
Lab 8. Measurement of Voltaic Cell Potentials & Electrolytic Reduction of Cu^{2+}

Prelab Assignment

Before coming to lab:

- This exercise does *not* require a report in your lab notebook. Record your data, observations, calculations and analysis in the spaces provided on the report pages of this handout, [pages 5-8](#). Use pen for the recording of all experimental data and observations directly into tables 1 – 3 during your lab period.
- Read the lab thoroughly and answer the pre-lab questions that appear at the end of this lab exercise. Background information for this lab can be found in [Chapter 21](#) in your textbook ([Silberberg 5th ed.](#)).

Purpose

This lab activity has two parts. In **part 1** you will measure the potential of several voltaic cells and use your results to generate an activity series. Your measured cell potentials will be compared to the accepted values that can be found in the tables in your textbook. In **part 2** 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!

Introduction

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 calculator, contain oxidizing and reducing substances. As the electrons are transferred they are “tapped” 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. 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.

In order for a redox reaction to serve as a source of power, the reaction must be spontaneous. What if the reaction isn't spontaneous? In this case, we can use electricity to make the reaction “go”. 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. See [section 21.7](#) in your textbook for a more thorough discussion of electrolytic cells.

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.

In this lab activity you will measure the voltage of several voltaic cells. A typical **voltaic cell**, such as the one in figure 1 on the following 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 functions as the **anode**, where the **oxidation** reaction takes place

($M_{(s)} \rightarrow M^{n+}_{(aq)} + ne^{-}$), while the other functions as the *cathode*, where the *reduction* reaction occurs ($M^{n+}_{(aq)} + ne^{-} \rightarrow M_{(s)}$). Electrons flow from the anode to the cathode via the wire. The *salt bridge* allows migration of ions to prevent imbalance of charge from building up as electrons leave the anode and move to the cathode.

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* 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 (e.g. *CRC Handbook of Chemistry and Physics*) contain tables of standard reduction potentials, which show the values of E° for various reduction half-reactions of the type $M^{n+}_{(aq)} + ne^{-} \rightarrow M_{(s)}$. The E° in the tables were obtained by assigning a potential of 0.00 V to the reduction reaction of 1 M H^{+} : $2H^{+}_{(aq)} + 2e^{-} \rightarrow H_{2(g)}$ and measuring the potentials of other standard half-cells coupled to this *standard hydrogen electrode*.

