



## AP<sup>®</sup> Chemistry 2003 Sample Student Responses Form B

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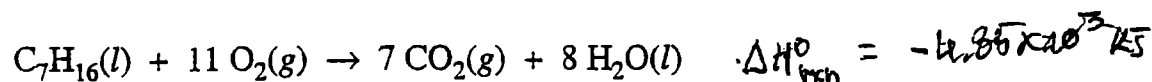
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3. In an experiment, a sample of an unknown, pure gaseous hydrocarbon was analyzed. Results showed that the sample contained 6.000 g of carbon and 1.344 g of hydrogen.

- (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<sup>-1</sup>.
- (i) Calculate the molar mass of the hydrocarbon.
- (ii) Determine the molecular formula of the hydrocarbon.

In another experiment, liquid heptane, C<sub>7</sub>H<sub>16</sub>(l), is completely combusted to produce CO<sub>2</sub>(g) and H<sub>2</sub>O(l), as represented by the following equation.



The heat of combustion,  $\Delta H_{\text{comb}}^\circ$ , for one mole of C<sub>7</sub>H<sub>16</sub>(l) is  $-4.85 \times 10^3$  kJ.

- (c) Using the information in the table below, calculate the value of  $\Delta H_f^\circ$  for C<sub>7</sub>H<sub>16</sub>(l) in kJ mol<sup>-1</sup>.

Compound	$\Delta H_f^\circ$ (kJ mol <sup>-1</sup> )
CO <sub>2</sub> (g)	-393.5
H <sub>2</sub> O(l)	-285.8

- (d) A 0.0108 mol sample of C<sub>7</sub>H<sub>16</sub>(l) is combusted in a bomb calorimeter.
- (i) Calculate the amount of heat released to the calorimeter.
- (ii) Given that the total heat capacity of the calorimeter is 9.273 kJ °C<sup>-1</sup>, calculate the temperature change of the calorimeter.

$$\begin{aligned} a - m_c &= 6.000 \text{ g} \left( \frac{1 \text{ mol C}}{12.011 \text{ g}} \right) = 0.4995 \text{ mol C} \\ m_H &= 1.344 \text{ g} \left( \frac{1 \text{ mol H}}{1.0079 \text{ g}} \right) = 1.333 \text{ mol H} \end{aligned} \quad \left. \begin{array}{l} \\ \end{array} \right\} \text{C}_{0.4995} \text{H}_{1.333} \Rightarrow \text{C}_{\frac{1}{2}} \text{H}_{\frac{4}{3}}$$

$$\Rightarrow \boxed{\text{C}_3 \text{H}_8 = \text{empirical formula}}$$

$$b - T = 273 + 25 = 298 \text{ K} \quad d = 1.96 \text{ g/L} \quad PV = nRT$$

$$P = 1.09 \text{ atm}$$

$$i) \text{ Molar mass} = \frac{dRT}{P} = \frac{1.96 \text{ g/L} \times (0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}) (298 \text{ K})}{1.09 \text{ atm}}$$

$$\text{Molar mass} = 43.97 \text{ g/mol}$$

$$\boxed{m_m = 44.0 \text{ g/mol}}$$

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ii) Molecular formula:

$$44.0g = k(3(12.011g) + 8(1.0079g))$$

$$44.0g = k(44.0962g)$$

$$k = \frac{44.0g}{44.0962g} = 0.998 \approx 1$$

Molecular formula is  $C_3H_8$  since  $k$  is 1.

$$c) -4.85 \times 10^3 \text{ kJ} = (8 \text{ mol} \times (-285.8 \text{ kJ/mol}) + 7 \text{ mol} \times (-393.5 \text{ kJ/mol})) - (0 + \Delta H_f^\circ \text{ of } C_7H_{16})$$

$$\Delta H_f^\circ \text{ of } C_7H_{16} = -2286.4 \text{ kJ} - 2754.5 \text{ kJ} + 4.85 \times 10^3 \text{ kJ}$$

$$\Delta H_f^\circ \text{ of } C_7H_{16} = -190.9 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{ of } C_7H_{16} = -191 \text{ kJ/mol}$$

$$d) \text{ Heat released} = 0.0108 \text{ mol } C_7H_{16} \left( \frac{-4.85 \times 10^3 \text{ kJ}}{\text{mol}} \right) = -52.38 \text{ kJ}$$

Thus 52.4 kJ released as heat to the surroundings.

$$ii) \Delta T = \frac{52.4 \text{ kJ}}{9.273 \text{ kJ}} \left( \frac{1^\circ \text{C}}{1} \right) = 5.6486^\circ \text{C}$$

$$\Delta T = 5.65^\circ \text{C}$$

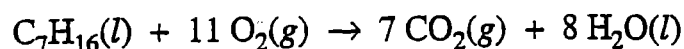
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3. In an experiment, a sample of an unknown, pure gaseous hydrocarbon was analyzed. Results showed that the sample contained 6.000 g of carbon and 1.344 g of hydrogen.

