AP Chemistry - Electrochemical Cells
Student Guide
IS 8016

INTRODUCTION

The tendency of oxidation-reduction (redox) reactions is to proceed to an equilibrium state. Redox reactions proceed in a concurrent nature, meaning that the number of electrons lost to oxidation must be gained through reduction. Certain metals, for example, can undergo redox reactions when placed in an ionic solution of another metal. Consider the following example, when zinc metal is placed in an ionic copper solution:

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

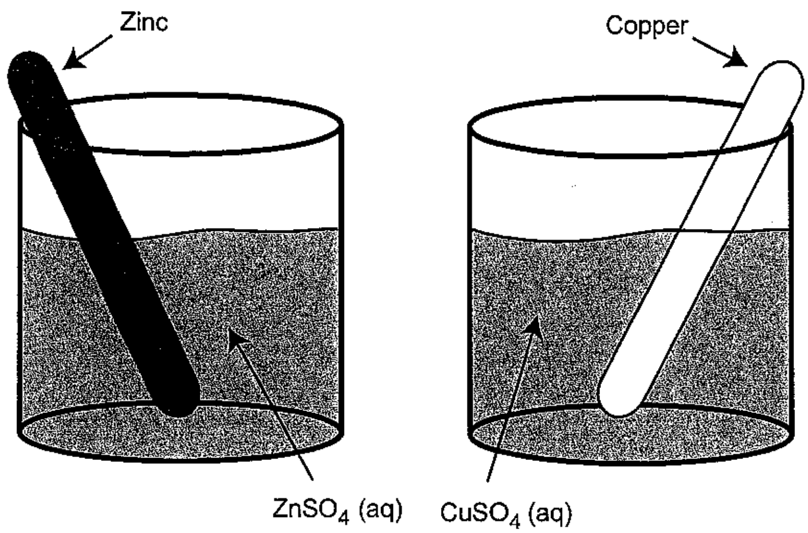
Notice that the zinc, when exposed to the copper solution, is oxidized. In other words, it is losing two electrons. Due to the concurrent nature of redox reactions, the copper is reduced, in turn gaining two electrons. When these reactions occur naturally, the energy generated is given off as heat, which quickly leaves the system. However, the energy from the redox reactions can be harnessed and used to perform work by constructing an electrochemical cell.

Electrochemical Cells

There are two primary types of electrochemical cells, both based on different redox reactions. Some redox reactions occur spontaneously. A cell that harnesses these reactions is called a galvanic, or voltaic, cell. In some cases, redox reactions do not occur spontaneously, and force (in the form of voltage) is applied to drive the redox reactions. These are known as electrolytic cells.

In galvanic cells, the redox reactions are occurring spontaneously. As mentioned above, the energy is quickly lost as heat. To harness the energy from the electron transfer, the cell is actually broken into two half cells. Each half cell is composed of a metal placed in a corresponding electrolytic solution (solution containing corresponding metal salts). One of the half cells contains the reduction part of the reaction and the other half cell contains the oxidation part of the reaction. Staying with the example of zinc and copper, the two half cells would be:

Figure 1: Zinc and Copper Half Cells

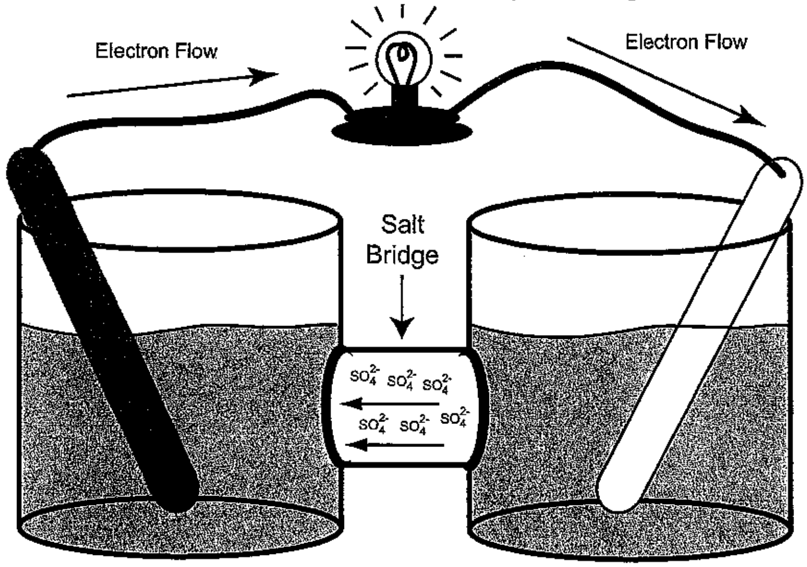


The half cell in which oxidation occurs, in this case zinc, is called the anode. The half cell where reduction occurs, copper, is called the cathode. As mentioned above, the reactions in a galvanic cell occur spontaneously. However, by separating the anode and cathode, the redox reactions will not proceed until both half cells are connected to each other. By connecting the zinc to the copper by a metal conductor, the reaction can proceed and the electrons that would normally be lost as heat are forced through the conductor, providing an electric force.

One problem when separating the components into half cells is that reduction and oxidation are occurring apart from each other. Over time, the cell containing the zinc anode will become more and more positive, as more and more zinc ions are produced. Conversely, the cathode is also becoming negative due to the removal of copper ions from the electrolytic solution containing the copper.

