Electrolysis: Splitting Water
Teacher 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.

California Science Content Standards:

- **3. Conservation of Matter and Stoichiometry: The conservation of atoms in chemical reactions leads to the principles of conservation of matter and the ability to calculate the mass of products and reactants.**
- 3a. Students know how to describe chemical reactions by writing balanced equations.
- 3b. Students know the quantity one mole is set by defining one mole of carbon 12 atoms to have a mass of exactly 12 grams.
- 3c. Students know one mole equals 6.02×10^{23} particles (atoms or molecules).
- 3d. Students know how to determine the molar mass of a molecule from its chemical formula and a table of atomic masses and how to convert the mass of a molecular substance to moles, number of particles, or volume of gas at standard temperature and pressure.
- 3e. Students know how to calculate the masses of reactants and products in a chemical reaction from the mass of one of the reactants or products and the relevant atomic masses.
- **3g. Students know how to identify reactions that involve oxidation and reduction and how to balance oxidation-reduction reactions.
- **6. Solutions: Solutions are homogeneous mixtures of two or more substances.**
- 6a. Students know the definitions of solute and solvent.
- **7. Chemical Thermodynamics: Energy is exchanged or transformed in all chemical reactions and physical changes of matter.**
- 7a. Students know how to describe temperature and heat flow in terms of the motion of molecules/atoms.

Prerequisites:

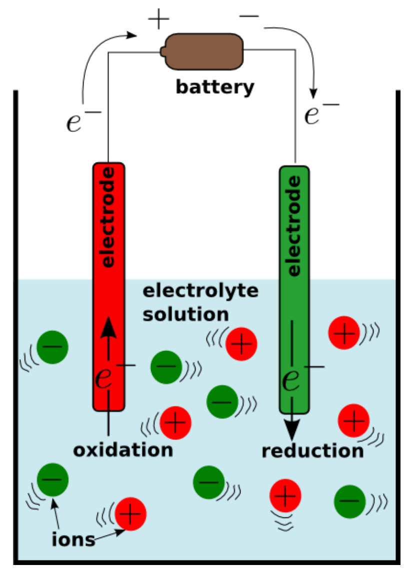
- Students are expected to know calculator arithmetic and unit conversions.
- Familiarity with chemical formulas, chemical reactions, and the periodic table is a plus.
- Some manual dexterity is required to build an apparatus.
- Suggested grade level: 9-12

Complete List of Materials (for each group):

- 9V battery
- 2 wires with alligator clips on either end
- 2 wooden pencils, sharpened at both ends (eraser removed)
- 2 test-tubes, wide enough to fit the pencils with room to spare (length doesn't matter)
- Shoebox, or comparable box
- Scissors with sharp tips
- Modeling clay
- Marker for labeling
- Large cup or bowl for mixing

- 2 plastic cups, 12-16 oz, preferably clear
- 2 shiny, clean pennies
- Ruler, with mm markings
- Epsom salt (about 6 tsp); this is sold at most drug stores, usually as a laxative
- Table salt (about 6 tsp)
- Water
- Optional: ½ cup and 1 tsp measuring spoon; exact measurements don't matter so can approximate

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.

Introductory Mini-Lecture:

Electrolysis refers to the use of electricity to drive a chemical reaction that would not normally occur on its own. In this lab, you will build an **electrolytic cell** – an apparatus for carrying out electrolysis and use it to produce various gases and other chemical compounds. A diagram of a basic electrolytic cell is shown here.

The battery drives an electric current through the cell. If the battery is not hooked up to anything, no electricity flows. In order for a current to flow, there must be a complete circuit of conducting material from one terminal of the battery to the other. In the setup shown, electrons flow out of the **negative terminal** of the battery, through a wire into the negative **electrode.** The electrode can be any good conductor (eg: a metal) that is stuck into the electrolytic cell solution. Electrons also flow out of the positive electrode into the **positive terminal** of the battery. But how is electric current carried from one electrode to the other? The electrolytic cell is filled with an **electrolyte solution** – a solution of **ions** (charged particles) in water. Because these charged particles can move, they are capable of carrying current across from one electrode to another. Thus, there is a complete circuit for current to flow!

