

# Electrochemistry

## Introduction and Summary

Galvanic Cells allow us to harness the electron flow in a [redox reaction](#) to perform useful [work](#). Such cells find common use as batteries, pH meters, and as fuel cells. The setup of the cell, as seen in Figure 2.1, requires that the [oxidation](#) and reduction half-reactions are connected by a wire and by a salt bridge or porous disk. Electrons will flow through that wire creating an electrical current. The salt bridge or porous disk allows the passage of ions in solution to maintain charge neutrality in each half-cell. Instead of drawing the setup of every cell like Figure 2.1, chemists have devised a shorthand [line notation](#) discussed in [Line Notation](#).

The direction of the current in a cell is determined by the standard reduction potential of each half-cell. For a reaction to be spontaneous, the overall cell potential must be positive (the rules for adding cell potentials are covered under [Adding potentials](#)). Therefore, the half-reaction with the greater reduction potential will be a reduction and the other half-reaction will be an oxidation. The electrode in the oxidation half-reaction is called the [anode](#). The electrode in the reduction half-reaction is called the [cathode](#).

## Redox Reactions

### Determining Oxidation States

In order to keep track of the electron flow in [redox](#) reactions we will define the concept of [oxidation state](#) and provide some rules for determining the oxidation states of atoms within a compound. The oxidation state of an atom in a covalent compound is an imaginary charge assigned to that atom if all the electrons in its bonds were completely given to the more electronegative atom in the bond. Of course, if two atoms of the same element are bonded together, then the two electrons in its bond are shared equally. In ionic compounds, the charge on the ion is equal to its oxidation number. Figure 1.1 shows, using hydrogen peroxide as an example, how one calculates the oxidation state of an atom in a covalent compound.

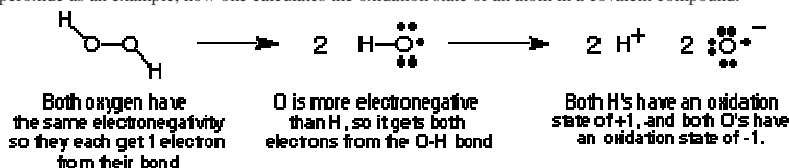


Figure 1.1: Determining the Oxidation States in a Covalent Compound

The above procedure for assigning oxidation states leads to the following useful observations about oxidation states which you should verify given the above discussion:

1. Atoms in elemental form have oxidation states of zero.
2. The charge on a monoatomic ion is equivalent to its charge.
3. Hydrogen has an oxidation state of +1 when bonded to non-metals and -1 when bonded to metals.
4. F, because it only forms one bond and is the most electronegative element, has an oxidation state of -1.
5. O, unless bonded to F or itself, has an oxidation state of -2.
6. The sum of all oxidation states in a compound must equal the total charge on the species. For hydrogen peroxide, we have seen that both O's have -1 oxidation states and both H's have +1 oxidation states, the sum of which is zero—the charge on hydrogen peroxide.

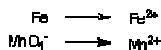
## Balancing Redox Reactions

[Oxidation](#)-reduction reactions, also called redox reactions, involve the transfer of electrons from one species to another. That electron transfer causes a change in oxidation state for both reactive partners. The reducing agent is oxidized, meaning that its oxidation number increases due to the loss of one or more electrons. The oxidizing agent is reduced, meaning that its oxidation number has decreased due to the gain of one or more electrons. For example, the following reaction 1.2 between permanganate and iron metal in acidic aqueous solution involves the transfer of five electrons to each permanganate (the oxidizing agent) from iron (the reducing agent).



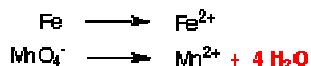
Balancing redox equations by inspection is quite difficult, as you must take into account not only the mass balance but also the charge balance in the equation. To aid in this task a set of rules, called the [Half-Reaction Method](#), has been devised. The following rules [work](#) for reactions performed in acidic or in basic solution.

1. Separate oxidation and reduction half-reactions:

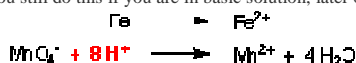


2. Balance all atoms except for hydrogen and oxygen in each half-reaction. In this example they are already balanced.

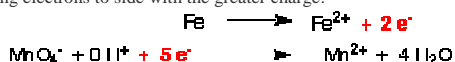
3. Balance oxygen by adding  $\text{H}_2\text{O}$  as needed:



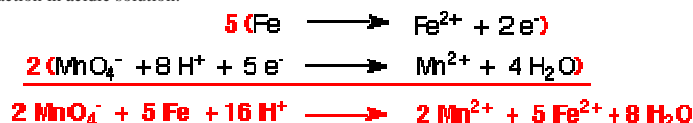
4. To balance hydrogen, add  $\text{H}^+$  as needed: (Note: You still do this if you are in basic solution, later on, you will add  $\text{OH}^-$  to "neutralize" the acid.)



