THE POTENTIAL OF ELECTROCHEMICAL CELLS

In studying oxidation-reduction reactions, we have compared the tendencies of metals to lose electrons. For example, a metal, X, was placed in a solution containing ions of another metal, Y^+. If a reaction occurred, it meant that X was more easily oxidized than Y. Therefore, $X^+ + e^- \rightarrow X$ would be placed below $Y^+ + e^- \rightarrow Y$ in a reduction table. However, we have not measured this tendency quantitatively.

The measurement of the voltage of a cell reaction is the quantitative measure of the tendency of a chemical substance to gain electrons. If, for example, a substance that has a strong attraction for electrons (oxidizing agent) is placed in one half-cell, and a substance having electrons which can be removed quite easily is placed in another half-cell, the voltage or **cell potential** resulting from the reaction should be measureable when the circuit is completed. The position of the half-reactions in the redox table helps us predict the size of the voltage produced by the reaction.

In this experiment, you will perform oxidation-reduction reactions in which the electrons are transferred through an external wire. The electrons in the wire will deflect the needle of the measuring device, a voltmeter. Though it is necessary to separate the two half-cell reactions, the electric circuit is completed by a conducting salt bridge or a porous cup's contact between the half-cells.

Objectives

In this experiment, you will

- measure the voltage of electrochemical cells,
- write half-cell reactions to indicate loss and gain of electrons, and
- calculate the theoretical voltages of the electrochemical cells and compare them to experimental values.

EQUIPMENT

goggles and apron
2 beakers (250 cm^3)
porous cup or salt bridge
voltmeter, DC
2 wires
alligator clips

PROCEDURE

As you read the procedure, list all data and observations to be recorded. Using your list, prepare an organized data table. Safety goggles and lab apron must be worn for this experiment.

1. Metal strips (1 cm × 8 cm) of copper, lead and zinc should be properly cleaned prior to use. Sand

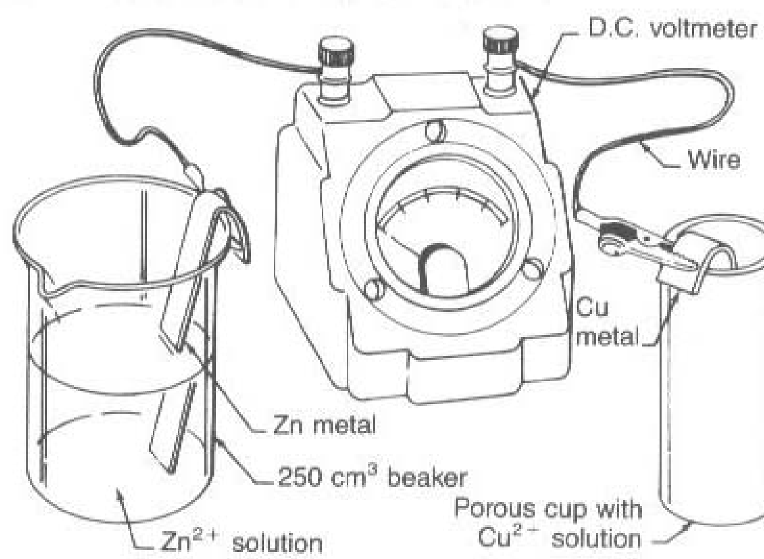

FIGURE 45-1. Separated half-cells with voltmeter.

FIGURE 45-2. Copper and zinc half-cells combined.

with a fine grade of sand paper or emery cloth to remove the outside coating.

2. Place about 20 cm^3 of Zn^{2+} solution in a 250 cm^3 beaker. Immerse a carefully cleaned strip of Zn metal in the solution as shown in Figure 45-1.
3. Obtain a porous cup and fill it to a height of 2 cm with Cu^{2+} solution. **CAUTION:** *Copper compounds are toxic. Do not inhale and avoid skin contact.* Place a clean strip of Cu metal in the solution to complete the half-cell.
4. Connect one wire to the Zn metal and another to the Cu metal, and then connect each of the wires to the voltmeter.
5. Pick up the porous cup and momentarily touch it to the solution in the beaker to see if the voltmeter is connected properly. The needle on the meter should show a positive deflection. If the deflection of the needle on the meter is not in the correct direction, reverse the wires on the meter. Immerse the porous cup in the Zn^{2+} solution and read the voltage immediately. (The electrodes become polarized and the voltage drops in a very short period of time.)
6. In another beaker prepare a $Pb|Pb^{2+}$ half-cell. **CAUTION:** *Lead compounds are poisonous. Do not inhale and avoid skin contact.* Thoroughly clean the Cu metal strip used for the electrode and rinse the outside of the porous cup with distilled water. Immerse the Cu^{2+} cup in the Pb^{2+} solution. Record the voltage of the copper-lead cell.
7. Return the Cu^{2+} solution to the stock bottle (or do as directed by your teacher). Rinse the porous cup inside and outside with distilled water. Set up the $Zn|Zn^{2+}$ half-cell in the porous cup. Immerse the Zn^{2+} cup with a clean strip of Zn in the Pb^{2+} solution. Record the potential difference between the Zn and Pb half-cells.
8. **Optional.** While the $Zn|Zn^{2+}$ half-cell is immersed in the $Pb|Pb^{2+}$ half-cell and attached to the voltmeter, add some 2*M* Na_2SO_4 a few cm^3 at a time to the Pb^{2+} solution. Watch the voltmeter and note any changes in the voltage or in the appearance of the solution.
9. Return the Zn^{2+} solution to the stock bottle. Return the metal strips to the reagent table. Return the contaminated Pb^{2+} solution to a container designated by your teacher.

ANALYSIS

1. Write each half-cell reaction and the total cell reaction for each cell prepared in this experiment. Indicate the substance oxidized and the substance reduced in each reaction.
2. Calculate the voltages, $E°$, for each cell reaction using the reduction potentials in Table A-12 of the Appendix.

CONCLUSIONS

1. Compare the calculated $E°$ values with your experimental results and give reasons for their differences, if any differences exist.
2. Draw a picture of two beakers connected with a salt bridge and show the electrodes wired together through an external wire. Label one half-cell as $Mn|Mn^{2+}$ and the other as $Au|Au^{3+}$. Indicate the flow of electrons, write the balanced cell reaction, and predict the cell voltage. Indicate what particles migrate toward the cathode and which to the anode.
3. What is standard oxidation potential?
4. **Optional.** Consider electrode half-reactions as equilibrium reactions. Explain your observations in Step 8.