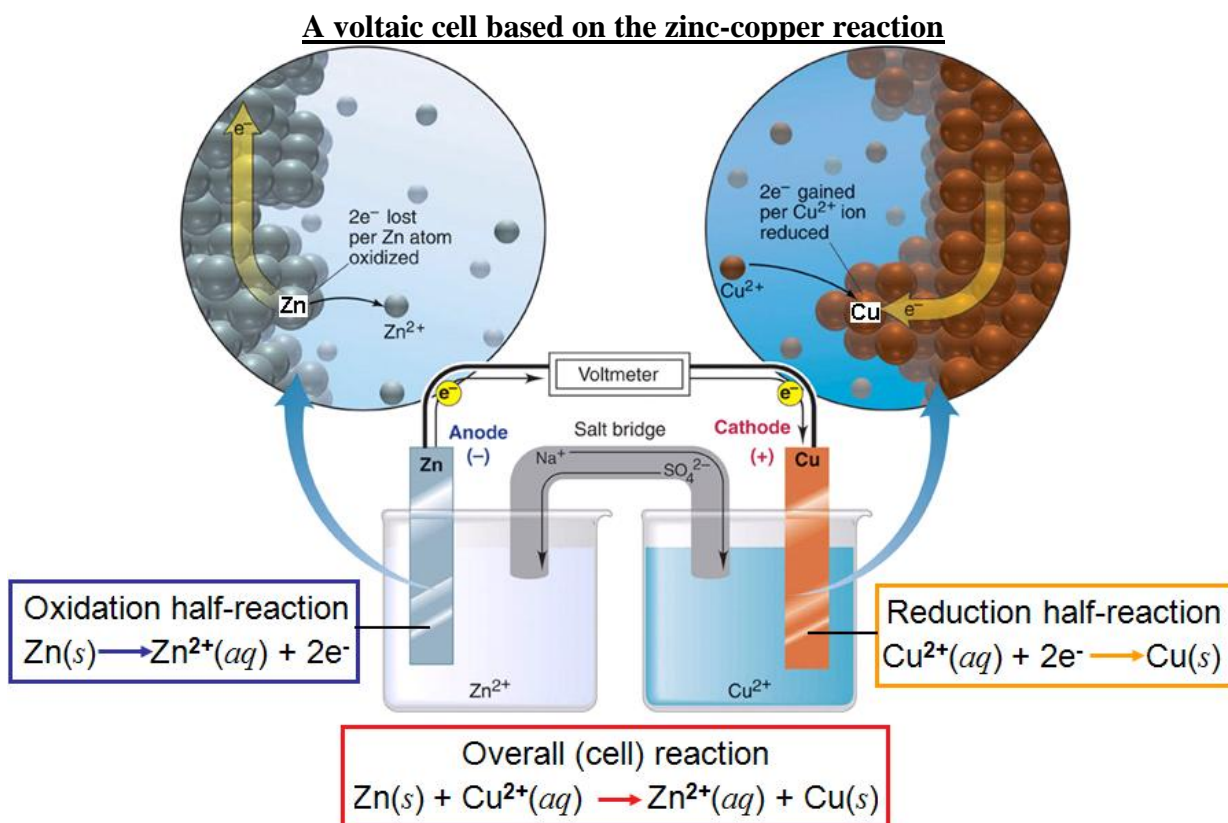


Figure 1. The *anode half-cell* consists of a zinc electrode submerged into a solution containing Zn^{2+} ions (e.g. aqueous $Zn(NO_3)_2$), while the *cathode half-cell* consists of a copper electrode submerged in a solution of Cu^{2+} ions (e.g. aqueous $Cu(NO_3)_2$). *Oxidation* occurs at the anode as each zinc atom loses two electrons to form aqueous 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 Cu^{2+} ions. Finally, *reduction* occurs on the surface of the cathode as the electrons react with the 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 (Na^{+}) and sulfate ions (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). Looking at reduction potentials, since the reduction potential of Zn^{2+} (-0.76 V) is more negative than that of Cu^{2+} (+ 0.34 V), zinc is more easily oxidized.

How to Predict the Potential of a Voltaic Cell

Calculation of the Cell Potential of Standard Voltaic Cells:

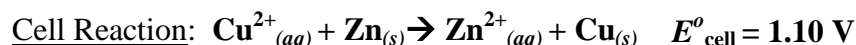
Whenever two standard half-cells are joined to create a voltaic cell as in figure 1, the one with the more negative E° will function as the anode since it is the metal that is most easily oxidized. The value for the standard potential of the cell, E°_{cell} , is the difference between the standard potential of the cathode and anode:

$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}$$

The use of standard reduction potentials to predict the potential of the voltaic cell in figure 1 is illustrated below.

Reduction Reactions	E°
$\text{Zn}^{2+}_{(aq)} + 2 e^{-} \rightarrow \text{Zn}_{(s)}$	-0.76 (more negative potential, therefore the anode)
$\text{Cu}^{2+}_{(aq)} + 2 e^{-} \rightarrow \text{Cu}_{(s)}$	+0.34 (more positive potential, therefore the cathode)

<u>Cathode:</u>	$\text{Cu}^{2+}_{(aq)} + 2 e^{-} \rightarrow \text{Cu}_{(s)}$	$E^{\circ}_{\text{cathode}} = +0.34 \text{ V}$
<u>Anode:</u>	$\text{Zn}_{(s)} \rightarrow \text{Zn}^{2+}_{(aq)} + 2 e^{-}$	$- E^{\circ}_{\text{anode}} = -(-0.76 \text{ V})$



Calculation of the Cell Potential of a Voltaic Cell NOT at Standard Conditions:

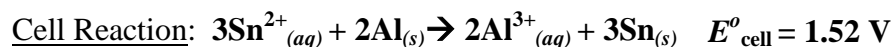
If the potential that is measured is *not a standard cell*, then the *Nernst equation* can be used to calculate the standard cell potential.

Nernst Equation	Where...
$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0591}{n} \log Q$	<ul style="list-style-type: none"> n = number of electrons transferred in the cell reaction Q_c = Reaction quotient E° = standard cell potential (1 M solutions at 298K and 1 atm) E = Cell potential at nonstandard conditions (both potentials are measured in volts)

Sample calculation of the standard cell potential for a voltaic cell made of a Sn electrode in 0.10 M Sn^{2+} in one half-cell and Al in Al^{3+} 0.10 M in the other.

