

## Experiment 11

# Electrochemistry

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## Pre-Lab Assignment

Before coming to lab:

- Read the lab thoroughly.
- Answer the pre-lab questions that appear at the end of this lab exercise. The questions should be answered on a separate (new) page of your lab notebook. Be sure to show all work, round answers, and include units on all answers.
- Follow the guidelines in the "Lab Notebook Policy and Format for Lab Reports" section of the lab manual to complete in your lab notebook the following sections of the report for this lab exercise: Title, Lab Purpose, and Procedure.

## Purpose

This lab activity has four parts. In Part A and B you will measure the potential of several voltaic cells and compare your measured values to the accepted values that can be found in the tables in your textbook. In Part C, you will construct a simple electrolytic cell, which you can then use to electroplate a nickel (or other coin) with copper. The plated coin will be a souvenir that you can take with you as your reward for successfully completing an entire year of chemistry! In Part D, you will make observations that will allow you to determine the reactions that are occurring in an electrolytic cell when there is more than one possibility.

## Background

As you saw in Chem1A, oxidation–reduction reactions involve a transfer of electrons. In a spontaneous redox reaction, electrons flow from the reactant that is oxidized (**reducing agent**) to the reactant that is reduced (**oxidizing agent**). If the two half–reactions are separated by a wire, this flow of electrons, instead of occurring at the surface of the metal, occurs through the external wire and an electric current is generated. This is called a **voltaic cell** (or galvanic cell), and is exactly how a battery works. Batteries, like the ones found in a flashlight or a calculator, contain oxidizing and reducing substances. As the electrons are transferred, they are run through the device in order to provide the voltage necessary to power the flashlight or calculator.

A good analogy for the flow of electrons is the flow of water. Water flows spontaneously downhill. Dams and waterwheels are examples of ways that the energy of flowing water is tapped to generate power. In order for a redox reaction to serve as a source of power, the reaction must be spontaneous. You will examine these processes in Parts A and B of the lab.

Sometimes we want water to flow uphill. In this case, we need to supply energy in the form of a pump to make this happen. Similarly, if a redox chemical reaction isn't spontaneous, we can use electricity to make the reaction occur. An **electrolytic cell** is a device that uses electricity to drive a non–spontaneous redox reaction. For example, water can be separated into hydrogen and oxygen gas (a non–spontaneous reaction) using electricity. You will examine these processes in Part C and D of the lab.

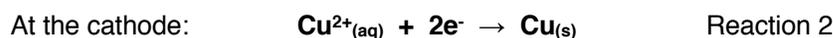
In short, the field of electrochemistry has two important applications- the use of spontaneous redox reactions to generate electricity, and the use of electricity to force non-spontaneous redox reactions to occur.

## Voltaic Cells

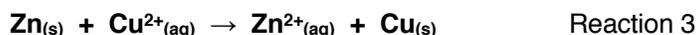
In Part A of this lab activity you will measure the potential of several voltaic cells. A typical **voltaic cell**, such as the one in the figure on the next page consists of two half-cells linked by a wire and a salt bridge. Each half-cell consists of metal electrode in contact with a solution containing a salt of that metal. One half-cell (think beaker in today's lab) functions as the **anode**, where the **oxidation** reaction takes place. Oxidation involves the loss of one or more electron(s) from the metal. For the first voltaic cell you will make and that is diagramed on the next page, the oxidation reaction is:



The other half-cell (think beaker in today's lab) functions as the **cathode**, where the **reduction** reaction takes place. For the first voltaic cell you will make and that is diagramed on the next page, the reduction reaction is:



