AP CHEMISTRY

TOPIC 11: ELECTROCHEMISTRY, PART B, EXAMPLES, PART II Day 126:

- Voltaic Cells (Galvanic Cells)

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1. Answer the following questions using the data from the Standard Reduction Potentials at 25^{0}C.
   a) Is Ni^{+2} (*aq*) capable of oxidizing Sr (*s*) to Sr^{+2} (*aq*), explain why.

$Ni^{+2} + 2\,e^{-1} \rightarrow Ni$	(- 0.25 V)
$Sr \rightarrow Sr^{+2} + 2\,e^{-1}$	(+ 2.39 V)
Yes !!!	+ 2.14 V

Yes, E^0 (for the reaction) has a positive value. When E^0 is positive, ΔG is negative (spontaneous).

   b) Is I^{-1} (*aq*) capable of reducing Ag^{+1} (*aq*) to Ag (*s*) , explain why. (recall, these reactions are NOT occurring in the same "container" (or chamber) – this is simply an exchange of electrons from one container to another – THIS IS NOT a single replacement reaction !

$2\,I^{-1} \rightarrow I_2 + 2\,e^{-1}$	(- 0.53 V)
$Ag^{+1} + e^{-1} \rightarrow Ag$	(+ 0.80 V)
YES !!!	+ 0.27 V

Yes, E^0 (for the reaction) has a positive value. When E^0 is positive, ΔG is negative (spontaneous). Don't get "worked up" that we have a metal with a non-metal. This is not a single replacement reaction – 2 chambers!

2. In a galvanic cell the concentration of Sn^{+2} is changed from 1.0 *M* to a 0.33 *M*, and the concentration of Al^{+3} is changed from 1.0 *M* to a 0.47 *M*. Temperature is at 25^{0}C.

a) First predict the oxidizing agent and the reducing agent for the spontaneous reaction in the cell. After that determine the anode and cathode for the galvanic cell.

$2\,(Al \rightarrow Al^{+3} + 3\,e^{-1})$	(+ 1.66 V)	Oxidized, Reducing Agent
$3\,(Sn^{+2} + 2\,e^{-1} \rightarrow Sn)$	(- 0.14 V)	Reduced, Oxidizing Agent
	+ 1.52 V	

The above half-reactions give the "***Best***" Voltage

b) Calculate the cell potential for the galvanic cell at ***standard conditions*** (use the Nernst Equation) Know STANDARD CONDITIONS: 25^{0}C and the solutions begins at 1.0 molar !!!

$$@\,25\ ^0C:\quad E_{cell} = E^0_{cell} - \frac{0.0592\ V}{n}\log Q \quad ; \quad Q = \frac{\left[Al^{+3}\right]^2}{\left[Sn^{+2}\right]^3}$$

$$E_{cell} = \left(+1.52\ V\right) - \left(\frac{0.0592\ V}{(6)}\log\left(\frac{(1)^2}{(1)^3}\right)\right) = +1.52\ V \quad ; \quad (\text{Recall, } \log 1 = 0)$$

$$E_{cell} = +1.80\ V - (0) = +1.80\ V$$

b) Calculate the cell potential for the galvanic cell under the new concentrations (changed concentrations).

$$E_{cell} = \left(+1.52\ V\right) - \left(\frac{0.0592\ V}{(6)}\log\left(\frac{(0.47)^2}{(0.33)^3}\right)\right)$$

$$E_{cell} = \left(+1.52\ V\right) - 0.07781\ V = +1.51\ V$$