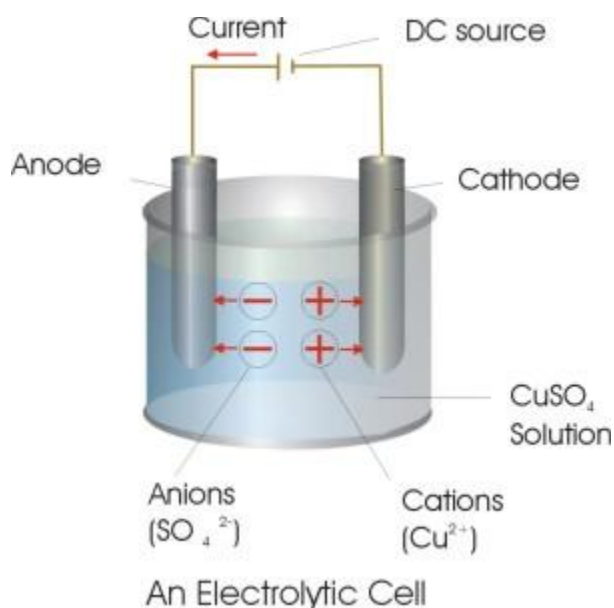


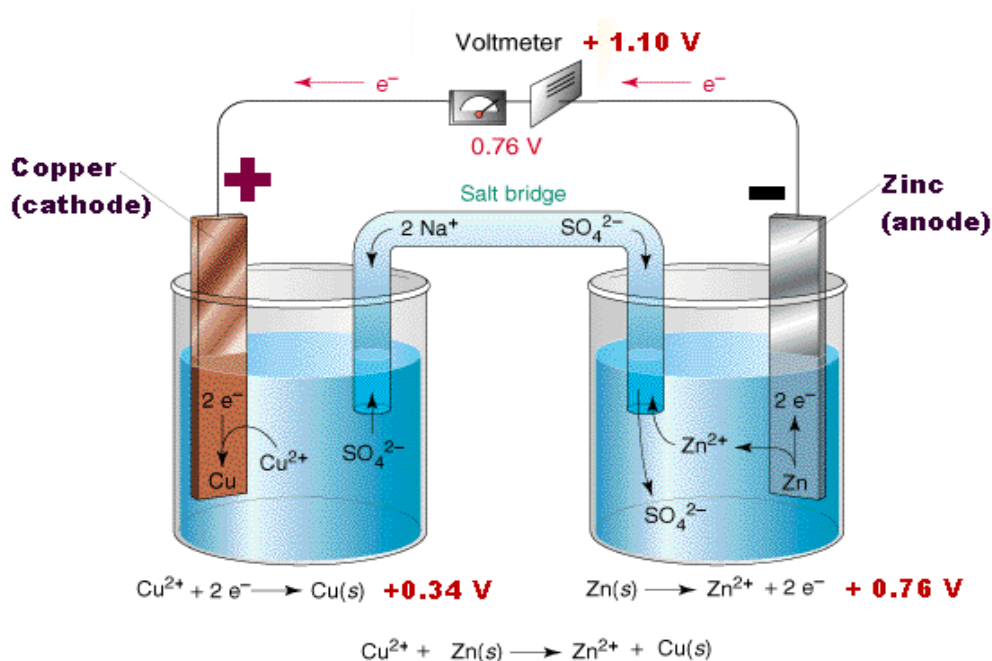
## Introduction

In this lab you will be able to observe changes when an electric current passes through a solution. In all electrochemical cells, electrons move through the wire (the external circuit) to the cathode, where a reduction reaction occurs, thus consuming electrons. The reaction results in charge transfer between the electrode and the electrolyte solution inside the cell. Within the cell, negative ions (anions) continue the negative current flow from the cathode toward the anode while positive ions (cations) migrate toward the cathode away from the anode. At the anode an oxidation reaction produces electrons, transferring charge between the internal circuit, the electrolyte and the external circuit (the wire).

Two kinds of cells will be observed. In electrolysis (*electrolytic cell*) an external power source is used to move charge through the circuit and force the chemical reaction at the electrodes.



In *galvanic* or *voltaic cells*, a spontaneous chemical reaction powers the charge movement. In either case reduction occurs at the cathode, while oxidation takes place at the anode.



