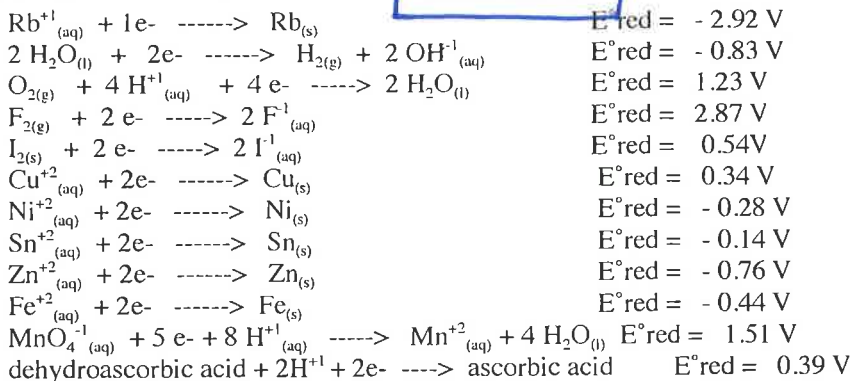


AP Chem Redox Test Review!

Redox Review!



1a. Rank these from strongest to weakest oxidizing agent:

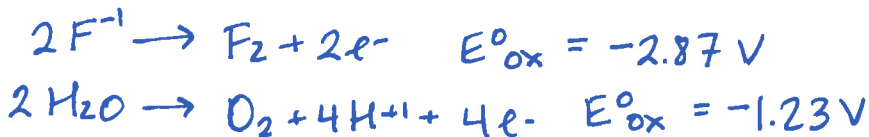
Ni^{+2}	Rb	Rb^{+1}	H_2O
$E^{\circ}_{\text{reduct.}}: -0.28 \text{ V}$		-2.92 V	-0.83 V
#1 (strongest)	#4 (weakest)	#3	#2

very negative

The strongest oxidizing agent is the substance that is most capable of being reduced; the substance with the highest (most +) reduction potential.

1b. Which is a stronger reducing agent: H_2O or F^{-1} ?

The stronger reducing agt. is the substance most capable of being oxidized; the most positive E°_{red} .



↑ had to reverse the reduction half rxns on the chart to get oxidation half rxns.

↑ more positive, so H_2O is the stronger reducing agt.

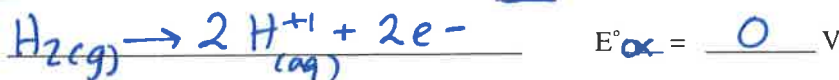
2a. For the "standard hydrogen electrode,"

a. What substances are present? At what pressures and/or concentrations are they present?

H_2 (gas) at 1 atm
 H^{+1} (aq) at 1 Molar

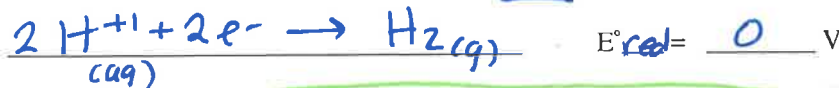
(both of these)

b. Write the half reaction that occurs at the anode, and fill in the half cell voltage.



oxidation occurs at the anode

c. Write the half reaction that occurs at the cathode, and fill in the half cell voltage.



reduction occurs at the cathode

d. Use the voltages given above to determine an activity series of metals and hydrogen.

Include rubidium, nickel, tin, iron, zinc, copper, hydrogen.

The higher on the activity series, the more capable the metal (or H) is of being oxidized; the higher the oxidation potential (E°_{ox})

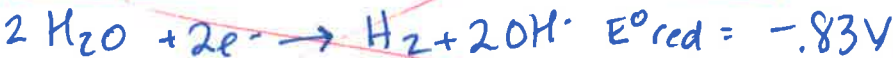
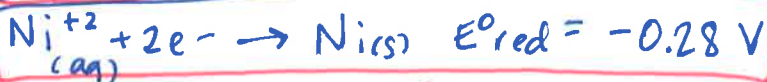
activity series

	E°_{ox}
Rb	+2.92 V
Zn	+0.76 V
Fe	+0.44 V
Ni	+0.28 V
Sn	+0.14 V
H (in acid)	0 V
Cu	-0.34 V

3. If an aqueous solution of nickel (II) fluoride is electrolyzed, what will be the half rxn at each electrode?

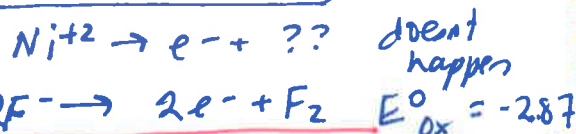
The solution contains $\text{Ni}^{+2}_{(aq)}$, $\text{F}^{-1}_{(aq)}$, and $\text{H}_2\text{O}_{(l)}$

