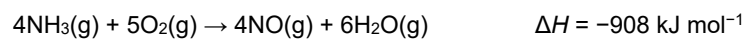


Enthalpy

1. Ammonia is used to make nitric acid. The first stage of the reaction is shown below.



Standard enthalpy changes of formation, $\Delta_f H^\ominus$, are given in the table.

Substance	$\Delta_f H^\ominus / \text{kJ mol}^{-1}$
$\text{NH}_3(\text{g})$	-46
$\text{O}_2(\text{g})$	0
$\text{H}_2\text{O}(\text{g})$	-242

- i. State the conditions of temperature and pressure used for standard enthalpy measurements.

Temperature

Pressure

[1]

- ii. Calculate the standard enthalpy change of formation for $\text{NO}(\text{g})$.

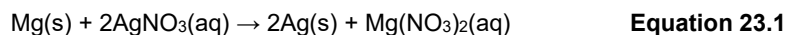
Give your answer to a **whole number**.

$\Delta_f H^\ominus$ for $\text{NO}(\text{g}) = \dots\dots\dots \text{kJ mol}^{-1}$ [3]

3.2.1 Enthalpy Changes

2. This question is about energy changes and rate of reaction.

Magnesium reacts with aqueous silver nitrate, $\text{AgNO}_3(\text{aq})$, as in **equation 23.1**.



A student carries out an experiment to determine the enthalpy change of this reaction, $\Delta_r H$.

- The student adds 25.0 cm^3 of $0.512 \text{ mol dm}^{-3}$ AgNO_3 to a polystyrene cup.
- The student measures the temperature of the solution.
- The student adds a small spatula measure of magnesium powder, stirs the mixture and records the maximum temperature

Temperature readings

Initial temperature	= $19.5 \text{ }^\circ\text{C}$
Maximum temperature	= $47.5 \text{ }^\circ\text{C}$

- i. Calculate $\Delta_r H$, in kJ mol^{-1} , for the reaction shown in **equation 23.1**.

Give your answer to an **appropriate** number of significant figures.

Assume that the density and specific heat capacity, c , of the solution are the same as for water and that all the aqueous silver nitrate has reacted.

$$\Delta_r H = \text{-----} \text{ mol}^{-1} \quad \begin{matrix} \text{kJ} \\ \mathbf{[4]} \end{matrix}$$

- ii. At the end of the experiment, the student adds a few drops of aqueous sodium chloride to the reaction mixture in the polystyrene cup to test whether all the aqueous silver nitrate has reacted.

Explain how the results would show whether all the aqueous silver nitrate has reacted. Include an equation with state symbols in your answer.

[2]

3.2.1 Enthalpy Changes

3(a). Hydrogen reacts with chlorine to form hydrogen chloride, HCl:

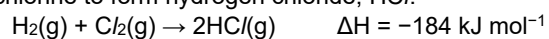


Table 3.1 shows bond enthalpies.

Bond	Bond Enthalpy / kJ mol^{-1}
H-H	+436
Cl-Cl	+243

Table 3.1

Calculate the bond enthalpy for the H-Cl bond from the information above.

bond enthalpy = _____ kJ mol^{-1} [2]

(b). 'Enthalpy change of vaporisation' is the enthalpy change when one mole of a substance changes from a liquid to a gas at its boiling point.

- i. Write an equation, including state symbols, to represent the enthalpy change of vaporisation of bromine.

----- [1]

- ii. Suggest whether the enthalpy change of vaporisation of bromine is exothermic or endothermic.

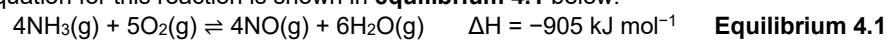
Explain your answer.

----- [1]

3.2.1 Enthalpy Changes

4. The reaction of ammonia, NH_3 , with oxygen to form nitrogen monoxide, NO , is an important industrial process.

The equation for this reaction is shown in **equilibrium 4.1** below.

