Exercise 03 **Name:** Key

Consider the combustion of methane in air:

$$CH_4\ (g) + 2\ O_2\ (g) \rightarrow CO_2\ (g) + 2\ H_2O\ (g)$$

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

Species	ΔH_f^o (kJ/mol)
$CH_4\ (g)$	−74.6
$H_2O\ (g)$	−241.8
$CO_2\ (g)$	−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 $\Delta H_f^o = 0\ \text{kJ/mol}$ for elements in their elemental states (e.g., O_2 gas).

$$\begin{aligned}\Delta H_{\text{rxn}} &= n_{CO_2}\Delta H^o_{f,CO_2} + n_{O_2}\Delta H^o_{f,O_2} - n_{CH_4}\Delta H^o_{f,CH_4} - n_{O_2}\Delta H^o_{f,O_2} \\ &= (1\ \text{mol CO}_2)\left(-393.5\frac{\text{kJ}}{\text{mol}}\right) + (2\ \text{mol H}_2\text{O})\left(-241.8\frac{\text{kJ}}{\text{mol}}\right) - (1\ \text{mol CH}_4)\left(-74.6\frac{\text{kJ}}{\text{mol}}\right) - \left(0\frac{\text{kJ}}{\text{mol}}\right) \\ \Delta H_{\text{rxn}} &= -802.5\ \text{kJ/mol}\end{aligned}$$

(b) The heat of reaction/combustion (ΔH_{rxn}) tells us that for every 1 mol CH_4 (*g*) combusted, 802.5 kJ of heat are released. From stoichiometry, we also know that 1 mol of CH_4 (*g*) produces 1 mol CO_2 (*g*). So, if we released 1.00×10^6 J:

$$1.00 \times 10^6\ \text{J} \times \frac{1\ \text{kJ}}{1000\ \text{J}} \times \frac{1\ \text{mol CH}_4}{802.5\ \text{kJ}} \times \frac{1\ \text{mol CO}_2}{1\ \text{mol CH}_4} = \mathbf{1.25\ mol\ CO_2}$$