

CHM152LL: Electrochemical Activity of Metals

Prelab: You do not need to do a prelab assignment for this experiment.

Introduction

The main objective of this experiment is to establish the relative electrochemical reactivity of solid copper, magnesium, lead, silver, and zinc. The activity ranking of the metals will be determined by two different methods. In Part I of this experiment, you will observe direct redox chemical reactions between the solid metals and metal nitrate solutions in order to establish an activity series based on your observations. In Part II of this experiment, you will construct voltaic cells and measure the cell voltages to determine the relative electrochemical activity for these metals.

In addition, you will label the different components of voltaic cell diagrams, and you will write the half reactions and overall reaction for each voltaic cell. The standard Gibbs free energy, ΔG° , and the equilibrium constant, K , will also be calculated for each cell.

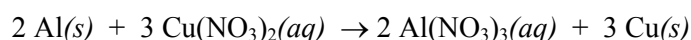
Part I: Electrochemical Activity from Chemical Reactions

An oxidation-reduction reaction occurs when a piece of aluminum metal is placed into a solution of copper(II) nitrate. The aluminum atoms lose three electrons and are oxidized to Al^{3+} ions, whereas the Cu^{2+} ions gain two electrons and are reduced to copper atoms. Hence, a coating of copper metal begins to form at the surface of the aluminum metal due to this electron transfer process.

In Part I of this experiment, you will look for similar evidence of reaction when you place several drops of a metal nitrate solution on top of a clean piece of a different metal. The metals that we will be using are Mg, Pb, Ag, Zn, and Cu; the corresponding metal nitrate solutions are $\text{Mg}(\text{NO}_3)_2$, $\text{Pb}(\text{NO}_3)_2$, AgNO_3 , $\text{Zn}(\text{NO}_3)_2$, and $\text{Cu}(\text{NO}_3)_2$. The least active metal in this series will not react with any of the other metal ions in solution, whereas the most active metal will react with all of the other metal ions. The remaining metals will react with either one, two, or three of the metal ions in solution, respectively. Hence, it will be possible to arrange these five metals in an order of increasing electrochemical activity. **Remember the more active metal prefers to be an ion in a compound (wants to be oxidized) whereas the less active metal prefers to be in its elemental state (wants to be reduced).**

Part II: Electrochemical Activity from Cell Potentials

The oxidation-reduction reaction that takes place between aluminum metal and copper(II) nitrate is a single replacement reaction. The balanced molecular equation for this reaction can be written as follows:

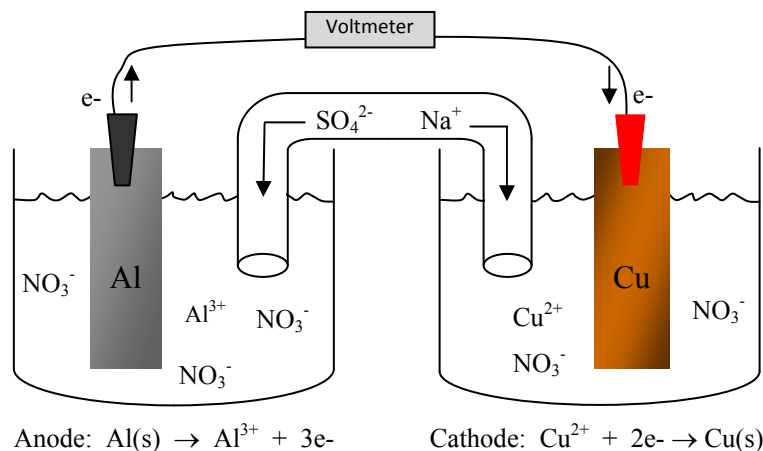


The net ionic equation for this reaction is: $2 \text{Al}(s) + 3 \text{Cu}^{2+}(aq) \rightarrow 2 \text{Al}^{3+}(aq) + 3 \text{Cu}(s)$

This reaction can also be broken down into an oxidation half-reaction and a reduction half-reaction as follows:



These half-reactions can be separated by constructing a voltaic cell consisting of two half-cells as shown below.



