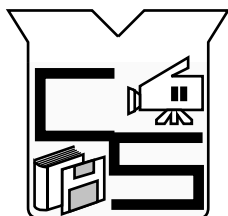
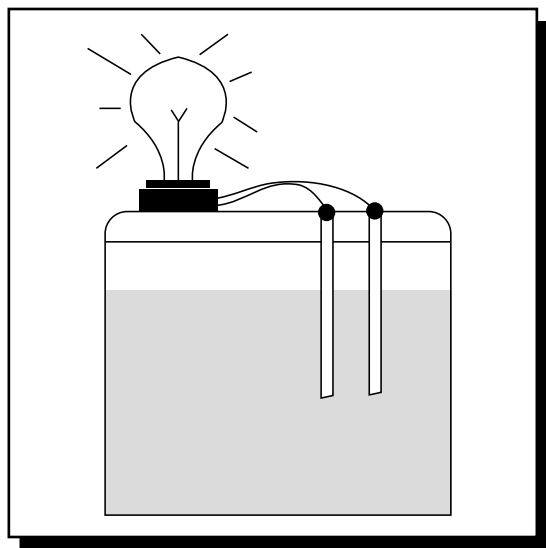


# A SourceBook Module

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*ChemSource Project Principal Investigator:  
Mary Virginia Orna, OSU  
Department of Chemistry  
College of New Rochelle  
New Rochelle, NY 10805  
Phone: (914) 654-5302  
FAX: (914) 654-5387*

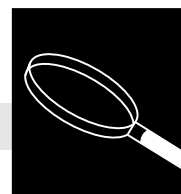


ChemSource

*Instructional Resources for Preservice and  
Inservice Chemistry Teachers*

ELECTROCHEMISTRY

# Topic Overview



## CONTENT IN A NUTSHELL

Like all of chemistry, electrochemistry is concerned with what electrons do. Electrons in atoms, molecules or ions are bound with a particular energy. When oxidation-reduction (redox) reactions occur, electrons are transferred from the substance being *oxidized* to the substance being *reduced*. In the process, energy is either released or absorbed, depending on the electron-binding energy difference between the reacting substances. A battery (a *voltaic cell*) is a device that allows the chemical energy released by a spontaneous oxidation-reduction reaction to do electrical work (*e.g.*, light a light bulb, power a radio.) An *electrolytic cell* can reverse this process by using external electrical energy to bring about a nonspontaneous redox reaction (*e.g.*, electroplating an automobile bumper.) In both cases, the electron transfer reactions occur at electrodes. The oxidation reaction occurs at the anode, and the reduction reaction occurs at the cathode.

Some substances give up or accept electrons easier than others and can be organized by their relative ability to undergo reduction reactions. A table of standard reduction potentials orders substances by their ability to accept electrons from a common donor (hydrogen gas). Among common substances, fluorine has the highest value (+2.87 V) for its reduction potential. It might be considered the *Tyrannosaurus Rex* of the Periodic Table because it can gobble electrons from anything. Lithium does not undergo reduction readily and has a very low reduction potential (−3.05 V). The difference between the reacting species in terms of reduction potential is an indication of the driving force for an electron-transfer reaction. In a voltaic cell, this difference defines the cell's potential and is the same as the electric potential for a standard single-cell battery. In an electrolytic cell, the cell potential is the electrical pressure (electric potential, expressed in volts) needed to drive the oxidation-reduction reaction.

Electrical charge is conducted through a solution by movement of cations and anions. In a voltaic cell, electrons are transferred from the cathode to the species being reduced. Cations must migrate to the cathode to offset the increase of negative charge near the electrode. Similarly, anions must migrate to the anode to offset the buildup of positive charge generated by the oxidation reaction at that electrode (see Transparencies 1 and 2 in the *Appendix*). The charge on the electrodes and the direction of movement of ions is just the opposite in electrolytic cells (see Transparencies 3 and 4).

The quantity of work that can be produced by a battery or the quantity of work needed to run an electrolytic cell can be calculated from the electric potential, the current, and the time the cell operates. Thus, electrochemistry is the study of the interconversion of electrical and chemical energy.

## PLACE IN THE CURRICULUM

The study of electrochemistry relies heavily on conceptual understanding of redox reactions. In essence, electrochemistry involves practical applications of redox chemistry. When students balance redox reactions by the *half-reaction method*, they are essentially separating the full reaction into two electrochemical half cells.

The structure and chemical properties of elements and molecular and ionic substances are governed by the chemical activity of their valence electrons. The energy that binds these electrons determines the value of a substance's standard reduction potential. Consequently, strong ties are apparent between electrochemistry and periodicity.

Whether an electrochemical cell operates in a voltaic or electrolytic mode depends on the relative free energy value for each half-cell. Observed phenomena in electrochemistry therefore directly relate to differences in thermodynamic functions of state (free energy, entropy, and enthalpy) and to temperature.

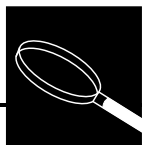
In electrochemistry, electrons are transferred in multiples of whole numbers. Therefore, there is a stoichiometric relationship between the quantity of charge transferred and the amounts of species oxidized and reduced.

1. Electrochemistry is a subtopic of oxidation/reduction.
2. Electrochemical cells contain several essential components:
  - a. Cathode, where reduction occurs
  - b. Anode, where oxidation occurs
  - c. Electrolyte (can be liquid or gel; is sometimes a salt bridge), which serves to conduct charge *via* moving ions in the cell
  - d. External circuit, where charge is conducted as a current of moving electrons
3. There are two kinds of electrochemical cells:
  - a. **Voltaic** (also called Galvanic)  
Reactions spontaneous  
Chemical energy transformed to electrical energy
  - b. **Electrolytic** (includes, but not limited to, electroplating)  
Reactions not spontaneous  
Electrical energy transformed to chemical energy
4. Electrochemical reactions occur at the surface of the electrodes.
5. Substances are categorized by half-reaction reduction potentials, which can be used to predict the spontaneity of oxidation-reduction reactions and the electric potential of batteries.
6. Technological applications of electrochemistry include batteries, fuel cells, electroplating, protection from corrosion, and chemical instrumentation.
7. Electrochemistry can be treated quantitatively at several levels of sophistication:
  - a. Calculation of cell potentials for batteries (essential)
  - b. Calculation of stoichiometric relationships (can be omitted if time is needed to treat other important chemistry topics)
  - c. Electrical work (should be treated only if time permits and if students have completed physics as a prerequisite)
  - d. Calculations using the Nernst equation or equations related to it or the Gibbs free energy equation (should not be treated in the general first-year high school course)

## CENTRAL CONCEPTS

1. Electrical energy
2. Electric potential energy difference (volts)
3. Chemical energy
4. Spontaneous and nonspontaneous reactions

## RELATED CONCEPTS



5. Conservation of charge
6. Anions, cations, and electrolytes
7. Electrostatic attraction/repulsion
8. Mole relationships in chemical reactions, balancing equations
9. Oxidation-reduction
10. Electronegativity, ionization energy, and electron affinity

## **RELATED SKILLS**

1. **Math/computation skills**  
Basic arithmetic computation  
Graphing  
Determining all possible combinations  
Proportional reasoning
2. **Laboratory skills**  
Basic manipulation of glassware and chemical substances  
Completing circuits according to a circuit diagram  
Using a balance  
Reading meters and other instruments

## **PERFORMANCE OBJECTIVES**

Upon completion of their study of electrochemistry, students will be able to:

1. distinguish between anions and cations.
2. define anode and cathode in terms of oxidation and reduction.
3. describe how a voltaic cell produces an internal ionic flow and an external electron flow.
4. write half-cell equations and total equations for voltaic cells.
5. in relation to a voltaic cell, define or explain: anode, cathode, electric potential (volts), salt bridge, internal circuit, external circuit.
6. describe how a battery produces electrical energy.
7. identify the substance being oxidized and the substance being reduced in an electrochemical cell.
8. describe the operation of an electrolytic cell.
9. define cathode, anode, and explain the charge on the cathode and anode in an electrolytic cell.
10. explain similarities and differences between voltaic and electrolytic cells.
11. explain the operation of an apparatus for electroplating with metals.
12. describe the zero-potential hydrogen half-cell.
13. given a table of standard reduction potentials: determine whether a redox reaction will occur; predict the electric potential of a voltaic cell made from two different half-cells; predict the products of an electrolysis reaction.
14. given the reduction potential of one half-cell and the electric potential of a voltaic cell, calculate the reduction potential of the other half-cell.



# Concept/Skills Development

## Activity 1: A Study of Voltaic Cells

### Purpose

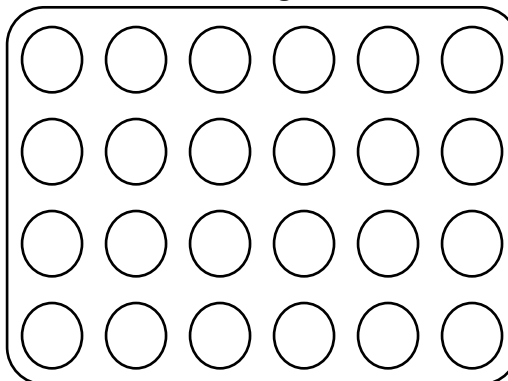
To generate an activity series by constructing and comparing several electrochemical cells.

### Safety

1. Wear protective goggles throughout the laboratory activity.
2. Any time you spill a solution on your clothes or body, wash it off immediately.
3. Silver nitrate spills on skin will cause dark stains and burns. Handle it with special care.
4. Dispose of all waste materials as your teacher directs.

### Procedure

1. Collect the following materials:  
24-Well plate  
Salt-bridge  
Voltmeter (High-impedance)  
1-2 cm Strips of Zn, Mg, Ag, Cu, and Sn  
Dropper each of zinc sulfate, magnesium sulfate, silver nitrate, copper(II) sulfate, and tin(II) chloride
2. Examine the well-plate. Use any two adjacent wells to make a voltaic cell. Use the diagram below to design the most efficient arrangement of half-cells so you can measure the electric potential (voltage) of every pair of half-cells. To be most efficient, use one well that is not on the edge of the plate to make your silver half-cell. The silver half-cell involves expensive materials; you will assemble only one of these. If you are unsure about your arrangement, have your teacher check it before you start.

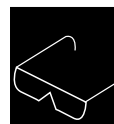


24 Well Plate

**Figure 1. Well plate.**

3. Fill each selected well with about 15 drops of appropriate metal ion solution.
4. Identify and label each of the five metals used in this activity according to directions provided by your teacher.
5. Clean each metal strip and wire with sandpaper. Place the cleaned metal samples on a piece of paper next to a chemical symbol that identifies each metal.
6. Make a table to record your data. For example, comparing Mg with Zn is the same as comparing Zn with Mg. You should find that 10 sets of measurements are needed to make all possible comparisons. However, if you are very clever and a little lucky, you'll be able to answer all the questions with a minimum of *four* comparisons.

## LABORATORY ACTIVITY: STUDENT VERSION





7. Select two half-reactions to be tested. Connect them with a salt bridge. Be sure each end of the salt bridge is immersed in the half-cell solution.
8. Place appropriate metal strips or wires in the half-cell. For example, place Zn metal in the zinc sulfate solution, Ag in the silver nitrate solution, and so forth. *NOTE: Always use the correct metal electrode for each solution.*
9. Connect the metal strips or wires to the voltmeter. Attach the voltmeter so the cathode of the voltaic cell is attached to the positive lead of the voltmeter. You can tell by seeing which connection of the voltaic cell to the voltmeter gives a positive electric potential. Record the electric potential (in volts) for that pair, and record which metal is the cathode.
10. Continue testing pairs until you can answer all the questions your teacher has assigned.
11. Dispose of all solutions as your teacher instructs. The salt bridges and metal strips or wires can be reused; do not throw them away. Clean them as your teacher directs.
12. Thoroughly wash your hands before leaving the laboratory.

### **Data Analysis and Concept Development**

1. Examine your data. List the half-reactions from highest reduction potential (most likely to be the cathode) to lowest reduction potential (least likely to be the cathode.)
2. If you have measured electric potential energy differences between pairs of half-cells and your teacher has provided the known reduction potential of one of the half-reactions, calculate reduction potentials for the other four half-reactions.
3. Compare your results with a table of standard reduction potentials. Are there any differences between either the order you observed or the values you calculated? If there are differences, what might be possible reasons? (Hint: How were the standard reduction potentials measured, as compared with your procedure?) The reasons you offer to explain differences should be specific enough to be tested in the laboratory, if time allows.
4. Why must the salt bridge be in contact with both solutions for the voltaic cell to generate current?
5. Will zinc metal react when immersed in a solution of copper(II) sulfate? Will silver metal react when immersed in a solution of copper(II) sulfate?

### **Implications and Applications**

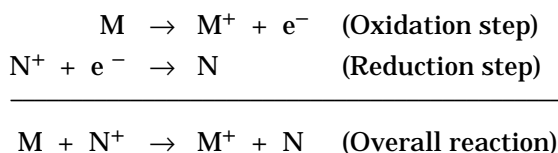
1. Cathodic protection is the method most often employed to protect buried iron fuel tanks, iron pipelines, and iron ship hulls. An active metal that oxidizes more readily than iron is attached to the iron object being protected. Which metals used in this activity would be useful for the cathodic protection of iron? The reduction potential for the  $\text{Fe}/\text{Fe}^{2+}$  half-reaction is  $-0.41\text{ V}$ , while the reduction potential for the  $\text{Sn}/\text{Sn}^{2+}$  half-reaction is  $-0.14\text{ V}$ .
2. In 1973, the wreckage of the Civil War ironclad USS *Monitor* was discovered near Cape Hatteras, North Carolina. (The *Monitor* and the CSS *Virginia*, formerly 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 *Monitor* metal hull. Describe how attaching zinc would protect the hull from further corrosion.

## Activity 1: A Study of Voltaic Cells

### Introduction

Among several major types of chemical reactions, one of the most important is oxidation-reduction. You may have studied the electron transport system and photosynthesis in a biology course, for example. Both systems are just a series of oxidation-reduction reactions, each involving a species that loses electrons (is oxidized), while another species gains electrons (is reduced).

Even though oxidation cannot occur without reduction and vice versa, it is often useful to consider oxidation-reduction reactions in two parts called half-reactions. Added together, the two half-reactions make up the overall oxidation-reduction reaction.



To decide whether a particular oxidation-reduction reaction will occur, it is helpful to think about the *reduction potential*. Reduction potential can be considered the driving force for a half-reaction to undergo reduction—that is, to go in the direction of the species gaining an electron. Consider a spontaneous oxidation-reduction reaction. The half-reaction with the larger reduction potential will occur as a reduction reaction. The half-reaction with the smaller reduction potential will run in the reverse direction—as an oxidation reaction in which electrons are lost rather than gained. For a more complete discussion of oxidation-reduction and half-reactions, refer to your textbook.

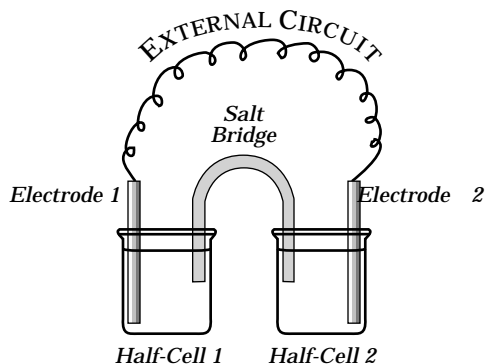
The standard way to compare half-reactions is in a voltaic cell. A voltaic cell is an electrochemical device that can produce electrical energy from spontaneous oxidation-reduction reactions. All electrochemical cells have two electrodes—a cathode and an anode. (An electrode supplies or accepts electrons from a chemical reaction.) Reduction reactions always occur at the cathode and oxidation reactions always occur at the anode. One easy way to avoid confusion is to remember that “oxidation” and “anode” both start with vowels, whereas “reduction” and “cathode” both start with consonants. In voltaic cells, the cathode is charged positively and the anode is charged negatively. The identity of the cathode and anode is determined by the relative reduction potentials of the half-reactions that make up the voltaic cell. The electrode in the half-reaction with the more positive reduction potential is always the cathode in a voltaic cell. The electrode in the half-reaction with the less positive reduction potential is always the anode in a voltaic cell.

Unlike reactions that occur in test-tubes when substances are mixed together, voltaic cells are arranged so that half-reactions are physically separated, often into different containers called half-cells. The half-cells are connected by a conducting wire between the two electrodes. The conducting wire is often called the external circuit. An electrical current passes through the external circuit. Electrical quantities such as current (rate of flow of electrons), electric potential (electric potential energy difference between the two half cells), and resistance can be measured in this external circuit. The unit for measuring electric potential is the *volt*, thus electric potential is often referred to as *voltage*. The electricity flowing through the external circuit can be used to provide energy to systems such as flashlight lamps, audio recorders, and science experiments. A system in the external circuit is often called a load.

