# Redox Reactions and Electrochemistry

An **oxidation-reduction reaction** (known as a *redox reaction*) is any reaction in which electrons are transferred from one reactant to another. You're already familiar with one redox reaction—the reaction that takes place in a hydrogen fuel cell. But you've probably talked about or even initiated many other redox reactions without even knowing it:

- If you've ever made a comment about rusty metal, you were talking about a type of redox reaction called corrosion.
- When you turn on a battery-powered device, you are setting in motion a redox reaction—and if you recharge those batteries, you are making the same reaction go in reverse.
- Have you ever lit a fire? That starts a type of redox reaction known as combustion.

You even have redox reactions going on inside your own body; for example, when your cells "burn" glucose (a sugar molecule), a redox reaction takes place that produces water, carbon dioxide, and energy.

So why "oxidation reduction"? Oxidation is the loss of electrons, and reduction is the gain of electrons. So in a redox reaction, one reactant loses electrons and one reactant gains electrons.

Redox reactions can be put to a lot of practical uses. Often this is done using a device called an electrochemical cell. **Electrochemical cells** have the following parts:

- Two electrodes, an anode and a cathode. An **electrode** is a solid substance on whose surface redox reactions take place. Oxidation (loss of electrons) takes place at the anode, and reduction (gain of electrons) takes place at the cathode.
- An **electrolyte**—a substance that contains ions and can conduct electricity.
- An electrical connection (such as a wire) from the anode to the cathode.

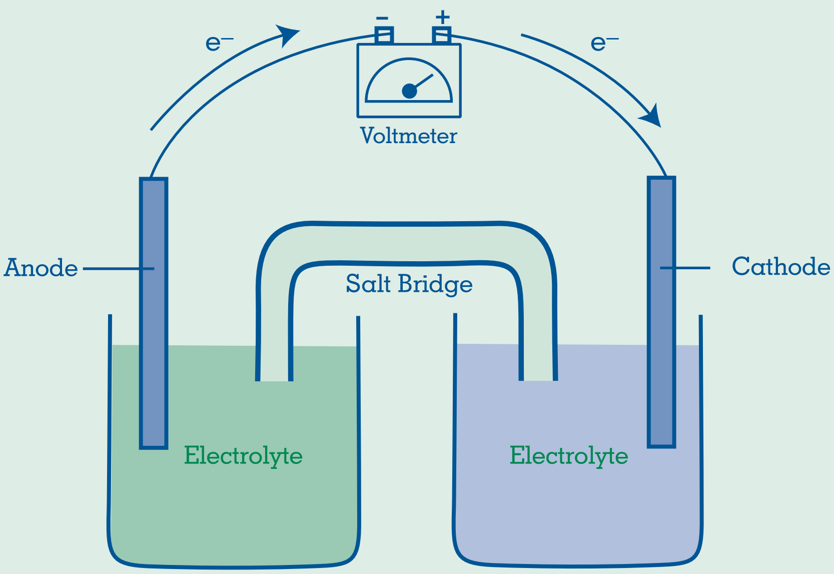
### WHAT'S IN A NAME?

The word *oxidation* might give the impression that oxygen is always involved in a redox reaction. That's not quite true, but it's close. Oxygen is the most plentiful element on Earth and can take electrons away from many other elements. As a result, many of the first redox reactions understood by early chemists involved oxygen, and the process of taking electrons away from an atom or molecule became known as oxidation.

