
The Energizer Lab

OBJECTIVES

- You will draw and label the parts of an electrochemical cell.
- You will identify and explain the reactions occurring at the anode and cathode.
- You will explain the flow of electrons in relation to the cell and the resulting voltage.
- You will explain the movement of cations and anions in regard to the anode, cathode and salt bridge.

PROBLEM / QUESTION

You will build a voltaic cell to observe the process macroscopically and measure the electric current. You will then work with a voltaic cell model to understand the process at a particulate level.

SAFETY

CuSO_4 -- Mildly toxic by ingestion. Irritant to skin, eyes and mucous membranes. Avoid contact with body tissues.

ZnSO_4 -- Mild body tissue irritant. Avoid body tissue contact.

MATERIALS

0.10 M CuSO_4

1.0 M ZnSO_4

test tube

Voltaic Cell Kit OR you can set up a cell using a 400 mL beaker, a porous cup and alligator clip wires.

1.0 M CuSO_4

2-10cm pieces of Zinc wire

Voltmeter

Electronic Balance

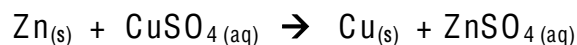
1-10cm copper wire

PRIOR KNOWLEDGE

Part I: Place a 10 cm piece of zinc wire into a small test tube containing approximately 5 mL of 0.10 M CuSO_4 . List your observations over a period of about five minutes. You may follow the procedure for Part II and then complete the observations.

Observations:

Part II: Given the following reaction:



Identify the oxidation and reduction half-reactions. Write the half-reactions below.

Oxidation: _____

Reduction: _____

Part III: Relate the oxidation and reduction half-reactions in Part II with the observations made in Part I. Explain the transfer of electrons that takes place in the oxidation and reduction half-reactions.

Oxidation: _____

Reduction: _____

Dispose of the chemicals in the waste container provided.

STOP HERE AND DISCUSS: Use the space below to record notes.

BACKGROUND INFORMATION

Electrochemistry deals with the relation of the flow of electric current to chemical changes. Refer to the half-reactions discussed in the Prior Knowledge section of this lab. If this reaction occurs in a beaker, electrons are transferred directly from the atoms at the surface of the zinc, which is acting as an electrode in this reaction, to the Cu^{2+} ions. This is a chemical change. If the electron transfer can occur through a wire connected to each electrode, the energy may be harnessed as an electric current. To do this, we separate the oxidation (which occurs at the anode) and reduction (which occurs at the cathode) half reactions into two separate containers called cells. This set-up is called a **galvanic cell** or **voltaic cell**. Even though the half-reactions are not in direct contact with each other, the electrons lost by the zinc atoms are able to flow along the wire to the other half-reaction cell where they are transferred to the Cu^{2+} ions. This flow of electrons through a wire is called an **electric current**. A salt bridge is necessary to balance out the charge of the cells so that the current may continue to flow as long as the other components are present. This salt bridge connects the two cells allowing for selective transfer of ions. The metal ions may not pass thru the salt bridge, but the anions may. A **battery** is two or more **electrochemical cells** connected to each other so that they operate as a unit.

What forces this transfer of electrons which results in an electrical current? When two different metal electrodes are placed in electrolytic solutions (a solution that conducts a current), electrons move from the electrode with the higher electrical potential energy to the electrode with the lower electrical potential energy. This difference in electrical potential energy is called the **potential difference**. This difference also referred to as the **electromotive force (EMF)** “drives” the electrons to travel from the anode to the cathode. The ability of a voltaic cell to move electrons through a wire is called the **cell potential**. The cell potential is measured in units of volts (V) and is called the cell voltage.

PREDICTION

Read through the procedure. Your group will make a voltaic cell using the reaction seen in Part II of PRIOR KNOWLEDGE. Predict if the copper and/or zinc strips will gain or lose mass. Explain the reasoning for your prediction below.

