

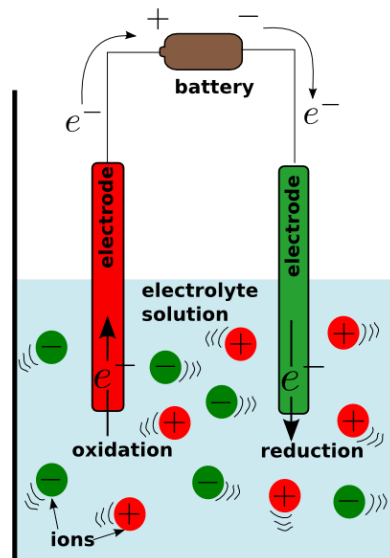
# Electrolysis: Splitting Water

## Student Version

In this lab you will use a battery to perform electrolysis, or chemical decomposition, of different aqueous solutions (like water) to produce gases (like hydrogen and oxygen in the case of water). You will measure the volumes of gas produced and compare this to the predicted ratios from chemical equations. Finally, you will explore an industrial application of electrolysis using metal electrodes.

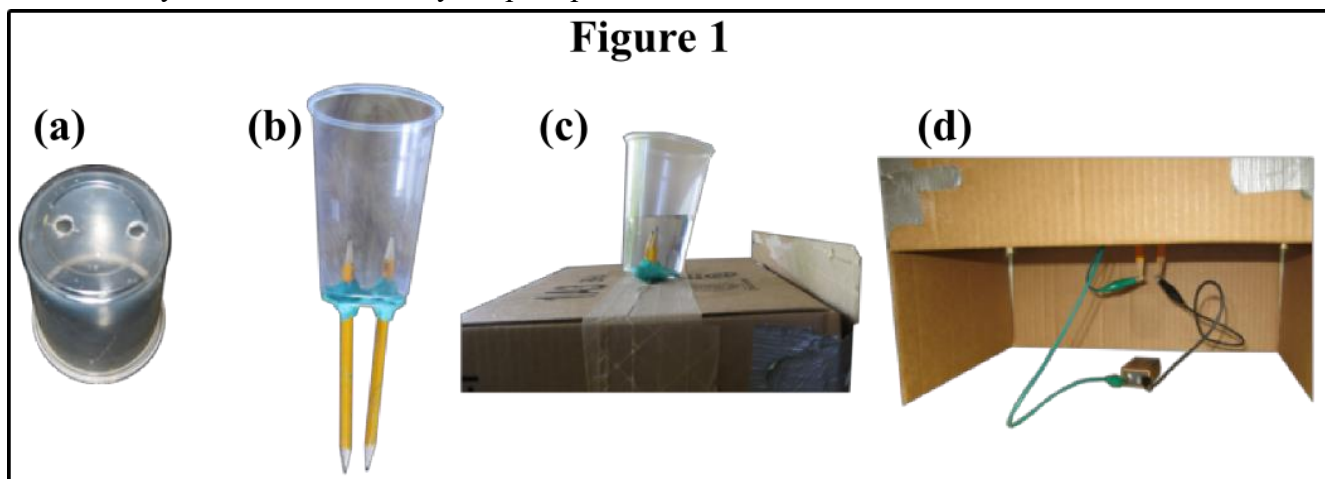
### Key Concepts:

- **Electrolysis** is the process by which electricity is used to drive a chemical reaction.
- A chemical reaction where some molecule gains electrons is known as a **reduction reaction**.
- A chemical reaction where some molecule loses electrons is an **oxidation reaction**.
- An **electrolyte solution** is one in which ions (charged particles) are dissolved in water.
- Electric current will flow only if there is a continuous circuit of conducting material to carry it.
- Which chemical reaction occurs in an electrolytic cell depends on what molecules are present and how easily they gain or accept electrons. A table of **standard reduction potentials** is useful for summarizing which molecules are most easily oxidized or reduced.
- Writing down balanced chemical reactions is useful for figuring out relative quantities of products formed.



## Part 1 – Electrolysis of water

In this section, you will use electricity to split apart water molecules!



1. Set up the electrolysis apparatus as shown in the picture.
  - a) Pierce two round holes in the bottom of a plastic cup. (Figure 1a)
  - b) Sharpen two wooden pencils on both ends (after pulling off the eraser). Insert the pencils into the holes in the cup. Here should be about 1 inch of each pencil sticking into the cup.

- c) Pack modeling clay around the pencils, both outside and inside the bottom of the cup to make a watertight seal. Make sure the cup can hold water without leaking. (Figure 1b)
- d) On the cup, use a marker to mark the side corresponding to one pencil as positive (+) and the other as negative (-).
- e) Cut one side off a shoebox.
- f) Pierce two holes in the top of the shoebox and insert the pencils, so that the cup can sit on top of the box. (Figure 1c)
- g) Connect an alligator clip to the bottom of each pencil (Figure 1d). The other ends of the wires will be used to connect to the battery, but do not attach the battery yet.

