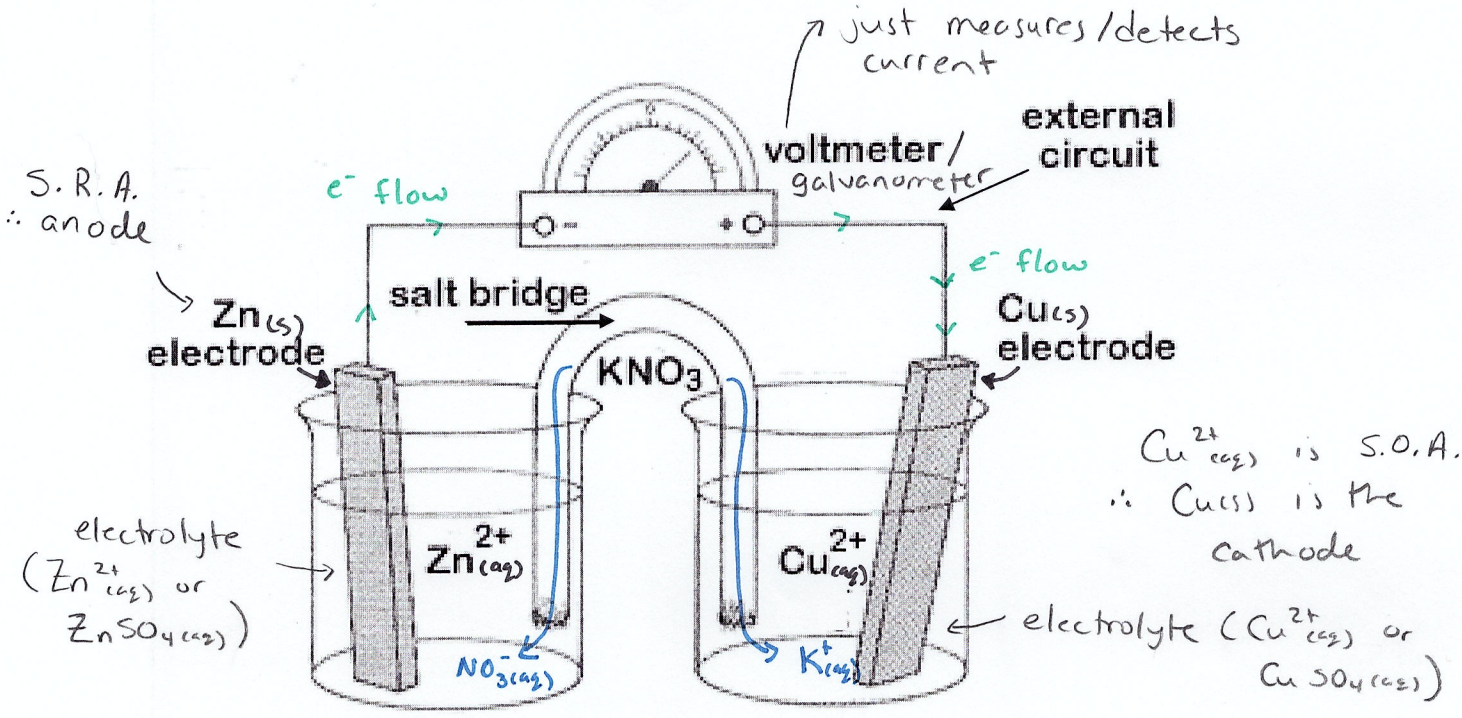
Voltaic Cells

- **Electrochemistry** is the study of the processes involved in converting chemical energy to electrical energy(ie. moving charges) and converting electrical energy to chemical energy
- A **voltaic cell** is a device that uses *spontaneous* redox reactions to transform chemical potential energy into electrical energy. In other words, a voltaic cell uses a spontaneous chemical redox reaction to produce an electrical current to flow.
- The following is a simple design of a voltaic call. Notice how the voltaic cell has two compartments/beakers.