## LABORATORY ACTIVITY: TEACHER NOTES



The half-cells are also connected by an internal circuit, often provided by a salt-bridge. The salt-bridge is composed of positive and negative ions that are free to move from one half-cell to the other but do not participate in the oxidation-reduction reaction. The salt-bridge is needed to keep charge from building up in the two half-cells. When electrons flow *from* a half-cell through the external circuit, negative ions travel to that half-cell to maintain electrical neutrality. Conversely, when electrons flow *into* the other half-cell, positive ions travel to that half-cell to ensure electrical neutrality there also.



**Figure 2. Voltaic cell.**

### Major Chemical Concepts

1. Electrochemistry represents a subset of oxidation-reduction.
2. Electrochemical cells include electrolytes, electrolyte bridge, electrodes labeled as anode and cathode, and an external circuit through which electrons flow.
3. Chemical energy is converted into electrical energy in a voltaic cell.
4. Half-cell potentials are relative, defined by arbitrary standards and an arbitrary zero value. (To successfully complete the activity and understand the concepts, it is not necessary to introduce the hydrogen half-cell as the arbitrary zero, but you may wish to do so.)
5. Charge is conserved in cells by the movement of ions.
6. Applications of the concepts developed in the activity include batteries and protection of artifacts from corrosion.
7. Cell potential or cell voltage—voltage associated with an electrochemical cell. It can be calculated from standard potentials and the Nernst equation.
8. The volt is the unit of electrical potential. It is defined as a Joule per Coulomb ( $1 \text{ V} = 1 \text{ J/C}$ ).

### Level

If you decide to have students *calculate* the half-reaction reduction potentials, this activity is most appropriate for general chemistry students. If students only measure the voltage, the activity may also be used with basic level students.

### Expected Student Background

The following concepts are prerequisites for successfully understanding concepts developed in this activity: oxidation and reduction; the general nature of energy including potential energy, chemical energy, and electrical energy; the nature and behavior of charged particles, conductivity in ionic solutions; spontaneity of chemical reactions; arbitrary standards for measurement. Students should have developed fundamental laboratory skills. If you elect to have students calculate reduction potentials of the half-reactions tested, students should have arithmetic and/or calculator skills.

## Time

If you prepare salt bridges, and dropper bottles of solutions in advance (see *Advance Preparation*), students should complete collecting data and clean up in 30-40 min. Very efficient students will be able to complete collecting data in considerably less time. Pre- and post-laboratory discussion times will vary, depending on how much concept development is tied to the laboratory experience.

## Safety

See safety instructions in the student instructions. Similar laboratory activities in commercially available curricula often use lead or nickel half-reactions. Because lead compounds are toxic and some nickel compounds have been identified as carcinogens, we recommend that these *not* be used.

## Materials (For 24 students working in pairs)

- Twelve 24-well plates
- 12 Salt bridges (see *Advance Preparation*)
- 12 Strips of Zn, Mg, Ag, Cu, and Sn, 1-2 cm in length
- 0.1 M Zinc sulfate,  $\text{ZnSO}_4$ , 100 mL (2.9 g  $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$  per 100 mL solution)
- 0.1 M Magnesium sulfate,  $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ , 100 mL (2.5 g  $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$  per 100 mL solution)
- 0.1 M Silver nitrate,  $\text{AgNO}_3$ , 100 mL (1.7 g  $\text{AgNO}_3$  per 100 mL solution)
- 0.1 M Copper sulfate,  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ , 100 mL (2.5 g  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$  per 100 mL solution)
- 0.1 M Tin(II) chloride,  $\text{SnCl}_2 \cdot 2\text{H}_2\text{O}$ , 100 mL (2.3 g  $\text{SnCl}_2 \cdot 2\text{H}_2\text{O}$  per 100 mL solution)
- 0.5 M Copper(II) sulfate,  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ , 100 mL (12.5 g  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$  per 100 mL solution)
- 0.5 M Tin(II) chloride,  $\text{SnCl}_2 \cdot 2\text{H}_2\text{O}$ , 100 mL (11.3 g  $\text{SnCl}_2 \cdot 2\text{H}_2\text{O}$  per 100 mL solution)
- High-impedance voltmeter (see section on instrumentation below for suggestions)
- Wires with alligator clips
- Polarity tester (optional—see *Extensions and Projects*)

In addition to 0.1 M solutions for students, prepare 0.5 M and 1.0 M copper(II) and zinc sulfate solutions for post-laboratory discussion. Also have available some test-tubes and strips of Zn and Cu. Use the metals and solutions recommended. Not all combinations work well under the conditions of this laboratory procedure.

## Advance Preparation

Check to ensure that all materials are available. Most materials are available in high school chemistry stockrooms. Silver wire can be ordered from a chemical supply company or may be available locally at a jeweler supply house.

## Salt Bridges

Prepare in advance the salt bridges. Materials needed include potassium nitrate, burner, razor blade, and thin-stemmed polyethylene pipets.

Quickly wave the pipet through burner flame to warm it. (You will be surprised at how quickly the tube softens.) While it is still soft, bend the tube into a “U” shape and allow it to cool. Do not kink the tube while bending it. Use a razor blade to cut off the bulb and excess stem. Leave about 2 cm of tube on each “arm” of the bridge. Make all the “U” tubes you will need before proceeding to the next step.



Mix 2 g potassium nitrate in 10 mL of water. Add 0.1 g agar. Boil the solution for 3 to 5 min. While the agar is cooking, stretch the end of an uncut, unbent polyethylene pipet to reduce the diameter of its tube. Do this by warming in the burner flame and stretching. Cut the end of the pipet so it is even. Remove the agar from the heat. Use the modified pipet to fill each “U” tube with the warm agar solution. Be certain to eliminate all air pockets. After the salt bridge has cooled, trim the ends with the razor blade. Store the salt bridges in a jar containing 2 M potassium nitrate. The salt bridges will last indefinitely if they do not dry out. Make several extras; those used with silver and tin half-reactions will plug up after several uses due to precipitation of silver chloride. This can be avoided by substituting tin(II) nitrate or acetate for tin(II) chloride, but the former are not commonly available in high schools.

If you do not wish to make permanent salt bridges, filter paper strips or pieces of cotton twine soaked in potassium nitrate solution work fine as temporary and easily disposable salt bridges.

### Pre-Laboratory Discussion

Pre-laboratory work will take two or three class periods if major concept development is tied to the activity. If your students have studied in advance all concepts to be developed by the activity, use the activity for concept review. The pre-laboratory would thus focus on demonstrating use of the apparatus. We recommend, however, that the activity be used for concept development.

Review prerequisite concepts briefly to ensure student understanding (see *Expected Student Background*). Ask certain students to explain each concept, followed by asking other students to elaborate, agree or disagree, correct, and provide examples. Your role is to serve as a “traffic director” for questions and responses, and to record on the board or overhead appropriate ideas to structure the review. You might administer an ungraded pretest quiz a day or so before the pre-laboratory discussion to help focus discussion. One useful way to do this is to provide a list of concept words in the domain being tested, such as oxidation-reduction or energy—and ask students to draw a concept map. (Concept maps are also useful after this discussion to summarize the review.)

Once prerequisite concepts have been reviewed, introduce laboratory equipment to be used. Demonstrate how voltmeter works by testing a small battery. Show students a 24-well plate with two solutions, metal wires or strips, and a salt bridge; connect the voltmeter.

If time permits, develop, *via* student discussion, the most efficient arrangement of wells for testing the various combinations and the minimum number of tests required for answering the questions. As students provide suggestions, ask them to explain how their arrangement will work and how data from their suggested tests can help answer the questions. Encourage students to propose tests that do not match up the silver and tin half-reactions. Although we recommend that *students* develop the well arrangement and test sequence, the arrangement in Figure 3 has been found to be very efficient. You can give it to students in advance to save time, but development of thinking skills will be compromised.

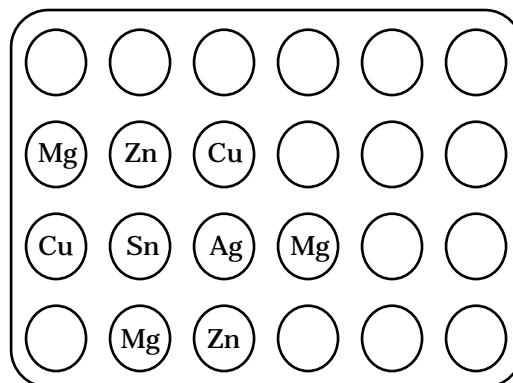


Figure 3. Well-plate arrangement.

## Teacher-Student Interaction

As students conduct tests, have them identify the cathode and the anode, telling you what observations support their answers. Ask them whether oxidation or reduction is occurring in a particular half-cell, to define whether this is gain or loss of electrons, and to write the predicted half-reaction equation. You may wish to ask only one question of each group as you move from group to group. In that event, you might wish to prepare a checksheet matrix of question type by group so you can ask each group a conceptually different question each time you stop by.

## Anticipated Student Results

Once laboratory errors such as placing the wrong metal into a well solution are taken into account, the half-reaction reduction potential order should be the same as that predicted by a standard table of reduction potentials. The actual values in volts for half reactions may differ somewhat from table values, depending on voltmeter impedance, meter reading errors, cleanliness of metal strip and wire surfaces, quality and size of salt bridge, and local concentrations of ions near the electrodes. Deviations from standard conditions under which tabled potentials are expressed should *not* much affect the results. (To understand why, review applications of the Nernst equation to cell potentials, as developed in college general chemistry textbooks.)

## Answers to Questions

### Data Analysis and Concept Development

1. See *Anticipated Student Results*.
2. See *Anticipated Student Results*.
3. Reasonable hypotheses include incorrect meter readings; using wrong metals in well plate solutions; and deviation from standard conditions of temperature, solution concentration, and/or pressure. Less likely to be proposed by students are surface coating of electrodes leading to changes in reduction potential, and lack of stirring in wells leading essentially to zero concentration of ions in well solutions near surfaces of electrodes. Accept any reasonable hypotheses as long as they are, in principle, testable.
4. For charge to be conserved in a half-cell, negative ions must replace lost electrons and positive ions must offset gained electrons.
5. Yes; No. (Be prepared to demonstrate this during post-laboratory discussion.)

## Implications and Applications

1. Of metals tested in this activity, Zn and Mg can be used to protect Fe. For economic reasons, Mg is not used.
2. The Zn undergoes oxidation more easily than does Fe. (It has a larger negative—therefore lower—reduction potential.) Therefore, when Zn is attached to the Fe hull, Zn will corrode via oxidation rather than the Fe.

## Post-Laboratory Discussion

Ask each group to report its results. Even if four of five groups obtain the same result, it is still possible that the *four* made a common mistake and the *one* completed the test correctly. Ask students to propose hypotheses to explain any discrepancies.



Ideally, student hypotheses can be tested directly at that time. A common hypothesis is that concentration makes a difference, so you should have available several solutions at higher concentrations to test that notion. By the close of this discussion, there should be general agreement on the observations—the cathode and anode in various combinations and the electric potential produced by those combinations—and reasons for any discrepancies. This discussion is an excellent opportunity to develop scientific thinking skills among your students.

Once observations are agreed on, process the information as a class. Repeat the questioning until an agreed order of reduction potential is established. If the activity's procedure is followed carefully, this order should agree with the sequence found in a table of standard reduction potentials. As part of this discussion ensure by questioning and explaining that students understand the concept of reduction potential.

After establishing the reduction potential order, give students the value from a standard table for one half-reaction used in the activity—any half-reaction potential will do. Once students know the half-cell potential for one half-cell, they can calculate the half-cell potential for another half-cell by finding the potential measured in the laboratory for that particular half-cell combination. Because a cell's electric potential is the sum of the potentials of its two half-cells, students can calculate a half-cell potential for the unknown half-cell. They can then proceed to calculate half-cell potentials for each half-cell studied in the activity. Our experience indicates that students learn to solve cell-potential problems much more readily by applying the relevant concepts to the laboratory *first*, followed by practicing textbook-type problems, rather than the reverse.

### **Extension**

Have students use their text and school library resources to find out what materials are used in commercially available batteries. Then ask students to use the table of half-reaction reduction potentials to explain why many batteries provide 1.5 V of electric potential. Radio Shack sells a book explaining the makeup of many common types of batteries. In addition to exploring radio batteries, students can also calculate the electric potential of car batteries based on half-cell potentials and the number of cells in a car battery. Additional activities in commercially available laboratory books are excellent for reinforcing the concepts developed in this activity. "Experiment 46: Corrosion" (Wilbraham *et al.*, 1987) is highly recommended, as is the essentially identical activity from Carmichael and Haines (1987).

### **Assessing Laboratory Learning**

#### **Laboratory Practical**

1. Provide students with two unknown metals and matching solutions. Ask them to determine which has the higher half-reaction reduction potential and/or the value of the cell electric potential. (The former task is more difficult conceptually.) The advantage of this assessment is its direct relation to laboratory procedures used by students. Disadvantages include set-up time, time needed for students to perform, and the large quantity of equipment required if individual students are to be tested simultaneously.
2. Provide students with a group of metals, test-tubes, and solutions. Without creating voltaic cells, have them determine the order, from lowest to highest, of half-reaction reduction potentials. Advantages include ease of set up, quick implementation by students, and the conceptual nature of the task. The primary disadvantage is that it does not duplicate the students' original laboratory procedure.

### Oral Questions

If you wish, you can use questions recommended for teacher/student interactions during the activity as an evaluation tool. If you wish to assess individual student progress, make a student-by-question matrix to ensure that you ask each student the same questions—or at least the same number of questions. Then, as you circulate through the laboratory, ask each student a different question, checking their relative success. By circulating back through the laboratory several times, you can ask each student several questions before the end of the laboratory period, referring to your matrix to see which questions a particular student has been asked. If student performance on questions is to be used for grading, inform students in advance. We recommend that questions asked during the activity not be used for grading purposes.

### Pencil and Paper

1. Provide students with written observations related to a similar laboratory activity. Ask students to interpret the information. **Advantage:** Duplicates the laboratory activity. **Disadvantage:** Requires considerable reading and therefore takes even good students a long time to complete.
2. Solve cell potential problems for half-cell potential, given the cell potential and the half-cell potential for one half-reaction. **Advantage:** Duplicates the laboratory activity. **Disadvantage:** Should probably not be used with basic level students.
3. Use questions recommended for teacher-student interactions during the laboratory activity later as written questions.

In the assessments described above, you may elect to allow students to refer to their notes, laboratory reports, and/or textbook.

### Other Laboratory Activity Ideas

Give students some bubble solution made of about three parts good quality liquid detergent such as Joy and about one part glycerol, a 9-V battery, some connecting wires, and several pieces of aluminum foil. Have students electrolyze the soap solution with the Al electrodes, holding the electrodes close together so that bubbles form containing a mixture of hydrogen and oxygen gas. Holding a lighted match to the bubbles gives a satisfying but safe explosion. [Edge, 1984]

### Demonstration 1: Voltaic Cells

#### *Purpose*

To demonstrate the ability of chemical reactions to produce electricity. The electric potential produced depends on the nature of the metal/metal ion half cell.

#### *Materials*

Petri dish

Filter paper

Dropper bottles of 0.1 M solutions of  $\text{KNO}_3$ , and  $\text{Mg}^{2+}$ ,  $\text{Zn}^{2+}$ ,  $\text{Sn}^{2+}$ , and  $\text{Cu}^{2+}$  as chloride or nitrate salt solutions.

Zn, Mg, Cu, and Sn strips, 1-2 cm

Wires with alligator clips

High-impedance voltmeter that will read 1-2 V

## DEMONSTRATIONS



### *Safety*

Use all normal safety precautions. Use accepted procedures for disposing of heavy-metal solutions.

### *Procedure*

Cut the filter paper in the shape of a cross. Place it in the Petri dish. Add some  $\text{KNO}_3$  solution to the center of the cross, allowing it to saturate the paper. Place one drop of each metal ion solution on each of the four corners. Attach two different metal electrodes to alligator clips wired to the voltmeter. Touch the proper electrode to its own metal ion solution. Substitute a different metal on the alligator clip until the electric potential of each of the six combinations is measured. This procedure avoids contamination and the need for a salt bridge, giving instant results.

## **Demonstration 2: Fruit and Vegetable Batteries**

### *Purpose*

To show that different combinations of electrodes produce different electric potentials and that some food materials can serve as electrolytes.

### *Materials*

Metal strip of zinc, copper, aluminum, iron, tin, lead, 1 cm x 3 cm  
Voltmeter  
2 Connecting wires with clips on both ends  
Firm fruit/vegetables (such as potato, lemon, grapefruit, orange)

### *Safety*

Do not eat these materials. Wash hands thoroughly when the demonstration is completed.