Combining these reactions and cancelling electrons, the overall reaction is



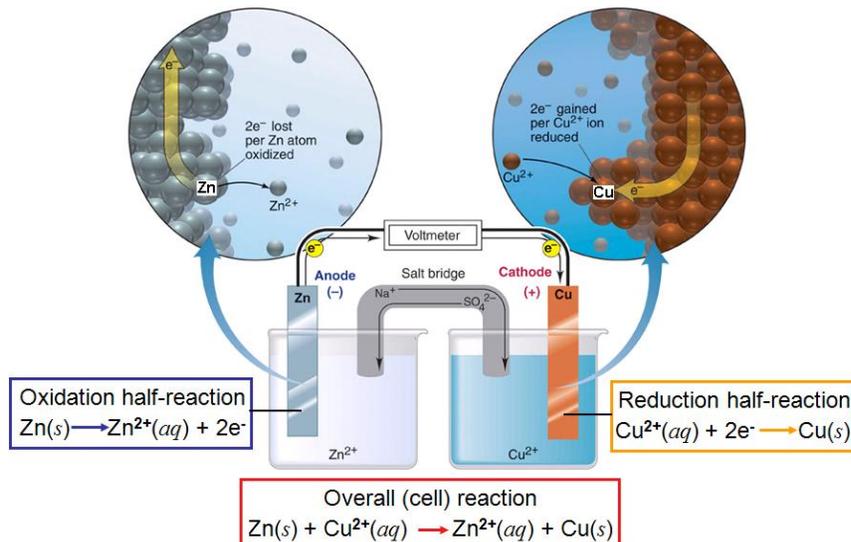
In this case, the zinc metal goes into solution as zinc ions, and the copper ions form solid copper. In addition, electrons flow from the anode (where they are released) to the cathode (where they are taken in) via the wire. As the figure on the following page shows, they are also flowing through a voltmeter. Inserting a voltmeter into the circuit between the half-cells permits a measurement of the voltage, or **potential difference**, between the half-cells. In general, this voltage is designated by the symbol, **E**. When the solutions are 1 Molar (1.0 M) with respect to the ions involved in the oxidation and reduction reactions, the cell is called a standard cell and its voltage is known as a **standard potential**, **E°**. Many textbooks and references books contain tables of standard reduction potentials which show the values of E° for various reduction half-reactions. The E° in the tables are all relative to the reduction of hydrogen as shown in the reaction:



The potential of this reaction has been assigned a potential of 0.00 Volts (V) and is called the **standard hydrogen electrode**.

The final piece of a voltaic cell is a salt bridge which allows migration of ions to prevent an imbalance of charge from building up as electrons leave the anode and move to the cathode.

Figure 1. A voltaic cell based on the zinc-copper reaction



**Figure 1.** The **anode half-cell** consists of a zinc electrode submerged into a solution containing  $\text{Zn}^{2+}$  ions (e.g. aqueous  $\text{Zn}(\text{NO}_3)_2$ ), while the **cathode half-cell** consists of a copper electrode submerged in a solution of  $\text{Cu}^{2+}$  ions (e.g. aqueous  $\text{Cu}(\text{NO}_3)_2$ ). **Oxidation** occurs at the anode as each zinc atom loses two electrons to form aqueous  $\text{Zn}^{2+}$ , thus decreasing the mass of the anode as oxidation occurs. These electrons move through the zinc electrode, the wire and voltmeter on to the copper electrode submerged in the solution of  $\text{Cu}^{2+}$  ions. Finally, **reduction** occurs on the surface of the cathode as the electrons react with the  $\text{Cu}^{2+}$  ions in solution to form solid copper metal which plates out on the surface of the cathode, thereby increasing the cathode's mass in the process. A **salt bridge** contains unreactive sodium ions ( $\text{Na}^+$ ) and sulfate ions ( $\text{SO}_4^{2-}$ ) that maintain neutral charge in the electrolyte solution: Anions in the salt bridge flow to the left, and cations to the right. The voltmeter measures the electrical output of the cell. Since zinc is above copper on the **activity series**, zinc is more easily oxidized than copper. Consequently, the electrical current flows from the zinc electrode (anode) to copper electrode (cathode).

## Predicting the Potential of a Voltaic Cell

For today's lab, you will be predicting the potential that the voltaic cells you construct should generate and then compare your predicted values to your measured values. To order to predict the voltage, you will need to use the table of Reduction Half-Reaction Potentials below. A more complete table is in your book on page 1064.

These ions like to be reduced. I know that since their  $E_{\text{red}}$  is a large positive number

Reduction Half-Reaction	$E_{\text{red}}^{\circ}$ , V
$\text{Au}^{3+}(\text{aq}) + 3e^{-} \rightarrow \text{Au}(s)$	+1.50
$\text{Pt}^{2+}(\text{aq}) + 2e^{-} \rightarrow \text{Pt}(s)$	+1.20
$\text{Ag}^{+}(\text{aq}) + e^{-} \rightarrow \text{Ag}(s)$	+0.80
$\text{Cu}^{2+}(\text{aq}) + 2e^{-} \rightarrow \text{Cu}(s)$	+0.34
$2\text{H}^{+}(\text{aq}) + 2e^{-} \rightarrow \text{H}_2(s)$	0.000
$\text{Pb}^{2+}(\text{aq}) + 2e^{-} \rightarrow \text{Pb}(s)$	-0.13
$\text{Sn}^{2+}(\text{aq}) + 2e^{-} \rightarrow \text{Sn}(s)$	-0.14
$\text{Ni}^{2+}(\text{aq}) + 2e^{-} \rightarrow \text{Ni}(s)$	-0.28
$\text{Fe}^{2+}(\text{aq}) + 2e^{-} \rightarrow \text{Fe}(s)$	-0.44
$\text{Zn}^{2+}(\text{aq}) + 2e^{-} \rightarrow \text{Zn}(s)$	-0.76
$\text{Al}^{3+}(\text{aq}) + 3e^{-} \rightarrow \text{Al}(s)$	-1.66
$\text{Mg}^{2+}(\text{aq}) + 2e^{-} \rightarrow \text{Mg}(s)$	-2.37
$\text{Na}^{+}(\text{aq}) + e^{-} \rightarrow \text{Na}(s)$	-2.71

