Chem 103 Lecture 18 PS 158 7715

Admin:
Repeat: Test #2 scheduled as indicated in syllabus: week 8, Fri Aug 10 NOT Aug 13.

Last time:
1) balancing redox reactions
2) electrochemical potential

Today:
0) Redox potential
1) Redox titration
2) electrochemical cells:Galvanic Cells
3) Group Quiz (extra credit)

Lecture:

0) **practice:** Consider
$E°_{AuCl4-/Au} = 1.00V$ and $E°_{O2/H2O}=1.23V$
what are the half reactions?

Balance them

What is the spontaneous redox reaction between these 2 redox couples?

What is the $\Delta E°$=? $\Delta G°$=?

2) **Redox titrations**.

Consider titration of Fe^{2+} with CrO_4^{2-}

The half reactions are: (from table)

$Cr_2O_7^{2-}+14H_3O^++ 6e^- \rightarrow 2 Cr^{3+}+21H_2O$ 1.33V

$Fe^{3+} +e^- \rightarrow Fe^{2+}$ -0.771V

So we see that indeed, Fe^{2+} can be oxidized by $Cr_2O_7^{2-}$.
The balanced equation:
$Cr_2O_7^{2-}+14H_3O^++6Fe^{2+} \rightarrow 2Cr^{3+}+6 Fe^{3+}+21H_2O$

Remember that at equivalence point, the

eq =# eq

#mol $Cr_2O_7^{2-}$ x 6 = #mol Fe^{2+}

for solution titrations:
$6M_{Cr2O7}V_{Cr2O7} = M_{Fe}V_{Fe}$

3) How much work can the galvanic cell do? (assuming standard conditions)

$W_{max} = \Delta G°_{rxn} = -nFE°_{rxn}$

Where F = Faraday's constant
= 96,500 coulombs/ mole electrons
$=9.65x10^4$C/mol

n = # electrons transferred.

Note that 1 Joule = 1 coul-volt

4) **Problem**:
25.0 mLs of $FeSO_4$ solution requires 20.0 mLs of 0.150 M potassium dichromate, $K_2 Cr_2O_7$. What is the concentration of the $FeSO_4$ solution?

$6 \times M_{Cr}V_{Cr} = M_{Fe}V_{Fe} \Rightarrow M_{Fe} = 6M_1V_1/V_2$

=6(.150)(20.0)/25.0 = .900(4/5)=0.720M Fe^{2+}

8) What is the E° and ΔG° for the above reaction?

E° = 1.33-(-.771) = 2.10 V

ΔG° = - (6)(96500C/mol)(2.10V)
= - 1.22 x 10^6 CV/mol (=J/mol)
= - 1.22x10^3 kJ/mol

5) Electrochemical Cells:
2 types:
a) **Voltaic** Cells or Galvanic Cells (spontaneous)
b) **Electrolytic** cells (nonspontaneous)

A) Voltaic Cells

Write a possible redox reaction based on the following half rxns.

1	$Cu^{+2}(aq) + 2e^{-} \rightarrow Cu(s)$	+0.34V
2	$Zn^{2+}(aq)$) $+2e^{-} \rightarrow Zn(s)$	-0.76

One has to be reduced, the other oxidized, so Reverse #2 and add to #1 :

(1) –(2): $Cu^{2+} + Zn \rightarrow Cu + Zn^{2+}$ .34+.76=1.10V

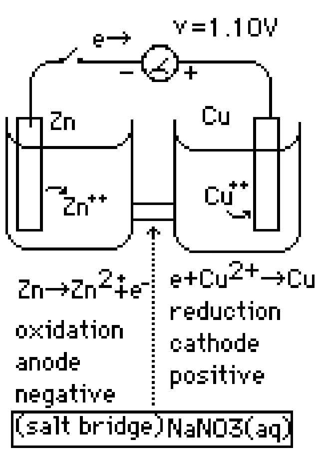
1.10 V is the redox potential?

$E°_{Cu2+/Cu} - E°_{Zn2+/Zn} = .34+.76=1.10V$

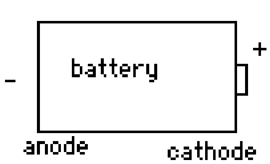
voltaic cells:
consider $Zn(s)+Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

Zn supplies e and Cu(II) ion accepts electrons.
What if we can force the e to go thru wire?

Consider the daniell cell and explain it.



Analogous to battery



Group Quiz

Consider the following standard reduction potentials:
$E°_{SO2/S}$ =0.45 V and $E°_{NO3-/NO}$=0.955
a) Write each reduction half reaction and balance each one (acidic conditions).
b) Write the balanced spontaneous redox reaction involving these two(2) redox couples.
c) Determine the reductant and oxidant.
d) Determine the reaction's standard redox potential for the spontaneous.
e) Determine the maximum work this reaction can do.

Key:
a) (1) $SO_2 + 4e^- + 4H^+ \rightarrow S(s) + 2H_2O$ .45V
(2) $NO_3^- + 3e^- + 4H^+ \rightarrow NO + 2H_2O$ .955V

b) reverse (1) and multiply it by 3
multiply (2) by 4 and add to the reversed (1)

$4NO_3^- + 12e^- + 16H^+ + 3\ S(s) + 6H_2O \rightarrow 4NO + 8H_2O + 3\ SO_2 + 12e^- + 12H^+$

cancelling redundant species:

$4NO_3^- + 4\ H^+ + 3\ S(s) \rightarrow 4NO + 2H_2O + 3\ SO_2$

c) oxidant = NO_3^- ; reductant = S(s)

d) E° = (.955-.45)= +.505

e) $W_{max} = \Delta G° = -nFE°$
=(-12)(96500C/mol)(.505V)
= 585,000 J = 585 kJ/mol