Chem 1062

Note Title 7/31/2008

* Extra Credit Opportunity
  * Give a 5-minute (±1minute) presentation on applied electrochemistry (5 points on Monday)
    - ✓ ~~1) Fuel cells~~
    - ✓ ~~2) Sacrificial Anodes / Cathodic Protection~~
    - ✓ ~~3) Rechargeable vs nonrechargeable batteries~~
    - ✓ ~~4) Electroplating~~

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* Extra Credit Opportunity
  * Give a 5-minute (±1minute) presentation on applied nuclear chemistry (5 points on Tuesday)
    - ✓ (1) Medical Applications of Radioisotopes
    - ✓ (2) Nuclear Reactors
    - ✓ (3) Radiation Counters
    - ✓ (4) Positron emission tomography

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ox: $Cu(s) \rightarrow Cu^{2+}_{(aq)} + \cancel{2e^-}$ $\quad E^\circ_{ox} = -0.34V$

red: $\cancel{2e^-} + 2Fe^{3+}_{(aq)} \rightarrow 2Fe^{2+}_{(aq)}$ $\quad E^\circ_{red} = 0.77V$

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overall: $Cu(s) + 2Fe^{3+}_{(aq)} \rightarrow 2Fe^{2+}_{(aq)} + Cu^{2+}_{(aq)}$

$E^\circ_{cell} = 0.43V$

$Cu(s) \,|\, Cu^{2+}_{(aq)} \,||\, Fe^{3+}_{(aq)}, Fe^{2+}_{(aq)} \,|\, C_{(graphite)}$

Determining the Strength of Oxidizing Agents and Reducing Agents

~~Will Cu react with Zn^{2+}?~~

Will Zn react with Cu^{2+}?

Find E°_{cell} for: $Zn(s) \mid Zn^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$

ox: $Zn(s) \rightarrow Zn^{2+} + 2e^-$ $\quad E^\circ_{ox} = +0.76V$

red: $2e^- + Cu^{2+}(aq) \rightarrow Cu(s)$ $\quad E^\circ_{red} = +0.34V$

$Zn(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Zn^{2+}(aq)$ $\quad E^\circ_{cell} = 1.10V$

~~$Cu(s) + Zn^{2+}(aq) \rightarrow Zn(s) + Cu^{2+}(aq)$~~ $\quad E^\circ_{cell} = -1.10V$

Which of the following is the weakest reducing agent?

✓ $Ag(s)$ $\quad E^\circ_{ox} = -0.80V$

$Cd(s)$ $\quad E^\circ_{ox} = +0.40V$

20.71 Calculate the standard emf from ΔG°_f:

→ E°_{cell}

$$Mg_{(s)} + 2Ag^+_{(aq)} \rightarrow Mg^{2+}_{(aq)} + 2Ag_{(s)}$$

ΔG_f° (kJ/mol): 0 $\quad 2(77.12)$ $\quad -454.8$ $\quad 2(0)$

$$\Delta G^\circ = 1(-454.8) + 2(0) - [1(0) + 2(77.12)]$$

$$= -609.0 \text{ kJ}$$

$1 J = 1 C \cdot V$

$$\frac{\Delta G^\circ}{-nF} = \frac{+nFE^\circ_{cell}}{+nF} = \frac{+609.0 \times 10^3 \cancel{J}}{+(2\cancel{mol})(9.65\times10^4 C/\cancel{mol})}$$

$$= \boxed{3.15 \text{ V}}$$

Last week:

Ch. 19 $\quad \Delta G^\circ = -RT \ln K$

Ch. 20 $\quad \Delta G^\circ = -nFE^\circ_{cell}$

$$\frac{+RT\ln K}{+nF} = \frac{-nFE^\circ_{cell}}{-nF}$$

$1 J = 1 C \cdot V$

$1 \frac{J}{C} = 1 V$

$\ln x = 2.303 \log x$

@ 25°C

$$E^\circ_{cell} = \frac{RT\ln K}{nF} = \frac{(8.315 \frac{J}{mol \cdot \cancel{K}})(298\cancel{K})\ln K}{n(mol)(9.65\times10^4 C/\cancel{mol})}$$

$$\boxed{E^\circ_{cell} = \frac{0.0257 V}{n}\ln K}$$

text $\rightarrow$

$$\boxed{E^\circ_{cell} = \frac{0.0592 V}{n}\log K}$$

(20.75) Find K @ 25°C:

$2Fe^{3+}_{(aq)} + Cu_{(s)} \rightarrow 2Fe^{2+}_{(aq)} + Cu^{2+}_{(aq)}$

ox: $Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e^-$ $\quad E^\circ_{ox} = -0.34V$

red: $2e^- + 2Fe^{3+}_{(aq)} \rightarrow 2Fe^{2+}_{(aq)}$ $\quad E^\circ_{red} = +0.77V$