These metals like to be oxidized. I know that because if I reverse the reaction, I would get a large positive value.

Notice that all reactions are listed as reduction reactions in the table. However, in voltaic cells there are always one oxidation half-reaction and one reduction half-reaction taking place. This means that we must “flip” one of the reactions in the table in order to calculate the potential created in the voltaic cell. Recall that voltaic cells are spontaneous processes ( $E_{\text{cell}} = \text{positive number}$ ) so nature keeps as a reduction reaction the reaction that has the most positive  $E_{\text{red}}$  and will “flip” the other reaction.

### Example: Calculation of the Cell Potential of Standard Voltaic Cells

Whenever two standard half-cells are joined to create a voltaic cell as in Figure 1, one of the half-cells will function as the anode where oxidation will occur. The other half-cell will serve as the cathode where reduction will occur. Consider the half-reactions for a voltaic cell composed of copper and zinc.

From the Reduction Half-Reaction Potentials Table:

$\text{Cu}^{2+} (\text{aq}) + 2\text{e}^{-} \rightarrow \text{Cu} (\text{s})$	+0.34
$\text{Zn}^{2+} (\text{aq}) + 2\text{e}^{-} \rightarrow \text{Zn} (\text{s})$	-0.76

When considering which one of these reactions will stay as the reduction reaction and which one will be the oxidation reaction, the species that generates the more positive potential prefers to be reduced. In this example, that is  $\text{Cu}^{2+}$ . Therefore, the top reaction will be the reduction half-reaction and the reaction involving zinc will be the oxidation reaction. To determine the voltage, we will use the copper reaction as written in the table, but reverse the zinc reaction.

$\text{Cu}^{2+} (\text{aq}) + 2\text{e}^{-} \rightarrow \text{Cu} (\text{s})$	+0.34
$\text{Zn} (\text{s}) \rightarrow \text{Zn}^{2+} (\text{aq}) + 2\text{e}^{-}$	+0.76

Notice that the potential for the reversed zinc reaction is the same value, but opposite in sign. The overall potential for this voltaic cell is the sum of these two values.

$$E^{\circ}_{\text{cell}} = 0.34 \text{ V} + 0.76 \text{ V} = 1.10 \text{ V}$$

### Calculation of the Cell Potential of a Voltaic Cell NOT at Standard Conditions (1.0 M, 298 K, 1atm)

If the potential that is measured is not a standard cell, then the **Nernst equation** can be used to calculate the standard cell potential.

$$\text{Nernst Equation: } E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{RT}{nF} \ln Q$$

Where...

**R**= 8.14 J/mol\*K

**F**= Faraday's constant (96485 Coulombs/mole  $\text{e}^{-}$ )

**T**= Temperature (K)

**Q** = Reaction quotient

**n** = number of electrons transferred

**$E^{\circ}$**  = standard cell potential (1 M solutions at 298K and 1 atm)

**$E_{\text{cell}}$**  = Cell potential at nonstandard conditions

### Example: Calculation of the Cell Potential for a Voltaic Cell at Non-standard Conditions

Consider a voltaic cell made of a Sn electrode in 0.10 M Sn<sup>2+</sup> solution in one half-cell and Al in 0.10 M Al<sup>3+</sup> solution in the other at a temperature of 273 K.

First find the potential at standard conditions ( $E^\circ$ )

$\text{Sn}^{2+} (\text{aq}) + 2\text{e}^- \rightarrow \text{Sn} (\text{s})$	-0.14
$\text{Al}^{3+} (\text{aq}) + 3\text{e}^- \rightarrow \text{Al} (\text{s})$	-1.66

Between these two reactions, Sn<sup>2+</sup> prefers to be reduced since it is "more positive" (closer to zero in this case). Therefore, the Al<sup>3+</sup> reaction will have to be reversed.

