Electrolysis of Water in Living Color

Electrolysis of Water

Introduction

Electrolysis is a great way to introduce the concepts of elements and molecules, chemical changes, as well as stoichiometry for higher-level chemistry students. With this simple apparatus and a DC power supply, a dazzling electrolysis experiment can easily be performed.

Concepts

- Electrolysis
- Oxidation–reduction reactions (redox)
- Half-reactions
- Balancing chemical equations

Materials

Bromthymol blue solution, 0.04%, 20 mL
Hydrochloric acid, 1 M, 10 mL
Sodium hydroxide solution, 1 M, 10 mL
Sodium sulfate solution, Na_2SO_4, 0.1 M, 200 mL
Beaker, 600-mL
Connector cords with alligator clips, 2
DC power supply (low voltage, low current)
Glass stirring rod
Hoffman apparatus, 27 cm × 10 cm
Lighter
Rubber stoppers, solid, size 0, 2
Support clamps, 2
Support stand
Test tubes 13 X 100 mm, 2
Wood splints

Safety Precautions

Do not operate the power supply with wet hands or in a wet area. Be sure that the working area is dry, and check the glass for chips and cracks before running the current. Do not use an AC power supply. AC will produce oxygen gas and hydrogen gas equally at both electrodes, which can be an explosive mixture. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Wash hands thoroughly with soap and water before leaving the laboratory. Follow all laboratory safety guidelines. Please review current Material Safety Data Sheets for additional safety, handling, and disposal information.

Procedure

Electrolysis

1. Set up a support stand with two support clamps.
2. Assemble the Hoffman apparatus.
3. In the 600-mL beaker mix the 200 mL of 0.1M sodium sulfate solution with enough of the 0.4% Bromthymol blue solution to give the solution a definite color. Ensure the solution is green. If it is blue, add 1 M hydrochloric acid drop-wise until a green color results. If it is yellow, add 1 M sodium hydroxide solution drop-wise until a green color results.
4. With the stopcocks open, add the electrolyte solution slowly through the mouth of the apparatus until the graduated side arms are filled up to the base of the stopcocks. *Note:* Do not allow liquid to flow into the glass tips above the stop-cocks. The center tube should *not* be filled completely—it will fill with more liquid as gases are produced at the elec-trodes and liquid is displaced from the side arms.
5. Close the stopcocks.
6. Connect the leads to the DC power supply. *Warning:* Make sure the DC power supply is *off* before attaching leads to

the wire electrodes. Do not use an AC power supply.

7. Connect the red alligator clip to the red lead of the Hoffman apparatus. This is the anode.
8. Connect the black alligator clip to the black lead of the Hoffman apparatus This is the cathode.
9. Turn on the DC power supply and adjust the voltage until a large amount of bubbles emanate from the electrodes.
10. Record observations such as color of the solution and bubbles formation at the electrodes.
11. Turn off the power supply to stop the electrolysis reaction and disconnect the alligator clips from the Hoffman apparatus.

Testing the Gases

12. Compare the volume of gas collected at each electrode.
13. Collect the hydrogen gas from the cathode that have been produced by placing a test tube over the stopcock, opening the stopcock to release the gas into the test tube
14. Immediately, place a rubber stopper in the end of the test tube.
15. Ignite the tip of a wood splint with a match or lighter. Let it burn until the tip has a small, steady flame.
16. Place the flaming wood splint near the stopper.
17. Quickly remove the rubber stopper and insert the flaming wood splint.
18. Observe what happens. What gas is this? (A pale blue flame and a loud "pop" or "bark" indicates the gas is hydrogen.)
19. Repeat steps 13–18 for the other test tube. Except this time blow out the flaming wood splint so that only the tip glows red. The oxygen gas test requires a *glowing* red splint, not a burning splint. (The glowing wood splint should burst into flames revealing the second gas to be oxygen.)

Disposal

Please consult your current *Flinn Scientific Catalog/Reference Manual* for general guidelines and specific procedures governing the disposal of laboratory waste. Sodium sulfate solution may be disposed of down the drain with an excess of water according to Flinn Suggested Disposal Method #26b.