### *Procedure*

Cut the fruit sample in half and insert two unlike metal strips into it. Connect the strips to the voltmeter with connecting wires. Read the electric potential in volts. Experiment with placing the metal strips at various distances from each other. Try other fruit/vegetables with the same metals, and other combinations of metal strips. To reinforce concepts, have students predict in each case which of the two metals will be the cathode and which will be the anode.

## **Demonstration 3: Electrolysis of Water in Color**

### *Purpose*

To demonstrate electrode half-reactions.

### *Materials*

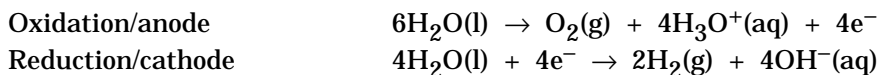
Hoffman apparatus (or U-tube fitted with 2-hole stoppers and platinum; Nichrome electrodes; Bent wires in inverted test-tubes will also work.)  
Bromthymol blue indicator (BTB)  
6-V to 9-V DC source, such as a transistor battery  
1 M HCl and 1 M NaOH, 5 mL each  
Wood splints  
2 Test-tubes

### *Safety*

Follow normal laboratory safety precautions.

*Procedure*

Set up the Hoffman apparatus with its attached power source. Prepare the volume of water needed for electrolysis by adding approximately 3 g of solid sodium sulfate and a squirt of BTB. Turn on the power and observe the color changes during electrolysis. While electrolysis is occurring, show students the color of BTB in HCl (yellow) and NaOH (blue). After each arm of the apparatus has collected at least 20 mL of gas, note that the volume in one arm is about half that in the other arm. Note the color of BTB in each arm. Collect samples of gas generated at each electrode and test with glowing splints and burning splints to determine which arm contains hydrogen gas and which contains oxygen gas.

*Reactions*

*NOTE: A U-tube set-up is preferred for seeing color differences, but withdrawing gas samples from a U-tube is difficult.*

**Demonstration 4: Electrolysis of Potassium Iodide***Purpose*

To show electrode half-reactions involving color changes

*Materials*

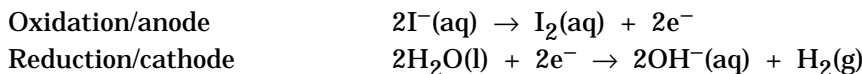
Potassium iodide, KI, several crystals  
 Petri dish  
 Mechanical pencil leads  
 2 Connecting wires with clips on both ends  
 9-V Transistor battery  
 1% Phenolphthalein (in 95% ethanol)  
 Plastic sheet  
 Overhead projector

*Safety*

Use normal laboratory safety precautions. The iodine produces a brown stain on skin and clothing.

*Procedure*

Place a Petri dish on a plastic sheet on the overhead projector. Half-fill the dish with water and dissolve about a match-head quantity of KI in it. Make carbon electrodes by using the connecting wires to connect a pencil lead to each battery electrode. Arrange the two pencil leads so that only carbon is in contact with the KI solution in the Petri dish. Add one drop of phenolphthalein. Observe the two electrodes for evidence of a reaction.

*Reactions*

Phenolphthalein turns pink as  $\text{OH}^-$  collects near the cathode. The region near the anode turns brown as  $\text{I}_2$  collects.

*NOTE: You can also make an electrolysis apparatus for this reaction from a 9-V battery, two pencils, two connecting wires, a 2-hole stopper, and some tape. Sharpen two pencils and push them through the stopper. Near the erasers, carve the wood away until you can make contact with the "lead" inside. (You needn't carve wood from all around the pencil to do this.) Tape the battery to the pencils. Connect each battery terminal to the pencil "lead" near the eraser via connecting wires.*



*This system is much less fragile than the one involving two mechanical pencil leads, although it does take more time to make. You can attach the device to a ring stand with a universal clamp on the stopper to avoid holding it during the reaction. You can show the presence of  $I_2$  by squirting a little starch solution near the anode to show the blue-black color created by the starch test. (First demonstrate the starch test to your students with a known iodine sample.) If you use the pencil lead electrolysis apparatus with a solution of  $SnCl_2$ , you can show students Sn needles that “grow” on the cathode.*

## **Demonstration 5: Electroplating Copper**

### *Purpose*

To demonstrate industrial uses of electrolysis.

### *Materials*

Copper strip, 1 cm x 5 cm  
Stainless steel object (spoon, paper clip, knife, *etc.*)  
18 M Sulfuric acid,  $H_2SO_4$ , 15 mL  
Beaker, 250-mL  
6 M Nitric acid,  $HNO_3$ , 50 mL  
1.5-V Battery  
2 Connecting wires with clips  
Copper(II) sulfate,  $CuSO_4$ , saturated solution, 200 mL

### *Safety*

Neutralize the acidic copper sulfate solution with an appropriate bicarbonate such as baking soda or a base like soda lime before disposing in accordance with local regulations. Dilute nitric acid can be flushed down the drain with ample water. Be careful to avoid contact with either acid. Copper salt solutions are toxic.

### *Procedure*

Clean the copper strip by dipping it into a beaker of dilute nitric acid and washing it thoroughly. Prepare a plating solution by carefully adding 15 mL concentrated sulfuric acid *slowly with continuous stirring* to about 200 mL saturated copper(II) sulfate solution. Attach the copper strip to the positive battery terminal using a connecting wire. Attach the object to be plated to the negative battery terminal with the other connecting wire. Place the object to be plated and the copper strip in the solution. Rotate frequently for even coating. After 3-5 min, remove the plated object and observe.

### *Reactions*

Oxidation/anode:  $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^{-}$   
Reduction/cathode:  $Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$

*NOTE: An interesting reversal of the reaction can be demonstrated by switching the battery leads.*

## **Demonstration 6: Making a Simple Battery: The Gerber Cell**

### *Purpose*

To demonstrate the ability of chemical reactions to produce electricity.

### *Materials*

Large baby food jar  
Dialysis tubing, 2-cm x 15-cm

- 1-Hole rubber stopper to fit jar
- 1 Strip each of copper and magnesium metal almost as long as the jar is deep and about half as wide as the dialysis tubing
- Voltmeter
- 2 Connecting wires with clips on the end
- 0.5 M Copper(II) sulfate,  $\text{CuSO}_4$ , 100 mL (12.5 g  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$  per 100 mL solution)
- 0.5 M Sodium sulfate,  $\text{Na}_2\text{SO}_4$ , 100 mL (16.1 g  $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$  per 100 mL solution)
- 6 M Hydrochloric acid,  $\text{HCl}$ , 50 mL
- 6 M Nitric acid,  $\text{HNO}_3$ , 50 mL

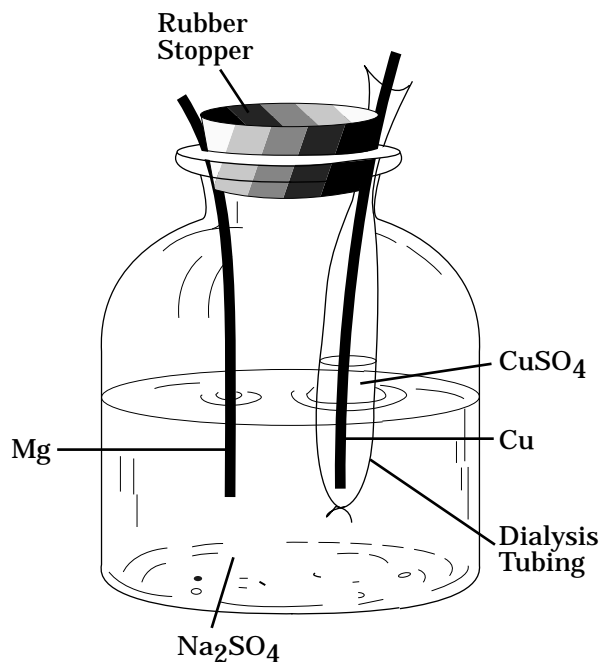
#### Safety

Use normal laboratory safety precautions and disposal procedures.

#### Procedure

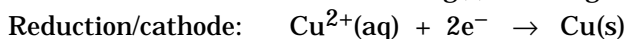
Clean the Cu strip by dipping it briefly in dilute  $\text{HNO}_3$  and then washing well with water. Clean the Mg strip by dipping it briefly in dilute  $\text{HCl}$  and then washing well with water. Fill the jar about 2/3 to 3/4 full of sodium sulfate solution. Wet the dialysis tubing and tie one end in a knot. Open the other end and fill to a depth slightly less than that of the jar with copper(II) sulfate solution. Place the copper strip in the dialysis tubing. Place the magnesium strip and the dialysis tubing in the jar. Then insert the stopper to hold them in place. Observe any reaction that may occur (see Figure 4).

**Figure 4. Simple battery apparatus.**



Attach the ends of the Cu and Mg strips to the voltmeter with the connecting wires.

#### Reactions



*NOTE: If you make six of these cells and hook them in series (anode to cathode), you can produce enough electric potential to operate a small radio that normally uses a 9-V battery.*

#### Reference

Summerlin and Ealy. (1985). *Chemical demonstrations, Vol. 1.* Washington, DC: American Chemical Society.

## Key Questions

1. What is the role of electrons in oxidation-reduction reactions? *[Electrons are the exchange particles of oxidation-reduction, moving from the oxidized to reduced species.]*
2. How can chemical reactions be used to produce electricity? *[If the half cells of a redox reaction are physically separated but connected by a salt bridge (to complete the internal circuit), electrons will flow through an external circuit of wire. The anion and electron movements are in opposite directions.]*

## GROUP AND DISCUSSION ACTIVITIES



3. How can electricity be used to drive a chemical reaction? *[If a source of electric potential (such as a battery) is attached to electrodes in a solution or ionic melt, the applied potential forces electrons through the system. The electrode connected to the negative terminal holds an excess of electrons, attracts cations, and reduces them. This electrode is the cathode. Simultaneously, oxidation is driven at the anode as anions are attracted, lose their excess electrons to the electrode, and complete the electrical circuit. Electrolysis and electroplating processes rely on such reactions.]*
4. How do materials compare in their relative ability to accept or donate electrons? *[Materials tend to be classified as electron donors if they are metallic and receivers if they are nonmetallic. However, any particular reaction depends on the relative tendencies of the materials. Thus we rank materials according to their reduction potential. Even though hydrogen in contact with hydrogen ions is assigned a half-cell potential of zero volts, it (like other half cells) can either involve oxidation or reduction. The process observed depends on the nature of the other reactant. If it has a higher reduction potential, hydrogen molecules will be oxidized; if not, hydrogen ions will be reduced.]*
5. How are electrochemical processes used in business and industry? *[As sources of energy (electrochemical cells) and to drive chemical processes (electrolytic cells).]*
6. How are electrochemical devices used in everyday life? *[Applications of electrochemical cells and batteries (groups of cells in series) are myriad. Some examples include flashlights, portable appliances, pacemakers, automobile cranking systems, and submarine propulsion units.]*

### **Counterintuitive Examples/Discrepant Events**

#### **1. Iron Can Be Protected from Corroding by Attaching It to Zinc**

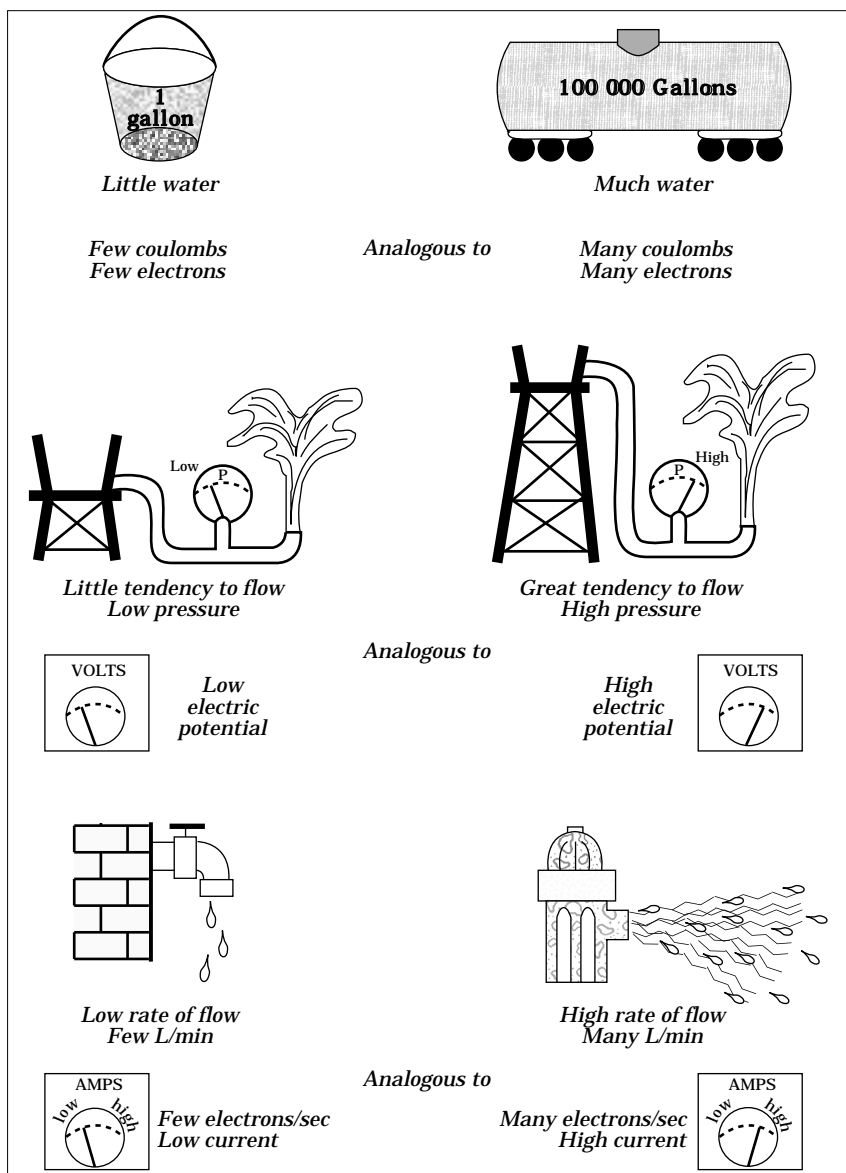
Iron is frequently used in construction and will corrode (oxidize) unless it is protected. One way to protect iron is to paint it. However, in many applications paint alone will not work because it is easily chipped or scratched. An effective way to protect iron is to attach a piece of zinc to it. This is frequently done to protect buried fuel tanks and hulls of ships. Zinc is easier to oxidize than iron and will lose its electrons (corrode) instead of iron. In such an application, zinc is a sacrificial cathode.

Have students observe the corrosion of 6-penny iron nails. Start with two identical 6-penny nails. Clean the surfaces of both nails with a piece of sandpaper. Wrap one nail with 10 cm of zinc wire, making sure to establish good physical contact between the two metals. Immerse nails in a beaker containing a 10% sodium chloride solution. Observe the nails over several days. The bare nail will show signs of corrosion while the one wrapped in zinc will be protected from corrosion. (Students should recognize that galvanized construction materials are made of iron protected by zinc.) This activity is useful for initiating discussion about electrochemistry prior to Activity 1 in this module.

### **Analogies and Metaphors**

1. Imagine an oxidation-reduction reaction as involving a hungry atom (draw a circle with chomping teeth in open mouth) eating an electron being fed to it by a nonhungry atom (closed mouth.) A voltaic cell just does the same thing, but the “hungry atom” has to be fed through a tube.

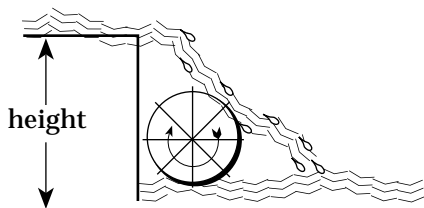
2. Use water analogies for electron transport.



**Figure 5. Electron transport/water analogies.**

3. Imagine that the quantity of work that can be performed by a galvanic cell is similar to the work obtainable from a water wheel powered by falling water.

The height of the waterfall is analogous to the cell potential. The higher the waterfall, the more potential energy it has. In the case of a voltaic cell, the higher the cell potential, the more the “driving force” for an electron. The quantity of water flowing over the waterfall is analogous to the current generated by the voltaic cell. The quantity of work that can be performed by a water wheel (equivalent to the voltaic cell) depends on both how much water flows each second (equivalent to current) and the height that water falls (equivalent to the electric potential.)



