Strong Acid-Weak Base Titrations

Qualitative Analysis

strong acid + weak base → acidic salt + water

$HCl_{(aq)} + NH_{3\,(aq)} \rightarrow NH_4Cl_{(aq)} + H_2O_{(l)}$

$HCl_{(aq)} + NH_{3\,(aq)} \rightarrow NH_4^+{}_{(aq)} + Cl^-{}_{(aq)} + H_2O_{(l)}$

Cl- does not have basic properties,
but $\mathbf{NH_4^+}$ is a weak acid
∴ solution will be slightly acidic
at the equivalence point.

Quantitative Analysis

**Remember,
all titrations need to be analyzed in two steps:**

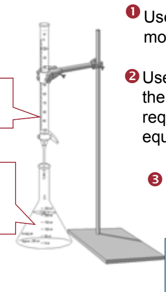
1. As **stoichiometry** problems:
   - How many **MOLES** of acid/base are in the solution? Which one is in excess, and how will that affect pH?
2. As **equilibrium** problems
   - In the case of **weak** acids or bases, what **CONCENTRATION** of acid/base will dissociate? This determines pH.

Example 3. $HCl_{(aq)} + NH_{3\,(aq)} \rightarrow NH_4Cl_{(aq)} + H_2O_{(l)}$

a) Use stoichiometry to calculate the volume of HCl that will be required to react completely with the sample of NH_3.

Titrant:
0.100 mol/L HCl

Sample:
20.0 mL of
0.100 mol/L NH_3



❶ Use n = c × v to find moles of NH_3.

n = 0.100 mol/L x 20.00 mL
= 2.00 mmol

❷ Use mole ratio to find the amount of HCl required to reach the equivalence point.

$n_{HCl} = 2.00 \text{ mmol } NH_3 \times \frac{1 \text{ mol HCl}}{1 \text{ mol } NH_3}$
= 2.00 mmol HCl

❸ Use the concentration to convert the amount of HCl to a volume.

$\mathbf{V_{HCl} = \frac{2.00 \text{ mmol}}{0.100 \text{ mol/L}} = 20.0 \text{ mL}}$

The reaction will have reached its equivalence point once 20.0 mL of acid has been added.

b) Before the titrant is added, what is the pH of the solution?

❶ **Stoichiometry**: How many moles of base are present?

No acid has been added, therefore no base has reacted yet. The solution contains the original amount of NH_3.

❷ Consider all species in the solution. Which one controls the pH?

$NH_{3\,(aq)}$, NH_4^+ (aq), $OH^-{}_{(aq)}$, $H_2O_{(l)}$

Even though NH_3 is a weak base, its K_b is much greater than the K_w of water ∴ it controls the pH:

$$\mathbf{NH_{3\,(aq)} + H_2O_{(l)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-{}_{(aq)}}$$

❸ **Equilibrium**: What is the concentration of OH^-?

NH_3 is a weak base ∴ it only partially ionizes. Use the K_b to find the concentration of OH^- at equilibrium.

	NH_3	H_2O	NH_4^+	OH^-
I	0.100	-	0	0
C	- x	-	+ x	+ x
E	0.100 – x	-	x	x

$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$ sub in $1.8 \times 10^{-5} = \frac{(x^2)}{(0.100-x)} \cong \frac{(x^2)}{(0.100)}$

$0.100/K_b > 100$ ∴ use simplifying assumption

$\mathbf{x \cong 1.3 \times 10^{-3} \text{ mol/L} = [OH^-]}$

x/0.300 = 0.43% ∴ assumption was valid

❹ Calculate pOH and pH

$pOH = -\log [OH^-]$

$= -\log (1.3 \times 10^{-3})$

$pOH = 2.88$

$pH = 14.00 - pOH$

$= 14.00 - 2.88$

pH = 11.11

Before any acid is added, the NH_3 solution has a pH of **11.11.**

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c) <u>At the equivalence point</u>: After 20.0mL of titrant is added to the sample, what is the pH?

❶ **Stoichiometry**: How many moles of acid/base are present?

At the equivalence point, exactly 3.00 mmol of HCl has been added, which reacts with the 3.00 mmol of base initially present.

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❷ Consider all species in the solution. Which one controls the pH?

All the H^+ and OH^- have been consumed.