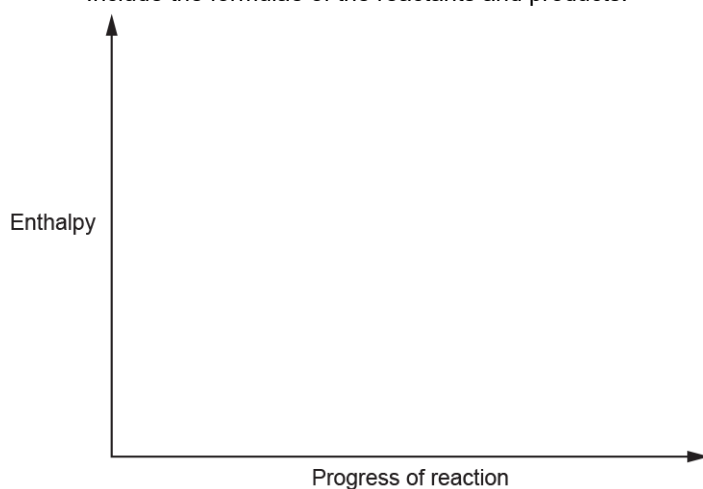


The forward reaction in **equilibrium 4.1** converts NH_3 into NO .

- i. Complete the enthalpy profile diagram for this reaction.

On your diagram:

- Label the activation energy, E_a
- Label the enthalpy change of reaction, ΔH
- Include the formulae of the reactants and products.



[2]

- ii. 5.10 tonnes of NH_3 are converted into NO .

Calculate the energy released, in kJ, for this conversion.

Give your answer in **standard form** and to an **appropriate** number of significant figures.

energy released = _____ kJ [4]

3.2.1 Enthalpy Changes

5. Bromine reacts with iodine to form iodine monobromide, IBr.

The table below lists some average bond enthalpies which are required in different parts of this question.

Bond	Average bond enthalpy / kJ mol ⁻¹
Br-Br	+193
I-I	+151
I-Br	+175

- i. Average bond enthalpy is the enthalpy change for the breaking of 1 mole of bonds in gaseous molecules.

Why do Br₂ and I₂ **not** exist in the gaseous state under standard conditions?

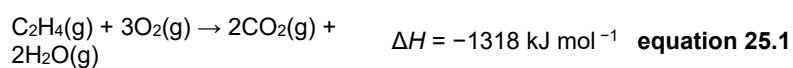
----- [1]

- ii. Calculate the enthalpy change of formation, $\Delta_f H$, for IBr.

$\Delta_f H = \dots\dots\dots$ kJ mol⁻¹ [2]

6. This question is about alkenes.

The combustion of ethene is shown in **equation 25.1** below.



- i. Explain, in terms of bond breaking and bond forming, why a reaction can be exothermic.

----- [1]

3.2.1 Enthalpy Changes

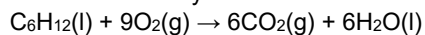
- ii. Average bond enthalpies are shown in the table.

Bond	Average bond enthalpy /kJ mol ⁻¹
O–H	+464
O=O	+498
C–H	+413
C=O	+805

Calculate the average bond enthalpy of the C=C bond.
Use the average bond enthalpies in the table and **equation 25.1**.

average bond enthalpy = kJ mol⁻¹ [3]

- 7(a). The equation for the complete combustion of cyclohexane is shown below.



Standard enthalpy changes of formation, $\Delta_f H^\ominus$, are shown in the table.

Substance	C ₆ H ₁₂ (l)	CO ₂ (g)	H ₂ O(l)
$\Delta_f H^\ominus / \text{kJ mol}^{-1}$	-156.3	-393.5	- 285.8

Calculate the standard enthalpy change of combustion, $\Delta_c H^\ominus$, of cyclohexane.

$\Delta_c H^\ominus = \dots\dots\dots$ kJ mol⁻¹ [3]

3.2.1 Enthalpy Changes

(b). Alkanes are saturated hydrocarbons with the general formula C_nH_{2n+2} .

A student carries out an experiment to measure the enthalpy change of combustion, $\Delta_c H$, of hexane.