**Figure 6. Water wheel analogy.**



## Other

### Memory Aids: Mnemonics, *etc.*

1. Use a cat picture with + signs for eyes to remember Cation is positive. Or use cat paw with + sign in the pad print to remind students that cations are “pawsitive.”
2. LEO the lion says GER. LEO = Loss of Electrons is Oxidation. GER = Gain of Electrons is Reduction.
3. Cathode and Reduction both start with consonants; Anode and Oxidation both start with vowels.
4. “An Ox and Red Cat” (ANode involves OXidation and REDuction involves CATHode)
5. OIL RIG = *Oxidation Is Loss (of electrons), Reduction Is Gain (of electrons)*

## TIPS FOR THE TEACHER

### Language of Electrochemistry

**anode** electrode at which an oxidation half-reaction occurs in an electrochemical cell

**battery** voltaic cell used to produce electricity through a chemical reaction

**cathode** electrode at which a reduction half-reaction occurs in an electrochemical cell

**cell potential** potential difference in volts of an electrochemical cell

**cell voltage** synonym for cell potential

**corrosion** oxidation of a metal by the action of air, water, and/or salt solutions

**electric potential** synonym for cell potential

**electrochemical cell** device in which electrons of a redox reaction pass through an electrical circuit

**electrolytic cell** electrochemical cell in which a nonspontaneous reaction is carried out by electrolysis

**electromotive force (EMF)** synonym for cell potential

**galvanic cell** electrochemical cell in which a spontaneous chemical reaction produces electricity

**half-cell** combination of reduced and oxidized forms of a given species upon which a redox equilibrium is established

**nonspontaneous** type of reaction when cell potential is negative

**salt bridge** device used to join two half-cells and containing a salt solution that permits the flow of ions between two half-cells

**spontaneous** type of reaction when cell potential is positive

**voltaic cell** synonym for galvanic cell

### Common Student Misconceptions

1. “Electrons can flow through solutions.”

In “conduction of electricity” through solutions, electrons themselves do *not* pass through the solutions. Rather, charge balance is maintained in the solution by movement of cations and anions toward the electrodes where charge transfer takes place at the solution interface.

2. **“Water is a good conductor of electricity.”**

Water is a very poor “conductor” of electricity. (Recall that the ionization constant for water is very small.) The reason it is dangerous to insert a light bulb while standing in a puddle of water is that water is a great solvent for ionic compounds. Tap water and fresh water typically contain dissolved ions in sufficient concentrations to enable the solution to be conductive. However, ions in solution carry the charge and are thereby responsible for the current, not the water itself.

3. **“Only batteries are electrochemical cells.”**

Both voltaic and electrolytic cells are electrochemical cells. Voltaic cells produce electrical energy from differences in chemical potential energy. Electrolytic cells use electrical energy to produce products of higher chemical potential energy.

### Problem Solving

1. Half-cell reduction potentials and voltaic cell potentials are intensive properties like density and temperature. Therefore, they are independent of the *quantity* of material present. That is the reason why multiplying half-reactions by a constant in balancing a redox equation does not affect the value of the potential for the reaction. Because the half-cell potential is an intensive property, it does not change when the amount of reacting material is changed. This is different from Hess’s Law calculations and other calculations of thermodynamic quantities such as enthalpy, entropy, and free energy. As extensive properties, they do depend on the amount of substance present.
2. The numeric value for an oxidation potential is the negative of the reduction potential. When you reverse the equation for a reduction reaction, it becomes an oxidation equation and the direction (*i.e.*, the sign) of the potential changes.
3. When obtaining a reduction potential value in a standard table of half-reactions, the equation of interest must be identical to the one represented in the table. This includes phase (solid, liquid, gas, aqueous solution, *etc.*) as well as cases in which a precipitate might be involved such as the reduction of  $\text{AgCl(s)}$  to  $\text{Ag(s)}$  and  $\text{Cl}^{-}\text{(aq)}$ . Often, standard tables contain several equations that are very similar but have different half-cell potentials. Also, be aware that the reduction potentials in standard tables are calculated from thermodynamic quantities. Actual values measured in the laboratory may vary from these standard values. If you are interested in reasons, refer to the concept of *overvoltage* in a college chemistry text.

**Michael Faraday (1791-1867).** Michael Faraday was one of ten children of a London blacksmith and his wife. He had little formal education and was originally apprenticed to a bookbinder when he was 14 years old. Faraday taught himself to read in the bookbinder’s shop and first learned about electricity, which was just being discovered during his teenage years, from an article in an encyclopedia brought to his employer for rebinding. Because electrical phenomena piqued his interest, he joined a group in London who attended lectures on science. He wrote complete notes on lectures he heard and had them bound in a book.

## HISTORY: ON THE HUMAN SIDE



In 1812, he attended lectures given by Humphry Davy at the Royal Institution. He was so interested in Davy's talk that he applied to Davy, the institute director, for a job as his assistant. He finally obtained the position after following the advice of friends and sending his bound notes to Davy. This marked the start of Faraday's illustrious scientific career and his prolific collaboration with Davy. He went on to become the principal lecturer at the Institution and did much to popularize science among the British upper class; his Christmas lectures on the chemistry of candles remains a classic. In 1825, he replaced Davy as director of the Institution and his reputation soon began to rival that of Davy.

Faraday's work in analytical chemistry resulted in the discovery of benzene and many compounds of chlorine and carbon. He was the first to liquefy several gases, including  $\text{CO}_2$ ,  $\text{H}_2\text{S}$ ,  $\text{HBr}$ , and  $\text{Cl}_2$ . However, he is most famous for his work with electricity and electrochemistry. He discovered the principles that underlie modern electrical generators when he observed that a current is induced in a coil of wire rotating in a magnetic field. His experimental and theoretical work in electricity and magnetism formed the basis of later developments by James Clerk Maxwell. Albert Einstein rated Faraday along with Newton, Galileo, and Maxwell as one of the greatest physicists of all time. Faraday went on to prove that electricity generated in a magnetic field was identical to electricity produced by an electrochemical cell and performed many experiments measuring changes taking place in electrochemical cells. His work established two laws of electrochemistry still named after him in chemistry books: (1) The amount of material deposited in an electrochemical cell is proportional to the current passing through it, and (2) the amounts of substances deposited and/or dissolved in electrochemical cells are proportional to their molar masses.

A man of strong religious beliefs, Faraday tried to live a simple life, accepting rather reluctantly the many honors that came to him. His beliefs allowed him to solve without uncertainty a moral problem that still faces scientists. During the Crimean War between England and Russia in the 1850s, the British government asked him to head an investigation of the possibility of preparing large quantities of poison gas for battlefield use. Faraday refused to consider the idea, and nothing came of it at that time. Poisonous gas was eventually used in warfare during World War I in the early part of the 20th century.

**Charles Hall (1863 - 1914) and Paul Heroult (1863 - 1914).** The process used today for aluminum metal manufacture was invented independently at almost the same time by a young Frenchman, Paul Heroult, and a young American, Charles Hall. Hall invented the process while he was still an undergraduate at Oberlin College. He was inspired by a professor's remark that anyone inventing a cheap process for mass-producing aluminum would make a fortune. After graduation, Hall set up a laboratory in a woodshed—the 19th century equivalent to a garage—using homemade and borrowed equipment. After a year, he found that cryolite,  $\text{Na}_3\text{AlF}_6$ , would dissolve alumina (aluminum oxide) to give a conducting solution from which aluminum could be deposited by electrolysis. He used an iron frying pan as a container for the cryolite-alumina mixture, which he melted over a blacksmith's forge. The electric current came from electrochemical cells he made from jars used for canning fruit. (See the *Gerber Cell* demonstration described in this module.) As a result of the discovery made by Hall and Heroult, large-scale production of aluminum became economically feasible for the first time, and it became a common and familiar metal. Hall and Heroult share an interesting history. They were born in the same year, they discovered aluminum independently of each other in the same year, and they died in the same year.

## HUMOR: ON THE FUN SIDE

1. Message on a T-shirt: I've got potential—Let's realize it!
2. "The patriotic ion went to the pole and volted."  
(*CHEM 13 NEWS, September 1980, p. 13*)
3. Word Search (see *Appendix* for master copy)

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A S E G D I R B T L A S N P E Z T
C U B Z N J T F O V F T N D D J C
O H A V O L T K L G O J O B C S Z
R S T S I Q V V Q T Y H I G U G M
R N L I N E F Q W U T Y T U Y O G
O F U J A N B S C A I T C Q L W M
S N G J B G A I C S Q U A T O K X
I D F O M H N Y C Z R D E V E L S
O G R J W A F G B R O N R I Q K Z
N S S H V R H N E K J O F R N K F
A Q W L A I T N E T O P L L E C R
L P A R W S T I Z F N W A E H R A
Y G L Z L F X P H V E L H H V O M
    
