



Titration 2: Strong Acid with Weak Base

A weak base, ammonia (NH_3 , 0.0500 M, 100.0 mL, $K_b = 1.8 \times 10^{-5}$), is titrated with a strong acid (HCl , 0.100 M). Find the pH after each addition of acid from the buret.

Vol HCl added (mL)	Total Volume (mL)	Calculation Workspace	pH
0.0	100.0	<p>Since NH_3 is a weak base, it will partially dissociate:</p> $\begin{array}{ccccccc} \text{NH}_3 & + & \text{H}_2\text{O} & \leftrightarrow & \text{NH}_4^+ & + & \text{OH}^- \\ 0.0500 & & - & & 0 & & 0 & \text{initial molarity} \\ -x & & - & & +x & & +x & \text{change in molarity} \\ \hline & & - & & & & & \text{final molarity} \end{array}$ <p>Use the K_b expression for NH_3 to solve for $[\text{OH}^-]$, then calculate pH.</p>	10.98
10.0	110.0	<p>The strong acid reacts completely with the weak base. Find the remaining moles after the reaction is complete:</p> $\begin{array}{ccccccc} \text{H}^+ & + & \text{NH}_3 & \rightarrow & \text{NH}_4^+ & & \\ 1.00 \times 10^{-3} & & 5.00 \times 10^{-3} & & 0 & & \text{initial moles} \\ -1.00 \times 10^{-3} & & -1.00 \times 10^{-3} & & +1.00 \times 10^{-3} & & \text{change in moles} \\ \hline & & & & & & \text{final moles} \end{array}$ <p>Convert the final moles to molarities using 110.0 mL as the volume:</p> <p>Do the weak base dissociation calculation using the new molarities:</p> $\begin{array}{ccccccc} \text{NH}_3 & + & \text{H}_2\text{O} & \leftrightarrow & \text{NH}_4^+ & + & \text{OH}^- \\ 3.64 \times 10^{-2} & & - & & 9.09 \times 10^{-3} & & 0 & \text{initial molarity} \\ -x & & - & & +x & & +x & \text{change in molarity} \\ \hline & & - & & & & & \text{final molarity} \end{array}$ <p>Use the K_b expression for NH_3 to solve for $[\text{OH}^-]$, then calculate pH.</p> <p>*Alternately, use the molarities calculated after the HCl addition in the Henderson-Hasselbalch equation to find the pOH, and then the pH.</p>	9.86
20.0		You try this one:	9.43

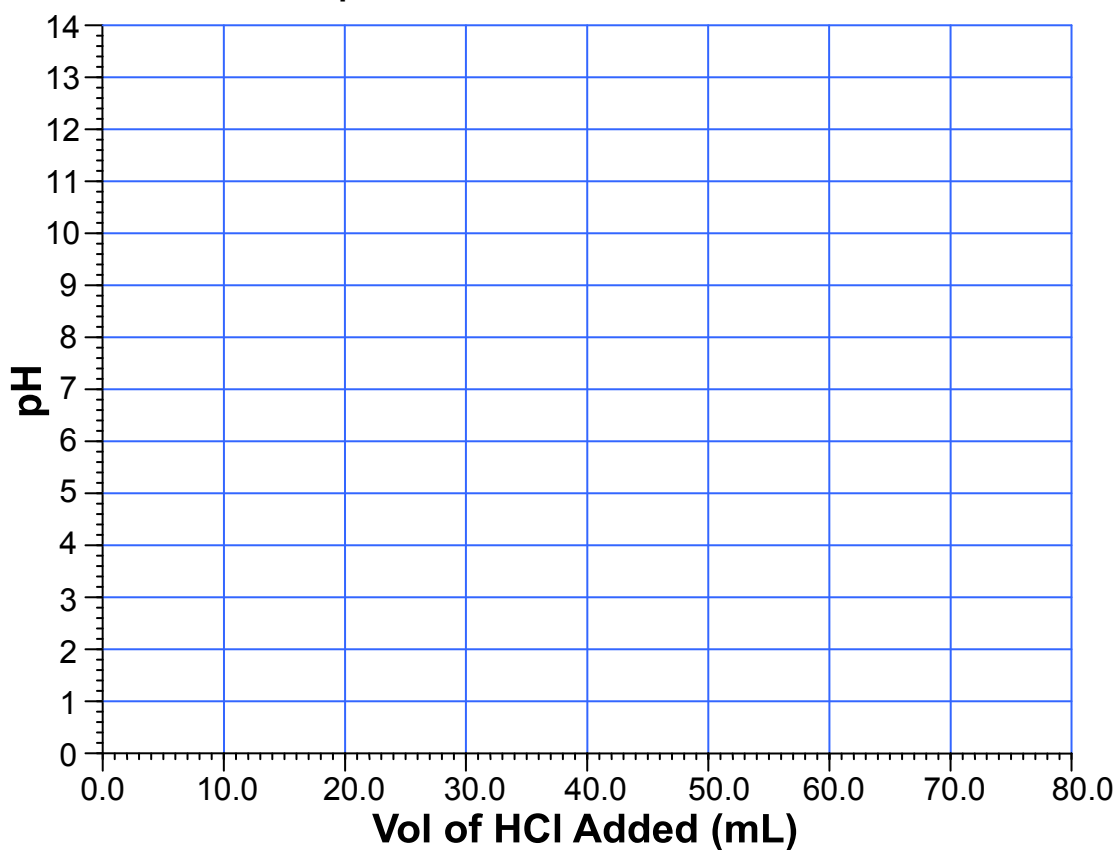
25.0	125.0	<p>Do the stoichiometric calculation:</p> $ \begin{array}{ccccccc} \text{H}^+ & + & \text{NH}_3 & \rightarrow & \text{NH}_4^+ & & \\ 2.50 \times 10^{-3} & & 5.00 \times 10^{-3} & & 0 & & \text{initial moles} \\ - 2.50 \times 10^{-3} & & - 2.50 \times 10^{-3} & & + 2.50 \times 10^{-3} & & \text{change in moles} \\ \hline & & & & & & \text{final moles} \end{array} $ <p>Convert the final moles to molarities using 125.0 mL as the volume:</p> <p>Do the weak base dissociation calculation for NH₃:</p> $ \begin{array}{ccccccc} \text{NH}_3 & + & \text{H}_2\text{O} & \leftrightarrow & \text{NH}_4^+ & + & \text{OH}^- \\ 2.00 \times 10^{-2} & & - & & 2.00 \times 10^{-2} & & 0 \text{ initial molarity} \\ - x & & - & & + x & & + x \text{ change in molarity} \\ \hline & & - & & & & \text{final molarity} \end{array} $ <p>Use the K_b expression for NH₃ to solve for [OH⁻], then calculate pH.</p>	9.26
40.0		You try this one:	8.65
50.0	150.0	<p>By now, 5.00 × 10⁻³ mol H⁺ reacts completely with the 5.00 × 10⁻³ mol NH₃ originally present. This is the equivalence point.</p> <p>Do the stoichiometric calculation:</p>	5.37

