

Calculate $\Delta H^\circ_{\text{rxn}}$ using Standard Enthalpy of Formation

When all the standard enthalpy values of formation for the species in a chemical reaction are known at standard conditions, the following equation can be used to calculate the enthalpy change for a reaction.

$$\Delta H^\circ_{\text{rxn}} = \sum \Delta H^\circ_{\text{f}(\text{products})} - \sum \Delta H^\circ_{\text{f}(\text{reactants})}$$

$\Delta H^\circ_{\text{f}(\text{products})}$ are the enthalpy of formation values for the products times the number of moles of product formed in the chemical equation. $\Delta H^\circ_{\text{f}(\text{reactants})}$ are the enthalpy of formation values for the reactants times the number of moles of reactant used in the chemical equation.

Standard conditions usually a pressure of 1 bar, temperature of 25 °C, concentration of 1M, and all species in their standard states. Refer to the information in the table for specific information about standard conditions. The $\Delta H^\circ_{\text{f}}$ is 0 for an element in its standard state. For example, $\text{O}_2(\text{g})$ and $\text{C}(\text{s, graphite})$ would have $\Delta H^\circ_{\text{f}}$ values of 0, but the $\Delta H^\circ_{\text{f}}$ value for $\text{C}(\text{s, diamond})$ would not be zero because diamond is not the standard state of carbon. The enthalpy value for $\text{C}(\text{diamond})$ is 1.8kJ/mol.

Example: Calculate the enthalpy change for the following reaction using standard enthalpy values found in a table of thermodynamic values.



Solution: Look up the values for $\text{C}_2\text{H}_6(\text{g})$, $\text{CO}_2(\text{g})$, and $\text{H}_2\text{O}(\text{g})$ in a table of thermodynamic values. The enthalpy of formation, $\Delta H^\circ_{\text{f}}$, for $\text{O}_2(\text{g})$ is 0 because it is an element in its standard state.

$$\Delta H^\circ_{\text{f}}[\text{C}_2\text{H}_6(\text{g})] = -83.85 \text{ kJ/mol}$$

$$\Delta H^\circ_{\text{f}}[\text{O}_2(\text{g})] = 0 \text{ kJ/mol}$$

$$\Delta H^\circ_{\text{f}}[\text{CO}_2(\text{g})] = -393.51 \text{ kJ/mol}$$

$$\Delta H^\circ_{\text{f}}[\text{H}_2\text{O}(\text{g})] = -241.83 \text{ kJ/mol}$$

$$\Delta H^\circ_{\text{rxn}} = \{6 \times \Delta H^\circ_{\text{f}}[\text{H}_2\text{O}(\text{g})] + 4 \times \Delta H^\circ_{\text{f}}[\text{CO}_2(\text{g})]\} - \{2 \times \Delta H^\circ_{\text{f}}[\text{C}_2\text{H}_6(\text{g})] + 7 \times \Delta H^\circ_{\text{f}}[\text{O}_2(\text{g})]\}$$

$$\Delta H^\circ_{\text{rxn}} = \{6 \text{ mol} \times -241.83 \text{ kJ/mol} + 4 \text{ mol} \times -393.51 \text{ kJ/mol}\} - \{2 \text{ mol} \times -83.85 \text{ kJ/mol} + 7 \text{ mol} \times 0 \text{ kJ/mol}\}$$

$$\Delta H^\circ_{\text{rxn}} = -2587.32 \text{ kJ}$$

Like all enthalpies of reaction, $\Delta H^\circ_{\text{rxn}}$ means $-2587.32 \text{ kJ} / 2 \text{ mol C}_2\text{H}_6$. Similar relationships apply for the other reactants and products (-2587.32 kJ per 7 mol of $\text{O}_2(\text{g})$ used, per 4 mol $\text{CO}_2(\text{g})$ formed, and per 6 mol $\text{H}_2\text{O}(\text{g})$ of formed).