Oxidation half-rxn

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

∴ $Zn_{(s)}$ metal (anode) decomposes & $Zn^{2+}_{(aq)}$ concentration increases

Reduction half-rxn

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

∴ $Cu^{2+}_{(aq)}$ decreases in concentration & $Cu_{(s)}$ metal is formed (anode builds up in mass)

physical observations based of half-rxns

Know components & functions! ✳

- A voltaic cell has several components
  - **Anode:** the electrode at which oxidation occurs (electrons leave the anode)
  - **Cathode:** the electrode at which reduction occurs (electrons flow into the cathode)
  - **Electrodes:** usually metal conductors that carry electrons in and out of the cell (ie. allows for electron transfer)
  - **Electrolyte:** a solution that contain ions; usually made from a soluble, ionic compound dissolved in water.
  - **External current:** a circuit outside the reaction vessel that allows electrons to flow
  - **Salt Bridge:** a U-shaped tube that contains an electrolytic solution (ie. salt) that allows ions to flow between the two beakers to maintain a charge balance, but keeps each half-reaction separate * not involved in redox rxn!

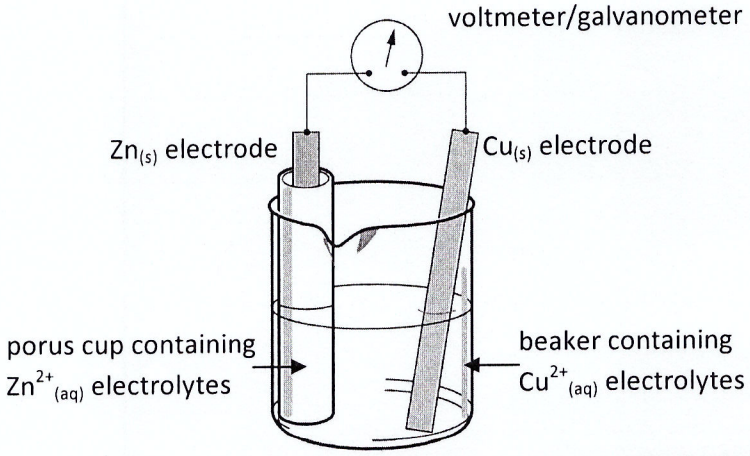
Important ✳

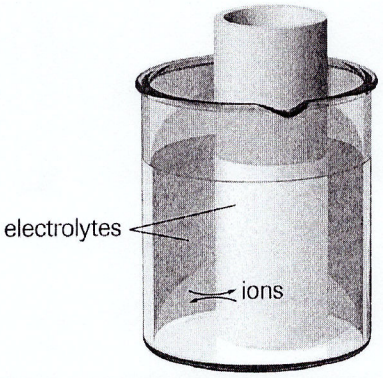
- The metal that acts as the cathode will be found in the half-cell that contains the strongest oxidizing agent (ie. has the most positive electric potential). The metal that acts as the anode will be found in the half-cell that contains the strongest reducing agent (ie. has the most negative electric potential).
  - This will ensure that the redox reaction will be spontaneous because the strongest oxidizing agent will be higher than the strongest reducing agent.

