In a titration a solution of accurately known concentration is added gradually to another solution of unknown concentration until the chemical reaction between the two solutions is complete.

**Equivalence point** – the point at which the reaction is complete

**Indicator** – substance that changes color at (or near) the equivalence point

A student was asked to determine the molecular weight of an unknown acid. A 0.1508 gram sample was dissolved in 25 mL of water and titrated with 0.105 M NaOH solution. An indicator was used to determine the endpoint which occurred after 12.32 mL of the NaOH solution was added. What is the molecular weight of the acid?

\[
\text{Moles of acid} = \frac{0.1508 \text{ gram}}{0.105 \text{ mole } / \text{L } \times 0.01232 \text{L}} = 117 \text{ g/mol}
\]

A 95 g/mol  B 102 /mol  C 117 g/mol  D 126 g/mol  E 152 g/mol

In doing this you will learn lots of practical chemistry.

**Titration Curves** – You should know how to draw them.

That means – you should understand what is happening: What is reacting with what? What equations do you use? In doing this you will learn lots of practical chemistry.

1. The initial Situation
2. The mid-point Situation
3. Intermediate-points (just before & after the mid-point)
4. The end-point Situation
5. Beyond the end point

\[ M_1V_1 = M_2V_2 \]
### Weak Acid-Strong Base Titration

**Problem**

Suppose you were carrying out a titration of a 20.0 mL solution of 0.200 M CH₂ClCO₂H using 0.200 M NaOH solution. Calculate the pH of the solution at the following points.

- **i.** Beginning
- **ii.** After 10 mL NaOH added
- **iii.** After 20 mL added
- **iv.** After 30 mL added

\[
[A^{-}] + [H^{+}] = K_{a} = \frac{[x][A^{-}]}{[A]} = 1.34 \times 10^{-3}
\]

\[x = \frac{1.34 \times 10^{-3}}{[A]} = 1.63 \times 10^{-2}
\]

Equal so \(pH = pK_{a}\)

\[\text{a} \ 1.4 \quad \text{b} \ 1.8 \quad \text{c} \ 2.9 \quad \text{d} \ 3.6 \quad \text{e} \ 7 \quad \text{f} \ 7.9 \quad \text{g} \ 8.1\]

### Weak Base-Strong Acid Titration

**Problem**

At equivalence point (\(pH > 7\)):

\[
CH_{3}COO^{-}(aq) + H_{2}O(l) \rightleftharpoons OH^{-}(aq) + CH_{3}COOH(aq)
\]

End Point Situation \(pH = 8.72\)

\[
CH_{3}COO^{-}(aq) + H_{2}O(l) \rightleftharpoons OH^{-}(aq) + CH_{3}COOH(aq)
\]

At the End Point situation \(pH = 5.28\)

\[
\text{B & BH}^{+}
\]

Salt of a weak base (conjugate) problem

\[0.10 \text{ M HCl added to 25 mL of 0.10 M NH}_{3}\]

\[\text{pOH} = 6.1 \quad \text{pH} = 7.9\]

\[\text{a} \ 3.6 \quad \text{b} \ 6.1 \quad \text{c} \ 6.9 \quad \text{d} \ 7 \quad \text{e} \ 7.2 \quad \text{f} \ 7.9 \quad \text{g} \ 8.1\]
Suppose you were carrying out a titration of a 20.0 mL solution of 0.200 M NH₃ using 0.200 M HCl solution. Calculate the pH of the solution at the following points.

i. Beginning ii. after 10 ml HCl added iii. After 20 ml added iv. After 30 ml added

\[
\text{[HA]}\frac{[OH^{-}]}{[A^{-}]^2} = K_b = \frac{[x][x]}{[0.2-x]} = 1.8 \times 10^{-5} \quad x = [OH^{-}] = 1.9 \times 10^{-3}
\]

Equal so \( pOH = pK_b \)

a. 4.7  b. 5.0  c. 5.1  d. 9.3  e. 11.3  f. 11.6

Weak Base-Strong Acid Titration

\( \text{H}^+(aq) + \text{NH}_3(aq) \rightarrow \text{NH}_4^+(aq) \quad \text{NH}_4^+ + \text{H}_2\text{O} \rightarrow \text{NH}_3(aq) + \text{H}^+(aq) \)

