

AP CHEMISTRY

TOPIC 9: THERMODYNAMICS, PART B,

Day 109:

- Hess's law
- Enthalpy of Formation
- Spontaneity
- Entropy
- Heat Capacity

A past AP question within the last five years: with some parts to help "jog the memory" again !!!

In an experiment, a sample of an unknown, pure gaseous hydrocarbon was analyzed. Results showed that the sample contained 6.000 g of carbon and 1.344 g of hydrogen.

- (a) Determine the empirical formula of the hydrocarbon.

$$\frac{6.000 \text{ g}}{12.011 \text{ g}} \times \frac{1 \text{ mol C}}{1} = 0.50 \text{ mol C} ; \frac{0.50 \text{ mol}}{0.50 \text{ mol}} = 1 \times 3 = 3$$

$$\frac{1.344 \text{ g}}{1.008 \text{ g}} \times \frac{1 \text{ mol H}}{1} = 1.33 \text{ mol H} ; \frac{1.33 \text{ mol}}{0.50 \text{ mol}} = 2\frac{2}{3} \times 3 = 8$$



- (b) The density of the hydrocarbon at 25 °C and 1.09 atm is 1.96 g L⁻¹.

- (i) Calculate the molar mass of the hydrocarbon.

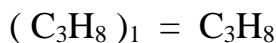
$$PV = nRT ; PV = \frac{m}{M}RT$$

$$M = \left(\frac{m}{V}\right) \times \frac{RT}{P} = \left(\frac{1.96 \text{ g}}{L}\right) \times \left(\frac{(0.0821 \text{ atm}\cdot\text{L})(298 \text{ K})}{(\text{mol}\cdot\text{K})1.09 \text{ atm}}\right) = 44 \frac{\text{g}}{\text{mol}}$$

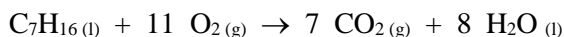
- (ii) Determine the molecular formula of the hydrocarbon.



$$n = \frac{\text{molecular formula mass}}{\text{empirical formula mass}} = \frac{44 \text{ g}}{44 \text{ g}} = 1$$



In another experiment, liquid heptane, $C_7H_{16(l)}$, is completely combusted to produce $CO_{2(g)}$ and $H_2O_{(l)}$ as represented by the following equation below:



The heat of combustion, ΔH^0_{comb} , for one mole of $C_7H_{16(l)}$ is - 4850 kJ

(c) Using the information in the table below, calculate the value of ΔH^0_f for $C_7H_{16(l)}$ in $kJ mol^{-1}$.

Compound	ΔH^0_f ($kJ mol^{-1}$)
$CO_{2(g)}$	- 393.5
$H_2O_{(l)}$	- 285.8
$O_{2(g)}$	0

$$\Delta H^0_{rxn} = \sum H^0_f \text{ (products)} - \sum H^0_f \text{ (reactants)}$$

$$\Delta H^0_{comb.} = [7 \Delta H^0_f (CO_2) + 8 \Delta H^0_f (H_2O)] - [\Delta H^0_f (C_7H_{16}) + 0] = - 4850 \text{ kJ mol}^{-1}$$

$$[7 (-393.5 \text{ kJ mol}^{-1}) + 8 (-285.8 \text{ kJ mol}^{-1})] - [\Delta H^0_f (C_7H_{16})] = - 4850 \text{ kJ mol}^{-1}$$

$$\Delta H^0_f (C_7H_{16}) = [7 (-393.5 \text{ kJ mol}^{-1}) + 8 (-285.8 \text{ kJ mol}^{-1})] + 4850 \text{ kJ mol}^{-1}$$

$$\Delta H^0_f (C_7H_{16}) = [- 5049 \text{ kJ mol}^{-1}] + 4850 \text{ kJ mol}^{-1} = - 190.9 \text{ kJ mol}^{-1}$$

$$\Delta H^0_f (C_7H_{16}) = - 190.9 \text{ kJ mol}^{-1} \quad ; \quad \Delta H^0_f (C_7H_{16}) = - 190.9 \text{ kJ mol}^{-1}$$

(d) A 0.0108 mol sample of $C_7H_{16(l)}$ is combusted in a bomb calorimeter.

(i) Calculate the amount of heat released to the calorimeter.

$$\Delta H^0_{combustion} (C_7H_{16}) = - 4850 \text{ kJ mol}^{-1}$$

$$\frac{0.0108 \text{ mol } C_7H_{16}}{1} \times \frac{-4850 \text{ kJ}}{\text{mol}} = -52.4 \text{ kJ}$$

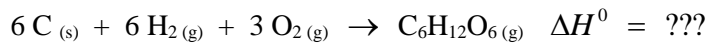
(ii) Given that the total heat capacity of the calorimeter is $9.273 \text{ kJ } ^\circ C^{-1}$, calculate the temperature change of the calorimeter.

The question wants the change for the calorimeter...

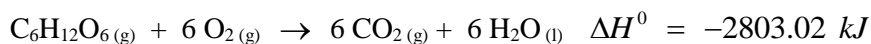
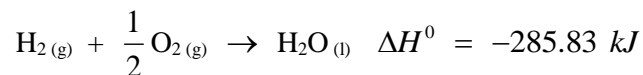
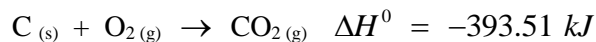
Amount of energy **ABSORBED** by the calorimeter was 52.4 kJ

$$C_p = \frac{\Delta H}{\Delta T} \quad ; \quad \Delta T = \frac{\Delta H}{C_p} = \frac{52.4 \text{ kJ } (^\circ C)}{9.273 \text{ kJ}} = 5.65 \text{ } ^\circ C$$

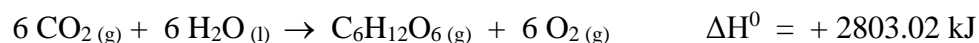
2) Calculate ΔH^0 for this reaction:



using the following equations:



Answer



CHANGES TO ...

Answer

