Spring 2006   CH302, Practice Exam 6.   Solution key answers in red.
(the correct answer is included below. A solution key will be posted on Monday.)

1. Using the half reactions in the table of standard reduction potentials in Worksheet 11b, what is the standard emf (cell potential in V) for $Pt|\ Li|\ Li^+(1M)||\ F_2$ (1 atm)$|F^-(1M)|Pt$

1. +5.91 V **correct**
2. -5.91 V
3. -0.17   V
4. 0.17   V

Solution: All compounds are standard state so simple use $E^o_{cell} = E^o_{cathode} - E^o_{anode}$ = 2.87-(-3.04) = 5.91 V. The nearly +6V spontaneous reaction is the largest cell reaction you can make with the half reactions found in the usual table of standard reduction potentials.

2. Consider the half-reactions in the table of standard reduction potentials in Worksheet 11b
Of the species listed, the strongest reducing agent is ? and the weakest oxidizing agent is ?

1. Li, Li^+ **correct**
2. Li, F_2
3. Li^+, F_2
4. F^-, F_2
Solution: From the table of standard reduction potentials, we want the strongest reducing agent which means we want the oxidation process that is the most positive--that is the Li oxidation in the upper right hand corner of the table (+3.04V). Next we want to find the weakest oxidizing agent which means we want the reduction half reaction compound with the most negative E^o, Li^+, in the upper left hand corner (-3.04V).
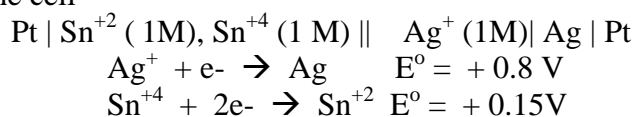
3. Consider the cell
$$Pt\ |\ Sn^{+2}\ (\ 1M),\ Sn^{+4}\ (1\ M)\ ||\quad Ag^+\ (1M)|\ Ag\ |\ Pt$$
$$Ag^+\ +\ e\text{-}\ \rightarrow\ Ag\qquad E^o =\ +0.8\ V$$
$$Sn^{+4}\ +\ 2e\text{-}\ \rightarrow\ Sn^{+2}\ \ E^o =\ +0.15V$$
at standard conditions. Calculate the value of ΔG for the reaction that occurs when current is drawn from this cell.
1. 125 kJ/mol
2. -125 kJ/mol   **correct**
3. 62 kJ/mol
4. -62 kJ/mol

Solution: This is a multi-step problem.

First we need to find E^o_{cell}. All the compounds are standard state so we simply use
$E^o_{cell} = E^o_{cathode} - E^o_{anode}$ = 0.8V -0.15V = 0.65 V where the silver electrode is cathode and tin electrode is anode.

Also note in balancing the reaction that the least common multiple is 2 which means this is an n=2 lectron process.

Finally, we find ΔG = -nFE = -(2)(9.65x10^4)(0.65V) = 125,000J $\rightarrow$ -125 kJ.

Note the E^o_{cell} is positive so free energy is negative, a spontaneous reaction.

4. An electrochemical process with a positive ΔG value has two terminals labeled positive and negative. In this cell, electrons would flow from the ? terminal to the ? terminal through the external circuit and ? reaction occurs at the positive terminal.
1. positive; negative; a reduction
2. positive; negative; an oxidation **correct**
3. negative; positive; a reduction
4. negative; positive; an oxidation
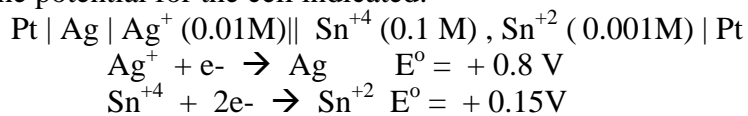5. positive; negative; an acid/base

Solution: Because the reaction is non-spontaneous (an electrolysis cell), by definition the anode is positive and the cathode is negative. Also by definition, electrons flow from the anode to cathode so the flow is from positive electrode to negative electrode. Finally, the positive terminal is the anode in electrolysis, so oxidation is occurring there.