Cathode (reduction)



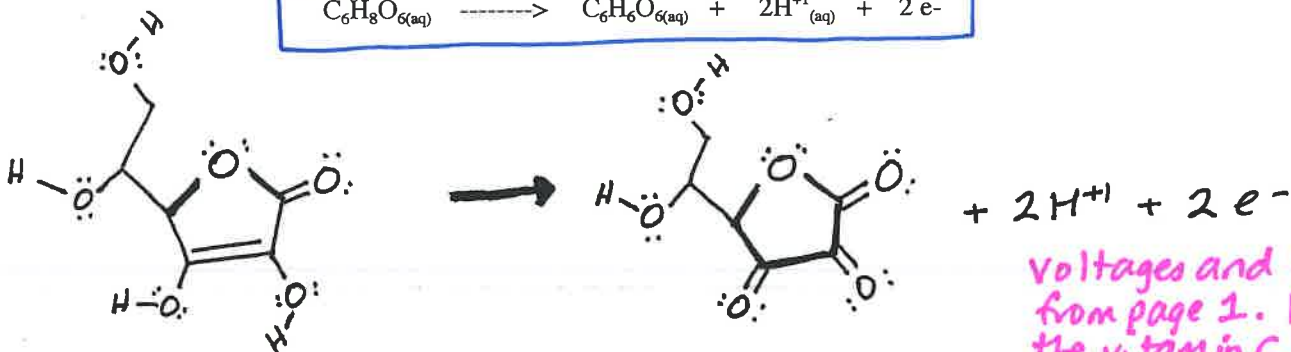
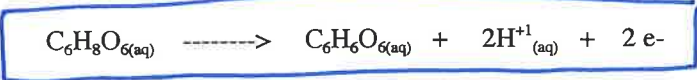
Since $-0.28 \text{ V} > -0.83 \text{ V}$

anode (oxidation)



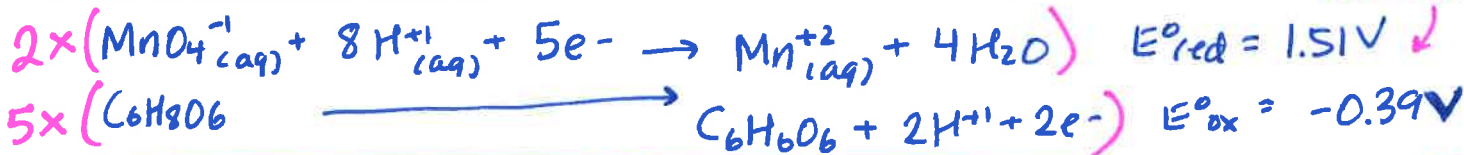
Since $-1.23 \text{ V} > -2.87 \text{ V}$

4. Vitamin C (ascorbic acid; $C_6H_8O_6$) can be oxidized to $C_6H_6O_6$ (dehydroascorbic acid) according to the following half reaction:



Voltages and rxns from page 1. I reversed the vitamin C voltage so it is oxidized.

a. Write the balanced redox reaction that occurs, if ascorbic acid is oxidized by permanganate ion, in acidic solution.



b. What is the value of "n" for this rxn? 10

c. Calculate ΔG° for this reaction.

$$E^{\circ}_{rxn} = E^{\circ}_{red} + E^{\circ}_{ox} = 1.51V + (-0.39V) = 1.12V$$

$$\Delta G^{\circ} = -1080 \text{ kJ/mole}$$

$$\Delta G^{\circ} = -nFE^{\circ} = -(10 \text{ mole } e^-) \left(\frac{96485 \text{ C}}{\text{mole } e^-} \right) (1.12 \frac{\text{J}}{\text{C}}) = -1080632 \text{ J/mole}$$

d. Calculate K_{eq} for this reaction at 25°C.

$$K_{eq} = e^{-\Delta G^{\circ}/RT} = e^{\frac{+1080632 \text{ J/mole}}{(8.314 \text{ J/mol}\cdot\text{K})(298\text{K})}} = e^{436.17} = 267 \times 10^{189}$$

e. Write an expression for K_c in terms of the appropriate concentrations.

$$K_c = \frac{[C_6H_6O_6]^5 [Mn^{2+}]^2}{[MnO_4^-]^2 [C_6H_8O_6]^5 [H^+]^6}$$

A tablet/pill is tested for its vitamin C content. A 0.60 gram pill is crushed and dissolved into a sulfuric acid solution. This vitamin C/acid solution is now titrated with 0.045 M $KMnO_4$ solution. 20.8 mL of $KMnO_4$ were required to reach the end point.

f. Determine the oxidizing and reducing agents in the titration.

