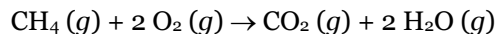


Exercise 03Name: _____ **Key**

Consider the combustion of methane in air:

(a) Determine the ΔH_{rxn} using the values of ΔH_f° given.

| Species | ΔH_f° (kJ/mol) |
|-------------------------|-----------------------------|
| $\text{CH}_4(g)$ | -74.6 |
| $\text{H}_2\text{O}(g)$ | -241.8 |
| $\text{CO}_2(g)$ | -393.5 |

(b) If 1.00×10^6 J of heat were released during a combustion of methane, how many moles of CO_2 gas were produced?

(a) We can determine the heat of the reaction (or heat of combustion) through the following expression.

Recall that $\Delta H_f^\circ = 0$ kJ/mol for elements in their elemental states (e.g., O_2 gas).

$$\begin{aligned}\Delta H_{\text{rxn}} &= n_{\text{CO}_2} \Delta H_{f,\text{CO}_2}^\circ + n_{\text{O}_2} \Delta H_{f,\text{O}_2}^\circ - n_{\text{CH}_4} \Delta H_{f,\text{CH}_4}^\circ - n_{\text{O}_2} \Delta H_{f,\text{O}_2}^\circ \\ &= (1 \text{ mol CO}_2) \left(-393.5 \frac{\text{kJ}}{\text{mol}} \right) + (2 \text{ mol H}_2\text{O}) \left(-241.8 \frac{\text{kJ}}{\text{mol}} \right) - (1 \text{ mol CH}_4) \left(-74.6 \frac{\text{kJ}}{\text{mol}} \right) - \left(0 \frac{\text{kJ}}{\text{mol}} \right) \\ \Delta H_{\text{rxn}} &= -802.5 \text{ kJ/mol}\end{aligned}$$

(b) The heat of reaction/combustion (ΔH_{rxn}) tells us that for every 1 mol $\text{CH}_4(g)$ combusted, 802.5 kJ of heat are released. From stoichiometry, we also know that 1 mol of $\text{CH}_4(g)$ produces 1 mol $\text{CO}_2(g)$. So, if we released 1.00×10^6 J:

$$1.00 \times 10^6 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} \times \frac{1 \text{ mol CH}_4}{802.5 \text{ kJ}} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol CH}_4} = \mathbf{1.25 \text{ mol CO}_2}$$