

AP CHEMISTRY

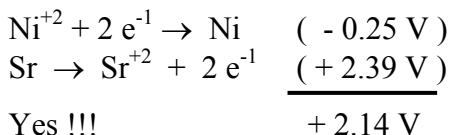
TOPIC 11: ELECTROCHEMISTRY, PART B, EXAMPLES, PART II

Day 126:

- Voltaic Cells (Galvanic Cells)

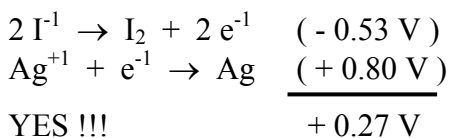
1. Answer the following questions using the data from the Standard Reduction Potentials at 25°C.

a) Is $\text{Ni}^{+2} (aq)$ capable of oxidizing $\text{Sr} (s)$ to $\text{Sr}^{+2} (aq)$, explain why.



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

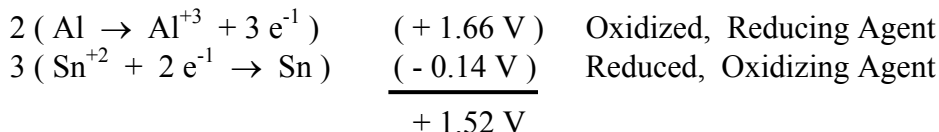
b) Is $\text{I}^{-1} (aq)$ capable of reducing $\text{Ag}^{+1} (aq)$ to $\text{Ag} (s)$, explain why. (recall, these reactions are NOT occurring in the same “container” (or chamber) – this is simply an exchange of electrons from one container to another – THIS IS NOT a single replacement reaction !



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°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.



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°C and the solutions begins at 1.0 molar !!!

$$@ 25^{\circ}\text{C} : E_{\text{cell}} = E_{\text{cell}}^0 - \frac{0.0592 \text{ V}}{n} \log Q ; Q = \frac{[\text{Al}^{+3}]^2}{[\text{Sn}^{+2}]^3}$$

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

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

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

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

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