

# Chemistry 163 Laboratory

## ELECTROCHEMICAL CELLS & THERMODYNAMICS

Adapted from Western Washington University

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### \*\*\*\*\* Important Notes \*\*\*\*\*

1. Because of the many details in this lab, you do not need to summarize the experiment in your lab notebook. Instead, you should bring a copy of this lab manual with you to lab.
2. On three separate lab notebook pages, you should show a separate electrochemical cell diagram patterned after the one shown in Figure 2 of this lab manual (see page 5).
3. On page 6 you will see the list of concepts that you will be expected to understand. This may require you to review relevant sections of your textbook in preparation for the lab (see sections below).
4. A conceptual pre-lab assignment is found on the last page of this document.

**\*\* This must be completed and pasted in your lab manual by the beginning of your lab period \*\***

### LEARNING OBJECTIVES AND TEXTBOOK REFERENCES

The learning objectives of this experiment are for you to:

- Learn to build various voltaic cells, measure their cell potentials and to relate them to fundamental redox half reactions.
- Be able to understand the “flow” of electrons, anions and cations in a voltaic cell.
- Be able to apply the Nernst equation to non-standard conditions.

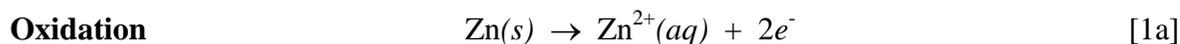
To understand the concepts in this lab, review the following sections of your textbook: 19.2 & 19.3 (basics of voltaic cells), 19.4 & 19.5 (cell potential) and 19.7 (concentration effects & the Nernst eqn.).

### BACKGROUND

Electrochemistry is an increasingly important topic area in modern materials. Batteries, photovoltaic materials, and chemical sensors are but a few examples. At the heart of electrochemical systems is the transfer of electrons, a process that can be understood with reduction-oxidation (“redox”) chemistry. In a **voltaic cell**, electrons, and thus electrical energy, are produced spontaneously. Most redox systems are reversible, and electrical energy can be provided to force a non-spontaneous process to occur; this is called an electrolytic cell. In this experiment you will investigate some properties of spontaneous processes using voltaic cells.

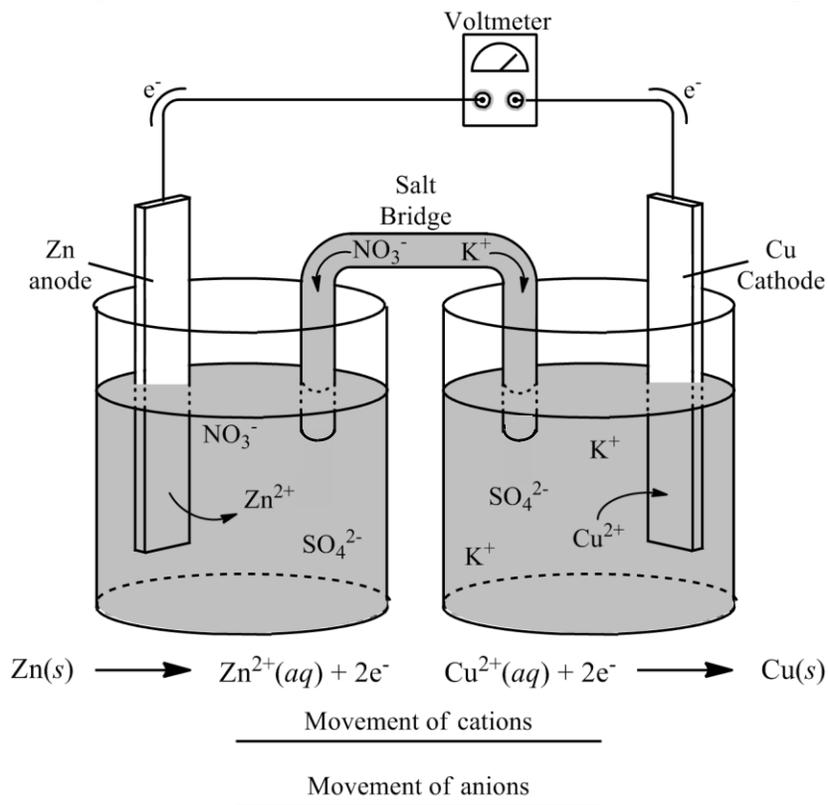
The overall redox reaction can be thought of as a combination of two half reactions. In the *oxidation half reaction*, the process of oxidizing a reactant generates electrons as products (the species losing electrons is said to be oxidized). The *reduction half reaction* uses these electrons as reactants; the species gaining electrons is said to be reduced.

When a piece of zinc metal is immersed in a solution containing copper(II) ions, the zinc is oxidized by the copper(II) ions. Zinc loses electrons and is oxidized while the copper(II) ions gain electrons and are reduced. A useful way to represent these two processes is by writing two half-reactions, which can be summed to give the overall reaction:



This reaction is spontaneous and produces electrical energy. In fact, any spontaneous redox reaction can be used to produce electrical energy; this is the basis of a voltaic cell. These two half reactions can be separated so that the electrons will flow through an external circuit. In order to maintain electrical charge neutrality, a salt bridge containing electrolytes must be provided. A voltaic cell based on Eqn [1] is shown in **Figure 1**. Here, the salt bridge contains  $\text{KNO}_3(aq)$ .

