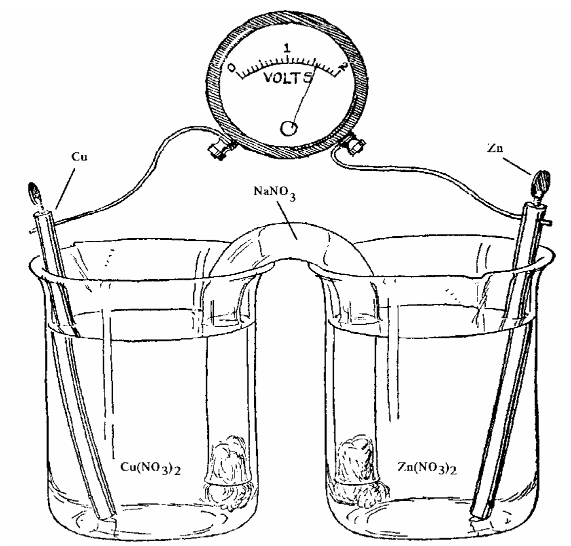
We all live under the same sky, but we don't all have the same horizon. --Konrad Adenauer

Galvanic Cells

One of the characteristics of redox reactions is the transfer of electrons (similar to the exchange of protons in acid/base reactions). If the electrons can be directed through a piece of metal during this transfer, then an *electric current* is created. We generally call such a set-up a *cell*. A group of cells hooked together is a battery.

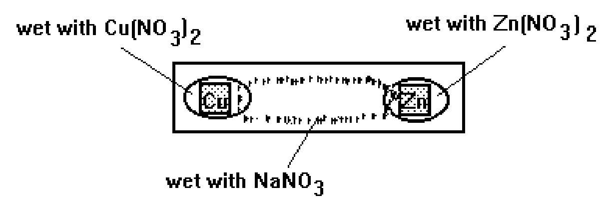
Each substance has a different tendency to exchange electrons with another substance. To a large extent this tendency is indicated by the position of a substance on the activity series. The actual phenomenon is more complex than that, but for our purposes here, the object is to place the metals in this experiment into a series from *most easily oxidized* to *most difficult to oxidize*.

A classic set-up for a galvanic or voltaic cell is shown below:



This *looks* very nice but in practice gives rather poor results and uses a lot of solutions. You can achieve far better results and use up less material by placing two pieces of metal at opposite ends of a strip of filter paper that has been wet with the appropriate solutions:

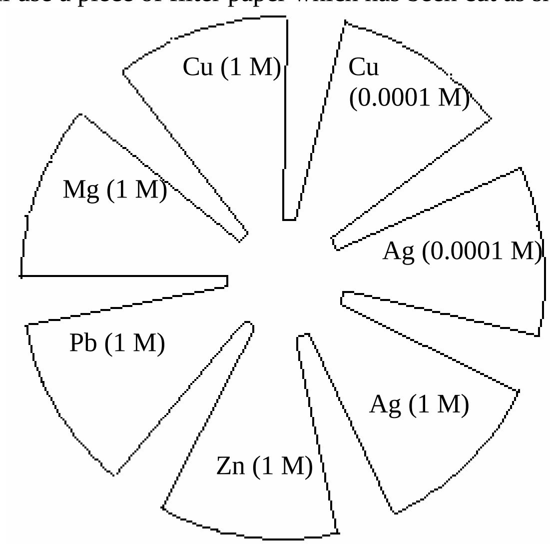
Micro-cell design adapted from: <u>Establishing a Table of Reduction Potentials: Micro-voltaic Cells</u>, Dan Holmquist, Jack Randall, Donald Volz, *Chemistry with CBL*, Vernier Software, 1995



[the $NaNO_3$ in the middle of the strip acts like the "salt bridge" or U-tube in the traditional set-up, allowing ions to migrate slowly across the paper, maintaining the electrical connection]

If the wires from a voltmeter are touched to the metal pieces, the resulting voltage reading represents the potential of electrons to flow from one metal to the other. Further, the polarity of the metals (which one is connected to the + or - terminals of the voltmeter) indicates their relative positions in a series from easily oxidized [-] (the **top** of the activity series) to difficult to oxidize [+] (the **bottom** of the series).

