AP Chemistry KEY

18 • Electron Transfer Reactions

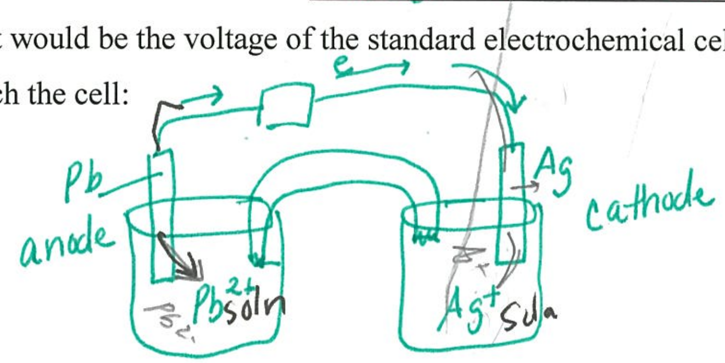
A VISIT TO ELECTROCHEMISTRYLAND

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

$Ag^+ + e^- \rightarrow Ag$ $\quad \mathcal{E}° = +0.80\ V$

$Pb^{2+} + 2e^- \rightarrow Pb$ $\quad \mathcal{E}° = -.13V$

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 smaller $\mathcal{E}°$ value
3. Write the reaction at the anode: $Pb \rightarrow Pb^{2+} + 2e^-$
4. Write the reaction at the cathode: $Ag^+ + e^- \rightarrow Ag$
5. What is the overall reaction? $2Ag^+ + Pb \rightarrow Pb^{2+} + 2Ag°$
6. What would be the voltage of the standard electrochemical cell? $.80 + .13 = .93V$
7. Sketch the cell:



8. Write the cell notation for the cell: $Pb \mid Pb^{2+} \parallel Ag^+ \mid Ag$ anode cathode
9. How many moles of electrons are involved in this reaction? n = 2

10. Find and copy down the Nernst Equation: $\mathcal{E} = \mathcal{E}° - \frac{RT}{nF} \log Q$
11. If the standard cell is allowed to run until the $[Ag^+] = 0.75$ M, the $[Pb^{2+}] =$ 1.125M
12. At these new concentrations, the cell voltage will be less (greater / less).
13. Use the Nernst equation to calculate the cell voltage with these new concentrations.

$$E_{cell} = .93\ V - \frac{(8.31\ V{\cdot}C)(298K)}{(2\ mole^-)(96500\ C\ mol^{-1})} \log \frac{1.125}{.75} V = .927$$

$2Ag^+ + Pb \rightarrow Pb^{2+} + 2Ag°$

1.00	1.00
- .25	+ .125
.75	1.125

Variables that appear in these equations are R and $\mathcal{F}$.
Values and units? R = 8.31 J/mol K $\mathcal{F}$ = 96500 coulombs/mole e^-

E° and K	ΔG° and E°	ΔG° and K
$\log K = nE°/ 0.0592$ @ 25°C $\ln K = \frac{nF E°}{RT}$	$\Delta G° = -nF E°$ Note: volt x coul = J ΔG° in kJ	$\Delta G° = -RT\ln K$

Problems:

1. An early method of producing aluminum metal was the reaction of aluminum salts with sodium metal:

$$Al^{3+} + 3Na(s) \rightleftharpoons Al(s) + 3Na^+ \qquad E° = +1.05\ V$$

What is ΔG° for this reaction

$\boxed{\text{a) -304 kJ}}$ d) +202 kJ
b) -101 kJ e) +304 kJ
c) +101 kJ

$\Delta G = -nF\mathcal{E}°$

$\Delta G = -(3mol)\left(\frac{96500\ C}{mole\ e^-}\right)(1.05V)\left(\frac{1J}{C\cdot V}\right)\left(\frac{1kJ}{1000J}\right) = \boxed{-304\ kJ}$

2. Calculate K_{eq} for this reaction.

$\log K = \frac{nE}{0.0592} = \frac{(3mol)(1.05V)}{.0592} = 53.2$

$K = 10^{53.2} = \boxed{1.6\times10^{53}}$

3. Calculate E° for the following reaction:

$Br_2 + 2e^- \longrightarrow 2Br^-$ +1.07

$I_2(s) + 2Br^-(aq) \rightarrow 2I^-(aq) + Br_2(l)$

$I_2 + 2e^- \longrightarrow 2I^-$ +.53

$2Br^- \longrightarrow Br_2 + 2e^-$ −1.07

$-1.07 + .53 = \boxed{-.54V}$

4. Calculate ΔG for the above reaction

$\boxed{\text{a) +105 kJ}}$ d) +52 kJ
b) -105 kJe) e) -312 kJ
c) +312 kJ

$\Delta G = -nF\mathcal{E}$

$-(2mol)\left(\frac{96500C}{mole}\right)(-.54V)\left(\frac{1J}{C\cdot V}\right)\left(\frac{1kJ}{1000J}\right)$

+105 kJ

5. If ΔG of the following reaction is -203 kJ, what is E°? $2Ag^+(aq) + Ni(s) \rightarrow 2Ag(s) + Ni^{2+}(aq)$

a) -1.05 V d) -0.011 V
b) +2.10 V $\boxed{\text{e) +1.05 V}}$
c) +0.0011 V

$\Delta G = -nF\mathcal{E}°$

$-203kJ\left(\frac{1000J}{1kJ}\right) = -2mol\left(\frac{96500\ C}{mole^-}\right)(\mathcal{E}°)$

$\mathcal{E} = +1.05V$

6. What is the equilibrium constant for the following reaction at 20°C?

$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$ E° = +0.78 V

a) 2.3 x 10^{26} d) 1.8 x 10^{28}
$\boxed{\text{b) 6.9 x } 10^{26}}$ e) 1.2 x 10^{-21}
c) 1.4 x 10^{27}

$\Delta G = -RT\ln K = -nF\mathcal{E}$

$\ln K = \frac{+nF\mathcal{E}}{+RT}$ @ any temp

$\ln K = \frac{(2mol)\left(\frac{96500C}{mole}\right)(+.78V)}{8.31\frac{V\cdot C}{mol\ K}(293K)} = 6.9\times10^{26}$

$\ln K = 61.8$

≈ 61.8

$C/s = AMP$

$96500\ C/\text{mole}\ e^-$

Stoichiometry

1. Calculate the quantity of electricity (Coulombs) necessary to deposit 100.00g of copper from a $CuSO_4$ solution.

$Cu^0 \quad Cu^{2+}$

$2e^- + Cu^{2+} \rightarrow Cu$

$$\frac{100.0\text{g Cu}}{} \Big| \frac{1\text{mol Cu}}{63.546\text{g}} \Big| \frac{2\text{ mole}^-}{1\text{mol Cu}} \Big| \frac{96500\text{ C}}{1\text{mole}^-} = \boxed{3.04 \times 10^5\text{ C}}$$

2. How many minutes will it take to plate out 40.00 g of Ni from a solution of NiSO4 using a current of 3.450amp?

$2e^- + Ni^{2+} \rightarrow Ni$

$$\frac{40.00\text{g Ni}}{} \Big| \frac{1\text{mol Ni}}{58.79\text{g Ni}} \Big| \frac{2e^-}{1\text{mol Ni}} \Big| \frac{96500\text{C}}{1\text{mole}} \Big| \frac{1\text{sec}}{3.450\text{C}} \Big| \frac{1\text{min}}{60\text{sec}} = \boxed{635.3\text{min.}}$$

$3.450\text{ Amp} = 3.450\ C/s$

3. What is the equivalent weight of a metal if a current of 0.2500 amp causes 0.5240g of metal to plate out a solution undergoing electrolysis in 1 hour? ** One mole of electrons will plate out one equivalent weight of metal

$$.2500\text{ AMP} = \frac{.2500\text{ C}}{\text{sec}} \Big| \frac{3600\text{sec}}{} \Big| \frac{1\text{mole}^-}{96500\text{C}} = 9.33 \times 10^{-3}\text{ mol}$$

$$\frac{.5240\text{g}}{9.33 \times 10^{-3}\text{mol}} = \boxed{56.2\text{ g/mol}}$$

4. How many hours will it take to plate out copper in 200.00ml of 0.15M Cu 2+ solution using a current of 0.200 amp?

$$.15M = \frac{x}{.200L} = 0.0300\text{ mol Cu} \Big| \frac{2\text{mole}^-}{1\text{ Cu}} \Big| \frac{96500\text{C}}{1\text{mole}^-} \Big| \frac{1\text{sec}}{.200\text{C}} \Big| \frac{1\text{hr}}{3600\text{sec}} = \boxed{8.04\text{ hrs}}$$

5. A constant electric current deposits 0.3650g of silver metal in 12960 seconds from a solution of silver nitrate. What is the current? What is the half reaction for the deposition of silver?

$1e^- + Ag^+ \rightarrow Ag$

$$\frac{.3650\text{g Ag}}{} \Big| \frac{1\text{mol Ag}}{107.87\text{g Ag}} \Big| \frac{1\text{mole}^-}{1\text{ mol Ag}} \Big| \frac{96500\text{C}}{1\text{mole }e^-} = \frac{326.48\text{C}}{12960\text{ sec}} = \boxed{0.252\text{ AMP}}$$