```

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Driving force in a voltaic cell. (2 words)
2. Either of two parts of an oxidation-reduction reaction. (2 words)
3. Electrical quantity measured in amperes.
4. Unit of electrical potential.
5. Electrode at which reduction occurs.
6. Negative ion.
7. Type of cell where chemical energy is converted to electrical energy spontaneously by a redox reaction
8. Process by which metals are oxidized in the atmosphere
9. State in which NaCl must be in order to conduct electricity
10. Device containing an electrolyte that connects the two compartments of a voltaic cell. (2 words)

Answers: 1. CELL POTENTIAL 2. HALF REACTION 3. CURRENT 4. VOLT  
5. CATHODE 6. ANION 7. GALVANIC 8. CORROSION 9. MOLTEN  
10. SALT BRIDGE

4. Electrochemistry Crossword Puzzle (see *Appendix*)
5. See cartoons at end of module.



## MEDIA

1. *The World of Chemistry* videotape "Number 15: The Busy Electron," can be used by chemistry teachers at all levels. World of Chemistry Videocassettes. Annenberg/CPB Project, P.O. Box 1922, Santa Barbara, CA 93116-1922; (800) 532-7637; World of Chemistry Series, Atlantic Video, 150 South Gordon Street, Alexandria, VA 22304; (703) 823-2800 or QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224.
2. *Electrical Interactions in Chemistry* is a film/video available from Ward's Natural Science Establishment, Inc., P.O. Box 92912, Rochester, NY 14692-9012; (800) 962-2660.

3. *Doing Chemistry*, videodisc set available from the American Chemical Society, 1155-16th St., N.W., Washington, DC 20036; (202) 872-4382.

This three-disc set contains large numbers of laboratory and demonstration activities presented both as motion and still shots. It can be supported with Macintosh computer programs that provide access to HyperCard stacks for lesson planning and handout preparation, and that enable teachers to create their own videotapes. Over 700 pages of print materials are also available. In addition to the other fine activities, it includes two pertaining directly to electrochemistry: "DEMO E30: Daniell Cell;" and "DEMO E31: Galvanic Cell Based on Aluminum Oxidation."

4. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 263-2837 (voice) or (608) 262-0381 (FAX).
  - a. For the Apple II computer: AP 603 (includes Faraday 2, a simulation of Faraday's Laws in a laboratory experiment and Faraday Aid, a data analysis program used in conjunction with Faraday 2)
  - b. For IBM PCs and PC-compatibles: PC 3101 (Faraday 2 and Faraday Aid)
5. Videodisc published by *JCE: Software*, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

"Copper-Zinc Electrochemical Cell," a chapter on *The World of Chemistry: Selected Demonstrations and Animations*: Disc I (double sided, 60 min.), Special Issue 3.

6. Software published by *JCE: Software*, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

*Frost Diagrams*, by James P. Birk and Heidi Hocker. Vol. IV C, No. 1, for the Apple Macintosh.

## Equipment

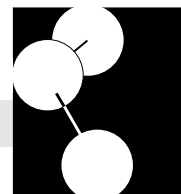
1. **Power Supplies for Electrolysis.** Try using AC adapters used for tape recorders or transistor radios. They are rated for different electric potentials and currents, depending on their applications. 9-V, 500-milliampere adaptors commonly available at Radio Shack stores work well for most applications. Also, 9-V radio batteries will drive most aqueous electrolysis processes of interest in first-year chemistry courses.
2. **Supporting Electrolytes.** Sodium sulfate, potassium nitrate, and sodium nitrate are all good, unreactive electrolytes to use in electrochemical cells. Chloride salts will react with any silver cations formed, and will also undergo oxidation to chlorine gas at the anode of electrolytic cells.
3. **Salt Bridges.** Salt bridges can be made with a variety of methods and materials. See the laboratory activity section of this module for instructions on making agar salt bridges. A cotton string or piece of filter paper soaked with an electrolyte solution will work as long as it remains wet.
4. **Voltmeters.** If you attempt to measure the potential of a voltaic cell directly with a voltmeter, be sure to use a meter with high impedance. Many inexpensive meters, especially those sold in science education supply catalogs, are low impedance and draw too much current from the cell to accurately determine the cell potential. Radio Shack, Sears, and others market relatively inexpensive multimeters of high impedance that can also measure current and resistance.
5. **Inert Electrodes.** Platinum wire makes excellent inert electrodes. However, it is very expensive. A recipe for preparing platinum electrodes is provided under *Instrumentation*. Graphite pencil “leads” also work well, although they sometimes offer too much resistance if the pencil lead diameter is too small. Thick pencils used by young children and thick square pencils used by artists are easier to handle. Just carve the wood away from the graphite at both ends. Paint on the pencil wood can be removed by sanding. Old dry cell batteries usually contain carbon electrodes as well. Because they are much thicker than pencil leads, they do not break as easily and offer much less resistance. To obtain them, dissect old batteries. (*NOTE: Avoid alkaline, lithium, NiCad, mercury, and other such batteries.*) The carbon electrodes in the center can be cleaned with any strong oxidizer that has a lower reduction potential than manganese dioxide/permanganate half-reaction. The electrodes can also be sanded clean.

## Instrumentation

Measuring the electric potential at which oxidation or reduction occurs under varying sets of conditions comprises a major subset of chemical instrumentation. You and your students can make simple electrochemical instruments for your laboratory. Directions for one such instrument are found in *Extensions and Projects*. Building any such instrument and making it work constitutes an excellent student project. The literature is full of references for building instruments. Specifically, look in *old* issues of the *Journal of Chemical Education*, “The Amateur Scientist” column in *Scientific American* when it was written by C. L. Strong, and *American Journal of Physics*. The electronics in those old instrument recipes are specified for vacuum tubes. These tubes can be duplicated by modern transistors and operational amplifiers. Find a ham radio operator, electrical engineer, electronic technician, or physics teacher to help you choose substitute parts. Instrument recipes as well as many other tips can be obtained at many science teacher meetings and conventions.

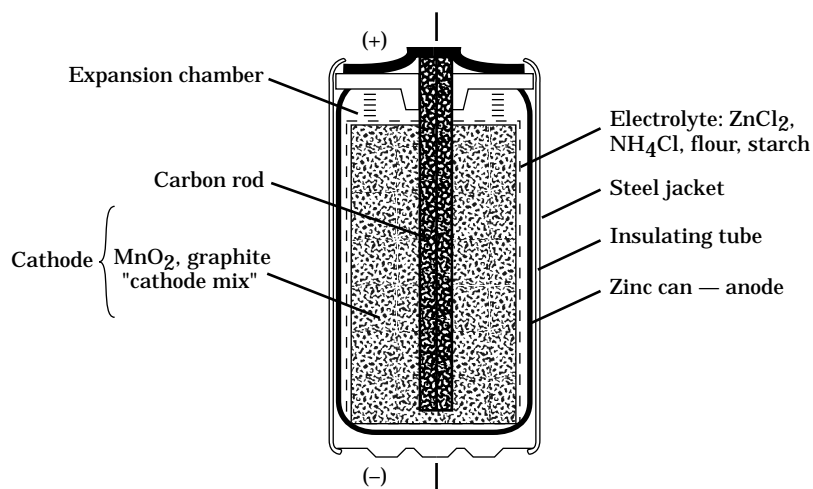
## EQUIPMENT AND INSTRUMENTATION

# Links/Connections



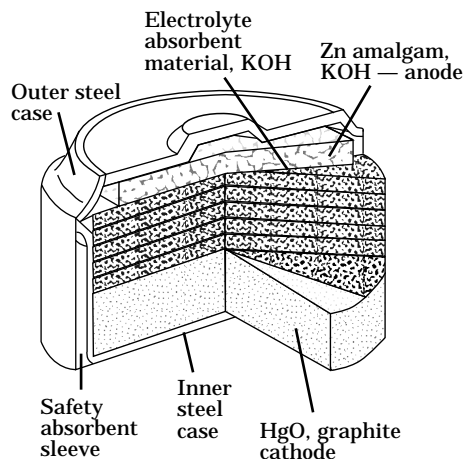
## WITHIN CHEMISTRY

1. **Stoichiometry and the Mole.** All electrochemical reactions are stoichiometric since an integral number of electrons must be transferred for each reacting atom, molecule, or ion. Many general chemistry texts include sections in which current in amperes and the time the current flows are used to calculate the total charge that has flowed through an electrochemical cell. This is, in turn, used to calculate moles of electrons transferred.
2. **Oxidation and Reduction: Commonly Used Batteries.** Dry cell batteries are a good example of the practical application of chemical principles. A standard design is to make the round “can” part of the battery of zinc metal and to serve as the anode. A carbon rod is inserted in the center to make contact with a mixture made of graphite carbon and manganese dioxide. This mixture serves as the cathode. An electrolyte/salt “bridge” mixture of moist zinc chloride, ammonium chloride, and nonconducting filler such as starch and flour surrounds the cathode mixture and connects it electrically to the zinc can. Not all features of the half-cell reactions are understood, although, in general, zinc loses electrons to become  $Zn^{2+}$  and manganese(IV) in manganese dioxide gains two electrons to become  $Mn^{2+}$  (see Figure 7).



**Figure 7. Dry cell battery.**

Alkaline batteries have numerous designs. A common “button”-type battery used for camera flashes and other small applications that demand high current uses mercury(II) oxide mixed with graphite as the cathode and powdered zinc amalgam (Zn and Hg) mixed with potassium hydroxide as the anode (see Figure 8). Because this type of battery contains mercury and mercury compounds, run-down batteries should be recycled to recover the mercury or treated to block entry of mercury compounds into the environment. Unfortunately, most people throw them unwittingly into the trash. A similar kind of battery uses silver oxide as the cathode and can be recycled for the silver.



**Figure 8. Alkaline “button”-type battery.**

The general characteristics of several varieties of dry cells are given in Figure 9:

Substance oxidized	Electrolyte mixture	Substance reduced	Generic name
Zinc	Zinc chloride Ammonium chloride Filler, Carbon black	Manganese dioxide	Dry cell
Zinc	Potassium hydroxide	Silver oxide	Silver cell
Zinc	Potassium hydroxide	Manganese dioxide	Alkaline cell
Zinc	Potassium hydroxide	Mercury(II) oxide	Mercury cell
Cadmium	Potassium hydroxide	Nickel oxide	Ni/Cad (Rechargeable)
Lithium	Lithium iodide	Iodine complex	Lithium

**Figure 9. Typical dry cells.**

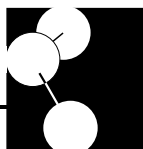
Details of lead-acid storage batteries for automobiles have been excluded from this module because explanations of this type of battery are included in most standard texts.

### High-Tech and Advanced Batteries

For those who design batteries, a number of engineering performance parameters are important. Some of these are cost, performance at low or high temperatures, shelf life, lifetime in use, rechargeability, size, safety, energy or power density (quantity of current that can be delivered per unit time at a given electric potential), constant discharge electric potential (known as a “flat” discharge because when electric potential vs. time is graphed, the line is “flat” rather than negatively sloped), ease of construction, availability of materials (especially if the battery has military applications), and resistance of electrodes to corrosion and to forming an unreactive layer. Advances in batteries are typically made by improving materials from which electrodes and electrolytes (salt bridges) are constructed. One trick is to make a previously impractical material practical.