$E^\circ_{cell} = +0.43V$

spontaneous

product-favored

$$\frac{nE^\circ_{cell}}{0.0592} = \frac{0.0592}{n}\log K$$

$$\frac{(2)(0.43V)}{0.0592V} = \log K$$

$$14.53 = \log K$$

$$10^{14.53} = 10^{\log K}$$

$$K = 3 \times 10^{14}$$

From Ch 19: $\Delta G = \Delta G^\circ + RT\ln Q$

$\Delta G = -nFE_{cell}$

$\Delta G^\circ = -nFE^\circ_{cell}$

$$\frac{-nFE_{cell}}{-nF} = \frac{-nFE^\circ_{cell}}{-nF} + \frac{RT\ln Q}{-nF}$$

Find E_{cell} for nonstandard conditions (@25°C)

$$E_{cell} = E^\circ_{cell} - \frac{RT}{nF}\ln Q$$

$$E_{cell} = E^\circ_{cell} - \frac{0.0257}{n}\ln Q$$

Nernst Equation → $E_{cell} = E^\circ_{cell} - \frac{0.0592}{n}\log Q$

Find E_{cell} for:

$$Al(s) \mid Al^{3+}(2.0M) \parallel Cu^{2+}(0.010M) \mid Cu(s)$$

ox: $2Al(s) \rightarrow 2Al^{3+}(aq) + 6e^-$ $\quad E^\circ_{ox} = +1.66V$

red: $6e^- + 3Cu^{2+}(aq) \rightarrow 3Cu(s)$ $\quad E^\circ_{red} = +0.34V$

$2Al(s) + 3Cu^{2+}(aq) \rightarrow 3Cu(s) + 2Al^{3+}(aq)$ $\quad E^\circ_{cell} = 2.00V$

$$E_{cell} = E^\circ_{cell} - \frac{0.0592V}{n} \log \frac{[Al^{3+}]^2}{[Cu^{2+}]^3} \qquad Q = \frac{[Al^{3+}]^2}{[Cu^{2+}]^3}$$

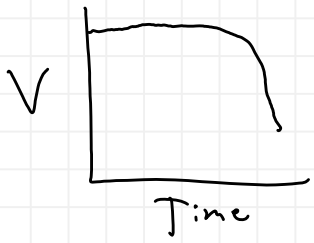
$$= 2.00V - \frac{0.0592V}{6} \log \frac{(2.0)^2}{(0.010)^3}$$

$$= 2.00V - \frac{0.0592V}{6} \log(4{,}000{,}000)$$

$$= 2.00V - \frac{0.0592V}{6}(6.602)$$

$$= 2.00V - 0.0651V$$

$$= \boxed{1.93V}$$



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* D2L Quizzes
* Lab Reports Returned (all except titration)

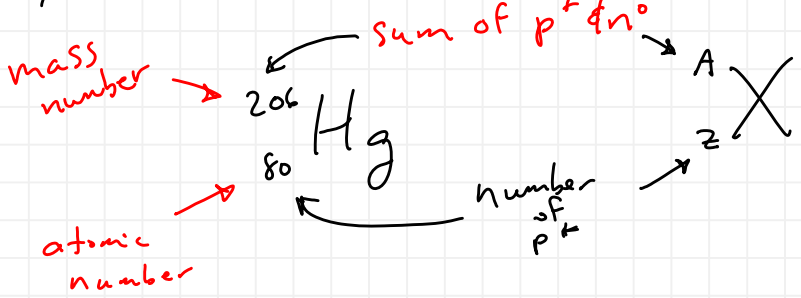
Applications of Electrochemistry

1) Batteries - Rechargeable vs Non rechargeable
2) Sacrificial Anodes / Cathodic Protection
3) Electroplating
4) Fuel Cells

Nuclear Chemistry

nuclide vs isotope

Nuclide Symbol



mercury-206

Hg-206

^{206}Hg

Types of Nuclear Reactions

1) Radioactive Decay - occurs when a nucleus spontaneously disintegrates, releasing radiation (α, β, γ rays)
2) Nuclear Bombardment (Transmutation) - occurs when a nucleus is bombarded or struck by another particle or nucleus

Types of particles

proton $\Rightarrow {}^{1}_{1}H$ or ${}^{1}_{1}p$

neutron $\Rightarrow {}^{1}_{0}n$

electron $\Rightarrow {}^{0}_{-1}e$ or ${}^{0}_{-1}\beta$

alpha $\Rightarrow {}^{4}_{2}He \equiv {}^{4}_{2}\alpha$

positron $\Rightarrow {}^{0}_{1}e \equiv {}^{0}_{1}\beta$

gamma photon $\Rightarrow {}^{0}_{0}\gamma$

Example:

$${}^{232}_{90}Th \rightarrow {}^{228}_{88}Ra + {}^{4}_{2}He$$

↑ alpha emitter

Beta emitter ↓

$${}^{32}_{15}P \rightarrow {}^{32}_{16}S + {}^{0}_{-1}e$$