

ELECTROCHEMICAL CELLS LAB

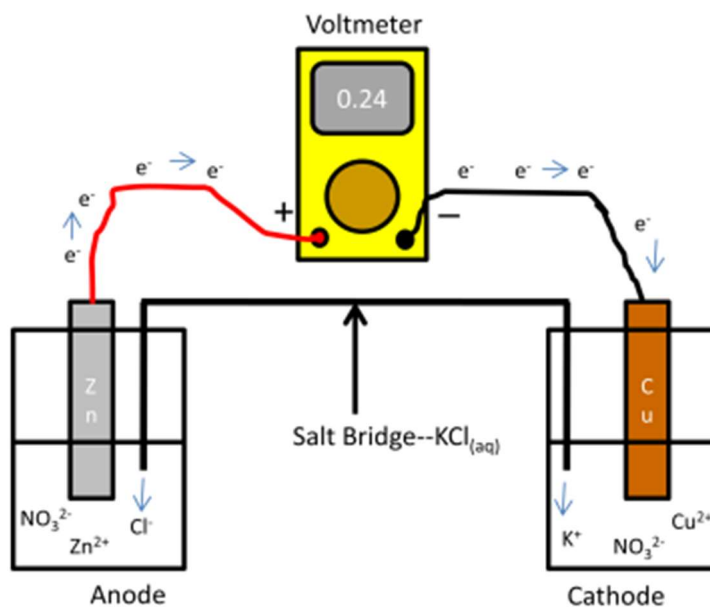
Purpose: The purpose of this lab is to demonstrate the ability of chemistry to make electric current using oxidation/reduction (REDOX) reactions, and to measure the electric current that can be harnessed via these reactions.

Background: In class, we have learned that a REDOX reaction is actually two $\frac{1}{2}$ reactions, a reduction $\frac{1}{2}$ reaction and an oxidation $\frac{1}{2}$ reaction. We have also learned that there are three definitions that define these reactions, gaining or losing oxygen, hydrogen, or electrons. In this lab, we will be looking at the definition depicting the movement of electrons. That is, oxidation is a loss of electrons and reduction is a gain of electrons.

We tend to think of electrons as being centered around a nucleus, but in fact, we can move electrons as a current along conductive materials (such as metals, graphite, and ionic salt solutions). This electric current can be used to power anything from cars to light bulbs. Because a redox reaction can involve the movement of electrons, we could use these reactions to produce an electric current.

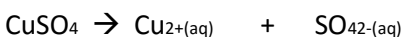
If we poured our oxidation materials into the same flask as our reduction materials, we will get a quick redox reaction, but we would be unable to harness the electrons in the form of a current. If we pour the oxidation reaction chemicals into one flask and the reduction reaction chemicals into another flask, they have no means of reacting with each other. If we could connect the two reactions using a conductive material so that the electron transfer could occur, without mixing the chemicals themselves, we could harness the electron current moving along the conductive substance. This is the case in an electrochemical cell (Figure 1). We can use an instrument called a voltmeter to measure the voltage, or electrochemical current that occurs with the transfer of electrons.

Figure 1: An electrochemical cell

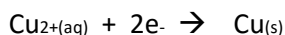


The components of an electrochemical cell are as follows:

- Anode—the site where an oxidation reaction happens, or a loss of electrons. The side with the anode will be gaining a positive charge from the surplus of protons.
- Cathode—the site where a reduction reaction happens, or a gain of electrons. The side with the cathode will be gaining a negative charge from the excess electrons.
- Salt bridge—a strip soaked with a non-reactive ionic solution. We will be using KCl in this experiment. The salt bridge will deposit chloride (Cl⁻) ions at the anode to neutralize the positive charge that forms, and it will deposit potassium (K⁺) ions at the cathode to neutralize the negative charge that forms.
- Electrode—A material through which electrons can flow. A conductive material (metals, graphite, etc...) which helps transport electrons in a redox reaction.
- Voltmeter—an instrument that measures electric current. The red cable going into the voltmeter (+) corresponds to the anode, and the black cable (-) corresponds to the cathode. When the voltmeter readings are positive, then the cables are aligned with the correct solutions. When the reading is negative, then the red cable is reading the cathode and the black cable is reading the anode—this requires you to switch the cables. The units measured are volts, V, which are a unit of electric potential.
- Salt solution—the cell requires the anode and cathode to be in salt solutions that will contain either the reactants or the products (whichever is a salt) of the reaction dissolved in water. The salts are **dissociated** in water into ions, which are good conductors of electric current. For example, the dissociation of copper sulfate in water gives copper II ions and sulfate ions:

