Measuring Cell Potentials

Introduction

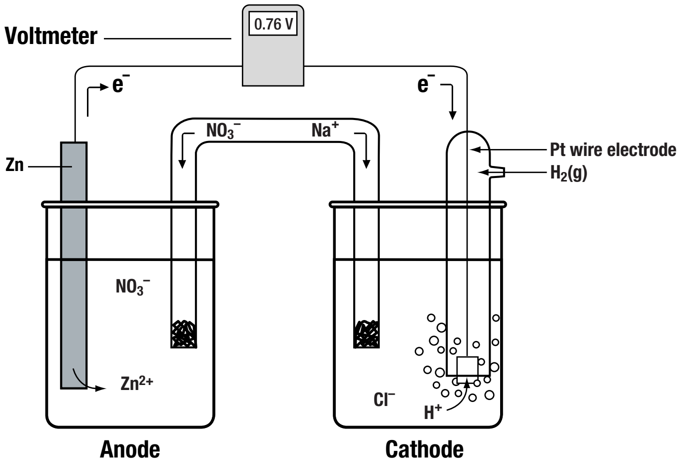
In an oxidation–reduction reaction, electrons flow from the substance that is oxidized, which loses electrons to the substance being reduced, which gains electrons. In a voltaic cell, the flow of electrons accompanying a spontaneous oxidation–reduction reaction occurs via an external pathway, and an electric current is produced. What factors determine the ability of a voltaic cell to produce electricity?

Concepts

- Oxidation–reduction
- Standard reduction potential
- Voltaic cell
- Activity series of metals

Background

The basic design of a voltaic cell is shown in Figure 1 for the net reaction of zinc and hydrochloric acid. The substances involved in each half-reaction are separated into two compartments connected by an external wire and salt bridge.



$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ $2H^+(aq) + 2e^- \rightarrow H_2(g)$

Figure 1. Components of an Electrochemical Cell.

Net Reaction: $Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$

Each half-reaction takes place at the surface of a metal plate or wire called an *electrode*. The electrode at which the oxidation occurs is called the *anode*, while the electrode at which the reduction occurs is called the *cathode*. Electrons flow spontaneously from the anode (the negative electrode) to the cathode (the positive electrode). Charge buildup at the electrodes is neutralized by connecting the half-cells internally by means of a *salt bridge*, a porous barrier containing sodium nitrate of another electrolyte. Dissolved ions flow through the salt bridge to either electrode, thus completing the electrical circuit.

The ability of a voltaic cell to produce an electric current is called the *cell potential* and is measured in volts. If the cell potential is large, there is a large "electromotive force" pushing or pulling electrons through the circuit from the anode to the cathode. The cell potential for a spontaneous chemical reaction in a voltaic cell is always positive. The *standard cell potential* ($E°_{cell}$) is defined as the maximum potential difference between the electrodes of an electrochemical cell under standard conditions—25 °C, 1 M concentrations of ions and 1 atm pressure (for gases).

It is impossible to directly measure the potential for a single electrode. The overall cell potential for an electrochemical cell may be expressed, however, as the difference between the *standard reduction potentials* ($E°_{red}$) for the reactions at the cathode and at the anode (Equation 1).

$$E°_{cell} = E°_{red}\,(\text{cathode}) - E°_{red}\,(\text{anode}) \qquad \textit{Equation 1}$$

The standard reduction potential is defined as the voltage that a *reduction* half-cell will develop under standard conditions when it is combined with the *standard hydrogen electrode* (SHE), which is arbitrarily assigned a potential of zero volts (Equation 2).

$$2H^+(aq, 1\ M) + 2e^- \rightarrow H_2(g, 1\ atm) \qquad \textit{Equation 2}$$
$$E°_{red}\ (SHE) = 0$$

For the zinc/hydrochloric acid voltaic cell shown in Figure 1, the measured cell potential is equal to 0.76 V. Substituting this value and the zero potential for SHE into Equation 1 gives a value of –0.76 V for the standard reduction potential of the Zn^{2+}/Zn half-cell.

$$E°_{red}\ (cathode) - E°_{red}\ (anode) = E°_{cell}$$
$$E°_{red}\ (SHE) - E°_{red}\ (Zn^{2+}/Zn) = 0.76\ V$$
$$0 - E°_{red}\ (Zn^{2+}/Zn) = 0.76\ V$$
$$E°_{red}\ (Zn^{2+}/Zn) = -0.76V$$

When two half-cells are combined in a voltaic cell, the reaction that has a more positive standard reduction potential will occur as a reduction, while the reaction that has a less positive (or negative) standard reduction potential will be reversed and will take place as an oxidation. In this experiment, electrochemical cells consisting of different metal ion/metal half-cells, i.e., copper(II) sulfate/copper metal versus zinc sulfate/zinc metal, will be tested. The "direction" of each reaction—the identity of the anode and the cathode—will be determined when a positive voltage is observed. A positive voltage means that the half-cells have been properly connected to the positive and negative leads on the voltmeter so that a spontaneous reaction will occur. Recall that in a voltaic cell, the positive electrode is the cathode (the site of the reduction) and the negative electrode is the anode (the site of oxidation).