The student finds that combustion of 1.29 g of hexane changes the temperature of 200 g of water from 20.5 °C to 65.5 °C.

i. Calculate the enthalpy change of combustion, $\Delta_c H$, of hexane, in kJ mol^{-1} .

Give your final answer to an **appropriate** number of significant figures.

$\Delta_c H$ kJ mol^{-1} [4]

ii. The calculated value of $\Delta_c H$ for hexane from this experiment is different from the data book value.

Suggest **two** reasons for this difference.

1

2

[2]

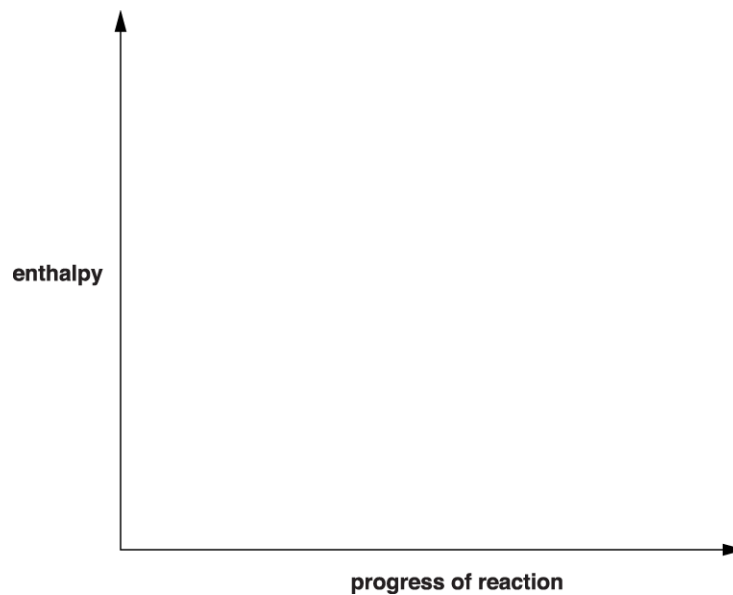
3.2.1 Enthalpy Changes

8. Catalysts can be used to change the rate of some chemical reactions.

- i. Zinc and sulfuric acid react together to form a solution of zinc sulfate, ZnSO_4 , and hydrogen gas. The reaction is exothermic.

The rate of the reaction increases when a catalyst is added.

- Complete the enthalpy profile diagram for this reaction using the formulae of the reactants and products.
- Label activation energies, E_a (without catalyst) and E_c (with catalyst).
- Label the enthalpy change of reaction, ΔH .



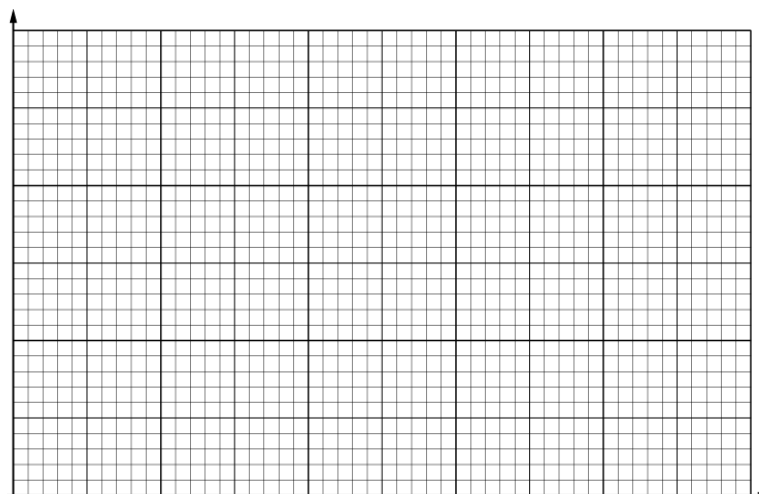
[3]

- ii. Using a Boltzmann distribution, explain how a catalyst increases the rate of a chemical reaction.

Include a labelled sketch of your Boltzmann distribution on the grid below. Label the axes and any other important features.



