Thermochemistry/Thermodynamics Study Guide

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

Elements ARE zero for: $\underline{\Delta H^\circ, \Delta G^\circ}$

Elements are NOT zero for: $\underline{S^\circ}$

Nature favors:

1. $\underline{+\Delta S}$ (increasing entropy, greater disorder)
2. $\underline{-\Delta H}$ (exothermic, release of energy)

- **Driven by enthalpy:** a very exothermic reaction ($\underline{-\Delta H}$) overcomes a decrease in entropy

- **Driven by entropy:** an endothermic reaction occurs spontaneously because of a highly $\underline{positive}$ ΔS

- You can use the signs of ΔS and ΔH (together with the reaction $\Delta G = \Delta H - T\Delta S$) to qualitatively predict the spontaneity of a given reaction. You can show this graphically, which leads to the best graph of all time (but SHHHH! It's a secret. 😊)
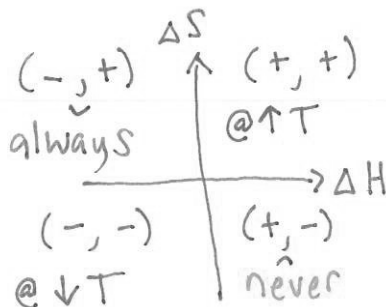
E^o_{cell}	ΔG^o	K	K & Q Relationship	Reaction Direction	Spontaneity in the Forward Direction (as written)
+	−	K > 1	K > Q	Forward	thermodynamically favorable
−	+	K < 1	K < Q	Reverse	NOT thermodynamically favorable
= 0	= 0	K = 1	K = Q	No reaction	n/a

At standard conditions (1 M, 1 atm, 298 K): $$\Delta G^o = -nFE^o_{cell}$$

n = number of moles of electrons transferred in a **BALANCED** redox reaction
F = faraday's constant = 96,485 C/ mol e⁻ (charge on one mole of electrons)

At standard conditions (1 M, 1 atm, 298 K): $$\Delta G^o = -RT \ln K$$

R = = 8.314 J mol⁻¹ K⁻¹   T = temperature (in Kelvin)   K = equilibrium constant

For both reactions, the units for $\Delta G^o = \dfrac{joules}{moles_{reaction}} = \dfrac{J}{mol_{rxn}}$

END-ing a bond (or IMF/attraction) is always END-othermic! (+ΔH)