ACIDS AND BASES – 6
ACID BASE TITRATIONS

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ACID BASE TITRATION - INTRODUCTION

- **Titration** – addition of a solution of known concentration to a solution of unknown concentration until end point. A pH curve can be drawn to gather information.
- **Standard solution** - solution with known concentration.
- **Titrant** – the solution that is place in the burette.
- **Equivalence point** – the point when stochiometrically equivalent amounts of acid-base have been added.
- **Endpoint** – the point at which titration is finished (usually with the help of an indicator).
- **Indicator** – a chemical that changes color at the end point at a specific pH range. Indicator is chosen either using a pH curve or by knowing the acid-base strength.

TYPES OF ACID BASE TITRATIONS

- Depending on the type of titration, the pH will be different for the "end point"; at equivalence point the pH will be 7.

- Strong acid – strong base

  $HCl + NaOH \longrightarrow NaCl + H_2O$

  final products are neutral (pH = 7)

  adding one drop more of the titrant will drastically change the pH.

- Weak acid – strong base

  $CH_3COOH + NaOH \rightleftharpoons CH_3COONa + H_2O$

  final products are basic (pH>7)

- Strong acid – weak base

  $HCl + NH_4OH \rightleftharpoons NH_4Cl + H_2O$

  final products are acidic (pH<7)

EXAMPLE: STRONG ACID-BASE TITRATION

Calculate the pOH and the pH of a solution in which 10.0 mL of 0.100 *M* HCl is added to 25.0 mL of 0.100 *M* NaOH.

Solution:

Calculate mols of HCl and NaOH then subtract the two; add the volumes; determine concentration and determine pH.

The overall reaction is: HCl + NaOH $\rightarrow$ H_2O + NaCl

Moles HCl = (0.100 *M*)(10.0×10^{-3} L) = 1.00×10^{-3} mol

Moles NaOH = (0.100 *M*)(25.0×10^{-3} L) = 2.50×10^{-3} mol

All of the HCl reacts, ($2.5 \times 10^{-3} - 1.00 \times 10^{-3}$) = 1.50×10^{-3} mol NaOH.

New volume is 10.0 mL + 25.0 mL = 35.0 mL = 0.0350 L

$$[OH^-] = \frac{1.50 \times 10^{-3}\ \text{mol}}{35.0 \times 10^{-3}\ \text{L}} = 4.29 \times 10^{-2}\ M$$

$$pH = 14.00 - pOH = 14.00 - (-\log[OH^-])$$

$$pH = 14.00 - 1.368$$

$$pH = 12.632$$

EXAMPLE: WA + SB TITRATION

What is the [OH-] and the pH at equivalence point in the titration of a 500.0 mL 0.100 M propionic acid (HPr) and 0.050 M $Ca(OH)_2$. $K_a = 1.3 \times 10^{-5}$

Solution:

a) Calculate mols of propionic acid; b) find volume of $Ca(OH)_2$ needed for titration; c) find new volume; d) use eq. method to find pH.

Mols propionic acid = $(0.10\ M)(500. \times 10^{-3}\ L) = 0.050$ mol HPr

$2HPr + Ca(OH)_2 \rightarrow Ca(Pr)_2 + 2H_2O$ *(write equation for mol ratio!)*

Mols $Ca(OH)_2$ = $0.050\ \text{mol HPr}\ \dfrac{1\ \text{mol}\ Ca(OH)_2}{2\ \text{mol HPr}} = 0.025\ \text{mol}\ Ca(OH)_2$

$$\text{Volume}\ Ca(OH)_2 = \frac{0.025\ \text{mol}}{0.050\ \frac{\text{mol}}{\text{L}}} = 0.50\ \text{L}$$

The new volume is $500. \times 10^{-3}$ L + 0.50 L = 1.00 L

The mols of Pr^- is 0.050 mol (from initial HPr conc.)