Voltaic Cell Diagram

In this lab the only gases that would form at an electrode would be H<sub>2</sub> or O<sub>2</sub> from the water (solvent). Thus, gas bubbles at the anode would be O<sub>2</sub> from the oxidation of H<sub>2</sub>O, while bubbling at the cathode would imply H<sub>2</sub> from the reduction of water.

## Procedure

### Part 1 - Electrolysis (Electrolytic Cells)

1. Connect graphite electrodes to a 6 volt power source using the alligator clip leads. Fill a 50ml beaker about 2/3 full of distilled water and insert the electrodes, taking care that they do not contact each other. Apply power and note whether any changes take place. Disconnect power.
2. a) Add a spoonful of Na<sub>2</sub>SO<sub>4</sub> to the beaker and stir to dissolve it. A strip of stiff paper can be inserted to divide the beaker into two sections. Reconnect the electrodes to the power source and again insert them in the beaker. Record any observable changes.  
 b) Stir in a few drops of bromothymol blue and repeat the electrolysis long enough to note what changes occur. What is the function of the Na<sub>2</sub>SO<sub>4</sub> in this case? Write half-reactions for the decomposition of water into its elements, H<sub>2</sub> and O<sub>2</sub> and identify at which electrode each half-reactions occurs. Identify the anode and cathode leads in this cell.
3. Replace the anode with a strip of copper metal, add 2 ml 6N H<sub>2</sub>SO<sub>4</sub> and electrolyze. Record observable changes and write half-equations for the electrode reactions.
4. Using graphite for both electrodes, replace the solution in the beaker with 0.1M CuBr<sub>2</sub> and electrolyze for a few minutes. Note the initial color and observe what

changes occur at each electrode. Disconnect the power and *cautiously* note the odor at each electrode.

## Part 2 - Galvanic Cells

1. Direct reactions. Polish strips of the metals listed below with a little steel wool to make the surfaces shiny. Then dip each strip into a shell vial containing a little 1.0 M  $\text{CuSO}_4$ . Record any observable changes. Give the equation for the reaction. The metals to test: (a) Al, (b) Fe, (c) Zn, (d) Cu, (e) Pb.
2. Construct a galvanic cell by linking two half-cells as follows. Fold a filter paper to about  $\frac{1}{2}$  inch thickness and dip it into a solution of 0.5M  $\text{KNO}_3$ . Place 20 ml of the appropriate electrolyte into a 50ml beaker for each half-cell. Connect two half-cells by laying the strip of soaked filter paper with each end dipping into one of the solutions. Insert the appropriate electrode into each half-cell and connect them to the voltmeter. Record the voltage generated in each case. The cells to be used are
  - a) Cu in  $\text{CuSO}_4$  with Pb in  $\text{Pb}(\text{NO}_3)_2$
  - b) Cu in  $\text{CuSO}_4$  with Fe in  $\text{FeSO}_4$
  - c) Cu in  $\text{CuSO}_4$  with Zn in  $\text{ZnSO}_4$

For each of these galvanic cells, identify the electrode where oxidation occurs (anode), and where the reduction occurs (cathode). Write the half-reaction occurring at each electrode and using a standard reduction potential table, calculate  $E^\circ$  and  $E$  for each cell. In a table, compare the calculated  $E^\circ$  and  $E$  values to the measured  $E$  values.

When dismantling the cells, rinse the electrodes and return them to the containers provided. Solutions can be flushed down the drain with water.

### **Stock Room**

For each pair of students allocate the following:

- 1 voltmeter with
- 1 pair of leads: red and black
- alligator clips and
- probe connectors to the voltmeter
- 2-3 strips of each metal: Al wire; Zn, Cu and Pb sheet cut in strips; Fe nails
- drafting pencil leads for graphite electrodes
- qualitative filter papers
- 150 ml  $\text{CuSO}_4$ , 1M
- 50 ml  $\text{Pb}(\text{NO}_3)_2$ , 1M
- 50 ml  $\text{FeSO}_4$ , 1M (stabilized with 0.25 M  $\text{H}_2\text{SO}_4$  to prevent air oxidation)
- 50 ml  $\text{ZnSO}_4$ , 1M
- 0.5 M  $\text{KNO}_3$ , 200 ml per lab section
- $\text{Na}_2\text{SO}_4$
- 0.1 M  $\text{CuBr}_2$
- 6 N  $\text{H}_2\text{SO}_4$