Electrolysis: Splitting Water
Teacher 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.

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:

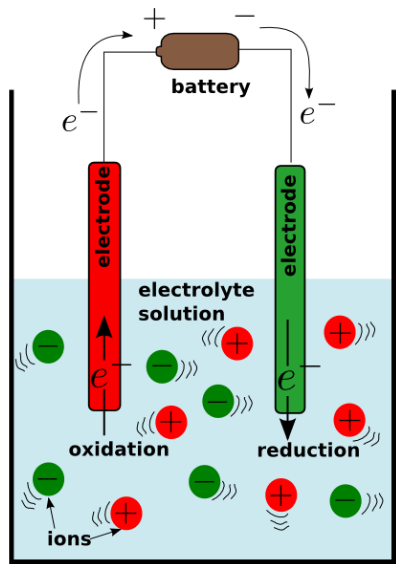
- Students are expected to know calculator arithmetic and unit conversions.
- Familiarity with chemical formulas, chemical reactions, and the periodic table is a plus.
- Some manual dexterity is required to build an apparatus.
- Suggested grade level: 9-12

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** (called the **cathode).** 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 (the **anode**) 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!

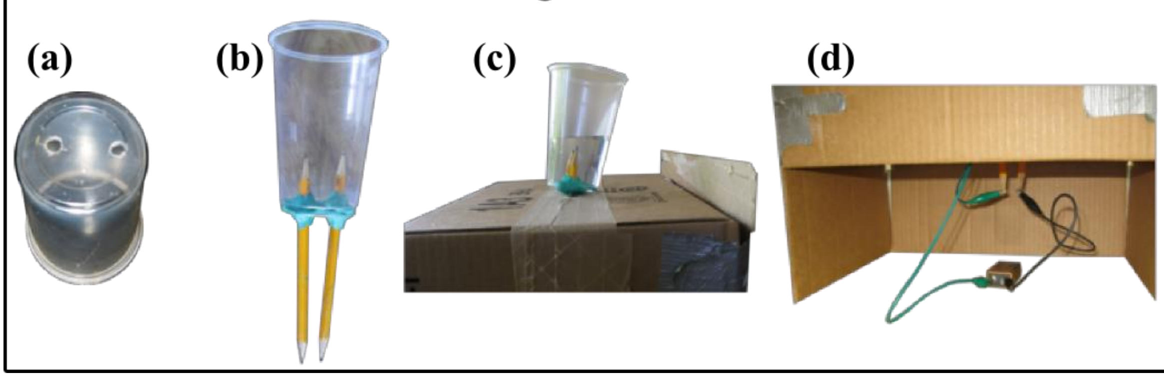
When we run current through an electrolytic cell, two chemical reactions occur. At the negative electrode, there is a **reduction reaction** – some molecule gains electrons. At the positive electrode, there is an **oxidation reaction** – some molecule loses electrons. These two chemical reactions can produce all sorts of useful products – like hydrogen gas for fuel cells, or purified metals from minerals. What specific chemical species is oxidized or reduced depends on how easily the various molecules present gain or accept electrons. This information is summarized in a table of **standard reduction potentials**. 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) Mark the side of the cup next to one pencil as (+) and the other side as (–)
   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

   Q1. What, if anything, do you observe around the pencil leads?
   Most likely nothing will be seen, or else very few bubbles.

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.
   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.

   Q2. What do you see now happening at the pencil tips?
   There should be noticeable bubbling at both 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?
   Ions in the water are necessary to make it a good conductor. The ions move around and so can carry charge across the solution from one electrode to the other. This completes the circuit and allows current to flow.

   c) Label your two test-tubes as (+) 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 (+) 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?
   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 directly to the pencil leads. Make sure to match up the (+) and (–) sides.
   h) 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.

$$\boxed{\_\_1\_\_\ Mg^{2+}(aq) + \_\_2\_\_\ e^- \rightarrow \_\_1\_\_\ Mg(s)}$$

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

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.