5. Balance the charge of each reaction by adding electrons to side with the greater charge:

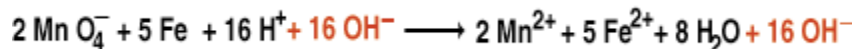


6. Multiply each half-reaction by the least integer factor that equalizes the number of electrons in each half-reaction. Then, add the half-reactions to obtain the overall balanced reaction in acidic solution:



If your redox reaction is in acidic solution, the above reaction is properly balanced. However, if the reaction you wish to balance is in basic solution, you need to add these three steps:

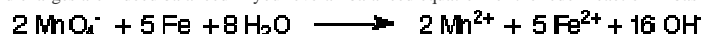
7. If the redox reaction is one in basic solution, then add  $\text{OH}^-$  to both sides of the equation to "neutralize" each  $\text{H}^+$ :



8. "React"  $\text{H}^+$  and  $\text{OH}^-$  to form  $\text{H}_2\text{O}$  and eliminate water molecules on both sides of the equation:



9. Make sure that all atoms and charges are indeed balanced in your overall balanced equation for the redox reaction in basic solution:



## Problems

### Problem 1.1:

Please provide the oxidation state of each atom in the following chemical species:

- $\text{Ag}^{2+}$
- $\text{F}_2$
- $\text{HMnO}_4$
- $\text{VO}_2^+$
- $\text{Cr}_2\text{O}_7^{2-}$
- $\text{CH}_3\text{OH}$

### [Solution]

- +2
- 0
- H is +1, each O is -2, and Mn is +7
- both O's are -2, V is +5
- each O is -2, both Cr's are +6
- O is -2, each H is +1, and C is -2.

### Problem 1.2:

Tell whether each of the following reactions are redox or not. If the reaction is a redox reaction, say what is oxidized and what is reduced.

- $\text{Na} + \text{Cl} \longrightarrow \text{NaCl}$
- $\text{Na}^+ + \text{Cl}^- \longrightarrow \text{NaCl}$
- $\text{CH}_3\text{OH} + \text{O}_2 \longrightarrow \text{CO}_2 + 2 \text{H}_2\text{O}$
- $\text{HSO}_3^- + \text{H}^+ \longrightarrow \text{H}_2\text{SO}_3$
- $4 \text{Ag} + 8 \text{CN}^- + \text{O}_2 + 4 \text{H}^+ \longrightarrow 4 \text{Ag}(\text{CN})_2^- + 2 \text{H}_2\text{O}$
- $\text{O}_3 + \text{NO} \longrightarrow \text{O}_2 + \text{NO}_2$

### [Solution]

- redox, Na oxidized, Cl reduced
- not redox
- $\text{CH}_3\text{OH}$  oxidized,  $\text{O}_2$  reduced
- not redox
- redox, Ag oxidized,  $\text{O}_2$  reduced
- redox, NO oxidized,  $\text{O}_3$  reduced

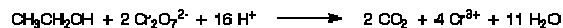
### Problem 1.3:

Balance the following redox reactions in both acidic and basic solution:

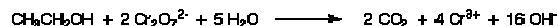
- $\text{CH}_3\text{CH}_2\text{OH} + \text{Cr}_2\text{O}_7^{2-} \longrightarrow \text{CO}_2 + \text{Cr}^{3+}$   
By the way, this is the reaction used by police in breathalyzers
- $\text{Cu} + \text{Fe}^{3+} \longrightarrow \text{Cu}^{2+} + \text{Fe}^{2+}$
- $\text{MnO}_4^- + \text{Fe} \longrightarrow \text{Mn}^{2+} + \text{Fe}^{2+}$
- $\text{VO}_2^+ + \text{Mg} \longrightarrow \text{VO}^{2+} + \text{Mg}^{2+}$

### [Solution]

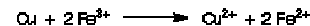
a. In acidic solution:



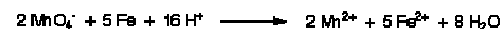
In basic solution:



b. In both acidic and basic solution:



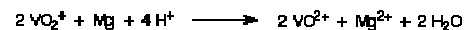
c. In acidic solution:



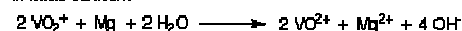
In basic solution:



d. In acidic solution:



In basic solution:



## Galvanic Cells

### Setting Up Galvanic Cells

Galvanic cells harness the electrical energy available from the electron transfer in a [redox](#) reaction to perform useful electrical [work](#). The key to gathering the electron flow is to separate the [oxidation](#) and reduction half-reactions, connecting them by a wire, so that the electrons must flow through that wire. That electron flow, called a current, can be sent through a circuit which could be part of any number of electrical devices such as radios, televisions, watches, etc.

The figure below shows two typical setups for galvanic cells. The left hand cell diagram shows an oxidation and a reduction half-reaction joined by both a wire and a porous disk, while the right hand cell diagram shows the same cell substituting a salt bridge for the porous disk.

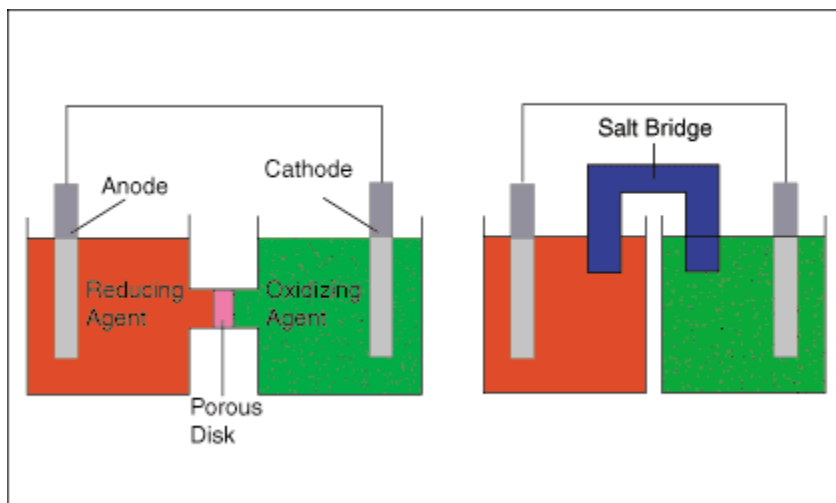


Figure 2.1: Diagram of a Galvanic Cell

The salt bridge or porous disk is necessary to maintain the charge neutrality of each half-cell by allowing the flow of ions with minimal mixing of the half-cell solutions. As electrons are transferred from the oxidation half-cell to the reduction half-cell, a negative charge builds in the reduction half-cell and a positive charge in the oxidation half-cell. That charge buildup would serve to oppose the current from [anode](#) to [cathode](#)-- effectively stopping the electron flow--if the cell lacked a path for ions to flow between the two solutions.

The above figure points out that the electrode in the oxidation half-cell is called the anode and the electrode in the reduction half-cell is called the cathode. A good mnemonic to help remember that is "**The Red Cat ate An Ox**" meaning reduction takes place at the cathode and oxidation takes place at the anode.

The anode, as the source of the negatively charged electrons is usually marked with a minus sign (-) and the cathode is marked with a plus sign (+). Physicists define the direction of current flow as the flow of positive charge based on an 18th century understanding of electricity. As we now know, negatively charged electrons flow in a wire. Therefore, chemists indicate the direction of electron flow on cell diagrams and not the direction of current. To make that point clear, the direction of electron flow is indicated on figure 2.2 with an arrow and the symbol for an electron,  $e^-$ .

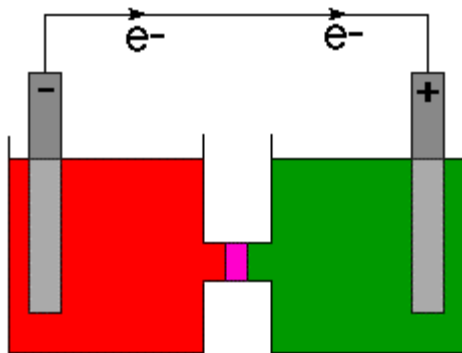
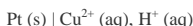


