More Acid-Base Reactions and Equilibria

<u>Possible Acid/Base Reactions:</u>

1. Strong Acid/Strong Base

$$HCl + NaOH \rightarrow H_2O + NaCl$$

– Net ionic eqn: $H^+ + OH^- \rightarrow H_2O$
– Reaction proceeds until one of the two reagents is consumed!
– pH is determined by:

2. Strong Acid/Weak Base (or Weak Acid/Strong Base)

$$HCl + NH_3 \rightarrow NH_4^+ + Cl^-$$
$$HCN + NaOH \rightarrow NaCN + H_2O$$

– Net ionic eqn(s): $H^+ + B \rightarrow HB^+$ $\quad K=K_b/K_w$
$HA + OH^- \rightarrow A^- + H_2O$ $\quad K=K_a/K_w$
– Reaction proceeds until one of the two reagents is consumed!
– pH is determined by:

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3. Weak Acid/Weak Base

$$HCN + NH_3 \rightleftarrows NH_4^+ + CN^-$$

– Net ionic eqn(s): $HA + B \rightleftarrows HB^+ + A^-$ $\quad K=K_aK_b/K_w$
– Extent of reaction depends on "strength" of HA and B
– pH is determined by:

Example: What is the pH of a solution prepared by adding 10 mL of 0.100 M NaOH to 25 mL of 0.100 M HCN ($K_a = 4.0 \times 10^{-10}$)?

	$HCN \quad + \quad OH^- \quad \rightarrow \quad CN^- \quad + \quad H_2O$
mol before rxn	
mol after rxn	
[X] after rxn	

- What will determine pH?

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Closer look at Acid/Base Equilibria

- Common Ion Effect: Le Chatelier's Principle restated!

- A solution that contains an acid and its conjugate base resists dramatic change in composition

$$HCN \rightleftarrows CN^- + H^+$$

$$K_a = \frac{[H^+][CN^-]}{[HCN]} = 4.0 \times 10^{-10}$$

- We can take advantage of this to prepare solutions that resist a change in [H+].

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Buffer Solutions

Go back to earlier example using HCN and NaOH. What is the new pH after adding an additional 5 mL of NaOH to the solution already at pH 9.22? How would this compare to a sol'n of NaOH @ pH 9.22?

	$HCN + OH^- \rightarrow CN^- + H_2O$
mol before rxn	
mol after rxn	
[X] after rxn	

- What is the new pH?

$$HCN \rightleftarrows CN^- + H^+ \qquad K_a = 4.0 \times 10^{-10}$$

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Buffer Solutions

The pH of a buffer solution depends on:

Henderson-Hasslebach Equation (know where this comes from)

$$pH = pK_a + \log \frac{[\text{conjugate base}]}{[\text{acid}]}$$

*Can write similar expression for K_b

Buffer Considerations

- Preparing buffer solutions:
- Selecting a buffer system:
- Buffer capacity:

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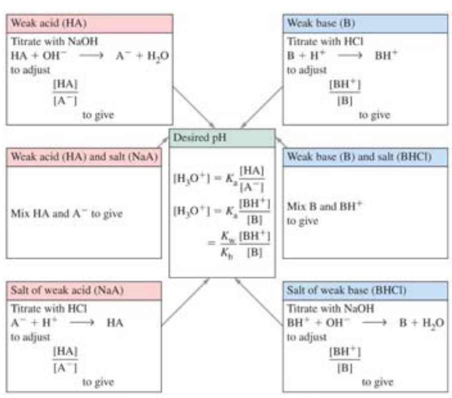
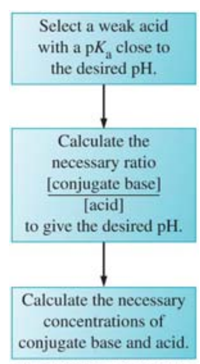
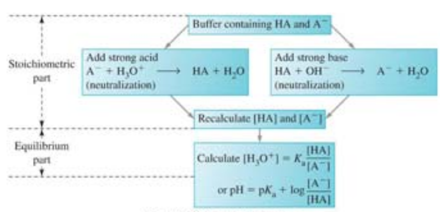
General Buffer Strategies







Buffer Examples

Example: What is the pH of a solution prepared by dissolving 15.0g $NaHCO_3$ and 18.0g Na_2CO_3 in 1.00L of H_2O?

