Comments on Exam 2: with free response, average dropped to ~65%
Grades will be curved at end of semester
To estimate performance on 2nd exam, may regard 65% as high C.

- equivalent - mass of acid that will deliver 1 mol H⁺ ions
  mass of base that will deliver 1 mol OH⁻ ions

  1 mol H₃PO₄ = 3 equivalents H₃PO₄

  Normality = \( \frac{\text{equiv.}}{\text{l of soln}} \) similar to molarity

ex. calculate # equivalents in 0.300 g Cu(OH)₂(s).

  mm = 97.6 g/mol

  \[
  0.300 \text{ g Cu(OH)₂} \times \frac{1 \text{ mol Cu(OH)₂}}{97.6 \text{ g}} \times \frac{2 \text{ mol Cu(OH)₂}}{1 \text{ mol Cu(OH)₂}} = 6.185 \times 10^{-3} \text{ eq.}
  \]

  molar mass - what about equivalent mass?

  \[
  \frac{97.6}{2} = 48.8 \text{ g/eq.}
  \]

  in titration, at equivalence point

  \# equivalents acid = \# equivalents base
Calculate normality & molarity of an HCl solution. 547.2 mL of it will neutralize 0.631 g of Fe(OH)$_3$ ($\text{mm Fe(OH)}_3 = 106.9 \text{ g/mol}$)

$$0.631 \text{ g Fe(OH)}_3 \times \frac{1 \text{ mol}}{106.9 \text{ g}} \times \frac{3 \text{ eq Fe(OH)}_3}{1 \text{ mol Fe(OH)}_3} = 17.71 \text{ eq Fe(OH)}_3$$

(eq acid = eq base at equivalence point)

$$17.71 \text{ eq Fe(OH)}_3 = 17.71 \text{ eq HCl}$$

Normality:

$$\frac{17.71 \text{ eq}}{472 \text{ mL}} = 0.375 \text{ N}$$

$$\text{M} = \frac{0.375 \text{ N}}{1} = 0.375 \text{ M}$$

**BONUS:** when working with normality & equivalents, you don't need a balanced equation.
If both acid & base are in solution, calculations are very easy.

\[ \# \text{eq acid} = \# \text{eq base} \]

\[ N_{\text{acid}} \times V_{\text{acid}} = N_{\text{base}} \times V_{\text{base}} \]

(qual for quiz)

Ex. 30.00 mL of 0.07500M H₃PO₄ required 22.50 mL of Ca(OH)₂ solution for complete neutralization. Calculate molarity of Ca(OH)₂ solution.

\[ N_A \times V_A = N_B \times V_B \]

\[ (M \times 3) \times V_A = (M \times 2) \times V_B \]

\[ (0.07500 \text{M}) \times (0.03 \text{L}) = (M \times 2) \times (0.0225 \text{L}) \]

\[ M_B = 0.1500 \text{M Ca(OH)₂} \]

Next: REDOX, review Ch 4 notes
**REDox reactions**

- Electrons transferred from one species to another, can be either covalent or ionic, change in oxidation state.

**ex. 1**

\[ 2 \text{Cs}_2^0 + \text{S}_2^0 \rightarrow \text{Cs}_2\text{S}_2^{1-} \]

(review oxidation state rules)

**ex. 2**

\[ \text{S}_2^0 + \text{O}_2^0 \rightarrow 2\text{SO}_2^{4-} \]

- Oxidized - loses electrons
- Reduced - gains electrons
- Assign to prev. equations

**Oxidizing agent** - helps oxidize something (is itself reduced)

**Reducing agent** - helps reduce something (is itself oxidized) - Assign to prev. equations

\# Electrons lost by oxidized species must be equal to \# electrons gained by reduced species.
Balancing redox reactions can get tricky because we have to make sure the number of electrons gained equals the number of electrons lost, but we don’t write that info in the formula unit eq.

Also, a lot of the time in solutions it’s just easier to deal with the net ionic equation, in which case we need to be really careful. The equation may look correct, stoichiometrically, but it may not be balanced properly in terms of electron transfer.

ex. $\text{Cr}^{6+}(aq) + \text{Fe}^{2+}(aq) \rightarrow \text{Cr}^{3+}(aq) + \text{Fe}^{3+}(aq)$

not balanced. How can we balance redox reactions reliably?