$Zn_{(s)}$ is S.R.A
∴ found at anode side & $Zn_{(s)}$ is anode

$Cu^{2+}_{(aq)}$ is S.O.A
∴ found at cathode side & $Cu_{(s)}$ is cathode

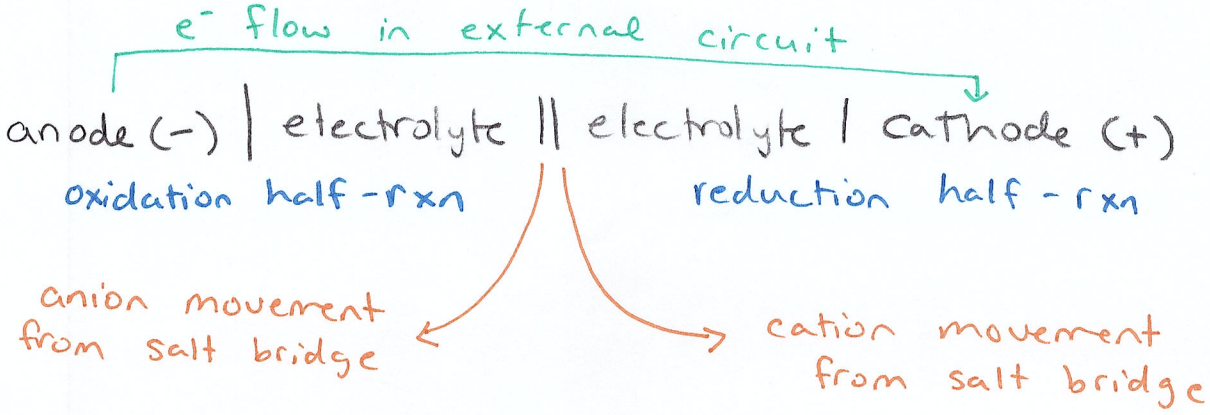
- Another simple design for a voltaic cell is to use a *porous cup* instead of a salt bridge
  - A **porous cup** is a porcelain cup that will allow ions to flow between solutions in order to maintain a charge balance, but keeps the bulk of the solutions from mixing





*****Demo: Build this Cell and Measure the Voltage*****

- The information pertaining to a voltaic cell can be condensed and summarized into a form called **cell notation**
  - Every single vertical line, | represents a phase boundary between the electrode and the solution in a half-cell
  - Every double vertical line, || represents the porous barrier or salt bridge between the half-cells



ie. $Zn_{(s)} \mid Zn^{2+}_{(aq)} \parallel Cu^{2+}_{(aq)} \mid Cu_{(s)}$

OR

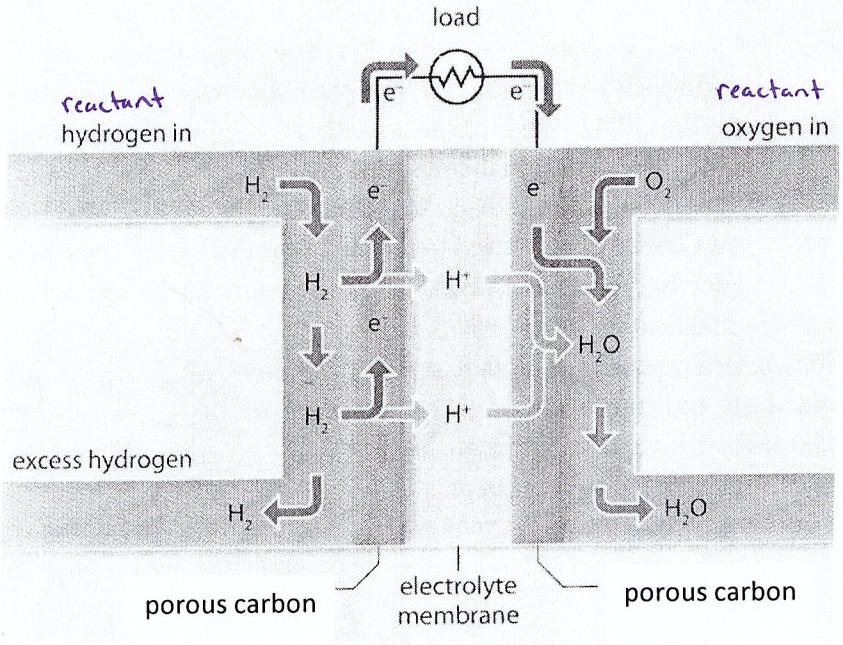
$Zn_{(s)} \mid ZnSO_{4(aq)} \parallel CuSO_{4(aq)} \mid Cu_{(s)}$

* might include spectator ions!

*****Now try Practice Problems #1, 2, 5 *****

- EXAMPLES: Identify the anode and cathode for the voltaic cell. Write out the individual half-reactions that occur at each electrode and write out the overall redox reaction that takes place in the cell.