In addition to determining the relative strengths of the metals as oxidizing agents, you can also investigate the effect on the voltage of changing the solution concentrations for the cell combination of Cu with Ag. To make all of this easier (and quicker) you will use a piece of filter paper which has been cut as shown below:



By placing metal pieces on drops of their ionic solutions and "bridging" the gap in the center with $NaNO_3$ you can easily set up a number of different reactions to investigate.

Preparing to experiment

You will be provided with the following materials:

1. pre-cut filter paper as shown on the previous page
2. pieces of the following metals: Cu, Zn, Ag, Pb, Mg
3. 1 M solutions of Cu^{2+}, Zn^{2+}, Ag^{+}, Pb^{2+}, Mg^{2+}
4. 0.0001 M solutions of Cu^{2+} and Ag^{+}
5. 1 M solution of $NaNO_3$
6. CBL voltage cable

Design an experiment to measure the voltage and determine the polarity in cell produced from all possible combinations of the metal supplied. [compare the 0.0001 M Cu^{2+} *only* with the 1 M Ag^{+} and the 0.0001 M Ag^{+} *only* with the 1 M Cu^{2+}]

BE SURE TO BRING YOUR TI-83/84 CALCULATOR TO CLASS FOR THIS EXPERIMENT. YOU WILL ALSO NEED A COPY OF THE HCHEM.83G FILES IN YOUR CALCULATOR MEMORY.

Technique

1. Measuring voltage with the CBL

For this experiment you can set up to read voltage continuously. Note that the RED connector on the voltage cable is + and the BLACK is -. A peculiar property of the voltage interface is that it indicates a small voltage even when nothing is attached to the wires!! Just ignore that. When you touch the leads to an actual source of voltage, you will get the correct reading <u>if</u> the negative (BLACK) connector is touching the metal at which oxidation occurs. Otherwise, you will get a very small (+ or -) voltage. Thus it is important to try reversing the leads if you get little response. None of the cells in this experiment should yield voltages in the mV range, so if yours do, try adding more solution around the metal or remoistening the "bridge" with $NaNO_3$--or both.

2. Preparation of metal samples for best results

Use the synthetic steel wool (<u>DRY</u>) to clean both sides of each metal piece. Note that you have 2 pieces of Cu and 2 pieces of silver. Prepare all of the metal pieces BEFORE you apply the solutions or the spots might be dry by the time you are ready.

3. Cleanup

The paper can be discarded and the plastic sheet should be rinsed and dried. Since the Ag^{+} solution will tend to spread out over the paper, we strongly recommend that you handle the used pieces of metal and the paper with tweezers to avoid stains from the silver solution. The metal pieces should be rinsed with a squirt of distilled water and <u>***dried***</u>.

Analysis

1. Make a table listing your metal combinations, the correct polarity for each metal in that combination and the recorded voltage [this may already be part of your data, in which case you should skip this and go on to #2].

2. Examine your results for the 1 M solutions and arrange the metals in a list from most easily oxidized (-) to most easily reduced (+). Compare this list with the activity series in your book. Do they agree?

3. What is the effect on the voltage of changing the concentrations to 0.0001 for half of the cell? Do your results suggest the *direction* in which the Cu/Ag reaction is spontaneous? (see below)

$$Cu(s) + 2\,Ag^{+}(aq) \leftarrow\ ?\ \rightarrow Cu^{2+}(aq) + 2\,Ag(s)$$

4. Refer to the table of standard reduction potentials in your text book (p. 629). Locate the $Cu^{2+}|Cu$ *half reaction* (+0.34 v). Also note the $Ag^{+}|Ag$ half reaction at +0.80 v [note, these are the voltages for 1 M concentrations only]. How would you combine these two voltages to get approximately the voltage you measured for this combination? After you figure *that* out, determine the rest of the "expected" voltages that you measured in the lab. How do your results compare?

5. Each of these combinations resulted in a *spontaneous* reaction (positive voltage). The metal that was oxidized (higher in the list you made earlier) is losing electrons and is called the *anode*. The other metal gains the electrons (becomes reduced) and is called the *cathode*. Note the relative positions of anodes and cathodes in the table on p. 615.

Finally, what do the actual reactions look like? You've seen one possible example above. Write the reactions that occurred in the correct direction for each cell combination.