For example, sodium/sulfur batteries can potentially store about four times as much energy as the common lead-acid battery currently used in automobiles. A practical sodium/sulfur battery could make electric automobiles feasible. However, sodium/sulfur batteries operate at high temperatures and pressures. Therefore, the electrolyte material which separates them must be able to withstand those conditions while maintaining its ability to allow ions to migrate through it. Currently, researchers are attempting to develop a tough conductive ceramic by treating alumina with zirconium oxide doped with a small amount of rare earth oxide such as yttrium oxide. Other researchers are testing a sodium ion-containing glass made from oxides of silicon, sodium, aluminum, and zirconium.

Other systems under investigation as potential batteries for electric automobiles include zinc/bromine, lithium/iron disulfide, lithium/iron monosulfide, and aluminum/air batteries. Of those currently under development, only the aluminum/air battery has a theoretical energy density greater than gasoline,



and is the only battery that could power an automobile for more than 300 miles without a recharge—the electrical equivalent of filling the tank. (The Ideal School Supply Company, Oak Lawn, IL 60453 sells an aluminum/air battery kit for educational purposes.) Scientists and engineers are constantly seeking new materials with interesting properties to accompany suitably high or low half-reaction reduction potentials. Other methods for improving batteries can include finding a coating for electrodes that allows them to function while protecting them from corrosion or the formation of an inert layer; redesigning the shapes of electrodes and their relationship to each other to lower internal resistance and to make recharging easier and more efficient; and spinning an electrode mechanically so that corrosion or surface layer formation is uniform.

Because batteries have so many applications and because the commercial potential of electric automobiles will be tremendous if an appropriate battery can be developed, large companies and entrepreneurial small companies invest heavily in electrochemistry research. For more information on advanced batteries see Brodd (1988) and Rawls (1985).

### **Fuel Cells**

Fuel cells resemble voltaic cells in that electrochemical reactions occur in half cells connected by an external circuit. Batteries contain all the chemical reactants internally with only electricity flowing out (discharge) or in (during recharging). Fuel cells, however, constantly produce electricity but require reactants to be constantly fed in as “fuel.” One simple fuel cell involves hydrogen fed into one electrode, oxygen fed into the other, producing electron flow through the external circuit and water.

Fuel cells are potentially much more efficient sources of electricity than are steam turbine-driven power plants. (See *Links to Physics*.) Several companies have built or are working on large fuel cells. Westinghouse Corporation has built a 1.5-MW fuel cell that uses phosphoric acid as the electrolyte and hydrogen and oxygen as reactants. It is called a PAFC Generator, where PAFC stands for Phosphoric Acid Fuel Cell.

Although fuel cells show promise as a system for generating electricity in the future, a number of technical problems must be solved. Most fuel cells use hydrogen as a fuel. A fuel cell that worked well with methyl alcohol would be more desirable because methyl alcohol is derived from plants more easily and cheaply than hydrogen gas can currently be produced. Many fuel cell combinations react too slowly to supply sufficient current for industrial uses. Catalysts that can speed up the electrode reactions are needed. Many noble metal (*e.g.*, Pt, Au, Ag, *etc.*) compounds work as fuel cell catalysts, but are expensive. Recent research has focused on using biological molecules such as porphyrins and phthalocyanines in conjunction with noble metal catalysts to increase the efficiency and decrease the amount of noble metal required. Fuel cells work best at high temperatures, but high temperatures are not possible for many applications; a polymer that conducts hydrogen ions (protons) at room temperature would be especially attractive as a solid electrolyte for many fuel cells (see *Potential Projects*).

One interesting application of the fuel cell idea is to turn a chemical plant into a fuel cell; for example, plants that make ammonia and methyl alcohol can be modified to become electricity-generating fuel cells. The opinion of many scientists and engineers is that fuel-cell technology will be a very profitable enterprise in the next few decades. For more information, see the 1988-89 series of articles by Lindstrom.

### Electrochemistry and Dental Fillings

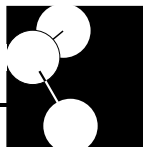
The material most commonly used for filling decaying teeth is called a dental amalgam. It is made by combining mercury with other metals such as silver or tin. If a person bites a piece of aluminum foil and the foil presses against a dental filling, a sharp, momentary pain often results. This result is due to a galvanic cell being created in the mouth. The aluminum serves as the anode and the filling is the cathode. Saliva acts as the electrolyte. A weak current flows between the electrodes and is detected by the sensitive nerve of the tooth as an unpleasant sensation.

Another type of reaction can occur if a filling makes contact with a gold inlay of a neighboring tooth. In this case, the dental filling made of mercury and tin acts as the anode and the gold inlay as the cathode. The dental filling starts to corrode. Prolonged corrosion will create the eventual need for a new filling. When this reaction occurs, tin(II) ions are released into the mouth, resulting in an unpleasant metallic taste. (See Transparency 5).

- 3. Kinetics/Rates of Reactions.** One feature of electrochemical reactions is that the rate of reaction, unlike reactions in solution, is limited by how fast reacting molecules or ions can travel through solution to get to the electrode surface. When the external circuit is first connected, reactions occur very quickly involving molecules that happen to be close to the electrode at the time. Therefore, there is essentially a “burst” of current for a voltaic cell or a “burst” of electroplating or electrolysis for an electrolysis cell. Once atoms and molecules near the electrode are used up, reacting species must travel to the electrodes. If those reacting species are charged, they will be attracted to the charged electrodes. If the reacting species are not charged, travel to the electrodes will be controlled by the rate of diffusion unless the reaction mixture is vigorously stirred. Even with vigorous stirring, the layer, just a few molecules thick, near the electrodes might not be stirred.

One way to stir the reaction at a molecular level if the species are charged is to place the electrochemical cell in a magnetic field, thereby forcing the ions to spiral toward the electrode, stirring the mixture on the way in. (This is analogous to the path taken by charged particles from the sun as they spiral into the earth at the poles under the influence of earth’s magnetic field.) A suggestion for a student research project based on this principle appears in *Extensions and Projects*. Since the current that a voltaic cell or fuel cell can deliver depends on the rate at which reactants reach the electrodes, this is a very important effect. Once reacting species reach the electrode, their reaction rate depends on all the features of reactions in solutions. Catalysts play a role in batteries at the electrode surface, lowering the activation energy for and thereby speeding up the electron transfer reactions. The chemical trick to placing catalysts on the electrode surface is, of course, to maintain or even improve half-reaction reduction potential characteristics. For a good description of how electrochemists modify the surface of electrodes, read Faulkner (1984).

- 4. Periodic Properties.** Although reduction potentials are more complex, they are similar in concept to the ideas of ionization potential and electronegativity, which are useful in discussions of periodicity. Where possible, periodic properties should be reinforced by comparing reduction potentials of elements within a single family.



## BETWEEN CHEMISTRY AND OTHER DISCIPLINES

5. **Producing Aluminum.** Aluminum is a very active metal and is therefore extracted from its compounds with difficulty. When aluminum metal first became available in the early 19th century, it was worth more than gold and other precious metals. All this changed with the advent of the Hall-Heroult electrolytic process, which is presently employed worldwide to produce commercial aluminum. The electrolysis is carried out in a carbon-lined vessel that serves as the cathode. The multiple anodes are made of graphite, and the electrolyte is purified bauxite ( $\text{Al}_2\text{O}_3$ ) dissolved in cryolite,  $\text{Na}_3\text{AlF}_6$ . It requires 0.75 lb of graphite to produce 1.0 lb of aluminum, which underscores the importance of aluminum recycling. A large aluminum plant produces  $4.0 \times 10^8$  kg ( $8.8 \times 10^8$  lb) of aluminum per year. It requires 200 kilowatt hours to produce 1000 kg ( $2.2 \times 10^3$  lb) of aluminum.
6. **Electroplating.** Electroplating is another important commercial process. Steel trash cans and steel siding are protected from rusting by electroplating a thin layer of zinc on the surface. A large silver anode, serving as a source of silver ions, is used to silver plate spoons, forks, and knives hung in a plating bath of silver cyanide. The plated objects serve as the cathodes in this electrolytic cell. A typical coating of silver has a thickness of about 0.1 mm.

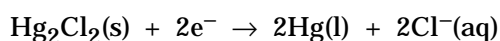
1. **Fuel Cells and Carnot Cycle Heat Engines—Relationship to Physics.** One fascination of fuel cells (see discussion below) is that they are not heat engines like current fossil or nuclear fuel-based electric energy plants and automobile engines. Heat engines are all limited in their efficiency by the Carnot cycle; the difference between the operating temperature of the engine (typically a steam turbine for electricity generation or the inside of the cylinder for autos) and the environment (typically the outside world) divided by the temperature of the environment gives the maximum theoretical efficiency of a heat engine if everything works perfectly and there is no friction.

For example, an automobile engine running at  $627^\circ\text{C}$  (9000 K) on a hot day of  $27^\circ\text{C}$  (300 K) would have an ideal maximum efficiency of 30%. Because real automobiles aren't frictionless or otherwise perfectly designed, the real fuel efficiency is less than 30%. This limit holds regardless of the kind of fuel used. By contrast, the efficiency of a fuel cell (and batteries, for that matter) are governed strictly by thermodynamic properties of the reactants—the electrochemical reaction's free energy change divided by its enthalpy change provides the theoretical efficiency of a fuel cell or battery.

