ELECTROCHEMICAL CELLS

- *Allessandra Volta* (1745-1827) invented the **electric cell** in 1800
  - A single cell is also called a *voltaic cell, galvanic cell or electrochemical cell*.
  - Volta joined several cells together to produce a **battery (***voltaic pile***).**

- These cells use a *spontaneous* redox reaction to generate electrical energy by facilitating the passage of electrons through an *external circuit*.
- They convert chemical energy into electrical energy

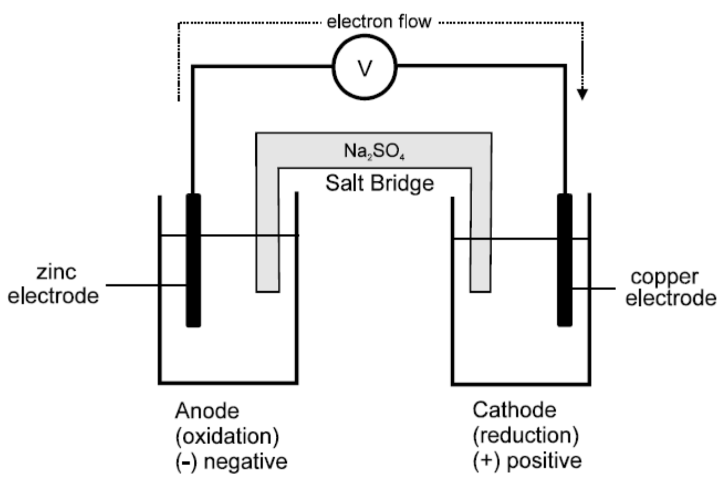
- Each cell is composed of:
  a) **2 electrodes** - solid conductors (anode and cathode); usually 2 metals
  b) **electrolytes** - aqueous conductors; usually neutral solutions of ionic compounds
  c) **salt bridge** - contains an electrolytic solution that doesn't interfere with the reaction
  d) **external circuit** - allows movement of electrons from anode to cathode; usually a wire that is connected to both electrodes
  e) **internal circuit** - allows movement of positive and negative ions in solution; usually a salt bridge or porous barrier

  Note: Together the internal and external circuits constitute a complete circuit, allowing an electrical current to be produced.

Example:

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

Zn and Cu are placed in different compartments (*half –cells*) each containing an electrolytic solution, $ZnSO_{4(aq)}$ and $CuSO_{4(aq)}$, respectively.



Anode – The site of oxidation and electron release

Cathode – The site of reduction and electron consumption

Memory Tip: Oxidation takes place at the anode, both begin with vowels(O and A).
Reduction takes place at the cathode, both begin with consonants (R and C).

How an Electrochemical cell works:

1. At the *anode,* $Zn_{(s)}$ atoms lose electrons to form $Zn^{2+}_{(aq)}$ ions. These ions are released into the zinc sulfate solution. (The electrons do not enter the solution).

   $Zn_{(s)} \rightarrow Zn^{2+} + 2e-$

   This reaction produces an excess of positive zinc ions in solution and an excess of electrons on the zinc electrode. This gives the zinc electrode a negative charge.

2. Electrons flow from the anode (Zn) where they are in excess, through the connecting wire (*external circuit*), to the cathode (Cu) where there is a shortage of electrons.

3. At the c*athode,* copper ions in solution gain electrons to form copper atoms. These copper atoms are deposited on the copper electrode.

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

   This reaction removes Cu^{2+} ions from the copper(II)sulfate solution leaving an excess of negative sulfate ions. It also results in a shortage of electrons on the Cu electrode, thus giving the electrode a positive charge (relative to the Zn electrode).

4. The aqueous ions in the *salt bridge* maintain the neutrality of the electrolytic solutions of the two half cells.
   a) Negative ions (SO_4^{2-}) move from the salt bridge into the half-cell that contains the anode and an excess of positive ions (Zn^{2+}) in solution
   ie. anions move toward the anode

   b) Positive ions (Na^+) move from the salt bridge into the half-cell that contains the cathode and an excess of negative ions (SO_4^{2-}) in solution
   ie. cations move toward the cathode

   The salt bridge also connects the two half-cells internally. This is necessary to complete the circuit and thereby maintain the current. Thus the current through the cell solutions is carried by moving ions.

   Only certain electrolytes can be used in the salt bridge. If the salt reacts with other ions in the half-cells or within the electrode, it will interfere with the redox reaction.

