

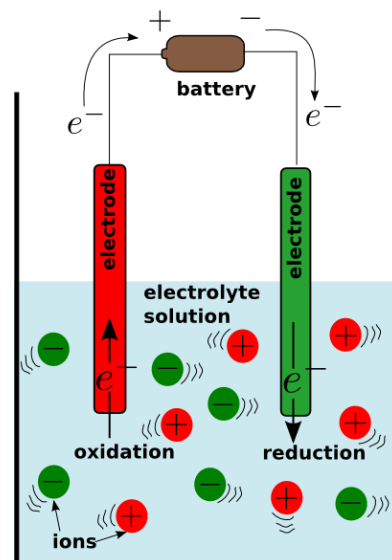
Electrolysis: Splitting Water

Student Advanced 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.

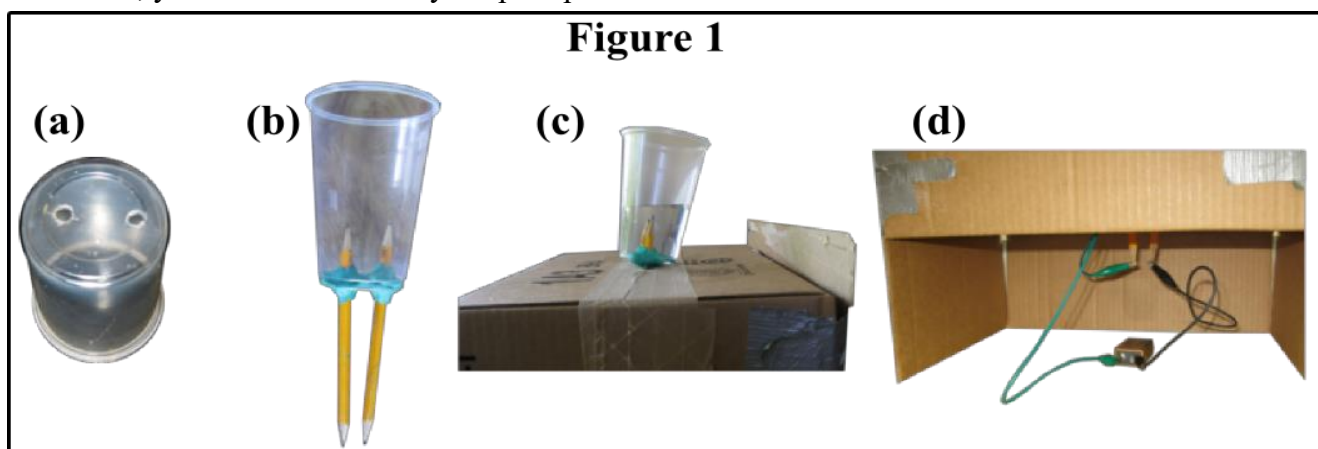
Key Concepts:

- **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)
- Mark the side of the cup next to one pencil as (+) and the other side as (-)
 - Cut one side off a shoebox.
 - 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)
 - 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.

- Pour enough **bottled water** into the cup to cover the pencil leads.
- Connect the alligator clips to the two terminals of the battery

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

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

3/2 cups water + 4 teaspoons Epsom Salt

4. Test apparatus using Epsom Salt electrolyte solution.

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

Q2. What do you see now happening at the pencil tips?

Q3. 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?

- Label your two test-tubes as (+) and (-).
- 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).
- Slip the test tube over one of the pencil leads inside the cup. Make sure the (+) and (-) 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.

Q4. Why is it important not to press the test-tube into the modeling clay, sealing off the bottom?

- 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!
- Hook up the battery directly to the pencil leads. Make sure to match up the (+) and (-) sides.
- Set a timer for **15 minutes**. When the timer goes off, unhook the battery. In the meantime, do the

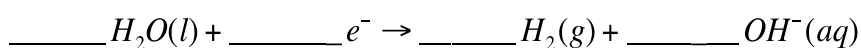
calculations in step 5.

