Unit 8, Lesson 05: Electrochemical (Galvanic) Cells

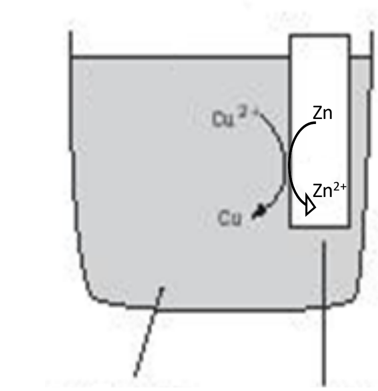
These youtube videos by Tyler DeWitt are excellent. Before reading the notes, please watch:
1. Introduction to Electrochemistry
2. Galvanic Cells (Voltaic Cells)

Tyler discusses chemical reactions that generate electricity (electrochemical cells). This is the material that we will cover in this unit. He also discusses using electricity to make chemical reactions happen that wouldn't happen otherwise (for example, electrolysis and electroplating). This is material for next year.

During a redox reaction, electrons are transferred from the substance being oxidized (the reducing agent) to the substance being reduced (the oxidizing agent). For example, when a piece of zinc is placed in a solution containing Cu^{2+} ions, the following reactions take place:

$$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2\,e^-$$
$$Cu^{2+}_{(aq)} + 2\,e^- \rightarrow Cu_{(s)}$$

Net ionic equation: $Cu^{2+}_{(aq)} + Zn_{(s)} \rightarrow Cu_{(s)} + Zn^{2+}_{(aq)}$

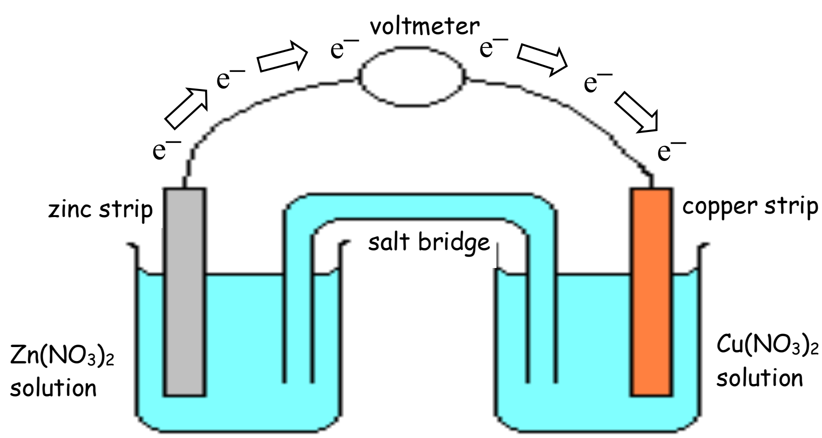


The oxidation and the reduction occur where the zinc atoms and Cu^{2+} ions are in contact with one another on the surface of the piece of zinc. The electrons are transferred directly from the zinc atoms to the Cu^{2+} ions.

In an electrochemical cell, the same reactions take place but the half-reactions are physically separated. The oxidation and reduction reactions take place at different surfaces called electrodes. The electrons get from one electrode to the other by travelling through a conducting wire that connects these electrodes, and this creates an electric current (flow of electrons).

The salt bridge contains an electrolyte (salt) solution such as potassium nitrate. This permits the flow of ions from one half-cell to the other and is necessary to maintain the electrical neutrality of the two solutions (to keep the charges balanced).

The construction of a simple electrochemical cell is illustrated below:



Anode	**Cathode**
- site of oxidation	- site of reduction
- negative electrode	- positive electrode
- electrons leave this electrode	- electrons flow toward this electrode

Callan Scott, a former student, came up with a clever way to remember which electrode is the anode:

'LEONA' which stands for **L**oss of **E**lectrons is **O**xidation at the **N**egative electrode, the **A**node

Using the activity series from Lesson #1 or page 470 of your text, we can state the following:

- zinc is above copper on the activity series, so zinc is more reactive than copper
- zinc loses its electrons more easily so it will donate electrons to the copper
- electrons will flow from the zinc half-cell toward the copper half-cell
- zinc will be oxidized and copper will be reduced
- the electrode at which oxidation occurs is called the anode, so zinc is the anode
- the electrode at which reduction occurs is called the cathode, so copper is the cathode
- electrons flow from the anode, so the anode (zinc) is the negative electrode
- electrons flow toward the cathode, so the cathode (copper) is the positive electrode

We can represent this cell using "Galvanic Cell Notation". The anode is always written on the left:

$$Zn\,(s) \;\;|\;\; Zn^{2+}\,(aq) \;\;\|\;\; Cu^{2+}\,(aq) \;\;|\;\; Cu\,(s)$$

$$\text{means:}\quad Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} \qquad \Uparrow \qquad Cu^{2+}_{(aq)} \rightarrow Cu_{(s)}$$

physical
separation

Example #1: A piece of tin metal is placed in a solution of tin (IV) nitrate. A piece of aluminum metal is placed in a solution of aluminum chloride. The pieces of metal are connected with a wire and a salt bridge is used to connect the solutions.

a) use the activity series to determine which metal is more reactive
b) write the two half reactions that will take place
c) determine which electrode will be the anode
d) write the galvanic cell notation for this electrochemical cell.

a) aluminum is above tin on the activity series, so aluminum will lose electrons and tin will accept them

b) the two half reactions are: $A\ell\,(s) \rightarrow A\ell^{3+}(aq) + 3\,e^{-}$ (oxidation)

$Sn^{4+}\,(aq) + 4\,e^{-} \rightarrow Sn\,(s)$ (reduction)

c) the aluminum electrode is losing electrons, so it is the anode (LEONA)

d) the galvanic cell notation is: $A\ell\,(s) \;|\; A\ell^{3+}\,(aq) \;\|\; Sn^{4+}\,(aq) \;|\; Sn\,(s)$

Some redox reactions include non-metal reactants in the gas phase. These reactions take place on the surface of an inert metal electrode such as platinum.

Example #2: An electrode is made by placing a piece of nickel metal in a solution of nickel (III) nitrate. Chlorine gas is bubbled over an inert platinum electrode which is placed in a solution chloride ions. The pieces of metal are connected with a wire and a salt bridge is used to connect the solutions.

Metal atoms tend to lose electrons while non-metals tend to gain them, therefore the nickel electrode will be the anode. (In the next lesson, we will see how to write the reaction involving chlorine)

The two half reactions are: $Ni\,(s) \rightarrow Ni^{3+}(aq) + 3\,e^{-}$ (oxidation)

$C\ell_2\,(g) + 2\,e^{-} \rightarrow 2\,C\ell^{1-}\,(aq)$ (reduction)

The galvanic cell notation is: $Ni\,(s) \;|\; Ni^{3+}\,(aq) \;\|\; C\ell_2\,(g) \;|\; C\ell^{1-}\,(aq) \;|\; Pt\,(s)$

the single lines represent species in different states or phases