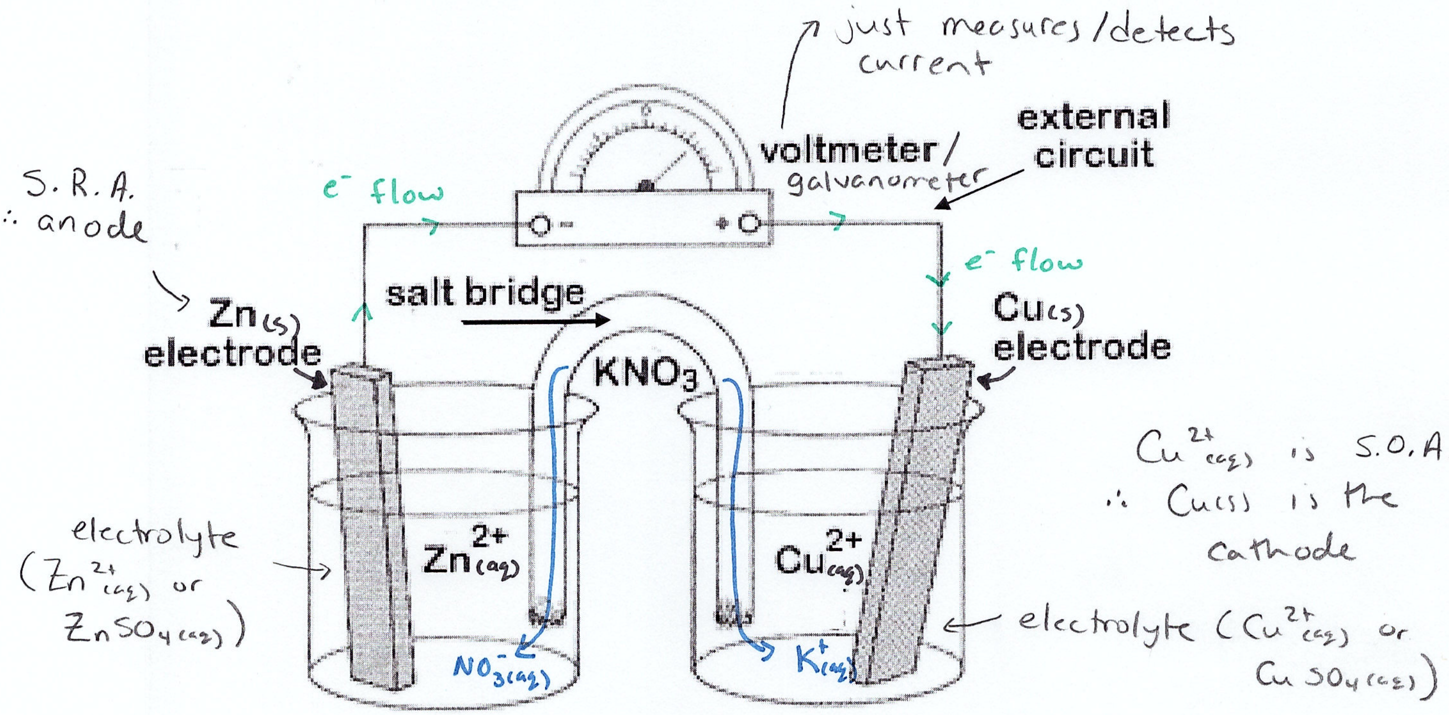
Voltaic Cells

- **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!

