

8.5: Standard Gibbs Energy Change, ΔG°

Learning Objectives

- Define Gibbs free energy, and describe its relation to spontaneity
- Calculate free energy change for a process using free energies of formation for its reactants and products
- Calculate free energy change for a process using enthalpies of formation and the entropies for its reactants and products

The Gibbs free energy (G), often called simply free energy, was named in honor of J. Willard Gibbs (1838–1903), an American physicist who first developed the concept. It is defined in terms of three other state functions with which you are already familiar: enthalpy, temperature, and entropy:

$$G = H - TS \quad (8.5.1)$$

Because it is a combination of state functions, G is also a [state function](#).

The criterion for predicting spontaneity is based on ΔG , the change in G , at constant temperature and pressure. Although very few chemical reactions actually occur under conditions of constant temperature and pressure, most systems can be brought back to the initial temperature and pressure without significantly affecting the value of thermodynamic state functions such as G . At constant temperature and pressure,

$$\Delta G = \Delta H - T\Delta S \quad (8.5.2)$$

where all thermodynamic quantities are those of the *system*. Under standard conditions Equation 8.5.2 is then expressed as

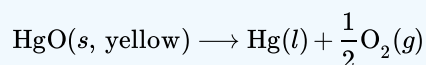
$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ \quad (8.5.3)$$

Since G is a state function, ΔG° can be obtained from the *standard free-energy of formation* values in [Table T1](#) (or [T2](#)) via the similar relationship used to calculate other state functions like ΔH° and ΔS° :

$$\Delta G^\circ = \sum n\Delta G_f^\circ (\text{products}) - \sum m\Delta G_f^\circ (\text{reactants}) \quad (8.5.4)$$

✓ Example 8.5.1: Calculation of ΔG_{298}°

Consider the decomposition of yellow mercury(II) oxide.



Calculate the standard free energy change at room temperature, ΔG_{298}° , using (a) standard free energies of formation and (b) standard enthalpies of formation and standard entropies. Do the results indicate the reaction to be spontaneous or nonspontaneous under standard conditions? The required data are available in [Table T1](#).

Solution

The required data are available in [Table T1](#) and are shown here.

Compound	ΔG_f° (kJ/mol)	ΔH_f° (kJ/mol)	S_{298}° (J/K·mol)
HgO (s, yellow)	-58.43	-90.46	71.13
Hg(l)	0	0	75.9
O ₂ (g)	0	0	205.2

(a) Using free energies of formation:

$$\begin{aligned} \Delta G_{298}^\circ &= \sum \nu G S_{298}^\circ (\text{products}) - \sum \nu \Delta G_{298}^\circ (\text{reactants}) \\ &= \left[1\Delta G_{298}^\circ \text{Hg}(l) + \frac{1}{2}\Delta G_{298}^\circ \text{O}_2(g) \right] - 1\Delta G_{298}^\circ \text{HgO}(s, \text{yellow}) \end{aligned}$$

$$= \left[1 \text{ mol}(0 \text{ kJ/mol}) + \frac{1}{2} \text{ mol}(0 \text{ kJ/mol}) \right] - 1 \text{ mol}(-58.43 \text{ kJ/mol}) = 58.43 \text{ kJ/mol}$$

(b) Using enthalpies and entropies of formation:

$$\begin{aligned} \Delta H_{298}^\circ &= \sum \nu \Delta H_{298}^\circ(\text{products}) - \sum \nu \Delta H_{298}^\circ(\text{reactants}) \\ &= \left[1 \Delta H_{298}^\circ \text{Hg}(l) + \frac{1}{2} \Delta H_{298}^\circ \text{O}_2(g) \right] - 1 \Delta H_{298}^\circ \text{HgO}(s, \text{ yellow}) \\ &= [1 \text{ mol}(0 \text{ kJ/mol}) + \frac{1}{2} \text{ mol}(0 \text{ kJ/mol})] - 1 \text{ mol}(-90.46 \text{ kJ/mol}) = 90.46 \text{ kJ/mol} \\ \Delta S_{298}^\circ &= \sum \nu \Delta S_{298}^\circ(\text{products}) - \sum \nu \Delta S_{298}^\circ(\text{reactants}) \\ &= \left[1 \Delta S_{298}^\circ \text{Hg}(l) + \frac{1}{2} \Delta S_{298}^\circ \text{O}_2(g) \right] - 1 \Delta S_{298}^\circ \text{HgO}(s, \text{ yellow}) \\ &= \left[1 \text{ mol}(75.9 \text{ J/mol K}) + \frac{1}{2} \text{ mol}(205.2 \text{ J/mol K}) \right] - 1 \text{ mol}(71.13 \text{ J/mol K}) = 107.4 \text{ J/mol K} \end{aligned}$$

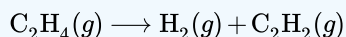
Now use these values in Equation 8.5.3 to get ΔG° :

$$\begin{aligned} \Delta G^\circ &= \Delta H^\circ - T \Delta S^\circ = 90.46 \text{ kJ} - 298.15 \text{ K} \times 107.4 \text{ J/K} \cdot \text{mol} \times \frac{1 \text{ kJ}}{1000 \text{ J}} \\ \Delta G^\circ &= (90.46 - 32.01) \text{ kJ/mol} = 58.45 \text{ kJ/mol} \end{aligned}$$

Both ways to calculate the standard free energy change at 25 °C give the same numerical value (to three significant figures), and both predict that the process is nonspontaneous (not spontaneous) at room temperature (since $\Delta G^\circ > 0$).

? Exercise 8.5.1

Calculate ΔG° using (a) free energies of formation and (b) enthalpies of formation and entropies (Appendix G). Do the results indicate the reaction to be spontaneous or nonspontaneous at 25 °C?



Answer

-141.5 kJ/mol, nonspontaneous



Calculating Gibbs Free Energy (Grxn) for a Reaction: <https://youtu.be/wmreE6zeFQo>

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