Figure 2.2: Diagram of a Galvanic Cell Showing Direction of Electron Flow

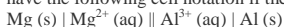
### Line Notation for Galvanic Cells

Instead of drawing a cell diagram such as figure 2.1 or figure 2.2 chemists have devised a shorthand way of completely describing a cell called [line notation](#). This notation scheme places the constituents of the cathode on the right and the anode components on the left. The phases of all reactive species are listed and their concentrations or pressures are given if those species are not in their standard states (i.e. 1 atm. for gasses and 1M for solutions). All phase interfaces are noted with a single line (|) and multiple species in a single phase are separated by commas. For example, a half-cell containing 1M solutions of CuO and HCl and a Pt electrode for the reduction of  $\text{Cu}^{2+}$  would be written as:



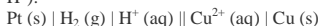
Note that the spectator ions, oxide and chloride, have been left out of the notation and that the anode is written to the far left.

The salt bridge or porous disk is shown in the notation as a double line (||). Therefore, a cell that undergoes the oxidation of magnesium by  $\text{Al}^{3+}$  could have the following cell notation if the anode is magnesium and the cathode is aluminum:



### Standard Reduction Potentials

One can measure the cell potential,  $E_{\text{cell}}$ , in volts, of any galvanic cell with the aid of a potentiometer. However, it is impossible to directly measure the potential of each individual half-cell. Chemists, however, have devised a method to measure the ability of a chemical species to reduce another by compiling tables of standard reduction potentials,  $E^\circ$  (the  $^\circ$  indicates that the reaction is at standard state). Arbitrarily assigning a value of exactly zero to the potential of the standard hydrogen electrode allows us to measure the  $E^\circ$  of any half-reaction. That measurement is made by constructing a galvanic cell between the SHE and the unknown half-cell at standard state conditions. For example, when the following cell is constructed (see [Heading](#) for a review of the line notation), an  $E^\circ_{\text{cell}}$  of 0.34 V is observed (note the setup of the SHE as the anode because  $\text{Cu}^{2+}$  has a greater reduction potential than  $\text{H}^+$ ):



Because the SHE has a potential of exactly zero volts, as defined above, the reaction:



has a value of 0.34 V for its  $E^\circ$  (recall that  $E^\circ_{\text{cell}} = E^\circ_{\text{SHE}} + E^\circ$ ). Fortunately, every important reduction potential has been measured and tabulated. Useful lists of reduction potentials are available in most introductory chemistry texts, including yours. In this SparkNote, all potentials will be given to you if you need them.

Those tables of standard reduction potentials list all half-reactions as reductions. Half-reactions with the largest reduction potential are placed at the top of the list and the smallest (most negative) reduction potentials are at the bottom. Those species on the left-hand side of the equations at the top of the list are the most easily reduced (like  $\text{F}_2$ , or  $\text{H}_2\text{O}_2$ ) and those at the bottom are the least readily reduced (like  $\text{Li}^+$ ).

Take a look at the list of standard reduction potentials in your chemistry text. An intuitive trend should be obvious when looking at the data-- electronegative species (those with the greatest attraction for electrons) are easily reduced, i.e. given an electron. The most electronegative atom, F, has the largest reduction potential whereas one of the least electronegative atoms, Li, has the smallest reduction potential.

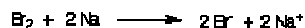
### Adding Standard Reduction Potentials

By compiling a list of standard reduction potentials of all possible reductions, one can, at least in theory, calculate the cell potential,  $E^\circ_{\text{cell}}$ , of any arbitrary redox reaction. By knowing the sign of  $E^\circ_{\text{cell}}$ , we can predict whether a reaction is spontaneous at standard conditions. If  $E^\circ_{\text{cell}}$  is positive, then the reaction is spontaneous. Conversely, if  $E^\circ_{\text{cell}}$  is negative, then the reaction is non-spontaneous as written but spontaneous in the reverse direction (see Thermodynamics, Electrical Work and Cell Potential for an explanation of why that is so).

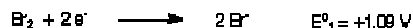
To compute the cell potential of a reaction at standard conditions,  $E^\circ_{\text{cell}}$ , you do not need to balance the equation of your redox reaction. However, as we will learn in Thermodynamics of Electrochemistry, if the reaction is not conducted at standard state, then it is essential to balance the redox reaction to compute its cell potential. For now, let's assume that all reactions are conducted at standard conditions unless otherwise specified.

When asked to compute the cell potential for a reaction, you will need to be able to separate the overall reaction into its oxidation and reduction half-reactions as in Figure 2.4.