Water (H_2O) is more easily reduced.

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

Hydrogen gas (H2)

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 an **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.

$$\boxed{\_\_2\_\_\ SO_4^{2-}(aq) \rightarrow \_\_1\_\_\ S_2O_8^{2-}(aq) + \_\_2\_\_\ e^-}$$

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

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.

Water (H_2O) is more easily oxidized

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

Oxygen gas (O_2)

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.

___***2***____ *units of gas at the cathode* **:** ___***1***___ *units of gas at the anode*

From the balanced equations, making one molecule of H_2 at the cathode requires 2 electrons and making one molecule of O_2 at the anode gives off 4 electrons. That means that for every 4 electrons that go through the circuit, we get 2 moles of H_2 and one mole of O_2 – a 2:1 ratio

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)	volume of gas (cm^3 = mL)
H_2 (cathode, negative terminal)	*Will be around 1 cm*	
O_2 (anode, positive terminal)	*Will be around 0.5 cm*	

   c) 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 H_2 to O_2 gas: _____*around 2:1*______________

Part 2 – Industrial Application

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.
   a) 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:

$$\boxed{\_\_1\_\_ Na^+(aq) + \_\_1\_\_ e^- \rightarrow \_\_1\_\_ Na(s)}$$

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

   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:

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

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

Using the table of reduction potentials, is one reaction much more likely than the other or are they pretty close?
According to the table, water is slightly more likely to be oxidized, but the numbers are 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)?
hydrogen gas (H_2)

Q14. What gas should be produced at the anode (positive terminal)? chlorine gas (Cl_2)

Q15. What would you expect for the ratio of gases produced at each terminal?
volume of gas at cathode : volume of gas at anode = ______1:1____________
It takes the same number of electrons to produce one molecule of H2 as one molecule of Cl2, so they should be produced in equal quantities.

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): _____about 1.5______ cm

Change in height at the anode (positive terminal): _____almost 0______cm

Do you see the ratio that you expected? No

Q17. Was one gas produced in much greater quantity than the other? Make a guess as to why this might be.
Much more hydrogen than chlorine is produced. That's because most of the chlorine gas dissolves in water.

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. Volume of H_2 gas produced at cathode **(**$V = \pi R^2 h$**):** *around 2 mL*

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.

$$(\_\_\_\_\_\_\_\_\text{mL } H_2) \times \left(\frac{1\text{ L}}{1000\text{ mL}}\right) \times \left(\frac{1\text{ mol } H_2}{22.4\text{ L}}\right) = \_\_\_\_\_\_\_\_\text{mol } H_2 \text{ gas produced}$$

about 8×10^{-5} mol H_2

Expected number of moles of Cl_2 gas produced (use the expected ratio of H_2:Cl_2) :
We expect the same amount of Cl_2 to be produced as H_2.

Look up the molar mass of Cl in the periodic table
Molar mass of Cl: _____35_______ g/mol

Calculate mass of chlorine gas produced (remember to multiply by 2 to get g/mol of Cl_2):

$$(\_\_\_\_\_\_\_\_\text{mol } Cl_2) \times \left(\_\_\_\_\_\_\_\_\frac{\text{g } Cl_2}{\text{mol } Cl_2}\right) = \_\_\_\_\_\_\_\_\text{g } Cl_2 \text{ produced}$$

should be about 0.006 g

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? _____about 10 to 15______ mL
Calculate mass of water in the test-tube:

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

will be about 0.015 kg water

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

$$(\_\_\_\_\_\_\_\_\text{kg water}) \times \left(\_\_\_\_\_\_\_\_\frac{\text{g } Cl_2}{\text{kg water}}\text{ solubility}\right) = \_\_\_\_\_\_\_\_\text{g } Cl_2 \text{ to saturate solution}$$

Q21. Did you make enough Cl_2 gas to saturate the solution?
No.

