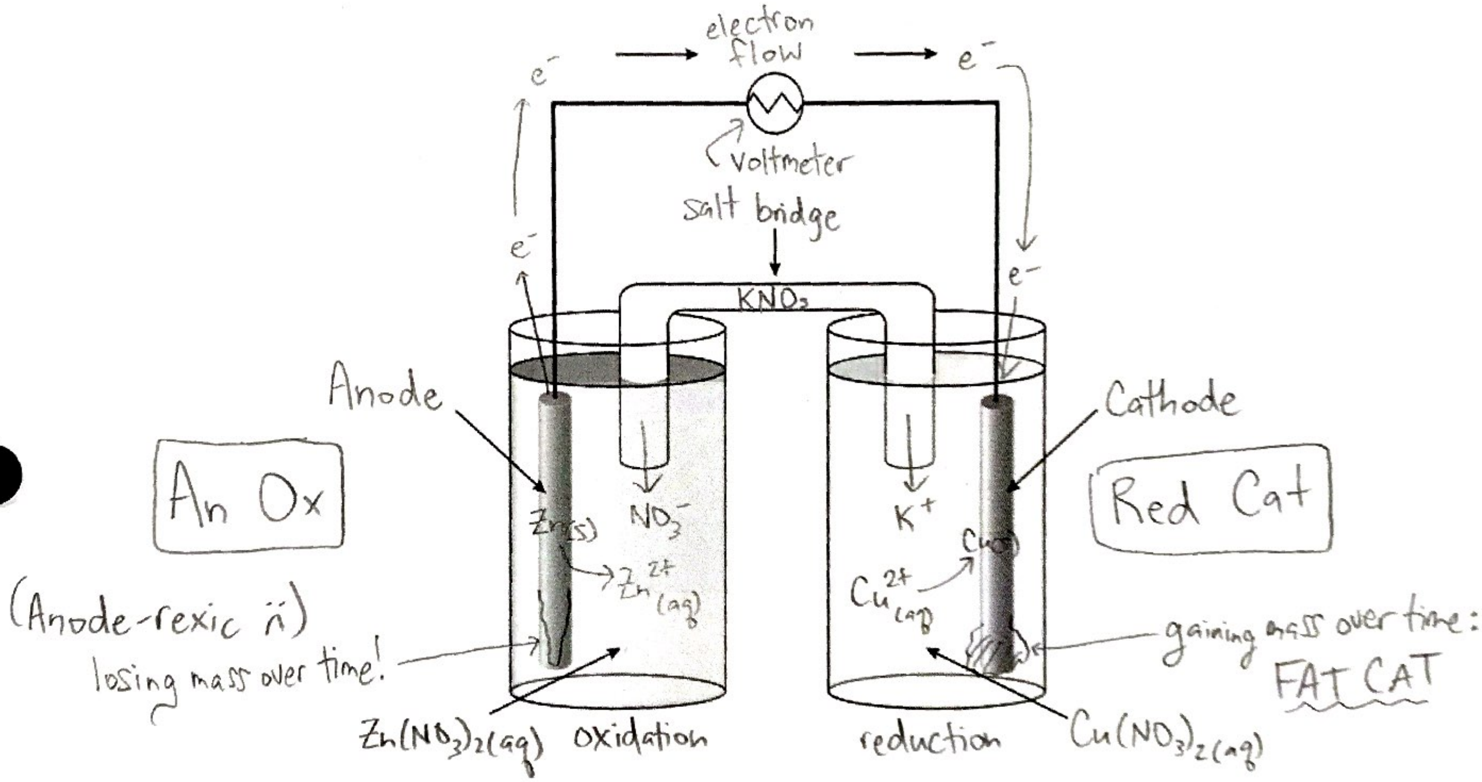
Totally Epic AP Chem Review: Galvanic/Voltaic Cells!

Galvanic/ Voltaic Cells

Redox reactions involve the transfer of electrons from one substance to another, and thus have the potential to generate an electric current (i.e. flow of electrons). To use that current, we need to Separate the place where oxidation is occurring from the place where reduction is occurring.

- Current is the number of electrons that flow through the system per second.
- Current is measured in amperes, or Amp (A) = 1 coulomb of charge per second.

This known as a **voltaic** (or **galvanic**) cell: the most common form of which is a battery! Galvanic (voltaic) cells are always thermodynamically favorable (spontaneous) and thus have a + E°_{cell}.