2. Initial test of apparatus with water.

- a) Pour enough bottled or tap water into the cup to cover the pencil leads.
- b) Connect the alligator clips to the two terminals of the battery. Make sure the side marked “+” gets connected to the positive terminal of the battery and the side marked “-” gets connected to the negative terminal.

Q1. *What, if anything, do you observe around the pencil leads?*

Q2. *Do you think any current is flowing out of the battery? Why or why not?*

3. Mix up an **electrolyte solution** (a solution of charged particles) by stirring together:

**3/2 cups water + 4 teaspoon Epsom Salt**

4. Test apparatus using Epsom Salt electrolyte solution.

- a) Pour out the bottled water currently in your cup. Pour in enough electrolyte solution to cover the pencil leads
- b) Hook up your battery to the pencil leads as before.

Q3. *What do you now see happening at the pencil tips?*

Q4. *Do you think there is now current flowing out of the battery?*

Q5. *Epsom Salt in solution breaks up into  $Mg^{2+}$  and  $SO_4^{2-}$  ions. Why did dissolving this salt in the water make a difference?*

- c) Disconnect the battery. Label your two test-tubes (+) and (-).
- d) Take a test-tube and fill it with more of your Epsom Salt solution. Place a finger (or cap) over the end of the tube, and flip it upside down into the cup. The test tube should remain mostly filled

with solution (a little air at the top is not a problem).

- e) Slip the test tube over one of the pencil leads inside the cup (make sure the +/- labels match up). You can set one edge in the modeling clay to prevent it from slipping, but be careful you do not make a seal that prevents water from flowing from the test-tube to the rest of the cup. Repeat with the 2<sup>nd</sup> test-tube.

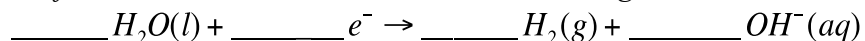
*Q6. Why is it important not to press the test-tube into the modeling clay, sealing off the bottom?*

- f) Use a thin marker to carefully mark off the current level of water inside each test-tube. Be as precise as possible, as you will be measuring very small changes!  
g) Hook up the battery with the alligator clips as before.  
h) Set a timer for 15 minutes. When the timer goes off, unhook the battery. In the meantime, answer the questions in step 5.

5. You have just set up an **electrolytic cell**. The graphite leads in the pencils serve as **electrodes**. They conduct electric current from the battery into the solution.

- a) At the negative terminal, the electrons are pumped through the pencil core and into solution. Once in solution, they react with water molecules. A reaction that involves adding electrons is called a **reduction reaction**.

*Q7. Balance the following chemical equation for the reduction of water. There must be the same number of each atom on each side, and the total charge must be the same on each side.*



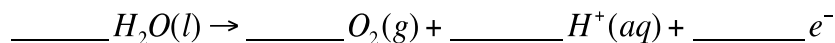
The symbols in parentheses tell you what form the molecule takes.

(l) means liquid, (g) means gas, (s) means solid, (aq) means **aqueous** or dissolved in water

*Q8. What are the bubbles being formed at the pencil connected to the negative terminal?*

- b) At the positive terminal, the battery works to suck electrons out of solution. At this terminal, water molecules lose their electrons. This type of reaction is called an **oxidation reaction**.

*Q9. Balance the chemical equation for the oxidation of water.*



*Q10. What are the bubbles being formed at the pencil connected to the positive terminal?*

*Q11. What would you expect for the ratio of how much gas is formed at the (-) terminal to how much is formed at the (+) terminal? Hint: Fill in the blanks below.*

*For every \_\_\_\_\_ electrons that go through the circuit, 1 molecule of O<sub>2</sub> is made. However, for the same number of electrons you can make \_\_\_\_\_ molecules of H<sub>2</sub>. Thus there will be \_\_\_\_\_ times as much hydrogen as oxygen.*

Ratio of gas at the (-) side to gas at the (+) side = \_\_\_\_\_

6. Check the results of the electrolysis.
- After you disconnect the battery, carefully mark the level of the water in each test-tube.
  - Take out the test-tubes. Measure the difference between the start and end marks, to the nearest 0.1 cm. Fill in the table below.

	change in height (cm)
<b>H<sub>2</sub> (negative terminal)</b>	
<b>O<sub>2</sub> ( positive terminal)</b>	

Measured ratio of H<sub>2</sub> to O<sub>2</sub> gas: \_\_\_\_\_

Q12. Was your prediction in step 5 correct?

- Empty out your electrolysis cup. Look carefully at the two pencil leads.