The measured cell potential (also called cell voltage or electromotive force or emf) will be displayed on the computer screen in units of volts. *A volt is a measure of electrical energy per unit charge.* The magnitude of the cell potential is directly related to the thermodynamic tendency of the reaction to occur. The cell potential is dependent on the constituents of the cell as well as their concentrations. Standard practice is to compare standard cell potentials ( $E^\circ_{\text{cell}}$ ). These potentials are measured under standard state conditions: gases at 1 atm, solutions at 1 M concentration, and temperature at 25°C.



**Figure 1.**

Just as the overall cell reaction may be regarded as the sum of two half-reactions, the overall cell potential can be thought of as the sum of two half-cell potentials. The overall cell potential is the sum of the voltage of the oxidation half-reaction (oxidation potential,  $E_{\text{ox}}$ ) and the voltage of the reduction (reduction potential,  $E_{\text{red}}$ ):

$$E_{\text{cell}} = E_{\text{ox}} + E_{\text{red}} \quad [2]$$

Selected standard reduction potentials are tabulated on page 10 of this lab manual.

What would be the  $E^\circ_{\text{cell}}$  in Figure 1 if it was kept at standard-state conditions?  
(Note that changing the direction of the half-reaction changes the sign of the corresponding voltage.)

Answer: \_\_\_\_\_ V

Draw arrows on the figure to indicate the flow of the anions, cations, and electrons.

The voltaic cell in **Figure 1** could be represented by a shorthand method, called a **cell diagram**.



Here, the anode is shown on the left, the double bar represents the salt bridge, and the cathode is shown on the right. It's as easy as **ABC** (**A**node **B**ridge **C**athode) for voltaic cells.

### Non-Standard Conditions and the Nernst Equation

The free energy change,  $\Delta G$ , associated with a chemical reaction is the maximum amount of work that can be obtained from a process. It can also be interpreted as a measure of the driving force or spontaneity of the process. If  $\Delta G$  is negative, the reaction will occur spontaneously in the direction indicated by the chemical equation. The cell potential of a redox process is related to  $\Delta G$  as follows:

$$\Delta G = -n \cdot F \cdot E_{cell} \quad [3]$$

In this equation,  $\Delta G$  is expressed in Joules per mol of completed reactions.  $F$  is the Faraday constant, the electrical charge of 1 mol of electrons ( $F = 96,485 \text{ C/mol } e^-$ ), and  $n$  represents the number of moles of electrons transferred in each mole of chemical reactions. For the case when both the products and reactants are in their standard states, Eqn [3] takes the following form:

$$\Delta G^\circ = -n \cdot F \cdot E^\circ_{cell} \quad [4]$$

The standard  $\Delta G^\circ$ , non-standard  $\Delta G$ , and reaction quotient  $Q$  are related as shown in Eqn [5].

$$\Delta G = \Delta G^\circ + RT \cdot \ln(Q) \quad [5]$$

Here  $R$  is the gas constant (8.3145 J/mol·K) and  $T$  is absolute temperature in Kelvin. The reaction quotient expression, like an equilibrium constant expression, is always written as “products over reactants.” For reaction 1c (also shown below),  $\text{Zn}^{2+}$  is a product and  $\text{Cu}^{2+}$  is a reactant, so  $Q$  is:

$$\text{Zn}(s) + \text{Cu}^{2+}(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{Cu}(s); \quad Q = \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]} \quad [6]$$

After substituting Eqn [3] and [4] into Eqn [5], and rearranging, one obtains:

$$E_{cell} = E^\circ_{cell} - \frac{RT}{nF} \ln Q \quad [7]$$

When applied at  $T = 298 \text{ K}$ , and when natural logarithm is converted to a conventional base-10 logarithm (*i.e.*,  $\ln[x] = 2.3026 \cdot \log[x]$ ), and values of  $R$  and  $F$  are substituted, Eqn [7] becomes:

$$E_{cell} = E^\circ_{cell} - \frac{0.0592}{n} \log Q \quad [8]$$

This is the Nernst equation, which allows us to relate the cell potential to non-standard, non-equilibrium concentrations of reactants or products.

## EXPERIMENTAL PROCEDURE

In your 4-part experiment, you will do the following:

- Measure the cell potential of an alkaline battery to understand the sign convention of voltage measurements and its relationship to the direction of electron flow in a voltaic cell.
- Measure cell potentials for various voltaic cells and compare them to standard cell potentials.
- Determine the effect of concentration on cell potential, and use the Nernst equation to determine the concentration in series of solutions.

**Part 0: Using Voltmeters and Interpreting Their Sign**

Before proceeding, separate pages 9 & 10 of this lab manual.