MnO_4^- = oxidizing agent (MnO_4^- is reduced)

Vitamin C ($C_6H_8O_6$) = reducing agent (it gets oxidized).

g. Calculate the number of moles of vitamin C that must have been dissolved in the solution.

$$(0.0208 \text{ L}) \left(\frac{0.045 \text{ mole } KMnO_4}{\text{L}} \right) = 0.000936 \text{ moles } KMnO_4$$

$$(0.000936 \text{ mole } KMnO_4) \left(\frac{5 \text{ mole } C_6H_8O_6}{2 \text{ mole } KMnO_4} \right) = 0.00234 \text{ moles} \rightarrow 0.0023 \text{ moles } C_6H_8O_6$$

h. Calculate the percent Vitamin (by mass) in the pill.

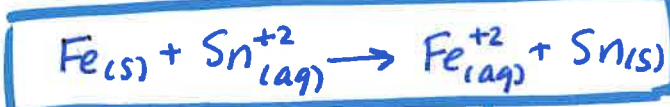
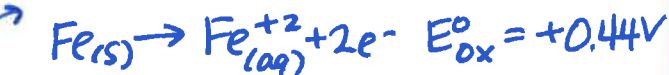
$$(0.00234 \text{ moles } C_6H_8O_6) \left(\frac{176.1256 \text{ g}}{\text{mole}} \right) = 0.41213 \text{ g}$$

$$\frac{0.41213 \text{ g}}{0.60 \text{ g}} \times 100 = 69\% \text{ Vit.C}$$

6. If you were going to make a [not very good] battery using the iron and tin half reactions on page 1,

a. What would be the overall reaction and cell voltage?

From p. 1:



$$E^\circ_{\text{rxn}} = 0.30\text{V}$$

we need to reverse one of these since both are reduction. since it is for a battery (or "galvanic cell" or "voltaic cell") it must have a positive E°_{rxn} so we'll ~~reverse~~ reverse the Fe rxn.

b. Calculate K_{eq} at 25°C (Answer: 1.4×10^{10})

$$\Delta G^\circ = -nFE^\circ = -\left(\frac{2\text{mole } e^-}{\text{mole rxn}}\right) \left(\frac{96485\text{C}}{\text{mole } e^-}\right) \left(\frac{0.30\text{J}}{\text{C}}\right) = -57891\text{ J/mole}$$

$$K_{\text{eq}} = e^{-\Delta G^\circ/RT} = e^{\frac{+57891\text{ J/mole}}{(8.314\text{ J/mol}\cdot\text{K})(298\text{ K})}} = e^{23.366} = 1.405 \times 10^{10}$$

$$K_{\text{eq}} = 1.4 \times 10^{10}$$

c. What substances would you put in all places in the battery?

d. Sketch the battery, and show the direction of flow of e^- in the external circuit, and the direction of ion flow in the salt bridge. Also label the anode and cathode. *c/d are on next page*

e. What would be the cell voltage at 25°C, if the concentration of Fe^{2+} is 0.010 M, and the concentration of Sn^{2+} is 2.0 Molar? How does E compare to E° , and how does this relate to the value of Q ?

$$Q = \frac{[\text{Fe}^{2+}_{(aq)}]}{[\text{Sn}^{2+}_{(aq)}]} = \frac{0.010}{2.0} = 0.0050$$

Since $Q < 1$
 $E > E^\circ$

$$E = E^\circ - \frac{RT}{nF} \ln Q = 0.30\text{V} - \frac{(8.314)(298)}{(2)(96485)} \ln(0.0050)$$

$$= 0.30\text{V} - (-0.068\text{V}) = 0.37\text{V}$$

f. What would be the cell voltage at 25°C, if the concentration of Sn^{2+} is 0.010 M, and the concentration of Fe^{2+} is 2.0 Molar? How does E compare to E° , and how does this relate to the value of Q ?

$$Q = \frac{(2.0)}{(0.010)} = 200 \quad E = E^\circ - \frac{RT}{nF} \ln Q = 0.30\text{V} - \frac{(8.314)(298)}{2(96485)} \ln(200)$$

$$E = 0.30\text{V} - 0.068\text{V} = 0.23\text{V}$$

Since $Q > 1$, $E < E^\circ$

g. If the concentration of Fe^{2+} is 5.0 M, at what concentration of Sn^{2+} would the cell emf be equal to zero? Explain.