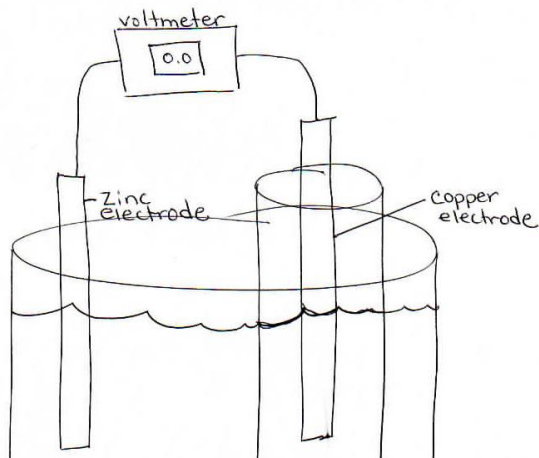


Figure 1: Voltaic Cell

Figure 1

PROCEDURE

Part A:

1. Mass the zinc and copper wires. Record the masses in the data section.
2. Set up the Voltaic Cell Kit or use the beaker and porous cup.
3. Add 100 mL of 1.0 M ZnSO_4 to a 400 mL beaker (or battery cell from kit).
4. Add 25 mL of 1.0 M CuSO_4 to the porous cup. Place the porous cup into the zinc sulfate solution.
5. Place the zinc wire in the ZnSO_4 solution. Attach the strip to the circuit. Record your description of the zinc strip. in the data section.
6. Place the copper wire in the CuSO_4 solution in the porous cup. Attach the strip to the circuit. Record your description of the copper strip in the data section.
6. Add a voltmeter to the voltaic cell (refer to Fig. 1).
7. Set the voltmeter to 2 volts DC range. Measure the voltage and record in the data section.
8. Allow the voltaic cell to continue running. Continue to Part B.

Note: By connecting the electrodes, the electrons lost by the zinc are allowed to travel through the wire to the copper electrode within the porous cup. These electrons can then interact with the Cu^{2+} in the solution.

Part B:
THE MODEL



- The colors of the circles are used to differentiate the atoms/ions during this procedure. In nature, the atoms and ions involved are colorless.
- When two electrons are inserted into the indentations in the circles, the cation is converted into an atom, i.e., $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}$
- When two electrons are removed from the circles, the atom becomes a cation.
- The relative sizes of the atoms and ions within the model are not accurate.
- In a real electrochemical cell containing a concentration of 1.0M CuSO_4 , there should be about 55 water molecules per each Cu^{2+} and SO_4^{2-} ion. For simplicity, very few ions are used and no water molecules are represented.
- The anode (the electrode through which electrons enter a metal) is shown on the left-hand side of this model. It can actually be shown on either side. It does not make a difference in the working of the model.
- In reality, the reactions occur simultaneously. During the working of the model, each step will be broken down.

