ELECTROCHEMISTRY
BUILD YOUR OWN HAND-HELD BATTERY

Introduction

Experience and learn the concepts you need to help you succeed on the AP* Chemistry exam with this guided-inquiry activity! This activity covers topics from Big Idea 3 in the AP Chemistry curriculum. Complete a thorough homework set before lab day to delve into the basic principles of electrochemistry, involving the similarities and differences between galvanic and electrolytic cells. Then, you are tasked with the challenge to build your very own hand-held battery out of a few simple materials. The object of the challenge is to successfully assemble the battery materials so the illumination of a red LED component completes the task. You'll love this safe and fun activity while gaining a deeper understanding of electrochemistry and its real-world connection to batteries!

Concepts

- Half-cell reaction
- Galvanic cell vs. electrolytic cell
- Oxidation–reduction reactions
- Standard reduction potential

Background

Galvanic Cells

An electrochemical cell results when an oxidation reaction and a reduction reaction occur, and the resulting electron transfer between the two processes occurs through an external wire. The oxidation and reduction reactions are physically separated from each other and are called half-cell reactions. A half-cell can be prepared from almost any metal in contact with a solution of its ions. Since each element has its own electron configuration, each element develops a different electrical potential, and different combinations of oxidation and reduction half-cells result in different voltages for the completed electrochemical cell.

The standard reduction potential is the voltage that a half-cell, under standard conditions (1 M, atm, 25 °C), develops when it is combined with the standard hydrogen electrode, that is arbitrarily assigned a potential of zero volts. A chart of reduction half-cell reactions, arranged in order of decreasing standard reduction potential, shows the relative ease of reduction of each substance listed. The more positive the reduction potential, the easier the reduction. A spontaneous cell (a battery) can be constructed if two half-cells are connected internally using a salt bridge and externally using a metallic connector. In an electrochemical cell, the reaction listed in the standard reduction potential chart with the more positive voltage occurs as a reduction, and the reaction listed with the less positive voltage reverses and occurs as an oxidation reaction. The cell voltage can be found by adding the voltages listed in the table, with the value of the voltage for the oxidation reaction becoming the negative of its reduction reaction voltage.

As an example, consider a cell made up of copper and aluminum half-cells.

$$Cu^{2+}(aq) + 2\ e^{-} \rightarrow Cu(s) \qquad E° = 0.34\ V$$

$$Al^{3+}(aq) + 3\ e^{-} \rightarrow Al(s) \qquad E° = –1.66\ V$$

The copper reaction has the more positive potential and remains a reduction reaction. The aluminum reaction with the less positive (more negative) potential is reversed and becomes an oxidation reaction. Its potential is now an oxidation potential:

$$Al(s) \rightarrow Al^{3+}(aq) + 3\ e^{-} \qquad E° = +1.66\ V$$

The reduction potential and the oxidation potential are added to find the cell voltage:

$$3Cu^{2+}(aq) + 2Al(s) \rightarrow 3Cu(s) + 2Al^{3+}(aq)$$

$$E°_{\text{cell}} = E°_{reduction} + E°_{oxidation}$$

$$E°_{\text{cell}} = \ 0.345\ V + 1.66\ V + 2.00\ V$$

A positive value for $E°_{\text{cell}}$ indicates the oxidation–reduction reaction, as written, is spontaneous.

A cell representation such as the following: Zn(s) | Zn^{2+}(1.0 M) || Cu^{2+}(0.0010 M) | Cu(s) means that a cell is constructed of zinc metal dipping into a 1.0 M solution of Zn^{2+}. The symbol "|" refers to a phase boundary. The symbol "||" indicates a salt bridge between the zinc ion solution and the copper ion solution. The second half-cell is copper metal dipping into a 0.0010 M solution of copper ions. The anode is on the left (where oxidation occurs) and the cathode is on the right (where reduction occurs).