Two methods in book:
- Change in oxidation number (electrons not written)
- Half-reaction method (electrons written)

[both useful, depending on what you are doing]
Dr. Heising’s handout on balancing redox reactions

Change in Oxidation Number Method

1. Write as much of unbalanced equation as possible.

2. Assign oxidation numbers & identify species undergoing changes in oxidation numbers.

3. Connect pairs of atoms (with changing oxidation numbers) on opposite sides of equation with brackets (or write a separate “mini” equation). Balance the oxidized/reduced species within the brackets.

4. Write the total # e- gained/lost next to bracket. Include subscripts, coefficients, etc, by multiplying.

5. Make the electrons gained/lost equal to one another by multiplying bracketed species with appropriate coefficients.

6. a. balance remaining atoms by inspection
   b. balance charge on each side of equation.

Half Reaction Method

1. Write as much of unbalanced equation as possible.

2. Assign oxidation numbers & construct unbalanced ‘half reactions’ for oxidized & reduced species.

3. Balance the half reaction atoms

4. Balance the half reaction charges (by adding electrons).

5. Balance the electron transfer (# e- gained/lost).

6. Add the half reactions together and cancel like terms. (electrons MUST cancel.)
ex. \[ \text{Cr}^{6+} + 3 \text{Fe}^{2+} \rightarrow \text{Cr}^{3+} + 3 \text{Fe}^{3+} \]

\[ 3e^- \text{ gained} \]

\[ 1e^- \text{ lost} \]

\[ \text{Cr}^{6+} + 3 \text{Fe}^{2+} \rightarrow \text{Cr}^{3+} + 3 \text{Fe}^{3+} \]

\[ 3e^- \text{ gained} \]

\[ 3(1e^-) \text{ lost} \]

\[ 12^+ \]

\[ 12^+ \]

has to balance both mass-wise & charge-wise.

balance charge on each side of eq.
\[
\begin{align*}
\text{Cr}_2\text{O}_7^{2-} + I^- + H^+ &\rightarrow \text{Cr}^{3+} + I_2 + H_2O \\
\text{at least balance pairs } &2(\text{e}^- \text{ lost}) \\
\begin{array}{c}
2(\text{e}^-) \\
\text{Cr}_2\text{O}_7^{2-} + 2I^- + H^+ \\
\text{lost} \\
\rightarrow \\
2\text{Cr}^{3+} + 2I_2 + H_2O \\
\text{gained} \\
2(\text{e}^-)
\end{array}
\end{align*}
\]

\[
\begin{align*}
\text{Cr}_2\text{O}_7^{2-} + 6I^- + H^+ &\rightarrow 2\text{Cr}^{3+} + 3I_2 + H_2O \\
\text{balance } H &6O \\
7O &\rightarrow 7H_2O \\
\text{need } 14H^+
\end{align*}
\]

\[
\begin{align*}
\text{Cr}_2\text{O}_7^{2-} + 6I^- + 14H^+ &\rightarrow 2\text{Cr}^{3+} + 3I_2 + 7H_2O(e) \\
\text{note: net } \text{charge on each side of eq. is same}
\end{align*}
\]

\[
\begin{align*}
2^+ + 6^- + 14^+ &\rightarrow 6^+ \\
2(3^+) + 0^- + 0^- &\rightarrow 6^+
\end{align*}
\]

oxidized (reducing agent) = I^- \\
reduced (oxidizing agent) = \text{Cr}^{6+}
ex. in acidic solution

\[
\begin{align*}
Mg_0 + NO_3^- & \rightarrow NH_4^+ + Mg^{2+} \\
& \text{net ionic eq.} \\
\text{\underline{8e}^{- gained}} \quad & \quad \text{Mg-oxidized} \\
\text{\underline{2e}^{- lost}} \quad & \quad \text{NO}_3^- - \text{reduced} \\
\end{align*}
\]

\[
4Mg_0 + NO_3^- \rightarrow NH_4^+ + 4Mg^{2+}
\]

problem - where do \( \text{H} \& \text{O} \) come from? need to mass balance.

see chart II-6 in acid solution -

- add \( 4H_2O \) for \( O \)
- add \( 4H^+ \) for \( H \)

\[
\begin{align*}
4Mg_0 + NO_3^- \rightarrow & \quad NH_4^+ + 4Mg^{2+} + 3H_2O(l) \\
& \quad + 10H^+ \\
O + 1^- + 10^+ & \rightarrow 9^+
\end{align*}
\]

see p. 417, textbook, for chart explaining adding H \& O under acidic/basic conditions.
Balance the following reaction in acidic solution:

\[
\text{MnO}_4^- + \text{SO}_3^- \rightarrow \text{Mn}^{2+} + \text{SO}_4^{2-}
\]