Reduction Reactions	E°
$\text{Sn}^{2+}_{(aq)} + 2 e^{-} \rightarrow \text{Sn}_{(s)}$	-0.14 V (more positive potential, therefore the cathode)
$\text{Al}^{3+}_{(aq)} + 3 e^{-} \rightarrow \text{Al}_{(s)}$	-1.66 V (more negative potential, therefore the anode)

<u>Cathode:</u>	$(\text{Sn}^{2+}_{(aq)} + 2 e^{-} \rightarrow \text{Sn}_{(s)}) \times 3$	$E^{\circ}_{\text{cathode}} = -0.14 \text{ V}$
<u>Anode:</u>	$(\text{Al}_{(s)} \rightarrow \text{Al}^{3+}_{(aq)} + 3 e^{-}) \times 2$	$- E^{\circ}_{\text{anode}} = -(-1.66 \text{ V})$



$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0591}{n} \log Q = 1.52 \text{ V} - \frac{0.0591}{6} \log \frac{[\text{Al}^{3+}]^2}{[\text{Sn}^{2+}]^3}$$

$$= 1.52 \text{ V} - \frac{0.0591}{6} \log \frac{[0.10]^2}{[0.10]^3} = 1.52 \text{ V} - 0.0099 \text{ V}$$

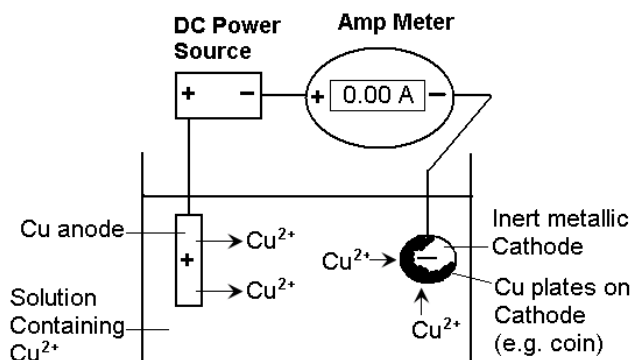
$E_{\text{cell}} = 1.51 \text{ V}$ (nearly the same as the standard cell potential!)

Procedure (work in teams of two)**Part 1—Measurement of Cell Potentials**

1. Obtain about 25 mL of 1 M KNO_3 from the lab cart. Cut up pieces of filter paper into small narrow strips. Soak the strips of filter paper in the KNO_3 . These will serve as your **salt bridges**.
2. Also obtain strips of Cu, Zn, Pb, Mg, and Sn from the lab cart. Clean the strips of metal by sanding and/or dipping in 3 M HCl.
3. Obtain a well plate from the lab cart. In one well put a piece of cleaned copper metal in contact with a solution of its ions, e.g., Cu metal in 0.10 M $\text{Cu}(\text{NO}_3)_2$. In another well put a piece of zinc metal in a solution of its ions e.g., Zn metal in 0.10 M $\text{Zn}(\text{NO}_3)_2$. Connect the cells using one of the salt bridges. Note: the solutions in all half-cells are 0.10 M M^{2+} .
4. Before making a measurement, **predict which metal will serve as the anode and which will serve as the cathode**—keep in mind the activity series! Then measure the potential of the cell by connecting the metals to the electrodes of a voltmeter, taking care to **keep the meter's metal contacts out of the solutions**. The black (negative) electrode should be on the anode, and the red (positive) electrode should be on the cathode. If you get a negative reading, try switching the electrodes. Record the measured cell potential in the appropriate cell in table 1 of the report sheet.
5. Measure the potential of all possible cell combinations using the method described above—see table 1 on the report sheet for all possible combinations. **Use a fresh salt bridge for each cell.**
6. 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 2—Electrolytic Reduction of Cu^{2+} to Copper on the Surface of a Coin

1. Place some copper sulfate solution in a 250 mL beaker such that you will be able to completely submerge a nickel or other coin in the solution when the coin is placed on end.
2. Determine the mass of the coin and record in table 3.
3. Set up the **electrolytic cell** as in the figure below. Use a strip of copper metal as one electrode, and a coin as the other. There is no need for a salt bridge in this experiment. Don't submerge the coin in the solution until you are ready to begin. **Keep the alligator clips out of the solution!**
4. Use a watch or a stopwatch to record the time from the moment the power is turned on. After turning the power on, set it to a moderate to low value—**if the voltage is too high you will see bubbles at each electrode indicating the electrolysis of water**. Allow the current to stabilize and record the current (in amperes, A) in table 3. Let the cell run for 5 minutes or so, but record the exact time in seconds in table 3.



5. Remove the coin from the solution, pat it dry and record the mass in table 3.

Lab 8 Report Sheet
Cell Potentials & Electrolytic Reduction of Cu^{2+}

Name _____
Team No. ___ Date _____ Section _____

Calculations and Analysis

Part 1—Measurement of Cell Potentials

1. Using Figure 1 on page 2 as your guide, make a diagram for cells 2, 6 and 9 (see table 1).

Cell 2: Cu/Pb

Cell 6: Zn/Mg

Cell 9: Pb/Sn

2. Identify each of the following for all cells observed in this lab activity and record your responses in table 1: anode, cathode, substance oxidized, substance reduced, oxidizing agent and reducing agent.