Electrolytic Cells

When an electric current is passed through an aqueous solution containing an electrolyte (Na_2SO_4), the water molecules break apart or decompose into their constituent elements, hydrogen and oxygen. The overall reaction occurs as two separate, independent half-reactions. Reduction of the hydrogen atoms to elemental hydrogen (H_2) occurs at the cathode (–), while oxidation of the oxygen atoms in water to elemental oxygen (O_2) occurs at the anode (+). Each half-reaction is accompanied by the production of OH^- or H^+ ions as shown below:

$$\text{Cathode:} \quad 4e^{-} + 4H_2O \rightarrow 2H_2(g) + 4OH^{-}$$

$$\text{Anode:} \quad 2H_2O \rightarrow O_2(g) + 4H^{+} + 4e^{-}$$

Experiment Overview

Gain pre-lab preparation by completing the following homework set to gain conceptual understanding of galvanic and electrolytic cells. Then, hit the ground running on lab day and build your own hand-held battery from simple components. You will have time to do a post hand-held battery build analysis. Draw and label your observations in your lab notebook. Did you build an electrolytic cell or a galvanic cell? Prove it.

Pre-Lab Homework Assignment

Complete the following homework set and turn in any graphs or figures you were asked to create. Use a separate sheet of paper, if necessary.

Redox Reactions

1. Identify equations a–d as redox or nonredox. Explain.
   a. $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

   b. $Zn^{2+}(aq) + Al(s) \rightarrow Zn(s) + Al^{3+}(aq)$

   c. $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

   d. $HF(g) + H_2O(l) \rightleftharpoons H_3O^+ + F^-(aq)$

2. Identify the oxidation half-reaction and the reduction half-reaction in *a–c.*
   a. $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

   b. $2Al(s) + 3CuCl_2(aq) \rightarrow 2AlCl_3(aq) + 3Cu(s)$

   c. $Zn(s) + V^{3+}(aq) \rightarrow V^{2+}(aq) + Zn^{2+}(aq)$

3. Go back to 2, a–c and identify:
   a. The oxidizing and reducing agents.

   b. The species that is oxidized and the species that is reduced.

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Electrolytic Cell

A student setup a Hoffman apparatus electrolysis experiment as shown in Figure 1. The Hoffman electrolysis apparatus is a type of electrolytic cell where an electric current passes through an aqueous solution containing an electrolyte. As a result, water molecules decompose into their constituent elements. Two independent half-reactions are observed at the cathode and the anode. For a free demonstration video on the Hoffman apparatus, visit flinnsci.com.

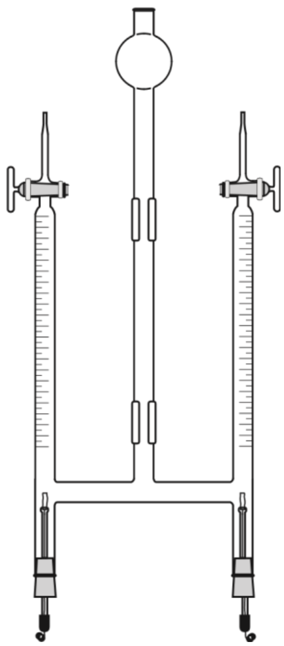

Figure 1.

4. The student added 0.5 M sodium sulfate electrolytic solution to the mouth of the Hoffman apparatus and then connected the battery leads to the electrodes. She immediately witnessed bubble formation and collected two gases on each side of the assembly. Answer questions *a–d.*

    a. Why did the student choose 0.5 M sodium sulfate for this experiment? Are there other electrolytic solutions she may test?

    b. Report the overall reaction and the independent half reactions. Identify the gases collected. Predict the volumes of each gas collected in the apparatus.

    c. Identify the cathode and anode and report where each occurs in the half reactions from question 4b.

    d. Can the student use the Hoffman apparatus without the battery? In other words, will the decomposition (formation of gases) spontaneously occur?

5. The student carefully opened each stopcock and collected each gas in separate test tubes. She inserted a lit wood splint into each.
   a. Predict what occurred to the lit wood splint of the gas collected at the anode.

   b. Predict what occurred to the lit wood splint of the gas collected at the cathode.

6. As an extension, 1-mL of universal indicator solution was added to the sodium sulfate solution while the Hoffman apparatus was connected. Predict the observations. *Hint:* Look up the exhibited color changes of universal indicator solution at various pH values.

