

Science Projects Energy and Science Projects For Students



Splitting Water

H₂O Electrolysis (Hydrolysis)

Electricity is "created" when certain chemicals react together. We use chemically- made electricity to power many machines from flashlights to a watch or sometimes a car. Yes, there are cars that run on electricity! The devices that store electricity are called batteries. Electricity can also be used to produce chemical changes.

Water is a simple chemical made from two gases -- hydrogen and oxygen. Every molecule of water has two atoms of hydrogen for every atom of oxygen. **H₂O** is the chemical formula for a molecule of water.

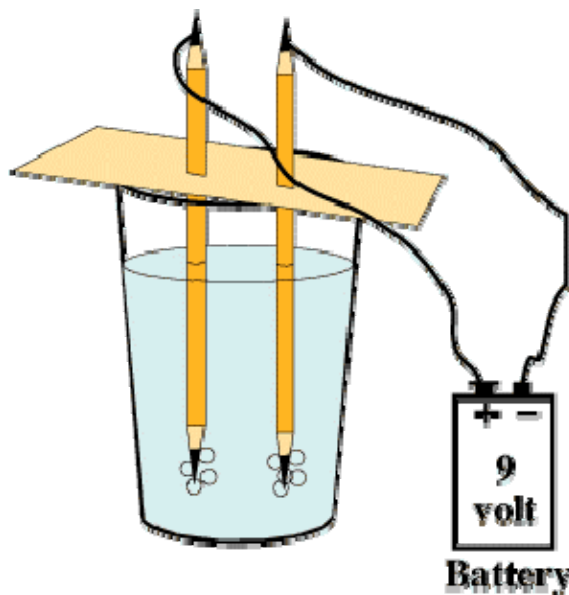
If an electrical current is passed through water between electrodes (the positive and minus poles of a battery), the water is split into its two parts: oxygen and hydrogen. This process is called electrolysis and is used in industry in many ways, such as making metals like aluminum. If one of the electrodes is a metal, it will become covered or **plated** with any metal in the solution. This is how objects are **silver plated**.

Try This!

You can use electricity to split water into its two gases -- oxygen and hydrogen.

What You'll Need

- A 9 volt battery
- Two regular number 2 pencils (remove eraser and metal part on the ends)
- Salt [Baking Soda prevents Chlorine from Forming]
- Thin cardboard
- Electrical wire
- Small glass
- Water



What to Do

1. Sharpen each pencil at both ends.
2. Cut the cardboard to fit over glass.
3. Push the two pencils into the cardboard, about an inch apart.
4. Dissolve about a teaspoon of salt into the warm water and let sit for a while [double if Baking Soda].
5. Using one piece of the electrical wire, connect one end on the positive side of the battery and the other to the black graphite (the "lead" of the pencil) at the top of the sharpened pencil. Do the same for the negative side connecting it to the second pencil top.
6. Place the other two ends of the pencil into the salted [or Baking Soda] water.

Results

As the electricity from the battery passes through and between the electrodes (the pencils), the water splits into hydrogen and oxygen, which collect as **very** tiny bubbles around each pencil tip.

Think About...

Our God thought about all of these things long before He created the World and everyone in the world. His miracles are everywhere and most of them we cannot even see unless we look very, very closely.

When we do our experiment, we may not immediately notice that there are twice as many bubbles coming from the graphite tip that connects to the negative or minus (-) terminal of the batter than from the positive or plus (+) terminal if you use the Baking Soda I suggested. This is because water is made up of hydrogen and oxygen in exactly the ratio of 2 to 1 which is why water is called H₂O. Hydrogen, which has a net positive charge when ionized in water, is attracted to the opposite polarity terminal which is negative. The oxygen is attracted to the opposite terminal but since there is half as much of it, there are fewer bubbles.

Pure water does not conduct much electricity so to allow current to flow in order to take apart these water molecules, we add some salt to the water. The reason we use a graphite pencil instead of putting a piece of wire into the salt water is that the salt will attack the piece of wire and make the solution messy. Otherwise, it would work just as well. Graphite does not get attacked easily by salt water plus the wooden pencil makes a nice holder for our experiment.

Special caution is needed to keep the top of the glass with the bubbles rising open so the bubbles can escape. The hydrogen and oxygen, if mixed together, could ignite (burn) and return back to water but giving out an unpleasant burst of heat.

This original experiment used salt which is sodium chloride (NaCl). While you definitely get bubbles from both electrodes (pencil tips), some chlorine gas is formed at the positive terminal. In fact, this is how chlorine gas is produced using sea water. The Baking Soda has no chlorine in it so no chlorine is produced and you should avoid getting that “swimming pool” smell of chlorine. Lye (NaOH) produces the best 2-hydrogen, 1-oxygen ratio but that chemical is too dangerous to work with in a home project.

