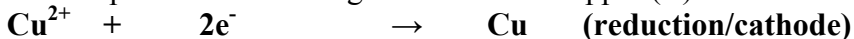


CH 302 Spring 2006
Worksheet 12b: Electrochemistry

1. (a) Calculate the mass of copper metal produced at the cathode during the passage of 2.50 amps of current through a solution of copper (II) sulfate for 50.0 minutes.



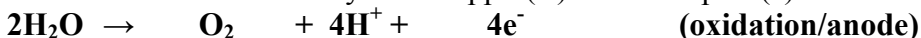
$$1 \text{ mol} \quad 2(6.02 \times 10^{23})\text{e}^- \quad 1 \text{ mol}$$

$$63.5 \text{ g} \quad 2(9.65 \times 10^4 \text{C}) \quad 63.5 \text{ g}$$

$$50.0 \text{ min} \times (60\text{s}/1 \text{ min}) \times (2.50\text{C}/\text{s}) = 7.50 \times 10^3 \text{C}$$

$$7.50 \times 10^3 \text{C} \times (1 \text{ mol e}^-/9.65 \times 10^4 \text{C}) \times (63.5 \text{ g Cu}/2 \text{ mol e}^-) = 2.47 \text{ g Cu}$$

(b) What volume of oxygen gas (measured at STP) is produced by the oxidation of water at the anode in the electrolysis of copper(II) sulfate in part (a)?

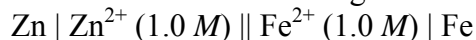


$$1 \text{ mol} \quad 4((6.02 \times 10^{23})\text{e}^-)$$

$$22.4 \text{ L} \quad 4(9.65 \times 10^4 \text{C})$$

$$7.50 \times 10^3 \text{C} \times (1 \text{ mol e}^-/9.65 \times 10^4 \text{C}) \times (22.4 \text{ L O}_2/4 \text{ mol e}^-) = 0.435 \text{ L O}_2$$

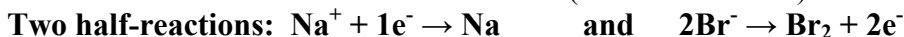
2. What is the E° for the following electrochemical cell where Zn is the cathode?



$$E^\circ (\text{Zn}) = -0.76 \quad E^\circ (\text{Fe}) = -0.44$$

$$E_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = -0.76 - (-0.44) = -0.32$$

3. For the electrolysis of molten sodium bromide, write the two half-reactions and show which electrode at which each occurs (cathode or anode).



Oxidation occurs at the anode, so the $2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^-$ reaction is at the anode, and

Reduction occurs at the cathode, so $\text{Na}^+ + 1\text{e}^- \rightarrow \text{Na}$ is at the cathode.

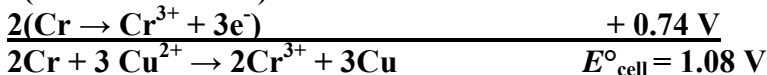
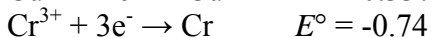
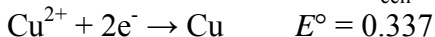
4. Calculate the potential, E , for the $\text{Fe}^{3+}/\text{Fe}^{2+}$ electrode when the concentration of Fe^{2+} is exactly five times that of Fe^{3+} .



$$Q = [\text{Red}]^y/[\text{Ox}]^x = [\text{Fe}^{2+}]/[\text{Fe}^{3+}] = 5[\text{Fe}^{3+}]/[\text{Fe}^{3+}] = 5$$

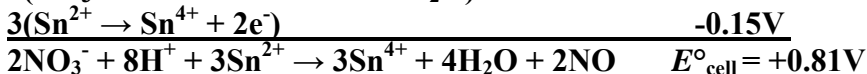
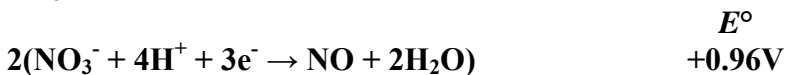
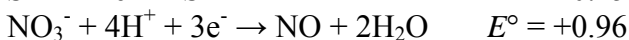
$$E = E^\circ - 0.0592/n \cdot \log Q = +0.771 - 0.0592/1 \cdot \log 5 = (0.771 - 0.041) \text{V} = 0.730 \text{ V}$$

5. At standard conditions, will chromium (III) ions, Cr^{3+} , oxidize metallic copper to copper (II) ions, Cu^{2+} , or will Cu^{2+} oxidize metallic chromium to Cr^{3+} ions? Write the cell reaction and calculate E°_{cell} for the spontaneous reaction.