Each half-cell consists of a metal electrode immersed in a solution containing its corresponding cation. The electrodes are connected by a wire through which electrons can flow to provide an electrical current. The half-cells are also connected by a salt bridge that permits the limited movement of ions into the two half-cells. This movement of ions completes the circuit and maintains the charge neutrality in each half-cell allowing the reaction to proceed. Note that the electrons do NOT travel through the salt bridge. For the voltaic cell shown above, negative ions from the salt bridge move into the anode half-cell on the left as aluminum atoms are oxidized to form Al^{3+} ions. (Anions move towards the anode.) Positive ions from the salt bridge move into the half-cell on the right as Cu^{2+} ions are reduced to copper metal. (Cations move towards the cathode.)

The electrode at which oxidation occurs is called the anode, and the electrode at which reduction occurs is called the cathode. Thus, electrons move through the external wire from the anode to the cathode. For the example above, the aluminum metal is the anode, and the copper metal is the cathode.

The transfer of electrons takes place through the external (wire) circuit because there is difference in electrical potential between the two electrodes that can be measured in volts. The cell potential is a measure of the tendency of the anode metal to be oxidized (lose electrons) and the tendency of the metal ions in the cathode compartment to be reduced (gain electrons). Thus, cell potentials vary with the composition of the substances being used as electrodes. Cell potentials also vary with the concentrations of ions and pressures of gases, as well as the temperature at which the reaction occurs. The standard cell potential, $\mathcal{E}^{\circ}_{\text{cell}}$, corresponds to 1 M solutions, gases at 1 atm pressure, and a temperature of 298 K.

Standard cell potentials can be considered to be the sum of the standard reduction potential for the cathode half reaction and the standard oxidation potential for the anode half reaction:

$$\mathcal{E}^{\circ}_{\text{cell}} = \mathcal{E}^{\circ}_{\text{red}} + \mathcal{E}^{\circ}_{\text{ox}}$$

In this equation, the oxidation potential for the anode reaction, $\mathcal{E}^{\circ}_{\text{ox}}$, is the negative of the reduction potential for the reverse reaction.

For example, if, $\text{Al}^{3+}(\text{aq}) + 3\text{e}^{-} \rightarrow \text{Al}(\text{s})$ $\mathcal{E}^{\circ}_{\text{red}} = -1.66 \text{ V}$

Then $\text{Al}(\text{s}) \rightarrow \text{Al}^{3+}(\text{aq}) + 3\text{e}^{-}$ $\mathcal{E}^{\circ}_{\text{ox}} = +1.66 \text{ V}$

By convention, the half-reaction $2\text{H}^{+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{H}_2(\text{g})$ is assigned a standard reduction potential of 0.00 volts, and values for the standard reduction potentials of other species are then determined by using the hydrogen electrode as a reference electrode.

In this experiment you will be using the reduction half reaction $\text{Zn}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Zn}(\text{s})$ as your reference point. It's standard reduction potential is -0.76 volts. You will measure the overall cell potentials for voltaic cells composed of zinc and each of the other metals in the series being investigated. You will then calculate the reduction potentials for the other metals from the total cell potential and the given reduction potential of -0.76 V for Zinc. You will compare your results from Part I and Part II.

Procedure

Part I. Electrochemical Activity from Chemical Reactions

Warning: Be careful when handling $\text{AgNO}_3(\text{aq})$ – it will cause dark stains on your skin after it is exposed to sunlight. Make sure to wash your hands thoroughly before you leave the lab!

Note: Do not put any of the $\text{Ag}(\text{s})$ pieces in the waste container – you must clean the silver pieces and return them to the original vial when you are finished with them!

1. From the shelves near the instructor station, obtain the plastic container with the special equipment needed for this lab.
2. Use sandpaper to briefly clean the surface of one strip each of copper, magnesium, and zinc to remove any oxide coatings. *Do not use sandpaper on the silver pieces.* **Cut the Cu, Mg, Pb, and Zn strips into four small pieces**, and place these pieces in the appropriate places on the laminated reaction table. Also place the square silver pieces on the laminated table. Note the difference in the labels of the columns and rows. One set denotes solid metals and the other denotes aqueous solutions. (*Don't use the long Ag wire, this will be used as the electrode in Part II*).
3. Place one drop of the appropriate metal nitrate solutions onto each piece of metal using the laminated table as your guide. **For example, you will place one drop of $\text{Cu}(\text{NO}_3)_2$ on the Zn, Pb, Mg, and Ag pieces.**

