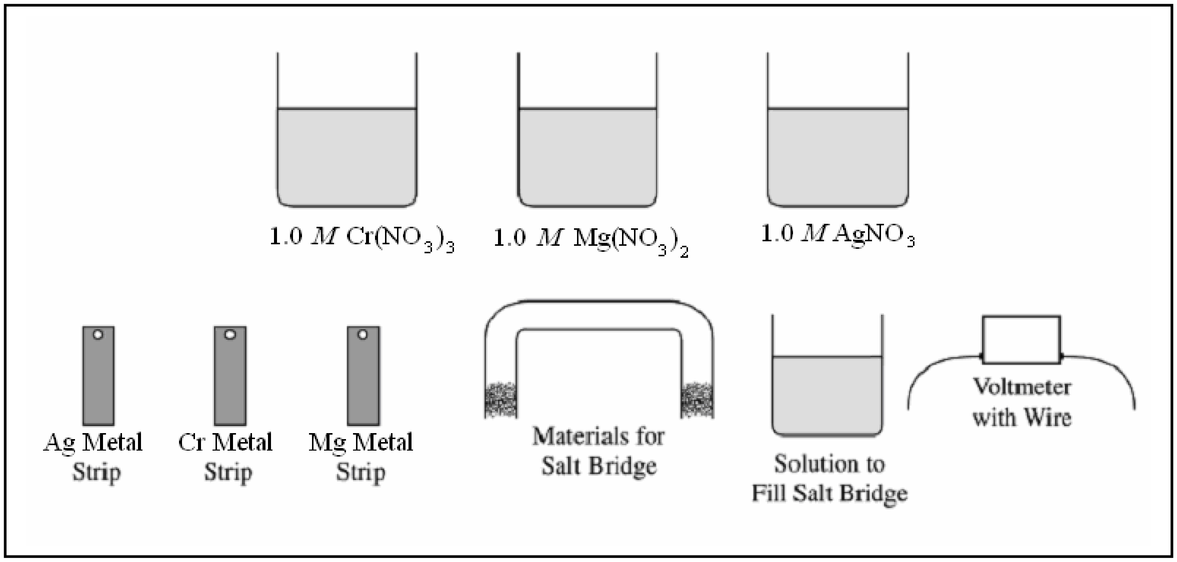
AP CHEMISTRY

TOPIC 11: ELECTROCHEMISTRY, EXAMPLES

Day 127:

CLEARLY SHOW THE METHOD USED AND THE STEPS INVOLVED IN ARRIVING AT YOUR ANSWERS

1. a) Several different electrochemical cells can be constructed using the materials shown below. Write the balanced net-ionic equation for the reaction that occurs in the cell that would have the greatest positive value of E_{cell}.



$Mg^{+2} + 2e^{-1} \rightarrow Mg$ $E^0 = -2.37$ V (reverse reaction to oxidize)

$Ag^{+1} + e^{-1} \rightarrow Ag$ $E^0 = +0.80$ V (best reduction reaction)

$\cancel{Cr^{+3} + 3e^{-1} \rightarrow Cr \quad E^0 = -0.74\text{ V}}$ (can only have one thing oxidize)

$$Mg \rightarrow Mg^{+2} + \cancel{2e^{-1}}$$

$$2\,(Ag^{+1} + \cancel{e^{-1}} \rightarrow Ag\,)$$

$$Mg + 2\,Ag^{+1} \rightarrow Mg^{+2} + 2\,Ag$$

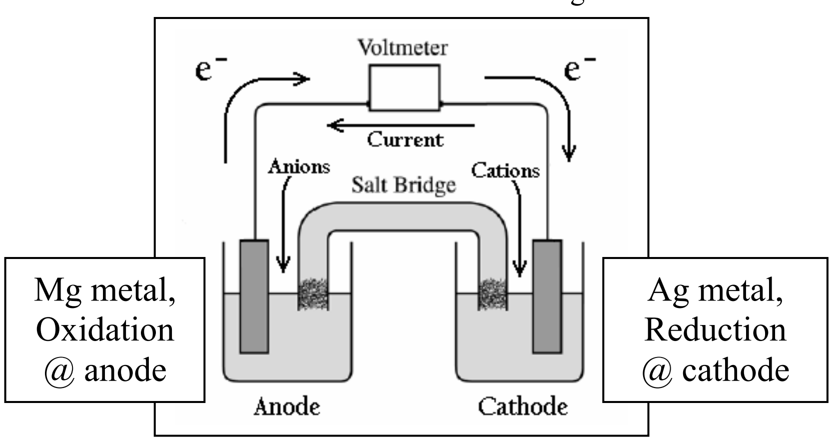
b) Calculate the standard cell potential, E^0_{cell}, for the reaction written in part (a).