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a) C<sub>x</sub>H<sub>y</sub> → 6g C ×  $\frac{1n}{12g} = 0.5nC$       (0.5 : 1.344) ÷ 0.5  
 → 1.344H ×  $\frac{1n}{1.008} = 1.344nH$       (1 : 2.688) × 3  
 3 : 8.064 ≈ 3 : 8  
 ∴ emp. formula = C<sub>3</sub>H<sub>8</sub>.

(b)(i) density =  $\frac{1.96g}{1L}$  emp. weight = 3(12) + 8(1) = 44g.  
 $M_p = n \cdot \text{mass} = g/n$   
 ∴ @ STP, 1 mol gas = 22.4L →  $\frac{1.96g}{1L} \times \frac{22.4L}{1n} = 43.904g/n$   
 (ii) emp. weight = 44g.  
 ∴ ratio (empirical weight : molecular weight) = 44 : 43.904  
 ≈ 1 : 1  
 ∴ molecular formula = C<sub>3</sub>H<sub>8</sub>.

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$$c) \Delta H_f^\circ = \Delta H(\text{product}) - \Delta H(\text{reactant}) \quad \Delta H_f^\circ(C_7H_{16})$$

$$-4.85 \times 10^3 = (7(-393.5) + 8(-285.8)) - X$$

$$-4.85 \times 10^3 = -5040.9 - X$$

$$-X = -4.85 \times 10^3 + 5040.9 = 190.9$$

$$X = -190.9 \text{ kJ/mol} = \Delta H_f^\circ(C_7H_{16})$$

$$d) (i) \Delta H_{\text{comb}} = -4.85 \times 10^3 \text{ kJ/mol}$$

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

$$(ii) C_p = \frac{\Delta H}{\Delta T} = (-52.38 \text{ kJ}) / X = \Delta T$$

$$\Delta X = \frac{\Delta H}{C_p} = (-52.38 \text{ kJ}) / (9.273 \text{ kJ/}^\circ\text{C}) = -5.65^\circ\text{C} = \Delta T$$

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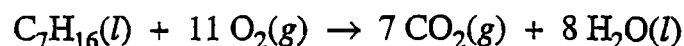
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3. (a) H<sub>x</sub>C<sub>y</sub> (C<sub>y</sub>H<sub>x</sub>)

1.344g    6.00g

C    H

= 0.15    1.34

= 1    2.68

= 5    13

C<sub>5</sub>H<sub>13</sub>

$$n(C) = \frac{6.00}{12.01} = 0.5 \text{ mol}$$

$$n(H) = \frac{1.344}{1.0079} = 1.34 \text{ mol}$$

(b) (i)  $PV = nRT$

$$d = \frac{m}{V} = \frac{1.96 \text{ g}}{L}$$

$$P \frac{m}{d} = \frac{n}{M_r} RT$$

$$M_r = RT \cdot \frac{d}{P}$$

$$= (0.0821) (25 + 273) \left( \frac{1.96}{1.09} \right) = 43.9 \approx 44.0 \text{ g}$$

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3C<sub>2</sub>

(i)  $C_5H_{13} = 739$

~~$73 = 44$~~

$44 \div 73 = 0.60 \approx \frac{3}{5}$

$C \ 5 \cdot \frac{3}{5} \ ; \ H \ 13 \cdot \frac{3}{5}$

$\approx C_3H_8$

(c)  $\Delta H_{comb}^{\circ} = [8(-285.8) + 9(-393.5)] - \Delta H_f^{\circ} C_7H_{16}$

$\Delta H_f^{\circ} C_7H_{16} = -5040.9 + 4.85 \times 10^3$

$= -190.9 \text{ kJ/mol}$

(d)

(i)  $\therefore -4.85 \times 10^3 \text{ kJ} = 0.0108 \text{ mol} \times x$

$\frac{-4.85 \times 10^3 \text{ kJ}}{0.0108 \text{ mol}} = x = (0.0108) (-4.85 \times 10^3)$

$= -52.38$

$\approx -52.4 \text{ kJ}$

$x = \frac{0.0108 \text{ mol} \times -4.85 \times 10^3 \text{ kJ}}{1 \text{ mol}}$

(ii)  $q = m C \Delta T$

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

$m(C_7H_{16}) \quad \Delta T = \frac{q}{m C}$

$= (0.0108) (100)$

$52.4 = \text{heat released} \times \frac{1}{0.0108}$

$\approx 1.089$

$= 0.566 \text{ kJ}$

heat released = 0.566 kJ

$(1.089) \left( \frac{9.273 \text{ kJ}}{^{\circ}\text{C}} \right)$

$= 0.0565 ^{\circ}\text{C}$

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