Galvanic Cell

See Figure 2 for the second experiment setup, a galvanic cell. In a galvanic cell, a spontaneous chemical reaction releases energy in the form of electricity (moving electrons).

7. Using arrows, label the parts of the galvanic cell in Figure 2 and answer questions *a–d*. Use this list of key words: electrodes, electrode storage/compartment, cathode, anode, and salt bridge.

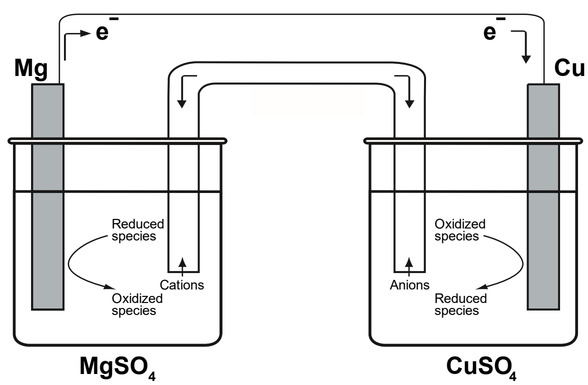

Figure 2. Galvanic Cell

   a. What would happen to the cell if the salt bridge was removed?

   b. Write the half-reactions that are taking place.

   c. Write the cell notation.

d. Calculate the standard cell potential, E°.

8. Seek educational resources and provide an example of a replacement electrode if the copper electrode was not available in Figure 2
9. The materials to build your very own hand-held battery include: 2 small squares of filter paper, 1 M copper(II) sulfate solution, 1 M sodium sulfate solution, magnesium ribbon, and an LED with copper tape attached to the positive terminal. Again, the challenge is to build a hand-held battery by successfully arranging the components to light the LED. *Helpful tips:*
   - *a.* Think safety, first. Make sure you have the proper PPE available to perform this lab, i.e., goggles, apron, and gloves.
   - *b.* Make a list of the equipment and glassware needed for this lab.
   - *c.* Once you successfully complete the experiment, draw a hand-held battery figure in your notebook and label the parts of your battery, i.e. the cathode, anode etc. Write the reactions occurring
   - *d.* How does this battery differ from those practiced in the homework set? How is it similar?
   - *e.* What are the half-cell reactions?
   - *f.* Inspect the separate components of the hand-held battery after connecting it and lighting the LED. Describe any observations.

Safety Precautions

The copper(II) sulfate solution is harmful if swallowed and causes serious skin and eye irritation. The sodium sulfate solution may be harmful if in contact with skin. Magnesium ribbon is a flammable solid. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant *apron. Wash hands thoroughly with soap and water before leaving the laboratory. Please follow all laboratory safety guidelines.*

Teacher's Notes
Electrochemistry: Build Your Own Hand-held Battery
AP* Chemistry Investigation
A Guided-Inquiry Wet/Dry Experiment

Materials Included in Kit *(for 12 groups of students)*

Copper(II) sulfate solution, $CuSO_4$, 1 M, 500 mL

Sodium sulfate solution, Na_2SO_4, 1 M, 375 mL

Copper foil conductive adhesive, Cu, 12″ piece, 2

Filter paper, 100 sheets

LEDs, clear, red, 24

Magnesium ribbon, Mg, 6 ft, 1

Sand paper, 1 (shared)

Additional Materials Required *(for each lab group)*

Weigh boats, medium, 24

Beakers, 50-mL, 24

Tweezers, 12

Scissors, 12

Graduated cylinders, 10-mL, 12

Deionized or distilled water

Time Required

This laboratory activity was specifically written, per teacher request, to be completed in one 50-minute class period. It is important to allow time between the Pre-Lab Homework Assignment and the *Lab Activity*. Prior to beginning the homework, show the students the hand-held battery materials—this will get the procedure thought process rolling.