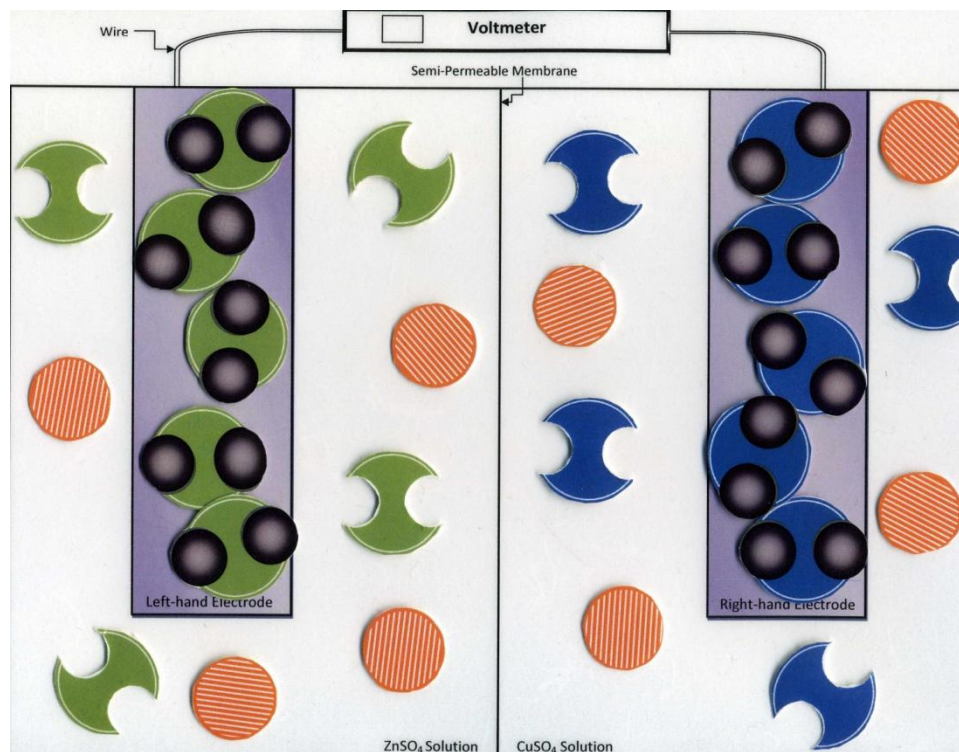


DIAGRAM #1: Voltaic Cell Model Kit

1. Set up the model to look like Diagram #1.

_____ Check Point -- Have your teacher check your model before you move on.

2. Begin using the model by finding a zinc atom at the zinc electrode. Remove two valence electrons from that zinc atom. Place the electrons onto the wire connecting the two electrodes.
3. The zinc cation must move off the zinc electrode into the ZnSO_4 solution.
4. The two electrons travel down the wire to the copper electrode. The negative charge attracts a copper cation to the electrode and the electrons are attached to that cation. There is now a copper atom added to the copper electrode.
5. Examine how the number of ions in the solutions has changed. Neutrality must be reestablished. Note that the sulfate ions present may now be transferred through the semi-permeable membrane (the porous cup) in any direction to reach electrical neutrality. Move the necessary ion(s) to reach neutrality.
6. Repeat the whole process as many times as possible, each time starting with the removal of two electrons from a zinc atom.

_____ Check Point -- Have your teacher check your understanding of the model before you move on.

Part C:

1. Disconnect the voltmeter.
2. Remove each electrode from the solutions and place on paper towel. Do NOT wipe the electrodes. Allow the electrodes to dry.
3. Mass the electrodes. Record your observations of each electrode in the data section.
4. Dispose of the chemicals in the waste container provided.

DATA AND OBSERVATIONS

Voltage: _____

Record the Mass and a Physical description of each of the following:

Zinc electrode before reaction

Zinc electrode after reaction

Copper electrode before reaction

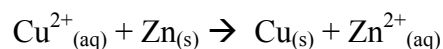
Copper electrode after reaction

Discuss the prediction you made before performing the procedure for Part A. Was your prediction correct or incorrect? Explain any difference in mass.

RESULTS AND ANALYSIS

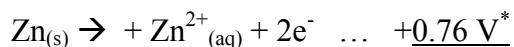
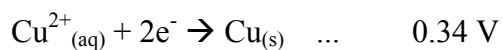
1. Did the electrodes in Part A change over time? Explain.
2. What conditions would cause the cell in Part A to stop working?

Look at the table of standard electrode potentials provided. You might notice that the elements are in the same order as they are on the activity series. The higher the element is in the reduction potential table, the more likely it is to be oxidized.



In a reaction like the one used in this activity, the direction of electron flow is determined by the fact that zinc is higher in the reduction potential table than copper is. That means that zinc is more likely to be oxidized (lose electrons) than copper (which will gain electrons). The voltage for the reduction of copper half-reaction is + 0.34. To look up the voltage for the oxidation half-

reaction, you need to first find the zinc reduction ($\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$) which is - 0.76. The oxidation voltage will be the same number with the opposite sign...+ 0.76. The sum of the individual voltages is the expected cell potential for the entire voltaic cell.



*Notice that the direction of the zinc half-reaction and the sign of the voltage are opposite the table. Once flipped, the new half-reaction explains what is happening to the zinc within the cell.