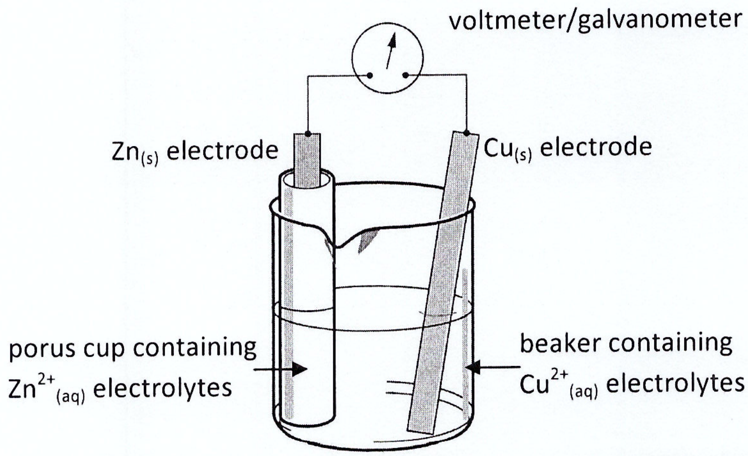
Important

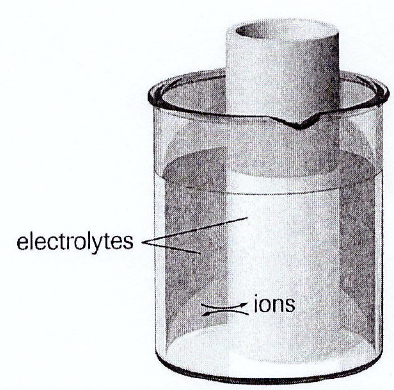
* - The metal that acts as the cathode will be found in the half-cell that contains the strongest oxidizing agent (ie. has the most positive electric potential). The metal that acts as the anode will be found in the half-cell that contains the strongest reducing agent (ie. has the most negative electric 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

e^- flow in external circuit

anode (−) | electrolyte || electrolyte | cathode (+)

oxidation half-rxn

reduction half-rxn

anion movement from salt bridge

cation movement from salt bridge

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

OR

$Zn_{(s)} | ZnSO_{4(aq)} || CuSO_{4(aq)} | Cu_{(s)}$

* might include spectator ions!

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

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

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)

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

$$\underset{0}{Pb_{(s)}} + \underset{+3\ -1}{2FeCl_{3(aq)}} \rightarrow \underset{+2\ -1}{2FeCl_{2(aq)}} + \underset{+2\ -1}{PbCl_{2(aq)}}$$

oxidized

reduced

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

reduction: (cathode side) $Fe^{3+} + e^- \rightarrow Fe^{2+}$ * no metal for cathode ∴ an inert electrode is needed!

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

inert electrode

EXAMPLE: Consider the following voltaic cell:

S.O.A ∴ cathode side

S.R.A ∴ anode side

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

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

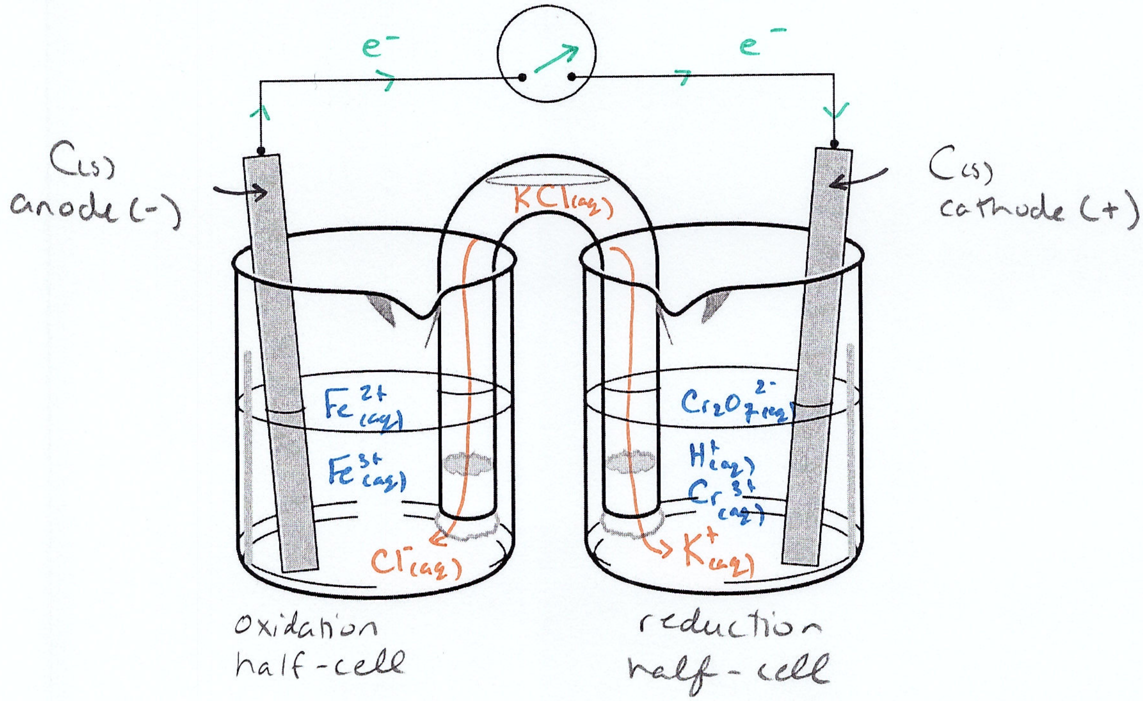
b. Draw a diagram of the cell and label the electrodes, electrolytes, direction of electron flow, and direction if ion movement from the salt bridge.