Brought to you by Mr. D, Children’s Church, Trinity Congregational Church, Bolton, MA 01740

The original document (sans “Think About” is available at:

http://www.reachoutmichigan.org/funexperiments/agesubject/lessons/energy/split_h2o.html

This Energy Education Project comes from the California Energy Commission



Electrolysis of Water – A Much Deeper Study

From Wikipedia, the free encyclopedia

Electrolysis of water is the decomposition of [water](#) (H₂O) into [oxygen](#) (O₂) and [hydrogen](#) gas (H₂) due to an [electric current](#) being passed through the water. This [electrolytic process](#) is used in some industrial applications when hydrogen is needed.

An electrical power source is connected to two [electrodes](#), or two plates, (typically made from some inert metal such as [platinum](#) or [stainless steel](#)) which are placed in the water. Hydrogen will appear at the [cathode](#) (the negatively charged electrode, where [electrons](#) are pumped into the water), and oxygen will appear at the [anode](#) (the positively charged electrode). The generated amount of hydrogen is twice the amount of oxygen, and both are [proportional](#) to the total [electrical charge](#) that was sent through the water.

Electrolysis of *pure* water is very slow, and can only occur due to the [self-ionization of water](#). Pure water has an [electrical conductivity](#) about one millionth that of seawater. It is sped up dramatically by adding an [electrolyte](#) (such as a [salt](#), an [acid](#) or a [base](#)).

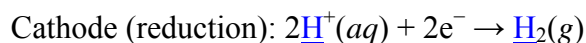
Historically, the first known electrolysis of water was done by [William Nicholson](#) and [Anthony Carlisle](#) in about 1800.

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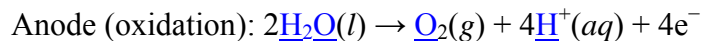
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Equations

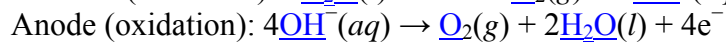
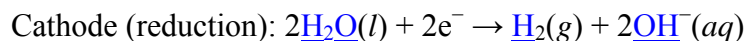
In the water at the negatively charged cathode, a [reduction](#) reaction takes place, with electrons (e[−]) from the cathode being given to hydrogen cations to form hydrogen gas (the half reaction balanced with acid):



At the positively charged anode, an [oxidation](#) reaction occurs, generating oxygen gas and giving electrons to the cathode to complete the circuit:



The same half reactions can also be balanced with base as listed below. Not all half reactions must be balanced with acid or base. Many do like the oxidation or reduction of water listed here. To add half reactions they must both be balanced with either acid or base.



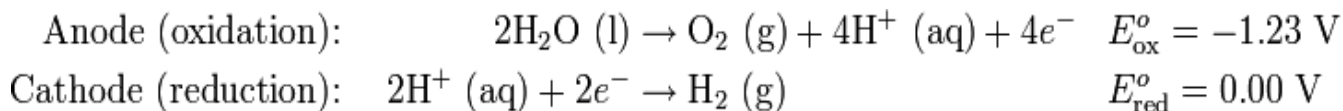
Combining either half reaction pair yields the same overall decomposition of water into oxygen and hydrogen:



The number of hydrogen molecules produced is thus twice the number of oxygen molecules. Assuming equal temperature and pressure for both gases, the produced hydrogen gas has therefore twice the volume of the produced oxygen gas. The number of electrons pushed through the water is twice the number of generated hydrogen molecules and four times the number of generated oxygen molecules.

Thermodynamics of the process


Decomposition of pure [water](#) into [hydrogen](#) and [oxygen](#) at [standard temperature and pressure](#) is not favorable in [thermodynamical](#) terms. This is because, $E(\text{cell}) = E(\text{Oxidation}) + E(\text{Reduction})$. If $E(\text{cell}) < 0$, reaction is not favorable.



Thus, the standard potential of the water electrolysis cell is -1.23 V at 25 °C from the [Nernst Equation](#).

The negative voltage indicates the [Gibbs free energy](#) for electrolysis of water is greater than zero for these reactions. This can be found using the $G = -nFE$ equation from chemical kinetics, where n is the moles of electrons and F is the [Faraday constant](#). The reaction cannot occur without adding necessary energy, usually supplied by an external electrical power source.



 Hoffman voltameter connected to a [direct current power supply](#).

Electrolyte selection

If the above described processes occur in pure water, H^+ [cations](#) will accumulate at the anode and OH^- [anions](#) will accumulate at the cathode. This can be verified by adding a [pH indicator](#) to the water: the water near the anode is acidic while the water near the cathode is basic. These charged ions will repel the further flow of electricity until they have [diffused](#) away, a slow process. This is why pure water conducts electricity poorly and why electrolysis of pure water proceeds slowly.

