

Simple Electrolysis

Electrolysis of Water

Concepts

- Electrochemistry
- Oxidation–reduction
- Anode vs. cathode

Background

Electrochemistry is the study of the relationship between electrical forces and chemical reactions. There are two basic types of electrochemical processes. In a voltaic cell, commonly known as a battery, the chemical energy from a spontaneous oxidation–reduction reaction is converted into electrical energy. In an electrolytic cell, electricity from an external source is used to “force” a nonspontaneous chemical reaction to occur. What chemical reaction will take place when an electric current flows through water?

The first electrochemical process to produce electricity was described in 1800 by the Italian scientist Alessandro Volta, a former high school teacher. Acting on the hypothesis that two dissimilar metals could serve as a source of electricity, Volta constructed a stacked pile of alternating silver and zinc plates separated by pads of absorbent material soaked in saltwater. When Volta moistened his fingers and repeatedly touched the top and bottom metal plates, he experienced a series of small electric shocks. The “voltaic pile,” as it came to be called, was the first battery—a chemical method of generating an electric current. Within months, William Nicholson and Anthony Carlisle in England attempted to confirm the production of electric charges on the upper and lower plates in a voltaic pile using an electroscope. In order to connect the plates to the electroscope, Nicholson and Carlisle added some water to the uppermost metal plate and inserted a wire to the electroscope. To their surprise, Nicholson and Carlisle observed the formation of a gas, which they identified as hydrogen. Nicholson and Carlisle then filled a small tube with river water and inserted wires from the voltaic pile into each end of the tube. Two different gases were generated, one at each wire—Nicholson and Carlisle had discovered electrolysis.

Materials

Bromthymol blue indicator solution, 0.04%, < 1 mL	Bunsen burner
Sodium sulfate, Na ₂ SO ₄ , 0.5 M, < 1 mL	Glass tubing, soft glass, 5 cm OD, 4 cm ID
Water, distilled or deionized	Pipet, Beral-type, thin-stem
Alligator cords, 2	Pencil lead, 2–3 cm, 2
Battery, 9-volt	Rubber band
Beaker, 100 mL	Styrofoam [®] cup

Safety Precautions

To extend the life of the battery, avoid connecting the positive and negative terminals to each other. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Wash hands thoroughly with soap and water before leaving the lab. Please review current Material Safety Data Sheets for additional safety, handling, and disposal information.

Preparation

1. Cut soft glass tubing into 16-cm long pieces.
2. Fire polish both ends of a 16 cm piece.
3. Heat four times the diameter of the tube by rotating the tube around where the bend will be placed. *Note:* If the tube has a 5 mm diameter, heat a 20 mm region in the middle of the tube where the bend will occur (see Figure 1).
4. Once an orange/red flame emerges from the glass it is ductile enough to bend.

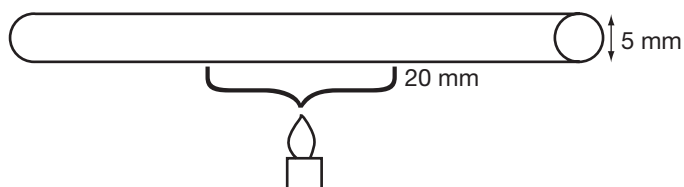


Figure 1.

5. Bend glass into a U-shape with the two open ends facing down as the bend is made.
6. Once the glass has cooled, turn the Styrofoam cup upside down and rubber band the U-tube to the cup (see Figure 2).

Procedure

1. Add 50 mL of distilled or deionized water to a 100-mL beaker.
2. Use a graduated pipet to add 0.5-1.0 mL of 0.5 M Na_2SO_4 to the 50 mL of water and stir.
3. Add 1–2 drops of bromthymol blue to the solution and stir.
4. Fill a thin-stem pipet with solution and insert the tip into the U-tube. *Note:* To reduce the number of air bubbles slowly begin filling from the bottom of the U-tube and gently move the pipet upwards until it is full.
5. Connect an alligator cord to the positive terminal of the battery. Connect the other end to a piece of pencil lead 2–3 cm long.
6. Repeat step 5 using the second alligator cord and the negative battery terminal.
7. Place one alligator cord holding pencil lead in one opening of the U-tube and the other in the second opening of the U-tube.
8. Allow current to flow through the solution for about 5 minutes or until chemical changes are evident.

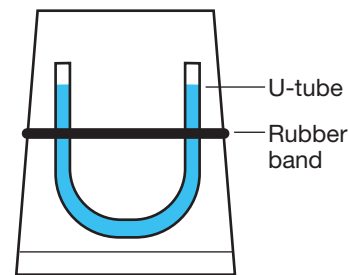


Figure 2.

Disposal

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

Tips

- Sodium sulfate is used as a source of dissolved ions to increase the current flow through the solution. In the absence of added electrolytes, no reaction will occur when the battery is connected to the pencil leads—there are no ions to “carry” the current through the solution. The rate of electrolysis increases as the concentration of sodium sulfate increases (compare 0.1 M and 1 M solutions). The conductivity of a sodium sulfate solution versus pure water may be demonstrated using a conductivity tester (available from Flinn Scientific, Catalog No. AP5355).
- The sulfate ion is an extremely weak base (the $\text{p}K_a$ for its conjugate acid, HSO_4^- , is 2.0). The initial indicator color for the electrolysis solution may be more blue-green rather than green. According to the *Merck Index*, the pH of a sodium sulfate solution is 6.0–7.6. Test a small amount of the sodium sulfate stock solution with bromthymol blue indicator before class—the solution should turn green. If the solution is blue, add one drop of 1 M hydrochloric acid to the stock solution. If the solution is yellow, add one drop of 1 M sodium hydroxide to the stock solution.
- The sodium sulfate electrolysis solution may be reused by several classes during the day. Recycling the solution in this way means the indicator will already be present—instruct students to omit step 3 in the *Procedure*. Adjust the pH of the electrolysis solution as described above, if necessary.
- A large, demonstration-scale version of this experiment uses the Hoffman electrolysis apparatus. Using the Hoffman apparatus makes it possible to measure the volume of gas generated at each electrode, collect the gases, and test their properties. If the time and current are measured, the amount of hydrogen gas collected may also be used for quantitative calculations of the Faraday constant.

Connecting to the National Standards

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

Unifying Concepts and Processes: Grades K–12

Constancy, change, and measurement

Content Standards: Grades 9–12

Content Standard A: Science as Inquiry

Content Standard B: Physical Science, structure and properties of matter, chemical reactions

Flinn Scientific—Teaching Chemistry™ eLearning Video Series

A video of the *Simple Electrolysis* activity, presented by John Little, is available in *Electrolysis of Water* and in *Decomposition Reactions*, part of the Flinn Scientific—Teaching Chemistry eLearning Video Series.

Materials for *Simple Electrolysis* are available from Flinn Scientific, Inc.

Catalog No.	Description
GP9020	Glass Tubing, Soft Glass, 24"
AP6052	Alligator Cords
AP1817	Pencil Leads
S0353	Sodium Sulfate Solution, 0.5 M, 500 mL
B0713	Bromthymol Blue Indicator Solution, 100 mL

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