Chem 103 Lecture 8a PS 158 77-15yppp-ll

Admin:
Reminder: Test #2 on Friday this week.
Review Session **Wednesday** 10-11am PS 607

Last time:
1) galvanic cells
2) balancing practice
3) adding 2 half rxns to get another half rxn
4) work, ΔG and K
5) started Nernst Equation.

Today:
1) Nernst equation
2) concentration cells
3) batteries
4) corrosion

1) the **Nernst Equation**

Nonstandard cells : What is the potential for a nonstandard cell?

Recall :

$\Delta G = \Delta G° + RT\ln Q$

and $\Delta G = - nFE$; $\Delta G° = -nFE°$

$\Delta G = \Delta G° + RT \ln Q$
$-nFE = -nFE° + RT \ln Q$

$E = E° - (RT/nF) \ln Q$
(*Nernst Equation*)

If assume T = 25°C, R = 8.314J/molK, F =Faraday's const, and use Log_{10} instead of ln:

$E = E° - 0.05916/n \text{ V} \log Q$

So if we have $[Cu^{2+}] = 0.010$ M & $[Zn^{2+}] = 1.00$M

Then $E = 1.10 - 0.06 \log\{[Zn^{2+}]/[Cu^{2+}]\} =$
$= 1.10 - .06(\log(1.00/.01)) = 0.98$ V

note that as rxn proceeds, $[Zn^{2+}]$ increases and E_{cell} decreases until it reaches 0. (example: that's how a battery runs down, potential (volts)

drops it gets used up)

Flow of electrons measured as current, (in Amperes, A). It can do work – such running a motor…

Practice:

Say that a Cu-Zn cell has a potential of 1.10 V. A Cu-Pb cell has a potential of 0.46 V. What is the potential of a Zn-Pb cell? Prove it by deriving the E°.

Concentration cells

Consider a galvanic cell with Cu half cells on both sides. Consider the left side being 1.00 M and the right 0.001M Cu^{2+}.

Draw the cell.

Draw the line diagram.

High conc side (i): $Cu^{2+} + 2e^- \rightarrow Cu(s)$
reduction, (+), cathode
Low conc side (ii): $Cu \rightarrow Cu^{2+}+2e^-$
Oxidation (-) ,anode

What is the potential?
$E_{cell} = E_{cath}-E_{anode} = E°-.06/2 \log(1/[Cu^{2+}])_i$
$- \quad E°-.06/2 \log(1/[Cu^{2+}])_{ii}$
$= 0 +.03\log\{[Cu^{2+}]_i/[Cu^{2+}]_{ii}\}=.03\log(.1/.001)$
$= +0.06$ V$=+60$ mV.

Note how concentration of one half cell can be monitored by means of measuring the voltage. The other half cell can be "fixed" and be the reference. Potentiometry.

How useful is it?

A potentiometer like pH meter.

Applications of galvanic cells for doing work:

Batteries:

Useful for: flash lights, portable computers, cell phones, ipods, ..electric cars!

Primary batteries: *can't be reversed

a) dry cell battery:

anode (ox): $Zn \rightarrow Zn^{2+} + 2e^-$

cathode (red): $2NH^+_4 + 2e^- \rightarrow 2NH_3(g) + H_2(g)$

H_2 removal: $2MnO_2 + H_2(g) \rightarrow Mn_2O_3(s) + H_2O(l)$

carbon rod electrode is used.

NH_3: $NH_3 + Zn^{2+} + 2Cl^- \rightarrow Zn(NH_3)_2Cl_2(s)$

Disadvantage: poor shelf life, can't draw current too fast.

E vs time, potential drops…

b) Hg battery: (used in calculators, cameras, watches, heart pacemakers,hearing aids …

anode: $Zn(s) + 2OH^- \rightarrow ZnO(s) + H_2O(l) + 2e^-$

cathode: $HgO(s) + H_2O(l) + 2e^- \rightarrow Hg(l) + 2OH^-$

net:

write the net rxn:

write the Nernst Equation

what is the E_{cell}=?

What is the characteristic E vs t curve?

Advantages: constant E

Disadvantages: Hg toxic!

Secondary Batteries: rechargeable batteries

Lead batteries:
Used in cars

Anode: $Pb(s) + SO_4^{2-} \rightarrow PbSO_4(s) + 2e^-$

Cathode: $PbO_2(s) + 4H_3O^+ + SO_4^{2-} + 2e^- \rightarrow PbSO_4(s) + 6\ H_2O(l)$

Net:

$Pb(s) + PbO_2(s) + 2\ H_2SO_4\ (aq) \rightarrow 2\ PbSO_4(s) + 2H_2O(l)$ $\quad E° = 2.041$ V

How do we get 12 V out of it?

What gets depleted as the battery runs down?

Advantage: simple, large current capability;
reversible!(can be recharged)

How to determine the strength of the battery?

By a hygrometer? (density)

Disadvantage: heavy! Pb toxic.

If you jump or quickly recharge, you may cause
electrolysis of water : O_2 and H_2 gases.
What happens if you have a spark???

Nickel-Cadmium batteries:

Anode: $Cd(s) + OH^- \rightarrow Cd(OH)_2 + 2e^-$

Cathode: $NiO(OH)(s) + H_2O + e^- \rightarrow Ni(OH)_2(s) + OH^-$