Q22. Why did you see so little gas at the anode (positive terminal)?
All the chlorine gas was able to dissolve in the water

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.
O_2 and H_2 are much less soluble in water (it takes a lot fewer grams of these gases to saturate the test-tube), so the gas that was produced remained mostly undissolved.

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, it was also 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.

   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.*

   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:
   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 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) H_2O (ii) Mg^{2+}

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

$$2\,H_2O(l) + 2\,e^- \rightarrow 1\,H_2(g) + 2\,OH^-(aq)$$

Q29. What do you expect to be oxidized at the anode (positive battery terminal)?
(i) H_2O (ii) SO_4^{2-} (iii) Cu(s)

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

$$\boxed{1Cu(s) \rightarrow 1Cu^{2+}(aq) + 2e^-}$$

Q31. Do you expect to see bubbles at the positive terminal, 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

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? Yes

Q33. 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.

Q34. 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 to form a **yellow** compound (Cu_2O).

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

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?
These metals are more easily oxidized than chlorine (they lose electrons more readily), so metal ions will be formed instead of chlorine gas.

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^{o}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.

loves electrons

easily reduced

hard to oxidize

Cathode (Reduction) Half-Reaction	Standard Potential E° (volts)
$S_2O_8^{2-}$(aq) + 2e$^-$ -> 2SO_4^{2-}(aq)	2.01
Cl_2(g) + 2e$^-$ -> 2Cl^-(aq)	1.36
O_2(g) + 4H^+(aq) + 4e$^-$ -> 2H_2O(l)	1.23
ClO^-(aq) + H_2O(l) + 2e$^-$ -> Cl^-(aq) + 2OH^-(aq)	0.90
ClO_2^-(aq) + H_2O(l) + 2e$^-$ -> ClO^-(aq) + 2OH^-(aq)	0.59
Cu^+(aq) + e$^-$ -> Cu(s)	0.52
ClO_3^-(aq) + H_2O(l) + 2e$^-$ -> ClO_2^-(aq) + 2OH^-(aq)	0.35
Cu^{2+}(aq) + 2e$^-$ -> Cu(s)	0.34
AgCl(s) + e$^-$ -> Ag(s) + Cl^-(aq)	0.22
ClO_4^-(aq) + H_2O(l) + 2e$^-$ -> ClO_3^-(aq) + 2OH^-(aq)	0.17
Cu^{2+}(aq) + e$^-$ -> Cu^+(aq)	0.16
2H^+(aq) + 2e$^-$ -> H_2(g)	0.00
2H_2O(l) + 2e$^-$ -> H_2(g) + 2OH^-(aq)	-0.83
Mg^{2+}(aq) + 2e$^-$ -> Mg(s)	-2.38
Na^+(aq) + e$^-$ -> Na(s)	-2.71