2. **High-Tech Batteries/Fuel Cells and High-Temperature Super Conductors.** A challenge for many advanced battery systems researchers is finding materials that will work at high temperature and pressure. One approach is to make conductive ceramic electrolytes that will allow ions to move through them. Many rare earth and heavy metal oxides that are used to prepare high temperature superconductors (*e.g.*, yttrium oxide) are also used in preparing conductive ceramic electrolytes for advanced battery systems.
3. **Biology and Electrochemistry.** As mentioned in the introduction to the laboratory activities in this module, biochemical systems are electrochemical in nature. Two examples include photosynthesis and the respiratory chain of the mitochondrion. There are also electric eels, ion currents and action potentials in nerves, brain electricity as measured by the electroencephalogram, or EEG, and the ion currents in heart muscle measured by the electrocardiogram or EKG.

If your students are mathematically oriented, it might be interesting for them to calculate human power output based on the oxidation of glucose as a fuel as described by Chirpich (1975). An outstanding resource on bioelectricity and a number of other interesting topics is the book by Becker and Selden (1985).

4. **Modern Electrochemical Instrumentation.** Electrochemical instruments operate both in voltaic and electrolytic modes. Like all electrochemical cells, a minimum of two electrodes are employed. In some cases, a third half-cell is used as an internal reference cell; this half-cell determines the potential from which all voltages are measured. The most common reference cell is the calomel half-cell, sometimes referred to as the calomel electrode. The half-reaction for the calomel electrode is:



Electrochemical instrumentation is commonly used to determine the concentration of ionic species in aqueous solution. In the simplest case, an electrode is held at a constant potential vs. the reference. The potential is chosen so that an electrolysis reaction occurs that involves the species of unknown concentration. The number of electrons transferred at this electrode (the current) is related to the concentration of the species of interest. This technique is called polarography.

In addition, quantitative measurements can also be made in voltaic cells. The cell potential of a voltaic cell depends on the solution concentration of the species undergoing reaction at the electrode. (See a discussion of the Nernst equation in any standard college chemistry text for further information on this relationship.) Electrochemical measurements of pH and various other specific ion concentrations are examples of this voltaic technique.

## Personal

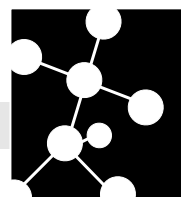
*Why an open flame is a hazard when charging an automobile battery.* The typical auto battery is a lead acid battery in which sulfuric acid is the electrolyte. When the battery is charged, hydrogen gas can form at the cathode. This is not a big problem if you use a “trickle” charge (*i.e.*, low-current battery charger) in a well-ventilated space. Hydrogen gas is produced at a slow rate; ventilation allows most of it to escape, thereby preventing an explosive mixture. However a “fast” charge (*i.e.*, high-current battery charger) produces hydrogen at a fairly rapid rate. In a poorly ventilated area, the concentration of hydrogen gas can increase to the point where explosive mixtures are formed.

## Community

1. **Field trips.** Electroplating companies, jewelry makers, electronic firms, chromeplating firms, battery companies
2. **Knowledgeable individuals.** Electroplating and electrolysis workers, individuals in battery industry, silversmiths
3. **Other community resources.** Local colleges/community colleges, any analytical chemistry laboratory using ion-specific electrodes

## TO THE CONTEMPORARY WORLD

# Extensions and Projects



## Extensions

1. **Electrochemical gadgets.** A number of suppliers provide gadgets that can be used to illustrate electrochemical principles. For example, toy and novelty stores often have clocks that can be powered by a copper/zinc voltaic cell. A “two potato” clock that uses potatoes as electrolytes and salt bridges is often available. A company that sells interesting gadgets like this is The Home Scientist/Jerryco, 601 Linden Place, Evanston, IL 60202. Another is Grand Junction, Ltd., Science and Surplus Center, 100 South Lynn Shores, Virginia Beach, VA 23452. See also information in the *Demonstrations* section.
2. **Using school-built batteries.** A small transistor radio or “walkman” can be powered by several voltaic cells connected in series. Any number of “homemade” cells can be used as long as their potentials add up to the proper electric potential. Fairly large electrodes should be used to pass sufficient current.
3. **Electroplating unusual objects.** Electroplating with copper is a fairly easy process. If the object to be plated is not conductive, first coat it with a film of conductive material. There are commercial sprays available for this purpose. Hook up an electrolysis bath using the object to be plated as the cathode and a piece of copper sheet as the anode. The electrolyte solution should be about 1 M copper(II) sulfate. *NOTE: Copper does not plate well on a zinc object.*

## Projects

Many chemistry teachers like to encourage students to complete research projects, either for their own edification or to enter into science fair and paper presentation research competitions. Below are a few suggestions for such projects.

1. **Gel electrophoresis.** (*This idea is by Sheila Kolb and Robert Roe, Jr., from the R. Arlotto et al. 1984 Dreyfus Institute Electrochemistry module.*) Electrophoresis is a method for separating mixtures that contain charges on molecules, or molecules with + and – charges like proteins, DNA, etc. The ions are attracted toward electrodes and forced to move through a medium such as gel. The combination of size, charge, and tendency to interact with the gel determine how fast various species travel and consequently how well and how quickly they can be separated from each other. Electrophoresis is not normally encountered in high school because of perceived expense and danger. Your local biology department may have an inexpensive unit. (Commercial firms use high electric potential.)

However, your students can make electrophoresis separators with 9-V transistor batteries, a short piece of clear plastic tubing such as Tygon, some grocery store gelatin, and a small syringe and needle. The gelatin is used to fill a length of tubing. The pH of the gelatin can be controlled by gelling it with various buffer solutions made from acetate, phosphate, and ammonium salts instead of water. Dissolve about 10 g of gelatin in about 200 mL of water or buffer solution (a 250-mL beaker nearly full) in a hot water bath (a 400-mL beaker with enough water to form a “jacket” around the 250-mL beaker). Fill the tubing with gelatin and stopper at both ends until the gelatin cools and sets. Inject a small sample of the mixture to be separated through the tubing into the middle of the gelatin using the needle and syringe. A mixture of bromocresol green and methyl red is a good trial mixture for practice. Attach wires to the batteries and stick the ends of the wires in the ends of the gelatin.

Here are some questions that can be answered with this system.

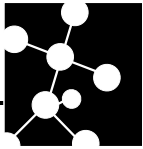
1. What mixtures of ions can be separated? Can other mixtures such as amino acids or DNA fragments be separated like this?
  2. What is the influence of pH on the separation?
  3. What is the influence of concentration of ions in the water used for making the gelatin?
  4. Does it matter how close together the electrodes are? Does the shape of the electrodes matter? For example, if you attached the wires to little flat electrodes that contacted the gelatin across the diameter of the tube, would the results be different?
2. **Making a Polymer that Conducts Hydronium Ions at Room Temperature.** Such a polymer was identified in *Links and Connections* as important for the economic development of fuel cells. Worthy (1985) reported news about such a polymer made from polyvinyl alcohol and phosphoric acid. One student might enjoy altering important variables such as reactant concentrations and preparation temperature to prepare such materials. The polymer is apparently made by mixing the two ingredients and evaporating the water. Details on testing the proton conductivity were not given in the article and might prove to be a challenge to a high school student. Before making a polymer, student(s) should look up the article and locate the research article by the original scientists by referring to Science Citation Index.
  3. **Effects of Magnetic Fields on Electrochemical Reactions.** Ions moving in electrochemical cells are just moving charged particles subject to all the laws of physics including the one describing the force a magnetic field places on them. The rates and perhaps other features of electrochemical reactions involving ions should therefore be affected by placing them in magnetic fields. Test-tube reactors placed between strong magnets such as those available from many science supply companies are good systems for testing these effects. Students can also build Helmholtz coils and, with a variable electric potential source, can even subject the electrochemical cells to magnetic fields of different strengths. If you are unfamiliar with Helmholtz coils, consult a high school or college physics teacher. One easy system to measure is electrolysis of water. How does the presence of the magnetic field affect the rate at which gas collects in a tube?

### Polarity Tester

A polarity tester is constructed of materials listed in Figure 10. It is used to determine which electrode in an electrochemical cell is the cathode and which is the anode.

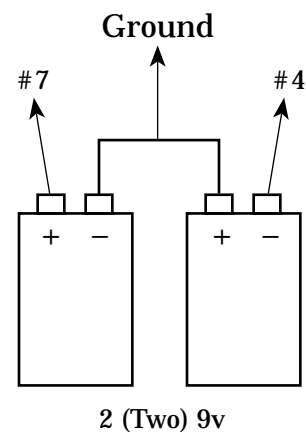
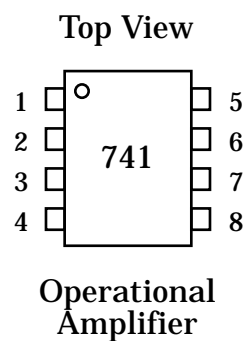
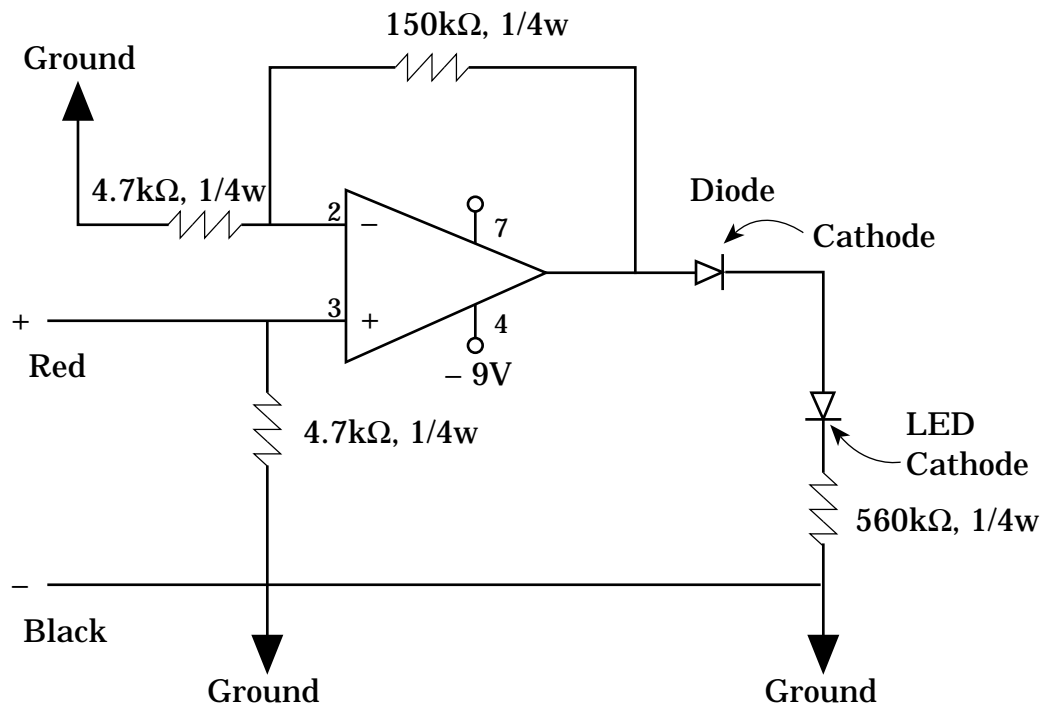
Quantity	Description	Radio Shack #
1	Operational amplifier (Op Amp) Type 741	276-007
1	Light-emitting diode	276-226
1	Signal diode IN 914/4148	276-1620
2	9 V Battery snap connector	270-325
2	Mini gator clips	270-374
1	Op amp socket	276-1995
2	4.7 k-ohm 1/4 W resistor	271-1330
1	150 k-ohm 1/4 W resistor	271-047
1	560 ohm 1/4 W resistor	271-020
1	Circuit board, 24 gauge black and red wire	276-159A

**Figure 10. Materials for a polarity tester.**



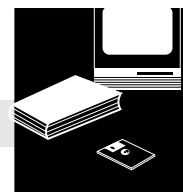