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

a) The negative terminal of the battery is connected to the **cathode**. Electrons are pumped into this electrode. Once they reach the solution, they participate in a **reduction reaction** – a chemical reaction where some chemical species gains electrons.

In the Epsom Salt solution, there are two species that could be grabbing the electrons: the Mg^{2+} ions or the water molecules themselves.

Q5. Balance the following possible reduction reactions. There must be the same number of each atom on each side, and the total charge must be the same on each side.



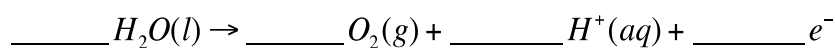
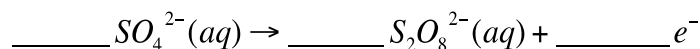
Look at the table of standard reduction potentials at the end of your lab. Which reaction is preferred? (i.e. is Mg^{2+} or H_2O more easily reduced)? Circle the appropriate equation above.

Q6. What are the bubbles being formed at the cathode (the pencil connected to the negative terminal)?

b) The positive terminal of the battery is connected to the **anode**. Electrons are pumped into this electrode. Once they reach the solution, they participate in a **oxidation reaction** – a chemical reaction where some chemical species loses electrons.

In the Epsom Salt solution, there are two species that could be giving up electrons: the SO_4^{2-} ions or the water molecules themselves.

Q7. Balance the following possible oxidation reactions.



Look at the table of standard reduction potentials at the end of your lab. Which reaction is preferred? (i.e. is SO_4^{2-} or H_2O more easily oxidized)? Circle the appropriate equation above.

Q8. What are the bubbles being formed at the anode (the pencil connected to the negative terminal)?

Q9. What would you expect for the ratio of how much gas is formed at the cathode to how much is formed at the anode? Hint: electrons are pumped into the cathode at the same rate as they are pumped out at the anode, so look at the relative number of electrons in your circled reactions.

_____ units of gas at the cathode : _____ units of gas at the anode

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)	volume of gas (cm ³ = mL)
H ₂ (cathode, negative terminal)		
O ₂ (anode, positive terminal)		

- Measure the diameter of the test-tube, to the nearest 0.1 cm. Divide by 2 to get the radius.
Radius of test tube = _____ cm

d) In the table, fill in the volume of each gas using: $V = \pi R^2 h$

e) Now compare to our predictions:

Q10. Ratio of H₂ to O₂ gas: _____

Part 2 – Industrial Application

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⁺ and Cl⁻ ions.
 - At the **cathode** (negative battery terminal) you will again have **reduction**, or addition of electrons to some chemical species.

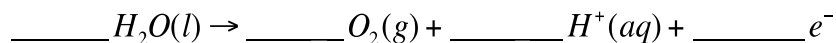
Q11. Balance the following possible reduction reactions:



Using the table of reduction potentials, circle which reaction is actually going to occur.

- b) At the **anode** (positive battery terminal) there will be **oxidation**, or removal of electrons from some chemical species.

Q12. Balance the following possible oxidation reactions:



Using the table of reduction potentials, is one reaction much more likely than the other or are they pretty close?

It turns out that in a concentrated NaCl solution, **chlorine is reduced** in preference to water. This has to do with more complicated non-equilibrium chemistry that isn't included in the table of reduction potentials.

Q13. What gas do you expect to be produced at the cathode (negative terminal)?

Q14. What gas should be produced at the anode (positive terminal)?

Q15. What would you expect for the ratio of gases produced at each terminal?

volume of gas at cathode : volume of gas at anode = _____

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.

Q16. Change in height at the cathode (negative terminal): _____ cm

Change in height at the anode (positive terminal): _____ cm

Do you see the ratio that you expected?

Q17. Was one gas produced in much greater quantity than the other? Make a guess as to why this might be.