Pre-Lab Preparation and Complete *Build Your Own Hand-held Battery* Procedure

1. Gently polish both LED terminals with the sand paper on all LEDs to be used by the students.
2. Cut 24 copper conductive adhesive tape pieces 2 cm in length.
3. Cover the positive terminal (the longer terminal) with the 2 cm piece of the adhesive conductive tape of each LED to be used by the students.
4. Cut 24 pieces of rectangular shaped filter paper. The sodium sulfate filter paper is the salt bridge and should be bigger in size than the copper(II) sulfate filter paper.
    a. Cut the 12 pieces to be submerged in copper(II) sulfate in about ½ cm^2 in size.
    b. Cut 12 pieces to be submerged in the sodium sulfate solution in about 1 cm^2 in size.

Build Your Own Hand-held Battery Procedure

Lead the students through steps 1–3 after providing the materials. At step 4, allow them to determine the correct component arrangements to successfully light the LED.

1. Measure 2 mL, each of the 1 M copper(II) sulfate solution and the 1 M sodium sulfate solution using a 10-mL graduated cylinder. Pour into separate medium sized weigh boats or small beakers (50-mL).
2. Using tweezers, dip the larger (1 cm^2) pre-cut filter paper into the sodium sulfate solution and the smaller (1/2 cm^2) pre-cut filter paper into the copper(II) sulfate solution. Dip long enough to completely coat each filter paper (about 10 seconds).
3. Place both on a separate weigh boat to let dry until the filter papers are damp with solution, not dripping. Students may hold each with tweezers and gently wave to decrease drying time.
4. While filter paper is drying, students make predictions. In their notebooks, students illustrate the experiment and identify the anode, cathode, the salt bridge, and evidence of electron flow; the lit LED. Students should include the half-cell reactions taking place.
5. See Figure 3 on how to correctly arrange the components. Lead the students into squeezing the LED between the thumb and index finger for component contact. Add a drop of DI water if necessary, but not too much to drown the battery. Darken the room or cup hand over LED to watch it light.

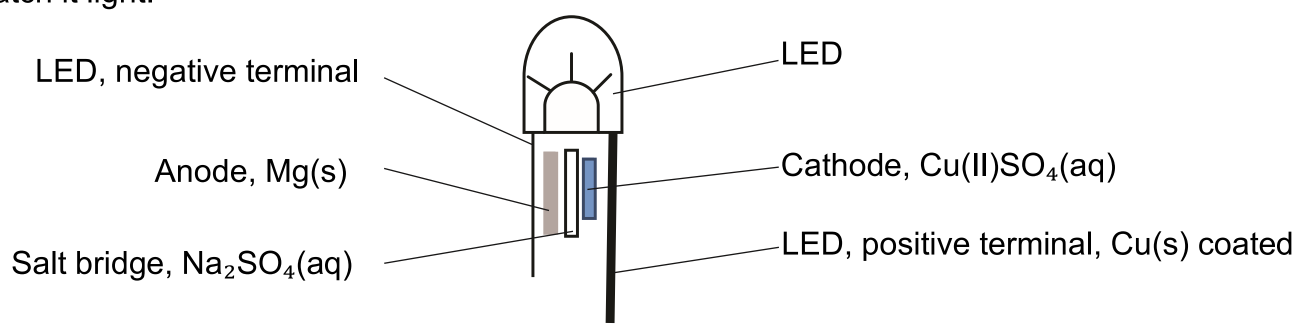

Figure 3. The correct arrangement of the hand-held battery. Squeeze with thumb and index finger on each side of the LED terminals to light the LED.

Safety Precautions

The copper(II) sulfate solution is harmful if swallowed and causes serious skin and eye irritation. The sodium sulfate solution may be harmful if in contact with skin. Magnesium ribbon is a flammable solid. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Wash hands thoroughly with soap and water before leaving the laboratory. Please follow all laboratory safety guidelines. Remind students to wash their hands thoroughly with soap and water before leaving the laboratory. Please review current Safety Data Sheets for additional safety, handling, and disposal information.

Disposal

Please consult your current *Flinn Scientific Catalog/Reference Manual* for general guidelines and specific procedures, and review all federal, state and local regulations that may apply, before proceeding. You may save all materials, including solutions, for future labs. Any leftover copper(II) sulfate and sodium sulfate solutions may be flushed down the drain with excess water according to Flinn disposal method #26b.