Materials

- Copper strips or foil, Cu, 1-cm^2
- Copper(II) sulfate solution, $CuSO_4$, 1 M, 1 mL
- Iron strips or sheet, Fe, 1-cm^2
- Iron(II) sulfate solution, $FeSO_4$, 1 M, 1 mL
- Magnesium ribbon, Mg, 1 cm
- Magnesium sulfate solution, $MgSO_4$, 1 M, 1 mL
- Silver foil, Ag, 1-cm^2
- Silver nitrate solution, $AgNO_3$, 1 M, 1 mL
- Sodium nitrate solution, $NaNO_3$, 1 M, 2 mL
- Zinc strips or foil, Zn, 1-cm^2
- Zinc sulfate solution, $ZnSO_4$, 1 M, 1 mL
- Beral-type pipets or eyedroppers, 6
- Computer interface system (optional)
- Data collection software (optional)
- Filter paper, quantitative, 9-cm
- Multimeter or Voltage sensor
- Petri dish or acetate transparency
- Sandpaper or steel wool
- Scissors
- Tweezers or forceps
- Wash bottle and distilled water
- White paper and pencil

Safety Precautions

Silver nitrate solution is a corrosive liquid and toxic by ingestion. It will stain skin and clothing. Copper(II) sulfate solution is toxic by ingestion. Iron(II) sulfate and zinc sulfate solutions are slightly toxic. Magnesium metal is a flammable solid; avoid contact with flames and heat. Metal pieces may have sharp edges—handle with care. Wash hands thoroughly with soap and water before leaving the laboratory. Follow all laboratory safety guidelines. Please review current Safety Data Sheets for additional safety, handling and disposal information.

Procedure

Part A. Cell Potentials versus Zinc as the Reference Electrode

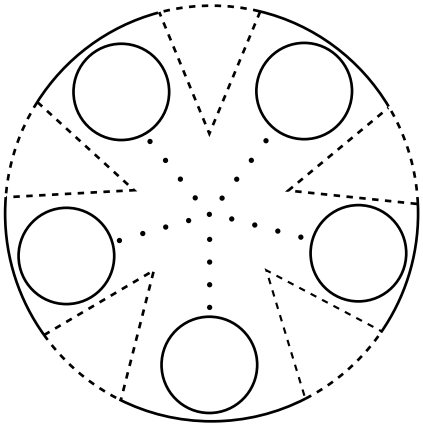


Figure 2.

1. Label a sheet of paper with the names of the five metals to be tested (copper, iron, magnesium, silver, and zinc). Obtain one small piece of each metal and place it on the paper.
2. Polish the metal pieces with sandpaper or steel wool, if necessary, to obtain fresh, shiny surfaces. Wipe the metal strips clean with paper towels to remove any bits of steel wool adhering to the metal.
3. Obtain a piece of filter paper. Using a *pencil*, draw five small circles in a symmetrical pattern on the filter paper, and connect the circles by means of a path of dots, as shown in Figure 2.
4. Using a pair of scissors, cut wedges between the circles and remove the wedges. Label the circles Cu, Fe, Mg, Ag, and Zn (see Figure 3).
5. Place the labeled filter paper in a Petri dish or on a sheet of acetate transparency or plastic wrap.
6. Using a separate Beral-type pipet or eyedropper for each solution, place 1–2 drops of metal ion solution onto its corresponding circle on the filter paper [copper(II) sulfate on Cu, iron(II) sulfate on Fe, etc.] *Note:* Add more drops of metal ion solution to each circle as needed to keep the paper moist during the course of the experiment.
7. Using forceps, place each metal on the "wet spot" in the appropriate circle on the filter paper. Wipe the forceps clean with a paper towel for each metal to avoid contaminating the metals with other metal ions.
8. Using a clean, Beral-type pipet, add several drops of sodium nitrate solution all along the path of dots connecting the metals. Be sure there is a continuous trail of sodium nitrate solution between each circle and the center. *Note:* It may be necessary to add more sodium nitrate along the path of dots as the paper dries out during the course of the experiment.
9. Zinc metal ($E°_{red}$ = –0.76 V) will be used as the "reference electrode" in Part A. Place the positive lead from a multimeter or voltage sensor on the piece of zinc and the negative lead on the piece of copper metal. Read the live voltage. *If the voltage drops to 0.00 V or if a negative voltage is displayed, reverse the leads—place the negative lead on the zinc and the positive lead on the copper.*