The cell voltage will be zero when the rxn is at eqm. In part (b) we found that $K_{\text{eq}} = 1.4 \times 10^{10}$... it will be at EQU when $Q = K = 1.4 \times 10^{10}$

$$Q = \frac{[\text{Fe}^{2+}]}{[\text{Sn}^{2+}]} \quad 1.4 \times 10^{10} = \frac{(5.0\text{M})}{[\text{Sn}^{2+}]}$$

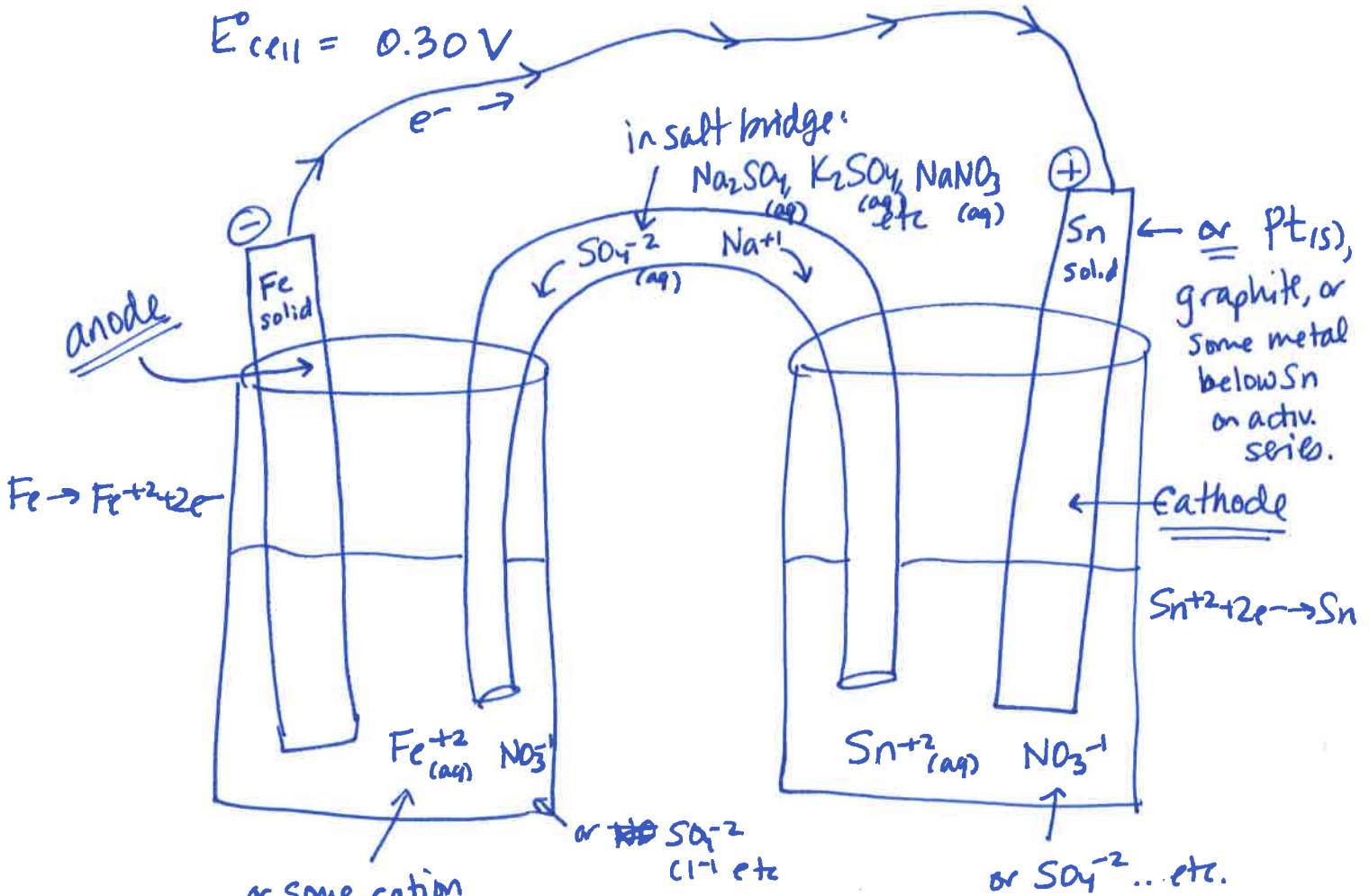
$$[\text{Sn}^{2+}] = 3.6 \times 10^{-10}\text{ M}$$

(or do $E = E^\circ - \frac{RT}{nF} \ln Q$ and solve for Q when $E = 0$)

#6 c/d



$$E^{\circ}_{\text{cell}} = 0.30 \text{ V}$$



or some cation
that won't react w/ solid Fe,
for example Na^{+} , K^{+} , ...

so use $\text{Fe}(\text{NO}_3)_2$
 NaNO_3
 K_2SO_4
 FeSO_4
etc

could use $\text{SnSO}_4(aq)$
 $\text{Sn}(\text{NO}_3)_2(aq)$
 $\text{SnCl}_2(aq)$
etc.