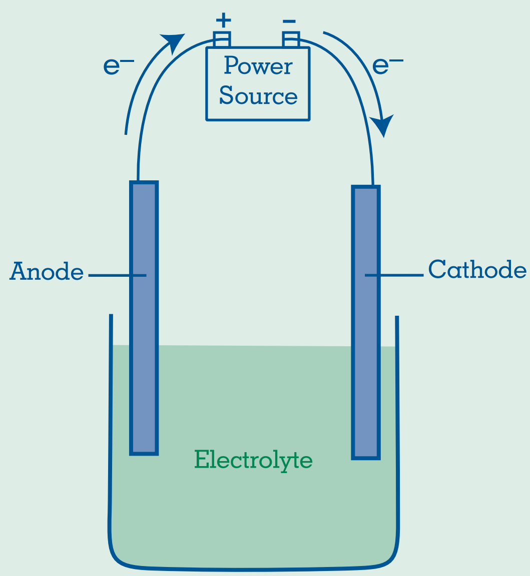
Do some of these words sound familiar? That's because fuel cells, which you learned about earlier, are a type of electrochemical cell. Batteries, both rechargeable and non-rechargeable, are also electrochemical cells. There are many other types of electrochemical cells, but they can all be divided into two main categories:

- In electrolytic cells, electricity is used to cause a non-spontaneous redox reaction to occur. A power source, such as a battery, pulls electrons from molecules at the anode, turning the molecules into positively charged ions. The power source pushes these electrons to the cathode, where they combine with positively charged ions in the electrolyte.
- In galvanic cells (also called voltaic cells, or sometimes just electrochemical cells), a spontaneous redox reaction occurs. The electrons at the anode are attracted to molecules at the cathode. Because free electrons cannot travel through the electrolyte, they get to the cathode by traveling through the external wire, producing an electrical current.

**Electrolytic Cell**




**Galvanic Cell**




The reactions that take place in electrochemical cells are often reversible. This means that a reaction that proceeds spontaneously, producing electric current, can be made to run in reverse by applying an electric current.

As an example, consider the reaction that occurs in hydrogen fuel cells to generate electricity. Just as it's possible to form water and energy from oxygen and hydrogen through a chemical reaction, it's possible to separate the hydrogen and oxygen molecules from water through a decomposition reaction. (As you've learned, a decomposition reaction is one in which a more complex molecule is broken into simpler parts.)

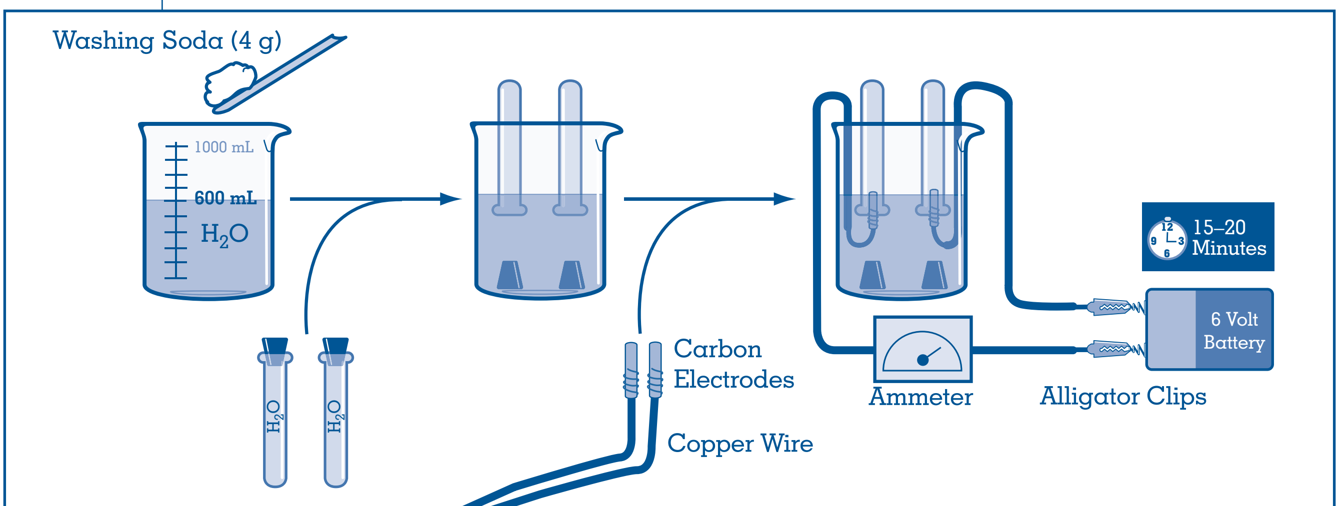
## On Your Own

Write a balanced chemical equation for the decomposition of water. The reactant is water ($H_2O$), and the products are oxygen gas ($O_2$) and hydrogen gas ($H_2$).