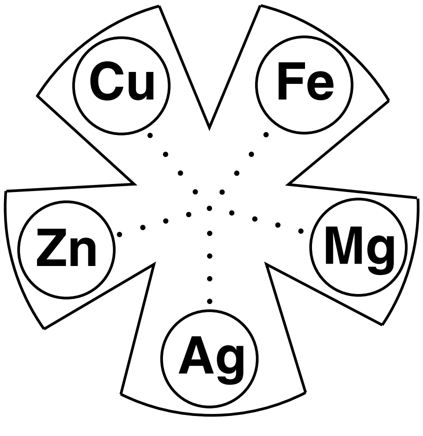


Figure 3.

10. Record which metal is the positive electrode (cathode) and which metal is the negative electrode (anode) on a separate sheet of paper when a positive voltage is obtained.
11. When the voltage reading stabilizes, record the positive voltage value.
12. Repeat steps 9–11 to measure the cell potentials for the other metals (iron, magnesium, and silver) versus zinc as the reference electrode. Remember to record which metal is the (+) electrode and which metal is the (–) electrode when a positive voltage is obtained. Replenish the metal ion solutions or the salt bridge (sodium nitrate) solution as needed if the paper dries out.
13. Using Equation 1 in the *Background* section, calculate the experimental value of the standard reduction potential ($E°_{red}$) for each metal and record the result. Recall that $E°_{red}$ for zinc is equal to –0.76V.

Part B.

14. Select two metals other than zinc (such as copper and magnesium) and use their experimental $E°_{red}$ values to predict the cell potential for a voltaic cell made up of these two half-cells. Show the calculation and record the predicted value of $E°_{cell}$. Note which metal should be the cathode and which metal should be the anode.
15. Measure the cell potential for the voltaic cell and record the value.
16. Repeat steps 14 and 15 for another combination of two metals, not including zinc. *Note:* There are six possible metal combinations (voltaic cells) in Part B. Different voltaic cells may be assigned to different groups so that all the data will be collected as a class.

Disposal

Please consult your current *Flinn Science Catalog Reference Manual* for general guidelines and specific procedures, and review all federal, state and local regulations that may apply, before proceeding. Excess iron(II) sulfate solution may be disposed of by rinsing it down the drain with excess water according to Flinn Suggested Disposal Method #26b. Save the other metal ion solutions in properly labeled bottles for future use. With the exception of magnesium, the metal pieces may be reused from lab to lab and from year to year. Rinse the metal pieces with distilled water, dry them thoroughly on paper towels, and store them in properly labeled zipper-lock bags for future use. Magnesium metal ribbon should be freshly cut each year; dispose of used magnesium pieces in the solid trash according to Flinn Suggested Disposal Method #26a.

Connecting to the National Standards

This laboratory activity relates to the following National Science Education Standards (1996):

Unifying Concepts and Processes: Grades K–12
Evidence, models, and explanation
Content Standards: Grades 5–8
Content Standard B: Physical Science, properties and changes of properties in matter
Content Standards: Grades 9–12
Content Standard B: Physical Science, structure of atoms, structure and properties of matter, chemical reactions

Tips

- Iron(II) sulfate solution is air and light sensitive—prepare the solution fresh for each lab.
- Either sodium nitrate or potassium nitrate may be used as the "salt bridge" electrolyte in this experiment. Do not use sodium or potassium chloride as the electrolyte if silver is one of the metals being tested.

References

Cesa, I. *Flinn ChemTopic™ Labs*, Vol. 17, *Electrochemistry;* Flinn Scientific: 2004; pp 11–22.

Materials for *Measuring Cell Potentials* are available from Flinn Scientific, Inc.

Catalog No.	Description
C0182	Copper Strips, 0.5″ × 6″
C0246	Copper(II) Sulfate Solution, 1 M, 500 mL
I0058	Iron Strips, 10/pk
F0016	Iron(II) Sulfate Heptahydrate, 500 g
M0139	Magnesium Ribbon, 12.5 g
M0016	Magnesium Sulfate Heptahydrate, 100 g
S0270	Silver Foil
S0304	Silver Nitrate Solution, 1 M, 25 mL
S0281	Sodium Nitrate, 100 g
Z0024	Zinc Strips, 10/pk
Z0031	Zinc Sulfate Solution, 1 M, 500 mL

Consult your *Flinn Scientific Catalog/Reference Manual* for current prices.