- Allow the metal pieces and solutions to remain in contact for at least 15 minutes and proceed with the next part of this experiment. Record your observations in the data table for Part I. (Note: Examine the Mg and $\text{Zn}(\text{NO}_3)_2$ mixture carefully for evidence of a reaction and **see your instructor** to confirm.)
- Have your instructor check your chart and results on the worksheet before you clean the laminated sheet!
- DO NOT PLACE THE SILVER PIECES IN THE WASTE CONTAINER!** Use forceps to remove the Ag pieces from the laminated table, place them in a small beaker and rinse with deionized water. Dry the silver pieces and return them to the original vial. Place the remainder of the metals and solutions in the waste container in the hood, rinse the laminated table with deionized water, and return the table to the special equipment container.

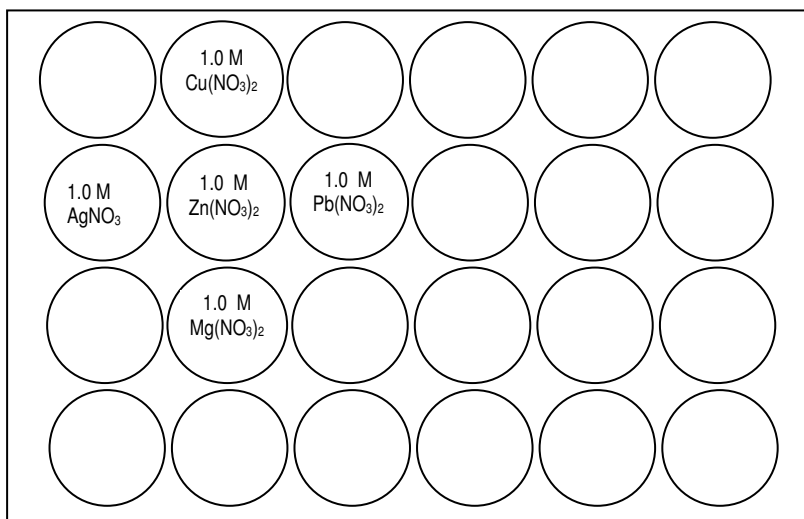
Part II. Electrochemical Activity from Cell Potentials

- Obtain and assemble the notebook computer and LabPro system.
- View the following YouTube animations that show the molecular-level processes occurring in a voltaic cell. Links for these can be found under the Electrochemistry link on the CHM152LL web page.

http://www.youtube.com/watch?feature=player_detailpage&v=J1ljxodF9_g

http://www.youtube.com/watch?feature=player_detailpage&v=C26pH8kC_Wk

- Attach the voltage probe to Channel 1 on the LabPro unit before you open the “Shortcut to GCC Chm Labpro” folder on the desktop. Open the “Electrochemical Cells” file.
- When the voltage probe leads are not in contact with a cell (or each other), a meaningless default voltage may be displayed. If you touch the two leads together, the voltage should drop to 0.00 V (or close to 0.00 V). Also check to make sure that the metal pieces at the end of the leads do not appear corroded. (You may use sandpaper to clean them if they look black due to corrosion.)
- Place enough $\text{Zn}(\text{NO}_3)_2$ solution in one of the center wells to fill it about one-half full, as shown in the figure below. Place about the same amount of 1.0 M AgNO_3 solution in an adjacent well.



- Clean a piece of Zn with sandpaper, then bend the piece near the end. Remove the piece of Ag wire from the vial – do not try to clean the Ag with sandpaper! Use your results from Part I to predict which metal (Zn or Ag) is more active and push the button on the end of the black lead to attach the black lead to this metal. Attach the red lead to the metal that is predicted to be less active. (The newer voltage probes have alligator clips instead of push buttons. If you have a set of these probes simply attach the alligator clip to the appropriate metal.) When the cell voltage is **positive**, as it should be for a spontaneous cell reaction, the positive (red) electrode will be connected to the cathode, and the negative (black) electrode will be connected to the anode.
- To prepare your salt bridge, place about 5-6 ml of 1.0 M NaNO_3 in a 10 ml beaker. Cut four pieces of string so that each piece is about two inches long, and immerse the pieces of string in the NaNO_3 solution. Remove one piece of

string, and place it so one end is inserted into the $\text{Zn}(\text{NO}_3)_2$ solution well and the other end is in the AgNO_3 solution well. **Immediately proceed to the next step.**

Note: Do not allow the metal clips on the black and red leads to come into contact with any solutions; this will corrode and eventually destroy the leads.