5. Which of the following batteries is rechargeable?
1. dry cell
2. copper/zinc
3. nickel-cadium battery **correct**
4. copper/carrot top

Solution. The famous rechargeable batteries mentioned in class are the lead acid storage battery found in your car and the nickel cadmium battery which is the kind of rechargeable battery you find in stores in standards sizes like AA and D.

6. Calculate the potential for the cell indicated:
$$Pt \mid Ag \mid Ag^+ (0.01M) \parallel Sn^{+4} (0.1\ M) , Sn^{+2} (0.001M) \mid Pt$$
$$Ag^+ + e\text{-} \rightarrow Ag \qquad E^o = +0.8\ V$$
$$Sn^{+4} + 2e\text{-} \rightarrow Sn^{+2}\ E^o = +0.15V$$
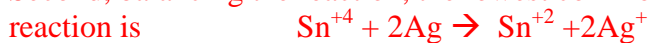
1. -0.65V
2. +0.65 V
3. -0.47 V **correct**
4. – 0.53 V
5. -0.29 V

Solution. The potentials in this reaction are nonstandard so we use the Nernst eqn

$E_{cell} = E^o{}_{cell} -.059/n \log Q$

Step by step
First: $E^o{}_{cell} = E^o{}_{cathode} - E^o{}_{anode} = 0.15 – 0.8\ V = \text{-0.65V}$
Second, balancing the reaction, the lowest common multiple is 2, so this is a 2 electron process. The balanced reaction is $Sn^{+4} + 2Ag \rightarrow Sn^{+2} +2Ag^+$
so $Q= (Sn^{+2})(Ag^+)^2/(Sn^{+4})$ (we don't include solid silver because it has unit activity.

Substituting the given concentrations, and remembering to square the Ag^+ we wind up with $Q = 10^{-6}$.

So $E_{cell} = E^o{}_{cell} -.059/n \log Q = -.65 - .059/2 \log\ 10^{-6} = \text{-0.47 V}$

7. What weight of Au is deposited on a piece of cheap zinc chain during the electrolysis of a Au^{+3} solution when a current of 1.3 mA amps flows for 1.2 hours? (Au = 197 g/mol)

1. 3.82 mg **correct**
2. 1.27 mg
3. 11.46 mg
4. 1.44 g
5. 43.2 g

Solution: This is a classic stoichiometry problem where we use an initial amount of charge instead of an initial mass. Otherwise the problem is the same.

To find the charge, in Coulombs, remember that current, i = Coulombs/second.
So Coulombs = i (in Amps) * seconds = 1.3×10^{-3}A * 1.2 hours (3600 seconds/hour) = 5.616 Coulombs.

Also note this is a three electron process because one mole of gold forms when one mole of Au^{+3} is reduced by three moles of electrons.

So: g of Au = 5.616 Coulombs (1 mole e⁻s/9.65×10^4 Coulombs)(1mole Au/3 moles e⁻s)(197 g Au/1 mole Au) = 3.82×10^{-3} grams.

8 Using the smallest possible integer coefficients to balance the redox equation
Fe + $Cr_2O_7^=$ → Fe^{+2} + Cr^{+3} in acidic solution, the coefficient for $Cr_2O_7^=$ is
1. 1 **correct**
2. 2
3. 3
4. 4
5. The correct coefficient is not given.

Solution.

This is a six electrons process, when the least common factor 6, is multiplied through the reaction.

$3Fe$ + $1Cr_2O_7^=$ → $3Fe^{+2}$ + $2Cr^{+3}$
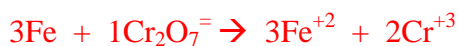
Note that to answer this problem you don't need to finish balancing the reaction—you only ned to know the coefficient of the dichromate.