$\text{Sn}^{2+} (\text{aq}) + 2\text{e}^- \rightarrow \text{Sn} (\text{s})$	-0.14
$\text{Al} (\text{s}) \rightarrow \text{Al}^{3+} (\text{aq}) + 3\text{e}^-$	+1.66

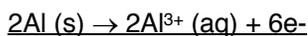
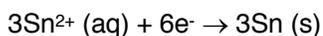
$$E^\circ_{\text{cell}} = -0.14 \text{ V} + 1.66 \text{ V} = 1.52 \text{ V}$$

The overall cell reaction after balancing and cancelling electrons will be



**same in**

**each reaction**



Substituting into the Nernst Equation

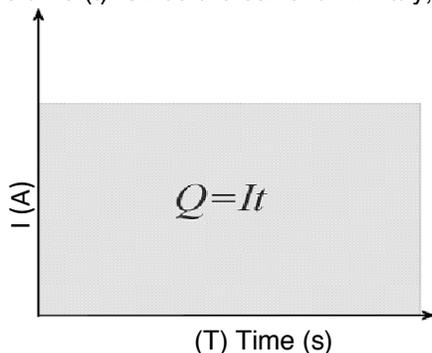
$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{RT}{nF} \ln Q = E^\circ_{\text{cell}} - \frac{RT}{nF} \ln \frac{[\text{Al}^{3+}]^2}{[\text{Sn}^{2+}]^3}$$

$$E_{\text{cell}} = 1.52 \text{ V} - \frac{(8.314 \frac{\text{J}}{\text{mol K}})(273 \text{ K})}{6 \text{ mol} \cdot 96485 \text{ C/mol}} \ln \frac{[0.1]^2}{[0.1]^3}$$

$$E_{\text{cell}} = 1.52 - 0.0099 \text{ V} = \mathbf{1.51 \text{ V}}$$
 (nearly the same as the standard cell potential!)

## Electrolytic Reduction of Cu<sup>2+</sup> to Copper on the Surface of a Coin

In Part C of the lab, copper plating will be accomplished by submerging both a coin and a strip of copper in a solution of copper(II) sulfate. At the anode, copper is oxidized ( $\text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{e}^-$ ) and at the cathode is reduced ( $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$ ). The overall process consists of simply moving copper atoms from one electrode to the other. Since this is not a spontaneous process, an electrical current is needed to drive the process. The total amount of charge used in the reaction can be found by multiplying the current (I) times the time (t). Since the current will vary, we will find the area under the line using the area tool in LabQuest.



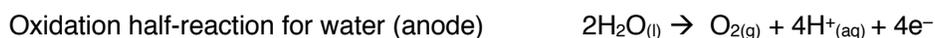
The amount of charge has a symbol of Q-not be confused with the reaction quotient which also has the symbol Q. Aren't there any more letters in the alphabet?

### Example

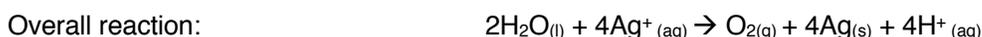
Suppose we plate out copper by supplying 6.30 A (C/s) of current through a  $\text{Cu}(\text{NO}_3)_2$  solution for 14.0 min. What mass of copper is plated?

$$14.0 \text{ min.} \times \frac{60 \text{ s}}{1 \text{ min}} \times \frac{6.30 \text{ C}}{1 \text{ s}} \times \frac{1 \text{ mole e}^-}{96485 \text{ C}} \times \frac{1 \text{ mol Cu}^{2+}}{2 \text{ mole e}^-} \times 63.55 \text{ g} = 1.74 \text{ g Copper}$$

In Part D of the lab, you will run an electrical current through a solution and make observations in order to determine the identity of the products of the electrolysis reaction. Since all the solutions contain water, one of the possibilities is that water will undergo electrolysis. The following half-reactions occur in the electrolysis of water:



Electrolysis of an aqueous solution **may** generate products other than oxygen or hydrogen if the electrolyte contains ions that are more easily oxidized or more easily reduced than water molecules. The electrolysis of aqueous silver nitrate ( $\text{AgNO}_3$ ), for example, produces oxygen at the anode and silver metal at the cathode. If you observe this reaction, you will observe the formation of Ag at the cathode since the reduction of silver ions ( $\text{Ag}^+$ ) to silver (Ag) occurs more readily than reduction of water. At the anion, bubbles are seen which indicates that  $\text{O}_2$  gas is being produced. The overall reaction is the sum of the oxidation and reduction half-reactions:



## Procedure

### Part A: Construction of Voltaic Cells and Measurement of Cell Potentials

1. Obtain about 25 mL of 1 M  $\text{KNO}_3$ . Cut up pieces of filter paper into six small narrow strips. Soak the strips of filter paper in the  $\text{KNO}_3$ . These will serve as your salt bridges.
2. Obtain strips of Cu, Zn, Pb, and Ag. Clean the strips of metal by sanding and/or dipping in 3 M HCl.
3. Place approximately 25 mL of 0.1 M  $\text{CuSO}_4$  in a 100 mL beaker. Place the strip of Cu in the beaker.
4. Place approximately 25 mL of 0.1 M  $\text{ZnSO}_4$  in another 100 mL beaker. Place the strip of Zn in this beaker.
5. Insert the salt bridge so that it is submerged in the solutions in both beakers.
6. Obtain a multimeter. Connect the black wire to the “-COM” terminal of the multimeter and the red wire to the “+V. $\Omega$ .mA” terminal. Turn the multimeter to the 2 Volt DC setting.
7. Before making a measurement, **predict which metal will serve as the anode (oxidation) and which will serve as the cathode (reduction)**—keep in mind the reduction table on page 3 of this lab.
8. Connect the black (negative) electrode to the anode metal strip, and the red (positive) electrode to the cathode metal strip. If you get a negative reading, you have connected the wires backwards and will need to reverse the order.
9. Allow the cell to stand until the voltage reading on the voltmeter has stabilized, then record the highest voltage
10. Measure the potential of all six possible cell combinations using the method described above— see the table below for all possible combinations. **Use a fresh salt bridge for each cell.**

Cell		Cell	
1	Cu / Zn	4	Zn / Pb
2	Cu / Pb	5	Zn / Ag
3	Cu / Ag	6	Pb/ Ag

11. Complete all the questions for each cell as described in the Data Tables and Calculations section of this lab guide.
12. When you have finished, dispose of the metal ion solutions and the filter paper strips in the waste containers provided. Clean the metal strips and return them to the lab cart.

### Part B: Effect of Concentration on Cell Potential

1. Prepare a copper/zinc voltaic cell as in Part A using 0.10 M  $\text{ZnSO}_4$  solution as before, but replace the 0.10 M  $\text{CuSO}_4$  solution with 0.005 M  $\text{CuSO}_4$  solution. Use a cleaned Zn metal strip in the  $\text{ZnSO}_4(\text{aq})$  beaker and a cleaned Cu metal strip in the  $\text{CuSO}_4(\text{aq})$  beaker. Record the voltage of the cell. How does the decrease in concentration of  $\text{CuSO}_4$  affect the voltage of the cell?
2. Prepare a copper/zinc voltaic cell as in Part A, using 0.10 M  $\text{CuSO}_4$  solution as before, but replace the 0.10 M  $\text{ZnSO}_4$  solution with 0.005 M  $\text{ZnSO}_4$  solution. Use a cleaned Cu metal strip in the  $\text{CuSO}_4(\text{aq})$  beaker and a cleaned Zn metal strip in the  $\text{ZnSO}_4(\text{aq})$  beaker. Record the voltage of the cell. Does the decrease in concentration of  $\text{Zn}^{2+}$  ion affect the voltage measured? Why?

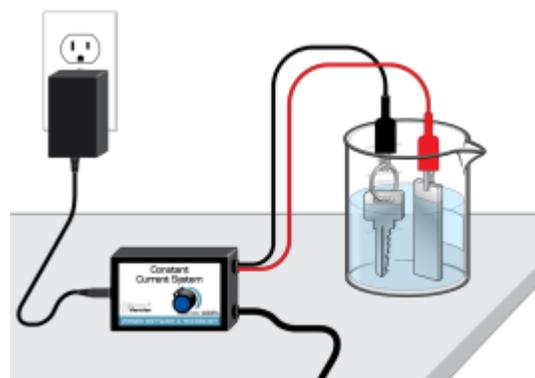
### Part C Electrolytic Reduction of $\text{Cu}^{2+}$ to Copper on the Surface of a Coin

1. Use steel wool to clean a strip of copper, which will be one of the electrode of the electrochemical cell.
2. Mix 3 g of NaCl with 15 mL of vinegar in a 250 mL beaker. Wash a quarter and the copper strip in this salt/vinegar solution. Rinse the coin and copper strip with distilled water and dry each metal piece.
3. Use an analytical balance to determine the mass of the quarter. Record this mass in your notebook.
4. Place 100 mL of 1.0 M copper(II) sulfate solution in a 250 mL beaker such that you will be able to mostly submerge the quarter in the solution when the coin is placed on end.