8. Insert the Zn electrode into the $\text{Zn}(\text{NO}_3)_2$ solution and the Ag electrode into the AgNO_3 solution, but make sure the black and red leads are not in direct contact with the solutions. After the reading has stabilized, record the voltage reading in the data table for Part II. The reading may fluctuate by about ± 0.01 V and the voltage will decrease with time, so it is important to record the voltage almost immediately after inserting the electrodes. (If the reading erratically fluctuates between (+) and (-) values, make sure there is good contact between the leads and metal electrodes, and that the black/red leads are not touching any solutions.)
9. You should also record which metal is the anode and which is the cathode. **The black (negative) electrode acts as the anode, and the red (positive) electrode acts as the cathode.** If you get a negative voltage, you guessed incorrectly and need to switch your electrodes.
10. Quickly remove the metal electrodes and the salt bridge from the well plate without contaminating the solutions in the wells, and disconnect the metal electrodes from the leads. Rinse the Ag wire electrode with deionized water and return it to the vial containing the silver pieces.
11. Place 1.0 M $\text{Mg}(\text{NO}_3)_2$ solution in a separate well that is adjacent to the $\text{Zn}(\text{NO}_3)_2$ reference well. Clean a new piece of Zn metal, and also clean a piece of Mg. Use a new salt bridge, and measure the Zn-Mg cell potential using the Zn and Mg electrodes. Record the potential in the data table for Part II. Also indicate which metal acts as the anode and which acts as the cathode. Disconnect the metal electrodes, and remove the salt bridge as soon as possible.
12. Check to see whether you need to add more $\text{Zn}(\text{NO}_3)_2$ solution to the reference cell to keep it one-half full. Then proceed in a similar way to measure the cell potentials of the Zn-Pb and Zn-Cu cells, making sure that you use clean metals and fresh salt bridges each time, and that you disconnect the electrodes and remove the salt bridge as soon as possible.

SET THE VIAL WITH THE SILVER PIECES IN THE SPECIAL EQUIPMENT CONTAINER AND RETURN IT TO THE SHELVES BY THE INSTRUCTOR'S STATION.

CHM 152LL Lab Report:

Electrochemical Activity of Metals

Name: _____

Partner(s): _____

Part I. Electrochemical Activity: Chemical Reactions

Indicate whether or not a reaction occurred by writing R or NR.

| Metal | Zn ²⁺ (aq) | Cu ²⁺ (aq) | Pb ²⁺ (aq) | Mg ²⁺ (aq) | Ag ⁺ (aq) |
|-------------------------|-----------------------|-----------------------|-----------------------|-----------------------|----------------------|
| 1.0 M Nitrate solutions | | | | | |
| Zn(s) | | | | | |
| Cu(s) | | | | | |
| Pb(s) | | | | | |
| Mg(s) | | | | | |
| Ag(s) | | | | | |

1. Place the **metals** Zn(s), Cu(s), Pb(s), Mg(s), and Ag(s) in order of **decreasing** electrochemical activity based on your observations from the chemical reactions.

A. _____ B. _____ C. _____ D. _____ E. _____
 most active least active

2. Write correctly **balanced** net ionic reactions for the following selected combinations based on your results for Part I. If there was no reaction simply write NR.

- _____ Zn²⁺(aq) + _____ Cu(s) → _____
- _____ Ag⁺(aq) + _____ Zn(s) → _____
- _____ Cu²⁺(aq) + _____ Mg(s) → _____
- _____ Mg²⁺(aq) + _____ Pb(s) → _____
- _____ Ag⁺(aq) + _____ Cu(s) → _____

