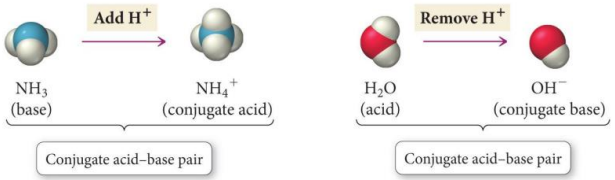
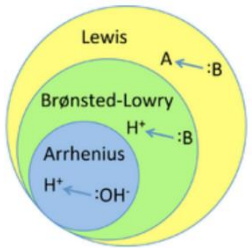
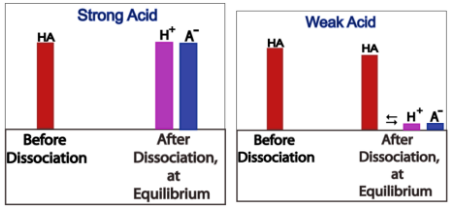
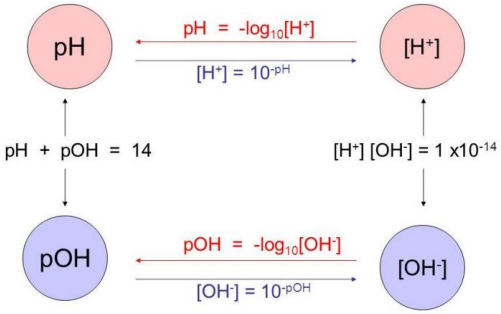
Unit 10 – ACID BASE – GLUE INS

N36 – Acid Base Review









N38 - Salts

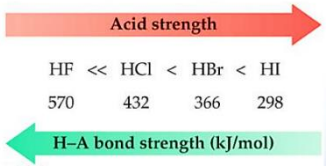
	Turns into a…	Hydrolyzes?
Strong Acid	Weaker conjugate base	No
Weak Acid	Stronger conjugate base	Yes
Strong Base	Weaker conjugate acid	No
Weak Base	Stronger conjugate acid	Yes

	Turns into a…	Hydrolyzes?	Ion makes sol'n
Strong Acid	Weaker conjugate base	No	Neutral
Weak Acid	Stronger conjugate base	Yes	Basic
Strong Base	Weaker conjugate acid	No	Neutral
Weak Base	Stronger conjugate acid	Yes	Acidic

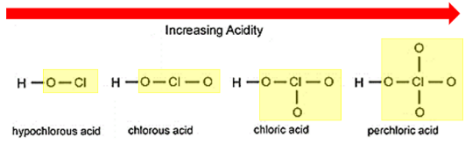
Makes the solution…	Compare Ka and Kb to determine which "wins"
Acidic	Acidic + Neutral
Basic	Basic + Neutral
Neutral	Neutral + Neutral
	Acidic + Basic

$Ka_{(ion)} > Kb_{(ion)}$	Acidic
$Ka_{(ion)} < Kb_{(ion)}$	Basic
$Ka_{(ion)} = Kb_{(ion)}$	Neutral

Strength of Binary Acids



Strength of Oxyacids (and other similar)

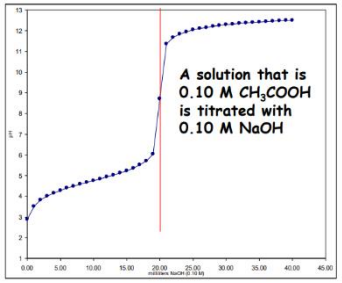


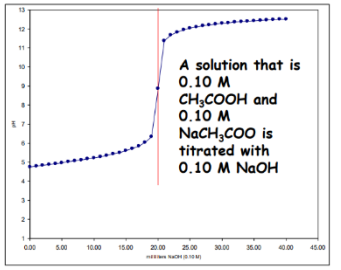
High electronegativity of the ***side group*** pulls electron density AWAY from the bond involving Hydrogen. Bond is therefore weakened so it breaks more easily, therefore more acidic.