1. A proton exchange membrane (PEM) fuel cell.



half-rxns copied from pg. 7 of data book

Strongest R.A.
Oxidation: (anode side)

$2H_{2(g)} \rightarrow 4H^+_{(aq)} + 4e^-$ (x2)

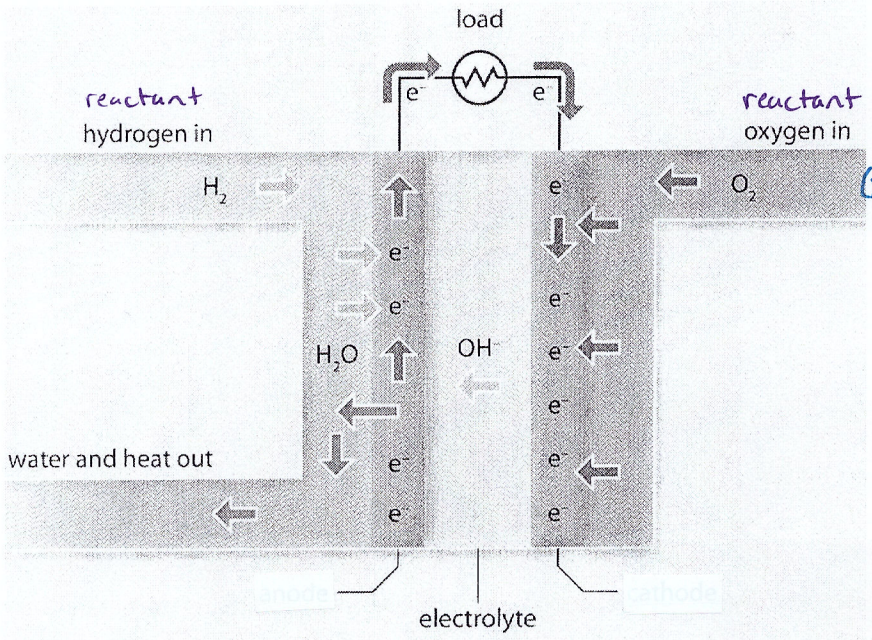
Strongest O.A.
reduction: (cathode side)

$O_{2(g)} + 4H^+_{(aq)} + 4e^- \rightarrow 2H_2O_{(\ell)}$

Overall redox:

$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(\ell)}$

2. An alkaline fuel cell



Strongest R.A.
Oxidation: (anode side)

(x2) $2H_{2(g)} + 4OH^-_{(aq)} \rightarrow 4H_2O_{(\ell)} + 4e^-$

4 − 2 = 2 H_2O

Strongest O.A.
Reduction: (cathode side)

$O_{2(g)} + 2H_2O_{(\ell)} + 4e^- \rightarrow 4OH^-_{(aq)}$

Overall redox:

$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(\ell)}$

INERT ELECTRODES

- For voltaic cells that contain half-reactions that don't involve a metal (ie. dissolved electrolytes or gases), it is necessary to use an *inert electrode*
  - An **inert electrode** is an electrode made of material that is neither a reactant nor a product of the redox reaction, but will allow for the flow of electrons
  - Most common types of inert electrodes are platinum (Pt) and carbon (C)

↗ ie. $O_{2(g)} + 2H^+_{(aq)} + 2e^- \rightleftharpoons H_2O_{2(e)}$

- Consider the following redox reaction that could possibly be used in a voltaic cell.

$$Pb_{(s)} + 2FeCl_{3(aq)} \rightarrow 2FeCl_{2(aq)} + PbCl_{2(aq)}$$

0 +3 −1 +2 −1 +2 −1

oxidized: Pb(s) → Pb²⁺(aq) + 2e⁻

reduced: Fe³⁺ + e⁻ → Fe²⁺

oxidation: $Pb_{(s)} \longrightarrow Pb^{2+}_{(aq)} + 2e^-$ (anode side) ∴ $Pb_{(s)}$ is anode

reduction: $Fe^{3+} + e^- \longrightarrow Fe^{2+}$ (cathode side)

* no metal for cathode ∴ an inert electrode is needed!

$$Pb_{(s)} \mid Pb^{2+}_{(aq)} \,||\, Fe^{3+}_{(aq)}, Fe^{2+}_{(aq)} \mid Pt_{(s)}$$

↑ inert electrode

S.O.A ∴ cathode side

EXAMPLE: Consider the following voltaic cell:

S.R.A ∴ anode side

inert electrode → $C_{(s)} \mid Fe^{2+}_{(aq)}, Fe^{3+}_{(aq)} \parallel Cr_2O_7^{2-}{}_{(aq)}, H^+_{(aq)} \mid C_{(s)}$ ← inert cathode

a. Write the half-reactions and the overall redox reaction that occur in the voltaic cell.