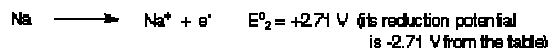
Overall Reaction:



Reduction Half-Reaction:



Oxidation Half-Reaction:



$$E^\circ_{\text{cell}} = E^\circ_1 + E^\circ_2 = 1.09 \text{ V} + 2.71 \text{ V} = 3.80 \text{ V}$$

**NOTE:  $E^\circ_{\text{cell}} \neq E^\circ_1 + 2 E^\circ_2$  because  $E^\circ$  is an intrinsic property**

Once those half reactions are separated, then find the reduction potential for the reaction written as a reduction. As you can see in Figure 2.4, one reaction is written as an oxidation. For that reaction, you need to calculate its oxidation potential. To calculate an [oxidation potential](#), simply reverse the sign of the  $E^\circ$  for the corresponding reduction reaction (just the oxidation written in the opposite direction). Simply add the reduction potential of the reduction and the oxidation potential of the oxidation to calculate the  $E^\circ_{\text{cell}}$ . It is important to note here that  $E^\circ$ 's are intrinsic properties of reactions and, therefore, do not depend on the stoichiometry of the reaction. That means that you **DO NOT** multiply the  $E^\circ$  of a reaction by the coefficient used to balance the overall redox reaction. A proof of that point is provided in Thermodynamics of Electrochemistry#. Multiplying the value of  $E^\circ$  for a half-reaction is the number one mistake made in calculating  $E^\circ_{\text{cell}}$ . Please, don't let that happen to you! Simply read off the values of  $E^\circ$  for the oxidation and reduction half-reactions and add those two values together, as in figure 2.4. /PARAGRAPH

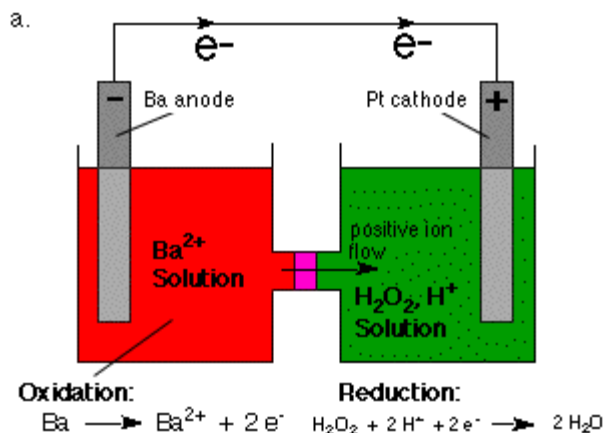
## Problems

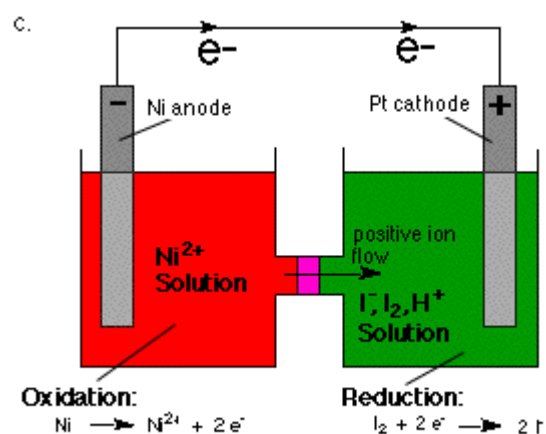
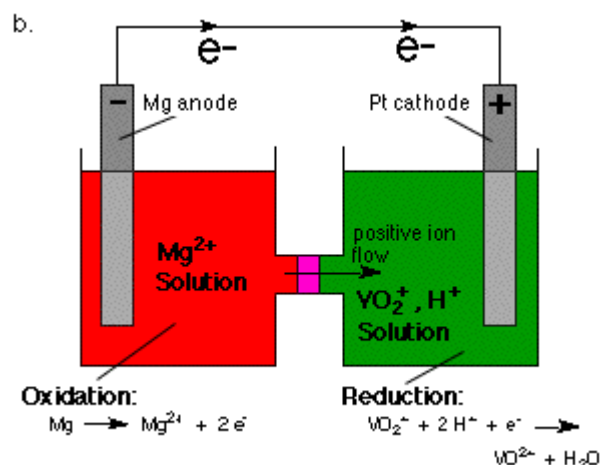
### Problem 2.1:

For each of the following redox reactions, sketch a diagram of a galvanic cell indicating the oxidation and reduction half-reactions, the anode, the cathode, the direction of electron flow in the wire and the direction of positive ion migration through the porous disk. Additionally, list the components of each half-cell. Assume each reaction proceeds in the forward direction as written.

