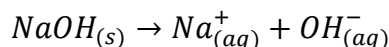


## Answers to Exercise 6.6 Free Energy Calculations

1. **Step 1: Write a balanced chemical equation for the reaction**



- Step 2: Calculate standard free energy change for the reaction from the standard free energies of formation**

$$\Delta_r G^\circ = \sum \Delta_f G^\circ(\text{products}) - \sum \Delta_f G^\circ(\text{reactants})$$

$$\Delta_r G^\circ = [\Delta_f G^\circ(Na^+_{(aq)}) + \Delta_f G^\circ(OH^-_{(aq)})] - \Delta_f G^\circ(NaOH_{(s)})$$

$$\Delta_r G^\circ = \left[ \left( -261.87 \frac{\text{kJ}}{\text{mol}} \right) + \left( -157.30 \frac{\text{kJ}}{\text{mol}} \right) \right] - \left( -379.53 \frac{\text{kJ}}{\text{mol}} \right)$$

$$\Delta_r G^\circ = -39.64 \frac{\text{kJ}}{\text{mol}}$$

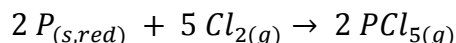
- Step 3: Check your work**

*Does your answer seem reasonable? Are sig. fig. correct?*

*The standard free energy change for this reaction is negative, so the reaction is thermodynamically allowed. Given that we know that sodium hydroxide dissolves in water at room temperature, this answer seems reasonable*

2. *Since the standard free energy values were not provided, the standard free energy change for the reaction will have to be calculated from the standard enthalpy change and the standard entropy change.*

- Step 1: Write a balanced chemical equation for the reaction**



- Step 2: Calculate standard enthalpy change for the reaction from the standard enthalpies of formation**

$$\Delta_r H^\circ = \sum \Delta_f H^\circ(\text{products}) - \sum \Delta_f H^\circ(\text{reactants})$$

$$\Delta_r H^\circ = 2 \Delta_f H^\circ(PCl_{5(g)}) - [2 \Delta_f H^\circ(P_{(s,red)}) + 5 \Delta_f H^\circ(Cl_{2(g)})]$$

$$\Delta_r H^\circ = 2 \left( -402 \frac{\text{kJ}}{\text{mol}} \right) - \left[ 2 \left( -17.6 \frac{\text{kJ}}{\text{mol}} \right) + 5 \left( 0 \frac{\text{kJ}}{\text{mol}} \right) \right]$$

$$\Delta_r H^\circ = -769 \frac{\text{kJ}}{\text{mol}} \quad -768.8$$

- Step 3: Calculate standard entropy change for the reaction from the standard entropies**

$$\Delta_r S^\circ = \sum \Delta_f S^\circ(\text{products}) - \sum \Delta_f S^\circ(\text{reactants})$$

$$\Delta_r S^\circ = 2 S^\circ(PCl_{5(g)}) - [2 S^\circ(P_{(s,red)}) + 5 S^\circ(Cl_{2(g)})]$$

$$\Delta_r S^\circ = 2 \left( 353 \frac{\text{J}}{\text{mol}\cdot\text{K}} \right) - \left[ 2 \left( 22.8 \frac{\text{J}}{\text{mol}\cdot\text{K}} \right) + 5 \left( 223.0 \frac{\text{J}}{\text{mol}\cdot\text{K}} \right) \right]$$

$$\Delta_r S^\circ = -455 \frac{\text{J}}{\text{mol}\cdot\text{K}} \quad -454.6$$

**Step 4: Calculate standard free energy change for the reaction from the standard enthalpy change and the standard entropy change**

$$\Delta_r G^\circ = \Delta_r H^\circ - T \Delta_r S^\circ$$

$$\Delta_r G^\circ = \left(-769 \frac{\text{kJ}}{\text{mol}}\right) - (298.15 \text{ K}) \left(-455 \frac{\text{J}}{\text{mol}\cdot\text{K}}\right) \left(\frac{1 \text{ kJ}}{1000 \text{ J}}\right)$$

$$\Delta_r G^\circ = \left(-769 \frac{\text{kJ}}{\text{mol}}\right) - \left(-136 \frac{\text{kJ}}{\text{mol}}\right)$$

$$\Delta_r G^\circ = -633 \frac{\text{kJ}}{\text{mol}}$$

**Step 5: Check your work**

*Does your answer seem reasonable? Are sig. fig. correct?*

*The standard free energy change for this reaction is negative, so the reaction is thermodynamically allowed.*

