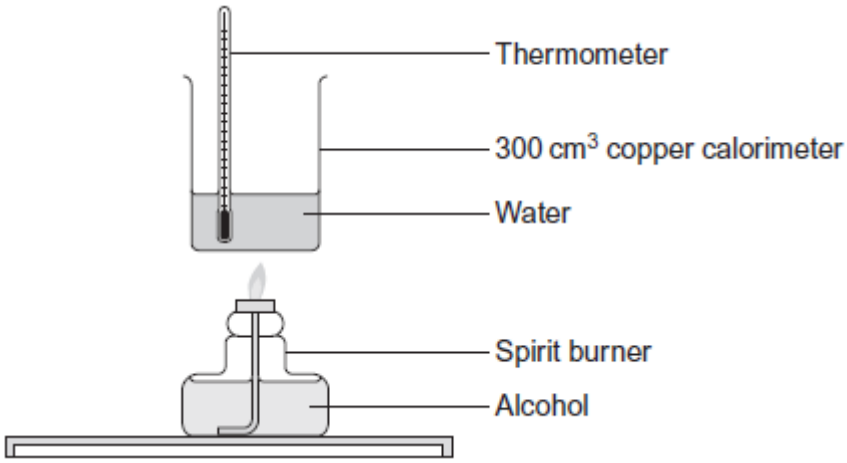


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**(Total 15 marks)**

## Q13.

A value for the enthalpy of combustion of an alcohol can be determined using the apparatus shown in the diagram. The calorimeter is held in position by a clamp.




This experiment can be repeated by using a different volume of water that would result in a more accurate value for the enthalpy of combustion because there would be a reduction in the heat lost.

State a change in the volume of water that would cause a reduction in heat loss and explain your answer.

Change in volume: _____

Explanation: _____

_____

**(Total 2 marks)**

## Q14.

Ethanol is an important fuel.

(a) A dilute aqueous solution of ethanol can be produced by the fermentation of an aqueous solution of glucose.
It is claimed that the ethanol obtained from this solution is a carbon-neutral biofuel.

Write an equation for this fermentation reaction.

Give **two** other essential conditions for this reaction to produce a good yield of ethanol.

Name a process used to produce a much more concentrated solution of ethanol from a dilute aqueous solution.

State the meaning of the term **carbon-neutral** in the context of this biofuel.

_____

_____



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**(5)**

(b) A student carried out a laboratory experiment to determine the enthalpy change when a sample of ethanol was burned. The heat produced was used to warm some water in a copper calorimeter. The student found that the temperature of 75.0 g of water increased by 5.50 °C when $2.40 \times 10^{-3}$ mol of pure ethanol was burned in air.

Use the student's results to calculate a value, in kJ $mol^{-1}$, for the enthalpy change when one mole of ethanol is burned.
(The specific heat capacity of water is 4.18 J $K^{-1}$ $g^{-1}$)

Deduce **two** reasons why the student's value for the standard enthalpy of combustion of ethanol is different from a Data Book value of –1279 kJ $mol^{-1}$.

**(5)**

(c) Mean bond enthalpies can be used to calculate enthalpies of reaction.

(i) Give the meaning of the term **mean bond enthalpy**.



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_____

_____

**(2)**

(ii) Consider the mean bond enthalpy data in the following table.

| | C—H | C—C | C—O | O=O | C=O | O—H |
|---|---|---|---|---|---|---|
| **Mean bond enthalpy / kJ mol$^{-1}$** | 412 | 348 | 360 | to be calculated | 805 | 463 |

Use the data in the table above and the equation shown to calculate a value for the bond enthalpy for the O=O double bond in an oxygen molecule.

$CH_3CH_2OH(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(g)$ $\Delta H = -1279$ kJ mol$^{-1}$

_____

_____

_____

_____

_____

_____

_____

_____

**(3)**

**(Total 15 marks)**

## Q15.

The alcohol 2-methylpropan-2-ol, $(CH_3)_3COH$, reacts to form esters that are used as flavourings by the food industry. The alcohol can be oxidised to produce carbon dioxide and water.

A student carried out an experiment on a pure sample of 2-methylpropan-2-ol to determine its enthalpy of combustion. A sample of the alcohol was placed into a spirit burner and positioned under a beaker containing 50 cm$^3$ of water. The spirit burner was ignited and allowed to burn for several minutes before it was extinguished.

The results for the experiment are shown in **Table 1**.

**Table 1**

| | |
|---|---|
| Initial temperature of the water / °C | 18.1 |
| Final temperature of the water / °C | 45.4 |
| Initial mass of spirit burner and alcohol / g | 208.80 |



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| | |
|---|---|
| Final mass of spirit burner and alcohol / g | 208.58 |

(a) Use the results from **Table 1** to calculate a value for the heat energy released from the combustion of this sample of 2-methylpropan-2-ol.
The specific heat capacity of water is 4.18 J $K^{-1}$ $g^{-1}$.
Show your working.

**(2)**

(b) Calculate the amount, in moles, of 2-methylpropan-2-ol burned in the experiment.
Hence calculate a value, in kJ $mol^{-1}$, for the enthalpy of combustion of 2-methylpropan-2-ol.
Show your working.

(If you were unable to calculate an answer to part (a), you should assume that the heat energy released was 5580 J. This is **not** the correct value.)

**(3)**

(c) An equation for the combustion of 2-methylpropan-2-ol is

$$(CH_3)_3COH(l) + 6O_2(g) \longrightarrow 4CO_2(g) + 5H_2O(l)$$

**Table 2** contains some standard enthalpy of formation data.

**Table 2**

| | $(CH_3)_3COH(l)$ | $O_2(g)$ | $CO_2(g)$ | $H_2O(l)$ |
|---|---|---|---|---|
| $\Delta H_f^{\ominus}$ **/ kJ mol$^{-1}$** | –360 | 0 | –393 | –286 |

Use the data from **Table 2** to calculate a value for the standard enthalpy of combustion of 2-methylpropan-2-ol. Show your working.



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**(3)**

(d) An accurate value for the enthalpy of combustion of 2-methylpropan-2-ol in which water is formed as a gas is –2422 kJ $mol^{-1}$.

Use this value and your answer from part (b) to calculate the overall percentage error in the student's experimental value for the enthalpy of combustion of 2-methylpropan-2-ol.

_____

_____

_____

**(1)**

(e) Suggest **one** improvement that would reduce errors due to heat loss in the student's experiment.

_____

_____

_____

**(1)**

(f) Suggest **one** other source of error in the student's experiment. Do **not** include heat loss, apparatus error or student error.

_____

_____

_____

**(1)**

**(Total 11 marks)**

## Q16.

The enthalpy of hydration for the chloride ion is −364 kJ $mol^{-1}$ and that for the bromide ion is −335 kJ $mol^{-1}$.

(a) By describing the nature of the attractive forces involved, explain why the value for the enthalpy of hydration for the chloride ion is more negative than that for the bromide ion.



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**(3)**

(b) The enthalpy of hydration for the potassium ion is −322 kJ $mol^{-1}$. The lattice enthalpy of dissociation for potassium bromide is +670 kJ $mol^{-1}$.

Calculate the enthalpy of solution for potassium bromide.

**(2)**

(c) The enthalpy of solution for potassium chloride is +17.2 kJ $mol^{-1}$.

(i) Explain why the free-energy change for the dissolving of potassium chloride in water is negative, even though the enthalpy change is positive.

**(3)**

(ii) A solution is formed when 5.00 g of potassium chloride are dissolved in 20.0 g of water. The initial temperature of the water is 298 K.



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Calculate the final temperature of the solution.

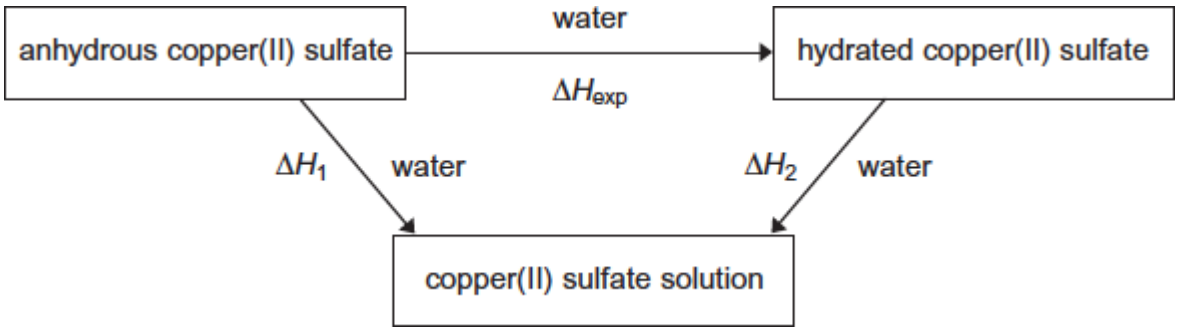
In your calculation, assume that only the 20.0 g of water changes in temperature and that the specific heat capacity of water is 4.18 J $K^{-1}$ $g^{-1}$.

**(5)**

**(Total 13 marks)**

## Q17.

A student used Hess's Law to determine a value for the enthalpy change that occurs when anhydrous copper(II) sulfate is hydrated. This enthalpy change was labelled $\Delta H_{exp}$ by the student in a scheme of reactions.




(a) State Hess's Law.

**(1)**

(b) Write a mathematical expression to show how $\Delta H_{exp}$, $\Delta H_1$ and $\Delta H_2$ are related to



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each other by Hess's Law.

_____

**(1)**

(c) Use the mathematical expression that you have written in part (b), and the data book values for the two enthalpy changes $\Delta H_1$ and $\Delta H_2$ shown, to calculate a value for $\Delta H_{exp}$

$\Delta H_1 = -156$ kJ mol$^{-1}$
$\Delta H_2 = +12$ kJ mol$^{-1}$

_____

_____

_____

_____

**(1)**

(d) The student added 0.0210 mol of pure anhydrous copper(II) sulfate to 25.0 cm$^3$ of deionised water in an open polystyrene cup. An exothermic reaction occurred and the temperature of the water increased by 14.0 °C.

(i) Use these data to calculate the enthalpy change, in kJ mol$^{-1}$, for this reaction of copper(II) sulfate. This is the student value for $\Delta H_1$

In this experiment, you should assume that all of the heat released is used to raise the temperature of the 25.0 g of water. The specific heat capacity of water is 4.18 J K$^{-1}$ g$^{-1}$.

_____

_____

_____

_____

_____

_____

_____

**(3)**

(ii) Suggest **one** reason why the student value for $\Delta H_1$ calculated in part (d)(i) is less accurate than the data book value given in part (c).

_____

_____

**(1)**

(e) Suggest **one** reason why the value for $\Delta H_{exp}$ **cannot** be measured directly.

_____



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_____

**(1)**

**(Total 8 marks)**

## Q18.

Methanol ($CH_3OH$) is an important fuel that can be synthesised from carbon dioxide.

(a) The table shows some standard enthalpies of formation.

| | $CO_2(g)$ | $H_2(g)$ | $CH_3OH(g)$ | $H_2O(g)$ |
|---|---|---|---|---|
| $\Delta H_f^{\ominus}$/kJ mol$^{-1}$ | – 394 | 0 | – 201 | – 242 |

(i) Use these standard enthalpies of formation to calculate a value for the standard enthalpy change of this synthesis.

$$CO_2(g) + 3H_2(g) \rightleftharpoons CH_3OH(g) + H_2O(g)$$

_____

_____

_____

_____

_____

_____

_____

**(3)**

(ii) State why the standard enthalpy of formation for hydrogen gas is zero.

_____

_____

**(1)**

(b) State and explain what happens to the yield of methanol when the total pressure is increased in this synthesis.

$$CO_2(g) + 3H_2(g) \rightleftharpoons CH_3OH(g) + H_2O(g)$$

Effect on yield _____

Explanation _____

_____

_____



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_____

_____

_____

**(3)**

(c) The hydrogen required for this synthesis is formed from methane and steam in a reversible reaction. The equation for this reaction is shown below.

$$CH_4(g) + H_2O(g) \rightleftharpoons C0(g) + 3H_2(g) \qquad \Delta H = +206 \text{ kJ mol}^{-1}$$

State and explain what happens to the yield of hydrogen in this reaction when the temperature is increased.

Effect on yield _____

Explanation _____

_____

_____

_____

_____

_____

**(3)**

(d) The methanol produced by this synthesis has been described as a carbon-neutral fuel.

(i) State the meaning of the term *carbon-neutral*.

_____

_____

_____

_____

**(1)**

(ii) Write an equation for the complete combustion of methanol.

_____

**(1)**

(iii) The equation for the synthesis of methanol is shown below.

$$CO_2(g) + 3H_2(g) \rightleftharpoons CH_3OH(g) + H_2O(g)$$

Use this equation and your answer to part (d)(ii) to deduce an equation to



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represent the overall chemical change that occurs when methanol behaves as a carbon-neutral fuel.

Equation _____

**(1)**

(e) A student carried out an experiment to determine the enthalpy change when a sample of methanol was burned.

The student found that the temperature of 140 g of water increased by 7.5 °C when 0.011 mol of methanol was burned in air and the heat produced was used to warm the water.

Use the student's results to calculate a value, in kJ $mol^{-1}$, for the enthalpy change when one mole of methanol was burned.
(The specific heat capacity of water is 4.18 J $K^{-1}$ $g^{-1}$).

_____

_____

_____

_____

_____

_____

_____

_____

**(3)**

**(Total 16 marks)**

## Q19.

(a) Anhydrous calcium chloride is not used as a commercial de-icer because it reacts with water. The reaction with water is exothermic and causes handling problems.

A student weighed out 1.00 g of anhydrous calcium chloride. Using a pipette, 25.0 $cm^3$ of water were measured out and transferred to a plastic cup. The cup was placed in a beaker to provide insulation. A thermometer was mounted in the cup using a clamp and stand. The bulb of the thermometer was fully immersed in the water.

The student recorded the temperature of the water in the cup every minute, stirring the water before reading the temperature. At the fourth minute the anhydrous calcium chloride was added, but the temperature was not recorded. The mixture was stirred, then the temperature was recorded at the fifth minute. The student continued stirring and recording the temperature at minute intervals for seven more



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minutes.

The student's results are shown in the table below.

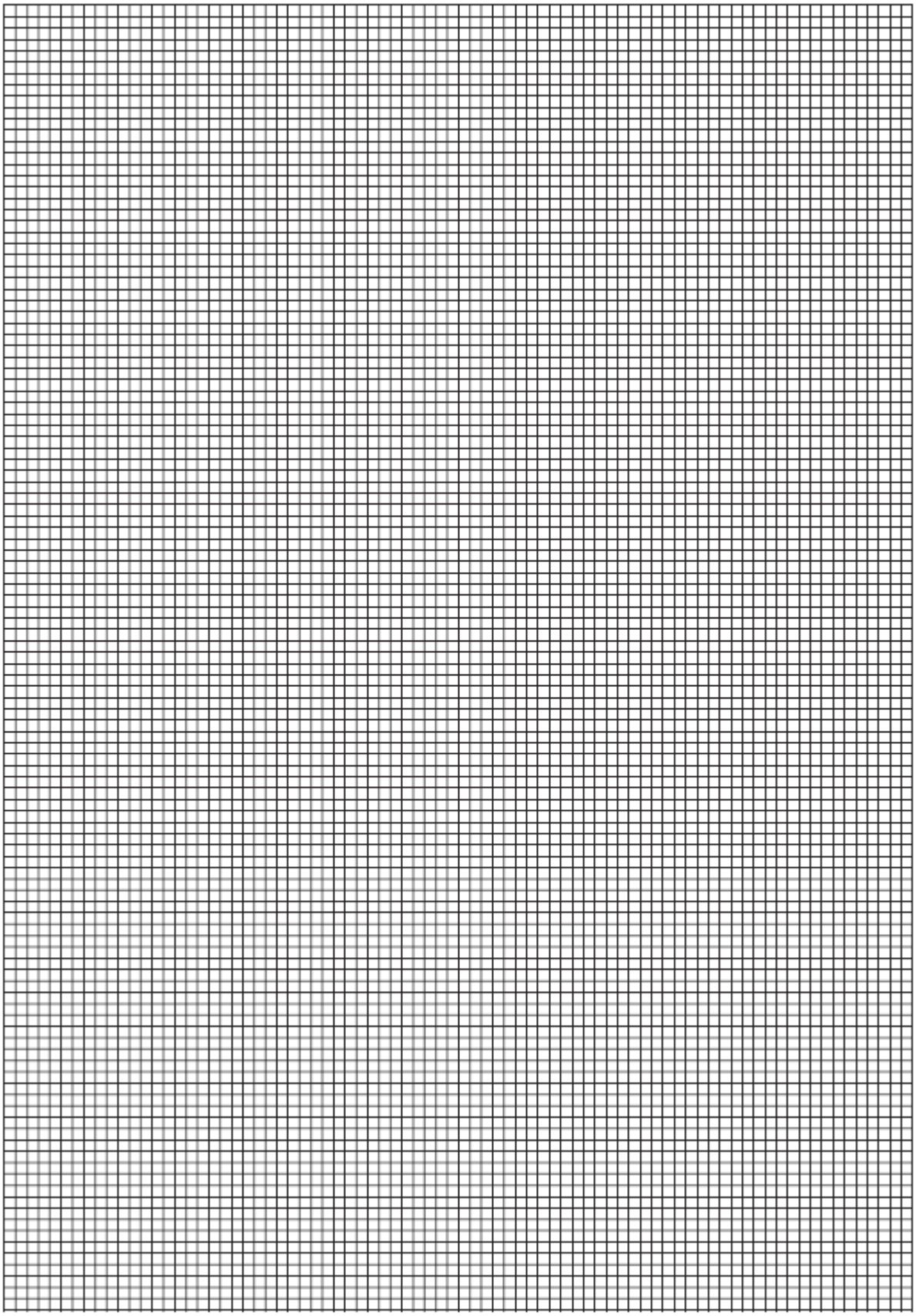
| Time / minutes | 0 | 1 | 2 | 3 | 4 |
|---|---|---|---|---|---|
| Temperature / °C | 19.6 | 19.5 | 19.5 | 19.5 | |

| Time / minutes | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | 12 |
|---|---|---|---|---|---|---|---|---|---|
| Temperature / °C | | 24.6 | 25.0 | 25.2 | 24.7 | 24.6 | 23.9 | 23.4 | 23.0 |

Plot a graph of temperature ($y$-axis) against time on the grid below.
Draw a line of best fit for the points before the fourth minute.
Draw a second line of best fit for the appropriate points after the fourth minute.
Extrapolate both lines to the fourth minute.