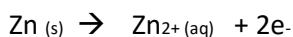


In Figure 1 above, the electrons are shown moving from the zinc anode to the copper cathode. The electrons will flow from the zinc to the copper solid and then into the solution, where they will attach to the copper II ion (Cu²⁺). This will give the following reaction, called a half-reaction, because it only shows the reduction half of the redox reaction:



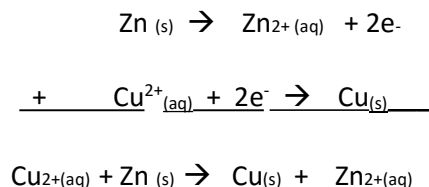
This is the half-reaction for the reduction of copper II into copper solid. Through the course of the reaction, the copper in solution will start depositing itself onto the solid copper cathode as crystals. Because negative charge is building up in the copper side of the electrochemical cell, potassium ions from the salt bridge will be flowing into the solution to balance the charge.

At the anode, where the solid zinc electrode is located, the lost electrons result in the ionization of zinc. When zinc loses electrons, it becomes a zinc cation, Zn²⁺. The half-reaction for the oxidation of zinc is:



The concentration of zinc ions in the solution will be steadily increasing, leading to the “disappearance” of zinc solid on the anode. The buildup of positive charge in the solution will be balanced by chloride ions from the salt bridge flowing into the solution.

The two half-reactions can be combined to give the overall reaction equation. The halfreactions are lined up and “added” together to give the net ionic equation as follows:



Note the cancellation of electrons from either side of the net ionic equation. When an overall reaction contains the same material on either side of the reaction arrow, it can be discarded in the net ionic equation. Similarly, the nitrate ions are not shown because they are unchanged in the reaction. Because they are unchanged they would appear both sides of the reaction, and would hence, cancel out.

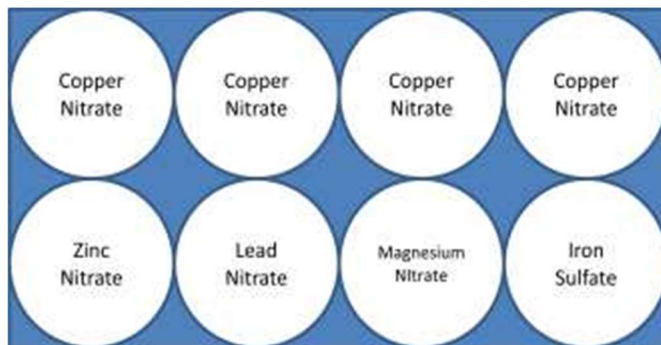
Materials:

- One 24-well plate
- Filter paper strips
- Pairs of electrodes (Cu/Zn, Cu/Mg, Cu/Pb, Cu/Fe)
- 0.1 M solutions of Cu(NO₃)₂, Zn(NO₃)₂, Mg(NO₃)₂, FeSO₄, Pb(NO₃)₂, KCl
- Sandpaper/steel wool
- Voltmeter with two wires with alligator clips
- Watch glass
- Disposable pipettes

Procedure:

***NOTE: THESE CHEMICALS ARE VERY REACTIVE WITH EACH OTHER. DO NOT CROSS CONTAMINATE YOUR EQUIPMENT! IT IS IMPORTANT TO KEEP ALL SOLUTIONS FROM TOUCHING EACH OTHER.

- I. You need to make up a salt bridge for your electrochemical cells. Obtain 4 pieces of filter paper. Place these on a watch glass and cover with the KCl solution from a pipette.
- II. Polish all electrode wires with a piece of steel wool/sandpaper.
- III. Carefully add 20 drops of the solutions from the pipettes into the appropriate 24-well plate as shown below. You will only be using 8 of your 24 wells.



