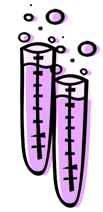
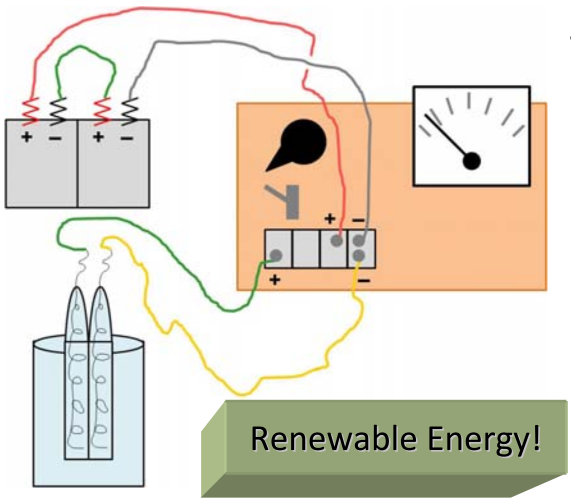




OBJECTIVES

Primary Objectives: Students will learn that a renewable fuel, Hydrogen (H_2), can be generated from water (H_2O) by electrolysis. This process can be powered by either *i*) batteries, or *ii*) solar panels. Students will be able to:

- Use an amperostat electrostation to carry out a chemical reaction
- Experimentally determine the stoichiometry of water splitting: $\mathbf{2\ H_2O \rightarrow \underline{2}\ H_2 + \underline{1}\ O_2}$
- Determine the effects of *reaction time* and *current reading* on the amount of H_2 and O_2 produced
- Use the Ideal Gas law to calculate the number of moles of H_2 and O_2 produced

Supplementary Objectives: In a second iteration of the experiment, solar panels (also provided on loan by Caltech, free of charge) can be used to drive the water-splitting reaction. In this experiment, students can learn about:

- The relationship between sunlight, energy and photons
- How much battery power (or many small solar panels) are needed to drive the reaction?
- Can room lighting provide enough energy to run the reaction? Or can only the sun provide enough photon flux for water-splitting?

Equipment and Supplies: ***all provided on loan from Caltech, free of charge!***

Relevance to California Chemistry Standards, Grades 9 - 12:

Chemical Bonds

2.a. *Students know...* atoms combine to form molecules by sharing electrons to form covalent bonds
2.b. ....chemical bonds between atoms in molecules such as H_2, H_2O, CH_4, N_2 are covalent
2.e. ....how to draw Lewis dot structures

Conservation of Matter and Stoichiometry

3.a. *Students know...* how to describe chemical reactions by writing balanced equations.
3.g. ....how to identify reactions that involve oxidation and reduction and how to balance redox reactions.

Gases and Their Properties

4.a. *Students know...* how to apply the ideal gas law to relations of temperature, pressure and volume.
4.d. ....the values and meanings of standard temperature and pressure
4.h. ....how to solve problems by using the ideal gas law in the form *PV=nRT*

Investigation & Experimentation:

Background:

The generation of hydrogen fuel (H_2) represents a new path in renewable energy production for the 21st century. Hydrogen fuel is extremely desirable because the combustion of H_2 with O_2 produces only water (H_2O). This is in stark contrast to the combustion of hydrocarbon fuels like natural gas (CH_4) or gasoline (long-chain hydrocarbons), which generate the carbon dioxide (CO_2) that is thought to be responsible for global warming and climate change.

Presently, 99% of H_2 is in fact produced from coal, which is not a renewable resource. An alternate and *renewable* means of H_2 production is by the chemical electrolysis of water. If the required electricity is provided by solar panels, the process is completely renewable. Overall, the reaction proceeds by a "water-splitting" mechanism, whose stoichiometry is outlined below:

$$\mathbf{2\,H_2O \rightarrow 2\,H_2 + 1\,O_2} \qquad \textbf{(Eq 1)}$$

This reaction can be effected by the application of a potential (voltage) across a conductive aqueous solution. Since water is non-conductive, an electrolyte, such as sodium chloride (NaCl) or sodium sulfate (Na_2SO_4), must be added to accelerate the reaction. Typically, the electrodes submerged in the solution are made of various types of metal or graphite. Expensive metals like platinum or iridium make the best catalyst/electrodes, but for this experiment we use more affordable nickel electrodes. Nickel electrodes are also in widespread use in industrial electrolysis due to their low cost and scalability.

