Key Concepts – (know these)

1. **Buffer pH Range:** pK_A ± 1

2. **Buffer Capacity:**
   The Acid and Conjugate Base concentrations determine the capacity to absorb strong acid or strong base.
   Large Concentrations → Large Capacity

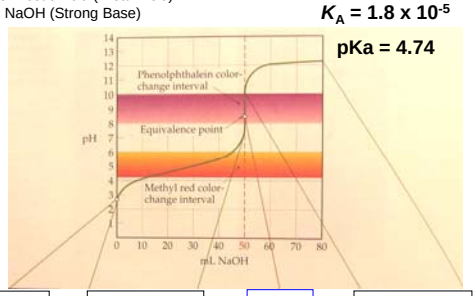
3. **Titrations** permit determinations of:
   (i) the amount of acid (or base) present,
   (ii) whether an acid (or base) is strong or weak, and
   (iii) the pK_A (or pK_B) in of a weak acid (or base)

Adding a Strong base to a Weak acid

Titration of Acetic Acid (Weak Acid) with NaOH (Strong Base)

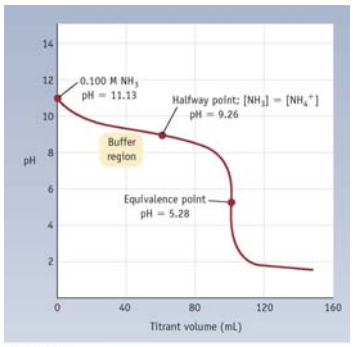
K_A = 1.8 x 10^{-5}

pKa = 4.74

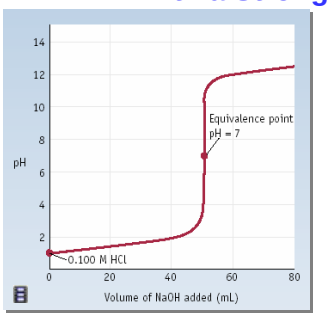


HAcO 0.100 M, 50 ml | HAcO + NaAcO | NaAcO 0.050 M | NaAcO + NaOH

Titration of a Weak Base with a Strong Acid



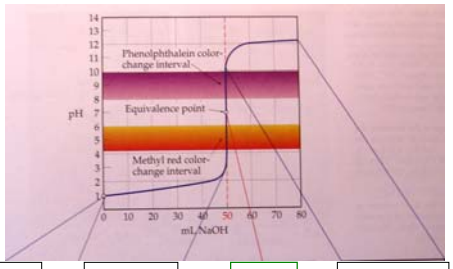
Titration of a Strong Acid with a Strong Base



50.0 mL of 0.100 M HCl titrated with 0.100 M NaOH

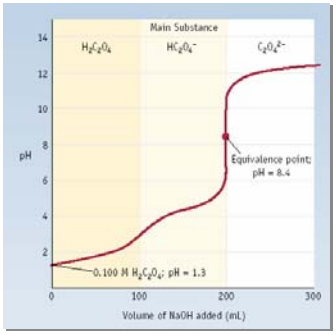
Amount of base added	pH
100.0	12.52
80.0	12.36
60.0	11.96
55.0	11.68
51.0	11.00
50.0	7.00
49.0	3.00
48.0	2.69
45.0	2.28
40.0	1.95
20.0	1.37
10.0	1.18
0.0	1.00
very large amount	13.00 (maximum)

Adding a Strong base to a Strong acid



HCl 0.100 M | HCl + NaCl | NaCl 0.050 M | NaCl + NaOH

Titration of a Weak Diprotic Acid (Oxalic acid) with a Strong Base



Two equivalents of base are required to complete the titration of a diprotic acid; moles of OH^- at equivalence (n_{OH-}) = 2 times the moles of acid (n_{Acid})

