GALVANIC CELLS

One of the most important types of chemical reactions is the oxidation-reduction reaction. This reaction involves the transfer of one or more electrons from one atom or molecule to another. The substance which is gaining electrons is being reduced. The substance which is losing electrons is being oxidized. All reduction reactions must occur with a corresponding oxidation reaction (i.e., the electrons must come from somewhere). It is often useful to consider oxidation-reduction reactions in two parts called half-reactions. Added together, these two half-reactions make up the overall oxidation-reduction reaction.

$$N^+ + e^- \longrightarrow N \qquad \text{(reduction)}$$
$$M \longrightarrow M^+ + e^- \qquad \text{(oxidation)}$$
$$M + N^+ \longrightarrow M^+ + N \qquad \text{(overall reaction)}$$

Reduction potentials are relative measures of the driving force for a half-reaction to undergo reduction. Let's consider a spontaneous oxidation-reduction reaction. The half-reaction with the larger reduction potential will proceed as written. The other half-reaction will proceed in reverse, as an oxidation. The driving force of the overall oxidation-reduction reaction is the difference between the reduction potentials of the the two half-reactions (E_{red} for the reduction reaction minus E_{red} for the oxidation reaction).

A galvanic cell is an electrochemical device that can produce electrical energy from spontaneous oxidation-reduction reactions. All electrochemical cells have two electrodes, a cathode and an anode. Reduction reactions always occur at the cathode and oxidation reactions always occur at the anode. In galvanic cells, the cathode is charged positive and the anode is charged negative. The identity of the cathode and anode is determined by the relative reduction potentials of the half-reactions which make up the galvanic cell. The electrode in the half-cell with the larger reduction potential is the cathode. The electrode in the half-cell with the smaller reduction potential is the anode.

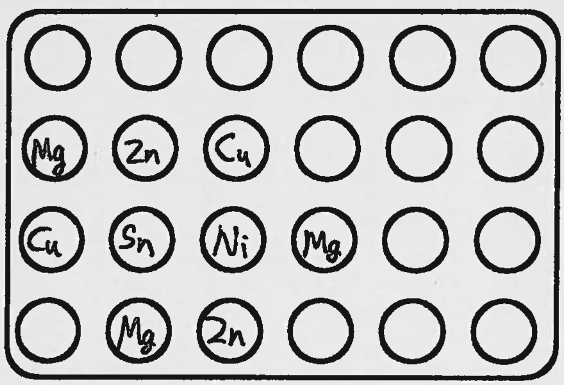
In this experiment, you will construct a series of galvanic cells. Each cell will consist of two half-cells, each containing a metal electrode and its corresponding ion in solution (such as copper wire in a Cu^{2+} solution). Pairs of half-cells will be connected together by a salt bridge that will supply inert cations and anions to each of the half-cells. By examining your results for a series of galvanic cells, you will be able to arrange five metal ions according to their ability to undergo reduction.

Materials

24-well plate (1)
filter paper, 1 x 2 cm (for salt bridge) (1)
1 M KNO_3
zinc, magnesium, nickle, copper, and tin metal strips, 1 to 2 cm (1 each)
15 drops of each of the following 0.1 M solutions:
    $ZnSO_4$, $MgSO_4$, $NiSO_4$, $CuSO_4$, and $SnCl_2$
sand paper (fine and extra fine) & steel wool (several pieces)

Procedure

1. Examine your 24-well plate. You can use any two adjacent wells to make a galvanic cell. Wells which are positioned diagonally will not work. Use the diagram below to design the most efficient arrangement of the half-cells so that you can test the polarity of every pair of half cells (you will need to use at least ten wells to cover all possible combinations). Your instructor will assist you if you get confused. Use this diagram to help you identify each of the solutions as you proceed through the experiment.



24-well Plate

2. Fill the wells that you have selected with 15 drops of the appropriate metal ion solutions.
3. Identify and label each of the five metals used in this experiment. Your instructor will explain how to identify each metal.
4. Clean each of the metal strips (or wires) with sandpaper and place them on a piece of paper next to a chemical symbol that identifies the metal.
5. Make a salt bridge by soaking a 1 x 2 cm strip of filter paper in 1 M potassium solution.

Before assembling any cells, consider your task at hand. You will be determining the relative reduction potentials of five half-reactions. It might appear at first that you will have to make 25 measurements (5^2). However, remember that you do not have to compare half-cells with themselves, and, if you compare the potential of half-cell "A" to half-cell "B," you also know the answer for "B" to "A."There are further data reduction possibilities as you proceed through the experiment. In fact, if you are clever and very lucky, you can obtain all the information about the relative reduction potentials with as few as four measurements (typically, you will need about ten measurements).

Use the following table to record your data as you proceed through this experiment.

CATHODES (RED)

A	black\red	Zn	Mg	Ni	Cu	Sn
N	Zn	X				
O	Mg		X			
D	Ni			X		
E	Cu				X	
S	Sn					X

6. Select two wells to be tested. Place the salt bridge so that it is immersed in both solutions.
7. Attach the alligator clips of the voltmeter to the metal strips or wires of the corresponding solutions.
8. Immerse the metal electrodes into their metal ion solutions and record the voltage. Reverse the connectors from the volt meter and record the voltage again. **Be sure to always use the correct metal electrode for each corresponding solution.**
9. Continue examining pairs of half-cells until you have completed the table either by experimentation or deduction.

10. Examine your completed table. The half-reaction with the highest reduction potential is the metal whose ion is easiest to reduce. Now deduce from your data the next easiest ion to reduce and continue until you have the correct order for all five ions.

Highest Reduction Potential ----------> Lowest Reduction Potential

(easiest to reduce) (hardest to reduce)

Compare your results with a table of standard electrochemical reduction potentials.

QUESTIONS

1. Cathodic protection is the method most often employed to protect buried fuel tanks and pipelines and the hulls of ships which are made of iron. An active metal which oxidizes more easily than iron is attached to the metal tank, pipeline, or ship hull. Which of the metals in this experiment would be useful for the cathodic protection of iron pipes?
2. Why must the salt bridge be in contact with both solutions in order for the galvanic cell to pass current?
3. In 1973, the wreckage of the Civil War ironclad USS *Monitor* was discovered near Cape Hatteras, North Carolina. (The *Monitor* and the CSS *Virginia*, formally the USS *Merrimack*, fought the first battle between iron-armored ships.) In 1987, investigations were begun to determine whether the ship could be salvaged. It was reported in *Time* (June 22, 1987) that scientists were considering attaching zinc anodes to the rapidly corroding metal hull of the *Monitor*. Describe how attaching zinc would protect the hull of the *Monitor* from further corrosion.
4. Will zinc metal react when immersed in a Cu^{2+} solution?
5. Will silver metal react when immersed in a Cu^{2+} solution?