Your answer needs to be clear and well organised using the correct terminology.

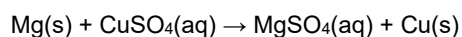


3.2.1 Enthalpy Changes

[4]

9(a). This question is about different ways of determining enthalpy changes.

A student carries out an experiment to determine directly the enthalpy change of reaction, ΔH_r , shown below.



- The student measures the initial temperature of 25.0 cm³ of 0.500 mol dm⁻³ CuSO₄(aq).
- The student adds an excess of magnesium powder and stirs the mixture.
- The student measures the maximum temperature of the solution.

Results

Initial temperature of solution = 21.5 °C

Maximum temperature of solution = 63.0 °C

Density of the solution = 1.00 g cm⁻³; specific heat capacity of the solution = 4.18 J g⁻¹ K⁻¹.

- i. Calculate the enthalpy change of reaction, ΔH_r , in kJ mol⁻¹.

Give your answer to **three** significant figures.

$\Delta H_r = \dots\dots\dots$ kJ mol⁻¹ **[4]**

- ii. The student weighed out enough magnesium so that it would be in excess by **at least** 25%. The student had access to a two decimal-place balance.

Calculate the minimum mass of magnesium that the student would need to weigh out on this balance.

mass = $\dots\dots\dots$ g **[1]**

3.2.1 Enthalpy Changes

- (b). Enthalpy changes of formation can be determined indirectly from standard enthalpy changes of combustion, ΔH_c^\ominus .

Three enthalpy changes of combustion are shown below.

Substance	$\Delta H_c^\ominus/\text{kJ mol}^{-1}$
C(s)	-394
H ₂ (g)	-286
C ₉ H ₂₀ (l)	-6125

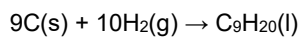
- i. Define *standard enthalpy change of combustion*.

Include the standard conditions that are used.

[3]

- ii. The equation that represents the enthalpy change of formation, ΔH_f , of nonane is shown below.

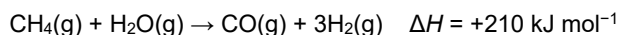
Calculate the standard enthalpy change of formation of nonane.



$\Delta H_f = \dots\dots\dots \text{kJ mol}^{-1}$ [2]

3.2.1 Enthalpy Changes

- (c). The bond enthalpy for the bond in carbon monoxide can be calculated from the information below.



Bond	Average bond enthalpy / kJ mol^{-1}
C–H	413
O–H	464
H–H	436

- i. What is meant by the term *average bond enthalpy*?

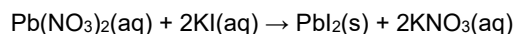
----- [2]

- ii. Calculate the bond enthalpy for the bond in carbon monoxide.

Show your working.

bond enthalpy = kJ mol^{-1} [3]

- 10(a). Aqueous lead(II) nitrate, $\text{Pb}(\text{NO}_3)_2(\text{aq})$, and aqueous potassium iodide, $\text{KI}(\text{aq})$, react together. The equation is shown below.



A student carries out an experiment to determine the enthalpy change of reaction, $\Delta_r H$, of this reaction.

The student follows the method outlined below.

- Add 50.0 cm^3 of 1.50 mol dm^{-3} $\text{Pb}(\text{NO}_3)_2(\text{aq})$ to a polystyrene cup.
- Measure out 50.0 cm^3 of a solution of $\text{KI}(\text{aq})$, which is in excess.
- Measure the temperature of both solutions.
- Add the $\text{KI}(\text{aq})$ to the polystyrene cup, stir the mixture and record the maximum temperature.

3.2.1 Enthalpy Changes

Temperature readings

Initial temperature of both solutions = 19.5 °C

Maximum temperature of mixture = 30.0 °C

Calculate $\Delta_r H$, in kJ mol^{-1} , for the reaction shown in the equation above.

Give your answer to an **appropriate** number of significant figures.

Assume that the density of all solutions and specific heat capacity, c , of the reaction mixture is the same as for water.