- Page 9 shows a brief and practical activity to help you understand how to use and interpret voltmeters. **It is to be done before parts I, II and III of the lab.**
- Page 10 shows a pattern for the half cells for the three parts of the lab. On this page you also find a list of standard reduction potentials that you will need for your data analysis.

**Part I: Standard Cells**

**CAUTION:** Handle all solutions with care. The aqueous  $\text{Fe}^{3+}/\text{Fe}^{2+}$  solution is acidic. Use goggles at all times. Ask your instructor or TA how to clean up spills.

For this portion of the lab, you will need to create a number of half-cells. The half-cells generally consist of a solid metal electrode and its corresponding solution containing the metal cation. These half-cells will then be connected to the leads of the voltmeter. The half-cell reaction does not always need to include the solid electrode; we will explore this by using a graphite electrode to replace the iron electrode. The graphite electrode is inert and will not participate in the reaction, but it does allow electrons to be transferred from the wire to the graphite and then to the solution. To complete the circuit, we need what is called a “salt bridge.” This allows ions to transfer while preventing the two half-cells from coming in direct contact. A salt bridge will be required to prevent the buildup of an electrical charge in the half cells (hindering the flow of electrons) and thus maintaining charge neutrality.

1. Before constructing the voltaic cells, connect the voltage probe to Channel 1 of the Vernier LabPro interface. (If you did the “using the voltmeter” activity, this step is already done)
2. Start the *Logger Pro* program on your computer. Open the file “20 Electrochemistry” from the *Advanced Chemistry with Vernier* folder.
3. “Zero” the voltmeter (like taring a balance) by touching the tips of the red and black alligator clips together. Then click the **Experiment** tab and select *Zero* from the pull-down menu.
4. You will take a 10-second average of your cell potential readings. To set up your data acquisition, click the **Experiment** tab, then choose *Data Collection*. Switch the mode to *Events with Entry*. Name the column “Cell Combo #”. Check the box for 10-second averaging, then click *Done*.
5. Wet a ~2 inch long strip of filter paper with 1.0 M  $\text{KNO}_3$  solution. Remove excess liquid gently by blotting it on a paper towel and fold the paper into an L-shape or bend into a U-shape; this will serve as your salt bridge. Place the salt bridge so that it will be immersed into each of two solutions in adjacent wells. **See the well-plate layout on the page 10 of the lab manual.**
6. Construct a voltaic cell by adding aqueous solutions of 0.10 M  $\text{CuSO}_4$  and 0.10 M  $\text{ZnSO}_4$  to two adjacent wells. Add 2 mL of each solution so that the well is about 1/3<sup>rd</sup> full. Do not immerse the solid copper and zinc at this time.
7. In order to get good results, you must “recondition” your electrodes *before each use*.

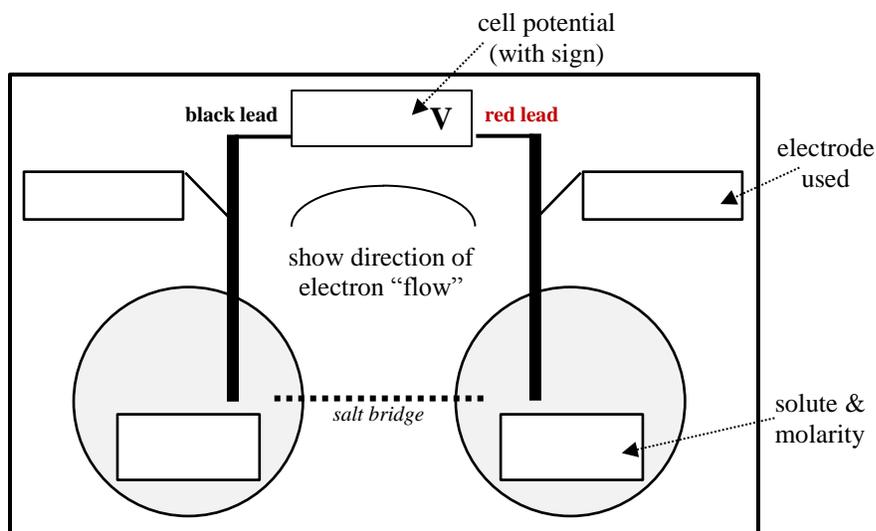
**Polish metal electrodes with steel wool and rinse with DI water immediately before use. Rinse graphite electrodes with DI water and tap dry on a KimWipe.**

Connect the black lead to the copper electrode, and the red lead to the zinc electrode.