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

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

* copied from pg. 7 of data book *

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



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

Practice Problems

1. Label each component of the following cell notation. The labels to correctly label include: phase boundary, porous barrier or salt bridge, anode, cathode, electrolyte in oxidation-half reaction, and electrolyte in reduction-half reaction.

anode

electrolyte in oxidation half-rxn

electrolyte in reduction half-rxn

$Sn_{(s)} \mid Sn^{2+}_{(aq)} \parallel Tl^{+}_{(aq)} \mid Tl_{(s)}$

phase boundary

porous barrier or salt bridse

cathode

2. Use voltaic cell notation to represent a voltaic cell that has a copper electrode in copper (II) sulfate solution and an iron electrode in iron(II) nitrate solution. Write the balanced half-reactions that are occurring at the cathode and anode. Write the balanced redox reaction that is occurring in the cell.

$Cu_{(s)}$, $Cu^{2+}_{(aq)}$, $Fe_{(s)}$, $Fe^{2+}_{(aq)}$

SOA ∴ cathode side

SRA ∴ anode side

$Cu^{2+} + 2e^- \longrightarrow Cu_{(s)}$ reduction at cathode

$Fe_{(s)} \longrightarrow Fe^{2+}_{(aq)} + 2e^-$ oxidation at anode

$Cu^{2+}_{(aq)} + Fe_{(s)} \longrightarrow Cu_{(s)} + Fe^{2+}_{(aq)}$

$Fe_{(s)} \mid Fe^{2+}_{(aq)} \parallel Cu^{2+}_{(aq)} \mid Cu_{(s)}$

3. Consider the following half-reactions:

Oxidation: $2I^-(aq) \rightarrow I_2(s) + 2e^-$ (x5)

Reduction: $MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(\ell)$ (x2)

a) Write the net cell reaction.

$10\,I^-_{(aq)} + 2MnO_4^-{}_{(aq)} + 16H^+_{(aq)} \rightarrow 5I_{2(s)} + 2Mn^{2+}_{(aq)} + 8H_2O_{(\ell)}$

b) A voltaic cell based on this reaction uses inert electrodes, such as graphite electrodes. Explain why.

None of the reactants or products are made of a solid material capable of transfering electrons.

c) Write the cell notation representation of this cell.

