CHAPTER 15

MORE EQUILIBRIUM

- Common Ion Effect
- Buffers
- Titration
- Solubility EQ
- Solubility & pH
- Complexion EQ/Solubility

CH 16 applies the math we learned in CH 15 (which applied math from CH 14)
- this week's lab looks at the complex ion (read ahead)

"Common Ion Effect"
- a result of the math we have been using.

If you put an ion into a solution @ ESL, it may disturb it!

Example: Vitamin C - already had H+ (pH 1.5) in the tissue

- How to handle? put it into the ICE EQN for initial

Henderson- Haselebich Equation
- super-duper derivation from Ka

\[
Ka = \frac{[H^+][A^-]}{[HA]}
\]

\[
\log(K_a) = \log([H^+]) + \log\left(\frac{[A^-]}{[HA]}\right)
\]

\[
\text{pH} = \text{p}K_a + \log\left(\frac{[A^-]}{[HA]}\right)
\]

\[
\text{products reactants !}
\]
Any time you add some of the acid & some of the conjugate base
\((\text{H}^+ \text{H})\), use the \(\text{H}^+ \text{H}\) eqn. : super-strong \((\text{except extreme ratios } \sim 10^7:1)\)

\[ \text{e.g. add } 0.2 \text{ M formic acid and } 0.25 \text{ M sodium formate, pH?} \]

\[ \text{HCO}_2\text{H} \rightleftharpoons \text{H}^+ + \text{CO}_2\text{O}^- \quad K_a = 1.7 \times 10^{-4} \]

\[ pK_a = -\log(K_a) = -\log(1.7 \times 10^{-4}) = 3.77 \quad \text{(2 sf)} \]

\[ [\text{HCO}_2\text{H}] = 0.2 \text{ M, } [\text{A}^-] = 0.25 \text{ M} \]

\[ \text{pH} = pK_a + \log\left(\frac{[\text{A}^-]}{[\text{HCO}_2\text{H}]}\right) = 3.77 + \log\left(\frac{0.25}{0.20}\right) = 3.87 \]

\[ \rightarrow \text{note that } pK_a \text{ is more useful than } K_a! \]

\[ 0.25 \text{ M ethylamine and } 0.05 \text{ M of its acid chloride (C}_6\text{H}_5\text{Cl)} \]

\[ \text{are added to a system; pH?} \]

\[ \text{this is a weak base. } \]

\[ \text{2 methods: } \]

\[ 1) \text{ derive } HH \text{ for base} \]

\[ \times \]

\[ 2) \text{ Reverse reaction and write } K_b \]

\[ \text{CH}_3\text{NH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{NH}_3^+ + \text{OH}^- \quad K_b = 4.4 \times 10^{-4} \]

\[ \text{reverse/}K_a: \]

\[ \text{CH}_3\text{NH}_3^+ + \text{OH}^- \rightleftharpoons \text{CH}_3\text{NH}_2 + \text{H}_2\text{O} \quad -pK_b = -3.36 \quad \text{< negative log} \]

\[ + \text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^- \quad pK_w = 14 \]

\[ \text{CH}_3\text{NH}_3^+ \rightleftharpoons \text{CH}_3\text{NH}_2 + \text{H}^+ \quad pK_a = pK_w - pK_b = 14 - 3.36 = 10.64 \]

\[ \text{pH} = pK_a + \log\left(\frac{[\text{A}^-]}{[\text{HCO}_2\text{H}]}\right) = 10.64 + \log\left(\frac{0.25}{0.05}\right) = 10.64 + \log(5) = 10.64 + 0.70 \]

\[ = 11.34 \]
pH Titrations

1. Next week we will repeat our titrations with the pH meter.
2. pH curve is different for different types of acid/base.

Strong Acid / Weak Base

- Titrated: 0.1 M NaOH
- Titrant: 0.1 M HCl

1. At Start: 
   
   \[ [\text{H}^+] = [\text{HCl}] = 0.1 \]  
   \[ \text{pH} = -\log(0.1) = 1 \]  

2. Add 0.05 M NaOH
   
   \[ \text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}, \text{ consumes} \]  
   \[ 0.05 \text{M H}^+ \Rightarrow 0.05 \text{M left over} \]  
   \[ \text{pH} = -\log(0.05) = 1.3 \]  

3. At 0.1 M NaOH (equal amounts H\textsuperscript{+}, OH\textsuperscript{-})
   
   \[ \text{pH} = 7 \]  

4. At 50\% excess NaOH, 0.05 M NaOH
   
   \[ \text{pH} = 1.3, \text{ pH} = 12.7 \]  


Weak acid / Strong Base

\[ \text{use H-H again.} \]

\[ \rightarrow \text{consider 2 solutions:} \]

**Soln 1:**

- 0.05 M NaAc
- 0.05 M HAC

\[ \text{pH} = \text{pK}_a + \log \left( \frac{\text{Ac}^-}{\text{HAC}} \right) \]

\[ \text{pH} = \text{pK}_a = 4.74 \]

**Soln 2:**

- 0.1 M HAC
- 0.05 M NaOH

\[ \text{HAC + NaOH \rightarrow H}_2\text{O + NaAc} \]

0.05 M NaOH consumes

0.05 M HAC making 0.05 M

HAC + NaAc

0.05 M