The generation *and quantification* of gaseous products (H_2 and O_2) from a liquid reactant will provide a platform to instruct your class about a number of chosen topics, including:

*Renewable/Non-renewable energy
*Phases (gases and liquids)
*Decomposition reactions (see **Eq 1**)
*Stoichiometry
*Electrochemistry (e.g., what is an electrode?)
*Ideal gas law
*Experimental techniques: reproducibility, taking averages, making data tables

In this experiment, the source of electricity for the water-splitting reaction will be 2 × 6 V batteries wired in series (12 V total). The provided electrostations will allow your students to carefully control the current running through the cell, i.e. the rate of the generation of H_2 and O_2. The gaseous products will be collected *quantitatively* in the provided electrode/housing apparatus. Thus, based on the stoichiometry, students can deduce which electrode is producing H_2 (2 equivalents) or O_2 (1 equivalent). Several example data tables are provided.

Using the supplied amperostat and with your guidance, students in groups of 2-4 can explore the effects of time (5, 10, 15 or 20 min) or current (50, 100, 150, 250 mA) on the observed electrolysis.

The ideal time period for preparation, explanation, performing the experiment, and pursuing follow-up questions is either one-and-a-half 90 minute periods, or 2 × 50 minute periods.

<u>Note:</u> Hydrogen is safe in the quantities generated here (<15 mL) and will dissipate without effect, but ***H_2 should not be generated in any greater quantity*** *than that suggested herein! The Na_2SO_4 used as electrolyte is non-corrosive, safe (if splashed on skin, simply rinse off) and can be disposed in a conventional drain.*

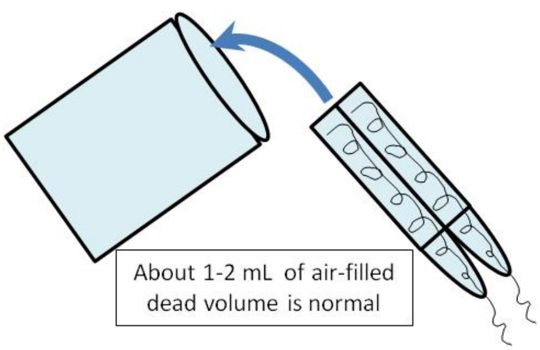
STUDENT LABORATORY PROCEDURE *[student handout]*

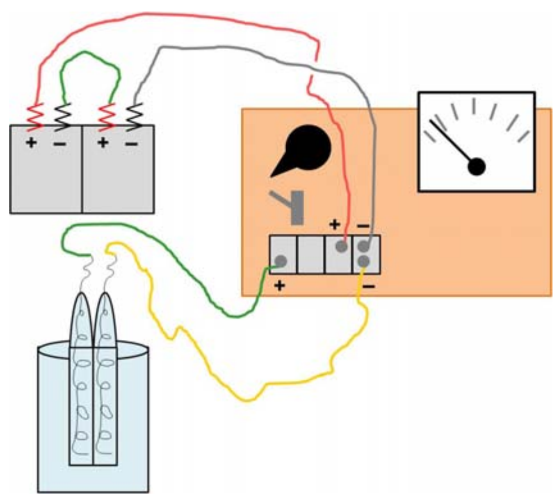
Materials and Supplies:

- Amperostat electrostation
- 6 Volt batteries (2)
- Tap water and sodium sulfate (Na_2SO_4)
- Electrode assembly, and tall 500 mL Beaker
- Clip leads (5)

<u>Procedure</u>: (to simplify and expedite the experiment, items in gray can be performed by the instructor)

