Making sense of ΔG and $\Delta G°$, when it comes to equilibrium

by Adrian Dingle, | Jan 31, 2012

Consider the two equations that deal with Delta G (ΔG).

Equation 1:

$$\Delta G° = – RT \ln K$$

Since K is the equilibrium constant, we *are* at equilibrium, the amounts of products and reactants in the mixture are fixed, and the sign of $\Delta G°$ can be thought of as a guide to the ratio of the amount of products to the amount of reactants at equilibrium and therefore the *thermodynamic favorability* of the reaction.

If it so happens that products and reactants are equally favored at equilibrium, then $\Delta G°$ is zero, **BUT $\Delta G°$ is not *necessarily* ZERO at equilibrium.**

Equation 2:

$$\Delta G = \Delta G° + RT \ln Q$$

Since Q is NOT the K, and we are NOT necessarily at the equilibrium position, the sign of ΔG can be thought of as a predictor about which way the reaction (that has reactants and products defined by Q), will go.

If $\Delta G°$ is negative at equilibrium, then we will have lots of products at equilibrium, meaning Q needs to be bigger (greater than 1) to approach K. As Q gets larger (i.e., as we get more products), the term 'RT ln Q' gets increasingly positive, and eventually adding that term to a negative $\Delta G°$**,** will make $\Delta G = 0$, equilibrium will be established and no further change occurs.

It is possible that Q could already be too large and therefore ΔG is positive. IF so, then the reaction will need to from more reactants, reduce the value of Q, and allow ΔG to reach zero, i.e., allow equilibrium to be established.

If $\Delta G°$ is positive at equilibrium, then we will have lots of reactants at equilibrium, meaning Q needs to be smaller (less than 1) to approach K. As Q gets smaller (i.e., as we get more reactants), the term 'RT ln Q' gets increasingly negative, and eventually adding that term to a positive $\Delta G°$**,** will make $\Delta G = 0$, equilibrium will be established and no further change occurs.

It is possible that Q could already be too small and therefore ΔG is negative, IF so, then the reaction will need more products, increase the value of Q, and allow ΔG to reach zero, i.e., allow equilibrium to be established.

In short, it is ΔG (NOT $\Delta G°$) that will be zero at equilibrium and the sign of *it* (generated by the combination of $\Delta G°$ and RT ln Q in Equation #2), will define which way the reaction proceeds.