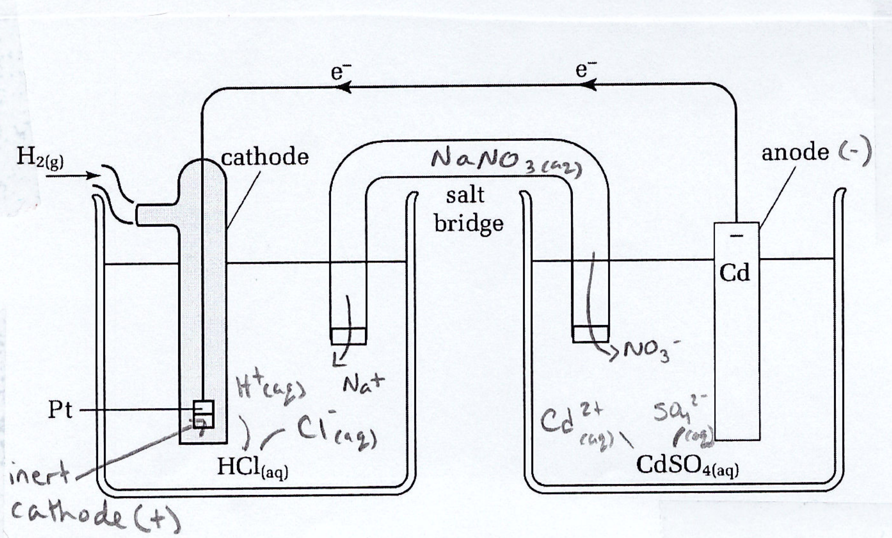
$Pt_{(s)} \,|\, I^-_{(aq)}, I_{2(s)} \,||\, H^+_{(aq)}, MnO_4^-{}_{(aq)}, Mn^{2+}_{(aq)} \,|\, Pt_{(s)}$

4. A voltaic cell operates by having $H^+(aq)/H_2(g)$ in one half cell and $Cd(s)/Cd^{2+}(aq)$ in the other half-cell.

a. Sketch a possible design for the cell. Label the specific cathode, anode, electrolytes, electron flow, and ion movement.

b. Write the half reactions that occur at each electrode.

c. Use the half reactions to determine (i) which electrode will decompose and (ii) whether the concentration of $H^+_{(aq)}$ will increase or decrease.



reduction (cathode): $2H^+_{(aq)} + 2e^- \rightarrow H_{2(g)}$

oxidation (anode): $Cd_{(s)} \rightarrow Cd^{2+}_{(aq)} + 2e^-$

∴ concentration of H^+ decreases

∴ anode decomposes

5. An electrochemical cell is represented by the cell notation $Ag^+_{(aq)}$, $NO_3^-{}_{(aq)}$

S.R.A ∴ anode side → Cd(s) | $Cd(NO_3)_2$(aq) || $AgNO_3$(aq) | Ag(s) S.O.A. ∴ cathode side

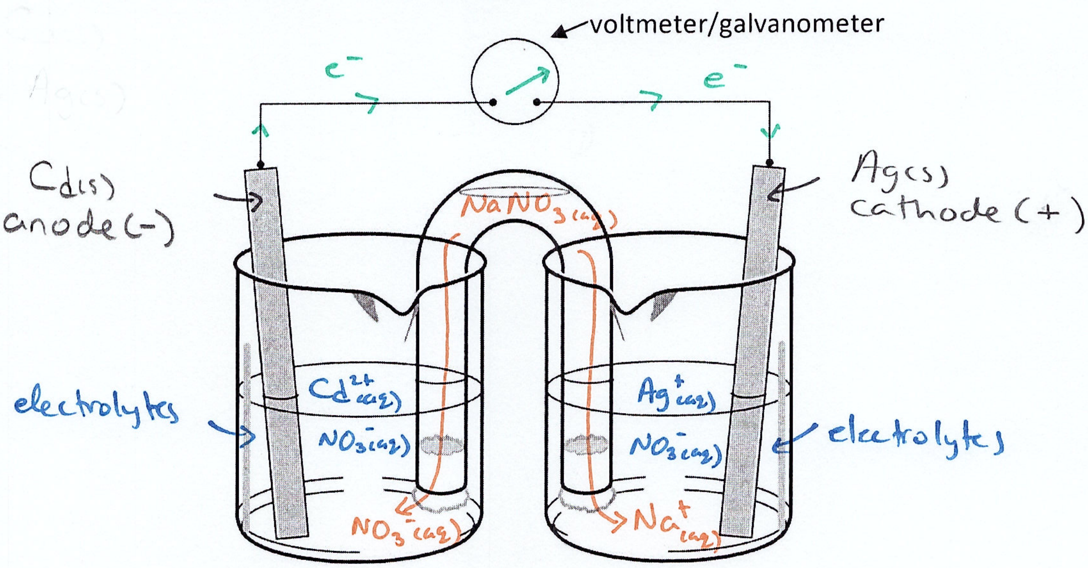
a. Identify what metal is the anode and what is the cathode.