Alignment with AP* Chemistry Curriculum Framework—Big Idea 3

Enduring Understandings and Essential Knowledge

Chemical reactions can be classified by considering what the reactants are, what the products are, or how they change from one into the other. Classes of chemical reactions include synthesis, decomposition, acid−base, and oxidation−reduction reactions. (Enduring Understanding 3B)

3B3: In oxidation−reduction (redox) reactions, there is a net transfer of electrons. The species that loses electrons is oxidized, and the species that gains electrons is reduced.

Chemical and physical transformations may be observed in several ways and typically involve a change in energy. (Enduring Understanding 3C)

3C3: Electrochemistry shows the interconversion between chemical and electrical energy in galvanic and electrolytic cells. Chemical equilibrium is a dynamic, reversible state in which rates of opposing processes are equal. (Enduring Understanding 6A)

6A4: The magnitude of the equilibrium constant, K, can be used to determine whether the equilibrium lies toward the reactant side or product side.

Chemical equilibrium plays an important role in acid−base chemistry and in solubility. (Enduring Understanding 6C)

6C3: The solubility of a substance can be understood in terms of chemical equilibrium.

Learning Objectives

3.8 The student is able to identify redox reactions and justify the identification in terms of electron transfer.

3.12 The student can make qualitative or quantitative predictions about galvanic or electrolytic reactions based on half-cell reactions and potentials and/or Faraday's laws.

3.13 The student can analyze data regarding galvanic or electrolytic cells to identify properties of the underlying redox reactions.

6.7 The student is able, for a reversible reaction that has a large or small K, to determine which chemical species will have very large versus very small concentrations at equilibrium.

6.21 The student can predict the solubility of a salt, or rank the solubility of salts, given the relevant K_{sp} values.

Science Practices

2.2 The student can apply mathematical routines to quantities that describe natural phenomena.

2.3 The student can estimate numerically quantities that describe natural phenomena.

4.3 The student can collect data to answer a particular scientific question.

5.1 The student can analyze data to identify patterns or relationships.

Lab Hints

- A pre-lab prep alternative to cutting filter paper squares: you may first dip the filter paper in the sulfate solutions, allow to dry, then cut into squares for use by the students.
- Depending the inquiry level of your students, they may perform all of the set-up steps from the Pre-Lab Preparation.

Teaching Tips

- Flinn Scientific has excellent video resources that enhance the teaching experience! Simply type in the key word electrolysis or Hoffman apparatus to pull up some great videos.
- The Colorful Electrolysis Demonstration is a great extension to this lab! This demonstration kit is available from Flinn (Catalog No. AP6467).

Answers to Pre-Lab Homework Assignment *(Student answers will vary.)*

Redox Reactions

1. Identify equations a–d as redox or nonredox. Explain.

   Electrons are gained or lost in redox reactions; a species is oxidized or reduced. 1 c and d are nonredox due to being acid base reactions.

   a. $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

   Redox

   b. $Zn^{2+}(aq) + Al(s) \rightarrow Zn(s) + Al^{3+}(aq)$

   Redox

   c. $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

   NonRedox

   d. $HF(g) + H_2O(l) \rightleftharpoons H_3O^+ + F^-(aq)$

   NonRedox

2. Identify the oxidation half-reaction and the reduction half-reaction in *a–c.*

   a. $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

   $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ (oxidation)
   $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ (reduction)

b. $2Al(s) + 3CuCl_2(aq) \rightarrow 2AlCl_3(aq) + 3Cu(s)$

$Al(s) \rightarrow Al^{3+} + 3e^-$ (oxidation)
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ (reduction)

c. $Zn(s) + V^{3+}(aq) \rightarrow V^{2+}(aq) + Zn^{2+}(aq)$

$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ (oxidation)
$V^{3+}(aq) + 2e^- \rightarrow V^{2+}(aq)$ (reduction)

3. Go back to 2, a–c and identify:

a. The oxidizing and reducing agents.

2a. *Zn is the reducing species and Cu^{2+} is the oxidizing species.*
2b. *Al is the reducing species and Cu^{2+} is the oxidizing species.*
2c. *Zn is the reducing species and V^{3+} is the oxidizing species.*

b. The species that is oxidized and the species that is reduced.