5. Gently turn the dial of the Constant Current System counterclockwise to confirm that it is in the minimum current position.

6. Use an alligator clip adapter to attach one end of a red patch cord to one end of the copper strip. This will be the anode.

7. Attach one end of black patch cord to the coin (the cathode).



**Important:** You will not place the electrodes in the cell until Step 9.

8. Plug the Constant Current System into a powered electrical outlet.

9. Connect the sensor cable into the USCB port of the LabQuest and choose New from the File menu.

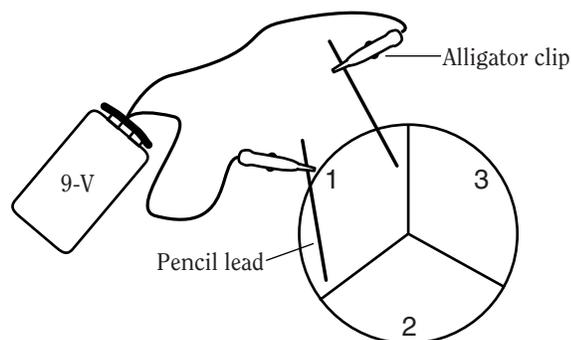
Place the coin and the copper strip into the 1.0 M  $\text{CuSO}_4$  solution in the beaker, and keep the two electrodes as far apart as possible. **Keep the alligator clips out of the solution or they will be plated too!**

10. Start data collection by pressing the green play button in bottom corner of LabQuest interface. Adjust the current to the 0.2–0.3 A range. Observe the electrolysis. Note the slow deposition of copper on the surface of the coin. Continue to collect data for 15 minutes.
11. After data collection has run for ~15 minutes, turn down power supply to 0 Amps and press stop on the LabQuest interface.
12. Carefully remove the copper strip and the coin from the electrolyte solution. Rinse the two metals with distilled water. Gently dry the copper strip and coin so as not to remove copper.
13. Measure and record the mass of the coin in your notebook.
14. Record the value for the area under the current versus time curve. This is the number of coulombs (ampere • seconds) of electricity. This can be done by clicking on the Analyze menu, then Integrate and then Current on the LabQuest interface.
15. Using the coulombs found, calculate the expected mass of copper produced. The electrolytic set-up we used is fairly crude, and there will be some inefficiency in the cell, so don't expect perfect agreement. Recall that the reaction is  $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$

- Determine the percent error between the calculated mass of Cu expected to be electroplated and the actual mass of Cu electroplated.
- Discard the electrolyte solution cell as directed.

### Part D: Electrolysis of Solutions

The purpose of this experiment is to identify the products obtained in the electrolysis of aqueous potassium iodide, copper(II) bromide, and sodium chloride solutions. The electrolysis reactions will be carried out in an electrolytic cell consisting of a Petri dish, a 9-V battery, and graphite electrodes (see figure to the right).



- Place the partitioned Petri dish on a sheet of white paper. Observe that the compartments or segments of the Petri dish are labeled 1, 2, and 3.
- Carefully pour about 8 mL of 0.5 M potassium iodide solution into the first compartment of the Petri dish until the compartment is one-third to one-half full.
- Add 3 drops of phenolphthalein solution and stir to mix.
- Connect the battery cap to the 6-V battery. Carefully attach a “pencil lead” electrode to each alligator clip lead. *Caution:* Do not allow the electrodes to touch each other.
- Hold the red (+) lead from the 6-V battery in one hand and the black (–) lead in the other hand. Keeping the electrodes as far apart as possible, dip the graphite electrodes into the potassium iodide solution.
- Let the electric current run for 1–2 minutes while observing any changes in the potassium iodide solution. Record all observations. See the Data Tables and Calculations for guidance.  
*Note: If you do not see any changes, you may need to use a new power source. Consult your instructor.*
- Remove the graphite electrodes from the electrolysis solution. Carefully rinse the electrodes with distilled water from a wash bottle and gently pat dry on a paper towel.
- Carefully pour about 8 mL of 0.5 M sodium chloride solution into the second compartment of the Petri dish. Add three drops of phenolphthalein indicator solution and stir to mix.
- Repeat steps 5–7 for the electrolysis of sodium chloride solution. Record observations as outlined in the Data Table and Calculations section.
- Carefully pour about 8 mL of 0.2 M copper(II) bromide solution into the third compartment of the Petri dish.
- Repeat steps 5–7 for the electrolysis of copper(II) bromide solution. Record observations as outlined in the Data Table and Calculations section.
- Remove the graphite electrodes from the alligator clips and disconnect the battery cap from the battery.