b. Identify the strongest oxidizing and reducing agents.

c. What happens to the concentration of $Cd^{2+}{}_{(aq)}$ when the cell is connected?

d. Write balanced half-reactions to represent what reaction is occurring at the cathode and anode. Use the two half-reactions to balance the redox reaction that is taking place in the cell.

e. Label what the specific cathode, anode, electrolytes, electron flow, and ion movement is on the cell diagram below.



a.) anode is $Cd_{(s)}$
cathode is $Ag_{(s)}$

b.) S.O.A. is $Ag^+_{(aq)}$
S.R.A. is $Cd_{(s)}$

c.) $Cd_{(s)} \rightarrow Cd^{2+}_{(aq)} + 2e^-$
∴ concentration of Cd^{2+} increases

d.) anode: (oxidation) $Cd_{(s)} \rightarrow Cd^{2+}_{(aq)} + \cancel{2e^-}$
cathode: (reduction) $2Ag^+_{(aq)} + \cancel{2e^-} \rightarrow 2Ag_{(s)}$ (x2)

redox: $Cd_{(s)} + 2Ag^+_{(aq)} \rightarrow Cd^{2+}_{(aq)} + 2Ag_{(s)}$

6. Use the following diagram of a voltaic cell to answer the next questions. You will need the following half-reaction which does not appear in your data book:

$2\,IO_3^-{}_{(aq)} + 12\,H^+{}_{(aq)} + 10\,e^- \rightarrow I_{2(s)} + 6\,H_2O_{(l)}$ $\quad E^\circ = +1.20\text{ V}$

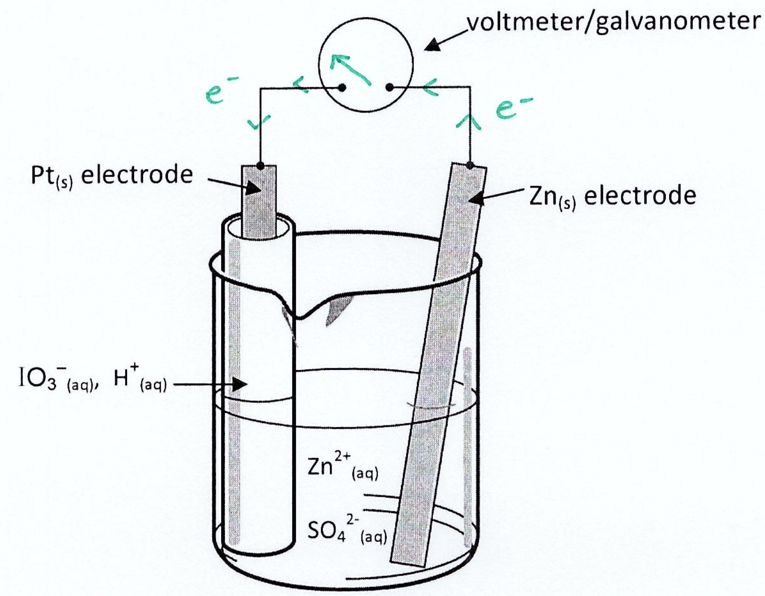
a. What is the anode and cathode?

b. What is the strongest oxidizing and reducing agent?

c. What happens to the pH of the solution in one of the half-cells? (Remember, as the concentration of $H^+{}_{(aq)}$ increases, pH decreases and vise-versa).

d. What happens to the mass of the $Zn_{(s)}$ electrode?

e. Label the direction of electron flow in the diagram below.