New conc. of $[Pr^-] = \dfrac{0.050\ \text{mol}}{1.0\ \text{L}} = 5.0 \times 10^{-2}\ M = 0.050\ M$

EXAMPLE: …CONTD…

Reaction:	Pr^- +	H_2O →	HPr	+	OH^-
Initial	0.05		0		0
Change	-x		+x		+x
Eq. conc	0.05 - x		x		x

$$K_b = \frac{K_w}{K_a} = \frac{1.00 \times 10^{-14}}{1.3 \times 10^{-5}} = 7.7 \times 10^{-10}$$

$$K_b = \frac{[HPr][OH^-]}{[Pr^-]}$$

$$7.7 \times 10^{-10} = \frac{x^2}{(0.050 - x)}$$

$$7.7 \times 10^{-10} = \frac{x^2}{0.050}$$

$$x^2 = 3.85 \times 10^{-11}$$

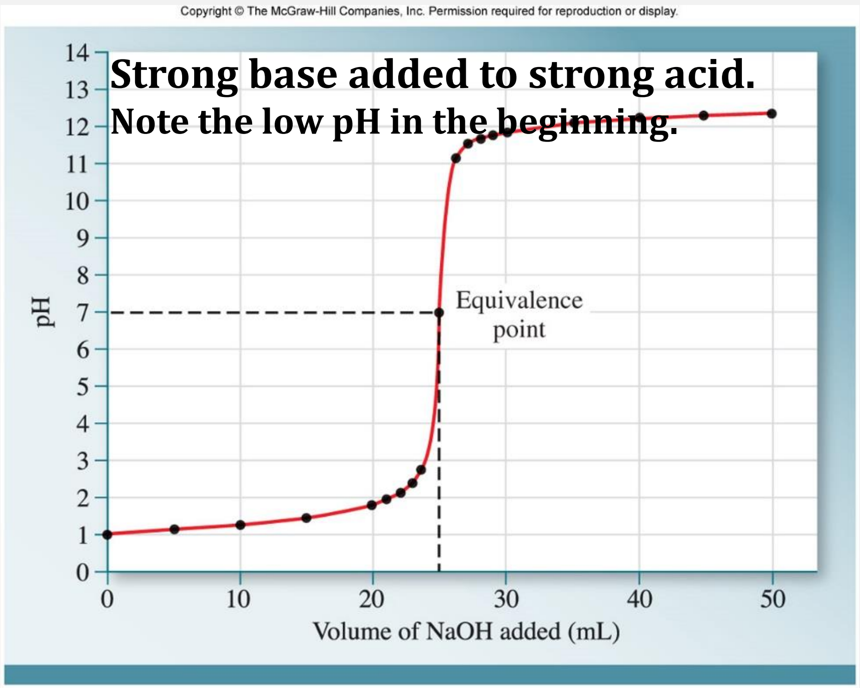
$$x = 6.20 \times 10^{-6}\ M$$

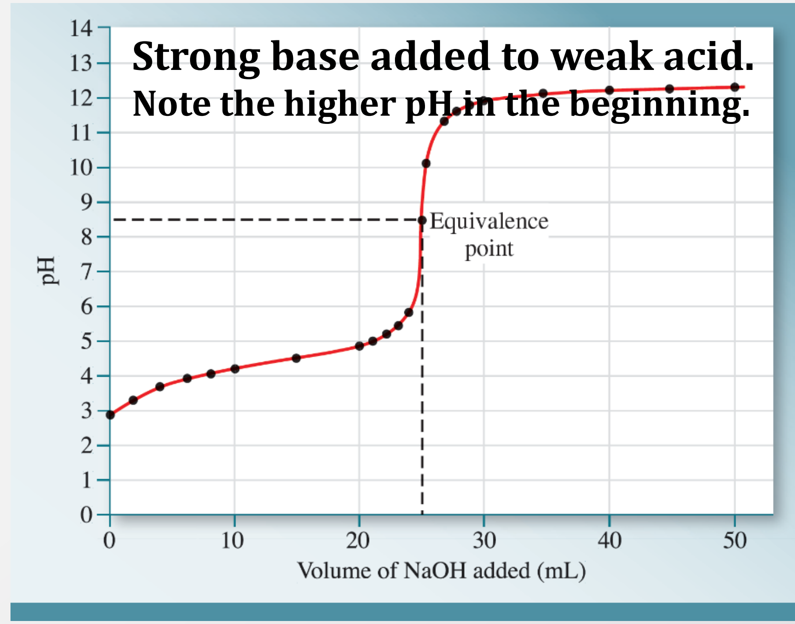
$$[OH^-] = x = 6.20 \times 10^{-6}\ M$$

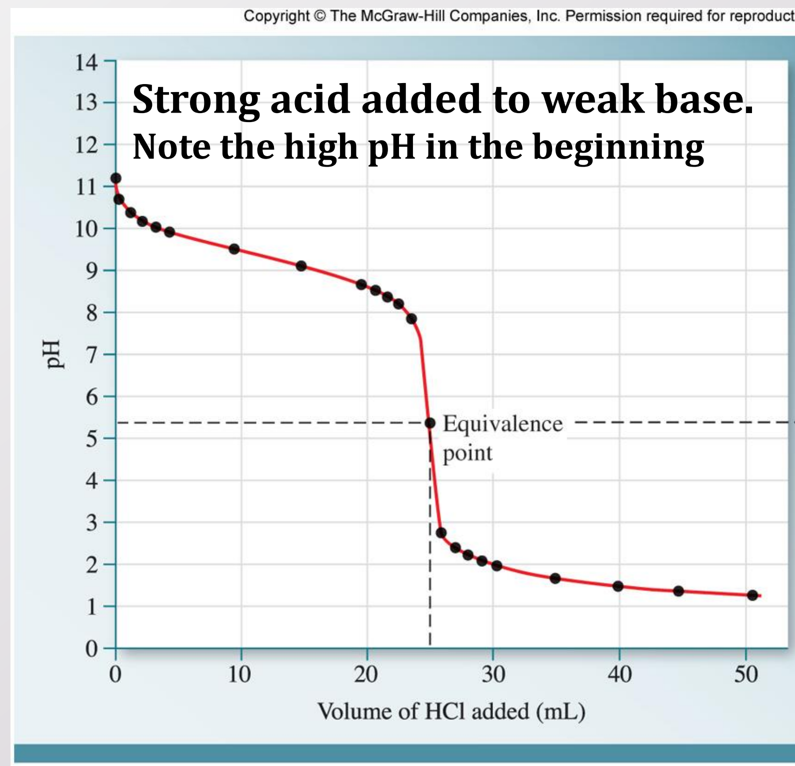
$$pOH = 5.21$$

$$pH = 8.79$$

TITRATION GRAPHS







INDICATORS

- Indicators are weak acids or bases that ionize in acidic or basic conditions and change colors with ionizations.
- pH range can be different for different indicators (see next slide)
- Indicators are key in identifying the end point of a titration.
- The equivalence point must fall in the pH range of the indicator.
- Cabbage juice indicator (shown below) has a variety of colors at the different pH levels.

INDICATORS

TABLE 17.3 Some Common Acid-Base Indicators

	Color		
Indicator	**In Acid**	**In Base**	**pH Range**
Thymol blue	Red	Yellow	1.2−2.8
Bromophenol blue	Yellow	Bluish purple	3.0−4.6
Methyl orange	Orange	Yellow	3.1−4.4
Methyl red	Red	Yellow	4.2−6.3
Chlorophenol blue	Yellow	Red	4.8−6.4
Bromothymol blue	Yellow	Blue	6.0−7.6
Cresol red	Yellow	Red	7.2−8.8
Phenolphthalein	Colorless	Reddish pink	8.3−10.0

KEY CONCEPTS

- Carry out titration calculations
- Calculate pH of a strong acid-base titration
- Calculate pH of strong/weak acid-strong/weak base titration
- Look at graphs and identify type of titrations