7. Concentration Cells:

a. Draw a picture of a concentration cell involving Co^{2+} . (What substances are present?)

b. Write a balanced equation for the reaction that occurs.



c. Answer these questions on your picture:

Which way will the e^- flow in the external circuit?

Which directions do anions/cations flow through the salt bridge?

Which side is the anode/cathode?

How will the solute concentrations change over time on each side?

How do the electrode masses change over time on each side?

d. What is the value of E°_{cell} ?

e. If one of the solutions is 0.30 M, and one is 2.5 M, calculate the E_{cell} at 25°C

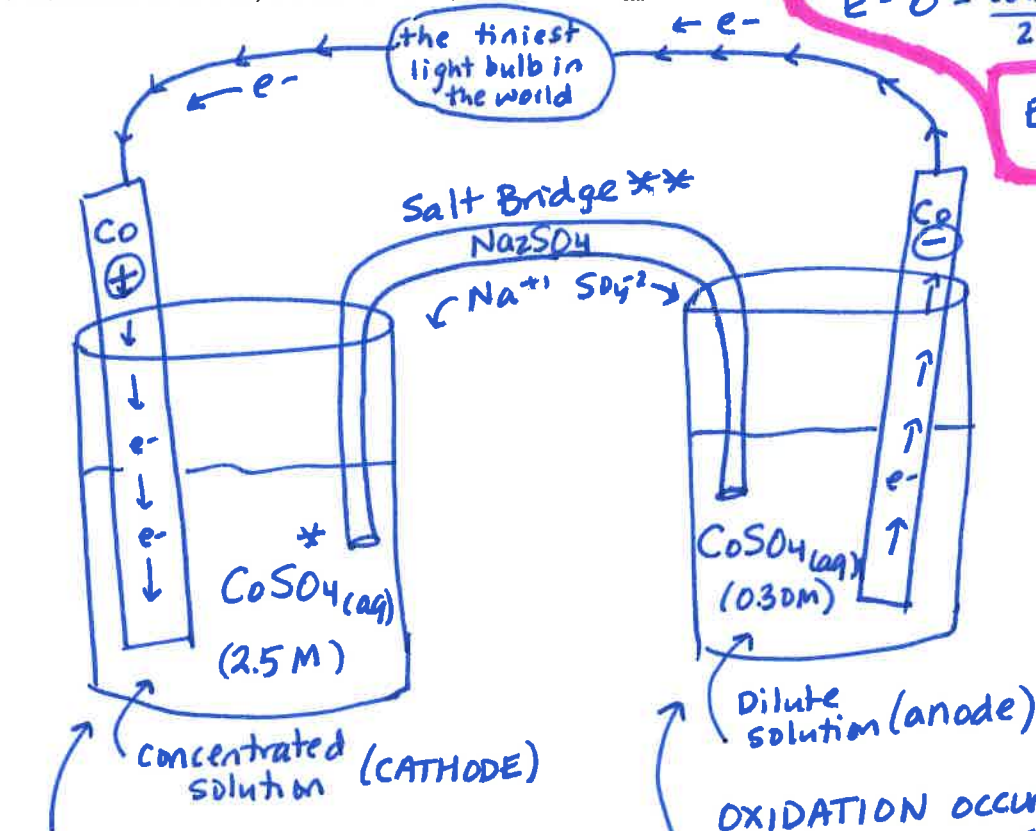
$$Q = \frac{[\text{Co}^{2+}(\text{aq})]_{\text{dilute}}}{[\text{Co}^{2+}(\text{aq})]_{\text{concentrated}}}$$

$$Q = \frac{0.30}{2.5} = 0.12$$

$$E = E^\circ - \frac{RT}{nF} \ln Q$$

$$E = 0 - \frac{(8.314)(298) \ln(0.12)}{2(96485)}$$

$$E = 0.027 \text{ V}$$



Reduction occurs at the CATHODE:



so on the cathode side, $[\text{Co}^{2+}(\text{aq})]$ decreases overtime and the solid Co electrode mass increases over time.

