3. 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.

(b) The density of the hydrocarbon at 25°C and 1.09 atm is 1.96 g L⁻¹.
   
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In another experiment, liquid heptane, \( C_7H_{16}(l) \), is completely combusted to produce \( \text{CO}_2(g) \) and \( \text{H}_2\text{O}(l) \), as represented by the following equation.

\[
C_7H_{16}(l) + 11 \text{O}_2(g) \rightarrow 7 \text{CO}_2(g) + 8 \text{H}_2\text{O}(l) \quad \Delta H_{\text{comb}}^\circ = -4.85 \times 10^3 \text{kJ}
\]

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(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.

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\[
\alpha = M_c = 6.000 \text{ g} \left( \frac{1 \text{ mol}}{12.011 \text{ g}} \right) = 0.4995 \text{ mol}
\]

\[
M_H = 1.344 \text{ g} \left( \frac{1 \text{ mol}}{1.0079 \text{ g}} \right) = 1.333 \text{ mol H}
\]

\[
C_{0.4995} \cdot H_{1.333} \Rightarrow C_{\frac{4}{7}} H_{\frac{8}{7}}
\]

\[
\Rightarrow C_7H_8 - \text{ empirical formula}
\]

\[
\text{b) } T = 253 + 25 = 278 \text{ K}
\]

\[
\text{d) } 1.96 \text{ g/L}
\]

\[
P = 1.09 \text{ atm}
\]

\[
\text{i) Molar mass } = \frac{dRT}{P} = 1.96 \frac{\text{g}}{\text{L}} \times \frac{0.08206 \text{ atm}}{1.09 \text{ atm}} \left(298 \text{ K} \right)
\]

\[
\text{Molar mass } = 43.97 \text{ g/mol}
\]

\[
\text{GO ON TO THE NEXT PAGE.}
\]
ii) Molecular formula:

\[ \text{H}_4 \text{O}_8 = \text{C}_3 \text{H}_8 + \text{H}_2 \text{O}_4 \]

\[ \text{H}_4 \text{O}_8 = \text{C}_3 \text{H}_8 + 4\text{H}_2 \text{O} \]

\[ k = \frac{4\text{H}_2 \text{O}}{\text{H}_4 \text{O}_8} = 0.998 \text{ or } 1 \]

Molecular formula is \( \text{C}_3 \text{H}_8 \) since \( k \approx 1 \).

(c) \( -1.85 \times 10^3 \text{kJ} = \left( 8\text{mol} \times \left(-225.8 \text{kJ/mol}\right) \right) + 2\text{mol} \times \left(-393.5 \text{kJ/mol}\right) - (0 + \Delta H^\circ_{\text{C}_3 \text{H}_8}) \)

\[ \Delta H^\circ_{\text{C}_3 \text{H}_8} = -2286.4 \text{kJ} - 787 \times 1.5 \text{kJ} + 4.85 \times 10^3 \text{kJ} \]

\[ \Delta H^\circ_{\text{C}_3 \text{H}_8} = -190.9 \text{kJ/mol} \]

\[ \Delta H^\circ_{\text{C}_3 \text{H}_8} = -191 \text{kJ/mol} \]

d) Heat released = 0.98 mol \( \Delta H^\circ_{\text{C}_3 \text{H}_8} \left(-1.85 \times 10^3 \text{kJ} \right) = -57.38 \text{kJ} \)

Thus, 57.4 kJ released as heat to the surroundings.

ii) \( \Delta T = \frac{52.4 \text{kJ}}{32738 \text{kg}} = 5.64^\circ\text{C} \)

\[ \Delta T = 5.64^\circ\text{C} \]

STOP

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The heat of combustion, \( \Delta H_{\text{comb}} \), for one mole of C₇H₁₆(l) is \(-4.85 \times 10^3 \) kJ.

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(ii) Given that the total heat capacity of the calorimeter is 9.273 kJ °C⁻¹, calculate the temperature change of the calorimeter.

\[
\begin{align*}
\text{a) } C₇H₁₆ & \rightarrow 6gC \times \frac{1\text{ mol}}{12g} = 0.5\text{ mol} C \\
& \rightarrow 1.23444 \times \frac{1\text{ mol}}{8.064} = 1.4424\text{ mol H}_.
\end{align*}
\]

\[
\begin{align*}
\text{Lp: emp. formula } &= C₂H₈. \\
\text{(i) density } &= \frac{1.962 g}{11} \text{ emp. weight } = 3(12) + 8(1) = 44 g. \\
\text{Molar mass } &= \frac{44 g}{11} = 4.00 g/mol \\
\text{(ii) emp. weight } &= 44 g. \\
\text{Emp. ratio } &= \frac{\text{empirical weight}}{\text{molecular weight}} = \frac{44}{43.904} \approx 1:1.
\end{align*}
\]

GO ON TO THE NEXT PAGE.
c) \( \Delta H^\circ = \Delta H (\text{product}) - \Delta H (\text{reactant}) \)

\[
-4.85 \times 10^3 = \pm (-393.5) + 8 (-265.8) - 2 \checkmark
\]

\[-4.85 \times 10^3 = -5040.9 - x.\]

\[-x = -4.85 \times 10^3 + 5040.9 = 190.9\]

\[x = -190.9 \text{ K} = \Delta H^\circ (\text{Cu}_1\text{O})\]

\[\Delta H^\circ (\text{product}) = -4.85 \times 10^3 \text{ kJ/mol} \]

\[
\Delta H = 0.0108n \times -4.85 \times 10^3 \text{ kJ/mol} = -52.38 \text{ kJ}
\]

\[\Delta H = \Delta U = -52.38 \text{ kJ} \]

\[(ii)\] \( CP = \frac{\Delta H}{\Delta T} = \frac{-52.38 \text{ kJ}}{10} \checkmark \]

\[\Delta T = \frac{\Delta U}{CP} = \frac{-52.38 \text{ kJ}}{10 \text{ kJ/mol} \cdot K} = -5.23 \text{ mol} \cdot \text{K} = \Delta T.\]

---

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\[ \rho V = nRT \]

\[ \rho \frac{\Delta T}{\Delta V} = \frac{nR}{V} \]

\[ M_r = \frac{RT}{\rho} \]

\[ = (0.0821) (25+273) \left( \frac{146}{11.09} \right) = 43.99 ^\circ C \]

GO ON TO THE NEXT PAGE.
\[ C_{5}H_{13} = \text{CH}_{3}CH(\text{CH}_{3})CH_{2}CH_{3} \]

\[ \frac{2}{3} + \frac{4}{4} = \frac{5}{3} \]

\[ 44 \div 93 = 0.48 \approx \frac{3}{5} \]

\[ C_{5}H_{13} \]

\[ C_{3}H_{8} \]

\[ \Delta H^\circ \text{comp} = [8(-285.8) + 9(-393.5)] J - \Delta H^\circ C_{7}H_{16} \]

\[ \Delta H^\circ C_{7}H_{16} = -5040.9 + 4085 \times 10^3 \]

\[ = -1901.9 \text{ MJ/mol} \]

\[ \Delta \text{H} = \text{heat released} \times \frac{1}{\text{mol}} \]

\[ m(C_{7}H_{16}) \cdot \Delta T = \frac{q}{m} \]

\[ = (0.0108 \text{ mol}) \cdot (108) \cdot \frac{521.4}{0.0108} \text{ J} \]

\[ = 0.566 \text{ KJ} \text{ released} = 0.566 \text{ KJ} \]

\[ (108)(9.293 \text{ KJ}) \]

\[ = 0.566 \text{ KJ} \]

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\[ \text{STOP} \]

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