8. Click **Collect** to start data collection. Immerse the electrodes in the solutions (avoiding contacting the salt bridge) and within a few seconds click **Keep** to capture the reading. Remove the electrodes after the reading has been obtained. Enter the “Cell Combo #” in numerical order (combination 1, then 2, *etc.*). If you must redo a reading, be sure to re-polish and rinse the metal electrodes (or just rinse the graphite electrode).
9. Repeat steps 5-7 for voltaic cells of *i*)  $C_{\text{graphite}}$  (in  $0.10\text{ M Fe}^{3+}$  &  $0.10\text{ M Fe}^{2+}$ ) connected to Cu (in  $0.10\text{ M CuSO}_4$ ) and *ii*)  $C_{\text{graphite}}$  (in  $0.10\text{ M Fe}^{3+}$  &  $0.10\text{ M Fe}^{2+}$ ) connected to Zn (in  $0.10\text{ M ZnSO}_4$ ). Use a freshly prepared filter paper strip for the salt bridge for each cell. Recondition the electrodes between measurements. Polish metal electrodes with steel wool and rinse with DI water. Rinse graphite electrodes with DI water and tap dry on a KimWipe.

<u>Black</u> -Lead Half-Cell	<u>Red</u> -Lead Half-Cell	$E_{\text{cell}}$ (V)	Direction of $e^-$ “flow”
<b>Cu</b>   $\text{Cu}^{2+}$	<b>Zn</b>   $\text{Zn}^{2+}$		black $\rightarrow$ red <u>or</u> red $\rightarrow$ black
<b>Cu</b>   $\text{Cu}^{2+}$	$C_{\text{graphite}}$   $\text{Fe}^{3+}, \text{Fe}^{2+}$		black $\rightarrow$ red <u>or</u> red $\rightarrow$ black
<b>Zn</b>   $\text{Zn}^{2+}$	$C_{\text{graphite}}$   $\text{Fe}^{3+}, \text{Fe}^{2+}$		black $\rightarrow$ red <u>or</u> red $\rightarrow$ black

10. Record the measured cell potential in the table above. You may find that the voltages recorded may be less than what you would expect based on standard reduction potentials. This is partly due to the presence of interfering oxide formed on the electrodes, a process which occurs very rapidly for readily oxidized metals. In addition, side reactions can sometimes occur, especially at high solution molarities; this is why you are using only  $0.10\text{ M}$  solutions instead of standard  $1.0\text{ M}$  solutions.
11. In your lab notebook, you should sketch cell illustrations (like the one shown below) for *each of the three cells that you analyzed*. Note that these do not necessarily show the anode on the left, since it is based on the connections of the red and black leads. On the diagram you should be able to show the direction of net flow of electrons (either  $\leftarrow$  or  $\rightarrow$ ), and direction of net flow of ions in the salt bridge.



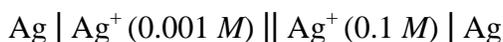
**Figure 2. Electrochemical cell diagram.**

In order to understand your data, you should be able to do the following:

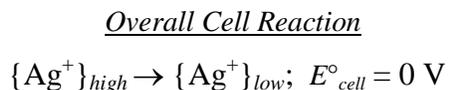
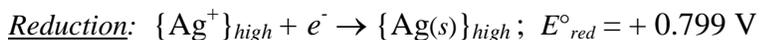
- identify the anode and cathode, and determine what is being oxidized and what is being reduced and where these processes are occurring;
- write the reduction and oxidation half-reactions and the overall cell reaction for all three systems;
- write a voltaic cell diagram (using ABC shorthand) to characterize each system;
- calculate the theoretical, *standard* cell potentials using a table of standard reduction potentials (see the last page of the lab manual for a partial list).

## Part II: Concentration Cell Prepared by Dilution

In this part of the lab, you will find the concentration in an unknown  $\text{Cu}^{2+}$  solution using the Nernst equation and a “concentration cell.” A concentration cell is a voltaic cell in which the anode and cathode compartments are the same except for the concentrations of species present. For example, the electrochemical cell represented by



would generate a *positive* cell potential. To understand why, let’s look at the half-reactions and the overall cell reaction (the half-cells are labeled “high” or “low” based on the concentration of  $\text{Ag}^+$ ).



Note that the *standard* cell potential is zero, but the non-standard cell potential is not zero because of the difference in concentrations. From the cell reaction, an expression for the reaction quotient,  $Q$ , can be written analogous to an equilibrium constant. In this case,  $Q = [\{\text{Ag}^+\}_{\text{low}}]/[\{\text{Ag}^+\}_{\text{high}}]$ .

For the  $\text{Ag}^+$  concentration cell above, the Nernst equation simplifies to:

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.0592}{n} \log Q \stackrel{0}{=} 0 - \frac{0.0592}{1} \log \left( \frac{0.001 \text{ M}}{0.1 \text{ M}} \right) = +0.118 \text{ V} \quad [9]$$

In order to minimize free energy, the  $[\text{Ag}^+]$  in the low-concentration cell ( $[\text{Ag}^+]_{\text{low}}$ ) will *increase* over time, and the  $[\text{Ag}^+]$  in the high-concentration cell ( $[\text{Ag}^+]_{\text{high}}$ ) will *decrease* over time. As the reaction proceeds, the  $[\text{Ag}^+]$  in the connected half-cells will get closer and closer to each other. Ultimately the concentrations will become equal, and the non-standard cell potential will drop to zero, as predicted by the Nernst equation.