a.) anode is $Zn_{(s)}$ & cathode is inert (ie. $Pt_{(s)}$ or $C_{(s)}$)

b.) S.R.A is $Zn_{(s)}$ & S.O.A. is $IO_3^-{}_{(aq)}$, $H^+{}_{(aq)}$

c.) $2IO_3^-{}_{(aq)} + 12H^+{}_{(aq)} + 10e^- \rightarrow I_{2(s)} + 6H_2O$

$\therefore$ $H^+{}_{(aq)}$ will decrease in concentration $\Rightarrow$ pH will increase

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

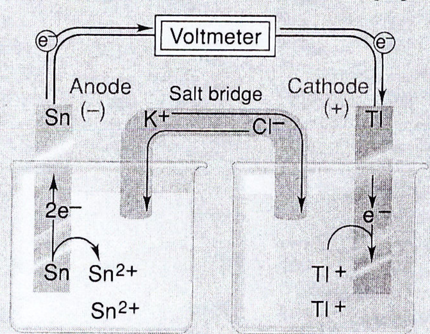
$\therefore$ mass of $Zn_{(s)}$ electrode will decrease

Student Textbook page 482

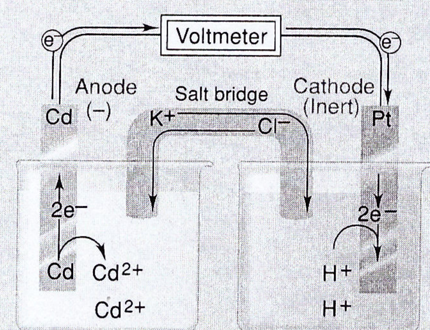
Q4. (a) In a test tube, the zinc would be the reducing agent and the copper(II) ions would be the oxidizing agent.

(b) In a Daniel cell, the copper(II) ions and the copper metal would be at the cathode while the zinc and zinc ions would be at the anode. Again students can determine this by remembering that electron flow is from anode to cathode, from reducing agent to oxidizing agent.

Q5. (a)



(b)



(a) anode: $Sn(s) \rightarrow Sn^{2+}(aq) + 2e^-$
cathode: $Tl^+(aq) + 1e^- \rightarrow Tl(s)$
Net: $Sn(s) + 2Tl^+(aq) \rightarrow Sn^{2+}(aq) + 2Tl(s)$

(b) anode: $Cd(s) \rightarrow Cd^{2+}(aq) + 2e^-$
cathode: $2H^+(aq) + 2e^- \rightarrow H_2(g)$
Net: $Cd(s) + 2H^+(aq) \rightarrow Cd^{2+}(aq) + H_2(g)$

Platinum is an inert electrode. A comma indicates that both species are in the same phase. There is a line instead of a comma to indicate that hydrogen gas is produced.

Q6. (a) The two half-reactions are:

$$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(\ell)$$

$$2I^-(aq) \rightarrow I_2(s) + 2e^-$$

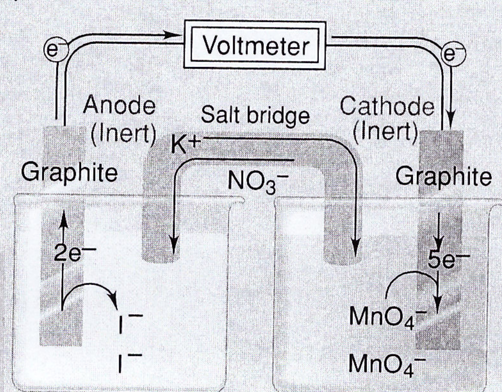
The overall balanced reaction is:

$$2MnO_4^-(aq) + 16H^+(aq) + 10I^-(aq) \rightarrow 5I_2(s) + 2Mn^{2+}(aq) + 8H_2O(\ell)$$

(b) The oxidizing agent is $MnO_4^-(aq)$ and the reducing agent is $I^-(aq)$.

(c) To form solid iodine from iodide ions, electrons must be lost (i.e. the ions must undergo oxidation). Since oxidation occurs at the anode, the solid iodine forms at the anode.

(d)



* Graphite is carbon!