CHEM 150

ACIDS AND BASES

Definitions:

- Bronsted-Lowry
  - Acid – Donate H^+
  - Base – Accept H^+
  - Ex. $HCl + NaOH \rightarrow H_2O + NaCl$
    A B
- Lewis
  - Acid – Accept a lone pair
  - Base – Donate a lone pair
  - Ex. $BF_3 + NH_3 \rightarrow BF_3THF$
    A B
- Arrhenius
  - Acid – Cause H^+ to form
  - Base – Cause OH^- to form
  - Ex. $CO_2 + H_2O \rightarrow HCO_3^- + H^+$
    A B
- Solvent system
  - Acid – Cause cation to form
  - Base – Cause anion to form
- Lux-Flood
  - Acid – Oxide acceptor
  - Base – Oxide donor
- Note:
  - Neutral metals tend to be neutral or basic (by Arrhenius def)
  - Metal ions tend to act as acids in the sense that they help increase H^+ concentration
  - Hard : more ionic character
  - Soft : more covalent character

Trends in strength:

- The stability of the Conj. Base will determine the strength of the acid
  - Bond dissociation energy is a bigger deciding factor
- Conj. Base
  - Increase size, increase stability
    - Better distribution of charge
  - Increase # of lone pairs, increase stability
    - Better distribution of charge
  - Increase EWG, increase stability
    - Electronegative elements pull e^- toward themselves, making H^+ easier to pull off
- In general, increase # of oxygens in acid, increase in acidity
  - Ex. $H_2SO_4 > H_2SO_3$
- Pauling's Rule: pKa ~ 9-7n where n=# unprotanated oxygen in neutral
  - Ex. $HClO_4$ (3 unprotanated O)
  - pKa ~ 9-(7)(3) = -12
- Lower pKa is the stronger acid