Instead of knowing the concentrations and predicting the potential, you could measure the cell potential and, if you knew the concentration in one half cell, you could use the Nernst equation to calculate the concentration of the unknown solution in the other half cell. You will apply this method to the following concentration cell. The unknown solution is one you create by dilution.



1. “Zero” the voltmeter as you did in Part I.
2. Obtain ~4 mL of 1.00 M  $\text{CuSO}_4$  in a clean small test tube. Add ~2 mL of this solution to two of the wells; see the well layout (**Figure 3**).
3. You will need two copper electrodes (obtain a new copper electrode and remove the copper electrode from the half-cell used in Part I). Clean and re-polish the electrodes with steel wool immediately before use. Rinse with DI water.

- Prepare a salt bridge and measure the potential difference between these two solutions. This should show 0.0 V because both solutions have the same concentration of  $\text{Cu}^{2+}$ .
- Use a disposal plastic pipette and add 2 drops of the 1.00 M  $\text{CuSO}_4$  to the empty well plate adjacent to one of the wells in step 1 (see the layout in **Figure 3**). Using a plastic pipette, dilute the 2 drops of  $\text{CuSO}_4$  solution by adding ~ 2 mL of deionized water to the well.
- Connect the two half-cells with a freshly prepared salt bridge. Connect the copper electrodes to the voltage probe leads. Attach the black lead to the electrode for the diluted solution, and the red lead to the 1.00 M  $\text{CuSO}_4$  solution, as shown in the table below. Measure and record the cell potential in the table below using the same technique as in Parts I and II.

<u>Black</u> Lead Half-Cell	<u>Red</u> Lead Half-Cell	$E_{\text{cell}}$ (V)
$\text{Cu} \mid \text{Cu}^{2+}$ ( <i>diluted; ? M</i> )	$\text{Cu} \mid \text{Cu}^{2+}$ (1.0 M)	

- Analyzing this system is like answering the question: What  $[\text{Cu}^{2+}]$  will give the cell potential that we measured? When you have completed your lab work, you should complete the following analysis on a page in your lab notebook.
  - Write the overall cell reaction; be sure to label the solutions so you can keep track of the two different  $\text{Cu}^{2+}$  concentrations. Use the “high” and “low” terminology in the Ag concentration cell above as an example.
  - Based on the cell reaction, write an expression for the reaction quotient,  $Q$ .
  - Once the value of  $\log(Q)$  is calculated from the Nernst equation, you can calculate the value of  $Q$  itself.
  - From there, it is straightforward to calculate the  $[\text{Cu}^{2+}]$  in the diluted solution.

### Part III: Concentration Cells Prepared by Reacting $\text{Cu}^{2+}$

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In this section you will prepare two concentration cell systems by chemically reacting  $\text{Cu}^{2+}$ .

- For this part of the lab, you will need the following:
  - two copper electrodes, as in part II,
  - one 1-mL plastic syringe (rinse well with DI water before using),
  - one wooden stir stick.
- “Zero” the voltmeter as you did in Part I.
- Be sure to read steps 3-4 below as these need to be completed in only a few minutes after step 2. You can make observations in your lab notebook at the end of these steps.**

To the well labeled “A” for part III (see **Figure 3**) and its partner just to the right of it, add 2 mL of 0.10 M  $\text{CuSO}_4(\text{aq})$ . *Do not use 1.0 M  $\text{CuSO}_4$ .* Initially, you will measure the cell potential between these two cells. What should the cell potential be if  $[\text{Cu}^{2+}]$  is the same?

Connect the two half-cells with a freshly prepared salt bridge and clean electrodes. Connect the copper electrodes to the voltage probe leads. Attach the black lead to the electrode with the solution in well plate A, and the red lead to the cell immediately to the right (contains 0.10 M  $\text{CuSO}_4$ ). Measure and record the cell potential in the table below.

A: <u>Black</u> Lead ½-Cell	<u>Red</u> Lead ½-Cell	<i>Initial</i> $E_{\text{cell}}$ (step 2)	$E_{\text{cell}}$ after $\text{NH}_3$ added	$E_{\text{cell}}$ after $\text{H}_2\text{SO}_4$ added
$\text{Cu} \mid \text{Cu}^{2+}$ ( <i>cell A</i> )	$\text{Cu} \mid \text{Cu}^{2+}$ (0.10 M)	V	V	V