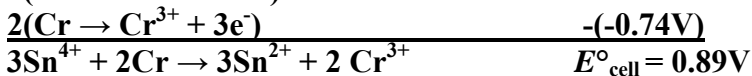
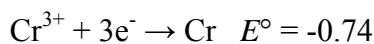
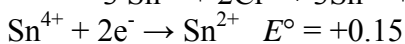
Cu^{2+} ions spontaneously oxidize metallic Cr to Cr^{3+} ions and are reduced to metallic Cu.

6. In an acidic solution at standard conditions, will tin(IV) ions, Sn^{4+} , oxidize gaseous nitrogen oxide, NO , to nitrate ions, NO_3^- , or will NO_3^- oxidize Sn^{2+} to Sn^{4+} ions? Write the cell reaction and calculate E°_{cell} for the spontaneous reaction.



Nitrate ions spontaneously oxidize tin(II) ions to tin(IV) ions and are reduced to nitrogen oxide in acidic solution.

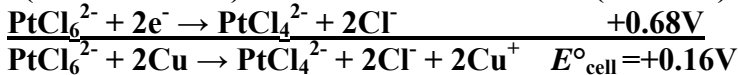
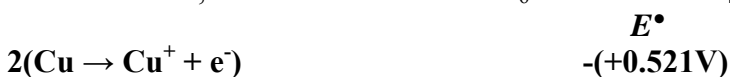
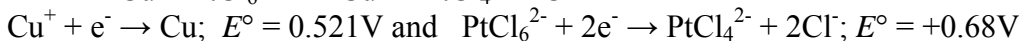
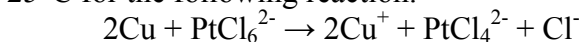
7. Calculate the Gibbs free energy change, ΔG° , in J/mol at 25°C for the following reaction:



$$\Delta G^\circ = -nF E^\circ_{\text{cell}} = -(6 \text{ mol e}^-/\text{mol rxn})(9.65 \times 10^4 \text{ J/V}\cdot\text{mol e}^-)(+0.89\text{V})$$

$$= -5.2 \times 10^5 \text{ J/mol rxn}$$

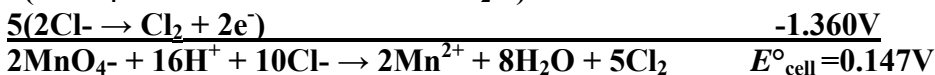
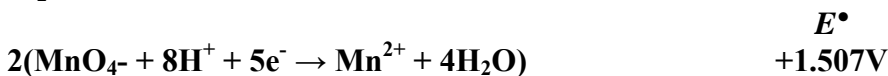
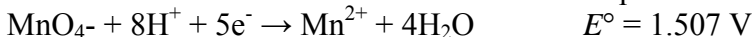
8. Use the standard cell potential to calculate the value of the equilibrium constant, K , at 25°C for the following reaction.



$$\ln K = -nF E^\circ_{\text{cell}}/RT = (2)(9.65 \times 10^4 \text{ J/V}\cdot\text{mol e}^-)(0.16\text{V})/(8.314\text{J/mol}\cdot\text{K})(298\text{K}) = 12.5$$

$$K = e^{12.5} = 2.7 \times 10^5$$

9. The following cell is maintained at 25°C . One half-cell consists of a chlorine/chloride, Cl_2/Cl^- , electrode with the partial pressure of $\text{Cl}_2 = 0.100 \text{ atm}$ and $[\text{Cl}^-] = 0.100 \text{ M}$. The other half-cell involves the $\text{MnO}_4^-/\text{Mn}^{2+}$ couple in acidic solution with $[\text{MnO}_4^-] = 0.100 \text{ M}$, $[\text{Mn}^{2+}] = 0.100 \text{ M}$, and $[\text{H}^+] = 0.100 \text{ M}$. Apply the Nernst equation to the overall cell reaction to determine the cell potential for this cell.



$$E_{\text{cell}} = E^\circ_{\text{cell}} - 0.0592/n \cdot \log \{ [\text{Mn}^{2+}]^2 (\text{P}_{\text{Cl}_2})^5 / [\text{MnO}_4^-]^2 [\text{H}^+]^{16} [\text{Cl}^-]^{10} \}$$

$$= 0.147\text{V} - (0.0592/10) \cdot \log [(0.100)^2 (0.100)^5 / (0.100)^2 (0.100)^{16} (0.100)^{10}]$$

$$= 0.147 \text{ V} - (0.0592/10) \cdot \log (1.00 \times 10^{21}) = 0.017\text{V}$$