Use Figure 11 as a guide in construction. Care is needed to not overheat any of the elements and to not cause a short by using too much solder in connections. Students interested in electronics may be induced to construct the polarity testers for you. If your school has an electronics course or a vocational education department, they might also help you build one. Once you have built a polarity tester or two, they take about 15-20 min to build. They are quite durable and should last many years with proper care.

Inexpensive multimeters of high impedance can also be purchased at many electronics stores (such as Radio Shack). These are good alternatives if you do not have the time to make your own polarity testers.



**Figure 11. Diagram for polarity tester.**

# References



Module developed by William Becker, Beatrice E. Epperson, and William Lamb, the Oregon team.

Arlotto, R., Beach, D., Bent, H., Domencich, D., Guitierrez, N., Herron, D., John, E., McHugh, K., Paroni, G., Schultz, E., Sievert, M., Smith, J., Stone, H., Wong, G., and Zmaczynski, R. (1984). *Electrochemistry*. (Available from Woodrow Wilson National Fellowship Foundation, P.O. Box 642, Princeton, NJ 08542).

A booklet including curriculum materials, enrichment materials, and laboratory activities developed by master chemistry teachers attending the Dreyfus Foundation Institute for High School Chemistry at Princeton University.

Becker, R. O., and Selden, G. (1985). *The body electric: Electromagnetism and the foundation of life*. New York, NY: Morrow.

This book details Becker *et al.*'s research on the biological effects of electricity, magnetism, and electromagnetic fields.

Brodd, R. J. (1985, October). Advanced batteries. *CHEMTECH*, 15(10), 612-621.

Article reviewing research and technology on advanced battery systems.

Carmichael, L. N., and Haines, D. F. (1987). *Laboratory chemistry* (Teacher Ed.). Columbus, OH: Merrill.

Chirpich, T. P. (1975). Electrochemistry in organisms: Electron flow and power output. *Journal of Chemical Education*, 52, 99-100.

A short article describing how to apply electrochemical calculations to biological systems.

Edge, R. D. (1984). *String and sticky tape experiments*. (Available from Department of Physics, University of South Carolina, Columbia, SC 29208).

A book of physics demonstrations published by the American Association of Physics Teachers. Many of the demonstrations are useful for chemistry concepts.

Faulkner, L. R. (1984, February 27). Chemical microstructures on electrodes. *Chemical and Engineering News*, 62(9), 28-45.

An article reviewing research on the effects of attaching various functional groups to electrode surfaces.

Freier, G. D., and Anderson, F. J. (1981). *Demonstration handbook for physics*. (Available from American Association of Physics Teachers, Graduate Physics Building, SUNY, Stony Brook, NY 11794).

A book of physics demonstrations published by the American Association of Physics Teachers that contain many chemistry concepts.

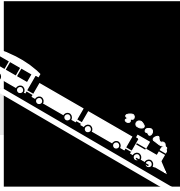
Hajian, H. G., and Pecsok, R. L. (1988). *Modern chemical technology* (Vol. 1). Englewood Cliffs, NJ: Prentice-Hall.

Kozawa, A., and Powers, R. A. (1972). Electrochemical reactions in batteries. *Journal of Chemical Education*, 49, 587-588.

An article proposing alternative modes for presenting electrochemical concepts to students.

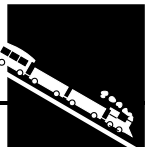


- Lauren, P. M. (1977). Sir Humphry Davy's battle with the sea. *Chemistry*, 50(7), 14-17.  
Historical account of Sir Humphry Davy, Director of the Royal Navy Institute in England in the early 1800s. Davy was the mentor of Michael Faraday.
- Lindstrom, O. (1988, August). That incredible fuel cell. Part 1. *CHEMTECH*, 18(8), 490-497.
- Lindstrom, O. (1988, September). Fuel cell power plants. Part 2. *CHEMTECH*, 18(9), 553-559.
- Lindstrom, O. (1988, November). Muscles, engines, and fuel cells. Part 3. *CHEMTECH*, 18(11), 686-693.
- Lindstrom, O. (1989, January). Fuel cell markets. Part 4. *CHEMTECH*, 19(1), 44-50.  
A series of articles describing research and development of fuel cells.
- Martin, C. R., and Freiser, H. (1980). Coated-wire ion selective electrodes and their application to the teaching laboratory. *Journal of Chemical Education*, 57, 512-514.  
An article that provides concepts, methods, and recipes for making ion selective electrodes and using them in laboratory activities.
- Mims, F. J. (1981, June). Experimenter's corner: Do-it-yourself batteries. *Popular Electronics*, 19(16), 84-87.  
An article that describes how to make batteries from readily available materials found around the home.
- Rawls, R. L. (1985, December 16). Materials developments highlight progress in batteries. *Chemical and Engineering News*, 63(50), 23-27.  
A news item about new materials used to develop better batteries.
- Rayovac Corporation. (1978). *Lithium batteries technical handbook*. (Available from Rayovac Technology Center, 630 Forward Drive, Madison, WI 53711).  
A handbook developed by Rayovac Corporation to assist the user of lithium batteries.
- Shugar, G. J., Shugar, R. A., Bauman, L., and Bauman, R. S. (1981). *Chemical technicians' ready reference handbook*. New York, NY: McGraw Hill.  
A standard reference listing recipes, techniques, and tricks of the trade—highly recommended for any chemistry teacher.
- Summerlin, L. R., Borgford, C. L., and Ealy, J. B. (1987). *Chemical demonstrations: A sourcebook for teachers* (Vol. 2). Washington, DC: American Chemical Society.  
A chemistry demonstration sourcebook that is highly recommended for all chemistry teachers.
- Wilbraham, A. C., Staley, D. D., Simpson, C. J., and Matta, M. S. (1987). *Addison-Wesley chemistry laboratory manual*. Menlo Park, CA: Addison-Wesley.
- Worthy, W. (1985, November 25). Polymer conducts protons at room temperature. *Chemical and Engineering News*, 63(47), 28-29.  
News item about a polymer that conducts an electric current using protons at room temperature.



# Appendix

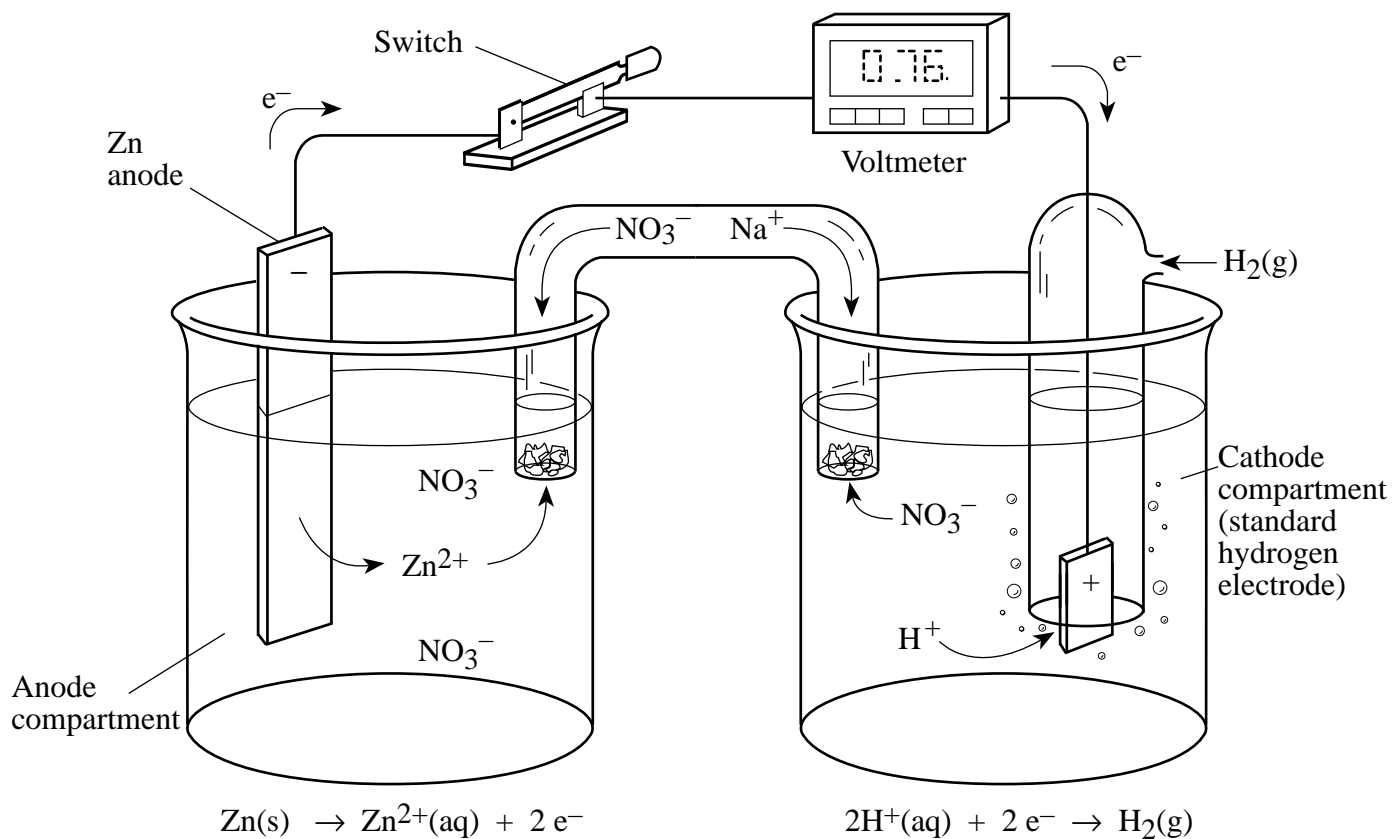
- **Transparency Masters**
  1. Voltaic Cell: Zinc/Hydrogen Electrodes
  2. Voltaic Cell: Zinc/Copper Electrodes
  3. Electrolysis of Molten Sodium Chloride
  4. Electrolysis of Water
  5. Electrochemistry and Dental Fillings
  6. Word Search
  7. Electrochemistry Crossword Puzzle
- **Humor**

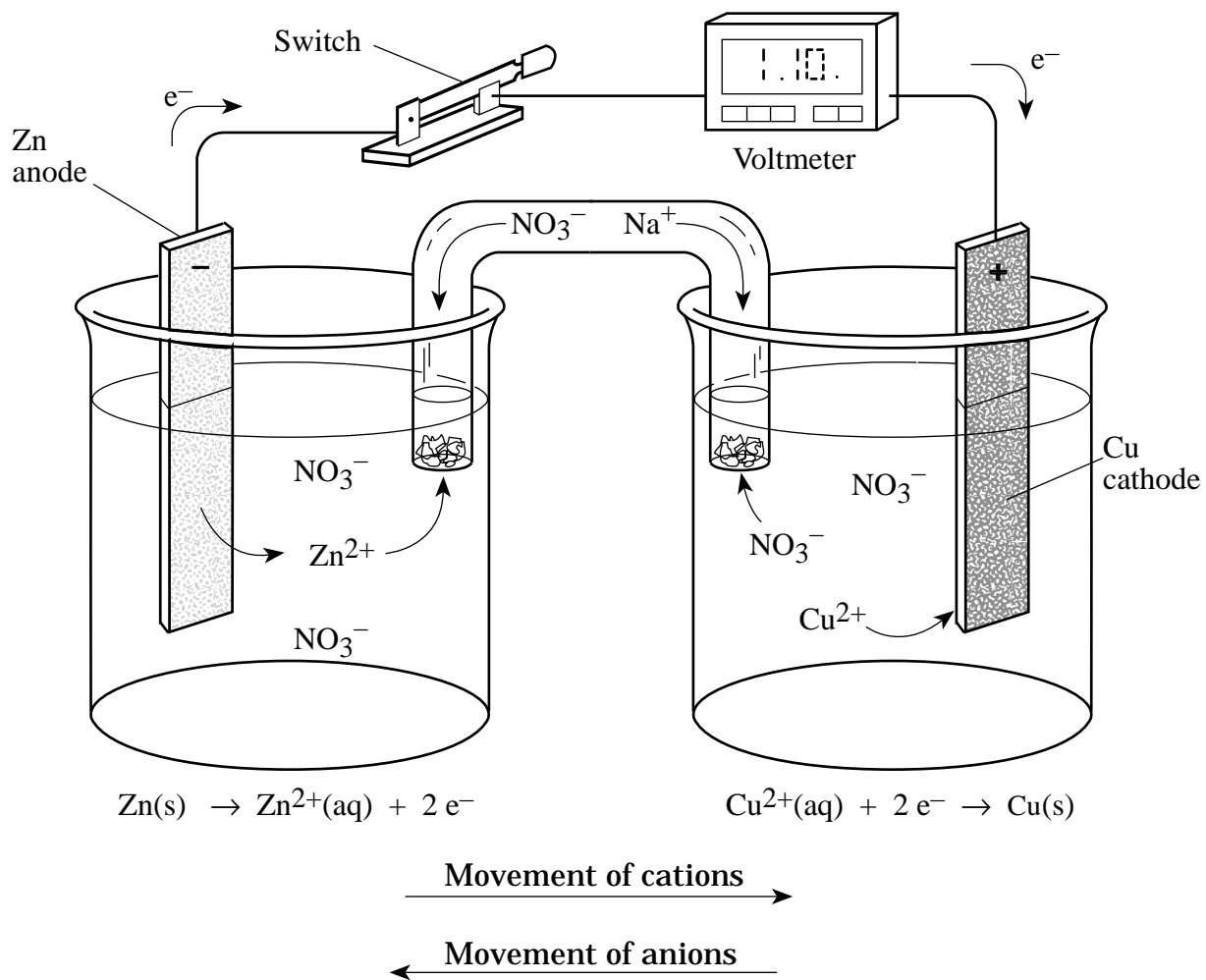


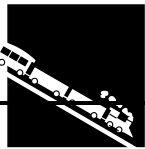
### Voltaic Cell: Zinc/Hydrogen Electrodes

$$E_{\text{cell}}^{\circ} = E_{\text{ox}}^{\circ} + E_{\text{red}}^{\circ}$$

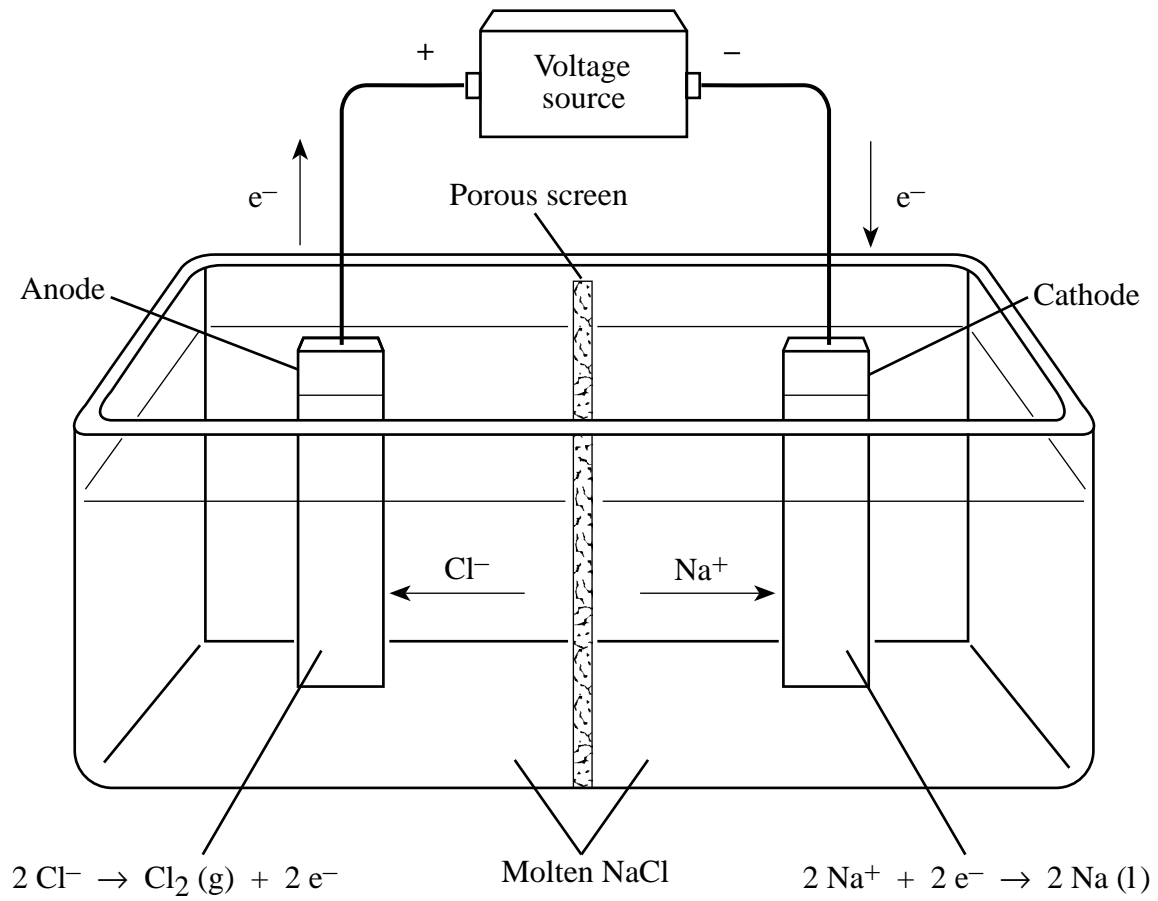
$$0.76\text{V} = E_{\text{ox}}^{\circ} + 0$$

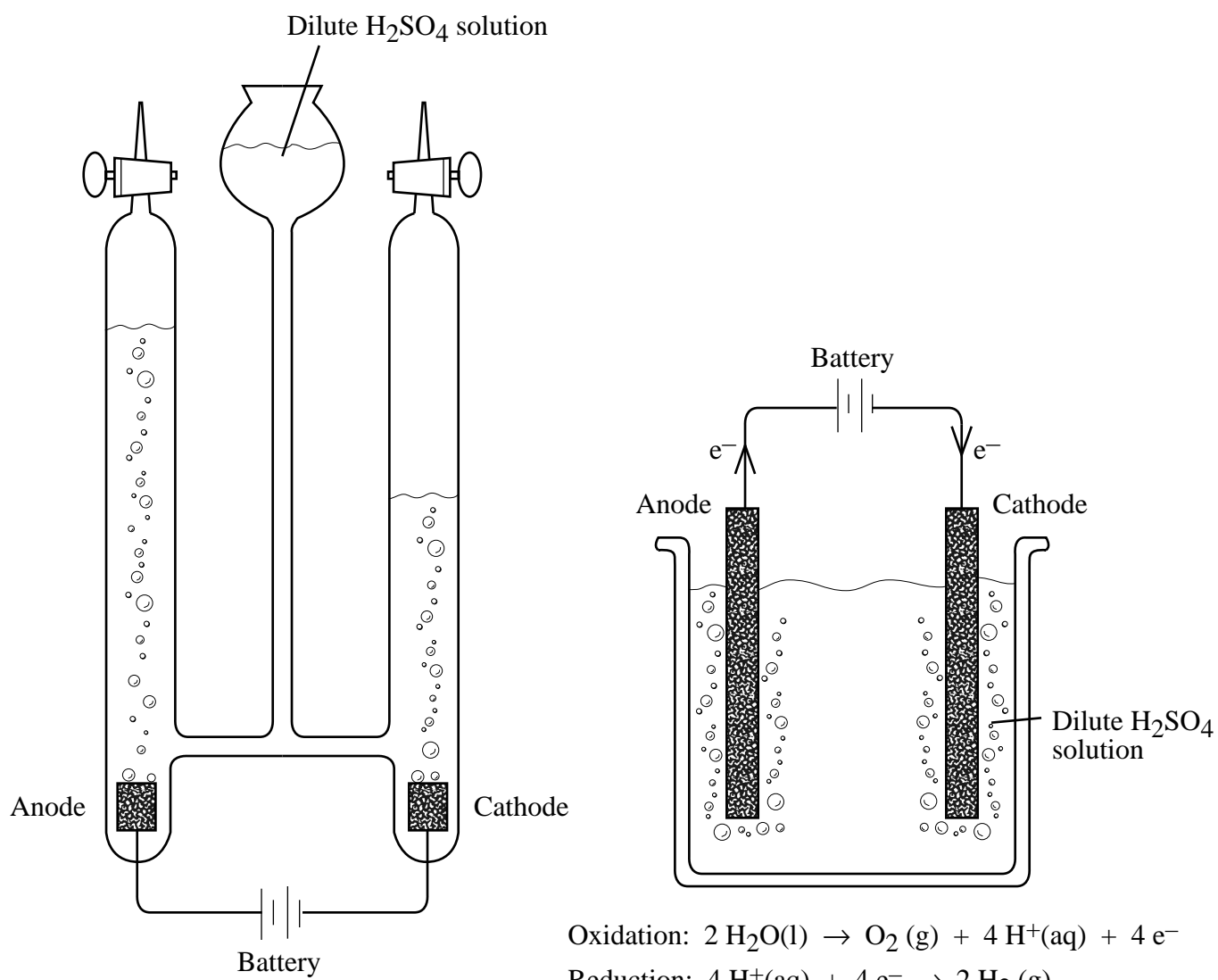


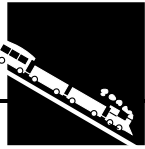
**Voltaic Cell: Zinc/Copper Electrodes**



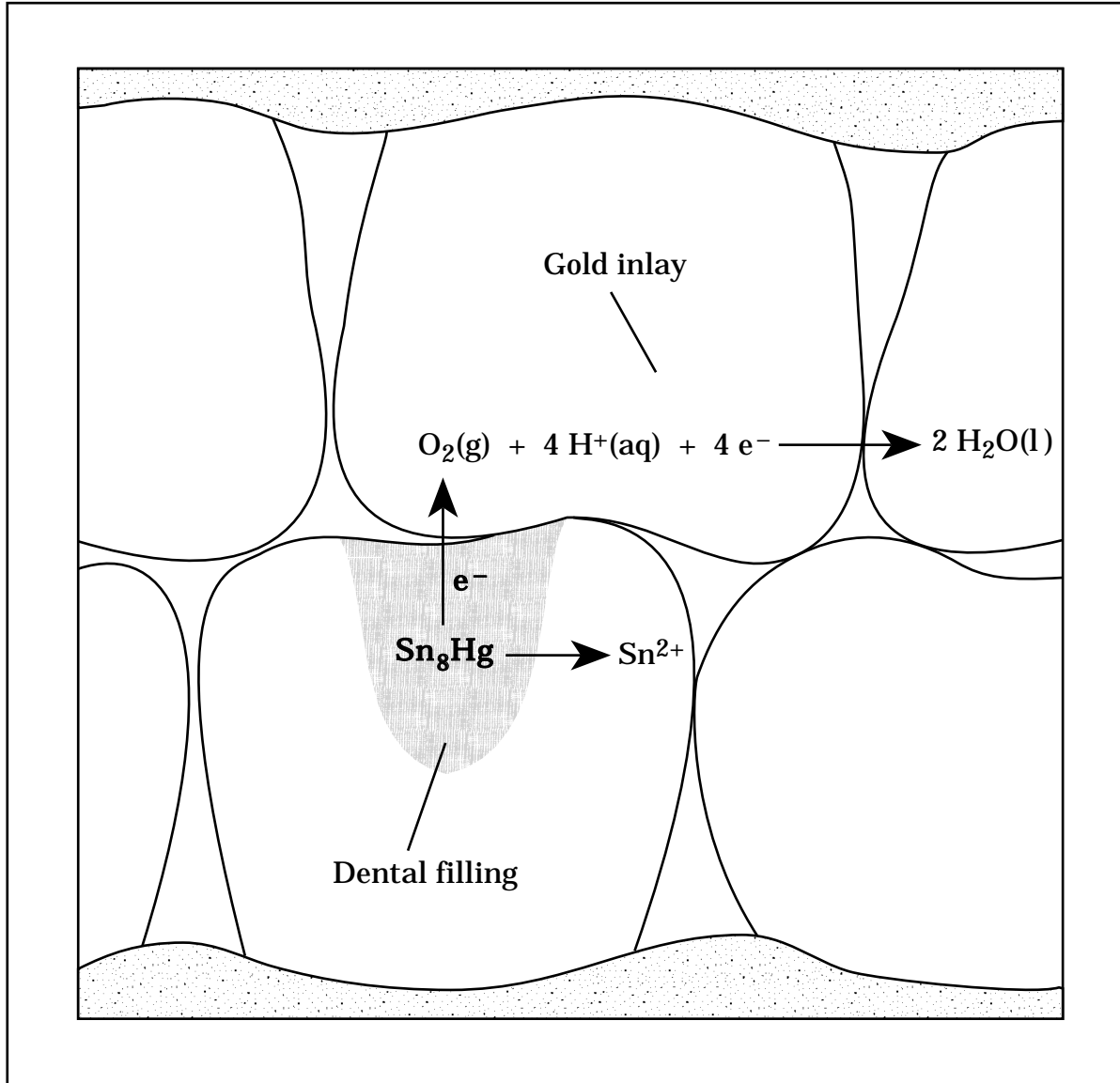
## Electrolysis of Molten Sodium Chloride



**Electrolysis of Water**



## Electrochemistry and Dental Fillings

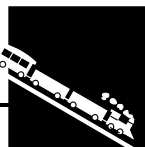


**Word Search**

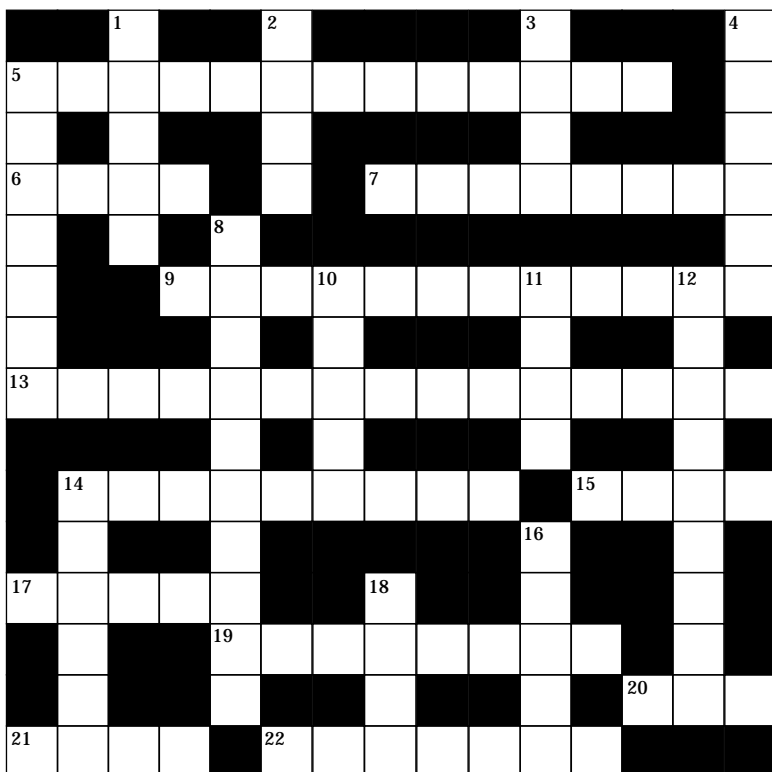
A S E G D I R B T L A S N P E Z T  
C U B Z N J T F O V F T N D D J C  
O H A V O L T K L G O J O B C S Z  
R S T S I Q V V Q T Y H I G U G M  
R N L I N E F Q W U T Y T U Y O G  
O F U J A N B S C A I T C Q L W M  
S N G J B G A I C S Q U A T O K X  
I D F O M H N Y C Z R D E V E L S  
O G R J W A F G B R O N R I Q K Z  
N S S H V R H N E K J O F R N K F  
A Q W L A I T N E T O P L L E C R  
L P A R W S T I Z F N W A E H R A  
Y G L Z L F X P H V E L H H V O M

Words about the concepts in this module can be obtained from the clues given.  
Find these words in the block of letters:

1. Driving force in a voltaic cell. (2 words)
2. Either of two parts of an oxidation-reduction reaction. (2 words)
3. Electrical quantity measured in amperes.
4. Unit of electrical potential.
5. Electrode at which reduction occurs.
6. Negative ion.
7. Type of cell where chemical energy is converted to electrical energy spontaneously by a redox reaction
8. Process by which metals are oxidized in the atmosphere
9. State in which NaCl must be in order to conduct electricity
10. Device containing an electrolyte that connects the two compartments of a voltaic cell. (2 words)



## Electrochemistry Crossword Puzzle



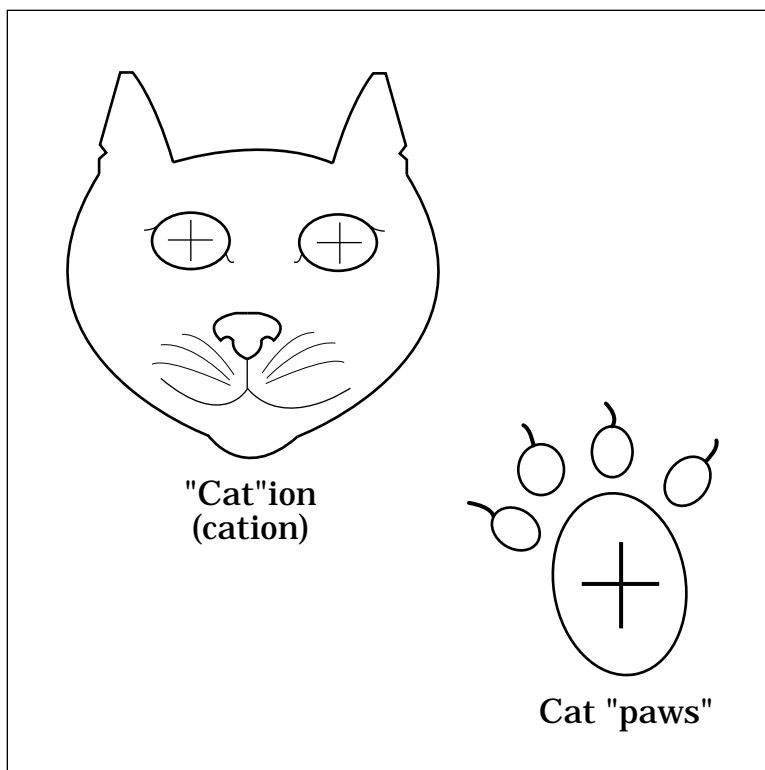
### Across

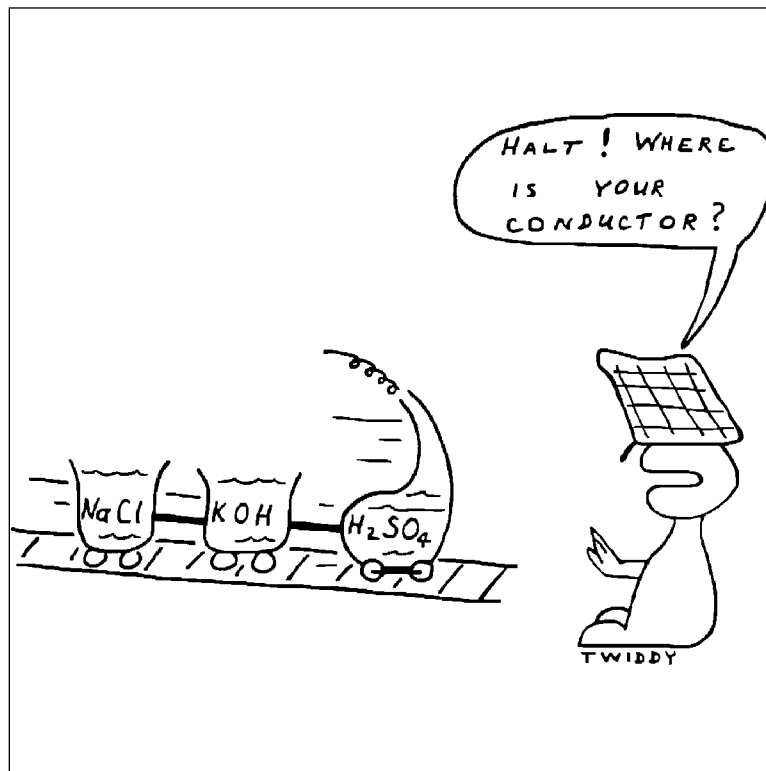
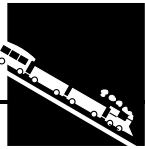
5. Driving force in a voltaic cell (2 words).
6. Fundamental quantity that must be measured in order to determine the rate of electron flow.
7. Substance added to fuel cells as fuel.
9. Either of two parts of an oxidation-reduction reaction (2 words).
13. Method of separating mixtures that contain charged molecular species.
14. Process by which metals are oxidized in the atmosphere.
15. Form of energy generated when electric current flows.
17. Word with ring denoting a support for laboratory devices.
19. Type of cell where chemical energy is converted to electrical energy spontaneously by a redox reaction.
20. Number of moles of electrons transferred when a mole of iron(II) is oxidized.
21. General term for a joule or coulomb, *e.g.*
22. Electrical quantity measured in amperes.

### Down

1. Laboratory holder for a test-tube.
2. Unit of electrical potential.
3. Element used in a Daniell Cell.
4. State in which pure NaCl must be in order to conduct electricity.
5. Electrode at which reduction occurs.
8. Device containing an electrolyte that connects the two compartments of a voltaic cell (2 words).
10. What an ion does when a voltaic cell discharges.
11. Center of a transformer or an apple.
12. Process whereby electrons are lost.
14. Positive ion.
16. Negative ion.
18. Word used to precede voltage to denote a condition where greater potential than calculated is needed to make a reaction proceed.

Down	Across
1. CLAMP	5. CELL POTENTIAL
2. VOLT	6. TIME
3. ZINC	7. REACTANT
4. MOLTEN	9. HALF REACTION
5. CATHODE	13. ELECTROPHORESIS
8. SALT BRIDGE	14. CORROSION
10. FLOWS	15. HEAT
11. CORE	17. STAND
12. OXIDATION	19. GALVANIC
14. CATION	20. ONE
16. ANION	21. UNIT
18. OVER	22. CURRENT

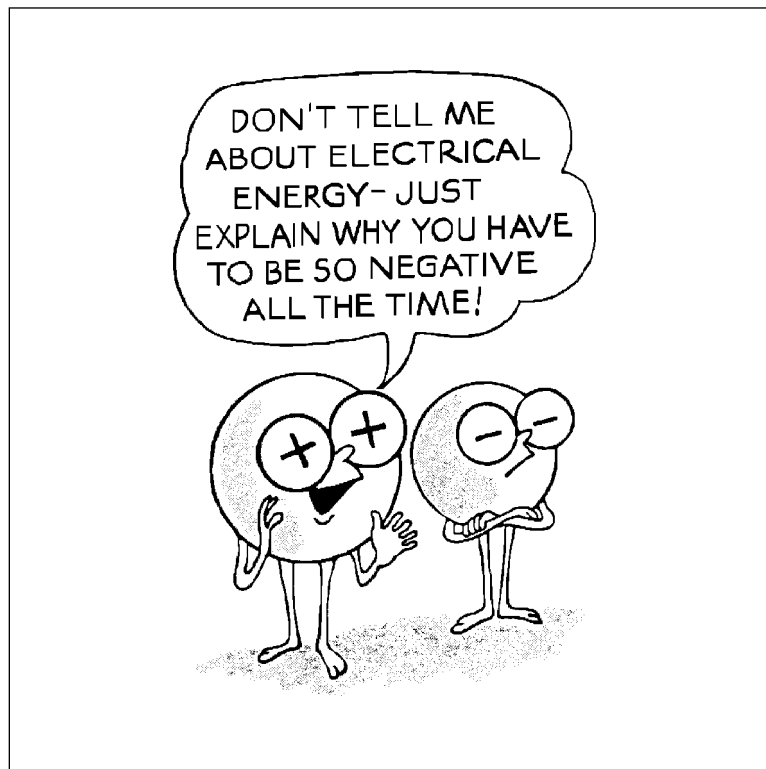
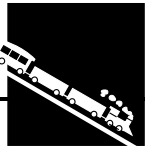




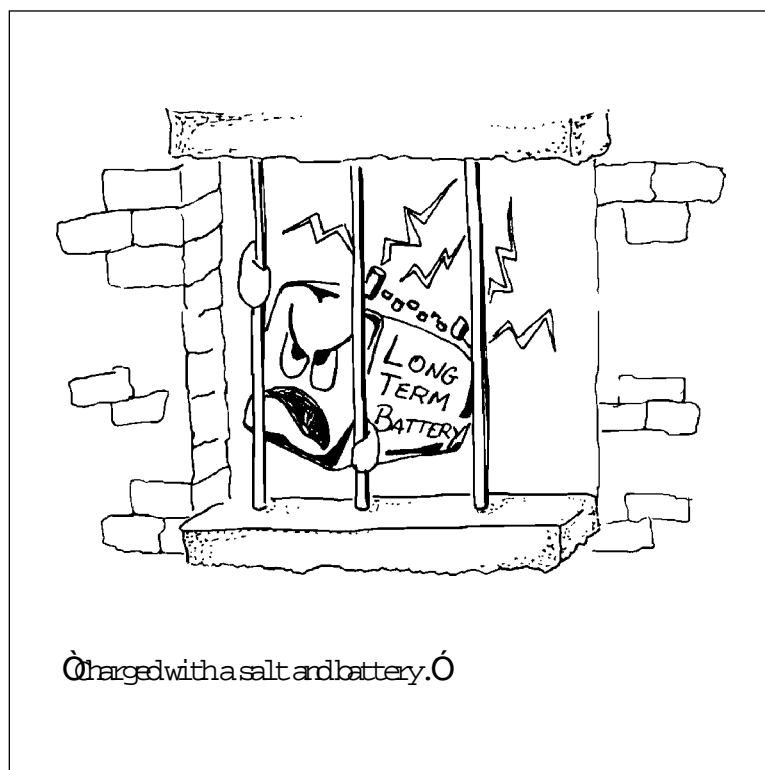
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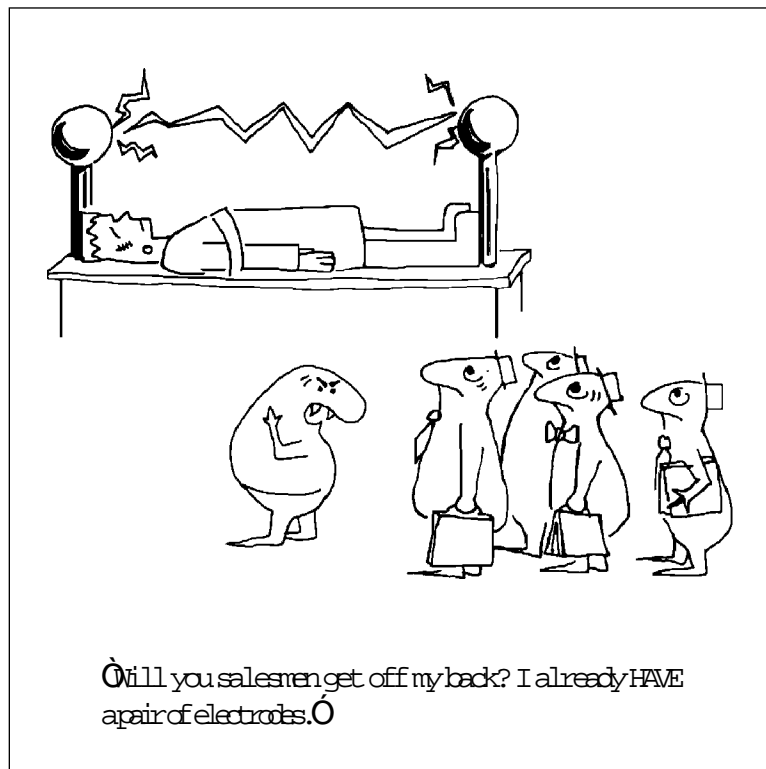
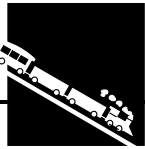
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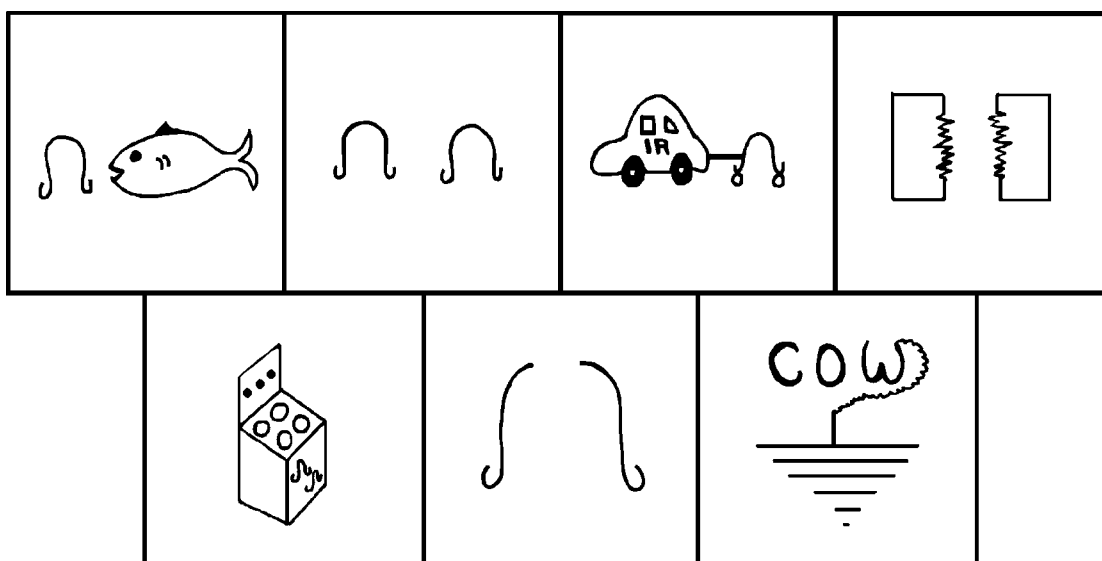


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CHEM 13 NEWS, March 1978, p. 1320. Reprinted with permission.

OH ME, OH MY, WHAT ARE THESE?



CHEM 13 NEWS, November 1980, p. 2. Reprinted with permission.

- The answers to the quiz are:
1. ohm eater
  2. ohm away from ohm
  3. voltswagen pulling a mobile ohm
  4. a pair of shorts
  5. ohm, ohm on the range
  6. product of a broken ohm
  7. ground beef