Entities in solution: $NH_4^+{}_{(aq)}$, $Cl^-{}_{(aq)}$, and $H_2O_{(l)}$

NH_4^+ is a weak acid and CAN affect pH. It dissociates:

$NH_4^+ \rightleftharpoons H^+ + NH_3$

Cl^- does not affect pH

To find the K_a of this weak conjugate base, use the K_b of its acid:

$K_w = K_a \times K_b$

$K_a = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}}$

$K_a = 5.6 \times 10^{-10}$

Even though the K_a for NH_4^+ is small, it is still much larger than the K_w for water.

Therefore, the NH_4^+ will control the pH.

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❸ **Equilibrium**:
What is the concentration of H^+?

$[NH_4^+] = \frac{3.00 \text{ mmol}}{20.0 \text{ mL} + 20.0 \text{ mL}} = 0.0750 \text{ mol/L}$

	NH_4^+	H^+	NH_3
I	0.0750	0	0
C	- x	+ x	+ x
E	0.0750 – x	x	x

$K_a = \frac{[H^+][NH_3]}{[NH_4^+]}$

$5.6 \times 10^{-10} = \frac{(x^2)}{(0.0750 - x)} \cong \frac{(x^2)}{(0.0750)}$

$0.0750/K_a > 100$ ∴ use simplifying assumption

$x \cong 6.5 \times 10^{-6}$ mol/L = $[H^+]$

$x/0.0750 = 0.0087\%$ ∴ assumption was valid

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❹ Calculate pH

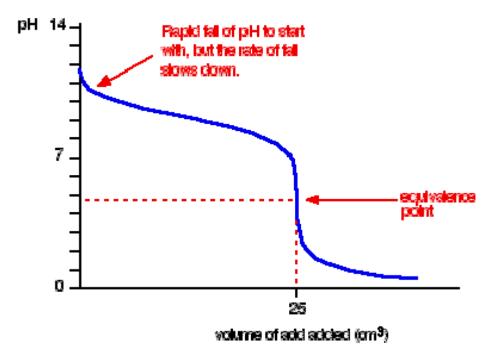
$pH = -\log [H^+]$

$= -\log (6.5 \times 10^{-6})$

pH = 5.19

In general, at the equivalence point of a titration between a strong acid and a weak base, the pH will be **slightly acidic**.

A typical pH curve for a strong acid-weak base titration



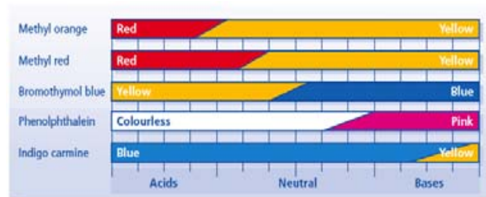
Learning Checkpoint

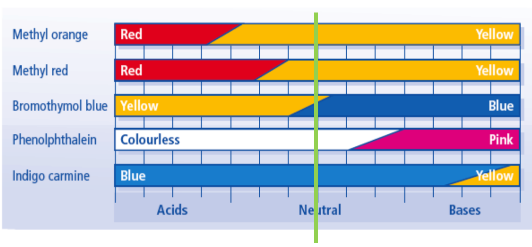
- In cases where either a weak acid or a weak base are involved in a titration, the pH will NOT be 7.00 at the equivalence point.

Acid-Base indicators are used to signal the endpoint of your titration reaction.

Ideally, you want your visual endpoint to match closely with the equivalence point of the reaction.

- Choose an indicator that changes its colour in the target pH range



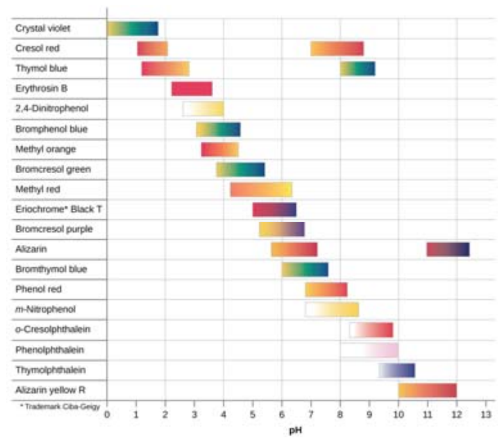


Match the indicator with the type of titration it is best suited for:

A. Strong acid - Strong base

B. Weak acid – Strong base

C. Strong acid – Weak base



Homework

- Pg. 557 #3, 7
- Worksheet: Weak Acids and Weak Bases