1.10 V

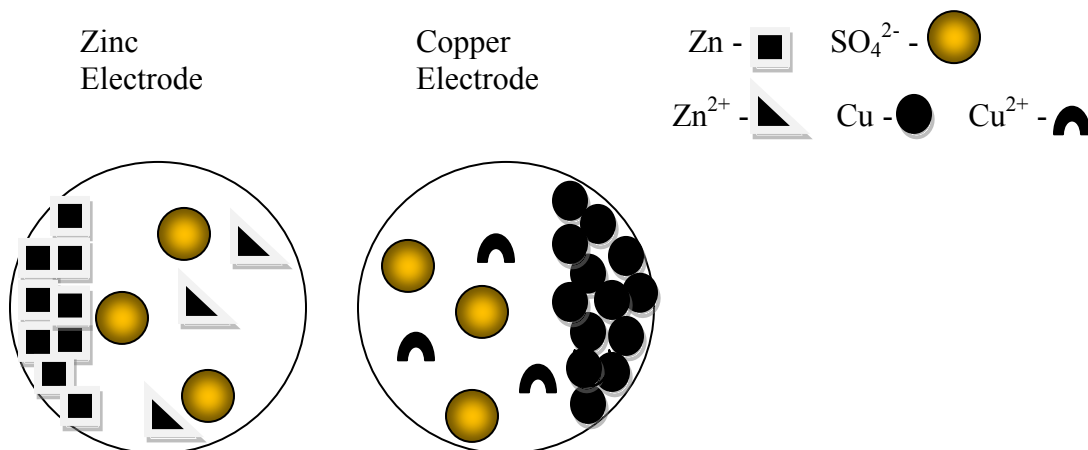
The sum of the two half-reactions must be positive for a working battery. With that knowledge, you may predict the oxidation (anode) and reduction (cathode) half-reactions before creating an electrochemical cell. It is important to note that the cell potential calculated may not be accurate. Notice that the table says the conditions for the measurement of the potentials are 25 °C, 1 atm and 1M solution. These exact conditions are unlikely in the laboratory.

3. If the following reaction occurs in a voltaic cell under standard conditions, calculate the voltage of the cell. $\text{Zn}_{(\text{s})} + 2 \text{AgNO}_{3(\text{aq})} \rightarrow 2 \text{Ag}_{(\text{s})} + \text{Zn}(\text{NO}_3)_{2(\text{aq})}$
(Note that the reduction half-reaction will have a coefficient of 2, but that does not mean that you need to double the reduction potential for that half-reaction.)

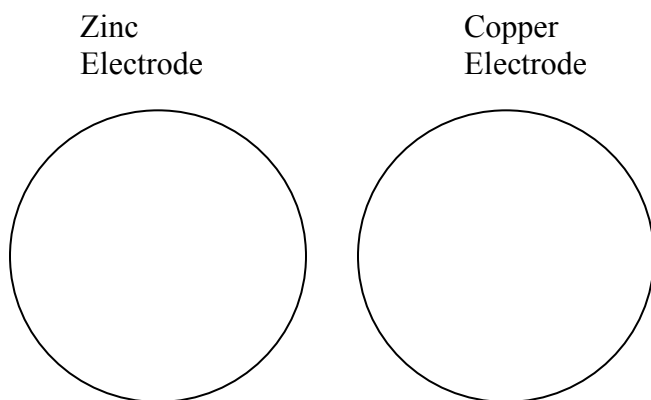
DISCUSSION

In the model provided, the cathode is on the right-hand side. It could be on either side, it makes no difference to the working of the model. For this reason, we often include an arrow in diagrams of voltaic cells to show the direction of flow of electrons.

1. The circles below represent the electrodes and surrounding solution at a particulate level before the reaction has begun.

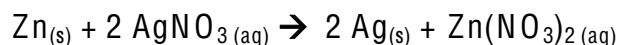


Draw a representation of the electrodes after the voltaic cell has been allowed to run for a long time.



2. How does a porous cup keep the electrolyte at each electrode electrically neutral?

3. Would the voltaic cell work if the semi-permeable membrane was replaced with a solid wall of metallic sheeting? Explain.
4. Discuss the similarities and differences (limitations) between the voltaic cell created in the lab and the model kit.
5. Look in your text or use another resource to find out what a “salt bridge” is. Draw and label a voltaic cell using the reaction below with a salt-bridge and discuss how that is similar to using a semi-permeable membrane. Label the anode and cathode. Designate the direction of electron flow. Write the appropriate half-reaction next to the related half-cell. Also calculate the cell potential.