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**(5)**

(b) Use your graph to determine an accurate value for the temperature of the water at the fourth minute (**before** mixing).

Temperature before mixing _____



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**(1)**

(c) Use your graph to determine an accurate value for the temperature of the reaction mixture at the fourth minute (**after** mixing).

Temperature after mixing _____

**(1)**

(d) Use your answers from parts (b) and (c) to determine an accurate value for the temperature rise at the fourth minute.
Give your answer to the appropriate precision.

Temperature rise _____

**(1)**

(e) Use your answer from part (d) to calculate the heat given out during this experiment. Assume that the water has a density of 1.00 g $cm^{-3}$ and a specific heat capacity of 4.18 $JK^{-1}$ $g^{-1}$. Assume that all of the heat given out is used to heat the water.
Show your working.

_____

_____

_____

_____

**(2)**

(f) Calculate the amount, in moles, of $CaCl_2$ in 1.00 g of anhydrous calcium chloride ($M_r$ = 111.0).

_____

**(1)**

(g) Use your answers from parts (e) and (f) to calculate a value for the enthalpy change,

in kJ $mol^{-1}$, for the reaction that occurs when anhydrous calcium chloride dissolves in water.

$$CaCl_2(s) + aq \longrightarrow CaCl_2(aq)$$

_____

_____

_____

_____

**(2)**

(h) Explain why it is important that the reaction mixture is stirred before recording each temperature.

_____

_____



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**(1)**

(i) Anhydrous calcium chloride can be prepared by passing chlorine over heated calcium.
To prevent unreacted chlorine escaping into the atmosphere, a student suggested the diagram of the apparatus for this experiment shown below.




(i) Suggest **one** reason why the student wished to prevent unreacted chlorine escaping into the atmosphere.

_____

_____

**(1)**

(ii) Suggest **one** hazard of using the apparatus as suggested by the student for this experiment.

_____

_____

**(1)**

**(Total 16 marks)**

## Q20.

A student calculated that a value for the enthalpy change of neutralisation is –51.2 kJ $mol^{-1}$.

The design of a possible hand-warmer using hydrochloric acid and sodium hydroxide was discussed. It was proposed that 500 $cm^3$ of hydrochloric acid should be used in a flexible, sealed plastic container with a breakable tube of solid sodium hydroxide also in the container. On breaking the tube, the sodium hydroxide would be released, react with the acid and produce heat.
A 40 °C temperature rise was thought to be suitable.

(a) Calculate the heat energy, in J, required to raise the temperature of the reaction mixture by 40 °C. Assume that the reaction mixture has a density of 1.00 g $cm^{-3}$ and a specific heat capacity of 4.18 J $K^{-1}$ $g^{-1}$.
Assume that all of the heat energy given out is used to heat the reaction mixture.

_____

_____

_____

_____

**(2)**



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(b) Use your answer from part (a) and the value for the enthalpy change of neutralisation of –51.2 kJ $mol^{-1}$ to calculate the minimum amount, in moles, and hence the minimum mass of sodium hydroxide required in the breakable tube.
(If you could not complete the calculation in part (a) assume that the heat energy required was 77 400 J. This is **not** the correct answer).

Show your working.

Moles of NaOH _____

_____

Mass of NaOH _____

_____

**(3)**

(c) Use the amount, in moles, of sodium hydroxide from part (b) to calculate the minimum concentration, in mol $dm^{-3}$, of hydrochloric acid required in the 500 $cm^3$ of solution used in the sealed container.

_____

_____

_____

**(1)**

(d) Suggest **one** possible risk to a person who uses a hand-warmer containing sodium hydroxide and hydrochloric acid.

_____

_____

**(1)**

(e) A commercial hand-warmer uses powdered iron sealed in a plastic container.
A valve allows air to enter the container, and oxygen in the air reacts slowly with the iron to form solid iron(III) oxide. The heat released warms the container.

(i) Write an equation for this reaction between iron and oxygen to form iron(III) oxide.

_____

**(1)**

(ii) One version of an iron-oxygen hand-warmer advertises that it is designed to stay warm for up to four hours.
Other than by increasing the amount of iron in the container, state **one** change to the iron in the hand-warmer that would increase this time.
Explain why this change to the iron might **not** be an advantage.

Change to the iron _____

_____

Explanation _____



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_____

_____

**(3)**

(f) Another type of hand-warmer uses sodium thiosulfate. Sodium thiosulfate is very soluble in water at 80 °C but is much less soluble at room temperature.
When a hot, concentrated solution of sodium thiosulfate is cooled it does not immediately crystallise. The sodium thiosulfate stays dissolved as a stable 'super-saturated' solution until crystallisation is triggered.
Heat energy is then released when the sodium thiosulfate crystallises.

(i) This type of hand-warmer is re-usable.
Suggest **one** environmental advantage that a sodium thiosulfate hand-warmer has over the other two types.

_____

**(1)**

(ii) Describe the **two** steps that you would take to make the sodium thiosulfate hand-warmer ready for re-use.

Step 1 _____

_____

Step 2 _____

_____

**(2)**

**(Total 14 marks)**

## Q21.

Glucose, produced during photosynthesis in green plants, is a renewable source from which ethanol can be made. Ethanol is a liquid fuel used as a substitute for petrol.
The processes involved can be summarised as follows.

| | |
|---|---|
| Process **1** | Photosynthesis in green plants<br>$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$ |
| Process **2** | Fermentation of glucose to form ethanol |
| Process **3** | Complete combustion of ethanol<br>$CH_3CH_2OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$ |

(a) State **three** essential conditions for the fermentation of aqueous glucose in Process **2**.

Write an equation for the reaction that takes place during this fermentation.

_____

_____

_____



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**(4)**

(b) It has been claimed that there is no net carbon (greenhouse gas) emission to the atmosphere when ethanol made by Process **2** is used as a fuel.

State the term that is used to describe fuels of this type.

Use the equations for Processes **1, 2** and **3** to show why it can be claimed that there is no net emission of carbon-containing greenhouse gases.

**(3)**

(c) Use the information from the equation for Process **3** above and the mean bond enthalpies from the table below to calculate a value for the enthalpy change for this process.

| | C–H | C–C | C–O | O–H | C=O | O=O |
|---|---|---|---|---|---|---|
| Mean bond enthalpy / kJ $mol^{-1}$ | +412 | +348 | +360 | +463 | +743 | +496 |

Give **one** reason why the value calculated from mean bond enthalpies is different from the value given in a data book.



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_____

_____

**(4)**

(d) A student carried out a simple laboratory experiment to measure the enthalpy change for Process **3**. The student showed that the temperature of 200 g of water increased by 8.0 °C when 0.46 g of pure ethanol was burned in air and the heat produced was used to warm the water.

Use these results to calculate the value, in kJ $mol^{-1}$, obtained by the student for this enthalpy change. (The specific heat capacity of water is 4.18 J $K^{-1}$ $g^{-1}$)

Give **one** reason, other than heat loss, why the value obtained from the student's results is less exothermic than a data book value.

_____

_____

_____

_____

_____

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_____

_____

**(4)**

**(Total 15 marks)**

## Q22.

A group of students devised an experiment which they believed would enable them to investigate the strength of the intermolecular forces between ethyl ethanoate molecules ($CH_3COOCH_2CH_3$) and trichloromethane molecules ($CHCl_3$).

They mixed exactly 0.10 mol of each of the two liquids in a copper calorimeter and recorded the following results. The starting temperature of both liquids was the same.

| | |
|---|---|
| Mass of 0.10 mol of ethyl ethanoate / g | 8.80 |
| Mass of 0.10 mol of trichloromethane / g | 11.95 |
| Increase in temperature ($\Delta T$) on mixing / K | 9.5 |

(a) (i) Write an expression for the heat change ($q$) which relates mass ($m$), specific heat capacity (c) and change in temperature ($\Delta T$).

_____

**(1)**

(ii) Calculate the amount of heat required to increase the temperature of 8.80 g of



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ethyl ethanoate by 9.5 K during the mixing process. (You should assume that *c* for ethyl ethanoate = 1.92 J $g^{-1}$ $K^{-1}$)

_____

**(1)**

(iii) Calculate the amount of heat required to increase the temperature of 11.95 g of trichloromethane by 9.5 K during the mixing process. (You should assume that *c* for trichloromethane = 0.96 J $g^{-1}$ $K^{-1}$)

_____

**(1)**

(iv) Using the values from parts (a) (ii) and (a) (iii), calculate the molar enthalpy change in kJ $mol^{-1}$ for the mixing process.

_____

_____

**(2)**

(b) The students deduced that the heat change was due only to the formation of intermolecular forces between ethyl ethanoate molecules and trichloromethane molecules.

Ignoring all experimental errors, give **one** reason why the students may have made an incorrect deduction.

_____

_____

**(1)**

**(Total 6 marks)**

## Q23.

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

_____

_____

_____

**(3)**

(b) Use the data in the table to calculate the standard enthalpy of formation of liquid methylbenzene, $C_7H_8$

| Substance | C(s) | $H_2(g)$ | $C_7H_8(l)$ |
|---|---|---|---|
| Standard enthalpy of combustion, $\Delta H_c^{\ominus}$ /kJ $mol^{-1}$ | –394 | –286 | –3909 |

$7C(s) + 4H_2(g) \rightarrow C_7H_8(l)$

_____

_____

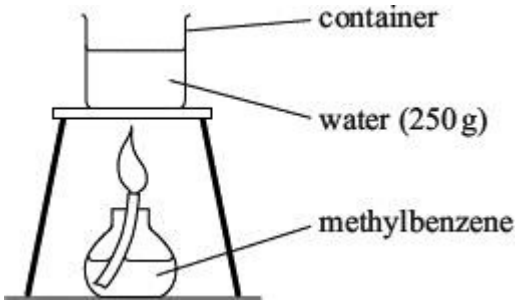


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_____

**(3)**

(c) An experiment was carried out to determine a value for the enthalpy of combustion of liquid methylbenzene using the apparatus shown in the diagram.




Burning 2.5 g of methylbenzene caused the temperature of 250 g of water to rise by 60°C. Use this information to calculate a value for the enthalpy of combustion of methylbenzene, $C_7H_8$

(The specific heat capacity of water is 4.18 J $K^{-1}$ $g^{-1}$. Ignore the heat capacity of the container.)

_____

_____

_____

_____

**(4)**

(d) A 25.0 $cm^3$ sample of 2.00 mol $dm^{-3}$ hydrochloric acid was mixed with 50.0 $cm^3$ of a 1.00 mol $dm^{-3}$ solution of sodium hydroxide. Both solutions were initially at 18.0 °C.

After mixing, the temperature of the final solution was 26.5°C.

Use this information to calculate a value for the standard enthalpy change for the following reaction.

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

In your calculation, assume that the density of the final solution is 1.00 g $cm^{-3}$ and that its specific heat capacity is the same as that of water. (Ignore the heat capacity of the container.)

_____

_____

_____

_____

**(4)**

(e) Give **one** reason why your answer to part (d) has a much smaller experimental error than your answer to part (c).



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_____

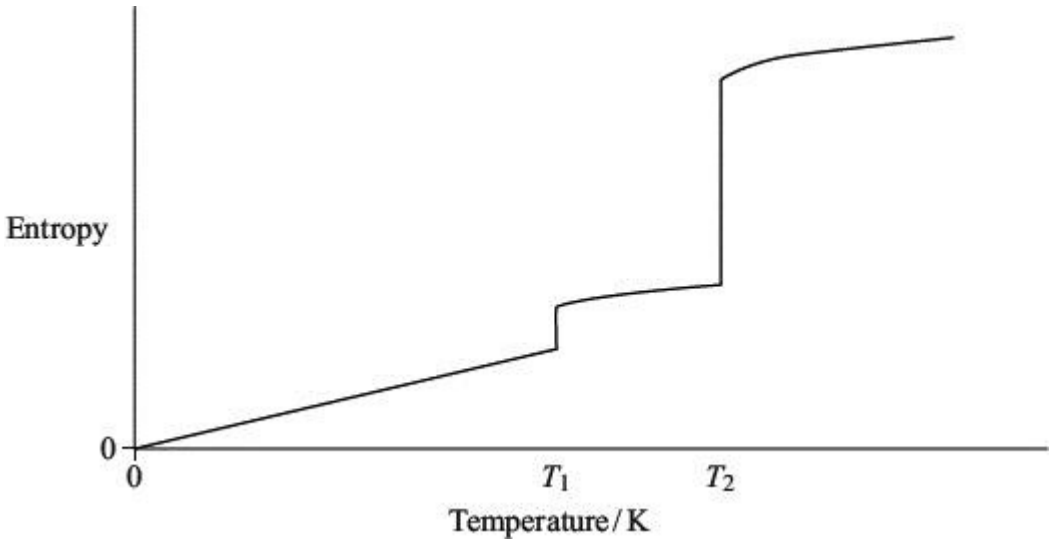
_____

**(1)**

**(Total 15 marks)**

## Q24.

The sketch graph below shows how the entropy of a sample of water varies with temperature.




(a) Suggest why the entropy of water is zero at 0 K.

_____

**(1)**

(b) What change of state occurs at temperature $T_1$?

_____

**(1)**

(c) Explain why the entropy change, $\Delta S$, at temperature $T_2$ is much larger than that at temperature $T_1$.

_____

_____

_____

**(2)**

(d) It requires 3.49 kJ of heat energy to convert 1.53 g of liquid water into steam at 373 K and 100 kPa.

(i) Use these data to calculate the enthalpy change, $\Delta H$, when 1.00 mol of liquid water forms 1.00 mol of steam at 373 K and 100 kPa.

_____

_____



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_____

_____

(ii) Write an expression showing the relationship between free-energy change, $\Delta G$, enthalpy change, $\Delta H$, and entropy change, $\Delta S$.

_____

(iii) For the conversion of liquid water into steam at 373 K and 100 kPa, $\Delta G = 0$ kJ mol$^{-1}$

Calculate the value of $\Delta S$ for the conversion of one mole of water into steam under these conditions. State the units.

(If you have been unable to complete part (d)(i) you should assume that $\Delta H = 45.0$ kJ mol$^{-1}$. This is not the correct answer.)

*Calculation* _____

_____

_____

*Units* _____

**(6)**

**(Total 10 marks)**

## Q25.

Methanol, $CH_3OH$, is a convenient liquid fuel.

(a) An experiment was conducted to determine the enthalpy of combustion of liquid methanol. The energy obtained from burning 2.12 g of methanol was used to heat 150 g of water. The temperature of the water rose from 298 K to 362 K. (The specific heat capacity of water is 4.18 J K$^{-1}$ g$^{-1}$)

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

_____

_____

_____

_____

_____

_____

**(3)**

(ii) Use the data above to calculate a value for the enthalpy of combustion of one mole of liquid methanol.

_____



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_____

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_____

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_____

_____

_____

**(4)**

(b) Methanol can be synthesised from methane and steam by a process that occurs in two stages.

*Stage 1* $CH_4(g) + H_2O(g) \rightleftharpoons 3H_2(g) + CO(g)$ $\Delta H^{\ominus} = +206 \text{ kJ mol}^{-1}$

*Stage 2* $CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$ $\Delta H^{\ominus} = -91 \text{ kJ mol}^{-1}$

(i) Explain why, in *Stage 1*, a higher yield of hydrogen and carbon monoxide is **not** obtained if the pressure is increased.

_____

_____

_____

_____

**(2)**

(ii) *Stage 2* is carried out at a compromise temperature of 500K. By considering what would happen at higher and lower temperatures, explain why 500 K is considered to be a compromise for *Stage 2*.

_____

_____

_____

_____

_____

_____

**(3)**

(c) The standard enthalpies of combustion of carbon monoxide and of hydrogen are –283 kJ mol$^{-1}$ and –286 kJ mol$^{-1}$, respectively. Use these data and the enthalpy change for *Stage 2* to calculate a value for the standard enthalpy of combustion of gaseous methanol.

_____



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**(3)**

**(Total 15 marks)**

## Q26.

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

**(3)**

(b) State Hess's Law and use it, together with the data given in the table below, to calculate the standard enthalpy change for the following reaction.

$MgO(s) + 2HCl(g) \rightarrow MgCl_2(s) + H_2O(l)$

| | MgO | HCl(g) | $MgCl_2$ | $H_2O$ |
|---|---|---|---|---|
| $\Delta H_f^{\ominus}$/kJ mol$^{-1}$ | –602 | –92 | –642 | –286 |



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**(4)**

(c) In an experiment, an excess of solid magnesium oxide was added to 50 $cm^3$ of 3.0 mol $dm^{-3}$ hydrochloric acid. The initial temperature of the solution was 21 °C. After reaction, the temperature had risen to 53 °C. (The specific heat capacity of water is 4.2 J $K^{-1}$ $g^{-1}$)

Use this information to calculate the enthalpy change for the reaction of one mole of magnesium oxide with hydrochloric acid. For your calculation you should assume that all the heat from the reaction is used to raise the temperature of 50 g of water.

**(8)**



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**(Total 15 marks)**

## Q27.

(a) Write an equation for the complete combustion of propanone, $C_3H_6O$, to form carbon dioxide and water.

_____

**(1)**

(b) In a laboratory experiment, 1.45 g of propanone were burned completely in oxygen. The heat from this combustion was used to raise the temperature of 100 g of water from 293.1 K to 351.2 K.

(i) Calculate the number of moles of propanone in the 1.45 g.

_____

_____

(ii) Calculate the heat energy required to raise the temperature of 100 g of water from 293.1 K to 351.2 K.
(The specific heat capacity of water is 4.18 J $K^{-1}$ $g^{-1}$)

_____

_____

_____

_____

(iii) Hence, calculate a value, in kJ $mol^{-1}$, for the enthalpy of combustion of propanone.

_____

_____

**(5)**

(c) In a similar experiment, the enthalpy of combustion of butanone, $C_4H_8O$, was found to be –1290 kJ $mol^{-1}$. A data book value for the same reaction is $\Delta H_c^{\ominus}$ = –2430 kJ $mol^{-1}$.

(i) Suggest one reason why the experimental value is very different from the data book value.