OXIDATION occurs at the ANODE



so on the anode side, $[\text{Co}^{2+}(\text{aq})]$ increases over time, and the mass of the solid Co electrode decreases over time.

* instead of $\text{CoSO}_4(\text{aq})$, we could use $\text{Co}(\text{NO}_3)_2(\text{aq})$ or $\text{CoCl}_2(\text{aq}) \dots$
 ** instead of Na_2SO_4 in the salt bridge, we could do K_2SO_4 , Li_2SO_4 , NaN_3 , LiNO_3 , $\text{KNO}_3 \dots$

8. a. A solution of zinc iodide is electrolyzed. Zinc ion and iodide ion both react.
 If a current of 2.00 A is applied over 5.0 hours, what mass of zinc can form?



$$(5.0 \text{ hr}) \left(\frac{3600 \text{ s}}{1 \text{ hr}} \right) \left(\frac{2.00 \text{ C}}{\text{s}} \right) \left(\frac{1 \text{ mole } e^-}{96485 \text{ C}} \right) \left(\frac{1 \text{ mole Zn}}{2 \text{ mole } e^-} \right) \left(\frac{65.38 \text{ g}}{1 \text{ mole}} \right) = 12.197$$

8a 12 grams Zn

b. Calculate E°_{cell} . $\text{Zn}^{2+}_{(aq)}$ and $\text{I}^-_{(aq)}$ both react acc to part (a)



I reversed the I_2/I^- rxn from page 1 since I^- needs to be a reactant.



c. What is the minimum energy required for part (a)?

$$(5.0 \text{ hr}) \left(\frac{3600 \text{ s}}{\text{hr}} \right) \left(\frac{2.00 \text{ C}}{\text{s}} \right) \left(\frac{1.30 \text{ J}}{\text{C}} \right) \left(\frac{1 \text{ kJ}}{1000 \text{ J}} \right) = 46.8 \rightarrow \text{47 kJ}$$

OR

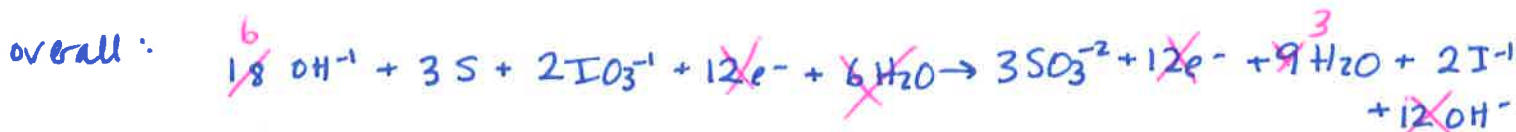
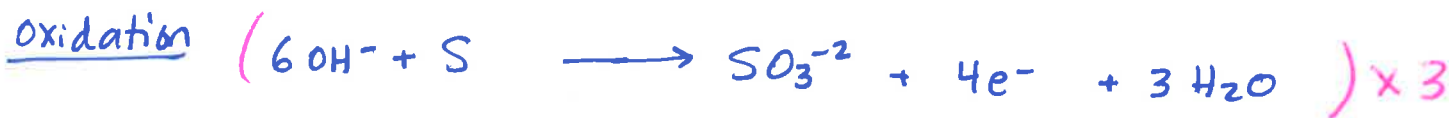
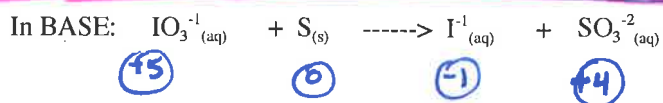
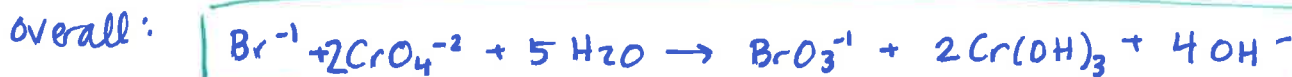
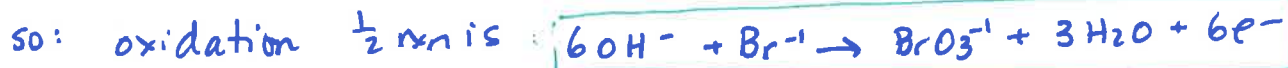
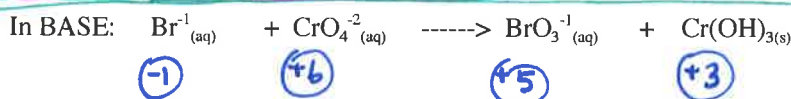
$$(12.197 \text{ g Zn}) \left(\frac{1 \text{ mole}}{65.38 \text{ g}} \right) \left(\frac{2 \text{ mole } e^-}{1 \text{ mole Zn}} \right) \left(\frac{96485 \text{ C}}{\text{mole } e^-} \right) \left(\frac{1.30 \text{ J}}{\text{C}} \right) \left(\frac{1 \text{ kJ}}{1000 \text{ J}} \right) = \text{47 kJ}$$

