

Electrochemical Cells and Electrolysis

An electrochemical, galvanic, or voltaic cell converts chemical potential energy into electricity. It relies on the difference in activity between the two metals that make up the cathode and anode or the two metal ions in the solutions.

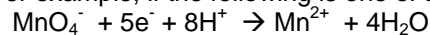
The electrons flow as a result of a **spontaneous** reaction. The spontaneity can be somewhat predicted based on the activity series. Species higher up on the chart will act as the anode while species lower on the chart will act as cathodes. A more detailed prediction of the output voltage can be calculated using Standard Reduction Potentials (which we are not going to cover in this class).

Factors that affect the voltage output of a cell are the activity difference in the two metals (emf), concentration, and temperature. A quantitative description is given by the Nernst equation (also something we're going to skip).

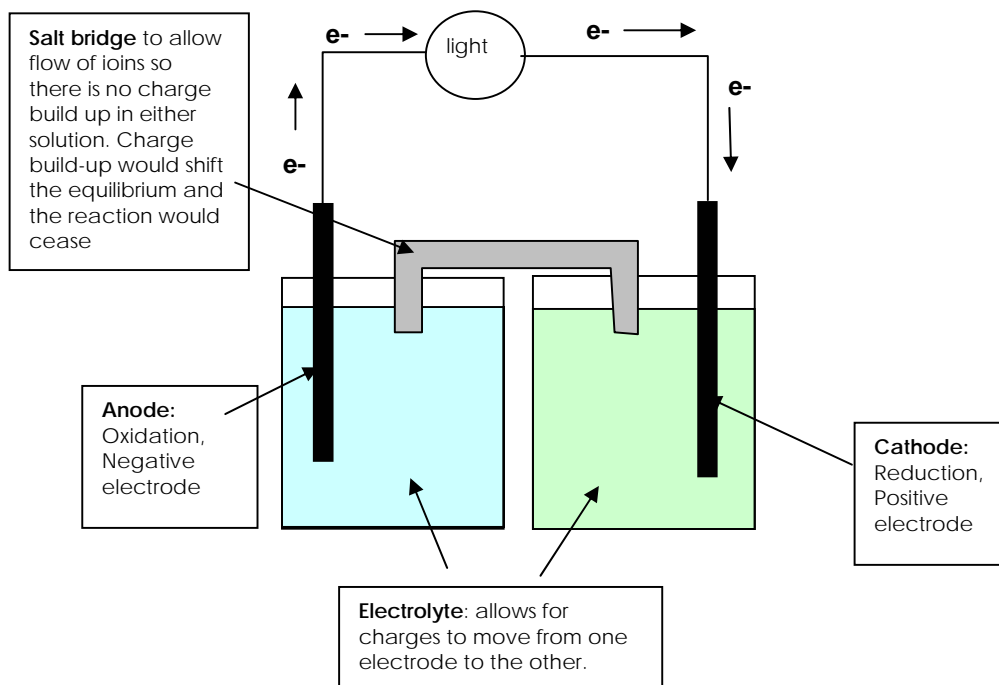
At the anode, oxidation takes place, the neutral metal atoms lose electrons which flow through the wire to the light bulb causing it to light up. The positive ions that form as a result of oxidation, dissolve in solution.

At the cathode, reduction takes place, the positive ions in solution gain electrons that have flowed from the anode. The ions become neutral atoms and precipitate out of solution onto the cathode.

Sometimes an inert but conducting material is used as an electrode when one or both of the species is non-conducting. (For example, if the following is one of the half-reactions in the cell:



Neither MnO_4^- nor Mn^{2+} can serve as electrodes since they are not solid so a Platinum electrode could be used)



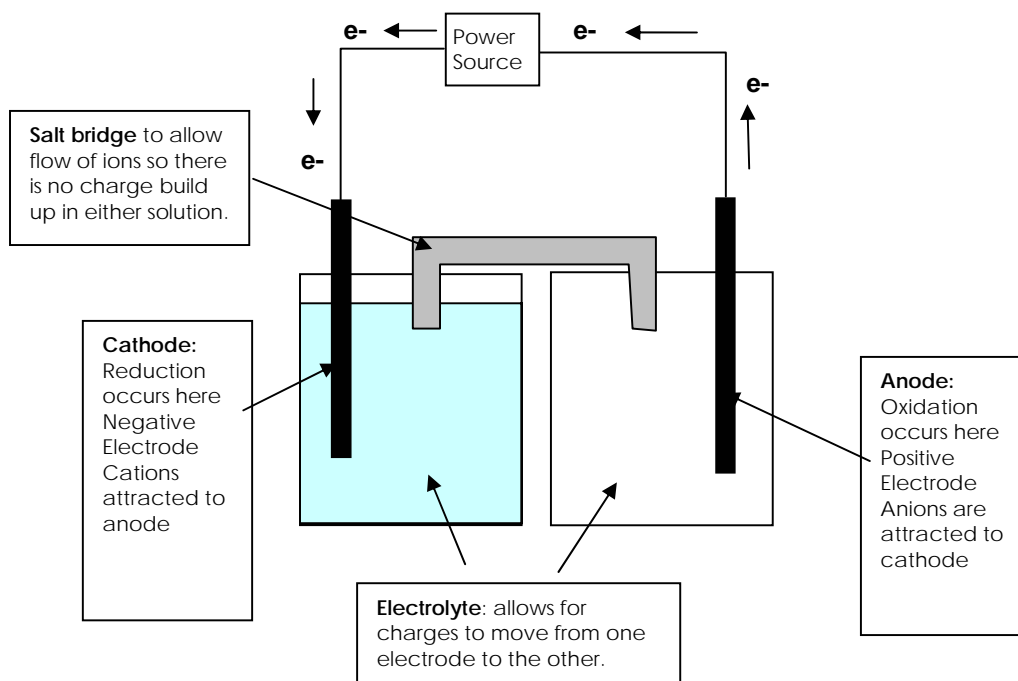
Electrolytic cells are the opposite of galvanic cells:
They convert electrical energy into chemical potential energy.

Electricity must be supplied in order to force a **non-spontaneous** reaction to occur. The spontaneity can be somewhat predicted based on the activity series.

Current flows from the species being oxidized to the species being reduced.

Major Differences:

- Direction of electron flow changes
- Cathode and anode switch so that oxidation still occurs at the anode and reduction still occurs at the cathode.
- Notice that cathode is now the negative electrode, it attracts cations
- Notice that the anode is now the positive electrode, it attracts anions
- Notice that there is a power supply instead of a light bulb



Examples of electrolysis in action are:

- making hydrogen and oxygen gas from water
 - $2\text{H}_2\text{O} \rightarrow 2\text{H}_2 + \text{O}_2$
 - $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$ (Reduction at negative electrode/cathode)
 - $\text{O}^{2-} \rightarrow \text{O}_2 + 4\text{e}^-$ (Oxidation at positive electrode/anode)
- Making sodium and chloride from molten salt
 - $\text{NaCl(l)} \rightarrow \text{Na} + \text{Cl}_2$
 - Cathode: $\text{Na}^+ + 1\text{e}^- \rightarrow \text{Na}$
 - Anode: $2\text{Cl}^- + 2\text{e}^- \rightarrow \text{Cl}_2$
- Electroplating precious metals onto less expensive metals