5. Chlorine gas is highly soluble in water. As the gas is produced, it will dissolve in the water until the test-tube is fully saturated. Let's calculate how much chlorine gas you produced and whether it was enough to fully saturate the tube of water.

Q18. Find the volume of H₂ gas produced at cathode ($V = \pi R^2 h$):

One mole of an ideal gas takes up 22.4 L at room temperature and pressure.

Q19. Calculate how many grams of chlorine gas were produced.

First calculate how much hydrogen gas was produced.

$$\left(\text{_____ mL H}_2 \right) \times \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) \times \left(\frac{1 \text{ mol H}_2}{22.4 \text{ L}} \right) = \text{_____ mol H}_2 \text{ gas produced}$$

Expected number of moles of Cl₂ gas produced (use the expected ratio of H₂:Cl₂):

Look up the molar mass of Cl in the periodic table

Molar mass of Cl: _____ g/mol

Calculate mass of chlorine gas produced (remember to multiply by 2 to get g/mol of Cl₂):

$$\left(\text{_____ mol Cl}_2 \right) \times \left(\frac{\text{g Cl}_2}{\text{mol Cl}_2} \right) = \text{_____ g Cl}_2 \text{ produced}$$

Q20. Calculate how much Chlorine would be needed to saturate the entire test-tube of water.

Approximately how much water is in the test-tube? _____ mL

Calculate mass of water in the test-tube:

$$\left(\text{_____ mL} \right) \times \left(\frac{1 \text{ g}}{1 \text{ mL}} \right) \times \left(\frac{1 \text{ kg}}{1000 \text{ g}} \right) = \text{_____ kg water}$$

Using the solubility chart in the back of your lab, how much chlorine would be required to saturate the entire test-tube?

$$\left(\text{_____ kg water} \right) \times \left(\frac{\text{g Cl}_2}{\text{kg water}} \text{ solubility} \right) = \text{_____ g Cl}_2 \text{ to saturate solution}$$

Q21. Did you make enough Cl₂ gas to saturate the solution?

Q22. Why did you see so little gas at the anode (positive terminal)?

Q23. In part 1, why did you see the oxygen and hydrogen gases actually produced – why didn't most of those dissolve into the water as well? Hint: look at the solubility charts provided.

6. 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.

We 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.

Q24. *What, if anything, do you smell?:*

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

Q25. *List some reasons producing chlorine gas on an industrial scale 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 are floating around everywhere.

Q26. *Which electrode creates the OH⁻ ions that form lye? Circle one: cathode / anode*

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.
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.

Use the reduction potential table at the back of your lab to make some predictions.

Q27. *What do you expect to be reduced at the cathode (negative battery terminal)?*

(i) H₂O (ii) Mg²⁺

Q28. *Write down the balanced chemical equation for the reduction reaction:*

Q29. *What do you expect to be oxidized at the anode (positive battery terminal)?*

(i) H₂O (ii) SO₄²⁻ (iii) Cu(s)

Q30. Write down the balanced chemical equation for the oxidation reaction:

Q31. Do you expect to see bubbles at the positive terminal, negative terminal, both or neither?

- 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.

Q32. Do you see bubbles at the electrode(s) where you expected to find them?

Q33. Why is it important that the pennies not be touching?

Q34. Do you see any color change in the solution?

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 to form a yellow compound (Cu_2O).

Q35. Did your oxidation reaction produce Cu^+ , Cu^{2+} or a combination of the two?

Q36. When industrial chemists want to make chlorine and lye through electrolysis, why can't they not use electrodes made of common metals, such as silver or iron or zinc?

Table of Standard Reduction Potentials in Aqueous Solution

This table tells you the electrical energy given off by the **reduction** reaction, under standard conditions (25°C, 1 M concentrations of everything). Negative values means energy has to be put in to drive the reduction reaction. High values of E^0 means the chemical species is very easily reduced.