Part II. Electrochemical Activity: Cell Potentials

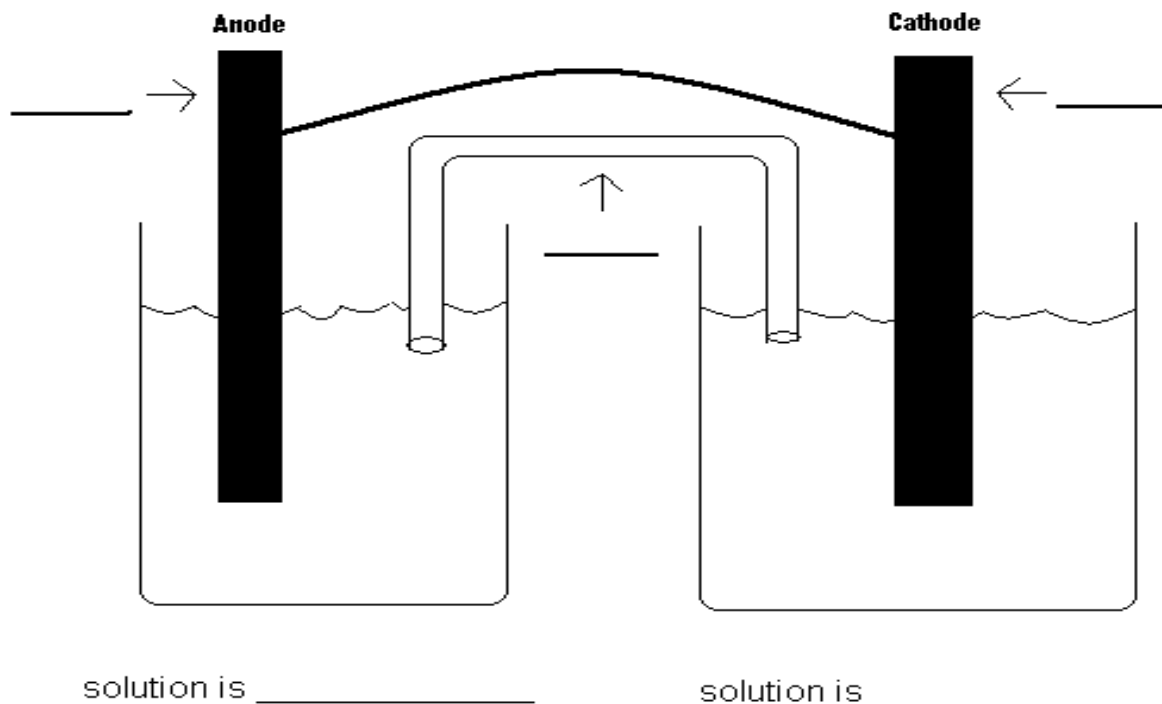
When the cell voltage is **positive**, as it should be for a spontaneous cell reaction, the positive (red) electrode will be connected to the cathode, and the negative (black) electrode will be connected to the anode. Do not dismantle your equipment until your instructor has checked and initialed your data!

| Data Table: Voltaic Cells | | Anode Metal | Cathode Metal | Measured Cell Potential, V |
|---------------------------|-------|-------------|---------------|----------------------------|
| #1 | Zn-Ag | | | |
| #2 | Zn-Mg | | | |
| #3 | Zn-Pb | | | |
| #4 | Zn-Cu | | | |

Instructor signature _____

Voltaic Cell #1: Zn and Ag

Complete the cell diagram below indicating the correct metals and chemicals you used in the experiment in their proper location, the direction of electron flow, and the direction of ion flow in the salt bridge.



Fill in the following information using **your measured experimental cell potential for $\mathcal{E}^{\circ}_{\text{cell}}$** .
 Given $\text{Zn}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Zn}(\text{s})$ $\mathcal{E}^{\circ}_{\text{red}} = -0.76 \text{ V}$, **calculate** the other half reaction potential.

1. half reaction in the anode: _____ $\mathcal{E}^{\circ}_{\text{ox}} =$ _____
2. half reaction in the cathode: _____ $\mathcal{E}^{\circ}_{\text{red}} =$ _____
3. Balanced cell reaction: _____ $\mathcal{E}^{\circ}_{\text{cell}} =$ _____

The standard Gibb's free energy, $\Delta G^{\circ}_{\text{rxn}}$, can be calculated from the cell potential by $\Delta G^{\circ}_{\text{rxn}} = -n F \mathcal{E}^{\circ}_{\text{cell}}$.
 Calculate $\Delta G^{\circ}_{\text{rxn}}$ for cell #1. Show your work. (Include the correct significant figures and units!)