13. The electrolysis products may include dilute halogen solutions (chlorine, bromine, and iodine). Working in the hood, carefully pour the contents of the Petri dish into a waste beaker containing sodium thiosulfate solution. Sodium thiosulfate will reduce the halogen waste products. Allow the beaker to stand in the hood overnight.

## Data Table and Calculations

### Part A: Construction of Voltaic Cells and Measurement of Cell Potentials

Cell		Cell	
1	Cu/ Zn	4	Zn/Pb
2	Cu/ Pb	5	Zn/Ag
3	Cu/Ag	6	Pb/Ag

Prepare and complete the following table for each of your six voltaic cells

**Example for Cell 1:**

Copper/Zinc cell

Oxidation half-reaction \_\_\_\_\_

Reduction half-reaction \_\_\_\_\_

Balanced Net reaction \_\_\_\_\_

Metal at the anode \_\_\_\_\_

Metal at the cathode \_\_\_\_\_

Experimental cell potential \_\_\_\_\_

Calculated standard cell potential \_\_\_\_\_

% difference \_\_\_\_\_

### Part B: Effect of Concentration on Cell Potential

For the 0.10 M ZnSO<sub>4</sub> / 0.005 M CuSO<sub>4</sub> cell

Anode \_\_\_\_\_

Cathode \_\_\_\_\_

Experimental cell potential \_\_\_\_\_

Calculated cell potential using the Nerst equation \_\_\_\_\_

For the 0.10 M CuSO<sub>4</sub> / 0.005 M ZnSO<sub>4</sub> cell

Anode \_\_\_\_\_

Cathode \_\_\_\_\_

Experimental cell potential \_\_\_\_\_

Calculated cell potential using the Nerst equation \_\_\_\_\_

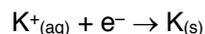
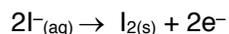
### Part C: Electrolytic Reduction of $\text{Cu}^{2+}$ to Copper on the Surface of a Coin

See the Lab Procedure section to determine the data and calculations needed.

### Part D: Electrolysis of Solutions

For the KI solution,

The following oxidation and reduction half-reactions are possible for the electrolysis of potassium iodide solution. The solution contains water molecules, potassium ions ( $\text{K}^+$ ), and iodide ions ( $\text{I}^-$ ).



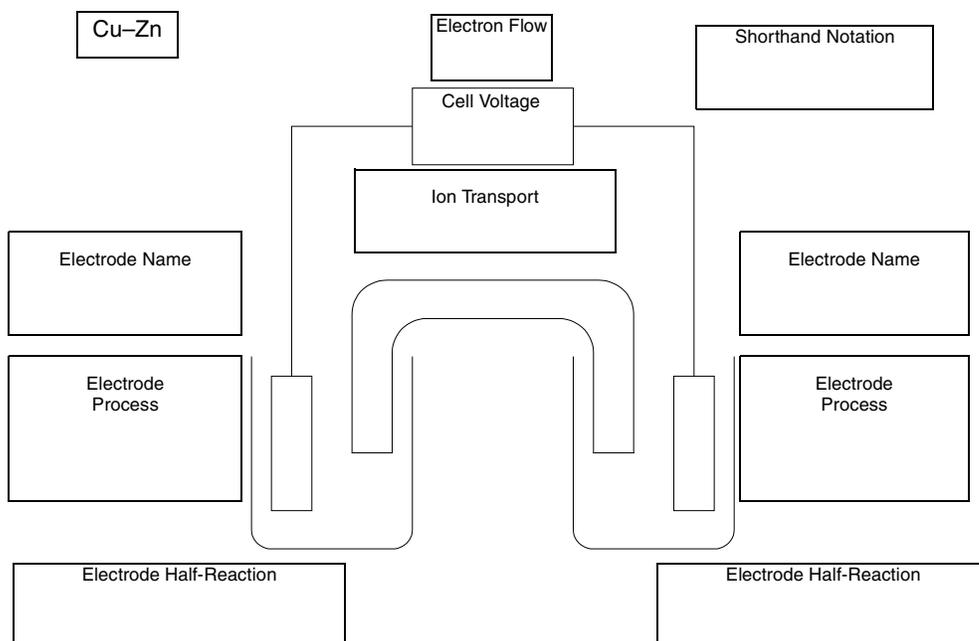
1. What product(s) was (were) formed at the anode in the electrolysis of potassium iodide solution? Explain citing specific evidence from your observations.
2. What product(s) was (were) formed at the cathode in the electrolysis of potassium iodide solution? Explain based on your observations.
3. Write the balance half-reaction that occurred at the anode.
4. Write the balanced half-reaction that occurred at the cathode.
5. Write the balanced chemical equation for the overall redox reaction in the electrolysis of aqueous potassium iodide. Hint: Remember to balance the electrons!