_____

(ii) This data book value of $\Delta H_c^{\ominus}$ for butanone (–2430 kJ $mol^{-1}$) refers to the formation of carbon dioxide gas and water in the gaseous state. How would this value differ if it referred to the formation of water in the liquid state? Explain your answer.

*Difference* _____

*Explanation* _____



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_____

**(3)**

(d) Calculate a value for the standard enthalpy of formation for liquid ethanethiol, $C_2H_5SH$. Use the equation given below and enthalpy of combustion data from the following table.

| Substance | $C_2H_5SH(l)$ | C(s) | $H_2(g)$ | S(s) |
|---|---|---|---|---|
| $\Delta H_c^{\ominus}$ / kJ mol$^{-1}$ | –1170 | –394 | –286 | –297 |

$2C(s) + 3H_2(g) + S(s) \rightarrow C_2H_5SH(l)$

_____

_____

_____

_____

**(3)**

**(Total 12 marks)**

## Q28.

A 50.0 cm$^3$ sample of a 0.200 mol dm$^{-3}$ solution of silver nitrate was placed in a polystyrene beaker. An excess of powdered zinc was added to this solution and the mixture stirred. Zinc nitrate, $Zn(NO_3)_2$, and silver were formed and a rise in temperature of 3.20 °C was recorded.

(a) Write an equation for the reaction between silver nitrate and zinc.

_____

**(1)**

(b) Calculate the number of moles of silver nitrate used in the experiment.

_____

_____

**(2)**

(c) Calculate the heat energy evolved by the reaction in this experiment assuming that all the energy evolved is used to heat only the 50.0 g of water in the mixture. (Specific heat capacity of water is 4.18 J g$^{-1}$ K$^{-1}$)

_____

_____

**(2)**

(d) Calculate the heat energy change for the reaction per mole of zinc reacted.

_____

_____



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**(2)**

(e) Explain why the experimental value for the heat energy evolved in this experiment is less than the correct value.

_____

_____

**(1)**

**(Total 8 marks)**



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## Q1.

This question is about energetics.

(a) Write an equation, including state symbols, for the reaction with an enthalpy change equal to the enthalpy of formation for iron(III) oxide.

_____

**(1)**

(b) **Table 1** contains some standard enthalpy of formation data.

**Table 1**

| | $CO(g)$ | $Fe_2O_3(s)$ |
|---|---|---|
| **$\Delta_f H^\ominus$ / kJ mol$^{-1}$** | −111 | −822 |

$Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(s) + 3CO_2(g)$ $\Delta H = -19$ kJ mol$^{-1}$

Use these data and the equation for the reaction of iron(III) oxide with carbon monoxide to calculate a value for the standard enthalpy of formation for carbon dioxide.

Show your working.

$\Delta_f H^\ominus$ _____ kJ mol$^{-1}$

**(3)**

(c) Some enthalpy data are given in **Table 2**.

**Table 2**

| **Process** | **$\Delta H$ / kJ mol$^{-1}$** |
|---|---|
| $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$ | −92 |
| $N_2(g) \longrightarrow 2N(g)$ | +944 |
| $H_2(g) \longrightarrow 2H(g)$ | +436 |



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Use the data from **Table 2** to calculate the bond enthalpy for N−H in ammonia.

N−H bond enthalpy _____ kJ mol$^{-1}$

**(3)**

(d) Give one reason why the bond enthalpy that you calculated in part (c) is different from the mean bond enthalpy quoted in a data book (388 kJ mol$^{-1}$).

_____

_____

_____

**(1)**

**(Total 8 marks)**

## Q2.

When alkanes are burned in an excess of oxygen they produce carbon dioxide and water.

(a) Write an equation for the complete combustion of propane in oxygen.

_____

**(1)**

(b) An expression can be derived using bond enthalpy data to estimate the enthalpy of combustion ($\Delta_c H$) of an alkane.

For an alkane with ***n*** carbon atoms: $\Delta_c H = -(496\boldsymbol{n} + 202)$ kJ mol$^{-1}$

The enthalpy of combustion of an alkane was calculated to be −6650 kJ mol$^{-1}$ using this expression.

Deduce the molecular formula of this alkane.

Show your working.



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Molecular formula of alkane _____

**(2)**

(c) Suggest **one** reason, other than the use of mean bond enthalpies, why a value for the enthalpy of combustion of a liquid alkane is different from the value obtained using the expression in part (b)

_____

_____

**(1)**

(d) Values of the enthalpy change for combustion of 1 g of some alkanes are shown in the table.

| | **methane** | **ethane** | **propane** | **butane** | **pentane** |
|---|---|---|---|---|---|
| Enthalpy change in kJ for combustion of 1 g | −55.6 | −52.0 | | −49.6 | −48.7 |

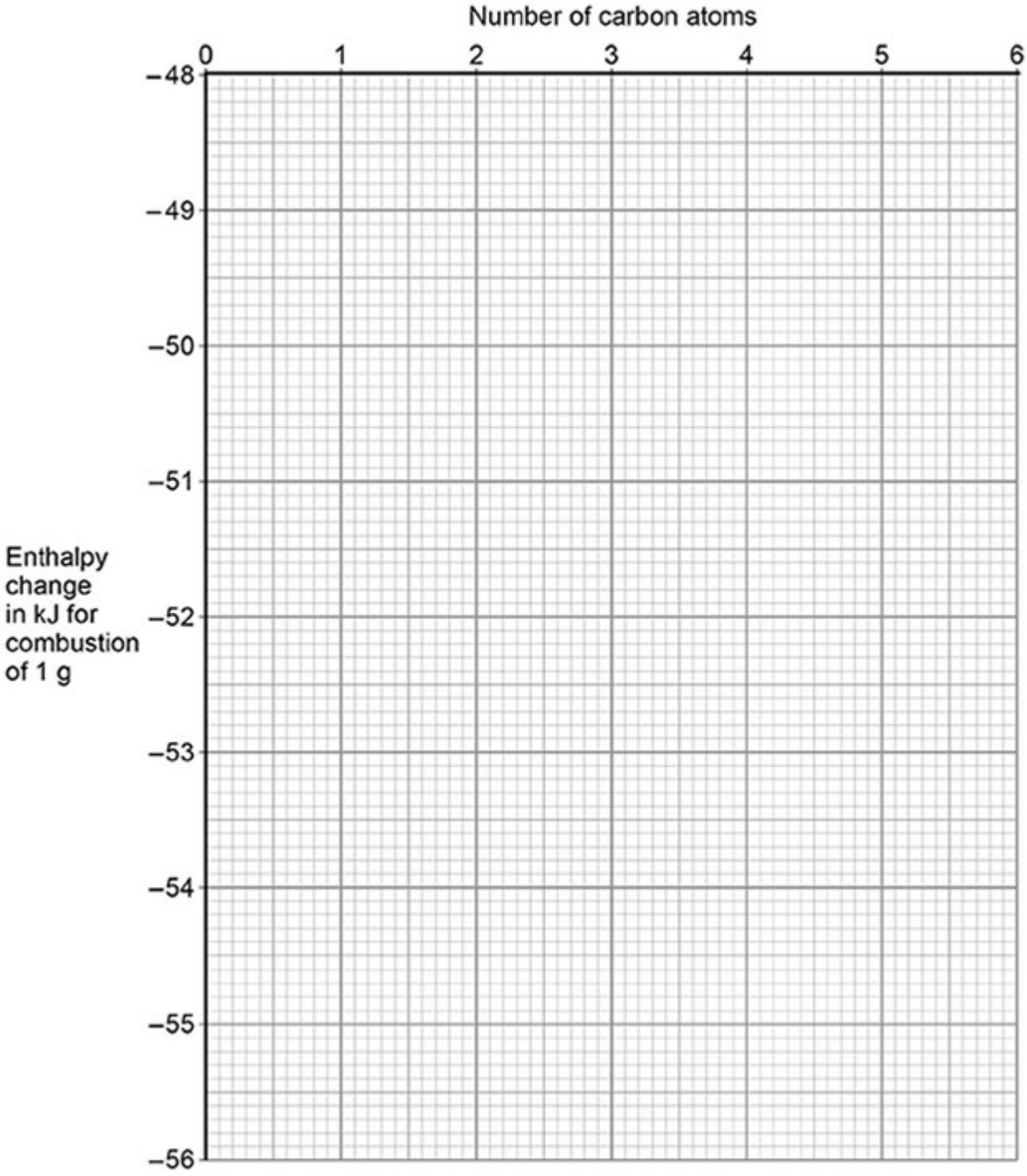
Plot the enthalpy change for the combustion of 1 g against the number of carbon atoms in the alkanes in the table.

Draw a best fit line and use this to estimate the enthalpy change for combustion of 1 g of propane.

Write your answer in the table.



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**(3)**

(e) Isooctane (2,2,4-trimethylpentane) is an important component of petrol used in cars.

When isooctane is burned, the enthalpy change is −47.8 kJ $g^{-1}$

Isooctane is a liquid at room temperature with a density of 0.692 g $cm^{-3}$

Calculate the heat energy released, in kJ, when 1.00 $dm^3$ of isooctane burns in excess oxygen.

Give your answer to the appropriate number of significant figures.



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Heat energy released _____ kJ

**(2)**

**(Total 9 marks)**

## Q3.

This question is about enthalpy changes.

(a) Write an equation, including state symbols, to show the reaction taking place when the standard enthalpy of combustion for ethanol is measured.

_____

**(2)**

(b) State the name given to the enthalpy change represented by the following chemical equation.

Explain why this enthalpy change would be difficult to determine directly.

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

Enthalpy change _____

Explanation _____

_____

_____

**(2)**

(c) Standard enthalpies of combustion for carbon and carbon monoxide are −393 kJ $mol^{-1}$ and −283 kJ $mol^{-1}$, respectively.

Use these data to calculate the enthalpy change for the reaction in part **(b)**.

Enthalpy change = _____ kJ $mol^{-1}$

**(2)**

(d) Use the following data to calculate a value for the Xe–F bond enthalpy in $XeF_4$



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$$Xe(g) + 2F_2(g) \rightarrow XeF_4(g) \quad \Delta H = -252 \text{ kJ mol}^{-1}$$
$$F_2(g) \rightarrow 2F(g) \quad \Delta H = +158 \text{ kJ mol}^{-1}$$

Xe–F bond enthalpy = _____ kJ mol$^{-1}$

**(3)**

(e) Suggest a reason why the value calculated in part **(d)** differs from the mean Xe–F bond enthalpy quoted in a data source.

_____

_____

_____

**(1)**

**(Total 10 marks)**

## Q4.

Calculate the enthalpy change, in kJ, for this dissociation of mole of propan-1-ol.

$$C_3H_7OH(g) \longrightarrow 3C(g) + 8H(g) + O(g)$$

| | C—H | C—C | C—O | O—H |
|---|---|---|---|---|
| Mean bond dissociation enthalpy / kJ mol$^{-1}$ | 412 | 348 | 360 | 463 |

**A** −4751 ○

**B** −4403 ○

**C** +4403 ○

**D** +4751 ○

**(Total 1 mark)**

## Q5.



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(a) Write an equation, including state symbols, for the reaction with enthalpy change equal to the standard enthalpy of formation for $CF_4(g)$.

_____

**(1)**

(b) Explain why $CF_4$ has a bond angle of 109.5°.

_____

_____

_____

_____

_____

**(2)**

(c) **Table 1** gives some values of standard enthalpies of formation ($\Delta_f H^\ominus$).

**Table 1**

| **Substance** | $F_2(g)$ | $CF_4(g)$ | HF(g) |
|---|---|---|---|
| **$\Delta_f H^\ominus$ / kJ mol$^{-1}$** | 0 | −680 | −269 |

The enthalpy change for the following reaction is −2889 kJ mol$^{-1}$.

$$C_2H_6(g) + 7F_2(g) \longrightarrow 2CF_4(g) + 6HF(g)$$

Use this value and the standard enthalpies of formation in **Table 1** to calculate the standard enthalpy of formation of $C_2H_6(g)$.

Standard enthalpy of formation of $C_2H_6(g)$ = _____ kJ mol$^{-1}$

**(3)**

(d) Methane reacts violently with fluorine according to the following equation.

$$CH_4(g) + 4F_2(g) \longrightarrow CF_4(g) + 4HF(g) \quad \Delta H = -1904 \text{ kJ mol}^{-1}$$

Some mean bond enthalpies are given in **Table 2**.

**Table 2**

| **Bond** | C−H | C−F | H−F |
|---|---|---|---|



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| **Mean bond enthalpy / kJ mol$^{-1}$** | 412 | 484 | 562 |
|---|---|---|---|

A student suggested that one reason for the high reactivity of fluorine is a weak F−F bond.

Is the student correct? Justify your answer with a calculation using these data.

_____

_____

**(4)**

**(Total 10 marks)**

## Q6.

Many chemical processes release waste products into the atmosphere. Scientists are developing new solid catalysts to convert more efficiently these emissions into useful products, such as fuels. One example is a catalyst to convert these emissions into methanol. The catalyst is thought to work by breaking a H–H bond.

An equation for this formation of methanol is given below.

$$CO_2(g) + 3H_2(g) \rightleftharpoons CH_3OH(g) + H_2O(g) \quad \Delta H = -49 \text{ kJ mol}^{-1}$$

Some mean bond enthalpies are shown in the following table.

| **Bond** | C=O | C–H | C–O | O–H |
|---|---|---|---|---|
| **Mean bond enthalpy / kJ mol$^{-1}$** | 743 | 412 | 360 | 463 |

(a) Use the enthalpy change for the reaction and data from the table to calculate a value for the H–H bond enthalpy.

H–H bond enthalpy = _____ kJ mol$^{-1}$

**(3)**

(b) A data book value for the H–H bond enthalpy is 436 kJ mol$^{-1}$.



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Suggest **one** reason why this value is different from your answer to part (a).

_____

_____

_____

**(1)**

(c) Suggest **one** environmental advantage of manufacturing methanol fuel by this reaction.

_____

_____

_____

**(1)**

(d) Use Le Chatelier's principle to justify why the reaction is carried out at a high pressure rather than at atmospheric pressure.

_____

_____

_____

_____

_____

_____

**(3)**

(e) Suggest why the catalyst used in this process may become less efficient if the carbon dioxide and hydrogen contain impurities.

_____

_____

_____

**(1)**

(f) In a laboratory experiment to investigate the reaction shown in the equation below, 1.0 mol of carbon dioxide and 3.0 mol of hydrogen were sealed into a container. After the mixture had reached equilibrium, at a pressure of 500 kPa, the yield of methanol was 0.86 mol.

$$CO_2(g) + 3H_2(g) \rightleftharpoons CH_3OH(g) + H_2O(g)$$

Calculate a value for $K_p$
Give your answer to the appropriate number of significant figures.
Give units with your answer.



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$K_p$ = _____ Units = _____

**(7)**

**(Total 16 marks)**

## Q7.

Hydrogen is produced by the reaction of methane with steam. The reaction mixture reaches a state of dynamic equilibrium.

$$CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3H_2(g) \qquad \Delta H = +206 \text{ kJ mol}^{-1}$$

Some enthalpy data is given in the table.

| **Bond** | C–H | O–H | H–H | C≡H |
|---|---|---|---|---|
| **Bond enthalpy / kJ mol$^{-1}$** | 413 | 463 | 436 | To be calculated |

Use the information in the table and the stated enthalpy change to calculate the missing bond enthalpy.

**A** 234 ⬭

**B** 1064 ⬭

**C** 1476 ⬭

**D** 1936 ⬭

**(Total 1 mark)**

## Q8.

Standard enthalpy of combustion data can be used to calculate enthalpies of formation.

(a) State the meaning of the term standard enthalpy of combustion.



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**(3)**

(b) The equation corresponding to the enthalpy of formation of propan-1-ol is shown.

$$3C(s) + 4H_2(g) + \frac{1}{2}O_2(g) \longrightarrow CH_3CH_2CH_2OH(l)$$

**Table 1** contains some standard enthalpy of combustion data.

**Table 1**

| | **C(s)** | **$H_2$(g)** | **$CH_3CH_2CH_2OH$(l)** |
|---|---|---|---|
| $\Delta H_c^{\ominus}$ / kJ mol$^{-1}$ | –394 | –286 | –2010 |

Use data from **Table 1** to calculate a value for the standard enthalpy of formation of propan-1-ol. Show your working.

**(3)**

(c) An equation for the complete combustion of gaseous propan-1-ol is shown.

$$CH_3CH_2CH_2OH(g) + 4\frac{1}{2}O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(g) \quad \Delta H = -1893 \text{ kJ mol}^{-1}$$

**Table 2** shows some bond enthalpy data.

**Table 2**

| | **C–H** | **C–O** | **O–H** | **C=O** | **O=O** |
|---|---|---|---|---|---|
| Bond enthalpy / kJ mol$^{-1}$ | 412 | 360 | 463 | 805 | 496 |



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Use data from **Table 2** and the enthalpy change for this reaction to calculate a value for the bond enthalpy of a C–C bond in propan-1-ol.

_____

_____

_____

_____

_____

_____

_____

_____

**(3)**

**(Total 9 marks)**

## Q9.

(a) Propanone can be formed when glucose comes into contact with bacteria in the absence of air.

(i) Balance the following equation for this reaction of glucose to form propanone, carbon dioxide and water.

.....$C_6H_{12}O_6$ ⟶ .....$CH_3COCH_3$ + .....$CO_2$ + .....$H_2O$

**(1)**

(ii) Deduce the role of the bacteria in this reaction.

_____

**(1)**

(b) Propanone is also formed by the oxidation of propan−2−ol.

(i) Write an equation for this reaction using [O] to represent the oxidising agent.

_____

**(1)**

(ii) State the class of alcohols to which propan−2−ol belongs.

_____

**(1)**

(c) A student determined a value for the enthalpy change when a sample of propanone was burned. The heat produced was used to warm some water in a copper calorimeter.
The student found that the temperature of 150 g of water increased by 8.0 °C when $4.50 \times 10^{-3}$ mol of pure propanone was burned in air.

Use the student's results to calculate a value, in kJ $mol^{-1}$, for the enthalpy change when one mole of propanone is burned.



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(The specific heat capacity of water is 4.18 J $K^{-1}$ $g^{-1}$)

**(3)**

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

**(3)**

(e) Use the mean bond enthalpy data in the table and the equation given below the table to calculate a value for the standard enthalpy change when gaseous propanone is burned.

| | **C−H** | **C−C** | **C−O** | **O−H** | **C=O** | **O=O** |
|---|---|---|---|---|---|---|
| **Mean bond enthalpy / kJ $mol^{-1}$** | 412 | 348 | 360 | 463 | 805 | 496 |

$$CH_3COCH_3(g) + 4O_2(g) \longrightarrow 3CO_2(g) + 3H_2O(g)$$



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**(3)**

(f) Suggest **two** reasons why the value obtained by the student in part (c) is different from the value calculated in part (e).

Reason 1 _____

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Reason 2 _____

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**(2)**

**(Total 15 marks)**

## Q10.

The table contains some bond enthalpy data.

| **Bond** | H−H | O=O | H−O |
|---|---|---|---|
| **Bond enthalpy / kJ mol$^{-1}$** | 436 | 496 | 464 |

(a) The value for the H−O bond enthalpy in the table is a mean bond enthalpy.

State the meaning of the term **mean bond enthalpy** for the H−O bond.

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**(2)**

(b) Use the bond enthalpies in the table to calculate a value for the enthalpy of formation of water in the gas phase.

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**(3)**

(c) The standard enthalpy of combustion of hydrogen, forming water in the gas phase, is almost the same as the correct answer to part (b).

(i) Suggest **one** reason why you would expect the standard enthalpy of combustion of hydrogen to be the same as the answer to part (b).

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**(1)**

(ii) Suggest **one** reason why you would expect the standard enthalpy of combustion of hydrogen to differ slightly from the answer to part (b).

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**(1)**

**(Total 7 marks)**

## Q11.

Ethanol is an important fuel.

(a) A dilute aqueous solution of ethanol can be produced by the fermentation of an aqueous solution of glucose.
It is claimed that the ethanol obtained from this solution is a carbon-neutral biofuel.

Write an equation for this fermentation reaction.

Give **two** other essential conditions for this reaction to produce a good yield of ethanol.

Name a process used to produce a much more concentrated solution of ethanol from a dilute aqueous solution.

State the meaning of the term **carbon-neutral** in the context of this biofuel.

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**(5)**

(b) A student carried out a laboratory experiment to determine the enthalpy change when a sample of ethanol was burned. The heat produced was used to warm some water in a copper calorimeter. The student found that the temperature of 75.0 g of water increased by 5.50 °C when $2.40 \times 10^{-3}$ mol of pure ethanol was burned in air.

Use the student's results to calculate a value, in kJ $mol^{-1}$, for the enthalpy change when one mole of ethanol is burned.
(The specific heat capacity of water is 4.18 J $K^{-1}$ $g^{-1}$)

Deduce **two** reasons why the student's value for the standard enthalpy of combustion of ethanol is different from a Data Book value of –1279 kJ $mol^{-1}$.

**(5)**

(c) Mean bond enthalpies can be used to calculate enthalpies of reaction.

(i) Give the meaning of the term **mean bond enthalpy**.



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**(2)**

(ii) Consider the mean bond enthalpy data in the following table.

| | C—H | C—C | C—O | O=O | C=O | O—H |
|---|---|---|---|---|---|---|
| **Mean bond enthalpy / kJ mol$^{-1}$** | 412 | 348 | 360 | to be calculated | 805 | 463 |

Use the data in the table above and the equation shown to calculate a value for the bond enthalpy for the O=O double bond in an oxygen molecule.

$CH_3CH_2OH(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(g)$ $\Delta H = -1279$ kJ mol$^{-1}$

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**(3)**

**(Total 15 marks)**

## Q12.

Ethanol is an important industrial compound.

(a) Ethanol can be produced by the hydration of ethene.
The equation for the equilibrium that is established is

$H_2C=CH_2(g) + H_2O(g) \rightleftharpoons CH_3CH_2OH(g)$ $\Delta H = -42$ kJ mol$^{-1}$

The operating conditions for the process are a temperature of 300 °C and a pressure of 7 MPa.
Under these conditions, the conversion of ethene into ethanol is 5%.

(i) Identify the catalyst used in this process.
Deduce how an overall yield of 95% is achieved in this process without changing the operating conditions.

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**(2)**

(ii) Use your knowledge of equilibrium reactions to explain why a manufacturer might consider using an excess of steam in this process, under the same operating conditions.

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**(3)**

(iii) At pressures higher than 7 MPa, some of the ethene reacts to form a solid with a relative molecular mass greater than 5000.

Deduce the identity of this solid.

Give **one** other reason for **not** operating this process at pressures higher than 7 MPa.
Do **not** include safety reasons.

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**(2)**

(b) Write an equation for the reaction that has an enthalpy change that is the standard enthalpy of formation of ethanol.

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**(2)**

(c) When ethanol is used as a fuel, it undergoes combustion.

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

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**(3)**

(ii) Consider these bond enthalpy data.

| | C–H | C–C | C–O | O=O | C=O | O–H |
|---|---|---|---|---|---|---|
| **Bond enthalpy / kJ mol$^{-1}$** | 412 | 348 | 360 | 496 | 805 | 463 |

Use these data and the equation to calculate a value for the enthalpy of combustion of gaseous ethanol.

$$CH_3CH_2OH(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(g)$$

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**(3)**

(d) Gaseous ethanol can be used to convert hot copper(II) oxide into copper.

(i) Deduce the role of ethanol in this reaction.

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**(1)**

(ii) Draw the structure of the organic compound with $M_r$ = 60 that is produced in this reaction.

**(1)**

**(Total 17 marks)**

## Q13.

Hydrazine ($N_2H_4$) decomposes in an exothermic reaction. Hydrazine also reacts exothermically with hydrogen peroxide when used as a rocket fuel.

(a) Write an equation for the decomposition of hydrazine into ammonia and nitrogen only.

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**(1)**

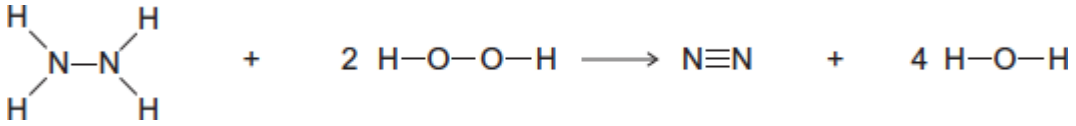
(b) State the meaning of the term *mean bond enthalpy*.

**(2)**

(c) Some mean bond enthalpies are given in the table.

| | N–H | N–N | N≡N | O–H | O–O |
|---|---|---|---|---|---|
| Mean bond enthalpy / kJ $mol^{-1}$ | 388 | 163 | 944 | 463 | 146 |

Use these data to calculate the enthalpy change for the gas-phase reaction between hydrazine and hydrogen peroxide.



**(3)**

**(Total 6 marks)**

## Q14.

This question is about bond dissociation enthalpies and their use in the calculation of enthalpy changes.

(a) Define *bond dissociation enthalpy* as applied to chlorine.



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**(2)**

(b) Explain why the enthalpy of atomisation of chlorine is exactly half the bond dissociation enthalpy of chlorine.

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**(1)**

(c) The bond dissociation enthalpy for chlorine is +242 kJ $mol^{-1}$ and that for fluorine is +158 kJ $mol^{-1}$. The standard enthalpy of formation of ClF(g) is −56 kJ $mol^{-1}$.

(i) Write an equation, including state symbols, for the reaction that has an enthalpy change equal to the standard enthalpy of formation of gaseous ClF

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**(1)**

(ii) Calculate a value for the bond enthalpy of the Cl – F bond.

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**(2)**

(iii) Calculate the enthalpy of formation of gaseous chlorine trifluoride, $ClF_3$(g). Use the bond enthalpy value that you obtained in part (c)(ii).

(If you have been unable to obtain an answer to part (c)(ii), you may assume that the Cl – F bond enthalpy is +223 kJ $mol^{-1}$. This is **not** the correct value.)

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**(3)**



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(iv) Explain why the enthalpy of formation of $ClF_3(g)$ that you calculated in part (c) (iii) is likely to be different from a data book value.

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**(1)**

(d) Suggest why a value for the Na – Cl bond enthalpy is **not** found in any data book.

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**(1)**

**(Total 11 marks)**

## Q15.

Ammonia can be manufactured by the Haber Process.

The equation for the reaction that occurs is shown below.

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

(a) The table below contains some bond enthalpy data.

| | N ≡ N | H–H | N–H |
|---|---|---|---|
| Mean bond enthalpy / kJ $mol^{-1}$ | 944 | 436 | 388 |

(i) Use data from the table to calculate a value for the enthalpy of formation for one mole of ammonia.

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**(3)**

(ii) A more accurate value for the enthalpy of formation of ammonia is –46 kJ $mol^{-1}$.
Suggest why your answer to part (a) (i) is different from this value.

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**(1)**



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(b) The table below contains some entropy data.

| | $H_2(g)$ | $N_2(g)$ | $NH_3(g)$ |
|---|---|---|---|
| $S\ominus$ / J $K^{-1}$ $mol^{-1}$ | 131 | 192 | 193 |

Use these data to calculate a value for the entropy change, with units, for the formation of one mole of ammonia from its elements.

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**(3)**

(c) The synthesis of ammonia is usually carried out at about 800 K.

(i) Use the $\Delta H$ value of –46 kJ $mol^{-1}$ and your answer from part (b) to calculate a value for $\Delta G$, with units, for the synthesis at this temperature.
(If you have been unable to obtain an answer to part (b), you may assume that the entropy change is –112 J $K^{-1}$ $mol^{-1}$. This is not the correct answer.)

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**(3)**

(ii) Use the value of $\Delta G$ that you have obtained to comment on the feasibility of the reaction at 800 K.

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**(1)**

**(Total 11 marks)**

## Q16.

Glucose, produced during photosynthesis in green plants, is a renewable source from which ethanol can be made. Ethanol is a liquid fuel used as a substitute for petrol.
The processes involved can be summarised as follows.

Process **1** Photosynthesis in green plants

$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$



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Process **2** Fermentation of glucose to form ethanol

Process **3** Complete combustion of ethanol
$CH_3CH_2OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$

(a) State **three** essential conditions for the fermentation of aqueous glucose in Process **2**.

Write an equation for the reaction that takes place during this fermentation.

**(4)**

(b) It has been claimed that there is no net carbon (greenhouse gas) emission to the atmosphere when ethanol made by Process **2** is used as a fuel.

State the term that is used to describe fuels of this type.

Use the equations for Processes **1, 2** and **3** to show why it can be claimed that there is no net emission of carbon-containing greenhouse gases.

**(3)**

(c) Use the information from the equation for Process **3** above and the mean bond enthalpies from the table below to calculate a value for the enthalpy change for this process.

| | C–H | C–C | C–O | O–H | C=O | O=O |
|---|---|---|---|---|---|---|
| Mean bond enthalpy / kJ $mol^{-1}$ | +412 | +348 | +360 | +463 | +743 | +496 |



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Give **one** reason why the value calculated from mean bond enthalpies is different from the value given in a data book.

**(4)**

(d) A student carried out a simple laboratory experiment to measure the enthalpy change for Process **3**. The student showed that the temperature of 200 g of water increased by 8.0 °C when 0.46 g of pure ethanol was burned in air and the heat produced was used to warm the water.

Use these results to calculate the value, in kJ $mol^{-1}$, obtained by the student for this enthalpy change. (The specific heat capacity of water is 4.18 J $K^{-1}$ $g^{-1}$)

Give **one** reason, other than heat loss, why the value obtained from the student's results is less exothermic than a data book value.

**(4)**

**(Total 15 marks)**

## Q17.

Hydrogen gas is used in the chemical industry.

(a) Tungsten is extracted by passing hydrogen over heated tungsten oxide ($WO_3$).

(i) State the role of the hydrogen in this reaction.



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**(1)**

(ii) Write an equation for this reaction.

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**(1)**

(iii) State **one** risk of using hydrogen gas in metal extractions.

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**(1)**

(b) Hydrogen is used to convert oleic acid into stearic acid as shown by the following equation.




oleic acid stearic acid

(i) Use your knowledge of the chemistry of alkenes to deduce the type of reaction that has occurred in this conversion.

_____

**(1)**

(ii) State the type of stereoisomerism shown by oleic acid.

_____

**(1)**

(c) Hydrogen reacts with nitrogen in the Haber Process. The equation for the equilibrium that is established is shown below.

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

(i) State Le Chatelier's principle.

_____

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**(1)**

(ii) Use Le Chatelier's principle to explain why an increase in the total pressure of this equilibrium results in an increase in the equilibrium yield of ammonia.

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**(2)**

(d) Hydrogen reacts with oxygen in an exothermic reaction as shown by the following equation.

$$H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(g) \qquad \Delta H = -242 \text{ kJ mol}^{-1}$$

Use the information in the equation and the data in the following table to calculate a value for the bond enthalpy of the H–H bond.

| | O–H | O=O |
|---|---|---|
| Mean bond enthalpy / kJ mol$^{-1}$ | + 463 | + 496 |

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**(3)**

**(Total 11 marks)**

## Q18.

A method of synthesising ammonia directly from nitrogen and hydrogen was developed by Fritz Haber. On an industrial scale, this synthesis requires a high temperature, a high pressure and a catalyst and is very expensive to operate.

(a) Use the data given below to calculate a value for the enthalpy of formation of ammonia

| Bond | N ≡N | H – H | N – H |
|---|---|---|---|
| Mean bond enthalpy/kJ mol$^{-1}$ | 945 | 436 | 391 |

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**(3)**

(b) A manager in charge of ammonia production wished to increase the daily production



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of ammonia and reduce the production costs. How would a chemist explain the factors that would influence the commercial efficiency of this production process?

**(8)**

**(Total 11 marks)**

## Q19.

(a) Define the term *standard enthalpy of combustion*, $\Delta H_c$⦵

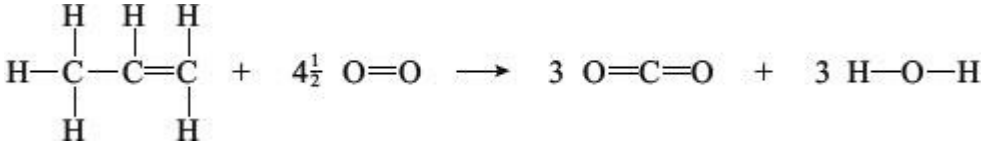
**(3)**

(b) Use the mean bond enthalpy data from the table and the equation given below to calculate a value for the standard enthalpy of combustion of propene. All substances are in the gaseous state.

| Bond | C == C | C—C | C—H | O == O | O == C | O—H |
|---|---|---|---|---|---|---|
| Mean bond enthalpy/ kJ $mol^{-1}$ | 612 | 348 | 412 | 496 | 743 | 463 |



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**(3)**

(c) State why the standard enthalpy of formation, $\Delta H_f^o$, of oxygen is zero.

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**(1)**

(d) Use the data from the table below to calculate a more accurate value for the standard enthalpy of combustion of propene.

| Compound | $C_3H_6(g)$ | $CO_2(g)$ | $H_2O(g)$ |
|---|---|---|---|
| Standard enthalpy of formation, $\Delta H_f^o$/ kJ mol$^{-1}$ | +20 | –394 | –242 |

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**(3)**

(e) Explain why your answer to part (b) is a less accurate value than your answer to part (d).

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**(2)**

**(Total 12 marks)**

## Q20.

(a) Explain the meaning of the terms *mean bond enthalpy* and *standard enthalpy of formation*.

*Mean bond enthalpy* _____

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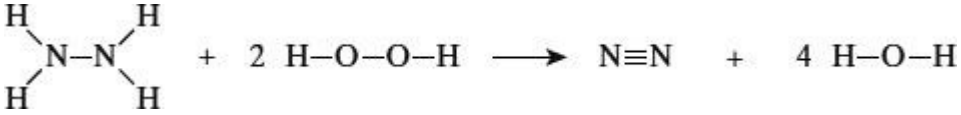
*Standard enthalpy of formation* _____

**(5)**

(b) Some mean bond enthalpies are given below.

| Bond | N–H | N–N | N≡N | H–O | O–O |
|---|---|---|---|---|---|
| Mean bond enthalpy/kJ mol$^{-1}$ | 388 | 163 | 944 | 463 | 146 |

Use these data to calculate the enthalpy change for the following gas-phase reaction between hydrazine, $N_2H_4$, and hydrogen peroxide, $H_2O_2$



**(3)**

(c) Some standard enthalpies of formation are given below.

| | $N_2H_4(g)$ | $H_2O_2(g)$ | $H_2O(g)$ |
|---|---|---|---|
| $\Delta H_f^\circ$ /kJ mol$^{-1}$ | +75 | –133 | –242 |

These data can be used to calculate the enthalpy change for the reaction in part (b).

$$N_2H_4(g) + 2H_2O_2(g) \rightarrow N_2(g) + 4H_2O(g)$$

(i) State the value of $\Delta H_f^\ominus$ for $N_2(g)$.

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(ii) Use the $\Delta H_f^\ominus$ values from the table to calculate the enthalpy change for this reaction.

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**(4)**

(d) Explain why the value obtained in part (b) is different from that obtained in part (c) (ii).

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**(1)**

**(Total 13 marks)**

## Q21.

(a) The table below contains some mean bond enthalpy data.

| Bond | H–O | O–O | O=O |
|---|---|---|---|
| Mean bond enthalpy/kJ mol$^{-1}$ | 463 | 146 | 496 |

The bonding in hydrogen peroxide, $H_2O_2$, can be represented by H–O–O–H. Use these data to calculate the enthalpy change for the following reaction.

$$H_2O_2(g) \rightarrow H_2O_2(g) + \tfrac{1}{2} O_2(g)$$

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**(3)**

(b) The standard enthalpy of formation, $\Delta H_f^{\ominus}$ for methane, is –74.9 kJ mol$^{-1}$. Write an equation, including state symbols, for the reaction to which this enthalpy change applies.

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**(2)**

(c) The enthalpy changes for the formation of atomic hydrogen and atomic carbon from their respective elements in their standard states are as follows.

$\tfrac{1}{2} H_2(g) \rightarrow H(g)$ $\quad \Delta H^{\ominus} = +218$ kJ mol$^{-1}$

$C(s) \rightarrow C(g)$ $\quad \Delta H^{\ominus} = +715$ kJ mol$^{-1}$

(i) By reference to its structure, suggest why a large amount of heat energy is required to produce free carbon atoms from solid carbon.