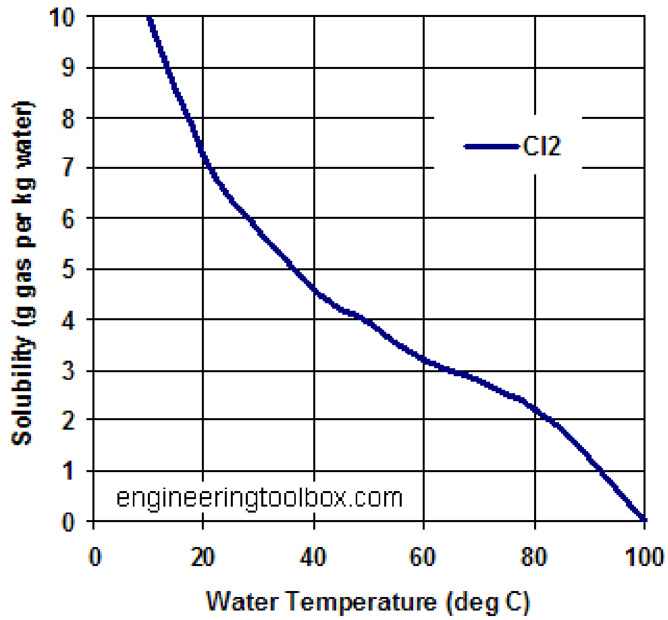
hard to reduce

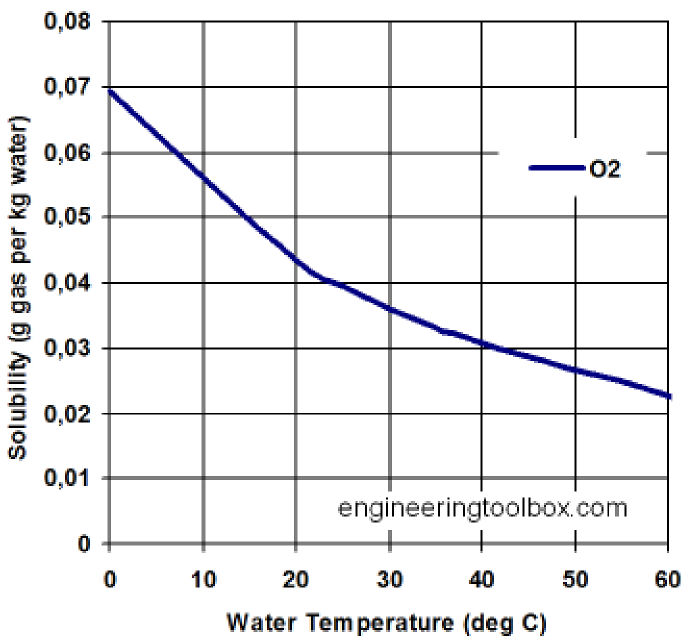
easily oxidized

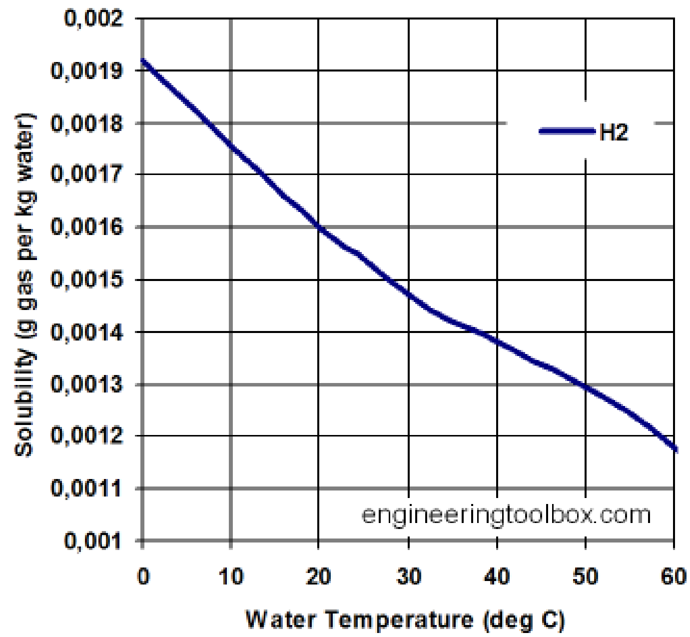
hates electrons

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 25ºC.