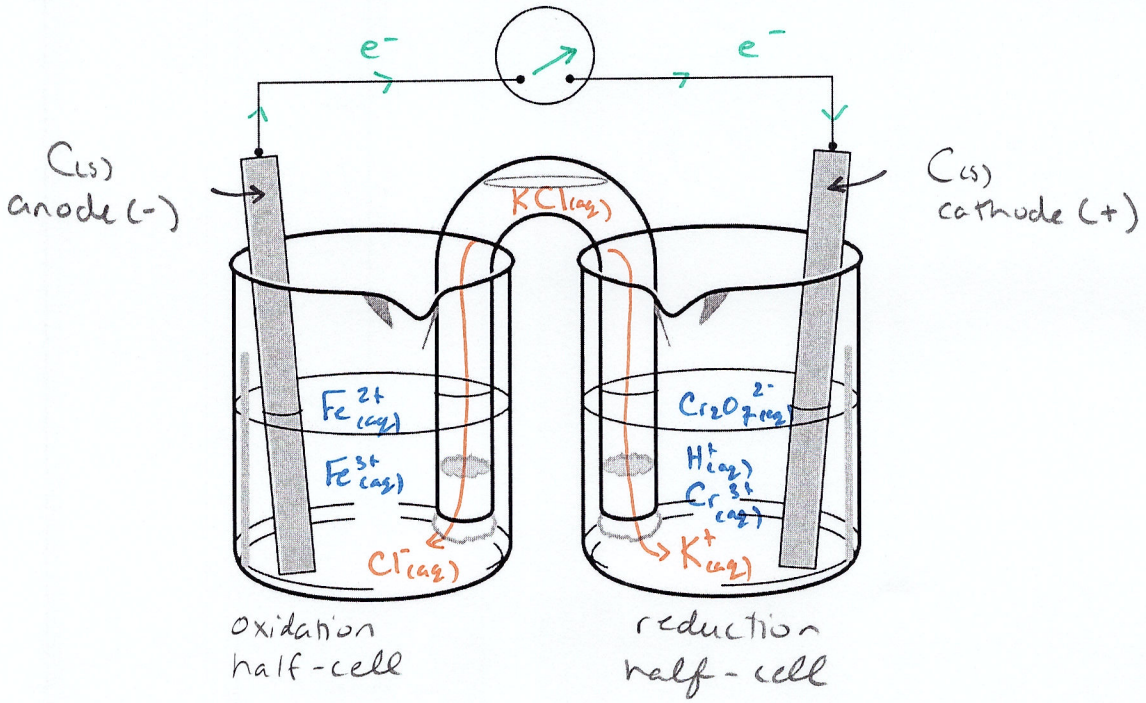
b. Draw a diagram of the cell and label the electrodes, electrolytes, direction of electron flow, and direction if ion movement from the salt bridge.

a.) $6Fe^{2+}_{(aq)} \longrightarrow 6Fe^{3+}_{(aq)} + 6e^-$ (x6)

$Cr_2O_7^{2-}{}_{(aq)} + 14H^+_{(aq)} + 6e^- \longrightarrow 2Cr^{3+}_{(aq)} + 7H_2O_{(\ell)}$

* copied from pg. 7 of data book *

$Cr_2O_7^{2-}{}_{(aq)} + 14H^+_{(aq)} + 6Fe^{2+}_{(aq)} \longrightarrow 6Fe^{3+}_{(aq)} + 2Cr^{3+}_{(aq)} + 7H_2O_{(\ell)}$



Now try pg. 482 #6 and Practice Problems #3, 4, 6