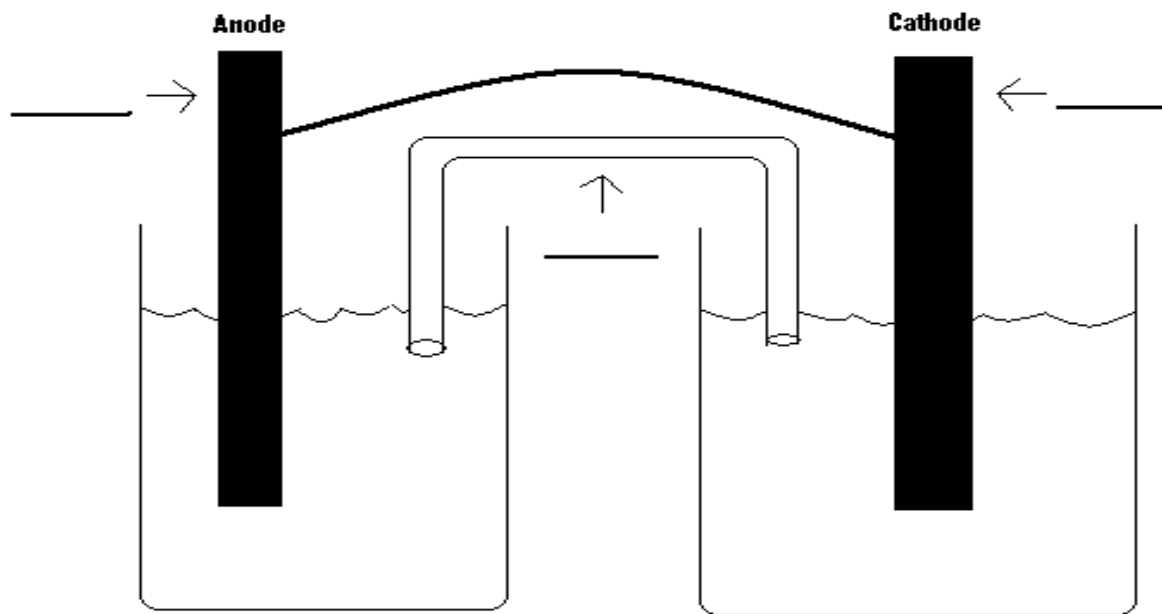
$$\Delta G^{\circ}_{\text{rxn}} = \underline{\hspace{2cm}}$$

The equilibrium constant, K , can be found for a cell reaction by $\mathcal{E}^{\circ}_{\text{cell}} = (RT/n F) (\ln K)$.
 Calculate K for cell #1 (with the correct significant figures). Assume $T = 298 \text{ K}$. Show your work.

$$K = \underline{\hspace{2cm}}$$

Voltaic Cell #2: Zn and Mg

Complete the cell diagram below indicating the correct metals and chemicals you used in the experiment in their proper location, the direction of electron flow, and the direction of ion flow in the salt bridge.



solution is _____

solution is _____

Fill in the following information using **your measured experimental cell potential for $\mathcal{E}^{\circ}_{\text{cell}}$** .
Given $\text{Zn}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Zn}(\text{s})$ $\mathcal{E}^{\circ}_{\text{red}} = -0.76 \text{ V}$, **calculate** the other half reaction potential.

1. half reaction in the anode: _____ $\mathcal{E}^{\circ}_{\text{ox}} =$ _____
2. half reaction in the cathode: _____ $\mathcal{E}^{\circ}_{\text{red}} =$ _____
3. Balanced cell reaction: _____ $\mathcal{E}^{\circ}_{\text{cell}} =$ _____

Calculate $\Delta G^{\circ}_{\text{rxn}}$ for cell #2. Show your work. (Include the correct significant figures and units!)

$$\Delta G^{\circ}_{\text{rxn}} = \underline{\hspace{2cm}}$$

Calculate K for cell #2 (with the correct significant figures). Show your work.

$$K = \underline{\hspace{2cm}}$$

Voltaic Cell #3: Zn and Pb (You do not need to draw the cell or show your work this time.) Fill in the following using your **measured cell potential** for $\mathcal{E}^{\circ}_{\text{cell}}$ and given $\text{Zn}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Zn}(\text{s})$ $\mathcal{E}^{\circ}_{\text{red}} = -0.76 \text{ V}$.