Parts of the Galvanic Cell

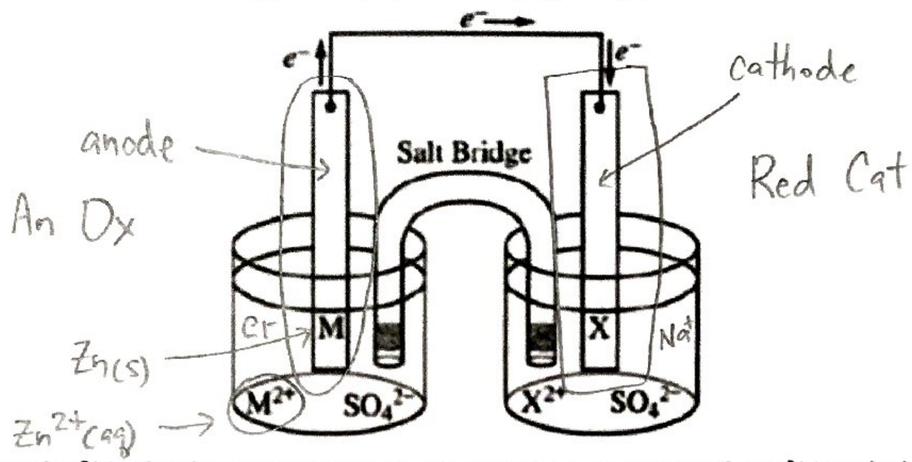
1. Anode (−): electrode where **oxidation** occurs (loses mass as reaction progresses)
2. Cathode (+): electrode where **reduction** occurs (gains mass as reaction progresses, metal 'plated')
3. Salt bridge (or disk): provides ions to balance the charge build-up in each cell
   - **Anions** (−) flow to the anode; **cations** (+) flow to the cathode
   - If the salt bridge is removed, current will slow and then Stop (V = 0) as charge builds up in half-cells.
4. Voltmeter: measures the cell potential (emf or E°) in volts
   - Over time, voltage in the galvanic cell will decrease as [reactants] ↓ and [products] ↑

Electron Flow: ALWAYS through the wire from anode to cathode (alpha order)

Ion-ion or Ion-Gas Redox: a voltaic cell can be constructed where the underlying redox reaction involves a gas or the conversion from one ion to another; requires an inert electrode: doesn't take part in the redox reaction but provides surface on which electrons can transfer; commonly made of platinum (expensive) or graphite (cheap)

1. The diagram below shows the experimental setup for a typical electrochemical cell that contains two standard half-cells. The cell operates according to the reaction represented by the following equation.

$$\overset{ox}{Zn(s)} + \overset{red}{Ni^{2+}(aq)} \rightarrow Zn^{2+}(aq) + Ni(s)$$



a. Identify M and M^{2+} in the diagram and specify the initial concentration for M^{2+} in solution.

b. Indicate which of the metal electrodes is the anode and which the cathode.

c. Write the balanced equation for the reaction that occurs in the half-cell containing the

   i. cathode. ii. anode

d. Circle the electrode that loses mass as the reaction progresses. As mass is "lost", where does it go?

e. Put a rectangle around the electrode that gains mass as the reaction progresses. Where does the mass come from?

f. If the salt bridge is made of NaCl(s), describe the ion flow that would occur in the salt bridge as the cell operates. Add the correct ions to each half-cell in the diagram above.

g. Describe what would happen if the salt bridge was removed. Justify your answer.

(a) $M = Zn_{(s)}$, $M^{2+} = Zn^{2+}_{(aq)}$ $\Rightarrow [M^{2+}] = 1.0\ M$

(c) anode: $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$ Cathode: $Ni^{2+}_{(aq)} + 2e^- \rightarrow Ni_{(s)}$

(d) As the anode loses mass, that mass goes into sol'n as $Zn^{2+}_{(aq)}$ ions,

(e) As the cathode gains mass, the mass comes from the sol'n ($Ni^{2+}_{(aq)}$) and is plated onto the cathode as solid nickel.

(f) Na^+ would flow to the cathode + Cl^- would flow to the anode to balance the charge build-up as the cell runs.

(g) Cell voltage would drop to zero over time if the salt bridge was removed, b/c there would be no ions to balance out the charge build-up in each half cell.