$$Mg \rightarrow Mg^{+2} + 2e^{-1} \qquad E^0 = +2.37\text{ V}$$

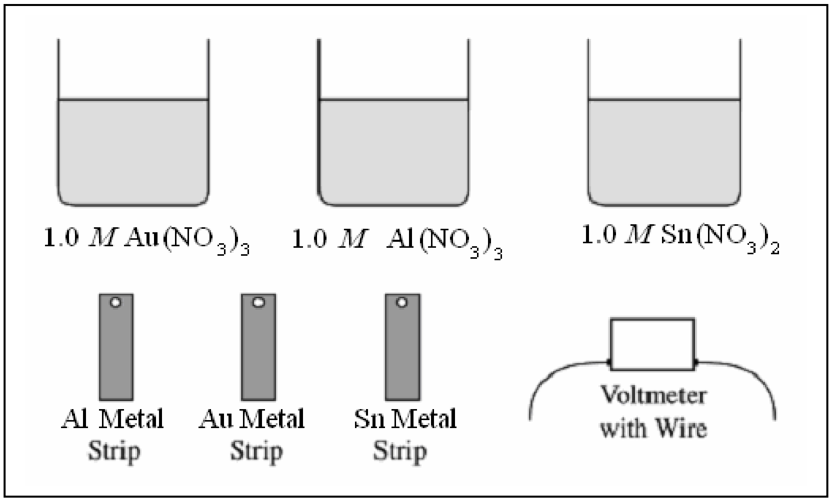
$$2\,(Ag^{+1} + e^{-1} \rightarrow Ag\,) \qquad E^0 = +0.80\text{ V}$$

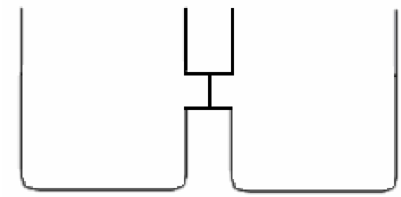
$$E^0 = +3.17\text{ V}$$

c) A cell is constructed based on the reaction in part (a) above. Draw and label the all the parts of the galvanic cell when put together correctly (place the anode on the left side). Also, show the flow of electrons, anions, cations, and current. Also, indicate where oxidation and reduction is occurring.



2. a) Several different electrochemical cells can be constructed using the materials shown below (notice there is no salt bridge used here, only the porous disk between the two chambers). Write the balanced net-ionic equation for the reaction that occurs in the cell that would have the greatest positive value of E_{cell}.




Two chambers with a porous disk – no salt bridge is to be used

$$Au^{+3} + 3\,e^{-1} \rightarrow Au \qquad E^0 = +1.50\text{ V} \quad \text{(best reduction reaction)}$$

$$Al^{+3} + 3\,e^{-1} \rightarrow Al \qquad E^0 = -1.66\text{ V} \quad \text{(reverse reaction to oxidize)}$$

$$\cancel{Sn^{+2} + 2\,e^{-1} \rightarrow Sn \qquad E^0 = -0.14\text{ V}} \quad \text{(can only have one thing oxidize)}$$

$$Al \rightarrow Al^{+3} + \cancel{3\,e^{-1}}$$

$$Au^{+3} + \cancel{3\,e^{-1}} \rightarrow Au$$

$$Al + Au^{+3} \rightarrow Al^{+3} + Au$$

b) Calculate the standard cell potential, E^0_{cell}, for the reaction written in part (a) at 25^{0}C.

$$Al \rightarrow Al^{+3} + 3\,e^{-1} \qquad E^0 = +1.66\text{ V}$$

$$Au^{+3} + 3\,e^{-1} \rightarrow Au \qquad E^0 = +1.50\text{ V}$$

$$E^0 = +3.16\text{ V}$$

c) Calculate the standard cell potential, E^0_{cell}, for the reaction when the concentrations are changed to (at 25^{0}C.): 0.75 M $Au(NO_3)_3$, 1.33 M $Al(NO_3)_3$, 0.55 M $Sn(NO_3)_2$. (Hint: we are still using the same cell potential from above).

$$E_{cell} = E^0_{cell} - \frac{0.0592\ V}{n}\log Q \quad ; \quad Q = \frac{\left[Al^{+3}\right]^1}{\left[Au^{+3}\right]^1}$$

$$E_{cell} = (\ +3.16\ V\ ) - \left(\frac{0.0592\ V}{(3)}\log\frac{(1.33)^1}{(0.75)^1}\right)$$

$$E_{cell} = (\ +3.16\ V\ ) - 0.00491\ V = +3.155\ V = +3.16\ V\,(sig\ \ figs)$$