

Integrative Problem of Hess' Law and Heat Evolution

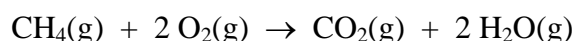
Problem 7-85

A particular natural gas consists, in mole percents, of 83.0%CH₄, 11.2%C₂H₆, and 5.8% C₃H₈. A 385-L sample of this gas, measured at 22.6°C and 739 mmHg, is burned at constant pressure in an excess of oxygen gas. How much heat, in kilojoules, is evolved in the combustion reaction?

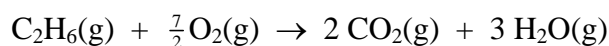
This multi-step problem requires calculation of:

- 1) $\Delta H_{\text{combustion}}$ for each reaction
- 2) total quantity of gas (in moles)
- 3) the quantity, in moles, of each gas
- 4) the quantity of heat produced by each gas and summing the individual heats together

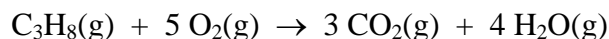
1) There's no table of $\Delta H_{\text{combustion}}$ in the textbook so use Hess' law to calculate the $\Delta H_{\text{combustion}}$ of each fuel gas from $\Delta H_{\text{formation}}$ (Table 7.2).



$$\Delta H_{\text{combustion}}^{\text{CH}_4} = (-393.5 \text{ kJ}) + (2 \times -241.8 \text{ kJ}) - (-74.81 \text{ kJ}) = -802.31 \text{ kJ}(/\text{molCH}_4)$$



$$\Delta H_{\text{combustion}}^{\text{C}_2\text{H}_6} = ((2 \times -393.5 \text{ kJ}) + (3 \times -241.8 \text{ kJ})) - (-84.68 \text{ kJ}) = -1427.7 \text{ kJ}(/\text{molC}_2\text{H}_6)$$



$$\Delta H_{\text{combustion}}^{\text{C}_3\text{H}_8} = ((3 \times -393.5 \text{ kJ}) + (4 \times -241.8 \text{ kJ})) - (-103.8 \text{ kJ}) = -2043.9 \text{ kJ}(/\text{molC}_2\text{H}_6)$$

2) Calculate the total moles of gas and 3) individual moles of each fuel.

$$V = 385 \text{ L} \quad T = 22.6^\circ\text{C} = 295.8 \text{ K} \quad P = 739 \text{ mmHg} = 0.9724 \text{ atm}$$

$$n = \frac{PV}{RT} = \frac{(0.9724 \text{ atm})(385 \text{ L})}{(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(295.8 \text{ K})} = 15.43 \text{ mol gas}$$

$$n_{\text{CH}_4} = 0.830 \times 15.43 \text{ mol} = 12.80 \text{ mol CH}_4$$

$$n_{\text{C}_2\text{H}_6} = 0.112 \times 15.43 \text{ mol} = 1.728 \text{ mol C}_2\text{H}_6$$

$$n_{\text{C}_3\text{H}_8} = 0.058 \times 15.43 \text{ mol} = 0.8949 \text{ mol C}_3\text{H}_8$$

4) Calculate the energy produced by the combustion of each and the total energy produced.

$$\Delta H_{\text{combustion}}^{\text{CH}_4} = -802.31 \text{ kJ/molCH}_4 \quad q_{\text{CH}_4} = n \cdot \Delta H_{\text{combustion}}^{\text{CH}_4} = 12.80 \text{ mol} \times -802.31 \text{ kJ/mol CH}_4 = -10,270 \text{ kJ}$$

$$\Delta H_{\text{combustion}}^{\text{C}_2\text{H}_6} = -1427.7 \text{ kJ/molC}_2\text{H}_6 \quad q_{\text{C}_2\text{H}_6} = n \cdot \Delta H_{\text{combustion}}^{\text{C}_2\text{H}_6} = 1.728 \text{ mol} \times -1427.7 \text{ kJ/molC}_2\text{H}_6 = -2,467 \text{ kJ}$$

$$\Delta H_{\text{combustion}}^{\text{C}_3\text{H}_8} = -2043.9 \text{ kJ/molC}_2\text{H}_6 \quad q_{\text{C}_3\text{H}_8} = n \cdot \Delta H_{\text{combustion}}^{\text{C}_3\text{H}_8} = 0.8949 \text{ mol} \times -2043.9 \text{ kJ/molC}_2\text{H}_6 = -1829 \text{ kJ}$$

$$q_{\text{total}} = -14,600 \text{ kJ}$$