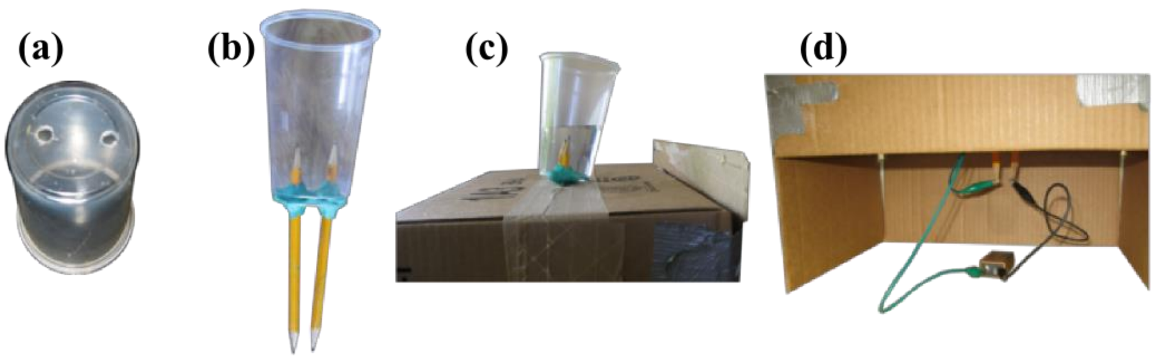
When we run current through an electrolytic cell, two chemical reactions occur. At the negative electrode, there is a **reduction reaction** – some molecule gains electrons. At the positive electrode, there is an **oxidation reaction** – some molecule loses electrons. These two chemical reactions can produce all sorts of useful products – like hydrogen gas for fuel cells, or purified metals from minerals. In this lab, we're going to start by using electrolysis to split water molecules (H_2O) into hydrogen and oxygen gas.

Incidentally, an electrolytic cell is the exact opposite of a chemical battery (see the Juice Battery Lab). In a battery, the oxidation and reduction reactions occur spontaneously to suck electrons out of one electrode and pump electrons into the other. This process forces current to flow, and so the battery becomes a source of electricity. In an electrolytic cell, we supply the power driving the current with an external battery and use that to force a chemical reaction to occur.

Part 1 – Electrolysis of water

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

Figure 1



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 <u>do not attach</u> the battery yet.

2. Initial test of apparatus with water.
   a) Pour enough **bottled 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?*
   Most likely students will see nothing; there may be a very small amount of bubbles near the pencil tips.

   Q2. Do you think any current is flowing out of the battery? Why or why not?
   No, because there are no ions to carry charge and complete the circuit.

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?
   There should be bubbles formed at the tips of both pencils.

   Q4. Do you think there is now current flowing out of the battery?
   Yes. The flow of electric current is what is causing the bubbles to be made.

   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?
   The ions in the water can move. This makes the electrolyte solution a good conductor of electricity and completes the circuit.

   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 2nd test-tube.

   Q6. Why is it important not to press the test-tube into the modeling clay, sealing off the bottom?
   Current would not be able to flow out of the test-tube, so the circuit would not be complete.

   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.

$$\boxed{\_\_2\_\_ H_2O(l) + \_\_2\_\_ e^- \rightarrow \_\_1\_\_ H_2(g) + \_\_2\_\_ OH^-(aq)}$$

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?
hydrogen gas (H_2)

b) At the positive terminal, the battery works to sucks 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.

$$\boxed{\_\_2\_\_ H_2O(l) \rightarrow \_\_1\_\_ O_2(g) + \_\_4\_\_ H^+(aq) + \_\_4\_\_ e^-}$$

Q10. What are the bubbles being formed at the pencil connected to the positive terminal?
oxygen gas (O_2)

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 ____4____ electrons that go through the circuit, 1 molecule of O_2 is made. However, for the same number of electrons you can make ____2____ molecules of H_2. Thus there will be ____2______ times as much hydrogen as oxygen.