Example: How would you prepare 1.00 L of a pH 5.00 acetic acid/sodium acetate buffer?

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Titrations:

Remember the purpose of a titration!

<u>Titration Types</u>

What controls the solution composition as titrant is added?

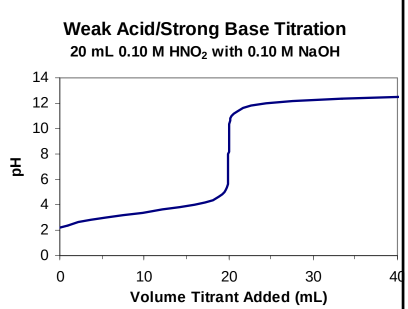
Strong Acid/Strong Base: Basically a stoichiometry problem

– pH depends on unreacted acid or base ("what's left")

Strong Acid/Weak Base: (or Weak Acid/Strong Base)

Monoprotic:

– 3 regions (types of calculations)
– pH is determined by "what's left"
– Consider stoichiometry component and pH-determination (equilibrium) component
  • "happy spots"



Titrations

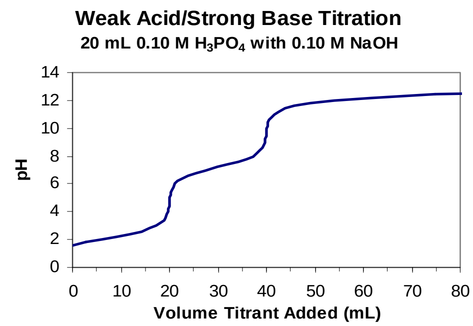
Polyprotic acid/base titrations:

- Same thought process, just a couple more regions
- Amphiprotic considerations

$H_3PO_4 \rightleftarrows H_2PO_{4-} + H^+$ $pK_{a1} = 2.12$

$H_2PO_4^- \rightleftarrows HPO_4^{2-} + H^+$ $pK_{a2} = 7.21$

$HPO_4^{2-} \rightleftarrows PO_4^{3-} + H^+$ $pK_{a3} = 12.32$



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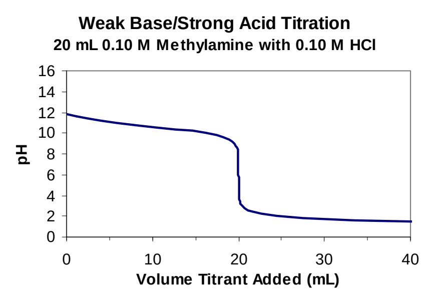
Titrations

- **Titrating bases with strong acids:**
  - Same general concepts, just work with K_b



$CH_3NH_2 + H_2O \rightleftarrows CH_3NH_3^+ + OH^-$ $pK_b = 3.36$

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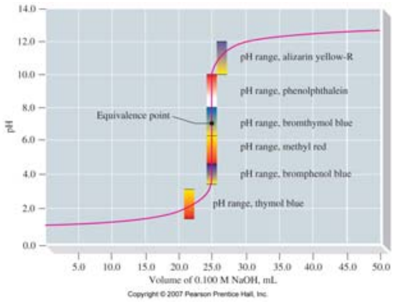
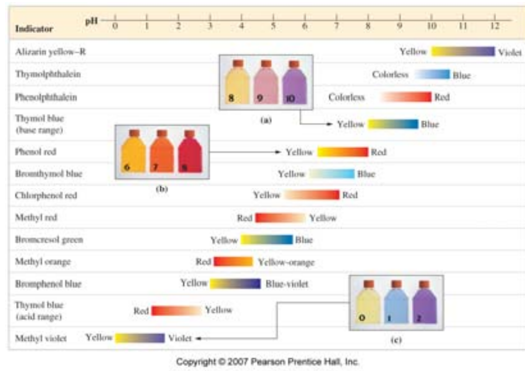
Titrations: End Point Determination

Indicators:

- Weak acids (bases) that change properties when H^+ is lost (or gained)
- Choosing an indicator
- Titration error





- **<u>Other means to determine endpoints</u>**

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