

Chapter 7

Acid-Base Equilibria and Titrations

- Principle of Titration—Titration Curve
- Detection of Equivalence Point—Indicator
- Features and Application

7.1 Principle

•Advantages of Titrimetric Method:

- Rapid,
- Convenient,
- Accurate,
- Readily automated

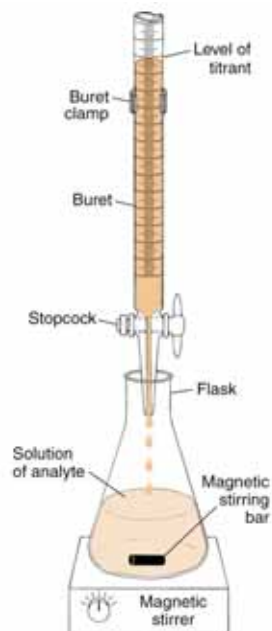
(1). Titration type:

- Volumetric: volume of standard is measured.
- Gravimetric: mass of standard is measured, eg. Mohr method(Cl-Ag-CrO_4).
- Coulometric: time at constant current is measured, i.e., amount of charge.

(2) Terms

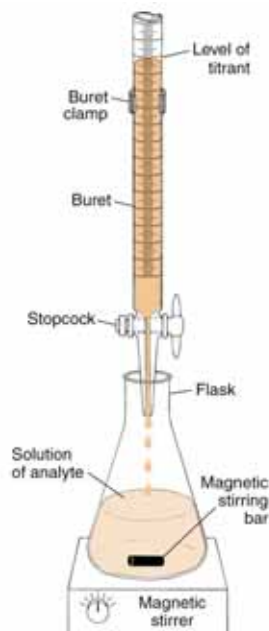
Standard Solution (Standard Titrant):

Reagent of known concentration that is used to carry out the analysis.



Titration:

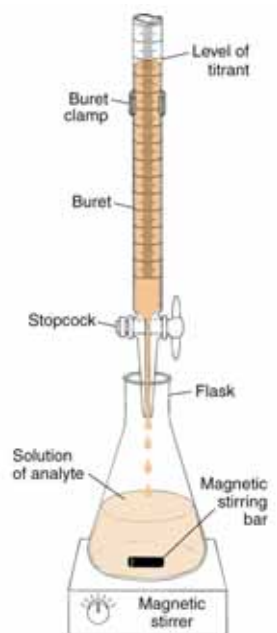
Performed slowly by adding standard solution from a buret until the reaction is judged to be complete.



Equivalence Point: The point when the amount of added titrant is chemically equivalent to the amount of analyte in the sample.

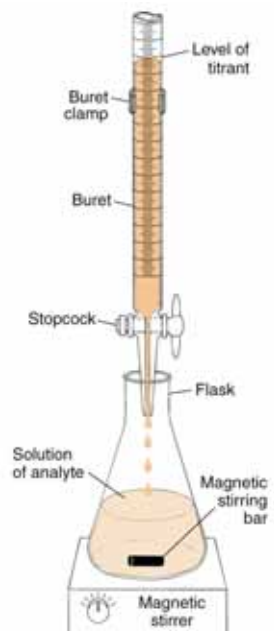
End Point: The point at which an indicator physically changes representing the estimated equivalence point (color or turbidity change).

Titration error = $V_{ep} - V_{eq}$



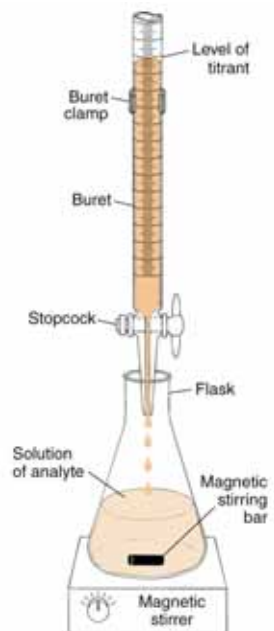
Indicators:

- Added to a titration to give an observable physical change at or near the equivalence point.
- A color or turbidity change.

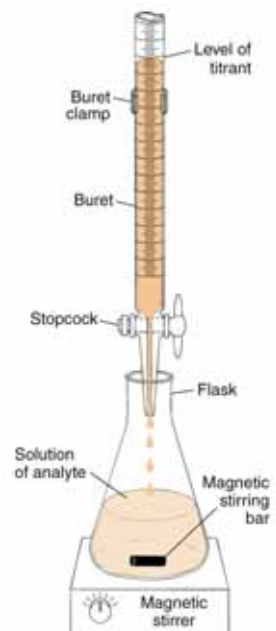


Back-titration:

Sometimes it is necessary to add an excess of standard titrant, and then a second standard titrant to determine how much was added in excess. May be required when a reaction is slow or the standard solution lacks stability.



Blank titration:



Standard Solutions (Titrants)

Desirable Properties of Standard Solutions

1. Sufficiently **stable** (only need to determine concentration once).
2. React **rapidly** with the analyte (less time between additions).
3. React more or less **completely** with the analyte (satisfactory end points).
4. Undergo a **selective** reaction with analyte described by a simple chemical equation

Standard Solutions (Titrants)

Methods for Establishing the Concentration of Standard Solutions

1. **Direct Method:** A carefully weighed quantity of primary standard is diluted to a known volume. First choice if possible.
2. **Standardization:** A titrant to be standardized is used to titrate a weighed quantity of a primary, secondary standard or a known volume of another standard solution.

Primary Standard

Highly purified compound that serves as a reference material.

- standard solution (direct method),
- to standardize a standard solution (standardization).

Primary Standard

1. High purity 100.02%
2. Stability toward air
3. Absence of hydrate water
4. Available at moderate cost
5. Soluble
6. Large F.W.

Secondary Standard

A reference material (a compound) whose purity has been established by a chemical analysis. Often used in lieu of a primary standard because there are not many primary standards.

Summary

titration

titrant

analyte

indicator

equivalence point vs. end point

titration error

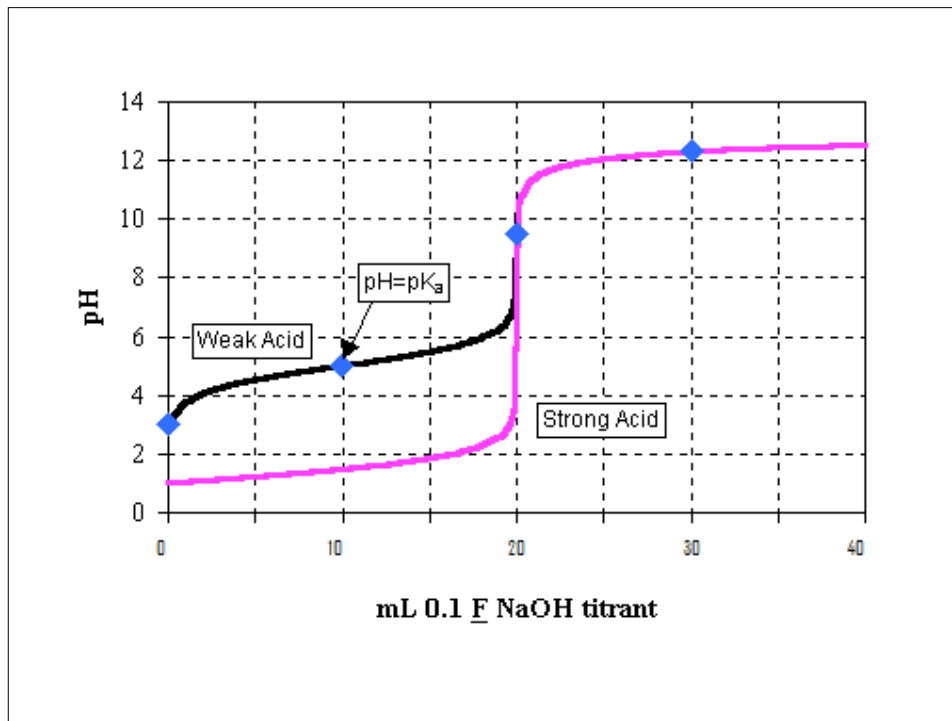
blank titration

Standard: Primary standard, second standard,
standard solution

7.2 Titration of Strong Acid and Strong Base

*How does the pH change as
titrant is added?*

Titration Curve



Strong Acid - Strong Base Titration Curve

EXAMPLE: Derive the titration curve for the titration of 20.00 mL of 0.1000 M HCl with 0.00, 10.00, 19.98, 20.00, 20.02 and 40.00 mL of 0.1000 M NaOH.