- With the electrodes and salt bridge still in place, dispense 4 drops of the the stock 3.0 M  $\text{NH}_3(aq)$  to cell A; do not add this to both cells. Stir with a wooden stick. Collect the cell potential data using the 10-s average as in part I. Write the cell potential in the table above and observations in your lab notebook. Note any evidence of a chemical reaction.
- Next, to cell A only, add 6.0 M  $\text{H}_2\text{SO}_4(aq)$  one drop at a time and stir with rinsed wooden stick that has been dried with a KimWipe. After a few drops the cell potential will stabilize. Record the voltage in the table in step 3. Make observations in your notebook.
- Now for cell B, repeat steps 2-4, but this time using  $\text{NaOH}(aq)$  instead of  $\text{NH}_3(aq)$ , according to the following steps.
  - To the well labeled “B” for part III (see **Figure 3**) and its partner just to the right of it, add 2 mL of 0.10 M  $\text{CuSO}_4(aq)$ . Clean the electrodes and prepare a fresh salt bridge.
  - Connect the copper electrodes to the voltage probe leads. Connect the black lead to the electrode with the solution in well plate B, and the red lead to the 0.10 M  $\text{CuSO}_4$  solution immediately to its right, from part II (see **Figure 3**). Measure and record the cell potential in the table below. Complete the next two steps in fairly rapid succession.
  - With the electrodes and salt bridge still in place, dispense 4 drops of 3.0 M  $\text{NaOH}(aq)$  to cell B only; do not add this to both cells. Stir with a wooden stick (rinsed & dried with a KimWipe). Collect the cell potential data; record in the table above.
  - To cell B only, add 6.0 M  $\text{H}_2\text{SO}_4(aq)$  one drop at a time and stir with a wooden stick (rinsed and dried with a KimWipe). After a few drops the cell potential will stabilize. Collect the cell potential data and record the voltage in the table. Make observations.

B: <u>Black</u> Lead ½-Cell	<u>Red</u> Lead ½-Cell	Initial $E_{\text{cell}}$ (B: 0.1 M $\text{Cu}^{2+}$ )	$E_{\text{cell}}$ after NaOH added	$E_{\text{cell}}$ after $\text{H}_2\text{SO}_4$ added
Cu   $\text{Cu}^{2+}$ (cell B)	Cu   $\text{Cu}^{2+}$ (0.10 M)	V	V	V

### 7. Lab Clean Up Steps

- Tip the contents of your well plate into the large plastic pitcher in the waste hood.
- Pour the contents of this pitcher into the large plastic waste jug (this prevents spills).
- Wash the well plates with soap and water, and return them to the dispensing bin.
- Electrodes can be cleaned and reused. Please return them to the dispensing area.
- The filter paper used for the salt bridges can be discarded into the waste bins.

- To analyze these data you will have to perform of number of steps in your lab notebook.
  - Write the half reactions and the overall cell reaction for:  $\text{Cu}|\text{Cu}^{2+}(?M)||\text{Cu}^{2+}(0.1M)|\text{Cu}$ .
  - Write an algebraic expression for the reaction quotient,  $Q$ , based on the stoichiometry of the cell reaction, and expressed in terms of the concentrations of the reactants and products. Note that the cell potential and the value of  $Q$  will be different for each half-cell combination measured:
    - $\text{Cu}^{2+}$  diluted with water
    - $\text{Cu}^{2+}$  reacted with  $\text{NH}_3^*$
    - $\text{Cu}^{2+}$  reacted with  $\text{OH}^*$ .

*\* You will calculate the “free”  $[\text{Cu}^{2+}]$  after either  $\text{NH}_3$  or  $\text{NaOH}$  was added but before  $\text{H}_2\text{SO}_4$  was added \**
  - However, the form of the Nernst equation (Eqn [8]) will be the same for all three systems because they all are, essentially, concentration cells. Assume that room temperature is close enough to 298K that the constant 0.0592 V/C in the Nernst equation does not require any modification.

**Part 0: Using Voltmeters and Interpreting Their Sign** (*to be done before Parts I, II & III*)

Before your experiment, familiarize yourself with using voltmeters and interpreting the sign of the measurement in the context of electrochemistry. Complete the brief activity below.

1. Login to your workstation computer. Connect the voltage probe to Channel 1 of the Vernier LabPro interface. Start the *Logger Pro* program on your computer.
2. When leads of the voltmeter are connected directly to each other, there should be no difference in electrical potential (or volts). To ensure this, “zero” the voltmeter (like taring a balance) by touching the tips of the red and black alligator clips together. While holding the leads together, click the **Experiment** tab and select *Zero* from the pull-down menu. Your reading should be 0.000 V;  $\pm$  deviation from this shows the instrument uncertainty.
3. Voltage measured with a voltmeter is for the red lead side relative to the black lead side.
  - If the voltage reading is “+”, then electrons flow from the black-lead side to the red-lead side.
  - In contrast, if the voltage reading is “-”, then electrons flow from the red-lead side to the black-lead side.
4. Try this out with an alkaline battery. Touch the alligator clip on the red lead to the “+” battery terminal, and touch the black lead to the “-” battery terminal. Record your voltage and sign below, in the left column.

Red lead: “+” terminal Black lead: “-” terminal	Red lead: “-” terminal Black lead: “+” terminal
Voltmeter reading:	Voltmeter reading:

Now reverse the leads on the battery so that the red lead is touching the “-” battery terminal and the black lead is touching the “+” battery terminal. Record your measurements above.