- V. Join your electrochemical cell pairs by placing a salt bridge strip across each pair. This means each strip will be dipped into two different wells (copper nitrate and zinc nitrate, copper nitrate and lead nitrate, etc.). Make sure the filter paper is in both of the solutions so your ions can flow into each well.
- VI. Take the Cu/Zn electrode pair and place in the corresponding wells (copper goes in copper nitrate, zinc goes in zinc nitrate). Attach the wires with alligator clips to the voltmeter (red (+) goes with red and black (-) goes with black). Attach one clip to the copper electrode and the other clip to the zinc electrode. Turn on the voltmeter to the DCV setting (direct current setting). If the reading on the voltmeter is negative, then you will need to switch the clips. Record the positive reading you obtain. Refer to the section above on voltmeters to determine which cell is being oxidized and which is being reduced to fill in the rest of the data sheet.
- VII. Using the appropriate electrode pair, measure the voltages for the other 3 solutions and record the current created in volts. Fill in the data sheet for these electrochemical cells.

References:

1. Blackburn, J. R.; Joesten, M. D.; Langford, P. B.; Craig, J. C.; Electrochemical Cells. In *Laboratory Manual to Accompany World of Chemistry*, 2nd ed.; Thomas Learning: Toronto, Canada, 1996; pp 247-253.
2. Harris, D. C. Fundamentals of Electrochemistry. In *Quantitative Chemical Analysis*, 6th ed.; W.H. Freeman: New York, 2003; pp 283-305.
3. Hill, W. H.; Kolb, D. K. Oxidation and Reduction. In *Chemistry for Changing Times*, 11th ed.; Pearson Prentice Hall: Upper Saddle River, NJ, 2007; pp 221-230.

ELECTROCHEMICAL CELLS LAB DATA SHEET

Name: _____

Partner: _____

1. Copper reaction:

- a. Draw the copper half reaction going from copper solid to copper II ion:

2. Copper-zinc redox reaction:

- a. Voltage (in volts, V) _____

- b. Zinc half-reaction from solid to Zn^{2+} :

- c. Net ionic equation:

- d. Cathode: _____ Anode: _____

3. Copper-magnesium redox:

- a. Voltage (V) _____

- b. Magnesium half-reaction from solid to Mg^{2+} :
- c. Net ionic equation:
- d. Cathode: _____ Anode: _____
4. Copper-lead redox:
- a. Voltage (V) _____
- b. Lead half-reaction from solid to Pb^{2+} :
- c. Net ionic equation:
- d. Cathode: _____ Anode: _____
5. Copper-iron redox:
- a. Voltage _____
- b. Iron half-reaction from solid to Fe^{2+}
- c. Net ionic equation:
- d. Cathode: _____ Anode: _____

ELECTROCHEMICAL CELLS LAB POST-LAB QUESTIONS:

Name: _____ **Partner:** _____

6. Using Figure 1 as a guide, draw a detailed copper-lead electrochemical cell. Show which ions from the salt bridge are entering each solution and the direction of electron flow.

7. What was the purpose of the salt soaked filter paper:
8. The activity of a metal is defined as its ability to be oxidized, or lose electrons. This can be measured by looking at the voltage in a redox reaction, where a greater voltage measured means a greater difference in the activity of the chemicals. The higher the number, the easier the anode is oxidized in comparison to the cathode material. List the 5 metals in this experiment below based off increasing activity (what is oxidized easiest). We can compare all of our results this way because we used a central copper electrode for each trial:

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9. Using the results in 8 above, show the balanced full reaction between the two metals that would give the largest voltage reading. Also, find the net-ionic equation.

a. Full reaction:

b. Net-ionic equation:

c. What chemical is being oxidized and what is being reduced in this case:

Oxidized: _____

Reduced: _____

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ELECTROCHEMICAL CELLS PRE-LAB QUESTIONS

Name: _____

After reading the introduction to this lab, answer the following questions to prepare yourself for the experiment. You may want to refer to lecture materials:

1. What are the chemical formulas for the following chemicals:
 - a. Copper nitrate _____
 - b. Zinc nitrate _____
 - c. Magnesium nitrate _____
 - d. Lead nitrate _____
 - e. Iron sulfate _____

2. Show the dissociation of the above chemicals into ions (what happens in water):
 - a. Copper nitrate

 - b. Zinc nitrate

 - c. Magnesium nitrate

 - d. Lead nitrate

 - e. Potassium chloride

 - f. Iron sulfate

3. Show the half-reactions for each metal ion becoming the solid metal:
 - a. Copper nitrate

 - b. Zinc nitrate

 - c. Magnesium nitrate

 - d. Lead nitrate

e. Iron sulfate