2a. *Zn is oxidized and Cu^{2+} is reduced.*
2b. *Al is oxidized and Cu^{2+} is reduced.*
2c. *Zn is oxidized and V^{3+} is reduced.*

Electrolytic Cell

A student setup a Hoffman apparatus electrolysis experiment as shown in Figure 1. The Hoffman electrolysis apparatus is a type of electrolytic cell where an electric current passes through an aqueous solution containing an electrolyte. As a result, water molecules decompose into their constituent elements. Two independent half-reactions are observed at the cathode and the anode. For a free demonstration video on the Hoffman apparatus, visit flinnsci.com.

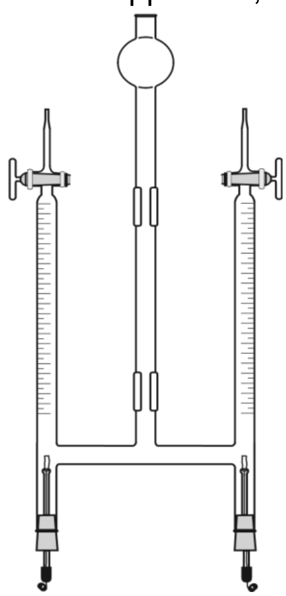

Figure 1.

4. The student added 0.5 M sodium sulfate electrolytic solution to the mouth of the Hoffman apparatus and then connected the battery leads to the electrodes. She immediately witnessed bubble formation and collected two gases on each side of the assembly. Answer questions *a–d.*

   a. Why did the student choose 0.5 M sodium sulfate for this experiment? Are there other electrolytic solutions she may test?

   Sodium sulfate is an electrolyte, it is conductive. The student can use solutions of potassium iodide or potassium chloride.

   b. Report the overall reaction and the independent half reactions. Identify the gases collected. Predict the volumes of each gas collected in the apparatus.

   Oxygen and hydrogen gases are produced/collected. Since hydrogen gas forms at twice the molar volume in the overall reaction, double the amount of hydrogen gas collects in the apparatus vs. oxygen gas.

   $H_2O \rightarrow 4H_2 + 2O_2$
   Cathode: $4e^- + 4H_2O \rightarrow 2H_2(g) + 4OH^-$
   Anode: $2H_2O \rightarrow O_2(g) + 4H^+ + 4e^-$

   c. Identify the cathode and anode and report where each occurs in the half reactions from question 4b.

   See 4b.

   d. Can the student use the Hoffman apparatus without the battery? In other words, will the decomposition (formation of gases) spontaneously occur?

   No, electrolysis reactions require external power where a non-favorable redox reaction occurs.

5. The student carefully opened each stopcock and collected each gas in separate test tubes. She inserted a lit wood splint into each.

   a. Predict what occurred to the lit wood splint of the gas collected at the anode.

   Oxygen gas is flammable, so it ignited the wood splint.

   b. Predict what occurred to the lit wood splint of the gas collected at the cathode.

   Hydrogen gas is also flammable, so it ignited the wood splint as well.

6. As an extension, 1-mL of universal indicator solution was added to the sodium sulfate solution while the Hoffman apparatus was connected. Predict the observations. Hint: Look up the exhibited color changes of universal indicator solution at various pH values.

   Beautiful colors will result. At the cathode: solution will be purple/blue due to the production of basic hydroxide ions. At the anode: solution will be pink/red due to the production of acidic protons.

Galvanic Cell

See Figure 2 for the second experiment setup, a galvanic cell. In a galvanic cell, a spontaneous chemical reaction releases energy in the form of electricity (moving electrons).