A way to compensate for this is the use of a salt bridge. A salt bridge is a barrier that separates the two half cells. The barrier does not allow the solutions of each half cell to mix but is porous enough to allow the movement of ions in either direction between the half cells. This maintains charge neutrality. As the redox reactions progress, ions (called cations) from the zinc side are attracted to the cathode and migrate through the salt bridge. The SO_4^{2-} ions (called anions) from the copper electrolyte will be attracted to the anode. The equal exchange of charge between Zn^{2+} and SO_4^{2-} maintains charge balance.

Figure 2: Half Cells Connected by Salt Bridge



The electrons traveling through the conductor can be made to work by attaching an electrical device in the conductor. This type of cell is the basis for batteries.

In some cells, the redox reactions are not occurring spontaneously. An external force, in the form of electricity, must be provided to drive the redox reactions. In an electrolytic cell, the system is not split up into half cells. This is because a redox reaction will not spontaneously occur, so both the anode and cathode can be placed in the same solution containing some sort of electrolyte. Though after examining galvanic cells, it may seem futile to create a cell that requires the input of energy, as opposed to producing energy, electrolytic cells are useful for performing electrolysis. Electrolysis (electro – electricity; lysis – break or split apart) is useful for splitting apart various substances. For example, placing a normally non-reactive anode and cathode into a solution of water and an electrolyte and applying enough electrical current to drive redox will cause an oxidation reaction at the anode. This reaction will produce oxygen gas and a reduction reaction at the cathode will produce hydrogen gas. The water has been effectively split (lysed) into its constituent components, hydrogen and oxygen, and the gas can be collected for other uses.

Objective

First, construct a simple chemical battery and determine from the standard reduction potentials what the output of the battery will be (if a voltmeter is available the actual and theoretical voltages can be compared). Second, construct an electrolytic cell and demonstrate how hydrogen and oxygen can be produced from the electrolysis of water.

Materials Included in the Kit

15 each	Copper metal strips
15 each	Magnesium metal strips
15 each	Dialysis tubes
2 X 500mL	Copper Sulfate solution
8 X 500mL	Sodium Sulfate solution
1 X 25mL	Bromothymol Blue Indicator solution

Materials Needed but not Supplied

30 each	250mL glass beakers
30 each	13 X 100 test tubes
15 each	1.5v flashlight bulbs
15 pairs	wire leads with alligator clips
15 each	9v batteries
	Voltmeters (optional)

Safety

Safety goggles
Rubber gloves
Chemical aprons

Safety note: *The hydrogen gas produced by electrochemical cells is extremely flammable. Be sure that no source of ignition is present during the electrolysis experiment.*

Procedure

Part I: The Chemical Battery

1. Soak the dialysis tube in water to soften. Tie a knot in one end of the tube and fill with about 50mL of the copper sulfate solution.

2. Place about 180mL of sodium sulfate solution in a 250mL beaker and then place the dialysis tube in the solution being careful not to allow any copper sulfate solution to enter the beaker.

3. If the copper and magnesium strips are tarnished or oxidized, clean them by dipping them in dilute HCl (1M). Be sure to rinse them thoroughly under running water before placing them in the cell.

4. Carefully place the copper strip in the dialysis tube filled with copper sulfate solution. The edges may be sharp so be careful not to puncture the tube. Place the magnesium strip into the beaker with the solution of sodium sulfate.

5. Connect the wire leads to the electrode strips and then to a 1.5v flashlight bulb or a voltmeter. Observe the results.

Part II: The Electrolysis Cell

1. Place about 200mL of sodium sulfate solution in a 250mL beaker. Add several drops of Bromothymol blue and mix.

2. Fill the two test tubes with solution from the beaker and carefully invert them into the beaker holding your finger over the opening to prevent solution from escaping from the test tubes.

3. Place a nine volt battery in the beaker and position one of the filled test tubes over each battery terminal to capture the gas that is generated. We will want to compare the volume of gas collected at each terminal so make sure the test tube is positioned completely over the terminal.

Optional: use wire leads with alligator clips attached to one end and a carbon or platinum electrode to the other. Place the electrodes completely inside the test tubes. Attach the alligator clips to the 9v battery. Observe the results.

Chemical disposal: *The solutions and metals should be disposed of by a licensed chemical waste contractor.*

Name:	Instructor:
Date:	Class/Lab Section:

DATA ANALYSIS

1. In your laboratory report include the balanced net ionic equations for the oxidation and reduction reactions occurring at each electrode. Indicate the flow of electrons in each electrochemical cell.

2. For the chemical battery, determine the theoretical voltage produced using standard reduction potentials found in chemical references (i.e. CRC Handbook of Chemistry and Physics). If a voltmeter is available, compare the theoretical value to the observed. Explain any difference between the two values.

3. In the electrolysis cell, explain the color change occurring at each electrode. What is the source of acid and base at the electrodes?

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DATA ANALYSIS

4. In comparing the gas generated during electrolysis, how many molecules of hydrogen are produced for every molecule of oxygen? How many electrons are liberated for every molecule of oxygen formed?

5. What is the potential industrial use of the electrolysis of water?