An **oxidation** reaction is the reverse of a reduction reaction. Flipping the reaction also flips the sign of the energy. So oxidation reaction towards the bottom of the table occur more easily than those towards the top. For instance reduction of Cl_2 to chloride ions gives off a lot of energy because chlorine loves electrons. Oxidation of Cl^- ions back to Cl_2 would instead require a lot of energy.

	Reduction Half-Reaction	Standard Potential E^0 (volts)
↑ loves electrons	$\text{S}_2\text{O}_8^{2-}(\text{aq}) + 2\text{e}^- \rightarrow 2\text{SO}_4^{2-}(\text{aq})$	2.01
	$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Fe}^{2+}(\text{aq})$	1.36
easily reduced	$2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$	1.23
	$\text{ClO}^-(\text{aq}) + \text{H}_2\text{O}(\text{l}) + 2\text{e}^- \rightarrow \text{Cl}^-(\text{aq}) + 2\text{OH}^-(\text{aq})$	0.90
	$\text{ClO}_2^-(\text{aq}) + \text{H}_2\text{O}(\text{l}) + 2\text{e}^- \rightarrow \text{ClO}^-(\text{aq}) + 2\text{OH}^-(\text{aq})$	0.59
	$\text{Cu}^+(\text{aq}) + \text{e}^- \rightarrow \text{Cu}(\text{s})$	0.52
	$\text{ClO}_3^-(\text{aq}) + \text{H}_2\text{O}(\text{l}) + 2\text{e}^- \rightarrow \text{ClO}_2^-(\text{aq}) + 2\text{OH}^-(\text{aq})$	0.35
	$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s})$	0.34
	$\text{AgCl}(\text{s}) + \text{e}^- \rightarrow \text{Ag}(\text{s}) + \text{Cl}^-(\text{aq})$	0.22
	$\text{ClO}_4^-(\text{aq}) + \text{H}_2\text{O}(\text{l}) + 2\text{e}^- \rightarrow \text{ClO}_3^-(\text{aq}) + 2\text{OH}^-(\text{aq})$	0.17
	$\text{Cu}^{2+}(\text{aq}) + \text{e}^- \rightarrow \text{Cu}^+(\text{aq})$	0.16
	$2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$	0.00
	$2\text{H}_2\text{O}(\text{l}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g}) + 2\text{OH}^-(\text{aq})$	-0.83
	$\text{Mg}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Mg}(\text{s})$	-2.38
	$\text{Na}^+(\text{aq}) + \text{e}^- \rightarrow \text{Na}(\text{s})$	-2.71