7. Using arrows, label the parts of the galvanic cell in Figure 2 and answer questions a–d. Use this list of key words: electrodes, electrode storage/compartment, cathode, anode, and salt bridge

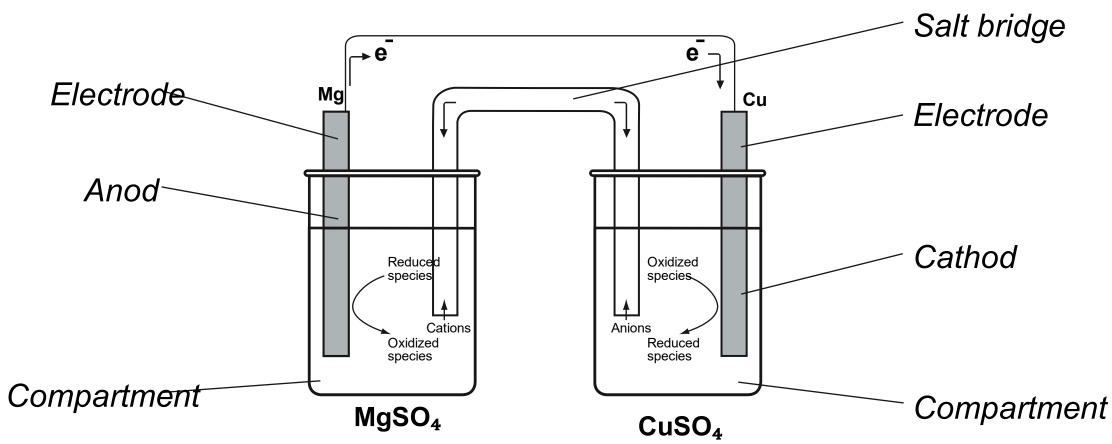

Figure 2. Galvanic Cell

a. What would happen to the cell if the salt bridge was removed?

   Build-up of positive and negative charges would occur; salt bridges are needed for chemical neutrality.

b. Write the half-reactions that are taking place.

   $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
   $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$

c. Write the cell notation.

   $Mg(s) \mid Mg^{2+}(1\ M) \parallel Cu^{2+}(1\ M) \mid Cu(s)$

d. Calculate the standard cell potential, E°.

$$E°_{cell} = E°_{cathode} + E°_{anode}$$

$$E°_{\mathrm{cell}} = 0.34 - (-2.37) = 2.71\ V$$

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8. Seek educational resources and provide an example of a replacement electrode if the copper electrode was not available in Figure 2.

   A standard, non-conductive electrode such as platinum or graphite are good options.

9. The materials to build your very own hand-held battery include: 2 small squares of filter paper, 1 M copper(II) sulfate solution, 1 M sodium sulfate solution, magnesium ribbon, and an LED with copper tape attached to the positive terminal. Again, the challenge is to build a hand-held battery by successfully arranging the components to light the LED. Helpful tips:

   a. Think safety, first. Make sure you have the proper PPE available to perform this lab, i.e., goggles, apron, and gloves.

   b. Make list of the equipment and glassware needed for this lab.

   c. Once you successfully complete the experiment, draw a hand-held battery figure in your notebook and label the parts of your battery, i.e. the cathode, anode etc. Write the reactions occurring.

   See Figure 3

   d. How does this battery differ from those practiced in the homework set? How is it similar?

   It is a galvanic cell—it does not require external power to work. The filter papers are the cell's compartments and there is not a magnesium sulfate solution present.

   e. What are the half-cell reactions?

   Cu tape was applied to the positive LED terminal of the battery as the conductive material to prevent reaction between Mg(s) and the bare LED terminal.

   $Mg(s) \rightarrow Mg^{2+} + 2e^-$ *(oxidation, anode)*
   $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ *(reduction, cathode)*

   f. Inspect the separate components of the hand-held battery after connecting it and lighting the LED. Describe any observations.

   Student answers will vary. Dark spots are seen on the copper(II) sulfate filter paper, which is copper metal. Without the conductive tape, the LED terminals can darken due to oxidation of Mg(s).

References

Eggen, P.; Skaugrud, B. *An Easy-to-Assemble Three-Part Galvanic Cell*. J. Chem. Educ. 2015, 92 (6), 1053–1055.
AP Chemistry Guided-Inquiry Experiments: Applying the Science Practices*; The College Board: New York, NY, 2013.

The *Electrochemistry: Build Your Own Hand-held Battery AP* Chemistry Investigation— A Guided-Inquiry Wet/Dry Experiment* is available from Flinn Scientific, Inc.

Catalog No.	Description
AP8496	Electrochemistry: Build Your Own Hand-held Battery AP* Chemistry Investigation—A Guided-Inquiry Wet/Dry Experiment

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