# WATER, WATER EVERYWHERE, BUT NOT A DROP OF HYDROGEN

Seventy percent of Earth's surface is covered by water, and every molecule of water contains two hydrogen atoms. That's *a lot* of hydrogen—so what's the big deal about obtaining it?

The problem is that hydrogen and oxygen stick together pretty tightly, so energy needs to be added to break them apart. However, the decomposition of water can be done fairly simply through **electrolysis**, the process of using electricity to split water molecules into hydrogen gas and oxygen gas (the electricity provides the energy needed to break the bonds). See this for yourself by conducting a Water Electrolysis Lab. Examine the following diagram and answer the **Questions for Reflection**. Then create a procedure for making hydrogen from water, using the diagram and your answers to the **Questions for Reflection**. You may use the materials listed—and any other materials that you think are necessary—to carry out your procedure.




## Questions for Reflection

1. What safety precautions will you take when you carry out this lab?
2. How will you keep the test tubes from falling over?
3. How will you know which test tube contains which electrode?
4. How will you measure the amount of gas formed in each test tube?
5. Based on the fact that the chemical formula for water is $H_2O$, how much hydrogen do you think you should produce relative to the amount of oxygen?
6. If you wanted to confirm which tube contained which gas, what test could you perform? (Hint: Remember the first balloon demonstration.)

7. The ammeter measures the electrical current flowing through the electrolysis circuit. How could you use this information to calculate the efficiency of electrolysis?

**Caution: In this lab, you will run an electric current through water. Be sure to take appropriate safety precautions, and do not touch any part of the electrolysis apparatus while the battery is connected to the system!**

## Materials

- Safety goggles
- Water
- 1000 mL beaker
- 4 g washing soda (sodium carbonate)
- Two test tubes (20 mL) with stoppers
- Grease pencil
- Two carbon electrodes approximately 4 cm long
- Two 30-cm long pieces of coated copper wire, with the coating removed from 4 cm of wire on both ends
- Two insulated alligator clips
- 6 V lantern battery
- Multimeter or ammeter

# MEASURING VOLTAGE

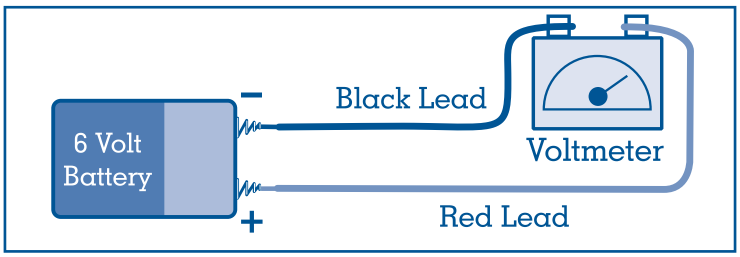
To measure the actual voltage supplied by the battery in the Water Electrolysis Lab, follow these steps:

1. Practice using the voltmeter (or a multimeter set to measure voltage) by measuring the voltage supplied by the 6 V battery alone:

   a. Set the voltmeter to measure direct current (DC voltage).

   b. Attach the voltmeter leads to the battery terminals as shown in **Figure 1**: red from the V (voltage) or "+" connection on the voltmeter to the positive terminal of the battery; black from the COM (common) or "–" connection on the voltmeter to the negative terminal of the battery.