N39 – Henderson-Hasselbalch

Weak Acid	Formula of the acid	Example of a salt of the weak acid
Hydrofluoric	HF	KF – Potassium fluoride
Formic	HCOOH	KHCOO – Potassium formate
Benzoic	C_6H_5COOH	NaC_6H_5COO – Sodium benzoate
Acetic	CH_3COOH	NaH_3COO – Sodium acetate
Carbonic	H_2CO_3	$NaHCO_3$ - Sodium bicarbonate
Propanoic	$HC_3H_5O_2$	$NaC_3H_5O_2$ - Sodium propanoate
Hydrocyanic	HCN	KCN - potassium cyanide

Weak Base	Formula of the base	Example of a salt of the weak acid
Ammonia	NH_3	NH_4Cl - ammonium chloride
Methylamine	CH_3NH_2	CH_3NH_3Cl – methylammonium chloride
Ethylamine	$C_2H_5NH_2$	$C_2H_5NH_3NO_3$ - ethylammonium nitrate
Aniline	$C_6H_5NH_2$	$C_6H_5NH_3Cl$ – aniline hydrochloride
Pyridine	C_5H_5N	C_5H_5NHCl – pyridine hydrochloride





One way of doing these calculations

Rearrange your Law of Mass Action:

$$Ka = \frac{[H^+][A^-]}{[HA]} \rightarrow [H^+] = Ka\frac{[HA]}{[A^-]} \rightarrow = Ka\frac{[Acid]}{[conj.Base]_{salt}}$$

$$Kb = \frac{[BH^+][OH^-]}{[B]} \rightarrow [OH^-] = Kb\frac{[B]}{[BH^+]} \rightarrow = Kb\frac{[Base]}{[conj.Acid]_{salt}}$$

Henderson-Hasselbalch Equation

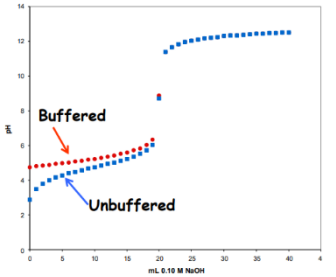
$$pH = pK_a + log\left(\frac{A^-}{HA}\right) \quad = pK_a + log\left(\frac{[Base]}{[Acid]}\right)$$

$$pOH = pK_b + log\left(\frac{BH^+}{B}\right) = pK_b + log\left(\frac{[Acid]}{[Base]}\right)$$

pKa = -log(Ka)
pKb = -log(Kb)

Just like pH = -log[H⁺]

The acids or bases may be conjugates from the salt!



Other ways to think about He-Ha

Acid with a buffer:

$$pH = pK_a + log\left(\frac{[salt]}{[Acid]}\right) = pK_a + log\left(\frac{[conj.Base]}{[Acid]}\right)$$

Base with a buffer:

$$pOH = pK_b + log\left(\frac{[salt]}{[Base]}\right) = pKb + log\left(\frac{[conj.Acid]}{[Base]}\right)$$

Glue-ins Continued on next page!

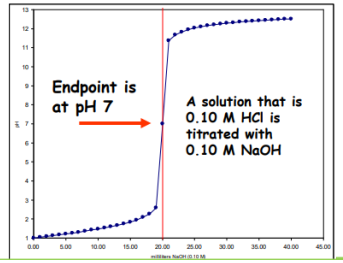
Suggestions…

Pick a method and stick to it. They all have pros and cons.

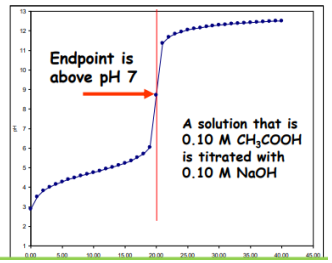
- **Ice tables**
  - Pro = familiar
  - Con = takes forever, lots of steps
- **He-Ha**
  - Pro = fast, on the AP eq. sheet
  - Con = Have to recognize to use it, not always solving for pH
- **Rearranging Law of Mass Action**
  - Pro = simple
  - Con = Have to recognize to use it, extra step to get to pH or pOH

