



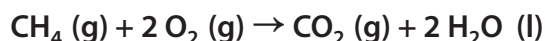
Connected Chemistry

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**

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

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

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

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

3. Use Hess's Law to calculate ΔH_{rxn}

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

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

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



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Student Appendix B: Enthalpy Data

Species	ΔH_f° kJ/mol
H ₂ (g)	0
O ₂ (g)	0
NH ₃ (g)	-45.95
NO (g)	90.25
CH ₄ (g)	-74.81
H ₂ O (g)	-241.83
H ₂ O (l)	-285.83
CO ₂ (g)	-393.509
C ₃ H ₈ (g)	-104.7
C ₅ H ₁₂ (l)	-173.5
Na ⁺ (aq)	-239.7
HCO ₃ ⁻ (aq)	-691.1
H ⁺ (aq)	0
C ₂ H ₃ O ₂ ⁻ (aq)	-486
H ₂ O ₂ (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.

