Name: ______________________________ Date __________

Lab Partner ____________________________

Galvanic and Electrolytic Cells

Overview:
Part A: Determining the Voltage of Electrochemical Cells
Part B: Examination of Electrolytic cell using carbon electrodes in a KI solution
Part C: Electroplating Pennies and calculating Faraday's Law

Part A: Determining the Voltage of Galvanic Cells

Galvanic (or Voltaic) Cells are electrochemical cells in which a spontaneous reaction generates an electric current. Galvanic cells have two half-cells connected, so that electrons flow from one metal electrode to another through an external circuit and the ions flow through an internal cell connection (or salt bridge). The half-cell in which a half-reaction occurs with a loss of electrons (oxidation) is the **anode**. The second half-cell, in which a half-reaction occurs with a gain of electrons (reduction), is the **cathode**. Oxidation always occurs at the **anode** and reduction always occurs at the **cathode**.

The objective of the first part of the lab is to determine the cell potential ($\mathbf{E_{cell}}$) for various galvanic cells and compare the data with the calculated $\mathbf{E_{cell}}$ values obtained by using the Nernst equation. See Chapter 2, sections 2-5 in the Ebbing/Gammon text.

$$E_{cell} = E_{cathode} - E_{anode}$$

$$E_{cell} = E^{\circ}_{cell} - (0.0592/n) \log Q$$

n = # of electrons
Q = reaction quotient

Procedure:

1. Place about 2 mL of solutions 0.1 M $Cu(NO_3)_2$, 0.1 M $Zn(NO_3)_2$, 0.1 M $Pb(NO_3)_2$, 0.1 M $FeCl_3$ and 0.1 M KNO_3 into the wells of your well plate as shown by Figure 1.

2. Clean the copper, zinc, lead, and iron electrodes using steel wool. Rinse with deionized water. Place each metal electrode in its corresponding ionic solution (It is important that the correct metal is in the correct solution, or your cell will not work properly).

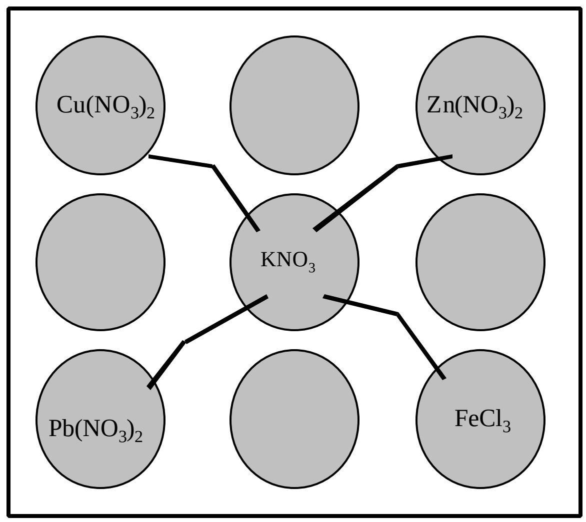
3. Obtain small strips of filter paper, to be used as salt bridges. Wet each strip with 0.1 M KNO_3 and insert one end into the KNO_3 solution and the other end into the $Cu(NO_3)_2$. Repeat for each of the three remaining salt solutions. (Hint: Make sure the strip is entirely wet. Any dry patches will inhibit the flow of ions through the bridge.)

4. Attach the alligator clip from the negative terminal of the voltmeter to one of the metal electrodes and attach the second clip from the positive terminal to a different metal electrode. If the voltmeter has a negative voltage, reverse the hookup so that each clip is now attached to the other metal in the pair.

5. Record the voltage of the electrochemical cell for the reaction occurring at each electrode on the Report Sheet Part A.

6. Repeat for the remaining cells.

Figure 1: Arrangement of ionic solutions for measuring cell potentials



Report Sheet Part A:

Cell	Measured E_{cell}	Anode	Cathode
Cu-Zn	________	________	________
Cu-Pb	________	________	________
Cu-Fe	________	________	________
Pb-Zn	________	________	________
Pb-Fe	________	________	________
Fe-Zn	________	________	________

Which of the four metals is the best oxidizing agent?

Which is the most effective reducing agent?

What purpose does the moist filter paper serve?

Complete the table below. Use the Nernst equation to calculate the potential E_{cell} for each overall cell reaction.

Cell	**Overall Cell Rxn**	**Measured Potential**	**Calculated Potential**	**% Error**
Cu-Zn	____________________	_________	__________	__________
Cu-Pb	____________________	_________	__________	__________
Cu-Fe	____________________	_________	__________	__________
Pb-Zn	____________________	_________	__________	__________
Pb-Fe	____________________	_________	__________	__________
Fe-Zn	____________________	_________	__________	__________

Show sample calculation:

Compare the measured potential and the calculated potential and provide a valid explanation of any major disagreement.

Part B: Electrolysis

Electrolysis takes place in an electrolytic cell. Electrolysis occurs when the passage of an electric current through a solution causes chemical reactions at the electrodes that would otherwise be nonspontaneous. The current from an outside source causes the nonspontaneous transfer of electrons.