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(ii) Parts (b) and (c) give enthalpy data for the formation of $CH_4(g)$, H(g) and C(g). Use these data and Hess's Law to calculate the value of the enthalpy change for the following reaction.

$$CH_4(g) \rightarrow C(g) + 4H(g)$$

(iii) Use your answer from part (c)(ii) to calculate a value for the mean bond enthalpy of a C–H bond in methane.

**(5)**

**(Total 10 marks)**

## Q22.

The table below contains some mean bond enthalpy data.

| Bond | H—H | C—C | C=C | N≡N | N—H |
|---|---|---|---|---|---|
| Mean bond enthalpy / kJ $mol^{-1}$ | 436 | 348 | 612 | 944 | 388 |

(a) Explain the term *mean bond enthalpy.*

**(2)**

(b) (i) Write an equation for the formation of one mole of ammonia, $NH_3$, from its elements.

(ii) Use data from the table above to calculate a value for the enthalpy of formation of ammonia.



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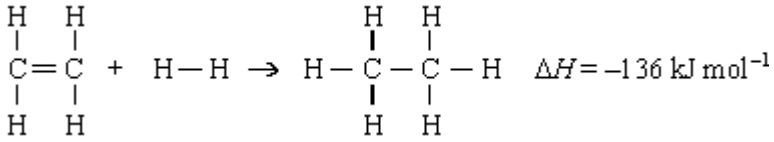
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**(4)**

(c) Use the following equation and data from the table above to calculate a value for the C–H bond enthalpy in ethane.




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**(3)**

**(Total 9 marks)**

## Q23.

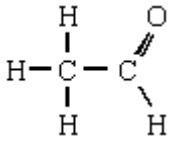
(a) State what is meant by the term *mean bond enthalpy*.

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**(2)**

(b) Ethanal has the structure 

Gaseous ethanal burns as shown by the equation

$$CH_3CHO(g) + 2\frac{1}{2}O_2(g) \rightarrow 2H_2O(g) + 2CO_2(g)$$

Use the mean bond enthalpy data given below to answer the following questions.

| Bond | Mean bond enthalpy/kJ mol$^{-1}$ |
|---|---|
| C—H | +413 |
| C—C | +347 |
| C==O | +736 |



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| | |
|---|---|
| O==O | +498 |
| O—H | +464 |

(i) Calculate the enthalpy change which occurs when all the bonds in the reactants shown in the above equation are broken.

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(ii) Calculate the enthalpy change which occurs when all the bonds in the products shown in the above equation are formed.

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(iii) Hence, calculate the enthalpy change for the complete combustion of ethanal as shown in the equation above.

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**(5)**

**(Total 7 marks)**

## Q24.

Given the following data

$C(s) + 2H_2(g) \rightarrow CH_4(g)$ $\Delta H = -75$ kJ mol$^{-1}$

$H_2(g) \rightarrow 2H(g)$ $\Delta H = +436$ kJ mol$^{-1}$

which one of the following is the enthalpy change, in kJ mol$^{-1}$, of the reaction below?

$CH_4(g) \rightarrow C(s) + 4H(g)$

**A** −947

**B** +511

**C** +797

**D** +947



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**(Total 1 mark)**

## Q1.

The data below refer to the industrial production of nitric acid from ammonia.

*Reaction 1* $4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(g)$ $\Delta H^{\ominus} = -909$ kJ mol$^{-1}$

*Reaction 2* $2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$ $\Delta H^{\ominus} = -115$ kJ mol$^{-1}$

*Reaction 3* $3NO_2(g) + H_2O(l) \rightleftharpoons 2HNO_3(aq) + NO(g)$ $\Delta H^{\ominus} = -117$ kJ mol$^{-1}$

The direct oxidation of ammonia to nitrogen dioxide can be represented by the equation

$$4NH_3(g) + 7O_2(g) \rightarrow 4NO_2(g) + 6H_2O(g)$$

for which the standard enthalpy change, in kJ mol$^{-1}$, is

**A** −1139

**B** −1024

**C** −794

**D** −679

**(Total 1 mark)**

## Q2.

(a) Write an equation for the complete combustion of propanone, $C_3H_6O$, to form carbon dioxide and water.

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**(1)**

(b) In a laboratory experiment, 1.45 g of propanone were burned completely in oxygen. The heat from this combustion was used to raise the temperature of 100 g of water from 293.1 K to 351.2 K.

(i) Calculate the number of moles of propanone in the 1.45 g.

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(ii) Calculate the heat energy required to raise the temperature of 100 g of water from 293.1 K to 351.2 K.
(The specific heat capacity of water is 4.18 J K$^{-1}$ g$^{-1}$)

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(iii) Hence, calculate a value, in kJ $mol^{-1}$, for the enthalpy of combustion of propanone.

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**(5)**

(c) In a similar experiment, the enthalpy of combustion of butanone, $C_4H_8O$, was found to be –1290 kJ $mol^{-1}$. A data book value for the same reaction is $\Delta H_c^{\ominus}$ = –2430 kJ $mol^{-1}$.

(i) Suggest one reason why the experimental value is very different from the data book value.

_____

(ii) This data book value of $\Delta H_c^{\ominus}$ for butanone (–2430 kJ $mol^{-1}$) refers to the formation of carbon dioxide gas and water in the gaseous state. How would this value differ if it referred to the formation of water in the liquid state? Explain your answer.

*Difference* _____

*Explanation* _____

_____

**(3)**

(d) Calculate a value for the standard enthalpy of formation for liquid ethanethiol, $C_2H_5SH$. Use the equation given below and enthalpy of combustion data from the following table.

| Substance | $C_2H_5SH(l)$ | C(s) | $H_2(g)$ | S(s) |
|---|---|---|---|---|
| $\Delta H_c^{\ominus}$ / kJ $mol^{-1}$ | –1170 | –394 | –286 | –297 |

$2C(s) + 3H_2(g) + S(s) \rightarrow C_2H_5SH(l)$

_____

_____

_____

_____

**(3)**

**(Total 12 marks)**

## Q3.

Using the information below, answer this question.

$Fe_2O_3(s) + 3H_2(g) \rightarrow 2Fe(s) + 3H_2O(g)$ $\Delta H^{\ominus}$ = +96 kJ $mol^{-1}$, $\Delta S^{\ominus}$ = +138 J $K^{-1}$ $mol^{-1}$



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| | $Fe_2O_3(s)$ | $H_2(g)$ | Fe(s) |
|---|---|---|---|
| $\Delta H_f^{\ominus}$ / kJ mol$^{-1}$ | −822.0 | 0 | 0 |
| $\Delta S^{\ominus}$ / J K$^{-1}$ mol$^{-1}$ | 90.0 | 131.0 | 27.0 |

The standard enthalpy of formation of steam is

**A** +286 kJ mol$^{-1}$

**B** +242 kJ mol$^{-1}$

**C** −242 KJ mol$^{-1}$

**D** −286 kJ mol$^{-1}$

**(Total 1 mark)**

## Q4.

The table below contains some standard enthalpy of formation data.

| Substance | C(s) | $N_2(g)$ | $H_2O(g)$ | $CO_2(g)$ | $NH_4NO_3(s)$ |
|---|---|---|---|---|---|
| $\Delta H_f^{\ominus}$ / kJ mol$^{-1}$ | 0 | 0 | –242 | –394 | –365 |

(a) Why are the values of the standard enthalpy of formation for carbon and nitrogen zero?

_____

**(1)**

(b) State Hess's Law.

_____

_____

_____

**(2)**

(c) Use $\Delta H_f^{\ominus}$ data from the table to calculate a value for the enthalpy change for the following reaction.

$$NH_4NO_3(s) + \frac{1}{2}C(s) \rightarrow N_2(g) + 2H_2O(g) + \frac{1}{2}CO_2(g)$$

_____

_____

_____



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_____

**(3)**

**(Total 6 marks)**

## Q5.

(a) What is the meaning of the term *enthalpy change*?

_____

_____

_____

**(2)**

(b) (i) Define the term *standard enthalpy of formation* of a compound.

_____

_____

_____

_____

(ii) Write an equation, including state symbols, for the formation from its elements of solid sodium sulphate, $Na_2SO_4$

_____

**(5)**

(c) State Hess's Law.

_____

_____

**(1)**

(d) Some standard enthalpy changes are difficult to measure directly but can be determined from standard enthalpies of combustion.
Maleic acid, $C_4H_4O_4$, reacts with oxygen to form carbon dioxide and water as shown by the following equation.

$$C_4H_4O_4(s) + 3O_2(g) \rightarrow 4CO_2(g) + 2H_2O(l)$$

Use the standard enthalpy of combustion data given below to calculate a value for the standard enthalpy change for the following reaction.

$$4C(s) + 2H_2(g) + 2O_2(g) \rightarrow C_4H_4O_4(s)$$

| | $C_4H_4O_4(s)$ | C(s) | $H_2(g)$ |
|---|---|---|---|
| $\Delta H_c^{\ominus}$ / kJ mol$^{-1}$ | –1356 | –393.5 | –285.8 |



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**(3)**
**(Total 11 marks)**

## Q6.

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

**(3)**

(b) State Hess's law.

**(1)**

(c) Propanone, $CH_3COCH_3$, burns in oxygen as shown by the equation

$$CH_3COCH_3(l) + 4\ O_2(g) \rightarrow 3H_2O(l) + 3CO_2(g)$$

Use the data given below to calculate the standard enthalpy of combustion of propanone.

| | $CO_2(g)$ | $H_2O(l)$ | $CH_3COCH_3(l)$ |
|---|---|---|---|
| $\Delta H_f^{\ominus}$/kJ mol$^{-1}$ | –394 | –286 | –248 |



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**(3)**
**(Total 7 marks)**

## Q7.

Using the data below, which is the correct value for the standard enthalpy of formation for $TiCl_4(l)$?

$C(s) + TiO_2(s) + 2Cl_2(g) \rightarrow TiCl_4(l) + CO_2(g)$ $\quad \Delta H^{\ominus} = -232$ kJ mol$^{-1}$

$Ti(s) + O_2(g) \rightarrow TiO_2(s)$ $\quad \Delta H_f^{\ominus} = -912$ kJ mol$^{-1}$

$C(s) + O_2(g) \rightarrow CO_2(g)$ $\quad \Delta H_f^{\ominus} = -394$ kJ mol$^{-1}$

**A** $-1538$ kJ mol$^{-1}$

**B** $-1094$ kJ mol$^{-1}$

**C** $-750$ kJ mol$^{-1}$

**D** $+286$ kJ mol$^{-1}$

**(Total 1 mark)**

## Q8.

When ethanamide ($CH_3CONH_2$) burns in oxygen the carbon is converted into carbon dioxide, the hydrogen is converted into water and the nitrogen forms nitrogen gas.

| Substance | ethanamide | carbon dioxide | water |
|---|---|---|---|
| Enthalpy of formation ($\Delta H_f^{\ominus}$) / kJ mol$^{-1}$ | −320 | −394 | −286 |

Using the data above, which one of the following is a correct value for the enthalpy of combustion of ethanamide?

**A** $-1823$ kJ mol$^{-1}$

**B** $-1183$ kJ mol$^{-1}$

**C** $-1000$ kJ mol$^{-1}$

**D** $-360$ kJ mo1$^{-1}$

**(Total 1 mark)**

## Q9.

Consider the reactions

$C_2H_4(g) + 2O_2(g) \rightarrow 2CO(g) + 2H_2O(g)$ $\quad \Delta H^{\ominus} = -758$ kJ mol$^{-1}$



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$2C(s) + 2H_2(g) \rightarrow C_2H_4(g)$ $\Delta H^{\ominus} = +52$ kJ mol$^{-1}$

$H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(g)$ $\Delta H^{\ominus} = -242$ kJ mol$^{-1}$

The enthalpy of formation of carbon monoxide is

**A** −111 kJ mol$^{-1}$

**B** −163 kJ mol$^{-1}$

**C** −222 kJ mol$^{-1}$

**D** -464 kJ mol$^{-1}$

**(Total 1 mark)**

## Q10.

Given the following data

$C(s) + 2H_2(g) \rightarrow CH_4(g)$ $\Delta H = -75$ kJ mol$^{-1}$

$H_2(g) \rightarrow 2H(g)$ $\Delta H = +436$ kJ mol$^{-1}$

which one of the following is the enthalpy change, in kJ mol$^{-1}$, of the reaction below?

$CH_4(g) \rightarrow C(s) + 4H(g)$

**A** −947

**B** +511

**C** +797

**D** +947

**(Total 1 mark)**

## Q11.

Nitric acid is produced industrially from ammonia, air and water using the following sequence of reactions:

(1) $4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$ $\Delta H = -909$ kJ mol$^{-1}$

(2) $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$ $\Delta H = -115$ kJ mol$^{-1}$

(3) $3NO_2(g) + H_2O(l) \rightarrow 2HNO_3(aq) + NO(g)$ $\Delta H = -117$ kJ mol$^{-1}$

Which is the enthalpy change (in kJ mol$^{-1}$) for the following reaction?

$4NH_3(g) + 7O_2(g) \rightarrow 4NO_2(g) + 6H_2O(g)$

**A** −679

**B** −794



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**C** −1024

**D** −1139

**(Total 1 mark)**



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## Q1.

Chlorine is used to decrease the numbers of microorganisms in water.

When chlorine is added to water, there is a redox reaction, as shown by the equation

$$Cl_2 + H_2O \rightleftharpoons HClO + HCl$$

(a) Deduce the oxidation state of chlorine in HClO and the oxidation state of chlorine in HCl

Oxidation state of chlorine in HClO _____

Oxidation state of chlorine in HCl _____

**(1)**

(b) Give two half-equations to show the oxidation and reduction processes that occur in this redox reaction.

Oxidation half-equation _____

Reduction half-equation _____

**(2)**

(c) Chlorine is reacted with cold, aqueous sodium hydroxide in the manufacture of bleach.

Give an equation for this reaction between chlorine and sodium hydroxide.

_____

**(1)**

(d) The concentration of $ClO^-$ ions in bleach solution can be found by reaction with iodide ions.

The overall equation for this reaction is shown.

$$ClO^- + 2I^- + 2H^+ \rightarrow I_2 + Cl^- + H_2O$$

A sample of bleach solution was found to contain $ClO^-$ ions with a concentration of 0.0109 mol $dm^{-3}$

Potassium iodide is added to a 20.0 $cm^3$ portion of this bleach solution.

Calculate the mass, in mg, of potassium iodide needed to react with all of the $ClO^-$ ions in the sample of bleach.
Give your answer to the appropriate number of significant figures.
Give **one** observation during this reaction.

Mass of potassium iodide _____ mg

Observation _____

_____

_____



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_____

**(4)**

(e) Potassium chlorate(VII), $KClO_4$, is used in fireworks. When potassium chlorate(VII) decomposes, it produces potassium chloride and oxygen.

Give an equation for the decomposition of potassium chlorate(VII).
Use the data in the table to calculate the enthalpy change for this reaction.

| **Substance** | **$\Delta_f H$ / kJ mol$^{-1}$** |
|---|---|
| $KClO_4(s)$ | – 434 |
| KCl(s) | – 436 |

Equation _____

Enthalpy change _____ kJ mol$^{-1}$

**(2)**

**(Total 10 marks)**

## Q2.

What is the enthalpy of formation of buta-1,3-diene, $C_4H_6(g)$?

| **Substance** | **Enthalpy of combustion / kJ mol$^{-1}$** |
|---|---|
| $C_4H_6(g)$ | –2546 |
| C(s) | –394 |
| $H_2(g)$ | –286 |



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**A** +112 kJ mol$^{-1}$ ⬭

**B** –112 kJ mol$^{-1}$ ⬭

**C** +746 kJ mol$^{-1}$ ⬭

**D** –746 kJ mol$^{-1}$ ⬭

**(Total 1 mark)**

## Q3.

This question is about energetics.

(a) Write an equation, including state symbols, for the reaction with an enthalpy change equal to the enthalpy of formation for iron(III) oxide.

_____

**(1)**

(b) **Table 1** contains some standard enthalpy of formation data.

**Table 1**

| | CO(g) | $Fe_2O_3$(s) |
|---|---|---|
| **$\Delta_f H^\ominus$/ kJ mol$^{-1}$** | −111 | −822 |

$Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(s) + 3CO_2(g)$ $\Delta H = -19$ kJ mol$^{-1}$

Use these data and the equation for the reaction of iron(III) oxide with carbon monoxide to calculate a value for the standard enthalpy of formation for carbon dioxide.

Show your working.

$\Delta_f H^\ominus$ _____ kJ mol$^{-1}$

**(3)**

(c) Some enthalpy data are given in **Table 2**.



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**Table 2**

| **Process** | **$\Delta H$ / kJ mol$^{-1}$** |
|---|---|
| $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$ | −92 |
| $N_2(g) \longrightarrow 2N(g)$ | +944 |
| $H_2(g) \longrightarrow 2H(g)$ | +436 |

Use the data from **Table 2** to calculate the bond enthalpy for N−H in ammonia.

N−H bond enthalpy _____ kJ mol$^{-1}$

**(3)**

(d) Give one reason why the bond enthalpy that you calculated in part (c) is different from the mean bond enthalpy quoted in a data book (388 kJ mol$^{-1}$).

_____

_____

_____

**(1)**

**(Total 8 marks)**

## Q4.

Anhydrous magnesium chloride, $MgCl_2$, can absorb water to form the hydrated salt $MgCl_2.4H_2O$

$$MgCl_2(s) + 4H_2O(l) \longrightarrow MgCl_2.4H_2O(s)$$

(a) Suggest **one** reason why the enthalpy change for this reaction cannot be determined directly by calorimetry.

_____

_____

**(1)**



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(b) Some enthalpies of solution are shown in Table 1.

**Table 1**

| Salt | Enthalpy of solution / kJ mol$^{-1}$ |
|---|---|
| $MgCl_2(s)$ | −155 |
| $MgCl_2.4H_2O(s)$ | −39 |

Calculate the enthalpy change for the absorption of water by $MgCl_2(s)$ to form $MgCl_2.4H_2O(s)$.

Enthalpy change _____ kJ mol$^{-1}$

**(2)**

(c) Describe how you would carry out an experiment to determine the enthalpy of solution of anhydrous magnesium chloride.

You should use about 0.8 g of anhydrous magnesium chloride.

Explain how your results could be used to calculate the enthalpy of solution.