0.00,	initial point
10.00 ml,	50% titration
19.98 ml,	half-drop before equivalence point
20.00 ml,	100% titration (equivalence point)
20.02 ml,	half-drop after equivalence point
40.00 ml,	200% titration (far-over)

EXAMPLE: Derive the titration curve for the titration of 20.00 mL of 0.1000 M HCl with 0.00, 10.00, 19.98, 20.00, 20.02 and 40.00 mL of 0.1000 M NaOH.

At 0.00 mL of NaOH added, initial point

$$[\text{H}_3\text{O}^+] = C_{\text{HCl}} = 0.1000 \text{ M}$$

$$\text{pH} = 1.00$$

EXAMPLE: Derive the titration curve for the titration of 20.00 mL of 0.1000 M HCl with 0.00, 10.00, 19.98, 20.00, 20.02 and 40.00 mL of 0.1000 M NaOH.

at 10.00 mL of NaOH added,

$$V_a * M_a > V_b * M_b$$

$$[\text{H}_3\text{O}^+] = \frac{(V_a M_a) - (V_b M_b)}{(V_a + V_b)}$$

$$= \frac{((20.00\text{mL}) * (0.1000\text{M})) - ((10.00\text{mL}) * (0.1000\text{M}))}{(20.00 + 10.00)\text{mL}}$$

$$= 3.33 \times 10^{-2} \text{M}$$

$$\text{pH} = 1.48$$

EXAMPLE: Derive the titration curve for the titration of 20.00 mL of 0.1000 M HCl with 0.00, 10.00, 19.98, 20.00, 20.02 and 40.00 mL of 0.1000 M NaOH.

at 19.98 mL of NaOH added,

$$V_a * M_a > V_b * M_b$$

$$[H_3O^+] = \frac{(V_a M_a) - (V_b M_b)}{(V_a + V_b)}$$

$$= \frac{((20.00\text{mL}) * (0.1000\text{M})) - ((19.98\text{mL}) * (0.1000\text{M}))}{(20.00 + 19.98)\text{mL}}$$

EXAMPLE: Derive the titration curve for the titration of 20.00 mL of 0.1000 M HCl with 0.00, 10.00, 19.98, 20.00, 20.02 and 40.00 mL of 0.1000 M NaOH.

at 20.00 mL of NaOH added

$$V_a * M_a = V_b * M_b, \text{ equivalence point}$$

at equivalence point of a strong acid - strong base titration

$$\text{pH} = 7.00$$

EXAMPLE: Derive the titration curve for the titration of 20.00 mL of 0.1000 M HCl with 0.00, 10.00, 19.98, 20.00, 20.02 and 40.00 mL of 0.1000 M NaOH.

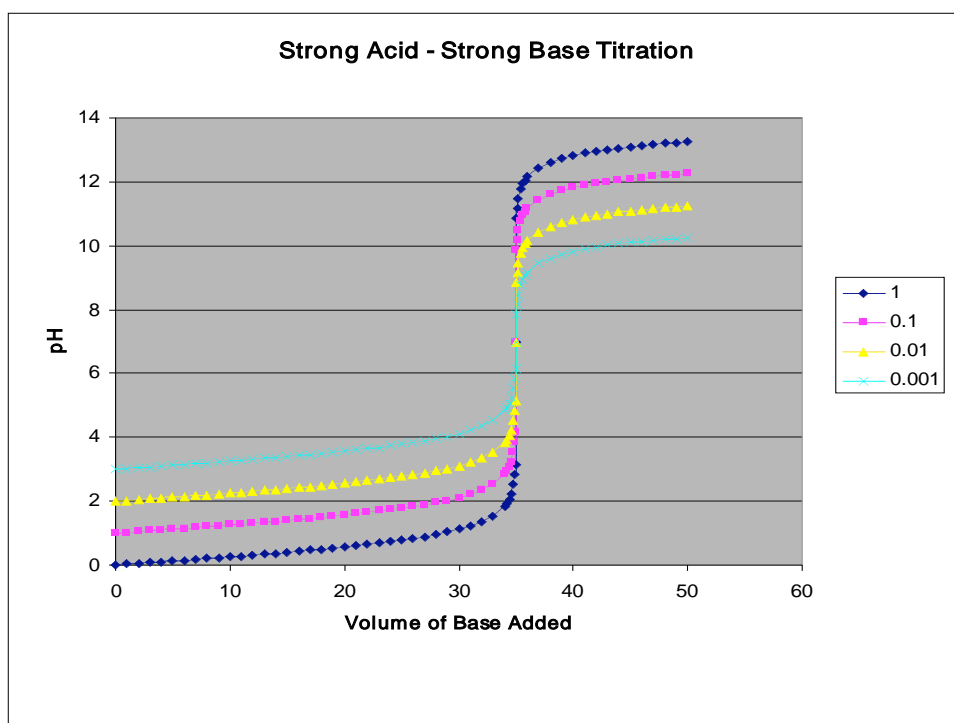
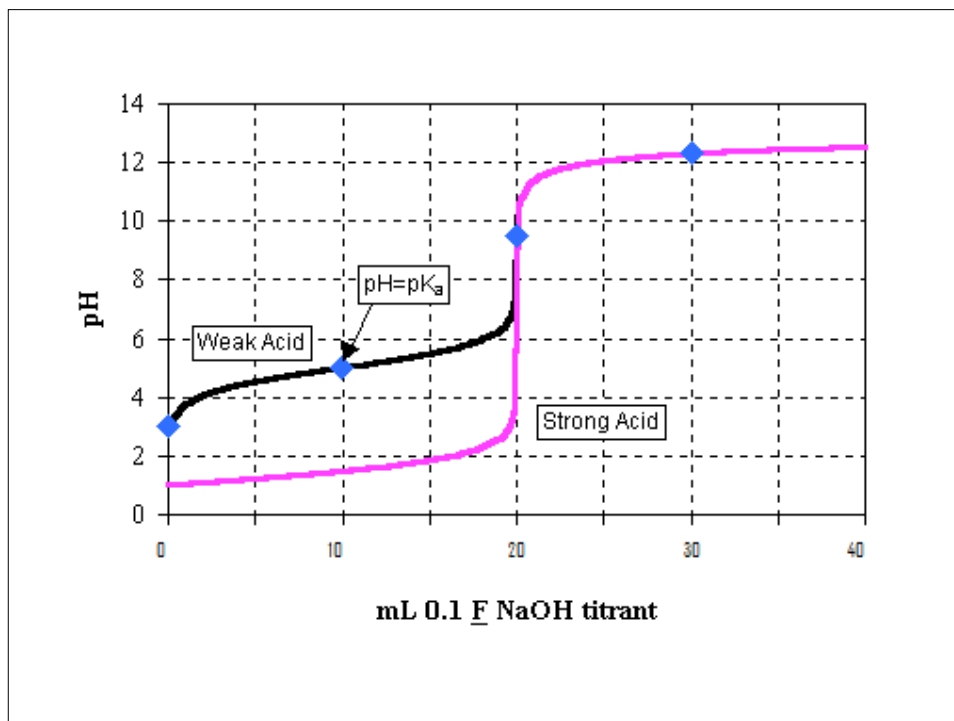
at 20.02 mL of NaOH added, $V_b * M_b > V_a * M_a$,
post equivalence point

$$\begin{aligned}
 [\text{OH}^-] &= \frac{(V_b * M_b) - (V_a * M_a)}{(V_a + V_b)} \\
 &= \frac{((20.02\text{mL}) * (0.1000\text{M})) - ((20.00\text{mL}) * (0.1000\text{M}))}{(20.02.00 + 20.00)\text{mL}} \\
 &= 5.0 \times 10^{-5}\text{M} \qquad \text{pOH} = 4.30 \\
 \text{pH} &= 9.70
 \end{aligned}$$

EXAMPLE: Derive the titration curve for the titration of 20.00 mL of 0.1000 M HCl with 0.00, 10.00, 19.98, 20.00, 20.02 and 40.00 mL of 0.1000 M NaOH.

at 40.00 mL of NaOH added, $V_b * M_b > V_a * M_a$,
post equivalence point

$$\begin{aligned}
 [\text{OH}^-] &= \frac{(V_b * M_b) - (V_a * M_a)}{(V_a + V_b)} \\
 &= \frac{((40.00\text{mL}) * (0.1000\text{M})) - ((20.00\text{mL}) * (0.1000\text{M}))}{(40.00 + 20.00)\text{mL}} \\
 &= 3.33 \times 10^{-2}\text{M}, \qquad \text{pH} = 1.48 \\
 \text{pH} &= 12.52
 \end{aligned}$$



The Effect of Concentration

- The change in pH at the equivalence point is less pronounced at decreased concentrations.
- The pH at the equivalence point is still 7.

