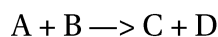


Chapter 6

Quantitative Aside 6.1—Biological Reactions and Free Energy

The equation $\Delta G = \Delta H - T\Delta S$ can be thought of as a quantitative expression of the second law of thermodynamics. It is useful because it can be used to predict the direction of a process: a process with a negative ΔG is spontaneous. For chemical reactions, this depends on conditions (temperature, pressure, pH) and relative concentrations of reactants and products. If we hold temperature and pressure constant, we call this the standard free energy, ΔG° . The change in free energy at nonstandard conditions depends on the standard free energy change, the conditions, and the relative concentrations of reactants and products. From all of this we can develop some useful equations.

Consider the reaction:



the equilibrium constant for this is:

$$K_{eq} = [C][D] / [A][B]$$

The ΔG for nonstandard conditions is:

$$\Delta G = \Delta G^\circ + RT \ln([C][D] / [A][B])$$

At equilibrium, $\Delta G = 0$

using K_{eq} above and rearranging we get:

$$\Delta G^\circ = -RT \ln K_{eq}$$

or in base 10 logs

$$\Delta G^\circ = -2.3 RT \log K_{eq}$$

This is useful in several ways. First it means that we can calculate the standard free energy for a reaction where we have an experimentally determined equilibrium constant. It also makes clear why a reaction with a negative ΔG would be spontaneous: if K_{eq} is >1 , then the ΔG will be negative. Lastly, it allows us to predict free energy changes under nonstandard conditions.