$n_{OH-} = 2n_{Acid}$

Notice that there are two inflection points, one for each acid form

Types of Chemical Reactions

- EXCHANGE REACTIONS:
  AB + CD → AD + CB
  - Acid-base:
    $CH_3CO_2H + NaOH \rightarrow NaCH_3CO_2 + H_2O$
  - Gas forming:
    $CaCO_3 + 2\ HCl \rightarrow CaCl_2 + CO_2 + H_2O$
  - Precipitation:
    $Pb(NO_3)_2 + 2\ KI \rightarrow PbI_2(s) + 2\ KNO_3$
- OXIDATION REDUCTION (Chapter 20)
  - $4\ Fe + 3\ O_2 \rightarrow 2\ Fe_2O_3$
- *Apply equilibrium principles to acid-base and precipitation reactions.*

Some Values of K_{sp}

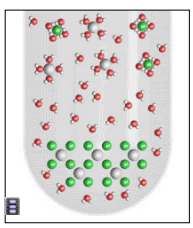
Table 18.2 Some Common Insoluble Compounds and Their K_{sp} Values*

Formula	Name	K_{sp} (25 °C)
$CaCO_3$	Calcium carbonate	3.4×10^{-9}
$MnCO_3$	Manganese(II) carbonate	2.3×10^{-11}
$FeCO_3$	Iron(II) carbonate	3.1×10^{-11}
CaF_2	Calcium fluoride	5.3×10^{-11}
AgCl	Silver chloride	1.8×10^{-10}
AgBr	Silver bromide	5.4×10^{-13}
$CaSO_4$	Calcium sulfate	4.9×10^{-5}
$BaSO_4$	Barium sulfate	1.1×10^{-10}
$SrSO_4$	Strontium sulfate	3.4×10^{-7}
$Ca(OH)_2$	Calcium hydroxide	5.5×10^{-5}

Solubility of Lead(II) Chloride

$$PbCl_2(s) \leftrightarrows Pb^{2+}(aq) + 2\ Cl^-(aq)$$

$$K_{sp} = 1.9 \times 10^{-5} = [Pb^{2+}][Cl^-]^2$$



Solubility of Silver Chloride

$Ag^+(aq) + NO_3^-(aq)$	*Soluble*
$+\ Na^+(aq) + Cl^-(aq)$	*Soluble*
$Na^+(aq) + NO_3^-(aq)$	*Soluble*
$Ag^+(aq) + Cl^-(aq)$	*Soluble(?)*

Dilute Solutions: No Precipitation

Concentrated Solutions: Precipitation

K_{sp}: Quantifies Solubility

$$AgCl(s) \leftrightarrows Ag^+(aq) + Cl^-(aq)$$

$$K_{sp} = [Ag^+]_{eq}[Cl^-]_{eq} = 1.9 \times 10^{-10}$$

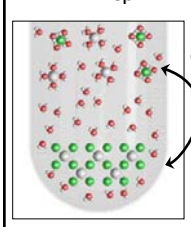
$Q_{sp} < K_{sp}$ → soluble, no precipitation

$Q_{sp} > K_{sp}$ → exceeds solubility, precipitation

Solubility of Lead(II) Iodide

$$PbCl_2(s) \leftrightarrows Pb^{2+}(aq) + 2\ I^-(aq)$$

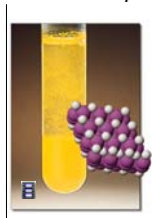
$$K_{sp} = 9.8 \times 10^{-9} = [Pb^{2+}]_{eq}[I^-]_{eq}^2$$



The equilibrium dissociated ions in solution and the precipitate is dynamic.

At equilibrium:

the rate of dissolution = the rate of precipitate formation



Lead Iodide

The value of K_{sp} depends on a 'competition' between interactions in the solid and ion solvation.

As the interactions in the solid become more favorable (and/or ion solvation becomes poorer), the substance becomes *more* insoluble. (K_{sp} becomes smaller.)

Some solubility guidelines

(useful but not necessary to memorize)

1. All sodium, potassium, and ammonium salts are soluble.
2. All nitrates, acetates and perchlorates are soluble.
3. All silver, lead and mercury(I) salts are insoluble.
4. All chlorides, bromides and iodides are soluble (except silver, lead and mercury(I) halides).
5. All carbonates, sulfides, oxides and hydroxides are insoluble.
6. All sulfates are soluble except calcium sulfate and barium sulfate.