↑
from part (a)

9. Balance these 3 redox reactions. Show the balanced oxidation and reduction half reactions too.



overall



(#10 is at the end of this key - on last page)

11. Each reaction below has a positive E°_{rxn} .

Therefore each reaction could be used to make a "battery", also known as a Voltaic cell or a galvanic cell.

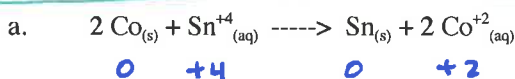
For each reaction, determine the value of "n" for the reaction, and determine a substance that could be used for the electrode at the anode, and a substance that could be used for the electrode at the cathode.

(In several cases, more than one substance could work. You can just list one substance though.)

The electrode substances have to be solid phase and have to conduct electricity.

The only substances that conduct as solids are metals and graphite (carbon)

If there is a metal on the reactant side, you need to use that metal as the electrode. (It would be the anode, since the metal would be oxidized in the rxn; metals will lose e^- when they react.)



$n=4$

Anode: must use Solid Cobalt (Co)

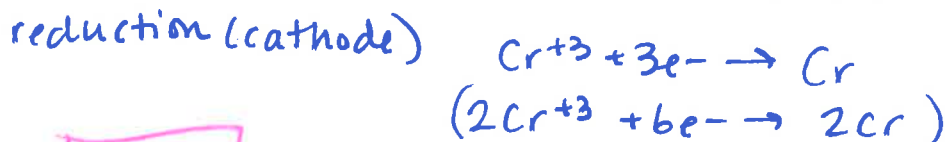
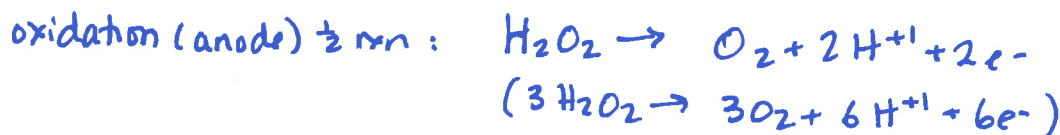
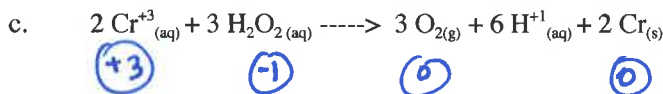
Cathode: could use solid Tin (Sn) or graphite or Pt
 (or any metal that is below Sn on the activity series)



$n=2$

anode: Pt or graphite

cathode: Fe or Pt or graphite

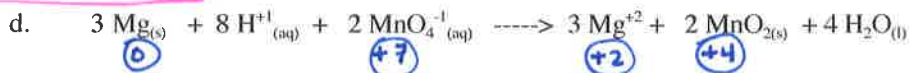


$n=6$

anode: Pt or graphite

cathode: Cr or Pt or graphite

11, contid

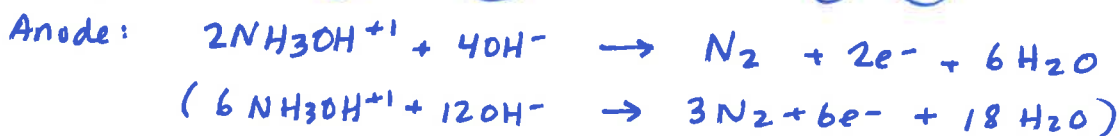
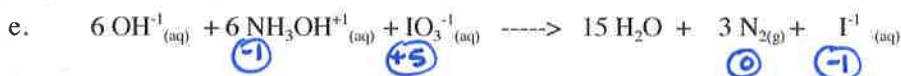


$n = 6$

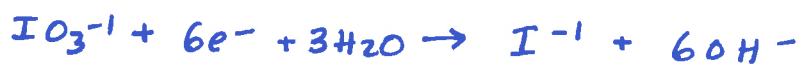


Anode: Mg (solid)