N41 – Titration

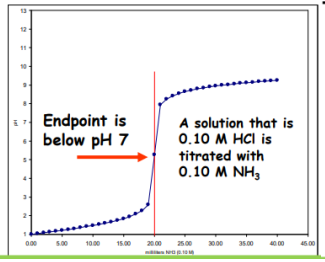
Strong Acid/Strong Base Titration



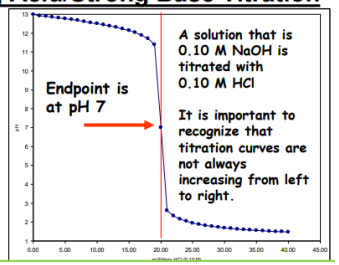
Weak Acid/Strong Base Titration



Strong Acid/Weak Base Titration

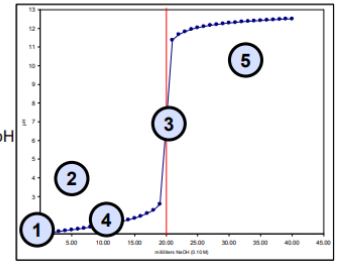


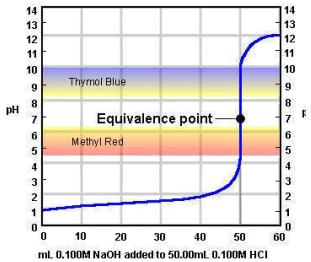
Strong Acid/Strong Base Titration

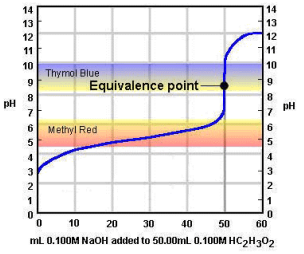


Calculations to Plot a Titration Curve

1. **Starting pH**
   - ICE table then pH
2. **Early on during titration**
   - Stoich then He-Ha
3. **Equivalence Point**
   - mol acid = mol base
   - No more buffer! Reverse rxn
   - Calc new K value - ICE then pH
4. **½ Way Point**
   - ½ moles @ eq.pt
   - pH = pKa
5. **Towards end of titration**
   - Extra titrant left over
   - Stoich then simple pH







Indicator	pH Range in which Color Change Occurs	Color Change as pH Increases
Crystal violet	0.0 - 1.6	yellow to blue
Thymol blue	1.2 - 2.8	red to yellow
Orange IV	1.4 - 2.8	red to yellow
Methyl orange	3.2 - 4.4	red to yellow
Bromcresol green	3.8 - 5.4	yellow to blue
Methyl red	4.8 - 6.2	red to yellow
Chlorophenol red	5.2 - 6.8	yellow to red
Bromthymol blue	6.0 - 7.6	yellow to blue
Phenol red	6.6 - 8.0	yellow to red
Neutral red	6.8 - 8.0	red to amber
Thymol blue	8.0 - 9.6	yellow to blue
Phenolphthalein	8.2 - 10.0	colourless to pink
Thymolphthalein	9.4 - 10.6	colourless to blue
Alizarin yellow	10.1 - 12.0	yellow to blue
Indigo carmine	11.4 - 13.0	blue to yellow

N39 – Henderson-Hasselbalch

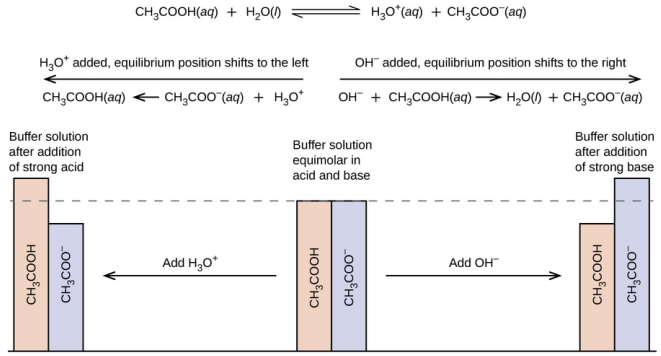


Figure 14.6.2: This diagram shows the buffer action of these reactions.

N40 - Ksp

[14] When solid BaF_2 is added to H_2O the following equilibrium is established.

$$BaF_2(s) \rightleftharpoons Ba^{2+}(aq) + 2\,F^-(aq) \quad K_{sp} = 1.5\times10^{-6} \text{ at } 25\ ^\circ C$$

a. Calculate the molar solubility of barium fluoride at 25 °C.

b. Explain how adding each of the following substances affects the solubility of BaF_2 in water.
   i. 0.10 M $Ba(NO_3)_2$
   ii. 0.10 M HNO_3

c. In an experiment to determine the K_{sp} of PbF_2 a student starts with 0.10 M $Pb(NO_3)_2$ and 0.10 M KF and uses the method of serial dilutions to find the lowest $[Pb^{2+}]$ and $[F^-]$ that form a precipitate when mixed. If the student uses the concentration of the ions in the combined solution to determine K_{sp}, will the value of K_{sp} calculated be too large, too small or just right? Explain.
   K_{sp} for $PbF_2 = 4.0 \times 10^{-8}$

d. i. In a solution of 0.010 M barium nitrate and 0.010 M lead(II) nitrate, which will precipitate first, BaF_2 or PbF_2, as NaF(s) is added? Assume volume changes are negligible. Explain (support your answer with calculations).
   ii. When the more soluble fluoride begins to precipitate, what is the concentration of the cation for the less soluble fluoride that remains in solution?