- half reaction in the anode: _____ $\mathcal{E}^{\circ}_{\text{ox}} =$ _____
- half reaction in the cathode: _____ $\mathcal{E}^{\circ}_{\text{red}} =$ _____
- Balanced cell reaction: _____ $\mathcal{E}^{\circ}_{\text{cell}} =$ _____
- As the reaction proceeds in this voltaic cell:
 - $[\text{Zn}^{2+}]$ \uparrow or \downarrow _____
 - $[\text{Pb}^{2+}]$ \uparrow or \downarrow _____
 - The mass of $\text{Zn}(\text{s})$ \uparrow or \downarrow _____
 - The mass of $\text{Pb}(\text{s})$ \uparrow or \downarrow _____
- Calculate $\Delta G^{\circ}_{\text{rxn}}$ and K for this reaction and write the answers in the summary table below.

Voltaic Cell #4: Zn and Cu (You do not need to draw the cell or show your work this time.) Fill in the following using your **measured cell potential** for $\mathcal{E}^{\circ}_{\text{cell}}$ and given $\text{Zn}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Zn}(\text{s})$ $\mathcal{E}^{\circ}_{\text{red}} = -0.76 \text{ V}$.

- half reaction in the anode: _____ $\mathcal{E}^{\circ}_{\text{ox}} =$ _____
- half reaction in the cathode: _____ $\mathcal{E}^{\circ}_{\text{red}} =$ _____
- Balanced cell reaction: _____ $\mathcal{E}^{\circ}_{\text{cell}} =$ _____
- For this voltaic cell:
 - the oxidizing agent is _____
 - the reducing agent is _____
- Calculate $\Delta G^{\circ}_{\text{rxn}}$ and K for this reaction and write the answers in the summary table below.

Summary for Part II.

- Fill in the following summary table with your data and calculated values.

| Cell | $\mathcal{E}^{\circ}_{\text{cell}}$ | $\Delta G^{\circ}_{\text{rxn}}$ | K |
|----------|-------------------------------------|---------------------------------|-----|
| #1 Zn-Ag | | | |
| #2 Zn-Mg | | | |
| #3 Zn-Pb | | | |
| #4 Zn-Cu | | | |

- Examine your data to answer the following question. As the **cell potential decreases**:

The standard Gibb's free energy \uparrow or \downarrow _____

K \uparrow or \downarrow _____

3. Write the appropriate reduction half-reaction for each metal ion in the table below and fill in the E°_{red} value from your experimental data.

| Reduction Half-Reaction Equation | | E°_{red} values from experiment | E°_{red} values from textbook for comparison |
|----------------------------------|---|---|--|
| Zn ²⁺ | Ex. Zn ²⁺ + 2e ⁻ → Zn | -0.76 V | -0.76 V |
| Cu ²⁺ | | | 0.34 V |
| Pb ²⁺ | | | -0.13 V |
| Mg ²⁺ | | | -2.37 V |
| Ag ⁺ | | | 0.80 V |

4. Place the **metals** Zn(s), Cu(s), Pb(s), Mg(s), and Ag(s) in order of decreasing electrochemical activity based on the experimental reduction potentials. (*Hint:* The more active metal is easier to oxidize.)

A. _____ B. _____ C. _____ D. _____ E. _____
 most active least active

5. Identify whether each process occurs at the **cathode or anode** for a voltaic cell. You may want to view the YouTube electrochemical cell animations again to answer these questions.

- Metal atoms are oxidized _____
- Metal ions are reduced _____
- Anions from the salt bridge move towards _____
- Mass of electrode decreases _____
- The concentration of the metal ions decreases _____
- The less active metal is the _____
- Electrons move from _____ to _____

6. Metals are **oxidized** into metal cations as they lose electrons and their charge increases. Thus metals are reducing agents. Metal cations are **reduced** to metal elements (solids except for Hg) as they gain electrons and their charge is reduced to zero. Thus metal cations are oxidizing agents. Answer the following questions about solid Zn, Cu, Pb, Mg, and Ag and their corresponding aqueous ions. Make sure to include (s) or (aq).

- Which is easiest to oxidize? _____
- Which is hardest to oxidize? _____
- Which is easiest to reduce? _____
- Which is hardest to reduce? _____
- Which is the best reducing agent? _____
- Which is the best oxidizing agent? _____