### Figure 1: Test the Battery Voltage

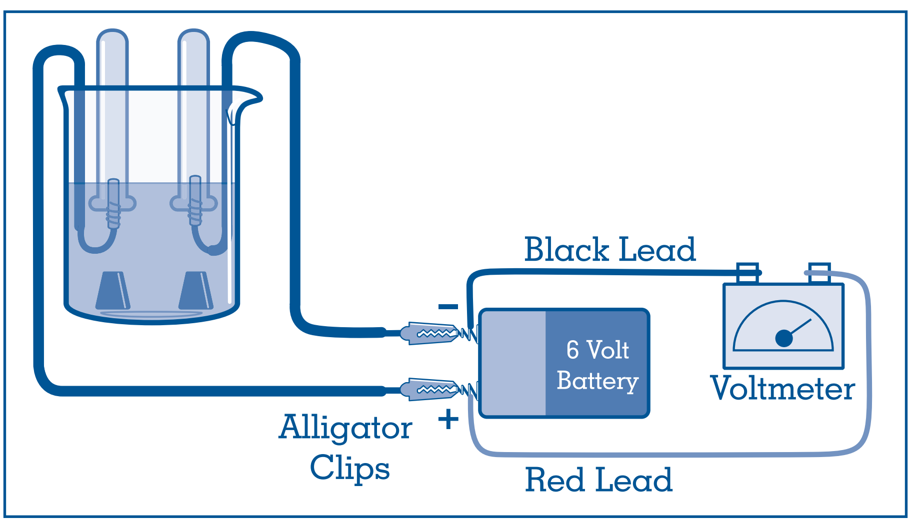



   Note: It is possible for the meter to show a value higher than the theoretical voltage of the battery (6 V).

2. Measure the voltage of the circuit while the electrolysis reaction is occurring:

   a. Connect the voltmeter in the electrolysis circuit across the battery, as shown in **Figure 2**.

### Figure 2: Setup to Measure the Battery Voltage




   b. Begin the electrolysis. Measure and record the voltage.

   c. Disconnect the voltmeter from the circuit.

# TURNING WATER INTO FUEL

With your Hydrogen Technology Research team, carry out the procedure you developed for the Water Electrolysis Lab. Be sure to record the length of time that you run the reaction, the current being applied, and the volume of hydrogen produced.

While the reaction is running, think about electrolysis as an *energy transformation*: The electrical energy produced by the battery is transformed into the chemical bond energy of the hydrogen and oxygen produced. In an energy transformation, the **efficiency** of the process is the ratio between the *energy out* (in this case, the chemical bond energy in the hydrogen gas) and the *energy in* (in this case, the electricity). As a team, discuss any ideas you have about how you might calculate the efficiency of the electrolysis reaction.

# EFFICIENCY OF ELECTROLYSIS

You've probably heard quite a bit about *energy efficiency*. Many advertisements promote products based on their energy efficiency. The energy company that supplies your home with electricity probably offers home energy audits to help homeowners figure out how to make their homes more energy-efficient, perhaps by adding insulation or replacing the windows with more energy-efficient ones. When people talk about something being energy-efficient, they usually mean that it uses less energy to perform the same task.

In science, energy efficiency means something a little different. Any time that energy is transformed from one form to another, some of the energy is "lost," meaning that it's not available to do useful work. For example, when a car burns gasoline, only 15–25 percent of the energy in the gasoline is converted into mechanical energy for moving the car; the rest is lost, much of it as heat. This form of energy efficiency is sometimes referred to as *energy conversion efficiency*.

# CALCULATING THE ENERGY EFFICIENCY OF ELECTROLYSIS

**Name:** ______________________________ **Date:** ______________

The energy efficiency of the electrolysis reaction can be expressed as:

$$\text{\% Efficiency of Electrolysis} = \frac{\text{Energy Content of } H_2 \text{ Produced}}{\text{Electrical Energy Supplied}} \times 100 \qquad \text{(Equation 1)}$$

To calculate the efficiency, first calculate the electrical energy supplied and the energy content of the hydrogen produced, as outlined in the following steps.

