CONCEPT: WEAK BASE-STRONG ACID TITRATIONS

Whenever you titrate a **WEAK** Species with a **STRONG** species you use an _____ CHART.

The following can be used as the roadmap for determining the pH for a Weak Base-Strong Acid Titration.

Equivalence Volume (V_e)

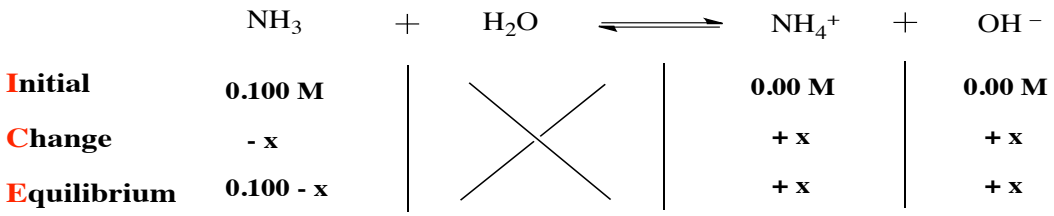
Calculate the equivalence volume, V_e, in order to determine the volume of titrant required to reach the equivalence point.

□ The titration of 100.0 mL of 0.100 M NH_3 with 0.20 M HCl

Before any Strong Acid is added

Before any of the strong acid titrant is added we only have a weak base initially.

□ The titration of 100.0 mL of 0.100 M NH_3 with 0.00 mL of 0.20 M HCl



$$K_b = \frac{x^2}{[\ ]_0 - x} \quad \therefore\ x = [OH^-] \quad \therefore\ pOH = -\log[OH^-] \quad \therefore\ pH = 14 - pOH$$

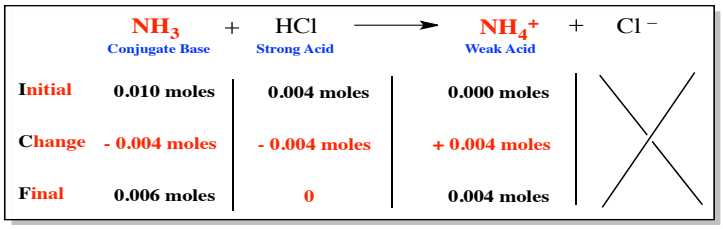
$$\frac{[\ ]_0}{K_b} > 500 \qquad \textbf{5\% Approximation Method}$$

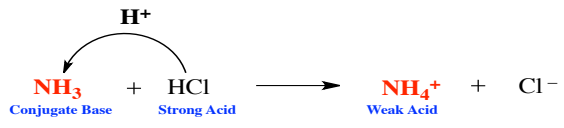
Before Equivalence Point

Once our acid and base begin to mix we use an ICF Chart to determine the pH.

□ The titration of 100.0 mL of 0.100 M NH_3 with 20.0 mL of 0.20 M HCl

moles = Liters × Molarity





- **Weak acid** and **Conjugate base** will be present at the end.
- Use the Henderson Hasselbalch Equation.

$$pH = pKa + \log\left(\frac{\text{Conjugate Base}}{\text{Weak Acid}}\right)$$

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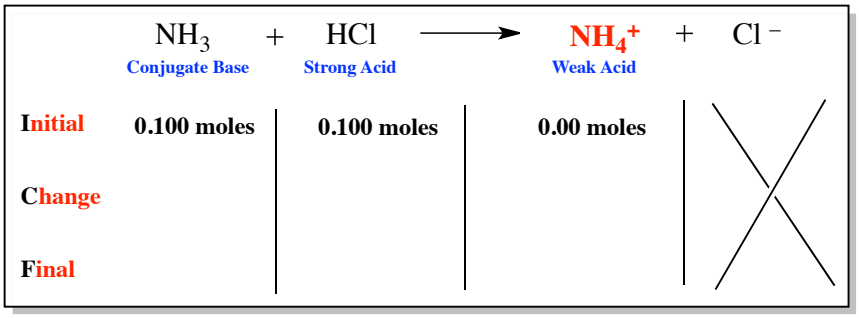
At Equivalence Point

At the equivalence point of a weak base–strong acid titration the solution is _____ and pH ___ 7.

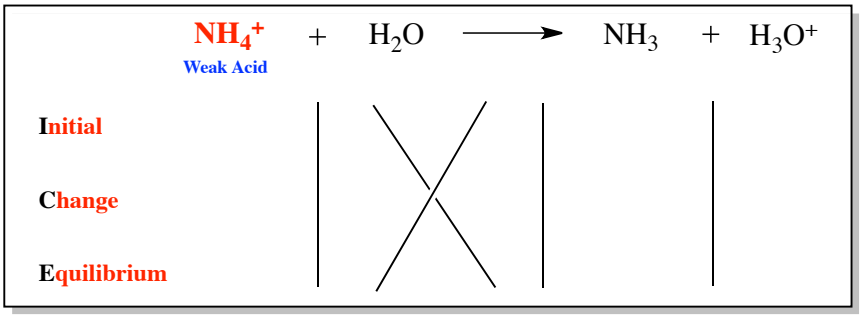
□ The titration of 100.0 mL of 0.100 M NH_3 with 50.0 mL of 0.200 M HCl



At the Equivalence Point

- Only **Weak Acid** will be present at the end.
- Use an ICE Chart to find pH.

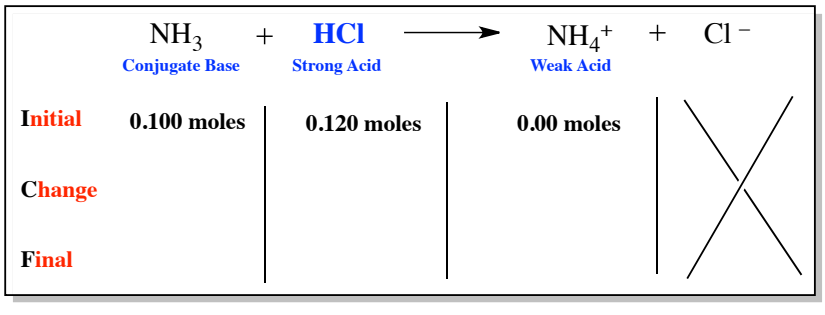
$$[WA] = \frac{\text{moles left}}{\text{Total Liters}} \quad \therefore \quad K_a = \frac{x^2}{[WA]}$$



After Equivalence Point

After the equivalence point of a weak base–strong acid titration we will have excess strong acid remaining.

□ The titration of 100.0 mL of 0.100 M NH_3 with 60.0 mL of 0.200 M HCl



After the Equivalence Point

- **Strong Acid** will be present at the end.

$$[SA] = \frac{\text{moles left}}{\text{Total Liters}} \quad \therefore \quad pH = -\log[SA]$$

PRACTICE: WEAK BASE-STRONG ACID TITRATIONS CALCULATIONS

EXAMPLE: Consider the titration of 50.0 mL of 0.150 M CH_3NH_2 ($K_b = 4.4.x\ 10^{-4}$) with 75.0 mL of 0.200 mL of 0.200 M HCl. Calculate the pH.

PRACTICE: Calculate the pH of the solution resulting from the mixing of 75.0 mL of 0.100 M $NaC_2H_3O_2$ and 75.0 mL of 0.150 M $HC_2H_3O_2$ with 0.0040 mole of $HClO_4$. The K_a value of acetic acid is 1.8×10^{-5}.