14-2 Titration of Weak Acid with Strong Base

EXAMPLE: Derive the titration curve for the titration of 35.00 mL of 0.1000 M $\text{HC}_2\text{H}_3\text{O}_2$ with 0.00, 15.00, 35.00, and 50.00 mL of 0.1000 M NaOH. $K_a = 1.75 \times 10^{-5}\text{M}$

EXAMPLE: Derive the titration curve for the titration of 35.00 mL of 0.1000 M HC₂H₃O₂ with 0.00, 15.00, 35.00, and 50.00 mL of 0.1000 M NaOH.

$$K_a = 1.75 \times 10^{-5} \text{M}$$

at 0.00 mL of NaOH added,
initial point

$$\begin{aligned} [\text{H}_3\text{O}^+] &= \sqrt{K_a C_{\text{HA}}} \\ &= \sqrt{(1.75 \times 10^{-5} \text{ M}) \times (0.1000 \text{ M})} = \\ &= 1.32 \times 10^{-3} \text{ M} \\ \text{pH} &= 2.878 \end{aligned}$$

EXAMPLE: Derive the titration curve for the titration of 35.00 mL of 0.1000 M HC₂H₃O₂ with 0.00, 15.00, 35.00, and 50.00 mL of 0.1000 M NaOH.

$$K_a = 1.75 \times 10^{-5} \text{M}$$

at 15.00 mL of NaOH added, $V_a M_a > V_b M_b$

$$\begin{aligned} [\text{HC}_2\text{H}_3\text{O}_2]_{\text{excess}} &= \frac{(V_a M_a) - (V_b M_b)}{(V_a + V_b)} \\ &= \frac{((35.00 \text{ mL}) * (0.1000 \text{ M})) - ((15.00 \text{ mL}) * (0.1000 \text{ M}))}{(35.00 + 15.00) \text{ mL}} \\ &= 4.000 \times 10^{-2} \text{ M} \end{aligned}$$

EXAMPLE: Derive the titration curve for the titration of 35.00 mL of 0.1000 M $\text{HC}_2\text{H}_3\text{O}_2$ with 0.00, 15.00, 35.00, and 50.00 mL of 0.1000 M NaOH.

$$K_a = 1.75 \times 10^{-5} \text{M}$$

at 15.00 mL of NaOH added, $V_a * M_a > V_b * M_b$

$$[\text{HC}_2\text{H}_3\text{O}_2]_{\text{excess}} = 4.000 \times 10^{-2} \text{M}$$

$$\begin{aligned} [\text{C}_2\text{H}_3\text{O}_2^-] &= \frac{(V_b * M_b)}{(V_a + V_b)} \\ &= \frac{(15.00 \text{ mL})(0.1000 \text{ M})}{(35.00 + 15.00) \text{ mL}} = 3.000 \times 10^{-2} \text{M} \end{aligned}$$

EXAMPLE: Derive the titration curve for the titration of 35.00 mL of 0.1000 M $\text{HC}_2\text{H}_3\text{O}_2$ with 0.00, 15.00, 35.00, and 50.00 mL of 0.1000 M NaOH.

$$K_a = 1.75 \times 10^{-5} \text{M}$$

at 15.00 mL of NaOH added, $V_a * M_a > V_b * M_b$

$$[\text{HC}_2\text{H}_3\text{O}_2]_{\text{excess}} = 4.000 \times 10^{-2} \text{M} \quad [\text{C}_2\text{H}_3\text{O}_2^-] = 3.000 \times 10^{-2} \text{M}$$

$$\begin{aligned} [\text{H}_3\text{O}^+] &= K_a \frac{[\text{HC}_2\text{H}_3\text{O}_2]_{\text{excess}}}{[\text{C}_2\text{H}_3\text{O}_2^-]} \\ &= 1.75 \times 10^{-5} \text{ M} \frac{4.000 \times 10^{-2} \text{M}}{3.000 \times 10^{-2} \text{M}} = 2.33 \times 10^{-5} \text{ M} \end{aligned}$$

EXAMPLE: Derive the titration curve for the titration of 35.00 mL of 0.1000 M $\text{HC}_2\text{H}_3\text{O}_2$ with 0.00, 15.00, 35.00, and 50.00 mL of 0.1000 M NaOH.

$$K_a = 1.75 \times 10^{-5}\text{M}$$

at 15.00 mL of NaOH added, $V_a M_a > V_b M_b$

$$[\text{HC}_2\text{H}_3\text{O}_2]_{\text{excess}} = 4.000 \times 10^{-2}\text{M}$$

$$[\text{C}_2\text{H}_3\text{O}_2^-] = 3.000 \times 10^{-2}\text{M}$$

$$[\text{H}_3\text{O}^+] = 2.33 \times 10^{-5} \text{ M} \quad \text{pH} = 4.632$$

EXAMPLE: Derive the titration curve for the titration of 35.00 mL of 0.1000 M $\text{HC}_2\text{H}_3\text{O}_2$ with 0.00, 15.00, 35.00, and 50.00 mL of 0.1000 M NaOH.

$$K_a = 1.75 \times 10^{-5}\text{M}$$

at 35.00 mL of NaOH added,

$$V_a * M_a = V_b * M_b, \text{ equivalence point}$$

EXAMPLE: Derive the titration curve for the titration of 35.00 mL of 0.1000 M $\text{HC}_2\text{H}_3\text{O}_2$ with 0.00, 15.00, 35.00, and 50.00 mL of 0.1000 M NaOH.

$$K_a = 1.75 \times 10^{-5}\text{M}$$

at 35.00 mL of NaOH added,
 $V_a * M_a = V_b * M_b$, equivalence point

$$[\text{OH}^-] = \sqrt{\frac{K_w V_a M_a}{K_a (V_a + V_b)}}$$

EXAMPLE: Derive the titration curve for the titration of 35.00 mL of 0.1000 M $\text{HC}_2\text{H}_3\text{O}_2$ with 0.00, 15.00, 35.00, and 50.00 mL of 0.1000 M NaOH.

$$K_a = 1.75 \times 10^{-5}\text{M}$$

$$[\text{OH}^-] = \frac{1.00\text{e-}14\text{M}^2 * (35.00\text{mL})(0.1000\text{M})}{1.75\text{e-}5 \text{ M} * (35.00 + 35.00)\text{mL}}$$

$$= 5.34 \times 10^{-6}\text{M}$$

$$\sqrt{\quad}$$

EXAMPLE: Derive the titration curve for the titration of 35.00 mL of 0.1000 M $\text{HC}_2\text{H}_3\text{O}_2$ with 0.00, 15.00, 35.00, and 50.00 mL of 0.1000 M NaOH.

$$K_a = 1.75 \times 10^{-5}\text{M}$$

at 35.00 mL of NaOH added,
 $V_a M_a = V_b M_b$, equivalence point
 $[\text{OH}^-] = 5.34 \times 10^{-6} \text{ M}$

$$\text{pOH} = 5.272 \quad \text{pH} = 8.728$$

EXAMPLE: Derive the titration curve for the titration of 35.00 mL of 0.1000 M $\text{HC}_2\text{H}_3\text{O}_2$ with 0.00, 15.00, 35.00, and 50.00 mL of 0.1000 M NaOH.

$$K_a = 1.75 \times 10^{-5}\text{M}$$

at 50.00 mL of NaOH added,
 $V_b M_b > V_a M_a$, post equivalence point

EXAMPLE: Derive the titration curve for the titration of 35.00 mL of 0.1000 M $\text{HC}_2\text{H}_3\text{O}_2$ with 0.00, 15.00, 35.00, and 50.00 mL of 0.1000 M NaOH.

$$K_a = 1.75 \times 10^{-5}\text{M}$$

at 50.00 mL of NaOH added,

$V_b M_b > V_a M_a$, post equivalence point

$$[\text{OH}^-] = \frac{(V_b * M_b) - (V_a * M_a)}{(V_a + V_b)}$$

EXAMPLE: Derive the titration curve for the titration of 35.00 mL of 0.1000 M $\text{HC}_2\text{H}_3\text{O}_2$ with 0.00, 15.00, 35.00, and 50.00 mL of 0.1000 M NaOH. $K_a = 1.75 \times 10^{-5}\text{M}$

at 50.00 mL of NaOH added, $V_b * M_b > V_a * M_a$

$$[\text{OH}^-] = \frac{(V_b * M_b) - (V_a * M_a)}{(V_a + V_b)}$$

$$= \frac{((50.00\text{mL}) * (0.1000\text{M})) - ((35.00\text{mL}) * (0.1000\text{M}))}{(35.00 + 50.00)\text{mL}}$$

$$= 1.765 \times 10^{-2}\text{M}$$

EXAMPLE: Derive the titration curve for the titration of 35.00 mL of 0.1000 M $\text{HC}_2\text{H}_3\text{O}_2$ with 0.00, 15.00, 35.00, and 50.00 mL of 0.1000 M NaOH.

