1/23/15
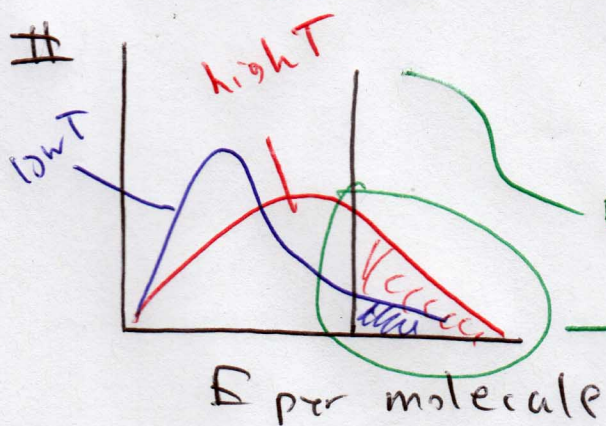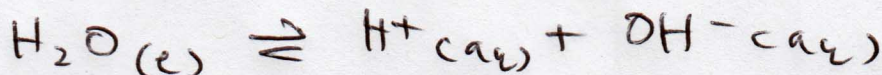
Arrhenius Acids + Bases

Acid - dissociates to produce H^+

Base - dissociates to produce OH^-

based on water

Auto-ionization of water

$$H_2O_{(\ell)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$$

"If equilibrium can happen, it will happen"

Bond dissociation energy

— fraction of molecules with enough energy to dissociate

E per molecule

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14} M^2$$ * Only true at one temperature

$$pH \equiv -\log_{10}[H^+]$$

In pure water at the right temperature, $[H^+] = [OH^-] = 1.0 \times 10^{-7} M$

$$pH = -\log_{10}(1.0 \times 10^{-7}) = 7 \longrightarrow neutral$$

neutral : equal concentration of $[H^+]$ and $[OH^-]$

not equal concentration of acid + base

pH = 7 is only neutral at one temperature since the value of K_w changes with temperature.

Brønsted-Lowry acids + bases

Acid - donates (H+) — proton

Base - accepts H^+

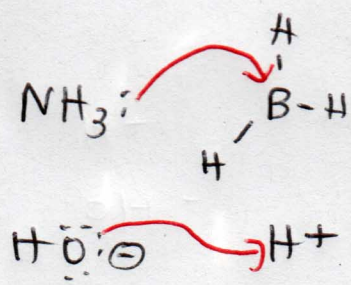NH_3 is a Brønsted-Lowry base but not an Arrhenius base (does not dissociate to produce OH^-)

All Brønsted-Lowry acids are Arrhenius acids and vice versa.
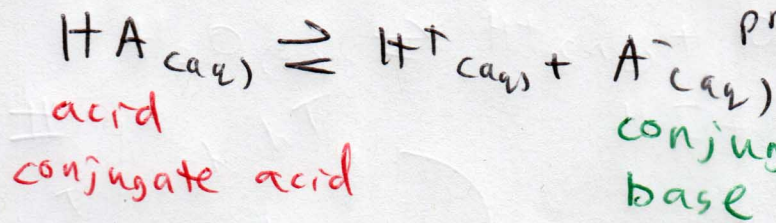
Lewis ○ acids + bases   ← electron

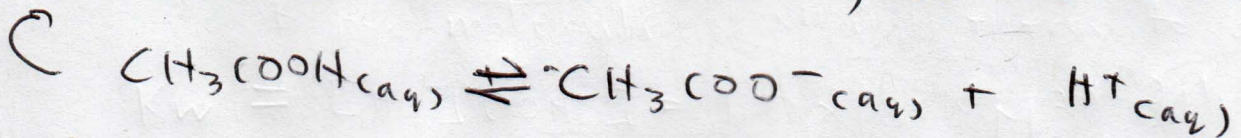Acid - ~~A~~ accepts e⁻ pair

Base - donates e⁻ pair

$NH_3 :$ ⤳ $B-H$ (with H above and H below)

$H\ddot{O} :^{\ominus}$ ⤳ H^+

K_a

Monoprotic – dissociates to produce only one H^+

$$HA_{(aq)} \rightleftharpoons H^+_{(aq)} + A^-_{(aq)}$$

acid

conjugate acid

conjugate base
base

$$K_a = \frac{[H^+][A^-]}{[HA]}$$ ← Acid dissociation constant

$K_a = 1.76 \times 10^{-5}$

$K >> 1$, $\Delta G << 0$ → Strong acid

$K << 1$, $\Delta G >> 0$ → weak acid

$$CH_3COOH_{(aq)} \rightleftharpoons CH_3COO^-_{(aq)} + H^+_{(aq)}$$

Given a 1.0 M aqueous sol'n of acetic acid, Calculate PH.

I   1.0   0 10^{-7} → Assumed to be zero if acid is strong enough or concentrated enough

C   -x   +x   +x

E   1.0-x   x   x

$$K_a = \frac{(x)(x)}{\boxed{1-x}} = \cancel{0}\, 1.76 \times 10^{-5}$$

— if x is small enough it can be ignored

$$K_a \approx \frac{x^2}{1} = 1.76 \times 10^{-5} \Rightarrow x = 4.20 \times 10^{-3}$$

since x is <1% of the concentration of the acid, it was acceptable to ignore x in the denominator,

$$pH = -\log_{10}(4.20 \times 10^{-3}) = 2.38$$

Given 50 mL of 0.2M acetic acid and
40 mL of 0.4M sodium hydroxide, calculate pH,
The approach taken for the neutralization of a
weak acid with a strong base is to assume the
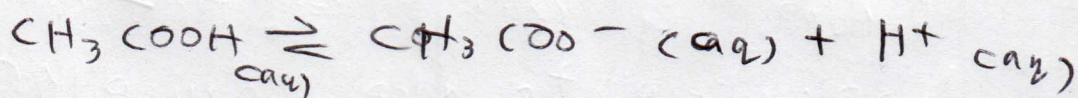two completely react then reachieve equilibrium.

· acetic acid : $0.050\,L \times 0.2M = .010$ mol acetic acid
  NaOH : $0.010\,L \times 0.4\,M = .004$ mol NaOH

acid left : $.0.010$ mol $- 0.004$ mol $= ,006$ mol
acetate produced $=$ mole NaOH used $= ,004$ mol

$$[CH_3 COOH]_i = \frac{,006\,mol}{.060L} = 0.1\underline{M}$$

total soln volume ↗

$$[CH_3 COO^-]_i = \frac{,004\,mol}{.060\,L} = 0.066\,M$$

$$CH_3 COOH \underset{(aq)}{\rightleftharpoons} CH_3 COO^- (aq) + H^+ (aq)$$

	0.1 M	.066M	0
I C	−x	+x	+x
E	⎯	⎯	⎯
	$.1-x$	$.066+x$	x

$$Ka = \frac{(x)(.066 + x)}{.1-x} = 1.76 \times 10^{-5}$$

$$X =$$

$$pH =$$

pH

buffer region

↓

+ equivalence point

Ɪ Ɪ Ɪ ½ equivalence point

0

volume (mL)
NaOH