If a water-soluble [electrolyte](#) is added, the conductivity of the water rises considerably. The electrolyte disassociates into [cations](#) and [anions](#); the anions rush towards the anode and neutralize the buildup of positively charged H^+ there; similarly, the cations rush towards the cathode and neutralize the buildup of negatively charged OH^- there. This allows the continued flow of electricity.^[1]

Care must be taken in choosing an electrolyte, since an [anion](#) from the electrolyte is in competition with the hydroxide ions to give up an [electron](#). An electrolyte [anion](#) with less [standard electrode potential](#) than hydroxide will be oxidized instead of the hydroxide, and no oxygen gas will be produced. A [cation](#) with a greater [standard electrode potential](#) than a hydrogen ion will be reduced in its stead, and no hydrogen gas will be produced.

The following [cations](#) have lower electrode potential than H^+ and are therefore suitable for use as electrolyte cations: [Li⁺](#), [Rb⁺](#), [K⁺](#), [Cs⁺](#), [Ba²⁺](#), [Sr²⁺](#), [Ca²⁺](#), [Na⁺](#), and [Mg²⁺](#). [Sodium](#) and [lithium](#) are frequently used, as they form inexpensive, soluble salts.

If an [acid](#) is used as the [electrolyte](#), the cation is H^+ , and there is no competitor for the H^+ created by disassociating water. The most commonly used [anion](#) is [sulfate](#) (SO_4^{2-}), as it is very difficult to oxidize, with the standard potential for oxidation of this ion to the [peroxodisulfate](#) ion being -0.22 volts.

Strong acids such as [sulfuric acid](#) (H_2SO_4), and strong bases such as [potassium hydroxide](#) (KOH), and [sodium hydroxide](#) (NaOH) are frequently used as electrolytes.


A solid polymer electrolyte can also be used such as [NAFION](#) and when applied with a special catalyst on each side of the membrane can efficiently split the water molecule with as little as 1.8 Volts.

Techniques

Fundamental Demonstration

Two [leads](#), running from the terminals of a battery, are placed in a cup of water with a quantity of electrolyte (not NaCl, anode creates [chlorine](#) gas) added to establish conductivity. Hydrogen and oxygen gases will stream from the oppositely charged [electrode](#). Oxygen will collect at the [anode](#) and hydrogen will collect at the [cathode](#).



 Match test used to detect the presence of hydrogen gas.

Hofmann voltameter

Main article: [Hofmann voltameter](#)

The Hofmann voltameter is often used as a small-scale electrolytic cell. It consists of three joined upright cylinders. The inner cylinder is open at the top to allow the addition of [water](#) and the [electrolyte](#). A [platinum](#) electrode is placed at the bottom of each of the two side cylinders, connected to the positive and negative terminals of a source of [electricity](#). When current is run through the Hofmann voltameter, gaseous [oxygen](#) forms at the [anode](#) and gaseous [hydrogen](#) at the [cathode](#). Each gas displaces water and collects at the top of the two outer tubes, where it can be drawn off with a stopcock.

Industrial electrolysis

Many industrial electrolysis cells are very similar to [Hofmann voltameters](#), with complex platinum plates or honeycombs as electrodes. Generally the only time hydrogen is intentionally produced from electrolysis is for specific point of use application such as is the case with [oxyhydrogen](#) torches or when extremely high [hydrogen purity](#) or oxygen is desired. The vast majority of hydrogen is produced from hydrocarbons and as a result contains trace amounts of [carbon monoxide](#) among other impurities. The carbon monoxide impurity can be detrimental to various systems including many [fuel cells](#).

High pressure electrolysis

Main article: [High pressure electrolysis](#)

High pressure electrolysis is the electrolysis of water with a [compressed hydrogen](#) output around 120-200 [Bar](#) (1740-2900 [psi](#))^[2]. By pressurising the hydrogen in the electrolyser the need for an external [hydrogen compressor](#) is eliminated, the average energy consumption for internal compression is around 3%^[3].

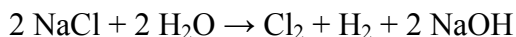
High-temperature electrolysis

Main article: [High-temperature electrolysis](#)

High-temperature electrolysis (also HTE or steam electrolysis) is a method currently being investigated for water electrolysis with a [heat engine](#). High temperature electrolysis is more efficient than traditional room-temperature electrolysis because some of the energy is supplied as heat, which is cheaper than electricity, and because the electrolysis reaction is more efficient at higher temperatures^{[4][5]}.