Tips

- In an electrolytic cell, the positive electrode is the anode (the site of oxidation), and the negative electrode is the cathode (site of reduction). The DC power supply acts as an electron pump "pushing" electrons into one electrode and "pulling" them from the other electrode. Electric current flows through the electrolysis solution via the migration of ions. Anions move towards the anode, cations move towards the cathode.
- When removing the test tubes from the stopcock, keep a finger over the hole in the stopper to prevent gas loss.
- The concentration of sodium sulfate is not critical. It also does not need to be made from anhydrous sodium sulfate. Hydrated forms will also work. A 0.1 *M* concentration of sodium sulfate is only one option; 0.5 *M*, 1 *M*, 2 *M*, 5 *M*, and even saturated solutions will also work. The higher the concentration the faster the reaction will proceed, because more current will be allowed to flow through the solution.
- Bromthymol blue indicator added to the sodium sulfate solution helps visualize where oxidation and reduction is occurring (change in color as a result of pH change). Add enough indicator solution to give the solution a deep green color. Make sure the solution starts at a neutral pH (green solution). During electrolysis the solution will turn yellow at the anode (where hydrogen ions form) and blue at the cathode (where hydroxide ions form).

Discussion

Electrolysis is a non-spontaneous redox reaction that is made to occur by passing an electric current through a solution under a sufficiently high voltage. For the electrolysis of water to occur, a high enough voltage (or electric potential) needs to be created that will add and remove the electrons from the water molecules, and a current needs to be generated so that electrons have the ability to flow to and from the electrodes. If there is a high enough voltage across the water solution, but the electric current is too low, then electrolysis will not occur. For this reason, electrolytes (ionic salts) are added to the water. The electrical resistance of pure water is too great to allow electrons to flow through it. Adding electrolytes (which dissociate to form positive and negative ions) reduces the resistance and creates an excellent environment for electrons (current) to flow. When a DC power supply is connected to the solution via electrodes, the electrons will flow from the power source to one electrode, which then becomes negatively charged. Since water is a very polar molecule, the negative charge will attract some water molecules. If the voltage across the system is high enough, the water molecules will be *reduced* (gain electrons) at the negative electrode to form hydrogen gas and hydroxide ions (Equation 1). In an electrolysis cell, the electrode where reduction occurs is called the cathode.

$$\text{Reduction at Cathode: } 2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq) \qquad \textit{Equation 1}$$

Meanwhile, the electrons in the other electrode are drawn toward the power supply and are removed, leaving this electrode with a positive charge. This positive charge also attracts some polar water molecules. If the voltage is great enough at this electrode, then electrons are removed from the water molecules and transferred to the positive electrode, producing oxygen gas and hydrogen ions (Equation 2). When water molecules lose electrons, they are said to be *oxidized*. The electrode where oxidation occurs is call the *anode*.

$$\text{Oxidation at Anode: } 2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^- \qquad \textit{Equation 2}$$

Equations 1 and 2 are known as *half-reactions*. These two equations can be added together to obtain the net equation for the electrolysis of water. (Equation 1 needs to be multiplied by 2 in order to balance the electrons produced by Equation 2 with the electrons consumed by Equation 1. The net reaction is a neutral solution. The four hydroxide ions and four hydrogen ions initially formed in the two half-reactions combine to give four water molecules which then cancel out with four water molecules from the reactant side.) The overall chemical equation for the electrolysis of water is shown in Equation 3.

$$\text{Electrolysis of Water: } 2H_2O(l) \rightarrow 2H_2(g) + O_2(g) \qquad \textit{Equation 3}$$

From Equation 3 it can be seen that two moles of water molecules form two moles of hydrogen molecules and one mole of oxygen molecules. This stoichiometric relationship explains why the volume of hydrogen gas produced is twice the volume of the oxygen gas.

Connecting to the National Standards

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

Unifying Concepts and Processes: Grades K–12
- Systems, order, and organization
- 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 and properties of matter, chemical reactions

Reference

Please refer to Electrochemistry, Volume 17, of the *Flinn ChemTopic*™ *Labs* series for additional electrolysis experiments.

Flinn Scientific—Teaching Chemistry™ eLearning Video Series

A video of the *Electrolysis of Water in Living Color* activity, presented by Penney Sconzo, is available in *Electrolysis of Water*, part of the Flinn Scientific—Teaching Chemistry eLearning Video Series.

Materials for *Electrolysis of Water in Living Color* are available from Flinn Scientific, Inc.

Catalog No.	Description
AP5879	Hoffman Electrolysis Apparatus
GP6010	Test Tubes, 13 × 100 mm
AP5375	Power Supply, Multiple Voltage Battery Eliminator
B0173	Bromthymol blue indicator solution, 100 mL
S0251	Sodium sulfate solution, 500 mL

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