- $\text{H}_2\text{O}_2 + 2 \text{H}^+ + \text{Ba} \longrightarrow 2 \text{H}_2\text{O} + \text{Ba}^{2+}$
- $2 \text{VO}_2^+ + \text{Mg} + 4 \text{H}^+ \longrightarrow 2 \text{VO}^{2+} + \text{Mg}^{2+} + 2 \text{H}_2\text{O}$
- $\text{I}_2 + \text{Ni} \longrightarrow 2 \text{I}^- + \text{Ni}^{2+}$

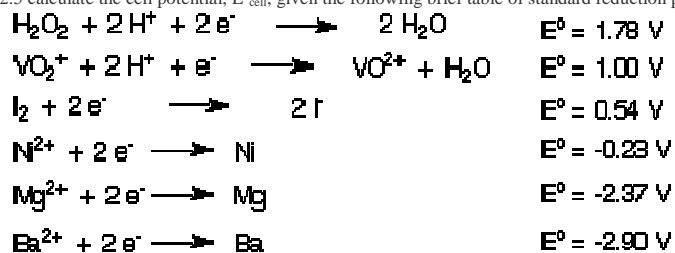
[\[Solution\]](#)





**Problem 2.2:**

For each of the above reactions 2.5 calculate the cell potential,  $E^\circ_{\text{cell}}$ , given the following brief table of standard reduction potentials.



**[Solution]**

a.  $1.78\text{ V} + 2.90\text{ V} = 4.68\text{ V}$

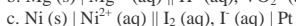
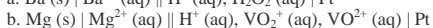
b.  $1.00\text{ V} + 2.37\text{ V} = 3.37\text{ V}$

c.  $0.54\text{ V} + 0.23\text{ V} = 0.77\text{ V}$

**Problem 2.3:**

For each of the above cells 2.5 provide the line notation that describes each. Assume that all species are in their standard states.

**[Solution]**



## Electrolysis

### Electrolytic Cells

The concept of reversing the direction of the spontaneous reaction in a [galvanic cell](#) through the input of electricity is at the heart of the idea of [electrolysis](#). See Figure 1.1 for a comparison of galvanic and electrolytic cells. If you would like to review your knowledge of galvanic cells (which I strongly suggest) before learning about [electrolytic cells](#), [click here](#).

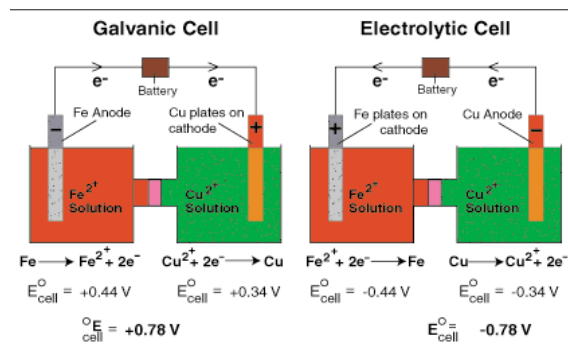


Figure 1.1: Comparison of Galvanic and Electrolytic Cells

Electrolytic cells, like galvanic cells, are composed of two half-cells—one is a reduction half-cell, the other is an oxidation half-cell. Though the direction of electron flow in electrolytic cells may be reversed from the direction of spontaneous electron flow in galvanic cells, the definition of both [cathode](#) and [anode](#) remain the same—reduction takes place at the cathode and oxidation occurs at the anode. When comparing a galvanic cell to its electrolytic counterpart, as is done in Figure 1.1, you should note that the direction of both half-reactions have been reversed. That changes the identity of the left-hand half-cell from an oxidation to a reduction half-cell and turns its anode into a cathode. A similar transformation, outlined in the diagram 1.1, occurs on the right-hand half-cell. Because the directions of both half-reactions have been reversed, the sign, but not the magnitude, of the cell potential has been reversed. Note that copper is spontaneously [plated](#) onto the copper cathode in the galvanic cell whereas it requires a voltage greater than 0.78 V from the battery to plate iron on its cathode in the electrolytic cell.