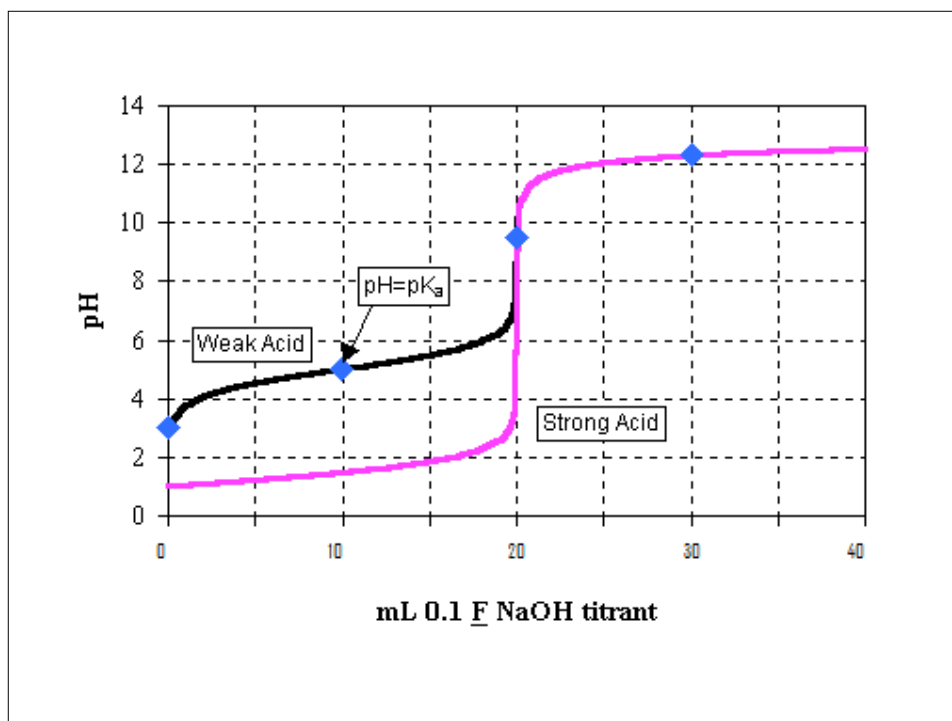
$$K_a = 1.75 \times 10^{-5} \text{M}$$

at 50.00 mL of NaOH added, $V_b * M_b > V_a * M_a$

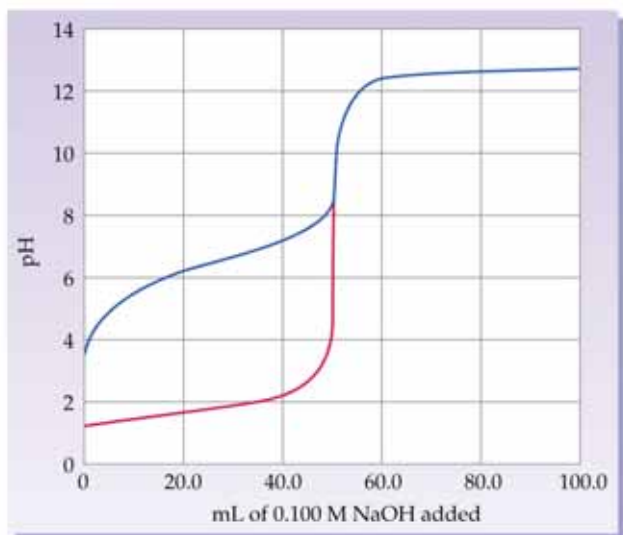
$$[\text{OH}^-] = 1.765 \times 10^{-2} \text{M}$$

$$\text{pOH} = 1.7533$$

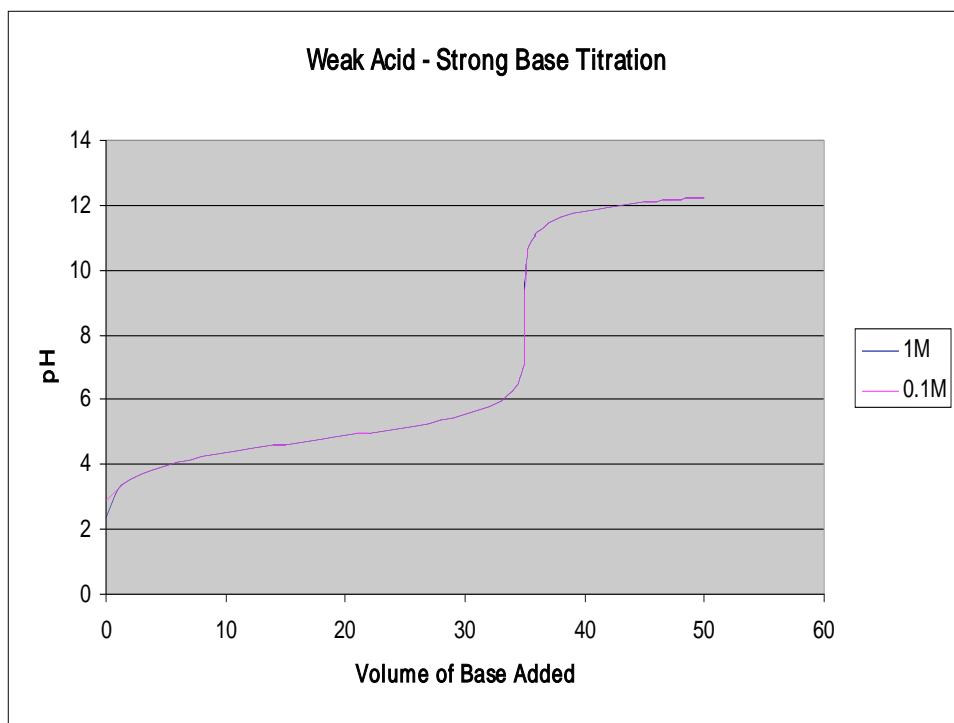
$$\text{pH} = 14.00 - 1.7533 = 12.25$$



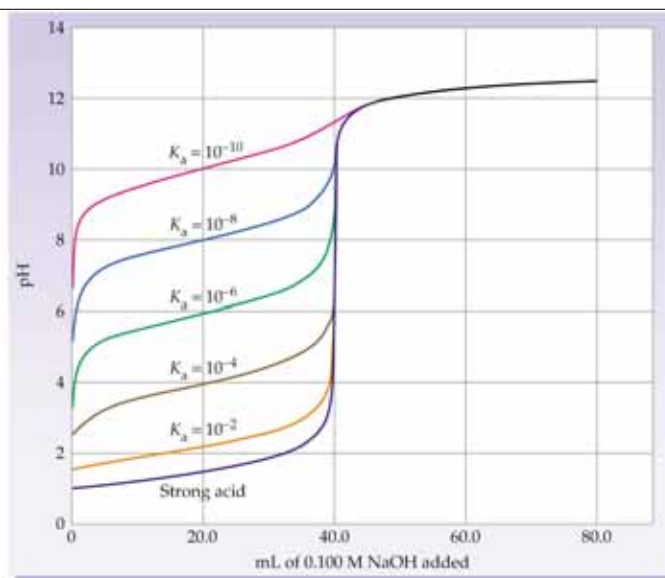
The following plot shows two titration curves, each representing the titration of 50.0 mL of 0.100 M acid with 0.100 M NaOH:



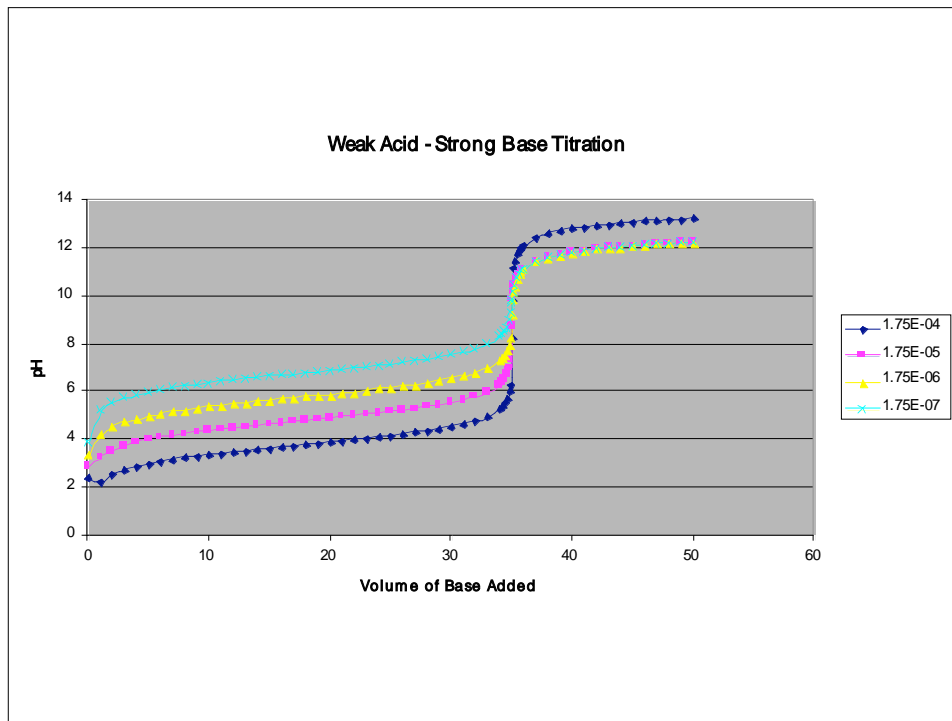
- Which of the two curves represents the titration of a strong acid, and which a weak acid?
- What is the approximate pH at the equivalence point for each of the acids?
- What is the approximate pK_a of the weak acid?



- The Effect of Concentration
- The pH of buffers is independent of dilution.
- pH at the equivalence point is lower for lower initial weak acid concentrations and initial pH is higher.

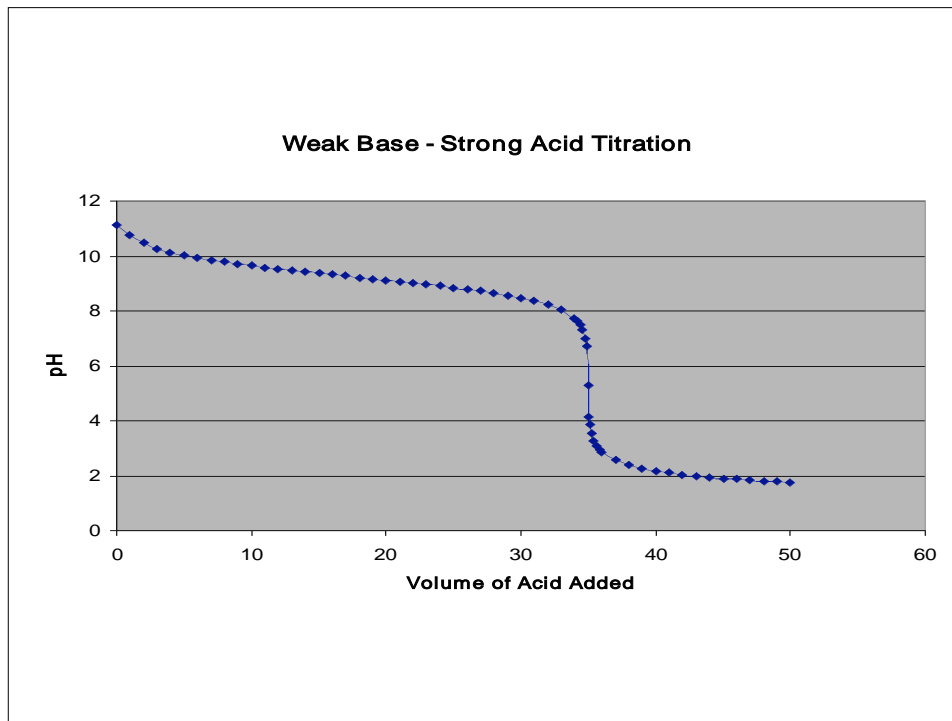


The pH change at the equivalence point decreases as the acid strength decreases



14.3 Titration of Weak Base with Strong Acid

Much the same shape, except the titration
would start at high pH and decrease as acid
is added

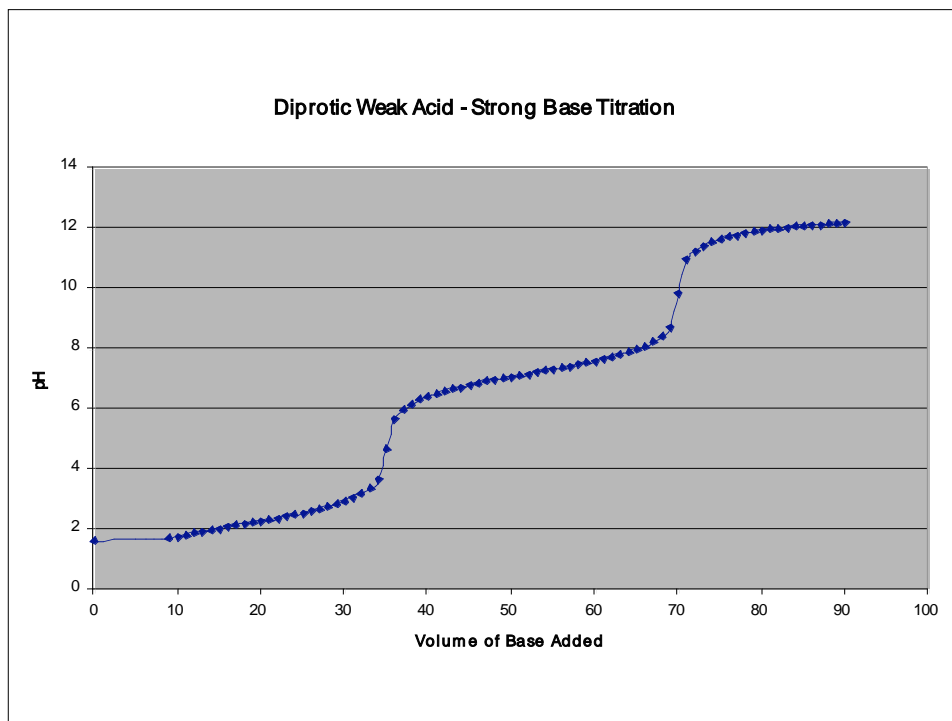


Diprotic Weak Acid - Strong Base Titration Curve

EXAMPLE: Derive the titration curve for the titration of 35.00 mL of 0.1000 M H_3PO_4 with 0.00, 15.00, 35.00, 50.00, 70.00, and 90.00 mL of 0.1000 M NaOH.

$$K_{a1} = 7.11 \times 10^{-3} \text{ M}$$

$$K_{a2} = 6.34 \times 10^{-8} \text{ M}$$



14.4 Solutions and Indicators for Acid/Base Titrations

- 1) Standard Solutions
- 2) Indicators

Standard Solutions

- Standard acid or base reagents are always strong acids or bases
- HCl, HClO₄, H₂SO₄, NaOH, and KOH
- They yield sharper end points.
- Weak reagents react incompletely.
- Nitric acid and hot sulfuric and perchloric yield undesirable ox/red side reactions.

Standard Solutions

- Preferred acids for neutralizations include HCl, HClO₄ and H₂SO₄.
- 1. All are stable to boiling.
 2. All may be prepared and standardized against a primary standard base

Standard Solutions

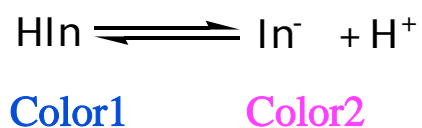
- Primary standards for acid standardization include
- Na_2CO_3
- Tris (THAM) – tris(hydroxymethyl)aminomethane
- Formula is $(\text{HOCH}_2)_3\text{CNH}_2$
- Sodium tetraborate decahydrate
- Formula is $\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$

Indicators

- Acid/Base Indicators
- pH Electrode

Indicators

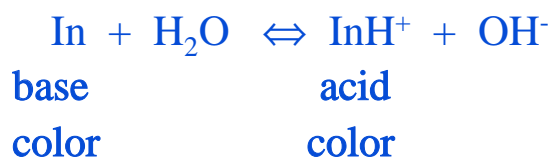
- Acid/Base Indicators
- Weak acids or bases
- The conjugate forms differs in color



$$\text{pH} = \text{pK}_{\text{a}_{\text{HIn}}} \pm 1$$

Indicators

Indicator Behavior



$$K_b = \frac{[\text{InH}^+][\text{OH}^-]}{[\text{In}]} \quad (2)$$

Indicators

Indicator Behavior

rearranging Equation (1) gives

$$\frac{[\text{In}^-]}{[\text{HIn}]} = \frac{K_a}{[\text{H}_3\text{O}^+]}$$

Indicators

Indicator Behavior

acid color shows when

$$\frac{[\text{In}^-]}{[\text{HIn}]} \ll \frac{1}{10} \quad \frac{[\text{H}_3\text{O}^+][\text{In}^-]}{[\text{HIn}]} = \frac{1}{10} * [\text{H}_3\text{O}^+] = K_a$$

base color shows when

$$\frac{[\text{In}^-]}{[\text{HIn}]} \geq \frac{1}{10} \quad \frac{[\text{H}_3\text{O}^+][\text{In}^-]}{[\text{HIn}]} = 10 * [\text{H}_3\text{O}^+] = K_a$$

Indicators

Indicator Behavior

acid color shows when

$$\text{pH} + 1 = \text{pK}_a$$

and base color shows when

$$\text{pH} - 1 = \text{pK}_a$$

Indicators

Indicator Behavior

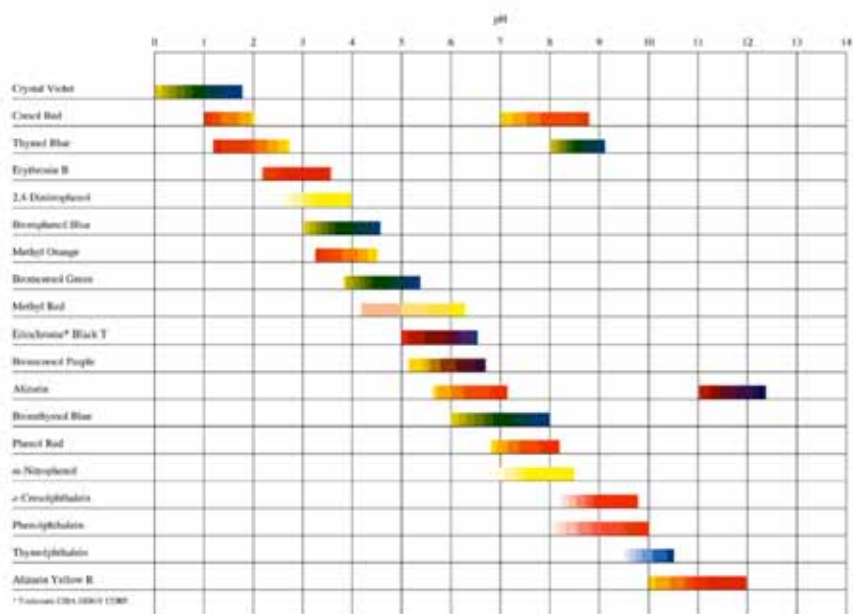
Color change range is

$$\text{pH} = \text{pK}_a \pm 1$$

Indicators

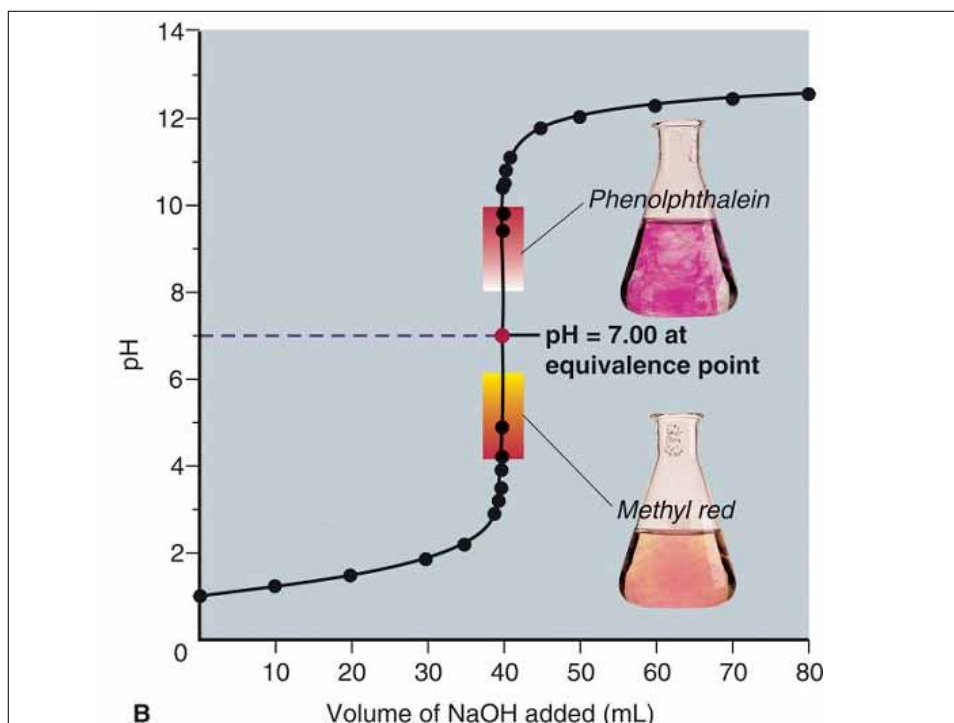
Indicator	pH transition range	color
Bromophenol Blue	6.2 - 7.6	yellow - blue
Methyl Orange	3.1 - 4.4	red - orange
Methyl Red	4.2 - 6.2	red - yellow
Bromothymol Blue	6.2 - 7.6	yellow - blue
Cresol Purple	7.6 - 9.2	yellow - purple
Phenolphthalein	8.3 - 10	colorless - red
Thymolphthaleine	9.3 - 10.5	colorless - blue
Alizerin Yellow GG	10 - 12	yellow - red