When electrolysis occurs in an aqueous solution, the oxidation and reduction of water are always possible reactions. These reactions are easily identified by gas evolution and pH changes in the solution near the electrode. See Chapter 20, Sections 10-11 in the Ebbing/Gammon text.

Reduction (cathode, -) half-reaction for water: $2H_2O(l) + 2e^-$ ® $H_2(g) + 2OH^-(aq)$

Oxidation (anode, +) half-reaction for water: $2H_2O(l)$ ® $O_2(g) + 4H^+(aq) + 4e^-$

1. Weigh 2.0 g of solid KI, in a 200 mL beaker. Add 100 mL of deionized water and mix until the salt dissolves.

2. Pour the solution into the U-tube provided. Using pH paper, measure the pH of the solution at each electrode.

3. Attach the alligator clips from the dc power supply (9V battery) to the carbon electrodes.

4. Record any observations (i.e. gas evolution, color changes).

5. Allow the electrolysis to proceed for 10 minutes.

6. Check the final pH at both electrodes. Remove the electrodes and examine. Record any observations and the probable identity of any substance that might be present.

Report Part B:

Initial pH of KI solution:____________

Initial $[OH^-]$:_______________

Initial time: _______________

Observations occurring within the U-tube during the reaction (i.e. color changes, gas evolution etc.), you may draw a picture if that helps your description:

Electrode observations:

Final time: ______________

Final pH: _______________

Final $[OH^-]$: ________________

What are the possible competing reactions that can occur at the anode? Based on your observations which reaction does occur?

What are the possible competing reactions that can occur at the cathode? Based on your observations which reaction takes place?

What is the overall reaction for the cell?

Calculate the average current that flowed through the electrolytic cell from the $\Delta[OH^-]$, the volume of the solution, and the time of electrolysis.

Part C: Electroplating Pennies

Faraday's Law: number of coulombs = (amperes) $\times$ (seconds)

Faraday's Constant = number of coulombs/mole of electrons

Procedure:

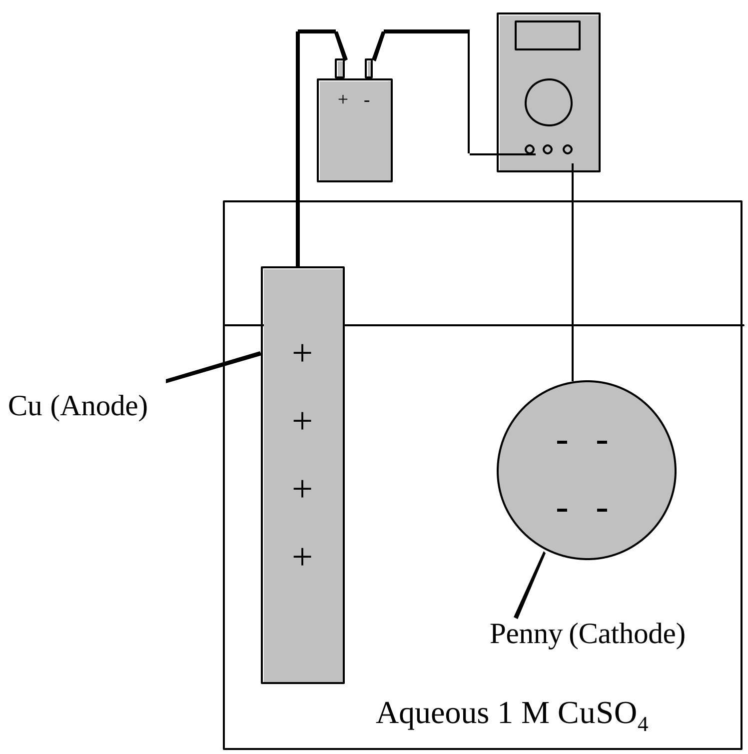
1. Polish a penny and a strip of copper. Measure their mass (+/- 0.001 g). Set up the apparatus as shown in Figure 2, making sure to connect the copper strip to the positive terminal (anode) and the penny to the negative terminal (cathode) of the 9V battery.

2. Add 50 mL of 1.0 M $CuSO_4$ to a 100 mL beaker. Start timing the electrolysis.

3. Allow electrolysis to proceed for 15 min. Record the amperage periodically.

4. After 15 minutes remove the copper strip and the penny and allow to air dry. Record their masses.

Figure 2.



Report Sheet (<u>show all calculations</u>):

Record observations made of both the electrodes and the solution:

Copper Strip:

Initial Mass: ________ Final Mass:___________ Weight Change: __________

Penny:

Initial Mass: ________ Final Mass:___________ Weight Change: __________

Length of time of electrolysis:_________________

Avg. Current (recorded in Amp): ________________

Amount of Cu reduced (recorded in mol):______________

Convert mol of Cu to mol of electrons (e^-):_____________

Using Faraday's Law calculate the number of Coulombs that passed through the electrolytic cell:______

Calculate Faraday's constant:______________

Compare your result with the literature Faraday value to find % error:_____________

Show above calculations: