

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.

You may cut and paste the "**Introduction**" through "**Skills**" section. Notice that the **chemical and mathematical equations** and large parts of the **procedure** are missing.

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. This lab is **due Wednesday, May 23, in class**.

Chemistry 160

Lab 5: Write a title describing, succinctly, what this lab was about

Purpose:

part one:

part two: Write a description of how you got a battery-powered clock to run

Introduction: Electrochemistry is the most obvious manifestation of redox chemistry – in these experiments you will be measuring the flow of electrons (**amperage**) and the "pressure" behind those electrons (**voltage**) using a multimeter. The multimeter is a digital display unit with controls and 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).

To measure amperage, the two leads have to be hooked up **in series** within the circuit you are measuring; in other words, all of the electrons at a given point in a circuit must flow through the multimeter to get an accurate measurement.

To measure voltage, the two leads have to be hooked up **in parallel** to the circuit you are measuring; in other words, virtually none of the electrons at the point in the circuit you are measuring should flow through the multimeter (a good example of this is putting the leads in an 2-prong wall outlet – no electrons flow through the multimeter, since it has a high resistance, which makes it perfect to measure voltage).

Skills:

- Using a multimeter to make voltage and amperage measurements
- Making a galvanic cell from solutions and electrodes
- Making a galvanic cell do work

Chemical equations: Fill in as needed.

Mathematical equation: Nernst equation

Chemicals needed:

- 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:

- steel wool (to remove oxide from metals)
- several clean test tubes that hold at least 5 mL of solution
- filter paper (cut into strips long enough to bridge two test tubes with 5 mL of solution in them apiece)
- disposable pipet
- multimeter
- thermometer
- battery-powered clock

Procedure (part one) – everything in **bold-face** you have to complete!

1. Obtain about 5 mL of each of the solutions in separate *clearly-labeled* test tubes. Don't forget to record concentrations and temperature.
2. Use the steel wool to remove the oxide from the more active metal strips, then place each metal strip in its corresponding ionic solution. Notice that nothing should be going into the potassium nitrate solution.
3. Note that each of the test tubes in step 2 is now a half-cell. The galvanic cell is made by arranging the half-cells and other equipment in the drawing shown below:

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

4. Measure and record the *potential* of the cell by placing the red lead of the multimeter on the.....**(I couldn't read this on the original of the lab description – you need to complete the writing of this instruction – it might help to know whether you need the multimeter to be in series or in parallel with the cell circuit.)**
5. Measure and record the *current* of the cell by placing the red lead of the multimeter on the.....**(I couldn't read this on the original of the lab description – you need to complete the writing of this instruction – it might help to know whether you need the multimeter to be in series or in parallel with the cell circuit.)**
6. Repeat the two steps above for every combination of half-cells.
7. Pick one of the solutions (**name it here**) from the set above (start with fresh solution) and dilute it so that it is *half* of its original concentration; repeat the voltage and current measurements for some cells using this solution.

Procedure (part two)

(this is the gist of the second purpose; describe in some detail, what your initial theoretical idea for what half-cells would be needed to run a battery-powered clock, then describe what your final setup that actually did run the clock turn out to be).

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 half-cell names, potentials and current (and the resistance setting it was at)

Analysis/calculations

1. Describe how the conditions of the experiment are non-ideal.
2. Calculate E° (units!) for each of the cells (you will need your textbook appendix for this). Include a sample calculation for one of the cells. You should make a table out of it, or at least start one, and call this column “ideal”.
3. Calculate the theoretical non-ideal potential (units!) for each of the cells using the Nernst equation. Include a sample calculation for one of the cells. You should make these numbers another column in the table started above.
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. This should be a couple more columns in your table.

Conclusion

Address the following points:

- Did you get the galvanic cells to generate **electrical current**? So 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.
- What did you expect would happen to the cell potential/current when the concentration of solution was **halved**? What actually did happen? Was your hypothesis correct? If not, what might have happened?
- 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?

- Is the minimum voltage or the minimum amperage more critical to know in making the clock run? (In other words, will a clock run with no current but lots of potential, or with no potential but lots of current, or is it some other scenario?) From your observations, what was the minimum voltage and minimum amperage necessary to run your clock?