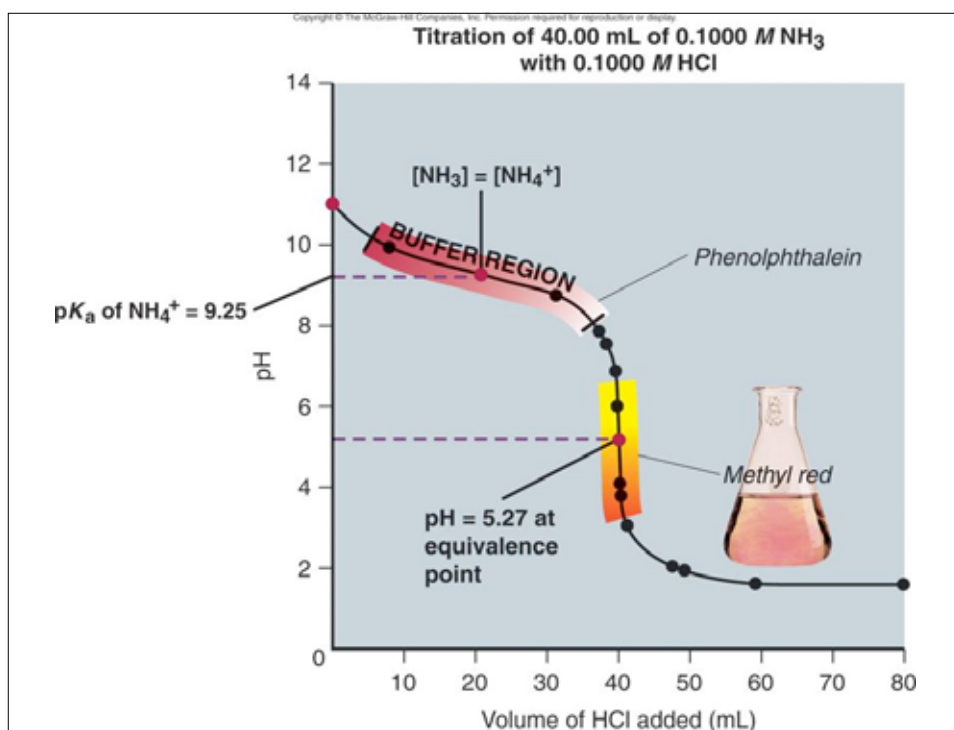
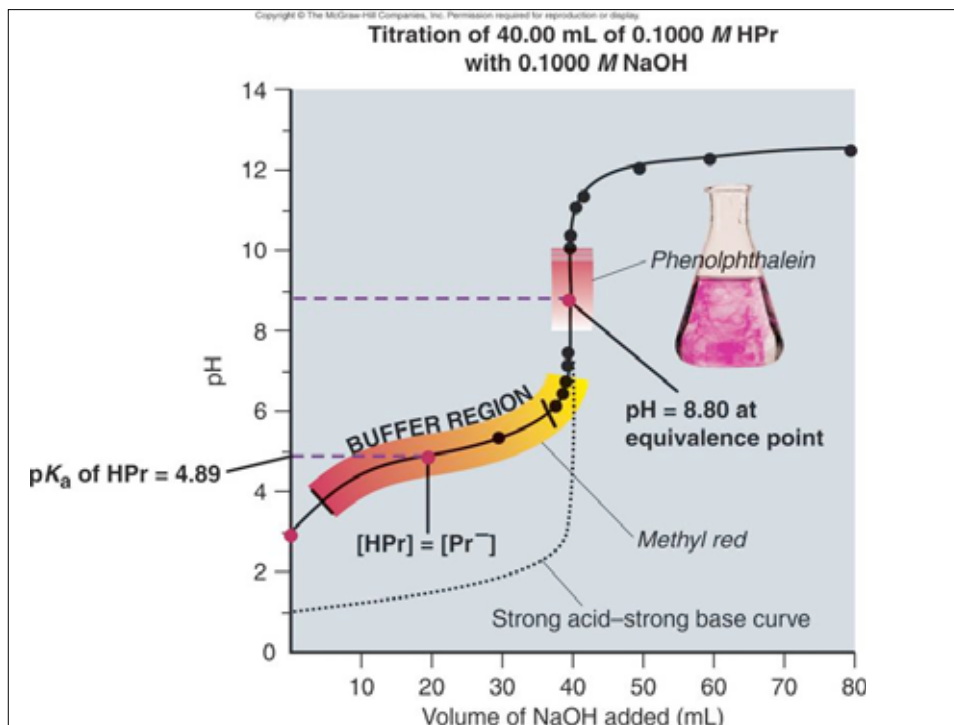
Common indicators.



Choosing the Proper Indicator

- Color change range should be in area where titration curve is most vertical



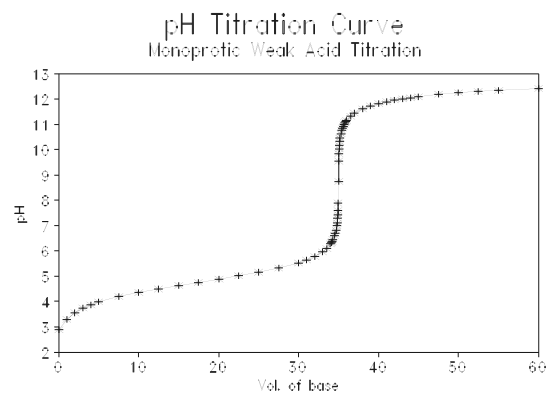


Indicators

Phenolphthalein

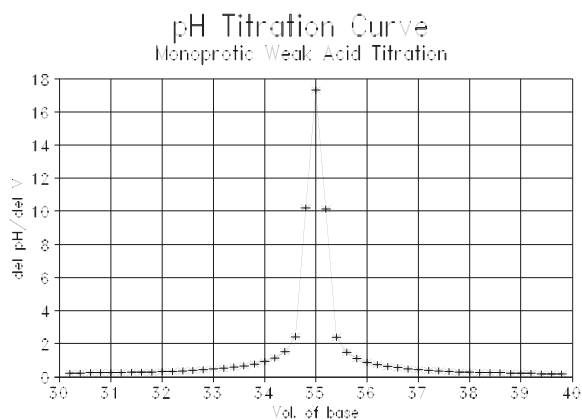


Finding the End Point with a pH Electrode



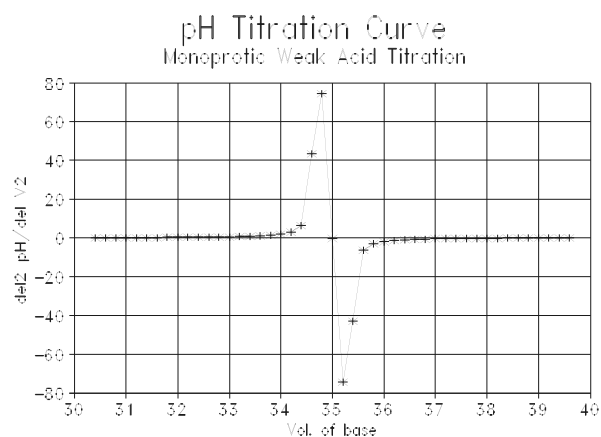
Titration curve for a weak acid - strong base titration.

Finding the End Point with a pH Electrode



First derivative of a weak acid- strong base titration

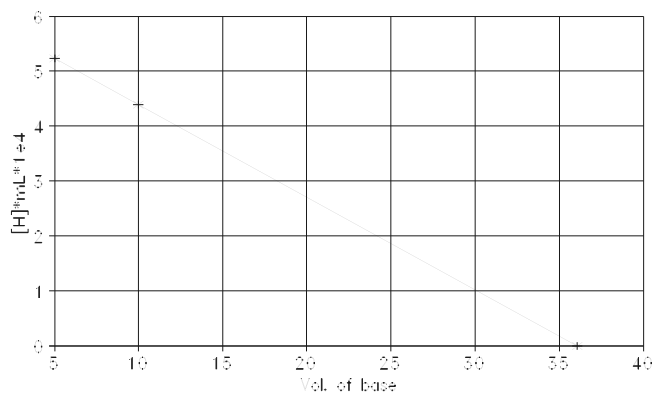
Finding the End Point with a pH Electrode



Second derivative of a weak acid - strong base titration

Using the Gran Plot to Estimate the End Point

Gran Plot
Acid-Base Titration



14.5 APPLICATION OF NEUTRALIZATION TITRATIONS

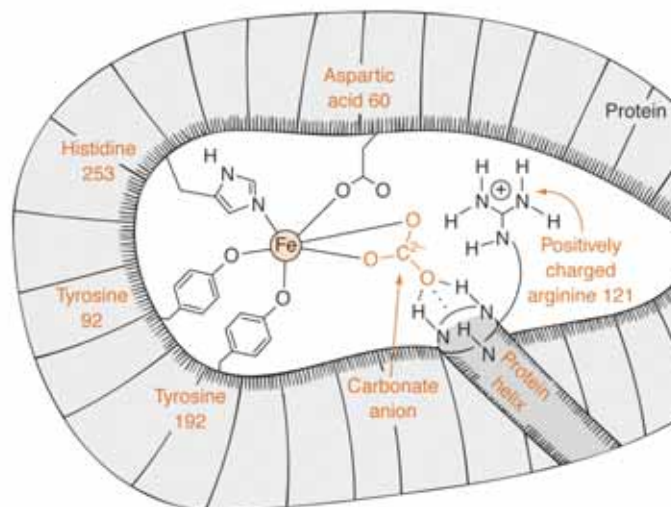
A. Neutralizations are widely used to determine the concentrations of analytes

B. Commonly analyzed species include acids, bases and any species that can be converted into an acid or base

14.5 APPLICATION OF NEUTRALIZATION TITRATIONS

A. Elemental analysis - Titrations can be performed for the analysis of N, S, C, Cl, Br, F, P and some other elements, by converting them to titratable forms. For example, the Keldahl analysis of nitrogen.

Figure 7.4



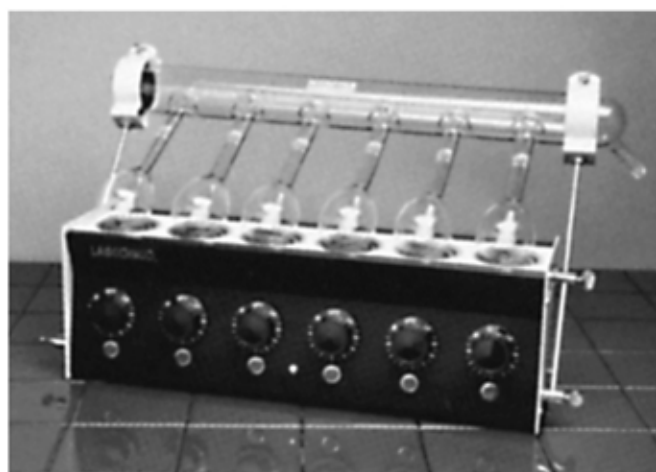
Kjeldahl Nitrogen Analysis

The sample is **digested** in boiling, concentrated sulfuric acid to which has been added some potassium sulfate using selenium-coated boiling chips as a catalyst.