_____

_____

_____

_____

_____

_____

_____

_____

**(6)**



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(d) Anhydrous magnesium chloride can be formed by direct reaction between its elements.

$$Mg(s) + Cl_2(g) \longrightarrow MgCl_2(s)$$

The free-energy change, $\Delta G$, for this reaction varies with temperature as shown in **Table 2**.

**Table 2**

| T / K | $\Delta G$ / kJ mol$^{-1}$ |
|---|---|
| 298 | −592.5 |
| 288 | −594.2 |
| 273 | −596.7 |
| 260 | −598.8 |
| 240 | −602.2 |

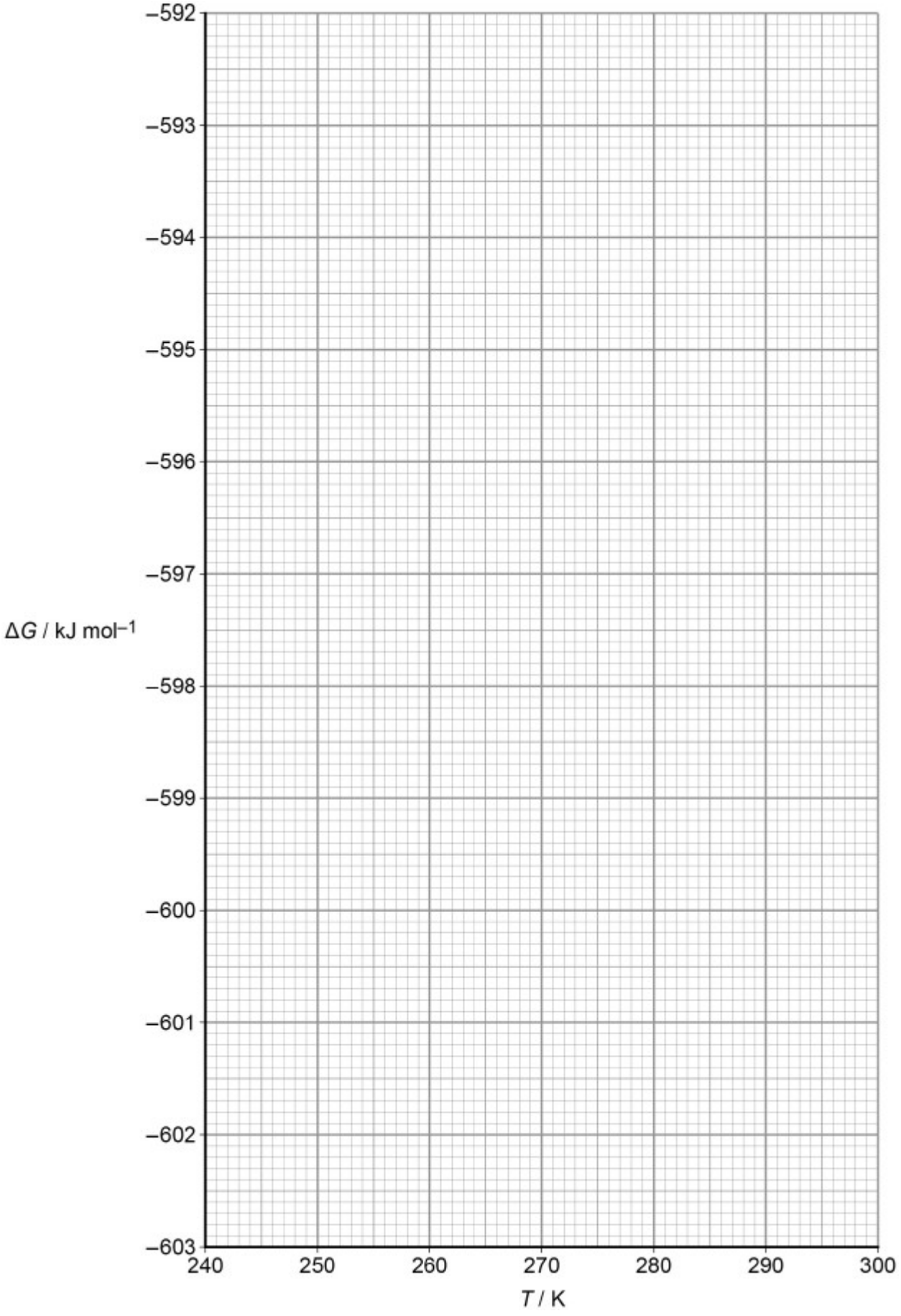
Use these data to plot a graph of free-energy change against temperature on the grid below.

Calculate the gradient of the line on your graph and hence calculate the entropy change, $\Delta S$, in J K$^{-1}$ mol$^{-1}$, for the formation of anhydrous magnesium chloride from its elements.

Show your working.



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ΔS _____ $JK^{-1}mol^{-1}$



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**(5)**
**(Total 14 marks)**

## Q5.

This question is about enthalpy changes.

(a) Write an equation, including state symbols, to show the reaction taking place when the standard enthalpy of combustion for ethanol is measured.

_____

**(2)**

(b) State the name given to the enthalpy change represented by the following chemical equation.

Explain why this enthalpy change would be difficult to determine directly.

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

Enthalpy change _____

Explanation _____

_____

_____

**(2)**

(c) Standard enthalpies of combustion for carbon and carbon monoxide are −393 kJ $mol^{-1}$ and −283 kJ $mol^{-1}$, respectively.

Use these data to calculate the enthalpy change for the reaction in part **(b)**.

Enthalpy change = _____ kJ $mol^{-1}$

**(2)**

(d) Use the following data to calculate a value for the Xe–F bond enthalpy in $XeF_4$



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$$Xe(g) + 2F_2(g) \rightarrow XeF_4(g) \quad \Delta H = -252 \text{ kJ mol}^{-1}$$
$$F_2(g) \rightarrow 2F(g) \quad \Delta H = +158 \text{ kJ mol}^{-1}$$

Xe–F bond enthalpy = _____ kJ mol$^{-1}$

**(3)**

(e) Suggest a reason why the value calculated in part **(d)** differs from the mean Xe–F bond enthalpy quoted in a data source.

_____

_____

_____

**(1)**

**(Total 10 marks)**

## Q6.

Hydrogen can be manufactured by the reaction of methane with steam. An equilibrium is established as shown by the equation.

$$CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3H_2(g)$$

(a) Use Le Chatelier's principle to predict the effect on the equilibrium yield of hydrogen if the overall pressure is increased.
Explain your answer.

Effect on yield _____

Explanation _____

_____

_____

_____

**(3)**

(b) Explain why the equilibrium yield of hydrogen is unchanged if a catalyst is used in



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the reaction.

_____

_____

_____

_____

_____

**(2)**

(c) The table shows the standard enthalpy of formation and the standard entropy for each substance in this equilibrium reaction.

| | $CH_4(g)$ | $H_2O(g)$ | $CO(g)$ | $H_2(g)$ |
|---|---|---|---|---|
| $\Delta_f H^\theta$ / kJ mol$^{-1}$ | −75 | −242 | −111 | 0 |
| Sθ / J K$^{-1}$ mol$^{-1}$ | 186 | 189 | 198 | 131 |

Use data from the table to calculate the standard enthalpy change for this equilibrium reaction.

Standard enthalpy change _____ kJ mol$^{-1}$

**(2)**

Use your answer from part **(c)** and the entropy data from the table above to calculate the minimum temperature, in °C, needed for this reaction to be feasible. Give your answer to the appropriate number of significant figures.

(If you did not complete part **(c)** you should assume a value of 120 kJ mol$^{-1}$ for the standard enthalpy change. This is **not** the correct value).

Minimum temperature _____ °C

**(5)**

**(Total 12 marks)**

## Q7.

The table shows the standard enthalpy of formation, $\Delta_f H^\theta$, for some of the substances in the reaction

$$C_2H_6(g) + 6F_2(g) \longrightarrow C_2F_6(g) + 6HF(g) \qquad \Delta H^\theta = -2898 \text{ kJ mol}-1$$

| | $C_2H_6(g)$ | $C_2F_6(g)$ |
|---|---|---|
| $\Delta_f H^\theta$ /kJ mol$^{-1}$ | −84 | −1344 |

What is the standard enthalpy of formation, in kJ mol$^{-1}$, for HF(g)?

**A** −1638 ○



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**B** −273 ⬭

**C** +273 ⬭

**D** +1638 ⬭

**(Total 1 mark)**

## Q8.

(a) Write an equation, including state symbols, for the reaction with enthalpy change equal to the standard enthalpy of formation for $CF_4(g)$.

_____

**(1)**

(b) Explain why $CF_4$ has a bond angle of 109.5°.

_____

_____

_____

_____

_____

**(2)**

(c) **Table 1** gives some values of standard enthalpies of formation ($\Delta_f H^\ominus$).

**Table 1**

| **Substance** | $F_2(g)$ | $CF_4(g)$ | HF(g) |
|---|---|---|---|
| **$\Delta_f H^\ominus$ / kJ mol$^{-1}$** | 0 | −680 | −269 |

The enthalpy change for the following reaction is −2889 kJ mol$^{-1}$.

$$C_2H_6(g) + 7F_2(g) \longrightarrow 2CF_4(g) + 6HF(g)$$

Use this value and the standard enthalpies of formation in **Table 1** to calculate the standard enthalpy of formation of $C_2H_6(g)$.

Standard enthalpy of formation of $C_2H_6(g)$ = _____ kJ mol$^{-1}$

**(3)**



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(d) Methane reacts violently with fluorine according to the following equation.

$$CH_4(g) + 4F_2(g) \longrightarrow CF_4(g) + 4HF(g) \quad \Delta H = -1904 \text{ kJ mol}^{-1}$$

Some mean bond enthalpies are given in **Table 2**.

**Table 2**

| **Bond** | C−H | C−F | H−F |
|---|---|---|---|
| **Mean bond enthalpy / kJ mol$^{-1}$** | 412 | 484 | 562 |

A student suggested that one reason for the high reactivity of fluorine is a weak F−F bond.

Is the student correct? Justify your answer with a calculation using these data.

_____

_____

**(4)**

**(Total 10 marks)**

## Q9.

Standard enthalpy of combustion data can be used to calculate enthalpies of formation.

(a) State the meaning of the term standard enthalpy of combustion.

_____

_____

_____

_____

**(3)**

(b) The equation corresponding to the enthalpy of formation of propan-1-ol is shown.

$$3C(s) + 4H_2(g) + \tfrac{1}{2}O_2(g) \longrightarrow CH_3CH_2CH_2OH(l)$$

**Table 1** contains some standard enthalpy of combustion data.

**Table 1**

| | **C(s)** | **$H_2$(g)** | **$CH_3CH_2CH_2OH$(l)** |
|---|---|---|---|



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| | | | |
|---|---|---|---|
| $\Delta H_c^{\ominus}$ / kJ mol$^{-1}$ | –394 | –286 | –2010 |

Use data from **Table 1** to calculate a value for the standard enthalpy of formation of propan-1-ol. Show your working.

**(3)**

(c) An equation for the complete combustion of gaseous propan-1-ol is shown.

$$CH_3CH_2CH_2OH(g) + 4\tfrac{1}{2}O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(g) \quad \Delta H = -1893 \text{ kJ mol}^{-1}$$

**Table 2** shows some bond enthalpy data.

**Table 2**

| | **C–H** | **C–O** | **O–H** | **C=O** | **O=O** |
|---|---|---|---|---|---|
| Bond enthalpy / kJ mol$^{-1}$ | 412 | 360 | 463 | 805 | 496 |

Use data from **Table 2** and the enthalpy change for this reaction to calculate a value for the bond enthalpy of a C–C bond in propan-1-ol.

**(3)**



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**(Total 9 marks)**

## Q10.

Vanadium is an important metal. Ferrovanadium, an alloy of iron and vanadium, is used to make a strong type of vanadium-steel. Pure vanadium is used in nuclear reactors.

(a) The table shows some standard enthalpy of formation data.

| | $V_2O_5(s)$ | CaO(s) |
|---|---|---|
| $\Delta H_f^\theta$ / kJ mol$^{-1}$ | −1560 | −635 |

In the oldest method of extraction of vanadium, $V_2O_5$ is reacted with calcium at a high temperature.

$$5Ca(s) + V_2O_5(s) \longrightarrow 2V(s) + 5CaO(s)$$

Use data from the table and the equation to calculate the standard enthalpy change for this reaction.

State the type of reaction that $V_2O_5$ has undergone.

Suggest **one** major reason why this method of extracting vanadium is expensive, other than the cost of heating the reaction mixture.

**(5)**

(b) Ferrovanadium is produced by the reaction of aluminium with a mixture of $V_2O_5$ and iron(III) oxide.

Write an equation for the reaction of aluminium with iron(III) oxide.

State the change in oxidation state of aluminium in this reaction.



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_____

_____

**(2)**

(c) Pure vanadium, for nuclear reactors, is formed by the reaction of hydrogen with purified $VCl_2$

Write an equation for this reaction in which the only other product is HCl gas.

Identify **two** hazards in this process, other than the fact that it operates at a high temperature.

Deduce why this process produces **pure** vanadium, other than the fact that purified $VCl_2$ is used.

_____

_____

_____

_____

_____

_____

_____

_____

**(4)**

**(Total 11 marks)**

## Q11.

Antimony is a solid element that is used in industry. The method used for the extraction of antimony depends on the grade of the ore.

(a) Antimony can be extracted by reacting scrap iron with low-grade ores that contain antimony sulfide ($Sb_2S_3$).

(i) Write an equation for the reaction of iron with antimony sulfide to form antimony and iron(II) sulfide.

_____

**(1)**

(ii) Write a half-equation to show what happens to the iron atoms in this reaction.

_____

**(1)**

(b) In the first stage of the extraction of antimony from a high-grade ore, antimony sulfide is roasted in air to convert it into antimony(III) oxide ($Sb_2O_3$) and sulfur dioxide.



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(i) Write an equation for this reaction.

_____

**(1)**

(ii) Identify **one** substance that is manufactured directly from the sulfur dioxide formed in this reaction.

_____

**(1)**

(c) In the second stage of the extraction of antimony from a high-grade ore, antimony(III) oxide is reacted with carbon monoxide at high temperature.

(i) Use the standard enthalpies of formation in the table and the equation given below the table to calculate a value for the standard enthalpy change for this reaction.

| | **$Sb_2O_3$(s)** | **CO(g)** | **Sb(l)** | **$CO_2$(g)** |
|---|---|---|---|---|
| **$\Delta H_f^{\ominus}$ / kJ mol$^{-1}$** | -705 | -111 | +20 | -394 |

$Sb_2O_3(s) + 3CO(g) \longrightarrow 2Sb(l) + 3CO_2(g)$

_____

_____

_____

_____

_____

**(3)**

(ii) Suggest why the value for the standard enthalpy of formation of liquid antimony, given in the table above, is **not** zero.

_____

_____

**(1)**

(iii) State the type of reaction that antimony(III) oxide has undergone in this reaction.

_____

**(1)**

(d) Deduce **one** reason why the method of extraction of antimony from a low-grade ore, described in part (a), is a low-cost process. Do **not** include the cost of the ore.

_____

_____

**(1)**



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**(Total 10 marks)**

## Q12.

The alcohol 2-methylpropan-2-ol, $(CH_3)_3COH$, reacts to form esters that are used as flavourings by the food industry. The alcohol can be oxidised to produce carbon dioxide and water.

A student carried out an experiment on a pure sample of 2-methylpropan-2-ol to determine its enthalpy of combustion. A sample of the alcohol was placed into a spirit burner and positioned under a beaker containing 50 $cm^3$ of water. The spirit burner was ignited and allowed to burn for several minutes before it was extinguished.

The results for the experiment are shown in **Table 1**.

**Table 1**

| | |
|---|---|
| Initial temperature of the water / °C | 18.1 |
| Final temperature of the water / °C | 45.4 |
| Initial mass of spirit burner and alcohol / g | 208.80 |
| Final mass of spirit burner and alcohol / g | 208.58 |

(a) Use the results from **Table 1** to calculate a value for the heat energy released from the combustion of this sample of 2-methylpropan-2-ol.
The specific heat capacity of water is 4.18 $J\ K^{-1}\ g^{-1}$.
Show your working.

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**(2)**

(b) Calculate the amount, in moles, of 2-methylpropan-2-ol burned in the experiment.
Hence calculate a value, in $kJ\ mol^{-1}$, for the enthalpy of combustion of 2-methylpropan-2-ol.
Show your working.

(If you were unable to calculate an answer to part (a), you should assume that the heat energy released was 5580 J. This is **not** the correct value.)

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**(3)**

(c) An equation for the combustion of 2-methylpropan-2-ol is

$$(CH_3)_3COH(l) + 6O_2(g) \longrightarrow 4CO_2(g) + 5H_2O(l)$$

**Table 2** contains some standard enthalpy of formation data.

**Table 2**

| | $(CH_3)_3COH(l)$ | $O_2(g)$ | $CO_2(g)$ | $H_2O(l)$ |
|---|---|---|---|---|
| $\Delta H_f^{\ominus}$ **/ kJ mol$^{-1}$** | –360 | 0 | –393 | –286 |

Use the data from **Table 2** to calculate a value for the standard enthalpy of combustion of 2-methylpropan-2-ol. Show your working.

**(3)**

(d) An accurate value for the enthalpy of combustion of 2-methylpropan-2-ol in which water is formed as a gas is –2422 kJ mol$^{-1}$.

Use this value and your answer from part (b) to calculate the overall percentage error in the student's experimental value for the enthalpy of combustion of 2-methylpropan-2-ol.

**(1)**

(e) Suggest **one** improvement that would reduce errors due to heat loss in the student's experiment.



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**(1)**

(f) Suggest **one** other source of error in the student's experiment. Do **not** include heat loss, apparatus error or student error.

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**(1)**

**(Total 11 marks)**

## Q13.

Group 2 metals and their compounds are used commercially in a variety of processes.

(a) Strontium is extracted from strontium oxide (SrO) by heating a mixture of powdered strontium oxide and powdered aluminium.

Consider these standard enthalpies of formation.

| | SrO(s) | $Al_2O_3(s)$ |
|---|---|---|
| **$\Delta H_f^{\ominus}$ / kJ mol$^{-1}$** | – 590 | – 1669 |

$$3SrO(s) + 2Al(s) \longrightarrow 3Sr(s) + Al_2O_3(s)$$

Use these data and the equation to calculate the standard enthalpy change for this extraction of strontium.

