• This type of titration has the \_\_\_\_\_ as the titrate and the \_\_\_\_ as the titrant.

□ Recall, when a weak species reacts with a strong species use an ICF (I\_\_\_\_\_, C\_\_\_\_, F\_\_\_\_) Chart.

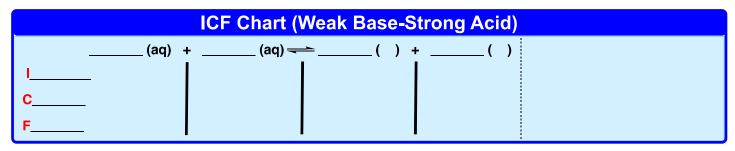
# **Before the Equivalence Point**

• In this part of the titration the moles of weak base is \_\_\_\_\_ the moles of strong acid.

□ As the Strong Acid neutralizes the Weak Base, some \_\_\_\_\_\_ is formed.

**EXAMPLE:** Calculate the pH of the solution resulting from the titration between 25.0 mL of a 0.100 M HClO<sub>3</sub> and 50.0 mL of a 0.100 M NH<sub>3</sub>. ( $K_b$  of NH<sub>3</sub> is 1.75 x 10<sup>-5</sup>).

Use STEPS 1 to 3 to setup the ICF Chart.



**STEP 4:** The Henderson-Hasselbalch Equation is used for a \_\_\_\_\_ to find the pH of a solution.

□ Using the FINAL ROW, use the moles of the \_\_\_\_\_ and \_\_\_\_ to find the pH.

# Henderson-Hasselbalch Equation $pH = pKa + log \frac{[CB]}{[WA]}$

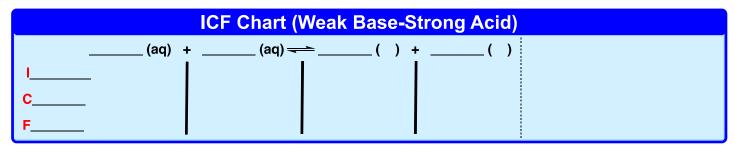
CONCEPT: TITRATIONS: WEAK BASE-STRONG ACID
<b>PRACTICE:</b> Calculate the pH of the solution resulting from the mixing of 75.0 mL of 0.100 M NaC <sub>2</sub> H <sub>3</sub> O <sub>2</sub> and 75.0 mL of 0.30 M HC <sub>2</sub> H <sub>3</sub> O <sub>2</sub> with 0.0040 moles of HBr.
<b>PRACTICE:</b> In order to create a buffer 7.321 g of potassium lactate is mix with 550.0 mL of 0.328 M lactic acid, $HC_3H_5O_3$ . What is the pH of the buffer solution after the addition of 300.0 mL of 0.100 M hydrobromic acid, HBr? The Ka of $HC_3H_5O_3$ is 1.4 x 10 <sup>-4</sup> .

## At the Equivalence Point

- In this part of the titration the moles of weak acid is \_\_\_\_\_ the moles of strong base.
  - ☐ The weak base and strong acid have been \_\_\_\_\_ and only the weak acid (conjugate acid) remains.

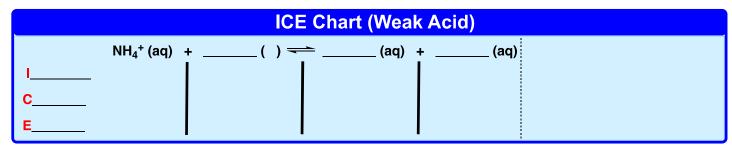
**EXAMPLE:** Calculate the pH of the solution resulting from the titration between 25.0 mL of a 0.100 M HClO $_3$  and 50.0 mL of a 0.050 M NH $_3$ . ( $K_b$  of NH $_3$  is 1.75 x 10 $^{-5}$ ).

Use STEPS 1 to 3 to setup the ICF Chart.



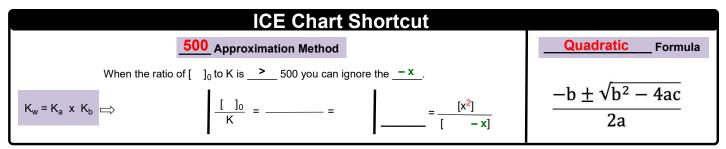
- STEP 4: Using the FINAL ROW, determine the concentration of the weak acid (conjugate acid).
  - □ Divide its final \_\_\_\_\_ by the total volume used in the chemical reaction.

STEP 5: If all that is left is a weak species then set up an ICE Chart that has it reacting with \_\_\_\_\_\_

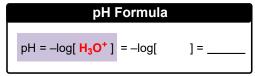


STEP 6: Using the EQUILIBRIUM ROW, setup the equilibrium constant expression with \_\_\_\_\_ and solve for \_\_\_\_.

□ Check if a shortcut can be utilized to avoid the \_\_\_\_\_ formula.



**STEP 7:** The \_\_\_\_\_ variable will equal [ ] and can be used to solve pH.



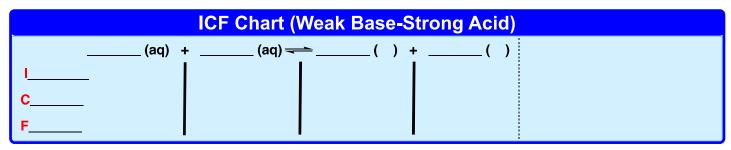
**PRACTICE:** Consider the titration of 100.0 mL of 0.100 M CH<sub>3</sub>NH<sub>2</sub> with 0.250 M HNO<sub>3</sub> at the equivalence point. What would be the pH of the solution at the equivalence point? The  $K_b$  of CH<sub>3</sub>NH<sub>2</sub> is 4.4 x 10<sup>-4</sup>.

## After the Equivalence Point

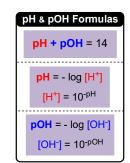
- In this part of the titration the moles of weak base is \_\_\_\_\_ the moles of strong acid.
  - □ There will be \_\_\_\_\_ strong acid remaining after it has neutralized the weak base.

**EXAMPLE:** Calculate the pH of the solution resulting from the titration between 125.0 mL of a 0.100 M HClO<sub>3</sub> and 50.0 mL of a 0.050 M NH<sub>3</sub>. ( $K_b$  of NH<sub>3</sub> is 1.75 x 10<sup>-5</sup>).

Use **STEPS 1 to 3** to setup the ICF Chart.



- STEP 4: Using the FINAL ROW, determine the concentration of the strong acid.
  - □ Divide its final \_\_\_\_\_ by the total volume used in the chemical reaction.
- **STEP 5:** Recall, the concentration of the strong acid will be equal to \_\_\_\_\_.



**PRACTICE:** A solution contains 100.0 mL of 0.550 M sodium nitrite, NaNO<sub>2</sub>. Find the pH after the addition of 180.0 mL of 0.400 M HClO<sub>4</sub>. The Ka of HNO<sub>2</sub> is 4.6 x 10<sup>-4</sup>.