1. Fill the tall beaker about 3/4 full with deionized or tap water, and dissolve about 20 g of Na_2SO_4.
2. Make sure the ON/OFF switch is set to **OFF**, and the dial is set to **MIN**. Using the included clip leads:
   A) connect battery #1 **(+)** terminal to the **Battery (+)** terminal on the electrostation
   B) connect Battery #2 **(-)** terminal to the **Battery (-)** on the electrostation
   C) between the two batteries, connect Battery #1 **(-)** to Battery #2 **(+)** to complete the circuit.
3. Fill the electrode assembly with the solution, and then submerge the electrode assembly by inverting them quickly at 45° angle into the beaker which is held also at 45°.





4. Again using clip leads, connect the two electrodes to the "**(+) Electrode (-)**" leads on the electrostation (in any order, it doesn't matter).
5. Note the water level in each electrode test tube, *and write it down* in your pre-made data table.
6. Turn the electrostation **ON**, and adjust the dial so that the meter reads 50, 100, 150 or 200 mA (depending on your instructor's direction).
7. Watch as gaseous bubbles evolve from the very tips of the electrodes. Run the reaction for 5-15 min as directed by your instructor.
8. In your data table, *write down the final water level in each electrode*.
9. Empty the solution from the test-tube electrodes back into the beaker, and re-do the experiment under different conditions (different time, different mA setting).

Effect of Current Setting on H_2 Production

5 Minute Rxn	Current Setting (mA)			
	50	**100**	**150**	**200**
Tube #1: *initial* Reading (mL_i)				
Tube #1: *final* Reading (mL_f)				
Gas produced (mL_f – mL_i)				
Tube #2: Initial Reading (mL_i)				
Tube #2: Initial Reading (mL_f)				
Gas produced (mL_f – mL_i)				

Effect of Reaction Time on H_2 Production

Set to 200 mA	**Time of Reaction (min)**			
	5	**10**	**15**	**20**
Tube #1: *initial* Reading (mL_i)				
Tube #1: *final* Reading (mL_f)				
Gas produced ($mL_f - mL_i$)				
Tube #2: Initial Reading (mL_i)				
Tube #2: Initial Reading (mL_f)				
Gas produced ($mL_f - mL_i$)				

Checking for Understanding: Analysis Questions to Ask Your Students.

1. The reaction you've just observed is given by the following equation: $\mathbf{2\ H_2O \rightarrow \underline{2}\ H_2 + \underline{1}\ O_2}$

   Based on the underlined coefficients above, which one of your tubes was generating $\mathbf{H_2}$? Which tube was generating $\mathbf{O_2}$? Why?

2. In your own words, describe the meaning of the word "*Stoichiometry*":

3. What effect did increasing the current level (mA reading on the meter) have on H_2 production? Why?

4. Draw the Lewis dot structures of H_2O, H_2 and O_2. <u>Bonus:</u> What is the shape of H_2O? Why?

Advanced Questions

5. During your reaction that produced the most H_2 (largest "Gas Produced" value in your table), how many moles of H_2 were produced? (<u>Hint:</u> remember the ideal gas law, *PV=nRT*)

6. The combustion of hydrocarbons or natural gas with O_2 proceeds with only 20% efficiency in mechanical (i.e. internal combustion) engines. In contrast, the combustion of H_2 with O_2 in a "Fuel Cell" produces electricity with 80% efficiency. At 80% efficiency, how many electrons does 15 mL of H_2 produce?

   (<u>Hint:</u> start by writing a balanced reaction for decomposing $\mathbf{H_2}$ into $\mathbf{H^+}$'s and $\mathbf{e^-}$'s)

7. Sunlight hits the earth with an intensity of about 1000 W/m^2. The experiment you just performed uses about 5 W and generated about 15 mL of H_2 in 15 minutes. (<u>Hint:</u> this is a problem of unit conversions)

   a) If instead of using batteries, you used a 1 m^2 solar panel, how much H_2 (in L) could you generate in 15 minutes?

   b) Unfortunately, most solar cells are only 20% efficient. In this case, how many mL of H_2 would be produced in 15 minutes?