Applications

About four percent of [hydrogen](#) gas produced worldwide is created by electrolysis. The majority of this hydrogen produced through electrolysis is a side product in the production of [chlorine](#).



The electrolysis of [brine](#) (saltwater), a water sodium chloride mixture, is only half the electrolysis of water since the [chloride](#) ions are oxidized to [chlorine](#) rather than [water](#) being oxidized to [oxygen](#). The hydrogen produced from this process is either burned, used for the production of specialty chemicals, or various other small scale applications.

The majority of hydrogen used industrially is derived from fossil fuels. One example is fossil fuel derived hydrogen used for the creation of [ammonia](#) for fertilizer via the [Haber process](#) and for converting heavy [petroleum](#) sources to lighter fractions via [hydrocracking](#). The production of this hydrogen usually involves the formation of [synthesis gas](#) a mixture of H₂ and CO. Synthesis gas can be hydrogen enriched through the [water gas shift](#) reaction. In this reaction the [carbon monoxide](#) is reacted with water to produce more H₂ with CO₂ byproduct.

Efficiency

Water electrolysis does not convert 100% of the electrical energy into the chemical energy of hydrogen. The process requires more extreme potentials than what would be expected based on the cell's total reversible [reduction potentials](#). This excess potential accounts for various forms of [overpotential](#) by which the extra energy is eventually lost as heat. For a well designed cell the largest [overpotential](#) is the [reaction overpotential](#) for the four electron oxidation of water to oxygen at the anode. An effective [electrocatalyst](#) to facilitate this reaction has not been developed. Platinum alloys are the default state of the art for this oxidation. The reverse reaction, the reduction of oxygen to water, is responsible for the greatest loss of efficiency in [fuel cells](#). Developing a cheap effective electrocatalyst for this reaction would be a great advance (see also^[6]).

The simpler two-electron reaction to produce hydrogen at the cathode can be electrocatalyzed with almost no [reaction overpotential](#) by platinum or in theory a [hydrogenase enzyme](#). If other, less effective, materials are used for the cathode then another large overpotential must be paid.

The [energy efficiency](#) of water electrolysis varies widely with the numbers cited below on the optimistic side. Some report 50–80%^[7] ^[8]. These values refer only to the efficiency of converting electrical energy into hydrogen's chemical energy. The energy lost in generating the electricity is not included. For instance, when considering a power plant that converts the heat of nuclear reactions into hydrogen via electrolysis, the total efficiency may be closer to 30–45%.^[9]

See also

- [Electrochemistry](#)
- [Electrolysis](#)
- [Hydrogen production](#)
- [Noryl](#)
- [Gas cracker](#)
- [Oxyhydrogen](#)
- [Water purification](#)
- [Timeline of hydrogen technologies](#)

References

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3. [2003-PHOEBUS-Pag.9](#)
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5. [WELTEMP-Water electrolysis at elevated temperatures](#)
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 9. "[Briefing Paper #73: Transport and the Hydrogen Economy](#)". *Uranium Information Centre's Briefing Papers*. Australian Uranium Association. January 2008. <http://www.uic.com.au/nip73.htm>. "[Fuel cells] catalyse the oxidation of hydrogen directly to electricity at relatively low temperatures and the claimed theoretical efficiency of converting chemical to electrical energy to drive the wheels is about 60% (or more). However, in practice about half that has been achieved, except for the higher-temperature solid oxide fuel cells - 46%."
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External links

- [EERE 2008 - 100 kgH2/day Trade Study](#)
- [NREL 2006 - Electrolysis technical report](#)

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Articles related to [electrolysis](#)

Principles of electrolysis

[Electrochemical cell](#) • [Electrolytic process](#) • [Faraday's laws of electrolysis](#) • [Half cell](#) • [High-temperature electrolysis](#) • [High pressure electrolysis](#) • [Regenerative fuel cell](#) • [Solid oxide electrolyser cell](#) • [Unitized regenerative fuel cell](#) • [Standard electrode potential](#)

Electrolytic processes

[Betts electrolytic process](#) • [Castner process](#) • [Castner-Kellner process](#) • [Chloralkali process](#) • [Downs cell](#) • [Electrolysis of water](#) • [Electrowinning](#) • [Hall-Héroult process](#) • [Hofmann voltameter](#) • [Kolbe electrolysis](#)

Materials produced by electrolysis

[Aluminium](#) • [Calcium metal](#) • [Chlorine](#) • [Copper](#) • [Electrolysed water](#) • [Fluorine](#) • [Hydrogen](#) • [Lithium metal](#) • [Magnesium](#) • [Potassium metal](#) • [Sodium metal](#) • [Sodium hydroxide](#) • [Zinc](#)

See also

[Electrochemistry](#) • [Standard electrode potential \(data page\)](#)