5. In a battery, which statement below is true? (*circle one below*)
  - Electrons travel from the “-” battery terminal to the “+” battery terminal.
  - Electrons travel from the “+” battery terminal to the “-” battery terminal.
6. Batteries are voltaic cells. These have oxidation half-cells (where electrons are generated) and reduction half-cells (where electrons are used for reduction). After considering your answer to question 5, choose which answer below is true for a battery? (*circle one below*)
  - Electrons are generated in the half cell associated with the “-” battery terminal.
  - Electrons are generated in the half cell associated with the “+” battery terminal.
7. Electrons are generated by oxidation half-reactions; this is called the *anode* in voltaic cells. Electrons are used as reactants in reduction half reactions; this is the *cathode*. Identify the **anode** and **cathode** on the figure below. If you connected an external circuit, which way would the electrons travel? (use an arrow on the wire represented below)



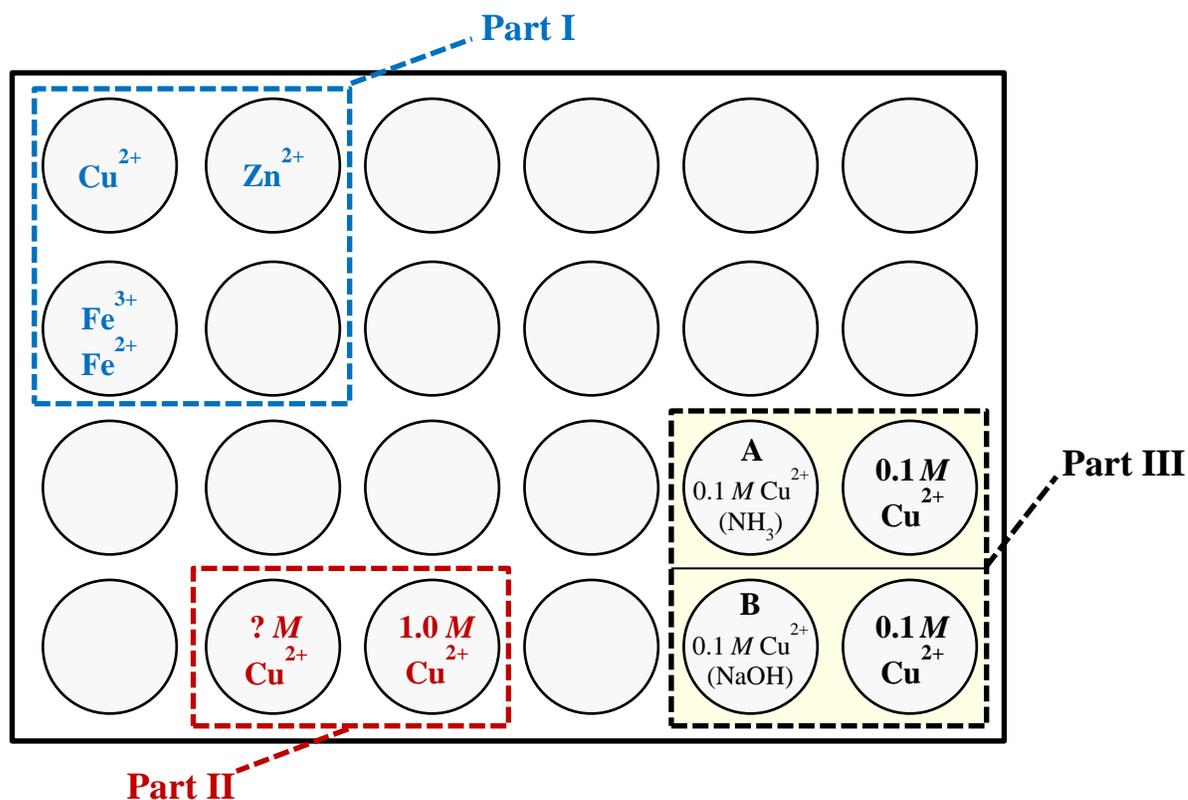


Figure 3. Suggested layout of half cells in the well plate.

