Indicators

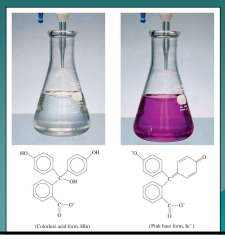
Homework

- Practice AP chapter 17

Acid Base indicators

- *Acid-base indicators* can mark the *end point* of a titration by changing color.
- The *equivalence point* is defined by the reaction stoichiometry.
- For strong acid-strong base titrations, color changes are sharp, allowing more flexibility in choosing an indicator.
- For weak acid titrations, less flexibility exists in choosing an indicator.

The Acid and Base Forms of the Indicator Phenolphthalein



How an acid base indicator works

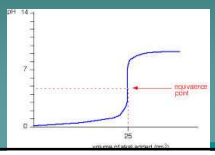
- **A generic indicator will follow this reaction, HId is the reactant indicator, and Id^- is its product**
- **$HId + H_2O \rightleftharpoons H_3O^+ + Id^-$**
- **The color differences are important, HId is one color and Id^- is a different color!**
- **in an acidic solution (add H_3O^+) you see reactant**
- **$HId + H_2O \rightleftharpoons$**
- **in a basic solution (remove H_3O^+) you see product**
- **$HId + H_2O \rightleftharpoons H_3O^+ + Id^-$**

Indicators

- So what is important is the **$[Id^-]$ /[HID]**
- It needs to be experimentally determined what ratio will cause your eyes to see different colors.
- The human eye sees a color change for most indicators when **$[Id^-]$ /[HID]** is between .1 and 10
- So we can use the Henderson Hasselbach equation
- $pH_{color\ change} = pK_a + \log [A^-]/[HA]$
- $pH_{color\ change} = pK_a \pm 1$

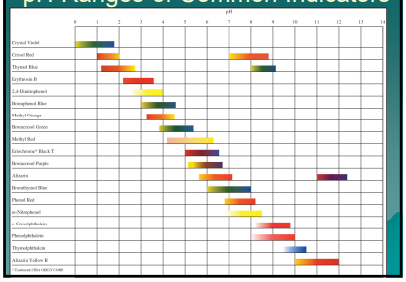
Choosing a suitable indicator

- You need to choose an indicator that will change closest to the equivalence point of your reaction.
- Which indicator should you use?



Indicator	pK_a
Methyl Red	5.5
Bromothymol Blue	7.1
Phenolphthalein	8.7

pH Ranges of Common Indicators



Solubility Equilibria and the Solubility Product

- K_{sp} is the solubility product constant is equilibrium between solid solute and dissolved ions.
- For $M_pX_q \rightarrow p\ M^{n+} + q\ X^{z-}$
- $K_{sp} = [M^{n+}]^p\ [X^{z-}]^q$
- or $Fe(OH)_3 \rightarrow Fe^{3+} + 3\ OH^-$
- $K_{sp} = [Fe^{3+}]\ [OH^-]^3$

Writing K_{sp} for Slightly Soluble Ionic Compounds

- Note, using solubility rules these compounds are all **insoluble**.
- K_{sp} would be very high for anything that is "soluble" and very low for anything that in "insoluble".
- In reality, nothing is completely soluble or insoluble. Everything dissolves to some extent before the solution becomes saturated, and then the rate of precipitation equals the rate of dissolution.

- *Insoluble* solutions become saturated with very **low concentrations**.
- *Soluble* solutions become saturated with **high concentrations**.

- Write the solubility product expression for:
- Magnesium carbonate
- Iron(II) hydroxide
- Calcium phosphate
- Silver sulfide

Calculating K_{sp} from Solubility

- Solubility is the concentration of original salt that has dissolved.

- The K_{sp} for copper(II) iodate, $Cu(IO_3)_2$, is 1.4×10^{-7} at 25° C. Calculate its solubility at 25° C.
- Copper(I) bromide has a measured solubility of 2.0×10^{-4} M at 25° C. Calculate its K_{sp} value.
- Calculate the K_{sp} value for bismuth sulfide (Bi_2S_3), which has a solubility of 1.0×10^{-15} M at 25° C.

Relative solubility

- Comparing how soluble one salt is to another can only be predicted by comparing K_{sp} values only for salts that produce the same total number of ions.
- $MX \rightarrow M^+ + X^-$
- $K_{sp} = [M^+][X^-]$ $\quad [M^+] = [X^-]$
- $K_{sp} = x^2$

- $MX_2 \rightarrow M^{2+} + 2\,X^-$
- $K_{sp} = [M^{2+}][X^-]^2$ $\quad [M^+] = 2\,[X^-]$
- $K_{sp} = 4x^3$

The Common Ion Effect

- This is dissolving a salt in a solution that contains one of the ions.
- You will need an stoichiometry chart.

- Calculate the solubility of solid CaF_2 ($K_{sp} = 4.0 \times 10^{-11}$) in a 0.025 M NaF solution.