Periodic Table of the Elements

Chemistry Reference Sheet

California Standards Test

Key

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

Period	1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1	1 **H** Hydrogen 1.01																	2 **He** Helium 4.00
2	3 **Li** Lithium 6.94	4 **Be** Beryllium 9.01											5 **B** Boron 10.81	6 **C** Carbon 12.01	7 **N** Nitrogen 14.01	8 **O** Oxygen 16.00	9 **F** Fluorine 19.00	10 **Ne** Neon 20.18
3	11 **Na** Sodium 22.99	12 **Mg** Magnesium 24.31											13 **Al** Aluminum 26.98	14 **Si** Silicon 28.09	15 **P** Phosphorus 30.97	16 **S** Sulfur 32.07	17 **Cl** Chlorine 35.45	18 **Ar** Argon 39.95
4	19 **K** Potassium 39.10	20 **Ca** Calcium 40.08	21 **Sc** Scandium 44.96	22 **Ti** Titanium 47.87	23 **V** Vanadium 50.94	24 **Cr** Chromium 52.00	25 **Mn** Manganese 54.94	26 **Fe** Iron 55.85	27 **Co** Cobalt 58.93	28 **Ni** Nickel 58.69	29 **Cu** Copper 63.55	30 **Zn** Zinc 65.39	31 **Ga** Gallium 69.72	32 **Ge** Germanium 72.61	33 **As** Arsenic 74.92	34 **Se** Selenium 78.96	35 **Br** Bromine 79.90	36 **Kr** Krypton 83.80
5	37 **Rb** Rubidium 85.47	38 **Sr** Strontium 87.62	39 **Y** Yttrium 88.91	40 **Zr** Zirconium 91.22	41 **Nb** Niobium 92.91	42 **Mo** Molybdenum 95.94	43 **Tc** Technetium (98)	44 **Ru** Ruthenium 101.07	45 **Rh** Rhodium 102.91	46 **Pd** Palladium 106.42	47 **Ag** Silver 107.87	48 **Cd** Cadmium 112.41	49 **In** Indium 114.82	50 **Sn** Tin 118.71	51 **Sb** Antimony 121.76	52 **Te** Tellurium 127.60	53 **I** Iodine 126.90	54 **Xe** Xenon 131.29
6	55 **Cs** Cesium 132.91	56 **Ba** Barium 137.33	57 **La** Lanthanum 138.91	72 **Hf** Hafnium 178.49	73 **Ta** Tantalum 180.95	74 **W** Tungsten 183.84	75 **Re** Rhenium 186.21	76 **Os** Osmium 190.23	77 **Ir** Iridium 192.22	78 **Pt** Platinum 195.08	79 **Au** Gold 196.97	80 **Hg** Mercury 200.59	81 **Tl** Thallium 204.38	82 **Pb** Lead 207.2	83 **Bi** Bismuth 208.98	84 **Po** Polonium (209)	85 **At** Astatine (210)	86 **Rn** Radon (222)
7	87 **Fr** Francium (223)	88 **Ra** Radium (226)	89 **Ac** Actinium (227)	104 **Rf** Rutherfordium (261)	105 **Db** Dubnium (262)	106 **Sg** Seaborgium (266)	107 **Bh** Bohrium (264)	108 **Hs** Hassium (269)	109 **Mt** Meitnerium (268)									

58 **Ce** Cerium 140.12	59 **Pr** Praseodymium 140.91	60 **Nd** Neodymium 144.24	61 **Pm** Promethium (145)	62 **Sm** Samarium 150.36	63 **Eu** Europium 151.96	64 **Gd** Gadolinium 157.25	65 **Tb** Terbium 158.93	66 **Dy** Dysprosium 162.50	67 **Ho** Holmium 164.93	68 **Er** Erbium 167.26	69 **Tm** Thulium 168.93	70 **Yb** Ytterbium 173.04	71 **Lu** Lutetium 174.97
90 **Th** Thorium 232.04	91 **Pa** Protactinium 231.04	92 **U** Uranium 238.03	93 **Np** Neptunium (237)	94 **Pu** Plutonium (244)	95 **Am** Americium (243)	96 **Cm** Curium (247)	97 **Bk** Berkelium (247)	98 **Cf** Californium (251)	99 **Es** Einsteinium (252)	100 **Fm** Fermium (257)	101 **Md** Mendelevium (258)	102 **No** Nobelium (259)	103 **Lr** Lawrencium (262)

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