



# STARTER FOR 10!!!

## 6.4. Using bond enthalpies

1. A student is carrying out a project to compare the theoretical and experimental value for the enthalpy change of combustion of ethanol. Using the data in the table, calculate a theoretical value for  $\Delta H_c^\ominus$  [ $\text{CH}_3\text{CH}_2\text{OH}(\text{l})$ ].

(**HINT** Remember to fully balance any equations before starting your calculations) (4 marks)

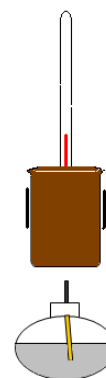
Bond	Mean bond enthalpy / $\text{kJ mol}^{-1}$
C—C	347
C—H	413
C—O	358
O—H	464
O=O	498
C=O	805

2. When the student shows his calculation to his teacher, she points out that mean bond enthalpies are only applicable for molecules in the gas state. Therefore the student must take into account the enthalpy change of vaporisation of ethanol [ $\text{CH}_3\text{CH}_2\text{OH}(\text{l}) \rightarrow \text{CH}_3\text{CH}_2\text{OH}(\text{g})$ ,  $\Delta H_{\text{vap}} +39 \text{ kJ mol}^{-1}$ ].

Use this value to correct your answer to Q1 (You may assume that the water formed from the combustion is in the gas state). (1 mark)

3. The student now wishes to determine an experimental value for the enthalpy of combustion of ethanol. He intends to burn approximately 1 g of fuel and measure the heat energy produced by heating up a known volume of water in a copper calorimeter (using the equipment shown).

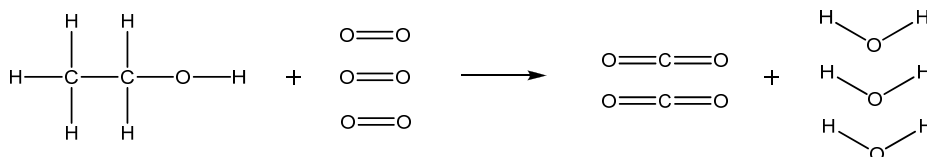
Using your answer to question 2, suggest a suitable volume of water for the copper calorimeter if he is aiming for a temperature rise of no more than  $40^\circ\text{C}$ ? (Specific heat capacity of water =  $4.2 \text{ J K}^{-1} \text{ g}^{-1}$ ) (4 marks)



4. The experimental value obtained by the student is considerably lower than the theoretical value calculated. Suggest one reason for this (other than experimental error). (1 mark)

## 6.4. Using bond enthalpies

1.



Bonds broken / kJ mol <sup>-1</sup>		Bonds made / kJ mol <sup>-1</sup>	
1 × C—C	1 × 347	4 × C=O	4 × 805
5 × C—H	5 × 413	6 × O—H	6 × 464
1 × C—O	1 × 358		
1 × O—H	1 × 464		
3 × O=O	3 × 498		
<b>Total energy in = 4728 kJ mol<sup>-1</sup></b>		<b>Total energy out = 6004 kJ mol<sup>-1</sup></b>	

(1 mark balanced equation)

(2 marks bonds broken /made)

$$\begin{aligned}\Delta H_c^\ominus (\text{CH}_3\text{CH}_2\text{OH}) &= \text{Total energy in} - \text{total energy out} \\ &= 4728 \text{ kJ mol}^{-1} - 6004 \text{ kJ mol}^{-1} \\ &= \mathbf{-1276 \text{ kJ mol}^{-1}}\end{aligned}$$

(1 mark correct sum)

2. 39 kJ mol<sup>-1</sup> of energy must be put in to the reaction to initially convert the liquid ethanol into gaseous ethanol. Hence, the total energy in becomes 4728 kJ mol<sup>-1</sup> + 39 kJ mol<sup>-1</sup> = 4767 kJ mol<sup>-1</sup>. Therefore the more correct  $\Delta H_c^\ominus [\text{CH}_3\text{CH}_2\text{OH}(\text{l})] = 4767 \text{ kJ mol}^{-1} - 6004 \text{ kJ mol}^{-1} = \mathbf{-1237 \text{ kJ mol}^{-1}}$

(1 mark)

3. Number of moles in 1 g = 1 g ÷  $M_r(\text{CH}_3\text{CH}_2\text{OH}) = 1 \text{ g} \div 46 \text{ g mol}^{-1} = \mathbf{0.022 \text{ moles}}$  (1 mark)

Theoretical heat transferred by 1 g = 0.022 moles × 1237 kJ mol<sup>-1</sup> = **26.9 kJ** (1 mark)

26891 J = mass of water × 4.2 J K<sup>-1</sup> mol<sup>-1</sup> × 40 K

∴ mass of water = 26891 J / (4.2 J K<sup>-1</sup> g<sup>-1</sup> × 40 K)

= 160 g (1 mark)

Density of water = 1 g cm<sup>-3</sup>, therefore 160 g has a volume of **160 cm<sup>3</sup>** (1 mark)

4. *either* Mean bond enthalpies are averages of the bond enthalpies in many different compounds. Therefore they are not exact for the specific bonds in ethanol

*or* There is considerably loss of heat to the environment / copper calorimeter that is not included in the calculations for the experimental enthalpy of combustion of ethanol. This loss of heat would result in an experimental value that is lower than the actual value.

*or* The heat capacity of the copper calorimeter has not been taken into account meaning that the heat transferred into the copper is not included in the calculation.