



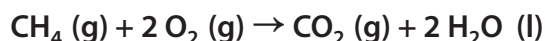
# Connected Chemistry

## Thermodynamics Unit

### Student Appendix A: Calculating Hess's Law

#### Sample Calculation

Methane gas is a small hydrocarbon fossil fuel which combusts when oxygen and a spark are added. The balanced equation for the combustion of methane and oxygen is



1. Look at table in "Appendix B: Enthalpy Data" on [page 93](#) for enthalpy values.

$$\Delta H_f^\circ \text{CH}_4 (\text{g}) = -74.81 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{ of O}_2 (\text{g}) = 0 \text{ kJ/mol (recall that pure oxygen is a diatomic element)}$$

$$\Delta H_f^\circ \text{ of CO}_2 (\text{g}) = -393.509 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{ of H}_2\text{O} (\text{l}) = -285.83 \text{ kJ/mol}$$

2. Make sure the equation is balanced and multiply by the coefficients. Sum the products and reactants separately.

#### Products

$$\Sigma \Delta H_{f \text{ products}} = \Delta H_f \text{ of CO}_2 (\text{g}) + \Delta H_f \text{ of H}_2\text{O} (\text{l})$$

$$\Delta H_f \text{ of CO}_2 (\text{g}) = \Delta H_f^\circ \text{ of CO}_2 (\text{g}) \times 1 \text{ mol of CO}_2 (\text{g}) = -393.509 \text{ kJ/mol} \times 1 \text{ mol CO}_2 (\text{g})$$

$$= -393.509 \text{ kJ}$$

$$\Delta H_f \text{ of H}_2\text{O} (\text{l}) = \Delta H_f^\circ \text{ of H}_2\text{O} (\text{l}) \times 2 \text{ mol of H}_2\text{O} (\text{l}) = -285.83 \text{ kJ/mol} \times 2 \text{ mol H}_2\text{O} (\text{l})$$

$$= -571.66 \text{ kJ}$$

$$\Sigma \Delta H_{f \text{ products}} = -393.509 \text{ kJ} + (-571.66 \text{ kJ}) = -965.17 \text{ kJ}$$

## Reactants

$$\Sigma \Delta H_{f \text{ reactants}} = \Delta H_f \text{ of CH}_4(\text{g}) + \Delta H_f^\circ \text{ of O}_2(\text{g})$$

$$\begin{aligned}\Delta H_f \text{ of CH}_4(\text{g}) &= \Delta H_f^\circ \text{ of CH}_4(\text{g}) \times 1 \text{ mol CH}_4(\text{g}) = -74.8 \text{ kJ/mol} \times 1 \text{ mol CH}_4(\text{g}) \\ &= -74.8 \text{ kJ}\end{aligned}$$

$$\begin{aligned}\Delta H_f \text{ of O}_2(\text{g}) &= \Delta H_f \text{ of O}_2(\text{g}) \times 2 \text{ mol O}_2(\text{g}) = 0 \text{ kJ/mol} \times 2 \text{ mol O}_2(\text{g}) \\ &= 0 \text{ kJ}\end{aligned}$$

$$\Sigma \Delta H_{f \text{ reactants}} = -74.8 \text{ kJ} + 0 \text{ kJ} = -74.8 \text{ kJ}$$

3. Use Hess's Law to calculate  $\Delta H_{rxn}$

$$\Delta H_{rxn} = \Sigma \Delta H_{f \text{ products}} - \Sigma \Delta H_{f \text{ reactants}}$$

$$-965.1 \text{ kJ} - (-74.8 \text{ kJ}) = -890.3 \text{ kJ}$$

$$\Delta H_{rxn} = -890.3 \text{ kJ}$$



# Connected Chemistry

## Thermodynamics Unit

### Student Appendix B: Enthalpy Data

Species	$\Delta H_f^\circ$ kJ/mol
H <sub>2</sub> (g)	0
O <sub>2</sub> (g)	0
NH <sub>3</sub> (g)	-45.95
NO (g)	90.25
CH <sub>4</sub> (g)	-74.81
H <sub>2</sub> O (g)	-241.83
H <sub>2</sub> O (l)	-285.83
CO <sub>2</sub> (g)	-393.508
C <sub>3</sub> H <sub>8</sub> (g)	-104.7
C <sub>5</sub> H <sub>12</sub> (l)	-146.44
Na <sup>+</sup> (aq)	-239.7
HCO <sub>3</sub> <sup>-</sup> (aq)	-691.1
H <sup>+</sup> (aq)	0
C <sub>2</sub> H <sub>3</sub> O <sub>2</sub> <sup>-</sup> (aq)	-483.4
H <sub>2</sub> O <sub>2</sub> (l)	-191.17

Kotz, J. C., Treichel, P., & Weaver, G. C. (2006). Chemistry & chemical reactivity. Belmont, CA: Thomson Brooks/Cole.

Masterton, Slowinski & Stanitski (1983). Chemical Principles. CBS Publishing.