## Step 1: Calculate the electrical energy supplied.

- The electrical energy supplied is equal to the electrical power ($P$) multiplied by the length of time ($t$) the power is supplied:

$$\text{Electrical Energy Supplied} = P \times t \qquad \text{(Equation 2)}$$

For an electrical circuit, power ($P$) is the voltage ($V$) multiplied by the current ($I$):

$$P = V \times I \qquad \text{(Equation 3)}$$

- Plug Equation 3 into Equation 2, and write the resulting equation:

$$\text{Electrical Energy Supplied} = \underline{\qquad\qquad\qquad\qquad} \qquad \text{(Equation 4)}$$

- Enter the data from your electrolysis reaction, with V in volts, I in amps, and t in seconds, and calculate the energy supplied:

$$\text{Electrical Energy Supplied} = \underline{\quad}\text{ V} \times \underline{\quad}\text{ A} \times \underline{\quad}\text{ s} = \underline{\quad}\text{ joules} \qquad \text{(Equation 5)}$$

- A joule (j) is a unit of energy. Your electric bill probably uses a different unit of energy, the kilowatt-hour (kWh). Convert joules to kilowatt-hours as follows:

$$\text{Electrical Energy Supplied (kWh)} = \frac{\text{Electrical Energy Supplied (j)}}{3{,}600} \qquad \text{(Equation 6)}$$

- Enter your result from Equation 5:

$$\text{Electrical Energy Supplied (kWh)} = \frac{\underline{\qquad\qquad}}{3{,}600} = \underline{\quad}\text{ kWh} \qquad \text{(Equation 7)}$$

## Step 2: Calculate the energy content of the hydrogen produced.

- For every kilogram of hydrogen that reacts with oxygen to produce water, 39 kWh of energy is released. So:

$$\text{Energy Content of } H_2 \text{ Produced} = m \times 39 \frac{\text{kWh}}{\text{kg}} \qquad \text{(Equation 8)}$$

  ($m$ is the mass of hydrogen in kg)

- You can calculate the mass of the hydrogen as follows:

$$m = \rho \times V \qquad \text{(Equation 9)}$$

  ( $\rho$ is the density of hydrogen gas and $V$ is the volume of hydrogen gas produced in milliliters)

- Plug Equation 9 into Equation 8, and write the resulting equation:

$$\text{Energy Content of } H_2 \text{ Produced} = \qquad \text{(Equation 10)}$$

- Using $\rho = 0.08988$ kg/mL, enter your result for the volume of hydrogen produced, and calculate the energy content:

$$\text{Energy Content of } H_2 \text{ Produced} = 0.08988 \frac{\text{kg}}{\text{mL}} \times \_\_\_\_ \text{ mL} \times 39 = \_\_\_\_ \text{ kWh} \times \frac{\text{kWh}}{\text{kg}} \qquad \text{(Equation 11)}$$

## Step 3: Calculate the efficiency.

- Plug in the results from Step 1 and Step 2:

$$\text{Energy Efficiency} = \frac{\text{Energy Content of } H_2 \text{ Produced}}{\text{Energy Supplied}} \times 100$$

$$\text{Energy Efficiency} = \_\_\_\_\_ \times 100 = \_\_\_\_\_\% \qquad \text{(Equation 12)}$$

## Step 4: Check your answer.

Does your answer make sense? The efficiency can't be less than 0 and it can't be greater than 100%. If your answer isn't between 0 and 100, double-check your calculations. Make sure that you entered the following:

- The time in seconds (which should be a big number)
- The voltage in volts
- The current in amps (not milliamps)
- The volume of hydrogen in milliliters

## HOMEWORK 3.5

Complete your lab report on the Water Electrolysis Lab.

## HOMEWORK 3.6

Look over the research that you have collected for your Hydrogen Production Background Memo. Identify any additional information you'll need to complete the memo.