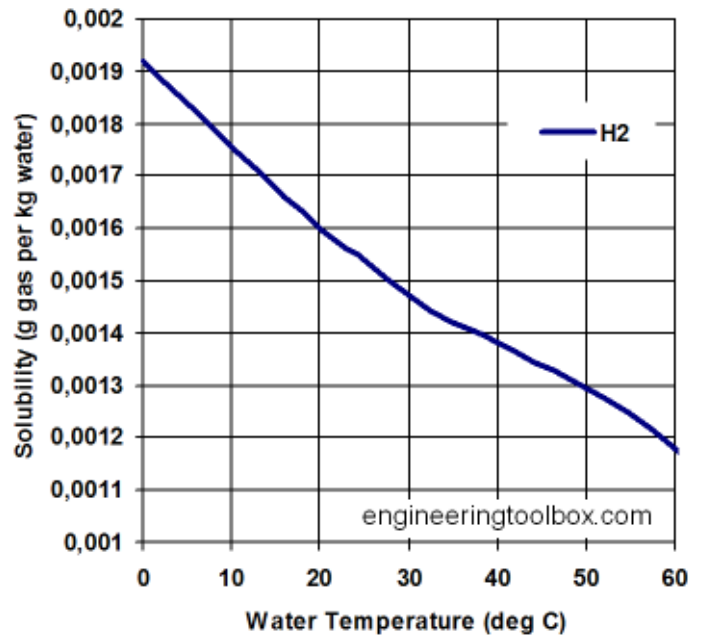
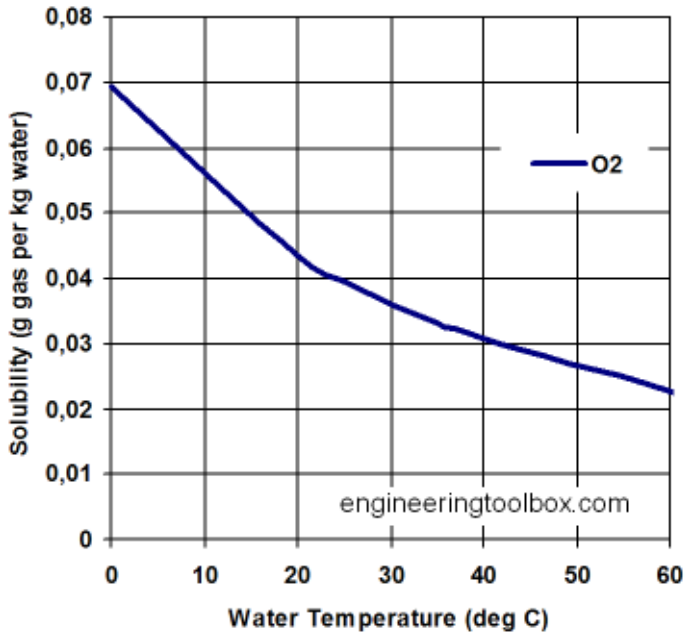
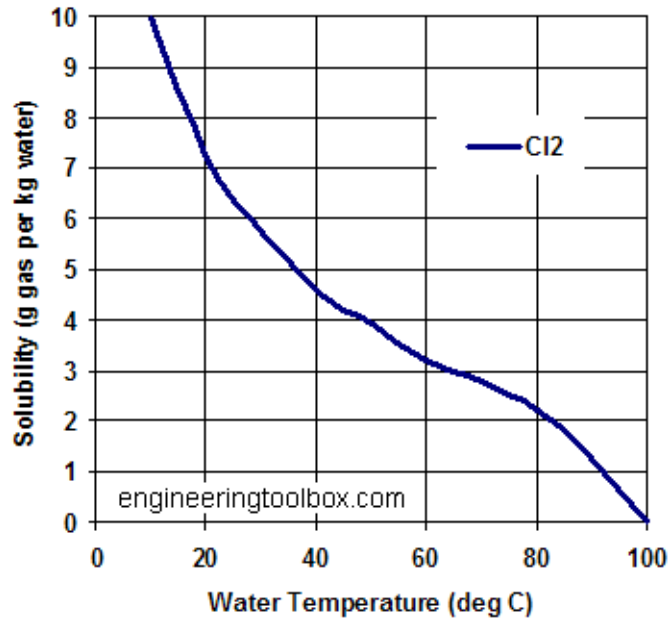
hard to reduce

easily oxidized

hates electrons

Solubility of Gases in Water

These plots tells you how many grams of gas are required to saturate 1 kg of water, at different temperatures. Keep in mind, room temperature is about 25°C.



1	1A	18
1	2	8A
Hydrogen 1.01	Helium 4.00	
2	3A	13
3	4A	14
Li 6.94	B 10.81	15
Be 9.01	C 12.01	16
12	N 14.01	17
11	O 16.00	18
Na 22.99	F 19.00	19
Mg 24.31	Ne 20.18	20
3	5A	35
3B	6A	36
21	7A	53
22	8A	86
23		
24		
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Key

11 — Atomic number
 Na — Element symbol
 Sodium — Element name
 22.99 — Average atomic mass *

* If this number is in parentheses, then it refers to the atomic mass of the most stable isotope.