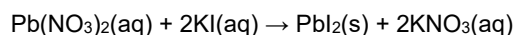
$\Delta_r H = \dots\dots\dots \text{kJ mol}^{-1}$ **[4]**

(b). Write an ionic equation for the reaction that the student carries out.

Include state symbols.

----- **[1]**

(c). The 50.0 cm^3 of KI(aq) used in the experiment contains 10% more KI than is needed to react with 50.0 cm^3 of 1.50 mol dm^{-3} $\text{Pb(NO}_3)_2\text{(aq)}$.



Calculate the concentration, in mol dm^{-3} , of KI that the student used.

concentration of $\text{KI} = \dots\dots\dots \text{mol dm}^{-3}$ **[2]**

3.2.1 Enthalpy Changes

11. Nitrogen forms several different oxides.

N_2O is a useful anaesthetic and NO has been linked to the depletion of ozone in the stratosphere.

The standard enthalpy changes of formation of N_2O and NO are given in the table.

Compound	$\Delta_f H^\ominus / \text{kJ mol}^{-1}$
$\text{N}_2\text{O (g)}$	+ 82.0
NO (g)	+ 90.2

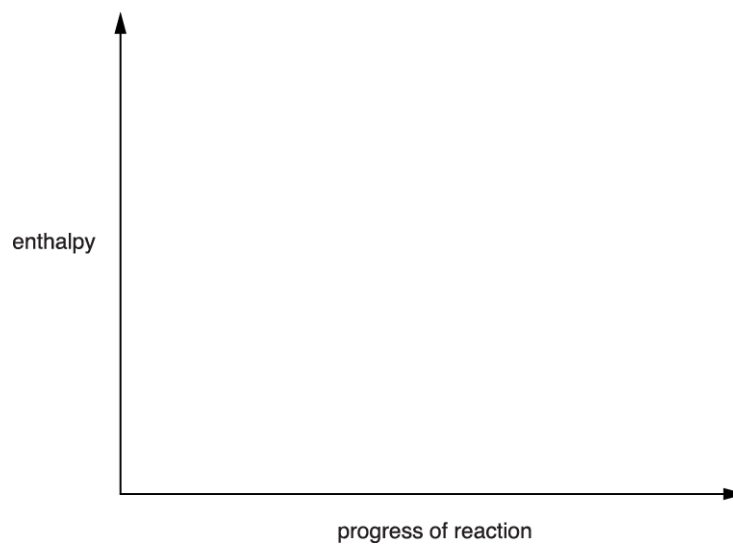
- i. Explain, in terms of bond breaking and bond making, why the enthalpy change of formation of NO is endothermic.

[1]

- ii. Draw a fully labelled enthalpy profile diagram to represent the enthalpy change of formation of N_2O .

The formulae, with state symbols, of the reactants and products should be included as part of the diagram.

You are **not** expected to show the activation energy for the reaction.

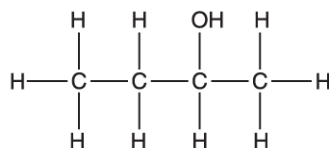
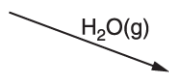
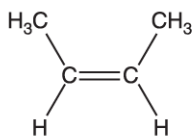


[2]

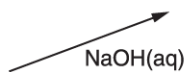
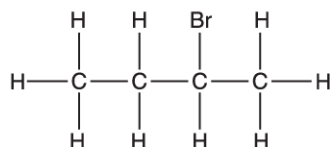
3.2.1 Enthalpy Changes

12. Butan-2-ol can be prepared using two different methods.

Method 1



Method 2



butan-2-ol

Average bond enthalpies can be used to calculate enthalpy changes.

- i. What is meant by the term *average bond enthalpy*?

[2]

- ii. Calculate the enthalpy change of reaction, ΔH_r , for preparing 1 mol of butan-2-ol by **Method 1**.

Average bond enthalpies are given below.