You should be asking yourself at this point how it is possible to make a non-spontaneous reaction proceed. The answer is that the electrolytic cell reaction is not the only one occurring in the system—the battery is a spontaneous [redox](#) reaction. By Hess's Law, we can sum the  $\Delta G$  of the battery and the electrolytic cell to arrive at the  $\Delta G$  for the overall process. As long as that  $\Delta G$  for the overall reaction is negative, the system of the battery and the electrolytic cell will continue to function. The condition for  $\Delta G$  being negative for the system (you should prove this for yourself) is that  $E_{\text{battery}}$  is greater than  $-E_{\text{cell}}$ .

### Electrolysis of Water

During the early history of the earth, hydrogen and oxygen gases spontaneously reacted to form the water in the oceans, lakes, and rivers we have today. That spontaneous direction of reaction can be used to create water and electricity in a galvanic cell (as it does on the space shuttle). However, by using an electrolytic cell composed of water, two [electrodes](#) and an external source [emf](#) one can reverse the direction of the process and create hydrogen and oxygen from water and electricity. Figure 1.2 shows a setup for the electrolysis of water.

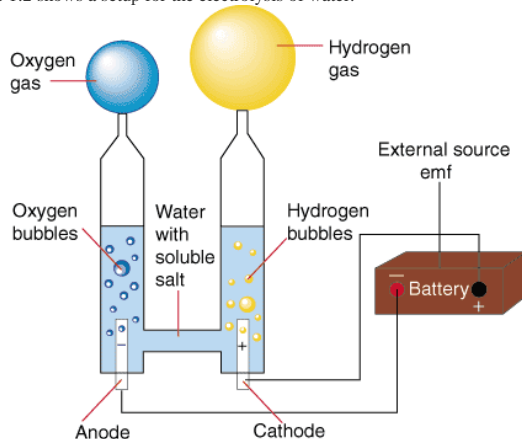


Figure 1.2: Setup for the Electrolysis of Water

The reaction at the anode is the oxidation of water to O<sub>2</sub> and acid while the cathode reduces water into H<sub>2</sub> and hydroxide ion. That reaction has a potential of -2.06 V at standard conditions. However, this process is usually performed with [H<sup>+</sup>] = 10<sup>-7</sup> M and [OH<sup>-</sup>] = 10<sup>-7</sup> M, the concentrations of hydronium and hydroxide in pure water. Applying the Nernst Equation to calculate the potentials of each half-reaction, we find that the potential for the electrolysis of pure water is -1.23 V. To make the electrolysis of water occur, one must apply an external potential (usually from a battery of some sort) of greater than or equal to 1.23 V. In practice, however, it is necessary to use a slightly larger voltage to get the electrolysis to occur on a reasonable time scale. Pure water is impractical to use in this process because it is an electrical insulator. That problem is circumvented by the addition of a minor amount of soluble salts that turn the water into a good conductor (as noted in Figure 1.2). Such salts have subtle effects on the electrolytic potential of water due to their ability to change the pH of water. Such effects from the salts are generally so small that they are usually ignored.

## Electroplating

Electroplating allows the production of metal coatings of such desirable commodities as silver and gold. People make fortunes gold or silver plating junk metal (usually aluminum) because they can sell gold plated necklaces for a comparable price to the real thing (or even pass them off as being solid gold). That's how [electrochemistry](#) can be used to rip you off! In our discussion of electroplating, we will discuss how you can set up a cell for electroplating, how you can calculate the amount of precious material consumed, and various other calculations you can perform with electroplating. In terms of the variety of electrochemistry problems possible to ask, this section, perhaps rivaled by Thermodynamics, is the richest.

The setup for electroplating is quite simple and the entire cell is usually conducted in a single solution as shown in Figure 1.3.

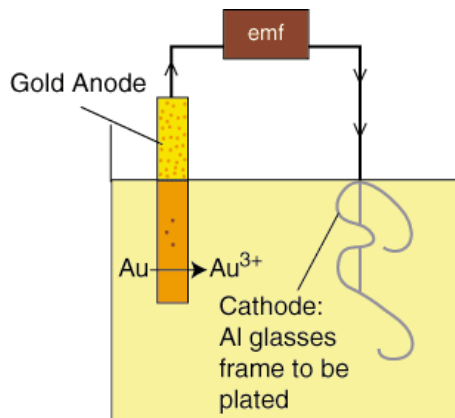


Figure 1.3: Electroplating Setup

The gold from the anode is oxidized and dissolves in solution as  $\text{Au}^{3+}$ . The electrons arriving at the aluminum glasses frame cathode reduce the  $\text{Au}^{3+}$  in solution to  $\text{Au}$  (s) on the surface of the frame cathode. We can calculate how long we should have our glasses frame in solution if we want a certain amount of gold to be plated.