Use an extra sheet of paper to the work for Question 7

Instructors Guide

[Review of key concepts]

1) <u>Stoichiometry:</u> The key aspect of the experiment is the **visual demonstration of stoichiometry**, in the volumes of H_2 and O_2 gases.

2) <u>Ideal Gas Law:</u> Leading naturally from the above observations, one can utilize the ideal gas law to prove that the volumes observed are linearly related to the moles of products. That is, the same number of moles of two different gaseous molecules occupy the same amount of volume.

3) <u>Unit Conversions:</u> As with most lab exercises, this experiment provides a chance for your students to practice unit conversions. Converting the volumes of H_2 and O_2 into #moles using the ideal gas law is the main example here.

[Extensions]

If more interest or time is available in class or in an after-school Science Club format: for additional experiments, instructors can make the experiment more creative by considering the following:

- have students bring in other metals of different type (household copper wire, zinc-plated nails, iron screws, brass screws, an old gold or silver ring) in different shapes to test for H_2 and O_2 evolving activity

- have students bring in different types of batteries that can be hooked up to the electrostation (either in series [high voltage] or in parallel [high current]) and used with the nickel electrodes to determine the effects of power on H_2 production.[<u>**Note:**</u> the electrostation can handle up to 24 V, and the resistivity of the aqueous solution will never allow more than 500 mA.]

- Caltech can provide you (on loan, no charge) with an array of small hobby solar panels to carry out the experiment outside on a bright sunny day (yes, it really does work pretty well if it's sunny!) [I think you need to have a rudimentary lesson plan about how to use and measure energy used with the solar panels. It helps make the connections more concrete to the teachers, and hopefully they will want to borrow the solar panels too.]

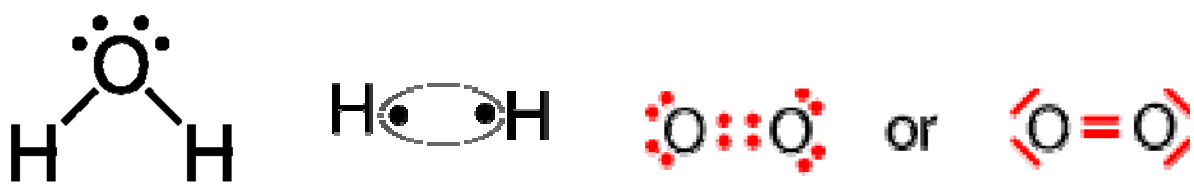
[Answers to Questions]

1. The tube with larger volume change is H_2, smaller volume change is O_2. The 2:1 stoichiometry of H_2:O_2 provides evidence of this.

2. <u>Stoichiometry:</u> the ratio of products to reactants, and among products and reactants so that the number of atoms on the left side is equal to the number of atoms on the right side.

3. Increasing the time of reaction and the current level increased H_2 production. More specifically, increasing the current level increased the *Rate of H_2 Production*, while increasing the time just increased the total amount produced.

4. <u>Bonus:</u> The shape of the water molecule is "bent"



5. As a standard: **15 mL of H_2 = 6.1×10^{-4} mol** (adjust accordingly according to students values).

6. The key here is to remember to multiply the mol of H_2 by **2**, because: $\underline{\mathbf{1}}\ \mathbf{H_2} \rightarrow \mathbf{2\ H^+ + \underline{2}\ e^-}$
   the answer is: $\mathbf{(6.1 \times 10^{-4}\ mol) \times 2 \times 0.80 = 9.8\ 10^{-4}\ mol\ of\ e^-}$
   $\mathbf{9.8\ 10^{-4}\ mol\ of\ e^- \times (6.02{\times}10^{23}) = 5.9 \times 10^{20}\ e^-}$

7. a) **1000W / 5W = 200**-fold more Watts
   therefore **15 mL × 200 = 3,000 mL (or 3 L)** of H_2
   b) **3 L × 0.2 = 0.6 L** at 20% efficiency