Electrochem Equilibrium Summary

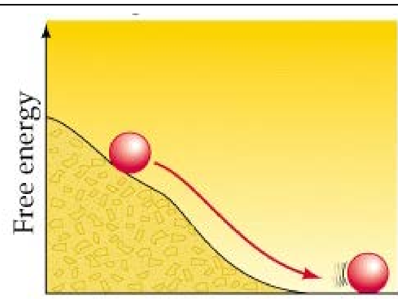
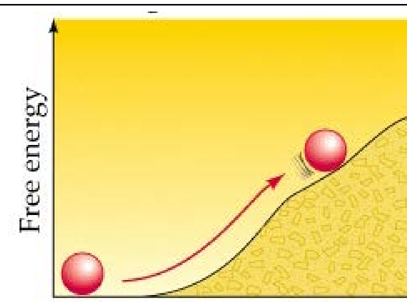
Given on formula chart:

$\Delta G^o = -nFE^o_{cell}$ n = e^- transferred, F = Faraday's constant

$\Delta G^o = -RT \ln K$ R = 8.314 J/(mol K), T = temp (K), ln(k) = natural log of K

Not given on formula chart:

$K = e^{-\Delta G^o/RT}$ e = 2.718, R = 8.314 J/(mol K), T = temp (K), ΔG must be in J/mol

Exergonic Reaction (−ΔG)	Endergonic Reaction (+ΔG)
Spontaneous (Thermodynamically Favorable)	Not Spontaneous (Thermodynamically Unfavorable)
$+E^o_{cell}$ = voltage created (battery)	$-E^o_{cell}$ = external power source needed
K > 1	K < 1
	

For Spontaneous Reactions: What if it's not at standard conditions?

Further from equilibrium	**At standard conditions** (1.0 M, 1.0 atm, 298K)	**Closer to equilibrium**	**At equilibrium**
Q < 1	Q = 1	Q > 1	Q >> 1
K >>> Q	K >> Q	K > Q	K = Q
[reactants] > [products]	[reactants] = [products]	[reactants] < [products]	[reactants] << [products]
Higher voltage (than standard E^o_{cell})	**Equal voltage** (to standard E^o_{cell})	**Lower voltage** (than standard E^o_{cell})	***No voltage!*** (dead battery)
$E_{cell} > E^o_{cell}$	$E_{cell} = E^o_{cell}$	$E_{cell} < E^o_{cell}$	E_{cell} = 0 V
How Does this Happen? • Increase reactants • Decrease products • Increase both, but increase REACTANTS more • Decrease both, but decrease PRODUCTS more	How Does this Happen? • Set [reactants] = 1.0 M • Set [products] = 1.0 M • Increase both, but increase BOTH the same • Decrease both, but decrease BOTH the same	How Does this Happen? • Increase products • Decrease reactants • Increase both, but increase PRODUCTS more • Decrease both, but decrease REACTANTS more	How Does this Happen? • Cell runs for a very long time • All reactants used up