Dougherty Valley HS AP Chemistry
Electrochemistry – Electrochemical cell

WORKSHEET #1 KEY

Name: **Date:** **Period:** **Seat #:**

Show all work

Consider the reduction potential chart. Find and copy the reduction equations for $Ag^{+} \rightarrow Ag°$ and $Pb^{2+} \rightarrow Pb°$

Silver reduction equation: $Ag^{+} + e^{-} \rightarrow Ag$	Potential Value: E° = + 0.80 V
Lead reduction equation: $Pb^{2+} + 2e^{-} \rightarrow Pb$	Potential Value: E° = - 0.13 V

1. Which metal ion has the greater reduction potential? **Ag**	2. If these two metals (and their solutions) were used to create a galvanic cell, which metal would be the anode? **Pb**
3. Write the reaction at the anode: $\mathbf{Pb \rightarrow Pb^{2+} + 2e^{-}}$	4. Write the reaction at the cathode: $\mathbf{Ag^{+} + e^{-} \rightarrow Ag}$
5. What is the overall reaction? $\mathbf{2Ag^{+} + Pb \rightarrow Pb^{2+} + Ag}$	6. What would be the voltage of the standard electrochemical cell? **E°** = + 0.80 V + 0.13 V = **0.93 V**
7. How many moles of electrons are involved in this reaction? n = **n = 2**	8. Find and copy down the Nernst Equation: $E_{cell} = E^{\circ}_{cell} - \left(\frac{0.0592}{n}\right) Ln\ Q$

9. If the standard cell is allowed to run until the $[Ag^{+}]$ = 0.50 M, the $[Pb^{2+}]$ = 2.0 M, the cell voltage will be **<u>LESS</u>** (greater / less)?

10. Use the Nernst equation to calculate the cell voltage with these new concentrations

$$E_{cell} = 0.93V - \left(\frac{0.0592}{2}\right) Ln\ \left(\frac{2.0\ M}{0.75M}\right) = \mathbf{0.91\ V}$$