Q13. Do you think either O<sub>2</sub> gas or H<sub>2</sub> gas reacted at all with graphite?

## Part 2 – Industrial Applications

In this part you will see how industrial chemists use electrolysis to produce useful chemicals.

- Mix up a different electrolyte solution by stirring together until fully dissolved:

**3/2 cup water + 2 tsp table salt**

- Test apparatus using table salt electrolyte solution.
  - Pour the saltwater solution into the cup with the pencils, and set up your electrolysis cell as before. Make a mark on each test-tube for the starting level of the water. Connect the battery.
  - Set a timer for 5 min. At the end of this time, disconnect the battery. Meanwhile, go to step 3.
- Table salt has the chemical formula NaCl. In water it breaks up into Na<sup>+</sup> and Cl<sup>-</sup> ions. The Cl<sup>-</sup> ions lose electrons more easily (they are more readily oxidized) than do water molecules.

Q14. Balance the chemical equation for the oxidation of chlorine ions



Q15. What gas do you expect to be produced at the negative terminal?

Q16. What gas should be produced at the positive terminal?

Q17. What would you expect for the ratio of gases produced at each terminal?

volume of gas at (+) terminal : volume of gas at (-) terminal = \_\_\_\_\_

4. After you disconnect the battery, take a look at how much gas was actually produced at each electrode. Make a mark for the current level of the water in each test-tube.

Q18. Change in height at the negative terminal:

Change in height at the positive terminal:

Did you see the ratio that you expected?

Q19. Was one gas produced in much greater quantity than the other?

Q20. Make a guess as to where the missing gas might have gone:

5. Electrolysis of salt water is used on an industrial scale to produce chlorine gas. Chlorine gas is toxic if breathed in large quantities. In fact, chlorine gas was used as a chemical weapon in World War 1, so be careful with this step!

DO NOT DO THIS STEP IF YOU HAVE ASTHMA OR RESPIRATORY PROBLEMS.

You will test that chlorine gas really is being produced at one of the electrodes.

Take the test-tube off the electrode that produced chlorine gas, keeping a finger over the top to minimize gas escaping. Hold the test-tube at arms length. And **using proper safe smelling technique**, release the gas and waft it towards you.

Q21. What, if anything, do you smell?:

(i) rotten eggs (ii) bananas (iii) bleach (iv) alcohol (v) nothing

6. Electrolysis of salt water (the reaction you just carried out) is used on an industrial scale to produce chlorine gas.

Q22. List some reasons producing chlorine gas is useful:

7. The same electrolysis setup is also used to produce **lye** (NaOH), which is used for oven cleaner, drain cleaner, and soap.

The sodium ions (Na<sup>+</sup>) are floating around everywhere.

Q23. Which electrode creates the OH<sup>-</sup> ions to form lye?



## Part 3 – Metal Electrodes

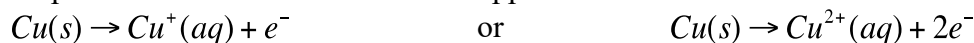
What happens if we use metal electrodes instead of the graphite core of pencils? We will try using pennies, which are coated with a layer of copper (Cu).

1. In a clean cup, mix up a different electrolyte solution.

**1/2 cup water + 3 tsp Epsom salt**

2. Setup and test the apparatus with metal electrodes.

Copper metal is more easily oxidized than water (copper does not hang on strongly to its electrons). There are two possible oxidation reactions for copper.



- a) Attach alligator clips to each of two shiny, clean pennies.
- b) Place the pennies in the Epsom Salt solution, making sure they do not touch.

*Q24. Do you expect to see bubbles at the penny connected to the positive terminal, the penny connected to the negative terminal, both, or neither?*

*Q25. Once you hook up the battery, why is it important that the pennies not be touching*

- c) Hook up the other ends of the alligator clips to the terminals of the 9V battery. Let the reaction run for 3-5 minutes.

*Q26. Do you see bubbles where you expected to find them?*

*Q27. Do you see any color change in the solution?*

4. When copper metal is oxidized, it can form both  $\text{Cu}^+$  and  $\text{Cu}^{2+}$  ions.  $\text{Cu}^{2+}$  ions are light blue.  $\text{Cu}^+$  can react with oxygen in the air to form a yellow compound ( $\text{Cu}_2\text{O}$ ).

*Q28. Did your oxidation reaction produce  $\text{Cu}^+$ ,  $\text{Cu}^{2+}$  or a combination of the two?*

5. Copper ions can combine with chloride ions ( $\text{Cl}^-$ ) to form yellow-brown copper-chloride complexes.

*Q29. Using materials in this lab, can you figure out a way to make these complexes? Set up the experiment – do you see the copper chloride being formed?*