At equivalence point (pH < 7):

\( \text{NH}_4^+ + \text{H}_2\text{O} \rightarrow \text{NH}_3 + \text{H}^+ \) Salt of a weak base (conj. acid) problem

\[
\text{HCl is dominant species}
\]

The first 20 mL is used up.
10 ml diluted to 50 ml so M of HCl is .04M

\( -\log(0.04) = \text{pH} = 1.4 \)

a. 0.7  b. 0.9  c. 1.4  d. 2.7  e. 4.7

Acid/Base Indicators:
substances that change color in a specific pH range and is used as a visual index of an approximate pH range.

Phenolphthalein

Acid Form = HIn  Conj. Base Form = In⁻

Colorless  Pink
The useful pH ranges for several common indicators

The pH Curve for Titration of HCl with NaOH

The titration curve is so steep between pH 4 and 10 that either indicator will give a satisfactory result.

The pH at the end point is 8.72. Methyl red will give a very diffuse color change around 40 mL rather than the correct 50 mL. Phenolphthalein will work well.

pH Curve for Titration of Acetic Acid with NaOH

Phosphoric acid & Phosphate Groups are Important

Industrially = About 10 million tons of phosphoric acid, $H_3PO_4$, are produced in the United States each year. Mainly used as fertilizer.

Biologically = ATP, ADP, DNA, RNA

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Suppose you had a liter solution of 1.0 M phosphoric acid. How many moles of solid NaOH would need to add to the solution to make a pH 7.0 buffer?

\[
\text{H}_3\text{PO}_4 \quad \text{pK}_{a1} = 2.12; \quad \text{pK}_{a2} = 7.21; \quad \text{pK}_{a3} = 12.32
\]

\[
K_A = \frac{[H^+][A^-]}{[HA]} \quad \text{Or} \quad \text{pH} = \text{pK}_a + \log \frac{[A^-]}{[HA]}
\]

\[
10^{-7.2} = \frac{[10^{-7}][A^-]}{[HA]} \quad \text{Or} \quad 7.0 = 7.2 + \log \frac{[A^-]}{[HA]}
\]

\[
\frac{[A^-]}{[HA]} = 10^{-2} = .63 \quad \frac{x}{1-x} = .63 \quad x = .387 \text{ mole} \quad + \text{1 mole to make H}_2\text{PO}_4^{-1}
\]

The strange one - Carbonic Acid

\[
\text{H}_2\text{CO}_3 \quad \text{K}_a \quad \text{pK}_a
\]

\[
\text{H}_2\text{CO}_3(aq) \quad \text{H}_2\text{CO}_3^{-1}(aq) \quad + \quad \text{H}^+ \quad 4.3 \times 10^{-7} \quad 6.4
\]

\[
\text{HCO}_3^{-1} \quad \text{CO}_3^{2-} \quad + \quad \text{H}^+ \quad 4.8 \times 10^{-11} \quad 10.3
\]

\[
\text{CO}_2(aq) \quad \text{H}_2\text{CO}_3(aq) \quad + \quad \text{H}^+ \quad 1.7 \times 10^{-3} \quad 6.4
\]

\[
\text{H}_2\text{CO}_3(aq) \quad \text{H}_2\text{CO}_3^{-1}(aq) \quad + \quad \text{H}^+ \quad 2.5 \times 10^{-4} \quad 3.6
\]

What is the pH of a 0.1 M solution of NaHCO3?

\[
\text{HCO}_3^{-1} + \text{H}_2\text{O} \quad \text{H}_2\text{CO}_3^{-1} + \text{OH}^- \quad \text{K}_b \quad \text{pK}_b
\]

\[
[\text{HCO}_3^{-1}] = \frac{2.3 \times 10^{-8}}{[1-x]} \quad \text{[OH]}^- = 4.8 \times 10^{-5}
\]

\[
\text{pOH} = 4.3 \quad \text{pH} = 9.7
\]