   Good choices: sulfates, nitrates and chlorides of alkali metals (Na, K, etc), NH_4^+,Zn or Cu

Cell Notation

· a method of describing an electrochemical cell without drawing a diagram

Anode / anode electrolyte / salt bridge / cathode electrolyte / cathode

$Zn / [Zn^{2+}] / / [Cu^{2+}] / Cu$

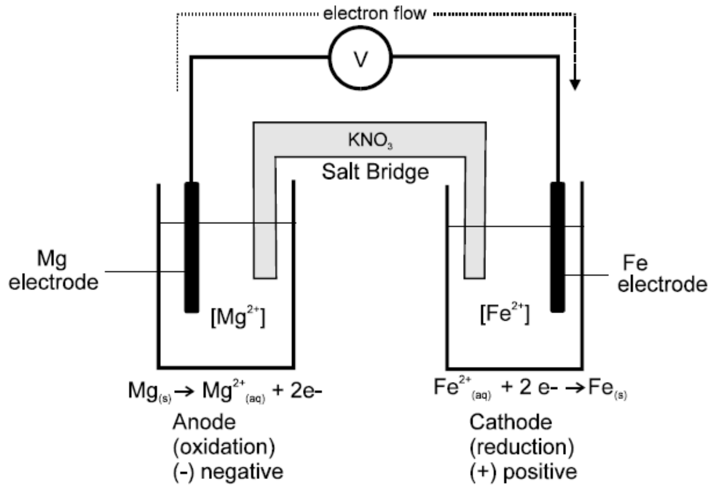
***Examples**:*

1. Sketch and label an electrochemical cell that makes use of the following spontaneous redox reaction. Include the cell notation as well.

   $Mg_{(s)} + Fe^{2+}_{(aq)} \rightarrow Mg^{2+}_{(aq)} + Fe_{(s)}$

   $Mg_{(s)} \rightarrow Mg^{2+}_{(aq)} + 2e-$ Oxidized at the anode
   $2e- + Fe^{2+}_{(aq)} \rightarrow Fe_{(s)}$ Reduced at the cathode

Electrochemical cell diagram



Cell notation: Mg / [Mg^{2+}] // [Fe^{2+}] / Fe

1. For the following reaction : $Fe_{(s)} + CuSO_{4\,(aq)} \rightarrow Fe_2(SO_4)_{3\,(aq)} + 3Cu_{(s)}$
   a) Balance the equation
   b) Give the cell notation
   c) Draw the schematic diagram of the galvanic cell

ELECTROCHEMISTRY WORKSHEET #4

1. An electrochemical cell used the two half-reactions

$$Cu^{2+} + 2e^- \rightarrow Cu_{(s)}$$
$$Sn_{(s)} \rightarrow Sn^{2+} + 2e^-$$

a.) Select a soluble compound that could be used to make a one molar solution for each of the half-cells and the salt bridge.
b.) Write the two half-reactions, labeling one reduction and one oxidation.
c.) Draw the diagram for the electrochemical cell and label all parts. Show the direction of electron movement.
d.) Write the net reaction.
e.) Draw the cell notation for this electrochemical cell.

2. Given the following overall reaction:

$$2I^-_{(aq)} + Fe^{3+}_{(aq)} \rightarrow Fe^{2+}_{(aq)} + I_{2\,(s)}$$

a.) Write the half-reactions
b.) Identify the oxidizing agent and the reducing agent.
c.) Sketch the galvanic cell and indicate the direction of electron flow.
d.) Pick a suitable salt bridge.
e.) Draw the cell notation for this galvanic cell.

3. A student in the lab, immersed some aluminum foil in a beaker containing 1.0 mol/L $Al(NO_3)_{3\,(aq)}$. Some silver foil was also placed in a beaker containing 1.00 mol/L $AgNO_{3(aq)}$. The two beakers were connected with a $KNO_{3\,(aq)}$ salt bridge and the metal foil electrodes were connected with a wire through a voltmeter. A reaction occurred.

   a.) Write the net ionic equation.
   b.) Sketch the galvanic cell and identify the anode and cathode.
   c.) Which direction will the electrons travel in the wire?
   d.) What cations are present in the cell?
   e.) Draw the cell notation for the cell.

4. For the following cell notation:

   $Na_{(s)}$ / Na^{+} // $Zn^{2+}_{(aq)}$ / $Zn_{\,(s)}$

   a.) Write the half-reactions
   b.) Write the overall reaction
   c.) Sketch the galvanic cell
   d.) Identify the cathode and anode

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