GOING FURTHER

Research how the Standard Reduction Potential Table was created using the standard hydrogen half-cell. Also, read about how it is used to predict if a single replacement reaction will occur. Be prepared to explain.

Standard Reduction Potentials at 25°C *

*For all half-reactions the concentration is 1 M for dissolved species and the pressure is 1 atm for gases. These are standard state values.

Half-Reaction	E ⁰ (V)
$\text{Li}^+_{(\text{aq})} + \text{e}^- \rightarrow \text{Li}_{(\text{s})}$	-3.05
$\text{K}^+_{(\text{aq})} + \text{e}^- \rightarrow \text{K}_{(\text{s})}$	-2.93
$\text{Ba}^{2+}_{(\text{aq})} + 2 \text{e}^- \rightarrow \text{Ba}_{(\text{s})}$	-2.90
$\text{Sr}^{2+}_{(\text{aq})} + 2 \text{e}^- \rightarrow \text{Sr}_{(\text{s})}$	-2.89
$\text{Ca}^{2+}_{(\text{aq})} + 2 \text{e}^- \rightarrow \text{Ca}_{(\text{s})}$	-2.87
$\text{Na}^+_{(\text{aq})} + \text{e}^- \rightarrow \text{Na}_{(\text{s})}$	-2.71
$\text{Mg}^{2+}_{(\text{aq})} + 2 \text{e}^- \rightarrow \text{Mg}_{(\text{s})}$	-2.37
$\text{Be}^{2+}_{(\text{aq})} + 2 \text{e}^- \rightarrow \text{Be}_{(\text{s})}$	-1.85
$\text{Al}^{3+}_{(\text{aq})} + 3 \text{e}^- \rightarrow \text{Al}_{(\text{s})}$	-1.66
$2 \text{H}_2\text{O} + 2 \text{e}^- \rightarrow \text{H}_{2(\text{g})} + 2 \text{OH}^-_{(\text{aq})}$	-0.83
$\text{Zn}^{2+}_{(\text{aq})} + 2 \text{e}^- \rightarrow \text{Zn}_{(\text{s})}$	-0.76
$\text{Cr}^{3+}_{(\text{aq})} + 3 \text{e}^- \rightarrow \text{Cr}_{(\text{s})}$	-0.74
$\text{Fe}^{2+}_{(\text{aq})} + 2 \text{e}^- \rightarrow \text{Fe}_{(\text{s})}$	-0.44
$\text{Co}^{2+}_{(\text{aq})} + 2 \text{e}^- \rightarrow \text{Co}_{(\text{s})}$	-0.28
$\text{Ni}^{2+}_{(\text{aq})} + 2 \text{e}^- \rightarrow \text{Ni}_{(\text{s})}$	-0.25
$\text{Sn}^{2+}_{(\text{aq})} + 2 \text{e}^- \rightarrow \text{Sn}_{(\text{s})}$	-0.14
$\text{Pb}^{2+}_{(\text{aq})} + 2 \text{e}^- \rightarrow \text{Pb}_{(\text{s})}$	-0.13
$2 \text{H}^+_{(\text{aq})} + 2 \text{e}^- \rightarrow \text{H}_{2(\text{g})}$	0.00
$\text{Sn}^{4+}_{(\text{aq})} + 2 \text{e}^- \rightarrow \text{Sn}^{2+}_{(\text{aq})}$	+0.13
$\text{Cu}^{2+}_{(\text{aq})} + \text{e}^- \rightarrow \text{Cu}^+_{(\text{aq})}$	+0.13
$\text{Cu}^{2+}_{(\text{aq})} + 2 \text{e}^- \rightarrow \text{Cu}_{(\text{s})}$	+0.34
$\text{I}_{2(\text{s})} + 2 \text{e}^- \rightarrow 2 \text{I}^-_{(\text{aq})}$	+0.53
$\text{Fe}^{3+}_{(\text{aq})} + \text{e}^- \rightarrow \text{Fe}^{2+}_{(\text{aq})}$	+0.77
$\text{Ag}^+_{(\text{aq})} + \text{e}^- \rightarrow \text{Ag}_{(\text{s})}$	+0.80

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