The University of Texas at Austin
Department of Chemistry
College of Natural Sciences

Galvanic Cell

Performer's Version

Safety Hazards

- Personal Protective Equipment
  - Safety glasses/goggles
  - Nitrile gloves
  - Chemical & flame retardant lab coat
- Chemical Hazards
  - Copper(II) sulfate is harmful if swallowed and causes serious skin irritation; very toxic to aquatic life with long lasting effects.
  - Zinc sulfate is harmful if swallowed and causes serious eye damage; very toxic to aquatic life with long lasting effects.
  - Sodium sulfate may cause skin, eye, and respiratory tract irritation.

Materials

- 75 mL 0.5M Copper(II) sulfate solution
- 75 mL 0.5M Zinc(II) sulfate solution
- 150 mL 0.5M Sodium sulfate solution
- Salt bridge
- 2x 150 mL glass beakers
- Zinc electrode
- Copper electrode
- Voltmeter/Multimeter
- Pipette and pipette tips
- Polypad (blue absorbent paper)

Safety Data Sheet(s)

- Copper(II) sulfate
- Zinc(II) sulfate
- Sodium sulfate

Procedure

1. To maximize visibility, use a document camera for this demonstration.
2. Place the polypad over the surface on which the galvanic cell will be set.
3. Place the salt bridge standing in the empty beakers, with one end in each beaker, or standing on the polypad. Using the pipette set to 5.00 mL, slowly add 0.5M Sodium sulfate to the salt bridge via the opening stem. Make sure it is entirely full (as depicted on the diagram)!
4. Pour 0.5M Copper(II) sulfate solution into one beaker.
5. Pour 0.5M Zinc sulfate solution to the other beaker.
6. Make sure the red wire is plugged into the middle plug of the multimeter, directly above the black one. Make sure the black wire is plugged into the bottom plug (black).
7. Turn the multimeter to the left until it's set to 2000 m. The arrow should be in the quadrant labeled 'V' – meaning that it will be measuring voltage. There shouldn't be any voltage initially, since the clips aren't attached to anything.
8. Carefully place the copper strip/electrode into the copper sulfate solution. Carefully place the zinc strip/electrode into the zinc sulfate solution.
9. Clip the red wire onto the top of the copper strip, and the black wire onto the top of the zinc strip. The multimeter should read a voltage ~1000 (Divide by 1000 to get actual voltage; should be roughly 1 V produced).

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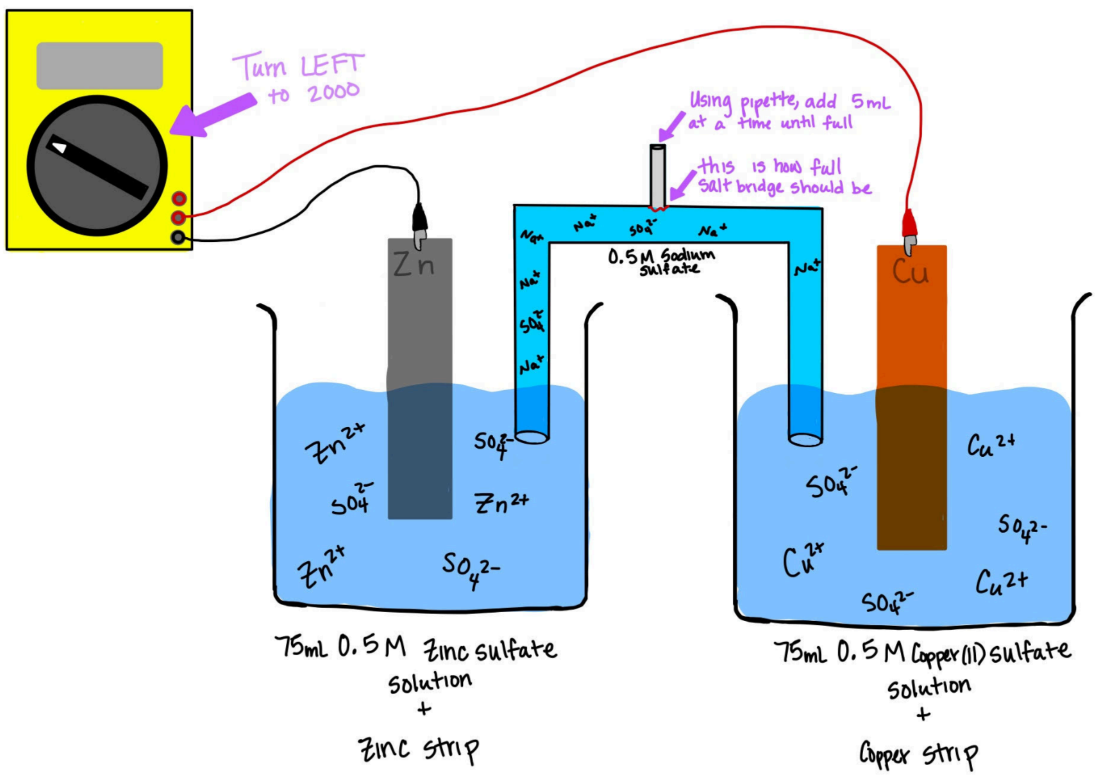
Set-Up References & Diagrams



Pedagogy & Supplemental Information

Electrochemical reactions are a cornerstone of chemistry, particularly evident in galvanic cells where chemical energy is converted into electrical energy. A classic demonstration of this concept involves constructing a galvanic/voltaic cell using zinc(II) sulfate ($ZnSO_4$) and copper(II) sulfate ($CuSO_4$) solutions, with a salt bridge containing sodium sulfate (Na_2SO_4) to complete the circuit. In chemistry, reduction is the gain of electrons by a molecule, atom, or ion, while oxidation is the loss of electrons. These processes often occur simultaneously in redox reactions, where one species is reduced and another is oxidized. This setup showcases the fundamental principles of redox reactions, where oxidation and reduction processes occur at different electrodes.

In the zinc-copper galvanic cell, zinc metal (Zn) serves as the anode and copper metal (Cu) as the cathode. At the anode, zinc undergoes oxidation, losing two electrons to form zinc ions (Zn^{2+}) that enter the solution:

$$Zn\ (s) \longrightarrow Zn^{2+}\ (aq) + 2e^{-}$$

Conversely, at the cathode, copper ions (Cu^{2+}) in the solution gain these electrons and are reduced to form copper metal on the electrode:

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

This transfer of electrons from zinc to copper through an external circuit creates an electric current, illustrating how chemical reactions can generate electrical energy.

The voltage produced by a galvanic cell results from the difference in potential energy between the electrons in the two metals. Zinc has a higher tendency to lose electrons (a more negative electrode potential) compared to copper. This difference in electrode potential drives the electrons from zinc to copper, creating a potential difference measured as voltage. The salt bridge, usually a U-shaped tube filled with an electrolyte like sodium sulfate, maintains electrical neutrality by allowing ions to flow between the two solutions, preventing charge buildup that would otherwise stop the reaction.

Real-world applications of galvanic cells are vast and significant. They form the basis of batteries, which power countless devices from flashlights to smartphones. Understanding the chemistry behind these reactions not only demystifies everyday technology but also underscores the importance of redox reactions in energy storage and conversion.