Announcements – 9/25/00

- Make sure you go over your exam -- see me with questions
- Problems for tomorrow's problem session?
- Quiz on Friday!

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Acid/Base Reactions

- Most common definition:

<u>Bronsted-Lowry</u>

Acid: *proton donor*

Base: *proton acceptor*

$$\underset{\text{Acid}}{HA} + \underset{\text{Base}}{BOH} \rightarrow \underset{\text{Salt}}{BA} + \underset{\text{Water}}{H_2O}$$

-an example of a *double displacement reaction (metathesis)*

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Bronsted-Lowry in Water: Arrhenius Acids and Bases

- In aqueous solution, we can use the Arrhenius definitions of acids and bases:

Acid: *substance yielding H^+ in water*
Base: *substance yielding OH^- in water*

$$HA \rightarrow H^+ + A^-$$

$$BOH \rightarrow B^+ + OH^-$$

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More B-L Acid/Base

- Bronsted-Lowry does not require reaction to be in aqueous solution:

$$HA + B \rightarrow A^- + BH^+$$

acid base conjugate base conjugate acid

Example:

$$HCl(g) + NH_3(g) \rightarrow NH_4Cl(s)$$

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Lewis Acids and Bases

- We can make this concept *broader* by defining acids and bases a bit differently:

Lewis Theory

Acid: electron pair *acceptor*
Base: electron pair *donor*

Fits with Bronsted-Lowry definition:

H^+ + $:O\text{-}H^-$ $\rightarrow$ $H:O - H$
e^- pair acceptor e^- pair donor

HCl + $:NH_3 \rightarrow Cl\text{-}H:NH_3$

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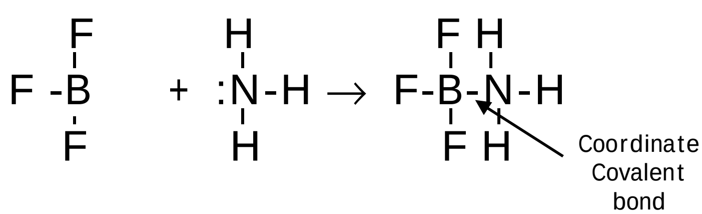
More Lewis Theory

- *MORE* reactions can be understood as acid/base reactions:

BF_3 + $NH_3 \rightarrow BF_3NH_3$

NOTE:
- No H^+



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Oxidation/Reduction Reactions

- Reactions in which electrons are transferred *individually* (not as pairs):

Oxidation: $M \rightarrow M^+ + e^-$ loses electron

Reduction: $X + e^- \rightarrow X^-$ gains electron

Net Reaction: $M + X \rightarrow MX$ *balanced # of e^- xfer'd*

Example: $2Na + Cl_2 \rightarrow 2NaCl$

-Na is *oxidized* to Na^+ (gives up an e^- to Cl)
-Cl is *reduced* to Cl^- (takes e^- from Na)

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Oxidation States

- We can assign a charge to atoms in a compound based on their propensity for donating or accepting valence electrons:

OXIDATION NUMBER

Examples: **NaCl - Na -> +1; Cl -> -1**
K_2O - K -> +1; O -> -2

- these assignments do not necessarily reflect the *actual distribution of electrons* in a compound (they are just a way to keep track of the valence electrons)

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Assigning Oxidation Numbers

- **Some rules**:
  - Sum of oxidation # = overal charge
  - **Group I, II**: always +1, +2 (ionic cmpds)
  - **Group VII**: F *always* -1, others often -1; could be + in cmpds with O or other Group VII elements
  - **Hydrogen**: always +1, except when with Group I, II (then it's -1; hydride)
  - **Oxygen:** always -2, except with F, hydrogen or Group I, II, or with itself (see exceptions in book)

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Oxidation #: examples

- **$SrBr_2$:** group II, group VII
  +2 -1
- **$Zn(OH)_4^{2-}$:** OH^- is -1, -2 charge for cmpd
  +2 -2 +1
- **$Cr_2O_7^{2-}$:** -2 = 7(-2) + 2(x)
  +6 -2 x = +6
- **MnO_4^-:** -1 = 4(-2) + x
  +7 -2 x =+7

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