

You'll need to write a **title** and a **purpose** for this experiment; read over the experiment first to figure out what to call it and what it's for.

Notice that the **chemical and mathematical equations** and large parts of the **procedure** are missing. You will need to supply these.

As always, you will need to use the hints to develop your **data, analysis** and **conclusion** sections. There is *no abstract* necessary for this lab writeup.

Photocopy the appropriate parts of your lab notebook; please make sure your name is legible.

Chemistry 163

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## Lab 8: Write a title describing, succinctly, what this lab was about

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### Purpose:

**Introduction:** Electrochemistry is the most obvious manifestation of redox chemistry – in these experiments you will be measuring the “pressure” behind those electrons (**voltage**) using a LabPro multimeter attachment. The attachment may be connected to a LabPro device or a laptop computer for a digital display and has two “leads” (a red one, which is usually hooked to the positive or hot side of the device, and a black one, negative or ground).

**Skills:**

- Making voltage measurements
- Making a galvanic cell from solutions and electrodes
- Making a solution cell

Chemical equations: Fill in as needed.

Mathematical equation: Nernst equation

Chemicals needed:

- 1.0 M copper (II) sulfate
- 0.1 M copper (II) sulfate
- 0.1 M iron (II) sulfate
- 0.1 M lead (II) nitrate
- 0.1 M zinc (II) nitrate
- potassium nitrate solution (concentration not critical)

- copper metal strip
- iron metal strip
- lead metal strip
- zinc metal strip

Equipment needed:

- 24-well plate
- 7 filter paper strips (to be used as salt bridges)
- disposable pipet
- multimeter attachment
- LabPro device or Laptop computer
- thermometer

**Procedure**– everything in **bold-face** you have to complete!

1. Obtain about 5 mL of each of the 0.1 M solutions in separate *clearly-labeled* test tubes. Don't forget to record concentrations and room temperature.
2. You will construct cells using wells in the well plate instead of beakers. Draw a schematic of the well plate so you can tell what is in a well. Do not place metal electrodes in the wells.
3. The appropriate metal electrodes will be clipped to the probes of the multimeter attachment and dipped into the appropriate solutions when a cell's voltage is measured.

**(Here in your notebook should be a drawing of one galvanic cell setup. Do not forget to label the apparatus parts)**

4. Measure and record the *potential* of the cell, indicating which metal is attached to the red probe and which to the black.
5. Repeat the step 4 for every combination of half-cells. There are six.
6. To prepare a concentration cell based on Cu, we first need to prepare solutions to use. You will perform a **serial dilution**, a method of making very dilute solutions. Starting with Solution A, 1.0 M  $\text{CuSO}_4$ , transfer 1.0 mL into a 100.00 mL volumetric flask and dilute to the mark with deionized water. Transfer the solution to another container. That is Solution B, 0.01 M  $\text{CuSO}_4$ . Rinse the pipet with few small portions of Solution B, holding the rinsing in a container until you have a chance to place them in the waste container. Prepare Solution C, 0.0001 M  $\text{CuSO}_4$

from Solution B in a similar fashion, and Solution D, 0.000001 M  $\text{CuSO}_4$  from Solution C.

7. Prepare your concentration cell using Solution A and Solution D. Record its voltage. In this case, both electrodes are copper. To indicate the direction of measure, note which solution the red probe was in contact with and which the black probe was in contact with.

**Waste disposal:** Place all waste (including metals) in the “aqueous waste” container in the hood.

### **Data section:**

Concentrations, temperature

List of appropriate ideal half-cell potentials (from the textbook appendix)

Table of cell reactions and measured potentials

### **Analysis/calculations**

1. Describe how the conditions of the experiment are non-ideal.
2. Calculate  $E^\circ$  (units!) for each of the cells. Include a sample calculation for one of the cells.
3. Calculate the theoretical non-ideal potential  $E$  (units!) for each of the cells using the Nernst equation. Include a sample calculation for one of the cells.
4. Compare the measured voltage for each of the cells versus its theoretical non-ideal potential and calculate a percent difference (include the sign) between them.
5. Summarize these results in a table with columns for cell,  $E$ , theoretical  $E$ , measured  $E$ , and percent difference.

### **Conclusion**

Address the following points:

- How come some of the time you got a **negative** voltage? Qualitatively use

the **free-energy expression** and discuss the spontaneity or non-spontaneity of the reaction.

- Was the error between the theoretical non-ideal potential you calculated for all of the cells and what you actually measured a **random** or a **systematic** one? How can you tell? If systematic, was it always high or always low? What might be the cause of this discrepancy?