*If you wrote the balanced chemical equation as  $\text{P}_{(\text{s,red})} + \frac{5}{2} \text{Cl}_{2(\text{g})} \rightarrow \text{PCl}_{5(\text{g})}$ , all of your answers should be equal to half of the ones listed in the answer key, but the overall conclusion should be the same.*

3.

(a) Yes. The standard free energy change for the reaction is negative.

(b)  $\Delta_r G^\circ = \sum \Delta_f G^\circ(\text{products}) - \sum \Delta_f G^\circ(\text{reactants})$

$$\Delta_r G^\circ = \Delta_f G^\circ(\text{C}_2\text{H}_6(\text{g})) - [\Delta_f G^\circ(\text{C}_2\text{H}_4(\text{g})) + \Delta_f G^\circ(\text{H}_2(\text{g}))]$$

$$\Delta_r G^\circ = \Delta_f G^\circ(\text{C}_2\text{H}_6(\text{g})) - \Delta_f G^\circ(\text{C}_2\text{H}_4(\text{g})) - \Delta_f G^\circ(\text{H}_2(\text{g}))$$

$$\Delta_f G^\circ(\text{C}_2\text{H}_4(\text{g})) = \Delta_f G^\circ(\text{C}_2\text{H}_6(\text{g})) - \Delta_f G^\circ(\text{H}_2(\text{g})) - \Delta_r G^\circ$$

$$\Delta_f G^\circ(\text{C}_2\text{H}_4(\text{g})) = \left(-32.89 \frac{\text{kJ}}{\text{mol}}\right) - \left(0 \frac{\text{kJ}}{\text{mol}}\right) - \left(-101.25 \frac{\text{kJ}}{\text{mol}}\right)$$

$$\Delta_f G^\circ(\text{C}_2\text{H}_4(\text{g})) = 68.36 \frac{\text{kJ}}{\text{mol}}$$

4.

(a) Yes. The standard free energy change for the reaction is negative.

(b) What is the standard free energy of formation for cyclohexane?

$$\Delta_r G^\circ = \sum \Delta_f G^\circ(\text{products}) - \sum \Delta_f G^\circ(\text{reactants})$$

$$\Delta_r G^\circ = \Delta_f G^\circ(\text{C}_6\text{H}_{12}(\text{l})) - [\Delta_f G^\circ(\text{C}_6\text{H}_6(\text{l})) + 3 \Delta_f G^\circ(\text{H}_2(\text{g}))]$$

$$\Delta_r G^\circ = \Delta_f G^\circ(\text{C}_6\text{H}_{12}(\text{l})) - \Delta_f G^\circ(\text{C}_6\text{H}_6(\text{l})) - 3 \Delta_f G^\circ(\text{H}_2(\text{g}))$$

$$\Delta_f G^\circ(\text{C}_6\text{H}_{12}(\text{l})) = \Delta_r G^\circ + \Delta_f G^\circ(\text{C}_6\text{H}_6(\text{l})) + 3 \Delta_f G^\circ(\text{H}_2(\text{g}))$$

$$\Delta_f G^\circ(\text{C}_6\text{H}_{12}(\text{l})) = \left(-97.6 \frac{\text{kJ}}{\text{mol}}\right) + \left(124.5 \frac{\text{kJ}}{\text{mol}}\right) + \left(0 \frac{\text{kJ}}{\text{mol}}\right)$$

$$\Delta_f G^\circ(\text{C}_6\text{H}_{12}(\text{l})) = 26.9 \frac{\text{kJ}}{\text{mol}}$$