Table 1. Summary table for Part 1—Measurement of Cell Potentials. *Enter all data using a pen!*

Cell No.	Electrode #1	Electrode #2	Observed Standard Cell Potential, E°_{cell} (V)	Expected Standard Cell Potential E°_{cell} (V) (show your work in each space, below)	Anode	Cathode	Substance Oxidized	Substance Reduced	Oxidizing Agent	Reducing Agent
1	Cu	Zn								
2	Cu	Pb								
3	Cu	Mg								
4	Cu	Sn								
5	Zn	Pb								
6	Zn	Mg								
7	Zn	Sn								
8	Pb	Mg								
9	Pb	Sn								
10	Mg	Sn								

3. Standard voltaic cells contain 1 M M^{2+} solutions, but we used 0.10 M M^{2+} solutions in all of our half-cells. Use the Nernst equation to prove that the cell potentials we measured in this lab, E_{cell} , are equal to the standard cell potential, E°_{cell} . Show your work in the space below.

4. Calculate the expected cell potential for each voltaic cell. Record your calculated values and show your work in the table 1.

5. Assume that the half-cell potential for Cu^{2+}/Cu is 0.34 V. Based on this value, and your experimental cell potentials, calculate the **reduction potentials** for all half-cells studied— Recall that $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$. Record your calculated values and show your work in the table 1.
- In the space below table 2 rank these potentials from highest (i.e. most positive) to lowest (most negative) and compare them to the values in your text. As a part of your analysis discuss the agreement between your experimental results and the actual values in terms of both relative ranking, and measured values.

Table 2. Experimental and Theoretic Standard Half-Cell Potentials. *Enter all data using a pen!*

Half-Cell Reduction Reaction	Experimental Value of E° (V) (show your work in each space below)	Literature Value of E° (V)	Percent Error
$\text{Cu}^{2+}_{(aq)} + 2 e^- \rightarrow \text{Cu}_{(s)}$		0.34	
$\text{Zn}^{2+}_{(aq)} + 2 e^- \rightarrow \text{Zn}_{(s)}$			
$\text{Pb}^{2+}_{(aq)} + 2 e^- \rightarrow \text{Pb}_{(s)}$			
$\text{Mg}^{2+}_{(aq)} + 2 e^- \rightarrow \text{Mg}_{(s)}$			
$\text{Sn}^{2+}_{(aq)} + 2 e^- \rightarrow \text{Sn}_{(s)}$			

Ranking of Reduction Potentials:

Analysis:

Calculations and Analysis

Part 2—Electrolytic Reduction of Cu^{2+} to Copper on the Surface of a Coin
Background Info:

The copper plating is accomplished by submerging both a coin and a strip of copper in a solution of copper sulfate. At one electrode copper is oxidized ($\text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{e}^-$) and at the other it is reduced ($\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$). Thus there is no net reaction. 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.

Table 3. Data for the Electrolytic Reduction of Cu^{2+} . *Enter all data using a pen!*

Mass of Coin after electroplating (g)	
Mass of Coin before electroplating (g)	
Mass of Copper plated on Coin (g)	
Expected mass of Copper plated on Coin (g)	
Time of Electrolysis (seconds)	
Average Current (A)	

- Determine the mass of copper plated out onto the coin and record in table 3.
- Using the average current in amperes and the time in seconds in which the electrolytic cell was in operation, calculate the expected mass of copper produced—[see pages 965-966](#) for an explanation and sample problems. The electrolytic set-up we used is fairly crude, and there will be some inefficiency in the cell, so don't expect perfect agreement! Show your work in the space below.

Lab 8 Prelab Questions

Cell Potentials & Electrolytic Reduction of Cu^{2+}

Name _____

Team No. ___ Date _____ Section _____

Instructions: Complete the following questions and hand in at the start of your lab period or when instructed by your instructor. Show your work with units and correct significant figures for all questions that involve a calculation.

1. A voltaic cell is constructed coupling a Cu/Cu^{2+} half-cell with a Ag/Ag^+ half-cell. The concentrations of Cu^{2+} and Ag^+ are each 0.010 M .
 - a. Which metal is the anode? How do you know?

 - b. At which electrode is reduction occurring? How do you know?

 - c. Write the half-reaction for the reduction reaction.

 - d. Write the half-reaction for the oxidation reaction.

 - e. Write the balanced chemical equation for the overall reaction for the voltaic cell.

 - f. If the observed potential of the cell, E_{cell} , is 0.40 V , use the Nernst equation to calculate the standard cell potential, E°_{cell} .

 - g. Calculate the standard reduction potential, E° , for the reduction of Ag^+ if the standard reduction potential, E° , for Cu^{+2}/Cu is 0.34 V .