Bond	Average bond enthalpy / kJ mol^{-1}
O—H	464
C—H	413
C—C	347
C—O	358
C=C	612

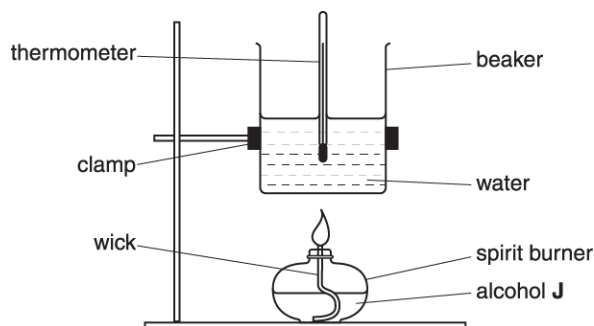
$\Delta H_r = \dots\dots\dots \text{kJ mol}^{-1}$ [3]

3.2.1 Enthalpy Changes

13(a). A branched-chain alcohol **J** is a liquid and has the molecular formula $C_5H_{12}O$.

A student does an experiment to measure the enthalpy change of combustion, ΔH_c , of alcohol **J**.

i. The student burns alcohol **J** using the apparatus below.



The student found that combustion of 1.54 g of alcohol **J** changes the temperature of 180 g of water from 22.8 °C to 75.3 °C.

The specific heat capacity of water is 4.18 J g⁻¹ K⁻¹.

- Calculate the amount, in mol, of alcohol **J** that burns.
- Calculate the enthalpy change of combustion, ΔH_c , of alcohol **J**, in kJ mol⁻¹.

Give your final answer to **three** significant figures.

$\Delta H_c = \dots\dots\dots$ kJ mol⁻¹

ii. The calculated value of ΔH_c from this experiment is different from the value obtained from data books.

Apart from heat loss, suggest **two** reasons for the difference.

Assume that the calculation has been carried out correctly.

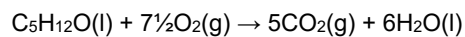
3.2.1 Enthalpy Changes

(b). The enthalpy change of combustion for alcohol **J** can also be determined indirectly from standard enthalpy changes of formation.

- i. Write an equation, including state symbols, for the chemical change that represents the standard enthalpy change of formation of the liquid alcohol **J**, $C_5H_{12}O$.

----- [1]

- ii. The equation for the complete combustion of alcohol **J** is shown below.



Enthalpy changes of formation, ΔH_f , are shown in the table.

Substance	$C_5H_{12}O(l)$	$CO_2(g)$	$H_2O(l)$
$\Delta H_f / \text{kJ mol}^{-1}$	-366	-394	-286

Calculate the enthalpy change of combustion, ΔH_c , of alcohol **J** from the information given above.

$\Delta H_c = \dots\dots\dots \text{kJ mol}^{-1}$ [3]

3.2.1 Enthalpy Changes

- ii. Write the equation, including state symbols, for the reaction that represents the standard enthalpy change of formation of $\text{NH}_4\text{C/O}_4(\text{s})$.

----- [2]

- iii. Calculate the enthalpy change of formation of $\text{NO}(\text{g})$ using the data above.

enthalpy change of formation of $\text{NO}(\text{g}) = \dots\dots\dots \text{kJ mol}^{-1}$ [3]

- 15(a).** Hydrogen iodide, HI, is a colourless gas that can be made from the reaction of hydrogen, H_2 , and iodine, I_2 .

This reversible reaction is shown in **equilibrium 3.1** below.



The activation energy for the forward reaction is 173 kJ mol^{-1} .

Calculate the activation energy, E_a , for the reverse reaction.

E_a (reverse reaction) = $\dots\dots\dots \text{kJ mol}^{-1}$ [1]

- (b).** When the reverse reaction takes place hydrogen iodide, HI, decomposes to form iodine and hydrogen.

Calculate the enthalpy change when 336 dm^3 of hydrogen iodide, measured at room temperature and pressure, decomposes.

Include the sign for enthalpy change in your answer.

enthalpy change $\dots\dots\dots \text{kJ}$ [2]

3.2.1 Enthalpy Changes

(c). Calculate the bond enthalpy for the H–I bond in **equilibrium 3.1**, given the following information.