The use of powdered strontium oxide and powdered aluminium increases the surface area of the reactants.
Suggest **one** reason why this increases the reaction rate.

Suggest **one** major reason why this method of extracting strontium is expensive.

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**(5)**

(b) Explain why calcium has a higher melting point than strontium.

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**(2)**

(c) Magnesium is used in fireworks. It reacts rapidly with oxygen, burning with a bright white light. Magnesium reacts slowly with cold water.

Write an equation for the reaction of magnesium with oxygen.

Write an equation for the reaction of magnesium with cold water.

Give a medical use for the magnesium compound formed in the reaction of magnesium with cold water.

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**(3)**

**(Total 10 marks)**

## Q14.

This question is about the extraction of metals.

(a) Manganese can be extracted from $Mn_2O_3$ by reduction with carbon monoxide at high temperature.



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(i) Use the standard enthalpy of formation data from the table and the equation for the extraction of manganese to calculate a value for the standard enthalpy change of this extraction.

| | $Mn_2O_3(s)$ | $CO(g)$ | $Mn(s)$ | $CO_2(g)$ |
|---|---|---|---|---|
| $\Delta H_f^{\ominus}$/ kJ mol$^{-1}$ | −971 | −111 | 0 | −394 |

$$Mn_2O_3(s) + 3CO(g) \longrightarrow 2Mn(s) + 3CO_2(g)$$

**(3)**

(ii) State why the value for the standard enthalpy of formation of Mn(s) is zero.

**(1)**

(b) Titanium is extracted in industry from titanium(IV) oxide in a two-stage process.

(i) Write an equation for the first stage of this extraction in which titanium(IV) oxide is converted into titanium(IV) chloride.

**(2)**

(ii) Write an equation for the second stage of this extraction in which titanium(IV) chloride is converted into titanium.

**(2)**

(c) Chromium is extracted in industry from chromite ($FeCr_2O_4$).

(i) In the first stage of this extraction, the $FeCr_2O_4$ is converted into $Na_2CrO_4$ Balance the equation for this reaction.

.....$FeCr_2O_4$ + .....$Na_2CO_3$ + .....$O_2$ $\longrightarrow$ .....$Na_2CrO_4$ + $2Fe_2O_3$ + $8CO_2$

**(1)**

(ii) In the final stage, chromium is extracted from $Cr_2O_3$ by reduction with aluminium.

Write an equation for this reaction.



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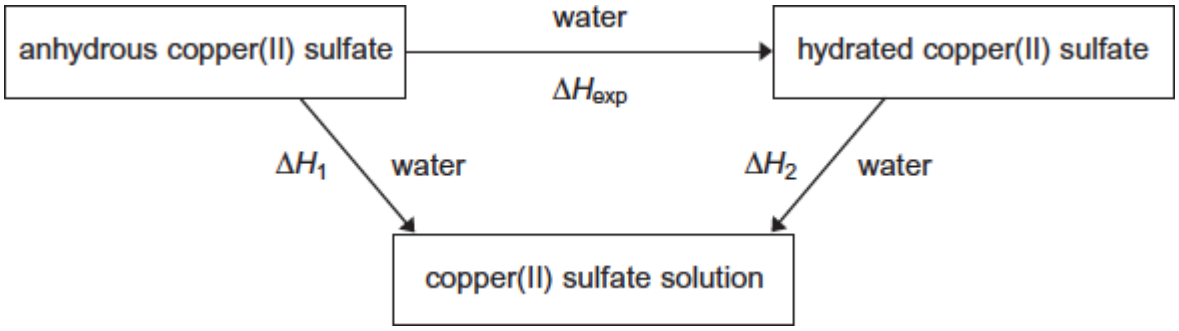
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**(1)**

**(Total 10 marks)**

## Q15.

A student used Hess's Law to determine a value for the enthalpy change that occurs when anhydrous copper(II) sulfate is hydrated. This enthalpy change was labelled $\Delta H_{exp}$ by the student in a scheme of reactions.




(a) State Hess's Law.

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**(1)**

(b) Write a mathematical expression to show how $\Delta H_{exp}$, $\Delta H_1$ and $\Delta H_2$ are related to each other by Hess's Law.

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**(1)**

(c) Use the mathematical expression that you have written in part (b), and the data book values for the two enthalpy changes $\Delta H_1$ and $\Delta H_2$ shown, to calculate a value for $\Delta H_{exp}$

$\Delta H_1 = -156$ kJ mol$^{-1}$
$\Delta H_2 = +12$ kJ mol$^{-1}$

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**(1)**

(d) The student added 0.0210 mol of pure anhydrous copper(II) sulfate to 25.0 cm$^3$ of deionised water in an open polystyrene cup. An exothermic reaction occurred and the temperature of the water increased by 14.0 °C.



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(i) Use these data to calculate the enthalpy change, in kJ $mol^{-1}$, for this reaction of copper(II) sulfate. This is the student value for $\Delta H_1$

In this experiment, you should assume that all of the heat released is used to raise the temperature of the 25.0 g of water. The specific heat capacity of water is 4.18 J $K^{-1}$ $g^{-1}$.

**(3)**

(ii) Suggest **one** reason why the student value for $\Delta H_1$ calculated in part (d)(i) is less accurate than the data book value given in part (c).

**(1)**

(e) Suggest **one** reason why the value for $\Delta H_{exp}$ **cannot** be measured directly.

**(1)**

**(Total 8 marks)**

## Q16.

This question is about bond dissociation enthalpies and their use in the calculation of enthalpy changes.

(a) Define *bond dissociation enthalpy* as applied to chlorine.

**(2)**

(b) Explain why the enthalpy of atomisation of chlorine is exactly half the bond



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dissociation enthalpy of chlorine.

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**(1)**

(c) The bond dissociation enthalpy for chlorine is +242 kJ $mol^{-1}$ and that for fluorine is +158 kJ $mol^{-1}$. The standard enthalpy of formation of ClF(g) is −56 kJ $mol^{-1}$.

(i) Write an equation, including state symbols, for the reaction that has an enthalpy change equal to the standard enthalpy of formation of gaseous ClF

_____

**(1)**

(ii) Calculate a value for the bond enthalpy of the Cl – F bond.

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**(2)**

(iii) Calculate the enthalpy of formation of gaseous chlorine trifluoride, $ClF_3(g)$. Use the bond enthalpy value that you obtained in part (c)(ii).

(If you have been unable to obtain an answer to part (c)(ii), you may assume that the Cl – F bond enthalpy is +223 kJ $mol^{-1}$. This is **not** the correct value.)

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**(3)**

(iv) Explain why the enthalpy of formation of $ClF_3(g)$ that you calculated in part (c) (iii) is likely to be different from a data book value.

_____

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**(1)**

(d) Suggest why a value for the Na – Cl bond enthalpy is **not** found in any data book.

_____

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**(1)**

**(Total 11 marks)**

## Q17.

Methanol ($CH_3OH$) is an important fuel that can be synthesised from carbon dioxide.

(a) The table shows some standard enthalpies of formation.

| | $CO_2(g)$ | $H_2(g)$ | $CH_3OH(g)$ | $H_2O(g)$ |
|---|---|---|---|---|
| $\Delta H_f^{\ominus}$/kJ mol$^{-1}$ | – 394 | 0 | – 201 | – 242 |

(i) Use these standard enthalpies of formation to calculate a value for the standard enthalpy change of this synthesis.

$$CO_2(g) + 3H_2(g) \rightleftharpoons CH_3OH(g) + H_2O(g)$$

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**(3)**

(ii) State why the standard enthalpy of formation for hydrogen gas is zero.

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**(1)**

(b) State and explain what happens to the yield of methanol when the total pressure is increased in this synthesis.

$$CO_2(g) + 3H_2(g) \rightleftharpoons CH_3OH(g) + H_2O(g)$$

Effect on yield _____



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Explanation _____

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**(3)**

(c) The hydrogen required for this synthesis is formed from methane and steam in a reversible reaction. The equation for this reaction is shown below.

$$CH_4(g) + H_2O(g) \rightleftharpoons C0(g) + 3H_2(g) \qquad \Delta H = +206 \text{ kJ mol}^{-1}$$

State and explain what happens to the yield of hydrogen in this reaction when the temperature is increased.

Effect on yield _____

Explanation _____

_____

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**(3)**

(d) The methanol produced by this synthesis has been described as a carbon-neutral fuel.

(i) State the meaning of the term *carbon-neutral*.

_____

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**(1)**

(ii) Write an equation for the complete combustion of methanol.

_____

**(1)**

(iii) The equation for the synthesis of methanol is shown below.



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$$CO_2(g) + 3H_2(g) \rightleftharpoons CH_3OH(g) + H_2O(g)$$

Use this equation and your answer to part (d)(ii) to deduce an equation to represent the overall chemical change that occurs when methanol behaves as a carbon-neutral fuel.

Equation _____

**(1)**

(e) A student carried out an experiment to determine the enthalpy change when a sample of methanol was burned.

The student found that the temperature of 140 g of water increased by 7.5 °C when 0.011 mol of methanol was burned in air and the heat produced was used to warm the water.

Use the student's results to calculate a value, in kJ $mol^{-1}$, for the enthalpy change when one mole of methanol was burned.
(The specific heat capacity of water is 4.18 J $K^{-1}$ $g^{-1}$).

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**(3)**

**(Total 16 marks)**

## Q18.

(a) Iron is extracted from iron(III) oxide using carbon at a high temperature.

(i) State the type of reaction that iron(III) oxide undergoes in this extraction.

_____

**(1)**

(ii) Write a half-equation for the reaction of the iron(III) ions in this extraction.



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_____

**(1)**

(b) At a high temperature, carbon undergoes combustion when it reacts with oxygen.

(i) Suggest why it is **not** possible to measure the enthalpy change directly for the following combustion reaction.

$$C(s,graphite) + \frac{1}{2}O_2(g) \longrightarrow CO(g)$$

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_____

**(1)**

(ii) State Hess's Law.

_____

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**(1)**

(iii) State the meaning of the term *standard enthalpy of combustion.*

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**(3)**

(c) Use the standard enthalpies of formation in the table below and the equation to calculate a value for the standard enthalpy change for the extraction of iron using carbon monoxide.

| | $Fe_2O_3(s)$ | $CO(g)$ | $Fe(l)$ | $CO_2(g)$ |
|---|---|---|---|---|
| $\Delta H_f$/ kJ mol$^{-1}$ | - 822 | - 111 | +14 | - 394 |

$$Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(l) + 3CO_2(g)$$

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**(3)**

(d) (i) Write an equation for the reaction that represents the standard enthalpy of formation of carbon dioxide.

_____

**(1)**

(ii) State why the value quoted in part (c) for the standard enthalpy of formation of $CO_2(g)$ is the same as the value for the standard enthalpy of combustion of carbon.

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**(1)**

**(Total 12 marks)**

## Q19.

This question is about the extraction of titanium from titanium(IV) oxide by a two-stage process.
The first stage in the process produces titanium(IV) chloride. In the second stage, titanium(IV) chloride is converted into titanium.
The enthalpy change for the second stage can be determined using Hess's Law.

(a) Give **one** reason why titanium is **not** extracted directly from titanium(IV) oxide using carbon.

_____

_____

**(1)**

(b) Give the meaning of the term *enthalpy change*.

_____

_____

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**(1)**

(c) State Hess's Law.

_____

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**(1)**

(d) Define the term *standard enthalpy of formation*.

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**(3)**

(e) The following standard enthalpy of formation data refer to the second stage in the extraction of titanium.

| | $TiCl_4(g)$ | Na(l) | NaCl(s) | Ti(s) |
|---|---|---|---|---|
| $\Delta H_f^\ominus$ / kJ $mol^{-1}$ | –720 | +3 | –411 | 0 |

(i) State why the value for the standard enthalpy of formation of Na(l) is **not** zero.

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**(1)**

(ii) Use data from the table to calculate a value for the standard enthalpy change of the following reaction.

$$TiCl_4(g) + 4Na(l) \longrightarrow 4NaCl(s) + Ti(s)$$

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**(3)**



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(iii) State the role of sodium in this reaction.

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**(1)**

**(Total 11 marks)**

## Q20.

Hess's Law is used to calculate the enthalpy change in reactions for which it is difficult to determine a value experimentally.

(a) State the meaning of the term *enthalpy change*.

_____

**(1)**

(b) State Hess's Law.

_____

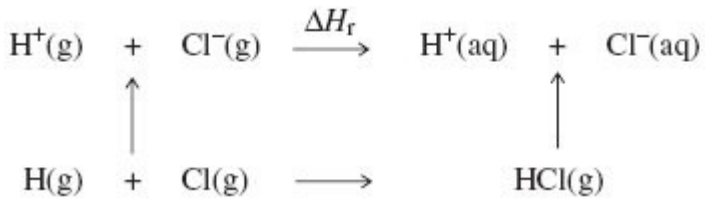
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**(1)**

(c) Consider the following table of data and the scheme of reactions.

| Reaction | Enthalpy change / kJ mol$^{-1}$ |
|---|---|
| $HCl(g) \rightarrow H^+(aq) + Cl^-(aq)$ | –75 |
| $H(g) + Cl(g) \rightarrow HCl(g)$ | –432 |
| $H(g) + Cl(g) \rightarrow H^+(g) + Cl^-(g)$ | +963 |




Use the data in the table, the scheme of reactions and Hess's Law to calculate a value for $\Delta H_r$

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**(3)**

**(Total 5 marks)**

## Q21.

Barium can be extracted from barium oxide (BaO) in a process using aluminium.
A mixture of powdered barium oxide and powdered aluminium is heated strongly.
The equation for this extraction process is shown below.

$$3BaO(s) + 2Al(s) \rightarrow 3Ba(s) + Al_2O_3(s)$$

Some standard enthalpies of formation are given in the table below.

| Substance | BaO(s) | $Al_2O_3(s)$ |
|---|---|---|
| $\Delta H_f^{\⦵}$/kJ mol$^{-1}$ | –558 | –1669 |

(a) (i) State what is meant by the term *standard enthalpy of formation*.

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**(3)**

(ii) State why the standard enthalpy of formation of barium and that of aluminium are both zero.

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**(1)**

(iii) Use the data to calculate the standard enthalpy change for the reaction shown by the equation above.

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**(3)**

(b) (i) Suggest the major reason why this method of extracting barium is expensive.

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**(1)**

(ii) Using barium oxide and aluminium powders increases the surface area of the



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reactants. Suggest **one** reason why this increases the rate of reaction.

_____

**(1)**

(c) (i) Write an equation for the reaction of barium with water.

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**(1)**

(ii) A solution containing barium ions can be used to test for the presence of sulfate ions in an aqueous solution of sodium sulfate.

Write the **simplest ionic** equation for the reaction which occurs and state what is observed.

Simplest ionic equation

_____

Observation _____

**(2)**

(iii) State how barium sulfate can be used in medicine. Explain why this use is possible, given that solutions containing barium ions are poisonous.

Use _____

Explanation _____

_____

**(2)**

**(Total 14 marks)**

## Q22.

The combustion of hydrocarbons is an important source of energy.

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

_____

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_____

**(3)**

(b) (i) Write an equation for the complete combustion of ethane, $C_2H_6$.

_____

(ii) Use the standard enthalpies of formation given below to calculate the standard enthalpy of combustion of ethane.

| Formula and state of compound | $C_2H_6(g)$ | $CO_2(g)$ | $H_2O(l)$ |
|---|---|---|---|



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| | | | |
|---|---|---|---|
| Standard enthalpy of formation (at 298 K)/kJ $mol^{-1}$ | –85 | –394 | –286 |

**(4)**

(c) A container and its contents of total heat capacity 120 J $K^{-1}$ were heated using a methane burner. Calculate the maximum theoretical temperature rise when 0.10 g of methane was completely burned. The standard enthalpy of combustion of methane is –890 kJ $mol^{-1}$.

**(4)**

**(Total 11 marks)**

## Q23.

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

**(3)**

(b) Use the data in the table to calculate the standard enthalpy of formation of liquid methylbenzene, $C_7H_8$

| Substance | C(s) | $H_2(g)$ | $C_7H_8(l)$ |
|---|---|---|---|
| Standard enthalpy of combustion, $\Delta H_c^{\ominus}$ /kJ $mol^{-1}$ | –394 | –286 | –3909 |

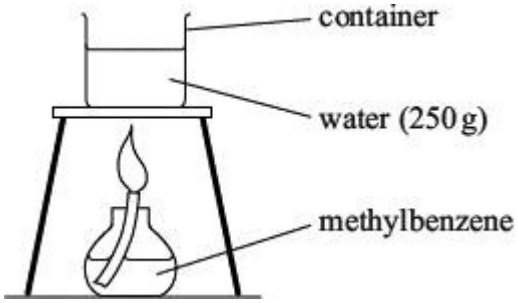
$7C(s) + 4H_2(g) \rightarrow C_7H_8(l)$

**(3)**

(c) An experiment was carried out to determine a value for the enthalpy of combustion of liquid methylbenzene using the apparatus shown in the diagram.



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Burning 2.5 g of methylbenzene caused the temperature of 250 g of water to rise by 60°C. Use this information to calculate a value for the enthalpy of combustion of methylbenzene, $C_7H_8$

(The specific heat capacity of water is 4.18 J $K^{-1}$ $g^{-1}$. Ignore the heat capacity of the container.)

**(4)**

(d) A 25.0 $cm^3$ sample of 2.00 mol $dm^{-3}$ hydrochloric acid was mixed with 50.0 $cm^3$ of a 1.00 mol $dm^{-3}$ solution of sodium hydroxide. Both solutions were initially at 18.0 °C.

After mixing, the temperature of the final solution was 26.5°C.

Use this information to calculate a value for the standard enthalpy change for the following reaction.

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

In your calculation, assume that the density of the final solution is 1.00 g $cm^{-3}$ and that its specific heat capacity is the same as that of water. (Ignore the heat capacity of the container.)

**(4)**

(e) Give **one** reason why your answer to part (d) has a much smaller experimental error than your answer to part (c).

**(1)**

**(Total 15 marks)**



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## Q24.