Ratio of gas at the (–) side to gas at the (+) side = ______2:1_______

6. Check the results of the electrolysis.
   a) After you disconnect the battery, carefully mark the level of the water in each test-tube.
   b) 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)
H_2 (negative terminal)	*Will be around 1 cm*
O_2 (positive terminal)	*Will be around 0.5 cm*

Measured ratio of H_2 to O_2 gas: __________*should be around 2*_____________

Q12. Was your prediction in step 5 correct?

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

Q13. Do you think either O_2 gas or H_2 gas reacted at all with graphite ?
The pencil tip on the (+) side will be somewhat worn down because O_2 reacts with graphite.

Part 2 – Industrial Applications

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

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

3/2 cup water + 2 tsp table salt

2. Test apparatus using table salt electrolyte solution.
   a) 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.
   b) Set a timer for 5 min. At the end of this time, disconnect the battery. Meanwhile, go to step 3.

3. Table salt has the chemical formula NaCl. In water it breaks up into Na^+ and Cl^- ions. The Cl^- 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

$$\boxed{\_\_2\_\_ Cl^-(aq) \rightarrow \_\_1\_\_ Cl_2(g) + \_\_2\_\_ e^-}$$

   Q15. What gas do you expect to be produced at the negative terminal?
   Hydrogen gas (H_2). Remember, reduction occurs at the negative terminal.

   Q16. What gas should be produced at the positive terminal?
   Chlorine gas (Cl_2), since chlorine ions are more easily oxidized than water.

   Q17. What would you expect for the ratio of gases produced at each terminal?
   volume of gas at (+) terminal : volume of gas at (–) terminal = _______1:1__________
   In this case, you need the same number of electrons going through the circuit to make one molecule of hydrogen gas as to make one molecule of chlorine gas. So they will be produced in equal amounts.

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: will be about 2 cm

   Change in height at the positive terminal: will be very small

   Did you see the ratio that you expected? No

   Q19. Was one gas produced in much greater quantity than the other?
   You will see much more H_2 than Cl_2

   Q20. Make a guess as to where the missing gas might have gone:
   The chlorine gas is dissolved in the water.

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 <u>be careful</u> 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 <u>(iii) bleach</u> (iv) alcohol (v) nothing

TEACHER NOTE
*SAFETY INSTRUCTIONS: Instruct students to hold the test-tube at arm's length, still keeping a finger over the top. Then release the gas and gently **waft** it towards their nose. Students should **not** put their nose directly over the tube to inhale the gas.*

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: it's used in bleach, pool disinfectant

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^+) are floating around everywhere.

   Q23. Which electrode creates the OH^- ions to form lye? The negative electrode

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 elecrodes.
   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.

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

   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?
   We expect to see bubbles only at the negative terminal, where hydrogen gas is still being formed. At the positive terminal, copper will be oxidized instead of water, so no gas is formed.

   Q25. Once you hook up the battery, why is it important that the pennies not be touching?
   If they are touching, the current will flow straight between the metal pennies and not through the solution at all. You will not see any chemical reaction, and you could easily short out the battery because pennies conduct electricity very well.

   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? Yes, hopefully

   Q27. Do you see any color change in the solution?
   The solution will likely turn some shade of yellow or light blue

4. When copper metal is oxidized, it can form both Cu^{+} and Cu^{2+} ions. Cu^{2+} ions are light blue. Cu^{+} can react with oxygen in the air to form a yellow compound (Cu_2O).

   Q28. Did your oxidation reaction produce Cu^{+}, Cu^{2+} or a combination of the two?
   Either or a combination is possible

5. Copper ions can combine with chloride ions (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?
   Just repeat the reaction in a solution of table salt (NaCl) instead of Epsom salt. Make sure to stir it after it runs. You should see a very noticeable brownish color appearing.