**In addition, answer the above five questions for the electrolysis of sodium chloride and electrolysis of copper (II) bromide trials.**

## Post-Lab Questions

1. Standard voltaic cells contain 1.0 M metal<sup>2+</sup> solutions, but we used 0.10 M metal<sup>2+</sup> solutions in Part A of the lab. Use the Nernst equation to prove that the cell potentials we measured in this lab,  $E_{\text{cell}}$ , are equal to the standard cell potential,  $E^{\circ}_{\text{cell}}$ . Show your work.

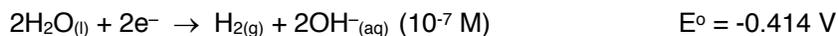
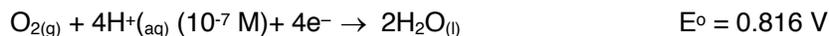
2. Using Figure 1 on page 2 as your guide, complete the diagram below for cells 2 **and** 6 (see table on pg 11). Your instructor will guide you through an example of this diagram so that you understand how to complete each “box”.



3. Rank the four metals used in Part A (Cu, Zn, Pb, and Ag) in terms of ease of oxidation from easiest to hardest.

4. Given three reasons why your measured voltages may differ from the theoretical voltage in Part A.

5. From the data below, determine what reaction will happen at the anode and what reaction will happen at the cathode for a 1.0 M  $\text{CdBr}_2$  solution. In addition, determine the minimum voltage required for the onset of the electrolysis reaction.



6. You only measured the mass gained by the coin in Part C. What do you think was the change in the mass of the copper electrode? Did it increase or decrease?

7. A voltaic cell is constructed with a  $\text{Cu}/\text{Cu}^{2+}$  half-cell and a  $\text{Ag}/\text{Ag}^+$  half-cell. If the observed potential of the cell,  $E_{\text{cell}}$ , is 0.40 V, calculate the ratio of  $[\text{Cu}^{2+}]$  to  $[\text{Ag}^+]^2$  present using the Nernst Equation.

## Pre-Lab Assignment

1. Given the reaction:  $\text{Al}_{(s)} + \text{Cu}^{2+}_{(aq)} \rightarrow \text{Al}^{3+}_{(aq)} + \text{Cu}_{(s)}$

- What is being oxidized?
- What is being reduced?
- Write the half-cell reaction for the oxidation (include electrons).
- What is the standard voltage for this half-cell reaction?
- Write the half-cell reaction for the reduction (include electrons).
- What is the standard voltage for this half-cell reaction?
- Write the balanced equation for the cell reaction.
- What is the standard cell voltage for the reaction?
- Is this reaction spontaneous under standard conditions?

2. The following data were measured using a nickel electrode as the standard:

$\text{Pt}^{2+}(\text{aq}) + 2e^- \rightarrow \text{Pt}(\text{s})$	+1.20
$\text{Ni}^{2+}(\text{aq}) + 2e^- \rightarrow \text{Ni}(\text{s})$	-0.28 V
$\text{Fe}^{2+}(\text{aq}) + 2e^- \rightarrow \text{Fe}(\text{s})$	-0.44 V
$\text{Al}^{3+}(\text{aq}) + 3e^- \rightarrow \text{Al}(\text{s})$	-1.66 V

- Which ion is most easily reduced?
- Which metal is most easily oxidized?
- The platinum and aluminum electrodes are connected in a battery.

1) Which is the anode? Cathode?

Anode: \_\_\_\_\_ Cathode: \_\_\_\_\_

2) Which is oxidized? Reduced?

Oxidized: \_\_\_\_\_ Reduced: \_\_\_\_\_

3) What will the battery voltage be?

4) Write a balanced net ionic equation for the reaction that takes place.

5) What happens to the mass of the anode electrode? Does it increase or decrease?

6) What happens to the mass of the cathode electrode? Does it increase or decrease?

3. Write the equation for the oxidation of water,

4. Write the equation for the reduction of water.

5. Calculate the mass of zinc that will be deposited if a current of 0.40 Amps is applied for 25 minutes to a  $\text{ZnSO}_4$  solution.