Agenda for October 14, 2008

jumbo

Function: *noun*
Inflected Form(s): *plural* jumbos
Etymology: *Jumbo,* a huge elephant exhibited by P. T. Barnum
Date: 1883
: a very large specimen of its kind

laconic
Pronunciation: l&-'kä-nik
Function: *adjective*
Etymology: Latin *laconicus* Spartan, from Greek *lakOnikos;* from the Spartan reputation for terseness of speech
Date: 1589
: using or involving the use of a minimum of words : concise to the point of seeming rude or mysterious

cum laude With honor
magna cum laude with great honor
summa cum laude with highest honor
(sine laude) without honor

mea culpa, mea maxima culpa

Creme de la crème

Politeness in class – please do not talk while I am talking.

Writing Project
You will write a 5 - 7 page " concept paper" on a topic from class. Imagine you are writing for your colleagues. Footnotes are not required. This is not to be a research paper but a clear exposition of a chemistry topic **from this course**.
12 point font, 1" margins, double spaced

October 19: Deadline for emailing me your topic.

November 9: Give your paper to a colleague in 461/561 to critique.

November 15: Paper with marked up and signed draft given to colleague is due. I will grade on technical accuracy, grammar and spelling and return with comments. You must write on a word processor. You may rewrite and turn in again, along with earlier versions.

Half Life

The half life of a process is defined as the amount of time required for half of the initial quantity to react. In our example above, it is about 7 seconds.

The half life is important because it is directly related to the rate constant.

$t_{1/2} = \frac{\ln 2}{k}$, where "ln 2" means the natural logarithm of 2, which is 0.693.

If you know the half life, then you really know the rate constant.

Predicting the amount of time

If you know the half life, how to predict how long it will take for a certain fraction of the reactant to react?

For our example, let us calculate how long it will take for 90% of A to react.

A) 90% reacted means that $[A](t) = 0.1*[A]_0$.

We want to solve for t.

$[A](t) = [A]_0\, e^{-k*t}$

$0.1*[A]_0 = [A]_0\, e^{-k*t}$ (we'll substitute in the numerical value of k later)

$0.1 \quad = e^{-k*t}$

How to solve for t?

Take the natural logarithm of both sides of the equation:

$\ln(0.1) = \ln e^{-k*t}$;

$\ln(0.1) = -k*t$.

$t = -\ln(0.1)/k$

NB: The natural logarithm and the exponential function are inverse functions of each other.

Let's look at a C14 dating problem:

Assuming a half life of 5700 years, determine the age of a sample that has 67% of the original C14.

Next Topic Chapter 16 Redox Reactions

What is a Redox reaction?

Reduction-Oxidation.

Instead of just transferring atoms between molecules, electrons are transferred.

Reduction = gain of electrons. When you add an electron to a species, its charge goes down, becomes more negative.

"Leo the lion goes Gerr"

Loss of electron(s) = oxidation
Gain of electron(s) = reduction

$Fe(s) \rightarrow Fe^{+2} + 2\ e^{-}s$ oxidation
$Fe^{+2} \rightarrow Fe^{+3} + 1\ e^{-}$
NB: Fe^{+2} = Fe(II); Fe^{3+} = Fe(III)

$Fe(III) + e^{-} \rightarrow Fe(II)$

Consider rusting

$2Fe + 3O_2 \rightarrow 2Fe_2O_3$

O is now "oxide" and has a charge of -2 so Fe has a charge of +3. Iron's oxidation state has increase (from 0 to +3) so it has been oxidized.

Oxygen is the *oxidizing agent* and it was reduced in the process because its oxidation state went from 0 to -2.
Iron is the *reducing agent*.

How to assign Oxidation States
See rules on p. 583. You are responsible for knowing these rules.

How to balance Redox reactions:….very carefully….

1) Separate into half reactions (Reduction and Oxidation)
2) Balance the mass, adding H^+ as needed if in acid solution, OH^-s if in base.
3) Balance the charge using electrons.
4) Multiple entire half reactions by coefficients such that when the half reactions are added together, the electrons will cancel out.

$Al(s) + Ag^+(aq) \rightarrow Al^{3+} + Ag\ (s)$

Reduction:

$Ag^{+-}(aq) + 1e^- \rightarrow Ag\ (s)$

Oxidation

$Al(s) \rightarrow Al^{+3}\ (aq) + 3\ e^-$

We need to multiply the Reduction half reaction by 3:

$3\ Ag^{+-}(aq) + 3e^- \rightarrow 3Ag\ (s)$

+

$Al(s) \rightarrow Al^{+3}\ (aq) + 3\ e^-$

$Al(s) + 3\ Ag^{+-}(aq) \rightarrow Al^{+3} + 3Ag\ (s)$

Here is an important reaction: The MRE "Flameless ration heater":

$Mg\ (s) + 2H_2O \rightarrow Mg(OH)_2\ (aq) + H_2\ (g)$

In acidic solution

$I^-(aq) + Cr_2O_7^-\ (aq) \rightarrow Cr^{3+}(aq) + I_2\ (s)$

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

(First balanced mass and then charge).

What is oxidation of Cr? 2x -14 = -2; x = 6+
Reduction:
$Cr_2O_7^{2-}\ (aq) \rightarrow Cr^{3+}(aq)$
Use H^+ to balance mass after balancing the Cr.

$Cr_2O_7^{2-}\ (aq) \rightarrow 2Cr^{3+}(aq)$

$Cr_2O_7^{2-}\ (aq) + 14H^+ \rightarrow 2Cr^{3+}(aq) + 7\ H_2O$
Now balance charge

$Cr_2O_7^{2-}\ (aq) + 14H^+ + 6 \rightarrow 2Cr^{3+}(aq) + 7\ H_2O$

On the left
-2 + 14 = 12
On the right side: 2 x 3+ = 6

We need to add 6 electrons to the left side. This makes sense because each of the 2 Cr^{+3} is being reduced from +3 to 0.

Overall:
$6I^-\ (aq) \rightarrow 3I_2\ (s) + 6e^-$

$Cr_2O_7^{2-}$ (aq) + $14H^+$ + 6 e^- → $2Cr^{3+}$(aq) + 7 H_2O

Net reaction:

$Cr_2O_7^{2-}$ (aq) + $14H^+$ + $6I^-$ (aq) → $3I_2$ (s) + $2Cr^{3+}$(aq) + 7 H_2O

Check yourself by checking the mass balance.

Electrochemical Cell: Getting work directly from a chemical reaction (battery)

Voltage (measured in volts) is a measure of how much energy the electron has in a chemical reaction (or circuit—analogous to the high from which water is dropping).

Current (measured in Amps) is a measure of how many electrons are moving through a circuit (so analogous to how much water is falling.)

To obtain current from a redox reaction, the reduction and oxidation reactions must be separated and the electrons transferred via electrodes and wires.

e.g. Zn (s) in when container connected by a wire to an electrode immersed in a solution of Cu^{2+}(aq). The two solutions are also connected by a 'salt bridge" (a gel filled with an *electrolyte*).

Zn is the anode and gives up electrons to the electrode, is oxidized and dissolves:

$Zn\ (s) \rightarrow Zn^{2+}(aq) + 2\ e^{-}$

At the cathode, Cu^{+2} ions are reduced by the electrons coming from the anode:

$Cu^{2+} + 2\ e^{-} \rightarrow Cu\ (s)$.

As it is reduced, the copper "plates out" onto the electrode.

When at least one of the reagents is completely consumed, the battery is "dead".

Doing chemistry using an electronic current