Cathode: Pt or graphite



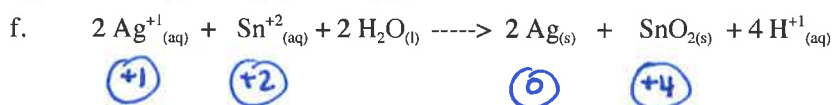
Cathode:



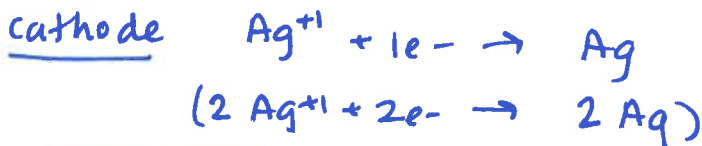
$n = 6$

Anode: Pt or graphite

Cathode: Pt or graphite

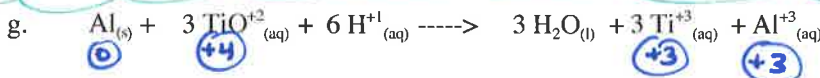


Anode: graphite or Pt



Cathode Ag solid or graphite or Pt

$n = 2$



$n = 3$



Anode: Al (solid)

Cathode: graphite or Pt

12. a. In reaction 11c, above, what mass of solid Chromium ^w could plate out, if the reaction produces a current of 0.044 amps over a time period of 3.0 minutes?



$$(3.0 \text{ min}) \left(\frac{60 \text{ s}}{1 \text{ min}} \right) \left(\frac{.044 \text{ C}}{\text{s}} \right) \left(\frac{1 \text{ mole } e^-}{96485 \text{ C}} \right) \left(\frac{1 \text{ mole } Cr}{3 \text{ mole } e^-} \right) \left(\frac{51.996 \text{ g}}{1 \text{ mole}} \right)$$

$$\uparrow$$

$$\text{or } \left(\frac{2 \text{ mole } Cr}{6 \text{ mole } e^-} \right)$$

$$= .001423 \rightarrow \boxed{0.0014 \text{ grams } Cr_{(s)} \text{ would form}}$$

b. In reaction 11e, above, how long would it take (in minutes) to produce 5.0 mL of nitrogen gas at 25°C and 0.91 atm, if the reaction is producing a current of 0.022 amps?

① First use ideal gas law to find moles $N_2(g)$

$$PV = nRT$$

$$n_{N_2} = \frac{PV}{RT} = \frac{(.91 \text{ atm})(0.0050 \text{ L})}{(.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(298 \text{ K})} = .00018597 \text{ moles } N_2$$

② use Rxn (stoichiometry) to convert moles N_2 to moles e^-



$$(.00018597 \text{ moles } N_2) \left(\frac{6 \text{ moles } e^-}{3 \text{ moles } N_2} \right) = .0003719 \text{ moles } e^-$$

③ Now use current of 0.022 A or .022 C/s, and the Faraday

$$(.0003719 \text{ moles } e^-) \left(\frac{96485 \text{ C}}{1 \text{ mole } e^-} \right) \left(\frac{1 \text{ second}}{.022 \text{ Coulombs}} \right) = 1631 \text{ seconds}$$

$$\text{or } \boxed{27 \text{ minutes}}$$

Reactants \rightarrow Products

10. Suppose that a reaction has an E°_{rxn} of 0.10.
As the reaction proceeds in the forward direction,

Does Q_{rxn} increase or decrease? Increase

$$Q = \frac{\text{products}}{\text{reactants}}$$

Does E_{rxn} increase or decrease? decrease
(becomes more negative)

$$E = E^{\circ} - \frac{RT}{nF} \ln Q$$

if $Q > 1$, then $\ln Q$ is \oplus so $E < E^{\circ}$

the larger Q becomes, the larger the " $\frac{RT}{nF} \ln Q$ " term becomes, the larger # you subtract from E° ;
the more negative the voltage becomes.

5. To protect nails from rusting, they are often "galvanized": They are coated in a more reactive metal, so that the more reactive metal gets oxidized instead of the iron. (The more reactive metal can be called the "sacrificial anode.") Based on the voltages on page 1, which metal should be used to coat Fe nails: Tin or Zinc?

From page 1 (we'll reverse the reduction $\frac{1}{2}$ rxns given):



Zinc! Zn(s) is oxidized more readily than Fe(s)

(Zn(s) has a more positive E°_{ox} than Fe(s))

so "galvanized" nails are coated in Zinc.

That way, if the nails come in contact with an oxidizing agent (for ex, $\text{H}^{+}_{\text{(aq)}}$ or $\text{O}_{2(\text{g})}$), then the Zinc coating will get oxidized instead of the iron that's on the inside of the nail.