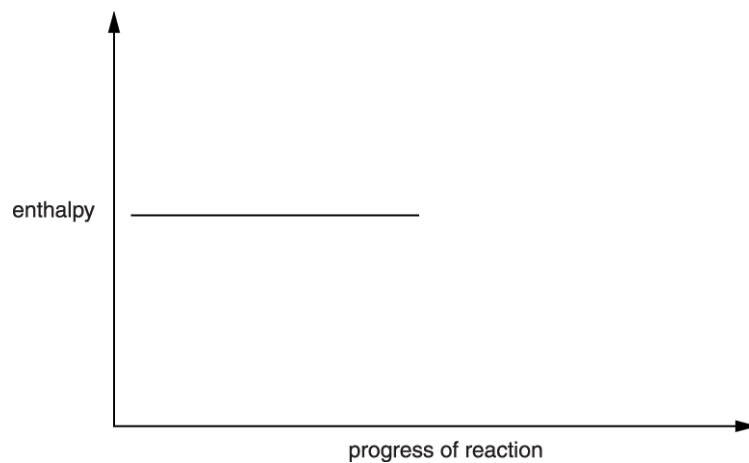
Bond	Bond Enthalpy / kJ mol^{-1}
H–H	436
I–I	151

bond enthalpy kJ mol^{-1} [2]

(d). Complete the enthalpy profile diagram below for the forward reaction in **equilibrium 3.1**.

On your diagram:

- Label the activation energy, E_a
- Label the enthalpy change of reaction, ΔH
- Include the formulae of the reactants and products.



[2]

3.2.1 Enthalpy Changes

16. This question looks at neutralisation reactions.

A student carries out an experiment to determine the enthalpy change for a neutralisation reaction.

The student measures out 35.0 cm³ of 2.40 mol dm⁻³ KOH and 35.0 cm³ of 1.20 mol dm⁻³ H₂SO₄.

The temperature of each solution is 19.5 °C.

The student mixes the solutions. The KOH is all neutralised and the maximum temperature reached is 36.0 °C.

- i. Write the overall equation for the reaction that takes place.

[1]

- ii. Calculate the enthalpy change for the reaction between 1 mol KOH and 1 mol HCl.

Assume that the density of the mixture is 1.00 g cm⁻³ and that the specific heat capacity for the solution is the same as for water.

$\Delta H = \dots\dots\dots$ kJ [3]

- iii. Explain, why the answer to (ii) is the enthalpy change of neutralisation.

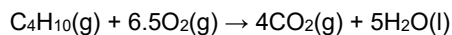
[1]

- iv. In this experiment, the student uses a thermometer with an uncertainty of ± 0.5 °C in each reading.

Calculate the percentage uncertainty in the temperature rise.

percentage uncertainty = $\dots\dots\dots$ % [1]

17(a). Butane, C₄H₁₀, is a highly flammable gas, used as a fuel for camping stoves. Butane reacts with oxygen as in the equation below:



Explain why this equation represents the standard enthalpy change of combustion of butane.

[1]

3.2.1 Enthalpy Changes

- (b). Butane can be produced from decane, $C_{10}H_{22}$, as shown in the equation below.
 $C_{10}H_{22}(l) \rightarrow C_4H_{10}(g) + 2C_3H_6(g)$

Standard enthalpy changes of combustion, $\Delta_c H^\ominus$, are shown in the table below.

Substance	$\Delta_c H^\ominus / \text{kJ mol}^{-1}$
$C_{10}H_{22}(l)$	-6778
$C_4H_{10}(g)$	-2877
$C_3H_6(g)$	-2058

Calculate the standard enthalpy change of reaction, $\Delta_r H^\ominus$, for the reaction. Include the sign.
 $C_{10}H_{22}(l) \rightarrow C_4H_{10}(g) + 2C_3H_6(g)$

$$\Delta_r H^\ominus = \dots\dots\dots \text{kJ mol}^{-1} \quad [2]$$

END OF QUESTION PAPER