Cathode ( <i>reduction</i> ) Half-Reaction	Std. Potential $E^\circ$ (V)
$\text{Mg}^{2+}(\text{aq}) + 2e^- \rightarrow \text{Mg}(\text{s})$	-2.360
$\text{Al}^{3+}(\text{aq}) + 3e^- \rightarrow \text{Al}(\text{s})$	-1.677
$2\text{H}_2\text{O}(\text{l}) + 2e^- \rightarrow \text{H}_2(\text{g}) + 2\text{OH}^-(\text{aq})$	-0.828
$\text{Zn}^{2+}(\text{aq}) + 2e^- \rightarrow \text{Zn}(\text{s})$	-0.762
$\text{Fe}^{2+}(\text{aq}) + 2e^- \rightarrow \text{Fe}(\text{s})$	-0.440
$\text{Ni}^{2+}(\text{aq}) + 2e^- \rightarrow \text{Ni}(\text{s})$	-0.236
$\text{Fe}^{3+}(\text{aq}) + 3e^- \rightarrow \text{Fe}(\text{s})$	-0.040
$2\text{H}^+(\text{aq}) + 2e^- \rightarrow \text{H}_2(\text{g})$	0.000
$\text{Cu}^{2+}(\text{aq}) + e^- \rightarrow \text{Cu}^+(\text{aq})$	0.161
$\text{Cu}^{2+}(\text{aq}) + 2e^- \rightarrow \text{Cu}(\text{s})$	0.339
$\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4e^- \rightarrow 4\text{OH}^-(\text{aq})$	0.40
$\text{Cu}^+(\text{aq}) + e^- \rightarrow \text{Cu}(\text{s})$	0.518
$\text{Fe}^{3+}(\text{aq}) + e^- \rightarrow \text{Fe}^{2+}(\text{aq})$	0.746
$\text{Ag}^+(\text{aq}) + e^- \rightarrow \text{Ag}(\text{s})$	0.799
$\text{NO}_3^-(\text{aq}) + 4\text{H}^+(\text{aq}) + 3e^- \rightarrow \text{NO}(\text{g}) + 2\text{H}_2\text{O}(\text{l})$	0.955
$\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4e^- \rightarrow 2\text{H}_2\text{O}(\text{l})$	1.229
$\text{Cl}_2(\text{g}) + 2e^- \rightarrow 2\text{Cl}^-(\text{aq})$	1.360
$\text{H}_2\text{O}_2(\text{aq}) + 2\text{H}^+(\text{aq}) + 2e^- \rightarrow 2\text{H}_2\text{O}(\text{l})$	1.763
$\text{O}_3(\text{g}) + 2\text{H}^+(\text{aq}) + 2e^- \rightarrow \text{O}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$	2.075

Source: Harris, Quantitative Chemical Analysis, 8<sup>th</sup> ed, 2010.

### Chem 123: Electrochemistry & Thermodynamics: Pre-Lab Concept Questions

*(this is due at the beginning of the lab period; it will be checked by your lab instructor)*

**Answer these questions and paste/tape this into your lab notebook.**

In part III of the lab (see p. 7), you perform a series of chemical reactions which represent the types of chemical reactions and equilibria that you have studied throughout Chem 123. These reactions are done in the well-plate locations labeled A and B in the lab instructions.

1. For half-cell A, when  $\text{NH}_3(aq)$  is added to  $\text{Cu}^{2+}(aq)$ , a complex ion will form via a Lewis acid-base reaction. The  $K_f$  for  $\text{Cu}(\text{NH}_3)_2^{2+}$  is  $2.1 \times 10^7$ .

a. Write the reaction for the formation of  $\text{Cu}(\text{NH}_3)_2^{2+}$ .

- b. Interpret the magnitude of  $K_f$ . When  $\text{NH}_3(aq)$  is added to  $\text{Cu}^{2+}(aq)$ , what will happen to the unreacted or “free”  $\text{Cu}^{2+}$  concentration? *Choose one.*

$[\text{Cu}^{2+}]$  will increase.    
   $[\text{Cu}^{2+}]$  will decrease.    
   $[\text{Cu}^{2+}]$  will not change.    
  It cannot be predicted.

- c. After the complex ion forms, the next step in the lab has you add  $\text{H}_2\text{SO}_4(aq)$ .

Write the reaction that will occur.

- d. When the  $\text{H}_2\text{SO}_4(aq)$  is added predict, what will happen to the “free”  $[\text{Cu}^{2+}]$ ? *Choose one.*

$[\text{Cu}^{2+}]$  will increase.    
   $[\text{Cu}^{2+}]$  will decrease.    
   $[\text{Cu}^{2+}]$  will not change.    
  It cannot be predicted.

2. For half-cell B, when  $\text{NaOH}(aq)$  is added to  $\text{Cu}^{2+}(aq)$ , a sparingly soluble solid,  $\text{Cu}(\text{OH})_2$ , will form.

a. Write the reaction for the formation of  $\text{Cu}(\text{OH})_2$ . The  $K_{sp}$  for  $\text{Cu}(\text{OH})_2$  is  $4.8 \times 10^{-20}$ .

- b. Interpret the magnitude of  $K_{sp}$ . When  $\text{NaOH}(aq)$  is added to  $\text{Cu}^{2+}(aq)$ , what will happen to the “free”  $[\text{Cu}^{2+}]$ ? *Choose one.*

$[\text{Cu}^{2+}]$  will increase.    
   $[\text{Cu}^{2+}]$  will decrease.    
   $[\text{Cu}^{2+}]$  will not change.    
  It cannot be predicted.

- c. After the nearly insoluble solid forms, the next step in the lab has you add  $\text{H}_2\text{SO}_4(aq)$ .

Write the reaction that will occur.

- d. When the  $\text{H}_2\text{SO}_4(aq)$  is added, predict what will happen to the “free”  $[\text{Cu}^{2+}]$ ? *Choose one.*

$[\text{Cu}^{2+}]$  will increase.    
   $[\text{Cu}^{2+}]$  will decrease.    
   $[\text{Cu}^{2+}]$  will not change.    
  It cannot be predicted.