(a) Define the term *standard enthalpy of combustion*, $\Delta H_c$⦵

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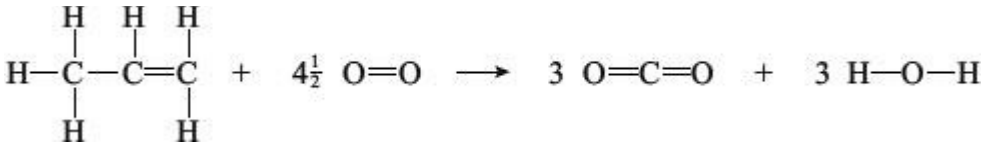
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**(3)**

(b) Use the mean bond enthalpy data from the table and the equation given below to calculate a value for the standard enthalpy of combustion of propene. All substances are in the gaseous state.

| Bond | C == C | C—C | C—H | O == O | O == C | O—H |
|---|---|---|---|---|---|---|
| Mean bond enthalpy/ kJ mol$^{-1}$ | 612 | 348 | 412 | 496 | 743 | 463 |



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**(3)**

(c) State why the standard enthalpy of formation, $\Delta H_f^{\circ}$, of oxygen is zero.

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**(1)**

(d) Use the data from the table below to calculate a more accurate value for the standard enthalpy of combustion of propene.

| Compound | $C_3H_6(g)$ | $CO_2(g)$ | $H_2O(g)$ |
|---|---|---|---|
| Standard enthalpy of formation, $\Delta H_f^{\circ}$/ kJ mol$^{-1}$ | +20 | –394 | –242 |

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**(3)**

(e) Explain why your answer to part (b) is a less accurate value than your answer to part (d).

_____

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**(2)**

**(Total 12 marks)**

## Q25.

(a) Explain the meaning of the terms *mean bond enthalpy* and *standard enthalpy of formation*.

*Mean bond enthalpy* _____

_____

_____

*Standard enthalpy of formation* _____

_____

_____

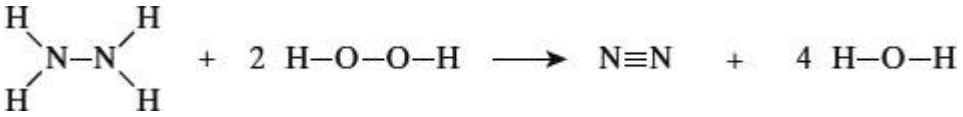
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**(5)**

(b) Some mean bond enthalpies are given below.

| Bond | N–H | N–N | N≡N | H–O | O–O |
|---|---|---|---|---|---|
| Mean bond enthalpy/kJ $mol^{-1}$ | 388 | 163 | 944 | 463 | 146 |

Use these data to calculate the enthalpy change for the following gas-phase reaction between hydrazine, $N_2H_4$, and hydrogen peroxide, $H_2O_2$



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**(3)**

(c) Some standard enthalpies of formation are given below.



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| | $N_2H_4(g)$ | $H_2O_2(g)$ | $H_2O(g)$ |
|---|---|---|---|
| $\Delta H_f^\circ$ /kJ mol$^{-1}$ | +75 | –133 | –242 |

These data can be used to calculate the enthalpy change for the reaction in part (b).

$$N_2H_4(g) + 2H_2O_2(g) \rightarrow N_2(g) + 4H_2O(g)$$

(i) State the value of $\Delta H_f^\ominus$ for $N_2(g)$.

_____

(ii) Use the $\Delta H_f^\ominus$ values from the table to calculate the enthalpy change for this reaction.

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**(4)**

(d) Explain why the value obtained in part (b) is different from that obtained in part (c) (ii).

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**(1)**

**(Total 13 marks)**

## Q26.

Methanol, $CH_3OH$, is a convenient liquid fuel.

(a) An experiment was conducted to determine the enthalpy of combustion of liquid methanol. The energy obtained from burning 2.12 g of methanol was used to heat 150 g of water. The temperature of the water rose from 298 K to 362 K. (The specific heat capacity of water is 4.18 J K$^{-1}$ g$^{-1}$)

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

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**(3)**

(ii) Use the data above to calculate a value for the enthalpy of combustion of one mole of liquid methanol.

**(4)**

(b) Methanol can be synthesised from methane and steam by a process that occurs in two stages.

*Stage 1* $CH_4(g) + H_2O(g) \rightleftharpoons 3H_2(g) + CO(g)$ $\Delta H^{\ominus} = +206 \text{ kJ mol}^{-1}$

*Stage 2* $CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$ $\Delta H^{\ominus} = -91 \text{ kJ mol}^{-1}$

(i) Explain why, in *Stage 1*, a higher yield of hydrogen and carbon monoxide is **not** obtained if the pressure is increased.

**(2)**

(ii) *Stage 2* is carried out at a compromise temperature of 500K. By considering what would happen at higher and lower temperatures, explain why 500 K is considered to be a compromise for *Stage 2*.



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**(3)**

(c) The standard enthalpies of combustion of carbon monoxide and of hydrogen are –283 kJ mol$^{-1}$ and –286 kJ mol$^{-1}$, respectively. Use these data and the enthalpy change for *Stage 2* to calculate a value for the standard enthalpy of combustion of gaseous methanol.

**(3)**

**(Total 15 marks)**

## Q27.

This question is about the reaction given below.

$$CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$$

Enthalpy data for the reacting species are given in the table below.

| Substance | CO(g) | $H_2O(g)$ | $CO_2(g)$ | $H_2(g)$ |
|---|---|---|---|---|
| $\Delta H_f^{\ominus}$ / kJ mol$^{-1}$ | −110 | −242 | −394 | 0 |

The standard enthalpy change for this reaction of carbon monoxide and steam is

**A** +42 kJ mol$^{-1}$

**B** −42 kJ mol$^{-1}$



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**C** +262 kJ mol$^{-1}$

**D** −262 kJ mol$^{-1}$

**(Total 1 mark)**

## Q28.

Use the information below to answer this question.

$$C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H^{\ominus} = -394 \text{ kJ mol}^{-1}$$

$$H_2(g) + \tfrac{1}{2}O_2(g) \rightarrow H_2O(l) \qquad \Delta H^{\ominus} = -286 \text{ kJ mol}^{-1}$$

$$4C(s) + 5H_2(g) \rightarrow C_4H_{10}(g) \qquad \Delta H^{\ominus} = -126 \text{ kJ mol}^{-1}$$

The standard enthalpy of combustion of butane, in kJ mol$^{-1}$, is

**A** −2880

**B** −2590

**C** −806

**D** −554

**(Total 1 mark)**

## Q29.

(a) The table below contains some mean bond enthalpy data.

| Bond | H–O | O–O | O=O |
|---|---|---|---|
| Mean bond enthalpy/kJ mol$^{-1}$ | 463 | 146 | 496 |

The bonding in hydrogen peroxide, $H_2O_2$, can be represented by H–O–O–H. Use these data to calculate the enthalpy change for the following reaction.

$$H_2O_2(g) \rightarrow H_2O_2(g) + \tfrac{1}{2}O_2(g)$$

_____

_____

_____

_____

**(3)**

(b) The standard enthalpy of formation, $\Delta H_f^{\ominus}$ for methane, is –74.9 kJ mol$^{-1}$. Write an equation, including state symbols, for the reaction to which this enthalpy change applies.

_____



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**(2)**

(c) The enthalpy changes for the formation of atomic hydrogen and atomic carbon from their respective elements in their standard states are as follows.

$\frac{1}{2}H_2(g) \rightarrow H(g)$ $\quad \Delta H^{\ominus} = +218$ kJ mol$^{-1}$

$C(s) \rightarrow C(g)$ $\quad \Delta H^{\ominus} = +715$ kJ mol$^{-1}$

(i) By reference to its structure, suggest why a large amount of heat energy is required to produce free carbon atoms from solid carbon.

_____

_____

(ii) Parts (b) and (c) give enthalpy data for the formation of $CH_4(g)$, H(g) and C(g). Use these data and Hess's Law to calculate the value of the enthalpy change for the following reaction.

$CH_4(g) \rightarrow C(g) + 4H(g)$

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(iii) Use your answer from part (c)(ii) to calculate a value for the mean bond enthalpy of a C–H bond in methane.

_____

**(5)**

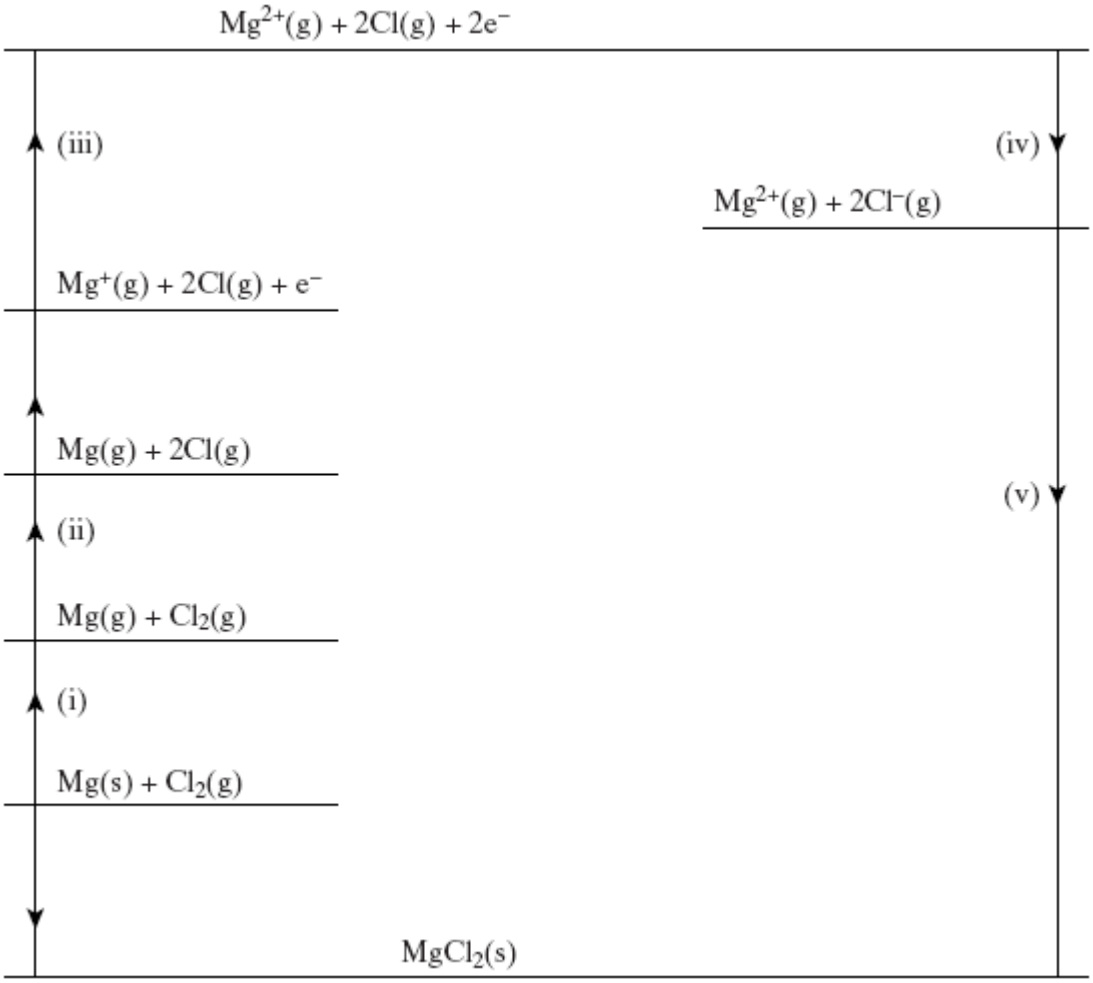
**(Total 10 marks)**

## Q30.

(a) A Born–Haber cycle for the formation of magnesium(II) chloride is shown below.



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Taking care to note the direction of the indicated enthalpy change and the number of moles of species involved, give each of the enthalpy changes (i) to (v) above.

*Enthalpy change (i)* _____

*Enthalpy change (ii)* _____

*Enthalpy change (iii)* _____

*Enthalpy change (iv)* _____

*Enthalpy change (v)* _____

**(5)**

(b) Write an equation for the decomposition of MgCl(s) into $MgCl_2(s)$ and Mg(s) and use the following data to calculate a value for the enthalpy change of this reaction.

$\Delta H_f^{\ominus}$ MgCl(s) = –113 kJ $mol^{-1}$

$\Delta H_f^{\ominus}$ $MgCl_2(s)$ = –653 kJ $mol^{-1}$

*Equation* _____

*Calculation* _____

_____



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**(4)**

(c) Use the data below to calculate a value for the molar enthalpy of a solution of $MgCl_2(s)$.

Lattice formation enthalpy of $MgCl_2(s)$ = –2502 kJ mol$^{-1}$

$\Delta H^{\ominus}_{hydration}$ of $Mg^{2+}(g)$ = –1920 kJ mol$^{-1}$

$\Delta H^{\ominus}_{hydration}$ of $Cl^-(g)$ = –364 kJ mol$^{-1}$

**(3)**

**(Total 12 marks)**

## Q31.

(a) The gaseous reactants **W** and **X** were sealed in a flask and the mixture left until the following equilibrium had been established.

$2W(g) + X(g) \rightleftharpoons 3Y(g) + 2Z(g)$ $\Delta H$ = –200 kJ mol$^{-1}$

Write an expression for the equilibrium constant, $K_p$, for this reaction.
State one change in the conditions which would both increase the rate of reaction and decrease the value of $K_p$. Explain your answers.



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**(7)**

(b) Ethyl ethanoate can be prepared by the reactions shown below.

**Reaction 1**

$CH_3COOH(l) + C_2H_5OH(l) \rightleftharpoons CH_3COOC_2H_5(l) + H_2O(l)$ $\Delta H^{\ominus} = -2.0$ kJ mol$^{-1}$

**Reaction 2**

$CH_3COCl(l) + C_2H_5OH(l) \rightarrow CH_3COOC_2H_5(l) + HCl(g)$ $\Delta H^{\ominus} = -21.6$ kJ mol$^{-1}$

(i) Give one advantage and one disadvantage of preparing ethyl ethanoate by **Reaction 1** rather than by **Reaction 2**.

**(2)**

(ii) Use the information given above and the data below to calculate values for the standard entropy change, $\Delta S^{\ominus}$, and the standard free-energy change, $\Delta G^{\ominus}$, for **Reaction 2** at 298 K.

| | $CH_3COCl(l)$ | $C_2H_5OH(l)$ | $CH_3COOC_2H_5(l)$ | $HCl(g)$ |
|---|---|---|---|---|
| $S^{\ominus}$/JK$^{1}$mol$^{1}$ | 201 | 161 | 259 | 187 |



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_____

**(6)**

**(Total 15 marks)**

## Q32.

Use the information below to answer this question.

$C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H^{\ominus} = -393.5$ kJ mol$^{-1}$

$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$ $\Delta H^{\ominus} = -285.8$ kJ mol$^{-1}$

$3C(s) + 4H_2(g) \rightarrow C_3H_8(g)$ $\Delta H^{\ominus} = -104.0$ kJ mol$^{-1}$

$4C(s) + 5H_2(g) \rightarrow C_4H_{10}(g)$ $\Delta H^{\ominus} = -125.2$ kJ mol$^{-1}$

The value in kJ mol$^{-1}$ of the enthalpy of thermal dissociation when butane forms propane, hydrogen and carbon is

**A** −26.3

**B** −17.5

**C** +17.5

**C** +21.2

**(Total 1 mark)**

## Q33.

Use the information below to answer this question.

$C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H^{\ominus} = -393.5$ kJ mol$^{-1}$

$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$ $\Delta H^{\ominus} = -285.8$ kJ mol$^{-1}$

$3C(s) + 4H_2(g) \rightarrow C_3H_8(g)$ $\Delta H^{\ominus} = -104.0$ kJ mol$^{-1}$

$4C(s) + 5H_2(g) \rightarrow C_4H_{10}(g)$ $\Delta H^{\ominus} = -125.2$ kJ mol$^{-1}$

The value in kJ mol$^{-1}$ for the enthalpy of combustion of propane is

**A** −211.7

**B** −419.7

**C** −2220

**C** −2878

**(Total 1 mark)**



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## Q34.

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

**(3)**

(b) State Hess's Law and use it, together with the data given in the table below, to calculate the standard enthalpy change for the following reaction.

$MgO(s) + 2HCl(g) \rightarrow MgCl_2(s) + H_2O(l)$

| | MgO | HCl(g) | $MgCl_2$ | $H_2O$ |
|---|---|---|---|---|
| $\Delta H_f^{\ominus}$/kJ mol$^{-1}$ | –602 | –92 | –642 | –286 |

**(4)**

(c) In an experiment, an excess of solid magnesium oxide was added to 50 cm$^3$ of 3.0 mol dm$^{-3}$ hydrochloric acid. The initial temperature of the solution was 21 °C. After reaction, the temperature had risen to 53 °C. (The specific heat capacity of water is 4.2 J K$^{-1}$ g$^{-1}$)

Use this information to calculate the enthalpy change for the reaction of one mole of magnesium oxide with hydrochloric acid. For your calculation you should assume that all the heat from the reaction is used to raise the temperature of 50 g of water.



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**(8)**

**(Total 15 marks)**

## Q35.

Chlorine is formed in a reversible reaction as shown by the equation

$$4HCl(g) + O_2(g) \rightleftharpoons 2Cl_2(g) + 2H_2O(g)$$

(a) Use the data below to calculate the standard enthalpy change, $\Delta H^{\ominus}$, and the standard entropy change, $\Delta S^{\ominus}$, for this reaction.

| Substance | HCl(g) | $O_2$(g) | $Cl_2$(g) | $H_2O$(g) |
|---|---|---|---|---|
| $\Delta H_f^{\ominus}$/kJ mol$^{-1}$ | –92 | 0 | 0 | –242 |
| $S^{\ominus}$/ J K$^{-1}$ mol$^{-1}$ | 187 | 205 | 223 | 189 |

*Standard enthalpy change*, $\Delta H^{\ominus}$ _____

*Standard entropy change*, $\Delta S^{\ominus}$ _____



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_____

_____

_____

**(6)**

(b) The data below apply to a different gas phase reversible reaction.

Standard enthalpy change, $\Delta H^{\ominus}$ = +208 kJ $mol^{-1}$
Standard entropy change, $\Delta S^{\ominus}$ = +253 J $K^{-1}$ $mol^{-1}$

(i) Deduce the effect of an increase in temperature on the position of the equilibrium in this reaction. Use Le Chatelier's principle to explain your answer.

*Effect* _____

*Explanation* _____

(ii) Calculate the minimum temperature at which this reaction is feasible.

_____

_____

_____

_____

**(7)**

**(Total 13 marks)**



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