Let's assume it takes 1.0 g of gold to provide an adequate coating for our glasses and also assume that we are using an emf sufficient to produce 10 amperes (A) of current (1 A = 1 coulomb per second). Let's now calculate how long it will take to plate that 1.0 g of gold.

$$\# \text{sec} = 1.0 \text{ g Au} \times \frac{1 \text{ mol Au}}{197 \text{ g Au}} \times \frac{3 \text{ mol e}^-}{1 \text{ mol Au}} \times \frac{96500 \text{ C}}{1 \text{ mol e}^-} \times \frac{1 \text{ sec}}{10 \text{ C}} = 146 \text{ sec}$$

Figure 1.4: Electroplating Setup

As you can see from the above calculation, such a problem only involves the use of unit cancellation. To calculate the time needed to deposit a certain amount of material, you need to start with the amount, converted to moles. Then, multiply by the number of electrons consumed in the reduction (in this case 3). Using the definition of a faraday, 96500 C per mole of electrons, you can convert between moles and charge. Finally, by using the definition of an ampere, 1 C per second, you can convert the amount of charge required to deposit the material into a time in seconds. There are various ways of phrasing this same problem such as "how much gold is deposited in 146 seconds at 10 A" or "what current is required to deposit 1.0 g of gold in 146 seconds." Don't be fooled by those permutations of the same problem, they all boil down to simple unit cancellation which you have been doing since you learned how to do [stoichiometry](#). Also note that in these problems, you do not need to know the cell potential. Students often try, incorrectly, to use the cell potential somewhere in that calculation. Furthermore, you need only know the number of electrons transferred--you could solve the same problem without even knowing what material was being plated (as long as you know its molar mass).

## Problems and Solutions

### Problem 1.1:

What volume of  $\text{O}_2$  (g) at  $0^\circ\text{C}$  and 760 mmHg is produced when 5 A of current is passed through a dilute aqueous solution of KCl for 5 minutes?

[\[Solution\]](#)

$$\begin{aligned} V \text{ O}_2 &= 5 \text{ minutes} \times \frac{60 \text{ s}}{1 \text{ min.}} \times \frac{5 \text{ C}}{1 \text{ s}} \times \frac{1 \text{ mol e}^-}{96500 \text{ C}} \\ &\times \frac{1 \text{ mol O}_2}{4 \text{ mol e}^-} \times \frac{22.4 \text{ L}}{1 \text{ mol O}_2} = 0.087 \text{ L O}_2 \end{aligned}$$

### Problem 1.2:

Molecular masses can be determined through electroplating. Determine the molecular mass and identity of a +2 metal, X, that plates 46.3g of X in 6.75 hours at a current of 2 A.

[\[Solution\]](#)

#### Solution for Problem 1.2

$$\# \text{g/mol} = \frac{46.3 \text{ g}}{6.75 \text{ h}} \times \frac{1 \text{ h}}{3600 \text{ s}} \times \frac{1 \text{ s}}{2 \text{ C}} \times \frac{96500 \text{ C}}{1 \text{ mol e}^-} \times \frac{2 \text{ mol e}^-}{1 \text{ mol X}} = 183.9 \text{ g/mol}$$

X is, therefore, tungsten.



**Problem 1.3:**

How long will it take for 1.25 L of a 1.0 M CuSO<sub>4</sub> solution being electrolyzed with a current of 3.40 A to reach a concentration of 0.25 M?

[\[Solution\]](#)

Solution for Problem 1.3

$$\# \text{mo Cu} = 1.25 \text{ L} \times \frac{1 \text{ mol}}{1 \text{ L}} = 1.25 \text{ mol Cu}$$

$$\# \text{mol Cu} = 1.25 \text{ L} \times \frac{0.25 \text{ mol}}{1 \text{ L}} = 0.31 \text{ mol Cu}$$

therefore, 0.94 moles of Cu have been plated onto the cathode

$$\# \text{h} = 0.94 \text{ mol Cu} \times \frac{2 \text{ mol e}^-}{\text{mol Cu}} \times \frac{96500 \text{ C}}{\text{mol e}^-} \times \frac{1 \text{ s}}{3.4 \text{ C}} \times \frac{1 \text{ h}}{3600 \text{ s}} = 14.8 \text{ h}$$