Figure 7.2 Kjeldahl Nitrogen Analysis

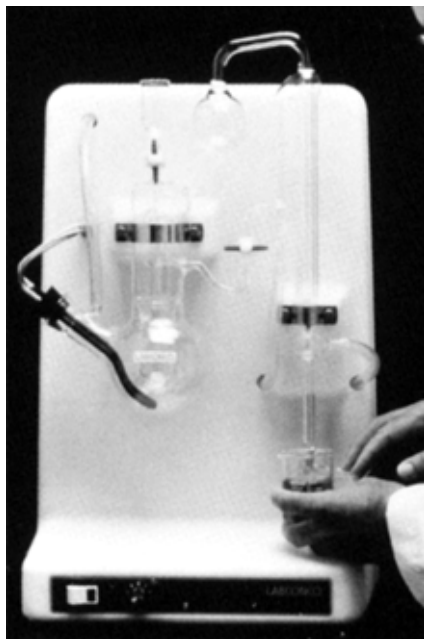


(a)



(b)

Figure 7.3
Modern
Kjeldahl
Nitrogen
Analysis



Kjeldahl Nitrogen Analysis

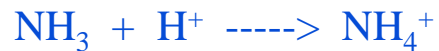
Kjeldahl digestion



The resulting NH_4^+ is neutralized with NaOH and the resulting $\text{NH}_{3(g)}$ is steam distilled into a flask containing an excess amount of a standardized HCl solution.

Kjeldahl Nitrogen Analysis

Kjeldahl digestion



The excess HCl is back titrated with a standard solution of NaOH. The number of moles of nitrogen in the sample equals the difference in the total number of moles of HCl and the number of moles of NaOH.

14.5 APPLICATION OF NEUTRALIZATION TITRATIONS

B. Determinations of inorganic substances -
Titrations can be performed to analyze for ammonia (via ammonium salts), nitrates, nitrites, carbonates, and other species.

Exercises

1. A solution containing 75.0 mL of 0.150 M strong acid, (HCl) is titrated with 75.0 mL of 0.300 M strong base (NaOH). What is the pH of the resulting solution?

- (a) 7.00
- (b) 13.2
- (c) 12.9

Exercises

2. Consider the titration of 25.0 mL of 0.150 M HBr with 0.0750 M NaOH. What is the volume of NaOH required to reach the equivalence point.

- (a) 25.0 mL
- (b) 50.0 mL
- (c) 12.5 mL

Exercises

3. A solution containing 50.0 mL of 0.1500 M $\text{C}_2\text{H}_3\text{COOH}$ is titrated with 0.1500 M NaOH. What is the pH of the solution after the addition of 25.0 mL of NaOH? pK_a for $\text{C}_2\text{H}_3\text{COOH}$ is 4.76

- (a) 4.76
- (b) 9.24
- (c) 1.30

Exercises

4. When an ant bites you, the reason it stings is because of an injection of formic acid. What is the pH when 50.0 mL of 0.1480 M formic acid (HCOOH) is titrated with 40.2 mL of 0.1841 M NaOH? pK_a of formic acid = 3.745

- (a) 2.42
- (b) 11.6
- (c) 11.7

Exercises

5. In a titration, 48.6 mL of 0.1280 M KOH is added to 45.0 mL of 0.1200 M hydroxyacetic($\text{HOCH}_2\text{CO}_2\text{H}$) acid. Calculate the pH of the solution. pK_a for $\text{HOCH}_2\text{CO}_2\text{H} = 3.831$.

- (a) 2.06
- (b) 2.38
- (c) 11.94

Exercises

6. In a titration, 25.0 mL of 0.154 M nitric acid (HNO_3) was added to 45.0 mL of 0.213 M ammonia(NH_3). What is the pH at this point? $\text{PK}_a \text{ NH}_3 = 9.244$

- (a) 9.42
- (b) 4.58
- (c) 9.07

Exercises

7. The weak base trimethyl amine ((CH₃)₃N) was titrated with hydrochloric acid (HCl). The titration required 25.30 mL of 0.150 M HCl to react with 35.15 mL of 0.108 molar trimethyl amine. What is the pH of the solution? pK_a for (CH₃)₃N = 9.80

- (a) 5.50
- (b) 2.70
- (c) 8.50

Exercises

8. A titration was performed by adding 35.85 mL of 0.1350 molar nitric acid (HNO₃) to 55.80 mL of 0.08673 molar potassium hydroxide (KOH). What is the pH of the solution?

- (a) 1.28
- (b) 7.00
- (c) 3.28

Exercises

9. 25.0 mL of 0.100 molar triethyl amine $((\text{CH}_3\text{CH}_2)_3\text{N})$ is titrated with 0.0500 molar hydrochloric acid

(HCl). The following data was obtained:

mL HCl added	pH
0.0	11.86
25.0	10.71
50.0	6.79

What is pK_a for the triethylammonium ion.

- (a) 11.86
- (b) 10.71
- (c) 6.79

Exercises

10. A flask containing 36.8 mL of 0.138 molar sodium acetate $(\text{NaC}_2\text{H}_3\text{O}_2)$ is titrated with 24.9 mL of 0.128 molar nitric acid (HNO_3) . Calculate the pH of the solution. pK_a for acetic acid = 4.757

- (a) 4.53
- (b) 9.47
- (c) 4.98

Exercises

11. In the titration of 30.0 mL of a 0.200 molar solution of a hypothetical compound NaH_2M , what is the pH of the solution after the addition of 0.00 mL of 0.100 molar NaOH ? For H_3M , $\text{pK}_{\text{a}1} = 3.00$, $\text{pK}_{\text{a}2} = 6.00$, $\text{pK}_{\text{a}3} = 9.00$

- (a) 1.85
- (b) 4.50
- (c) 3.35

Exercises

12. In the titration of 30.0 mL of a 0.200 molar solution of a hypothetical compound NaH_2M , what is the pH of the solution after the addition of 30.0 mL of 0.100 molar NaOH ? For H_3M , $\text{pK}_{\text{a}1} = 3.00$, $\text{pK}_{\text{a}2} = 6.00$, $\text{pK}_{\text{a}3} = 9.00$

- (a) 6.00
- (b) 3.00
- (c) 7.50

Exercises

13. In the titration of 30.0 mL of a 0.200 M solution of a hypothetical compound NaH_2M , what is the pH of the solution after the addition of 60.0 mL of 0.100 M NaOH? For H_3M , $\text{pK}_{\text{a}1} = 3.00$, $\text{pK}_{\text{a}2} = 6.00$, $\text{pK}_{\text{a}3} = 9.00$

- (a) 6.00
- (b) 9.00
- (c) 7.5

Exercises

14. In the titration of 30.0 mL of a 0.200 molar solution of a hypothetical compound NaH_2M , what is the pH of the solution after the addition of 70.0 mL of 0.100molar NaOH? For H_3M , $\text{pK}_{\text{a}1} = 3.00$, $\text{pK}_{\text{a}2} = 6.00$, $\text{pK}_{\text{a}3} = 9.00$

- (a) 8.30
- (b) 5.30
- (c) 12.0

Exercises

15. In the titration of 30.0 mL of a 0.200 molar solution of a hypothetical compound NaH_2M , what is the pH of the solution after the addition of 120 mL of 0.100 molar NaOH ? For H_3M , $\text{pK}_{\text{a}1}=3.00$, $\text{pK}_{\text{a}2}=6.00$, $\text{pK}_{\text{a}3}=9.00$

- (a) 7.50
- (b) 9.15
- (c) 8.8

Exercises

16. In the titration of 30.0 mL of a 0.200 molar solution of a hypothetical compound NaH_2M , what is the pH of the solution after the addition of 180 mL of 0.100 molar NaOH ? For H_3M , $\text{pK}_{\text{a}1} = 3.00$, $\text{pK}_{\text{a}2} = 6.00$, $\text{pK}_{\text{a}3} = 9.00$

- (a) 9.15
- (b) 8.73
- (c) 12.4