		$ \begin{array}{rcccc} \text{H}^+ & + & \text{NH}_3 & \rightarrow & \text{NH}_4^+ \\ 5.00 \times 10^{-3} & & 5.00 \times 10^{-3} & & 0 \\ - 5.00 \times 10^{-3} & & - 5.00 \times 10^{-3} & & + 5.00 \times 10^{-3} \end{array} $ <p style="text-align: right;">initial moles change in moles final moles</p> <p>Convert the final moles to molarities using 150.0 mL as the volume:</p> <p>The NH_4^+ produced in the neutralization is a weak acid, and will undergo an equilibrium reaction with water:</p> $ \begin{array}{rcccc} \text{NH}_4^+ & + & \text{H}_2\text{O} & \leftrightarrow & \text{NH}_3 & + & \text{H}_3\text{O}^+ \\ 3.33 \times 10^{-2} & & \text{—} & & 0 & & 0 \\ - x & & \text{—} & & + x & & + x \end{array} $ <p style="text-align: right;">initial molarity change in molarity final molarity</p> <p>Determine the K_a of NH_4^+ using $K_a = K_w \div K_b$. Use the K_b for NH_3.</p> <p>Use the K_a expression to solve for $[\text{H}_3\text{O}^+]$, then calculate pH. Or apply the Henderson-Hasselbalch equation to calculate pH.</p> <p>Note that for a weak base and a strong acid, the equivalence point occurs at $\text{pH} < 7$.</p>	
60.0	160.0	<p>Do the stoichiometric calculation using the neutralization reaction:</p> $ \begin{array}{rcccc} \text{H}^+ & + & \text{NH}_3 & \rightarrow & \text{NH}_4^+ \\ 6.00 \times 10^{-3} & & 5.00 \times 10^{-3} & & 0 \\ - 5.00 \times 10^{-3} & & - 5.00 \times 10^{-3} & & + 5.00 \times 10^{-3} \end{array} $ <p style="text-align: right;">initial moles change in moles excess moles</p> <p>NH_4^+ is a weak acid, but contributes little H^+ in comparison to the excess H^+ in solution. Therefore, the pH is mainly determined by the excess H^+ from the strong acid.</p> <p>Convert the final moles of H^+ to molarity using the total volume. $[\text{H}^+] =$</p> <p>Calculate pH.</p>	2.20
70.0		Try the last two.	1.93

80.0			1.78

After completing the table, a titration curve can be constructed. The pH is plotted against the volume of HCl added.

pH vs. Volume of HCl Added



When completed, the curve should be a decreasing curve, having a steeper section in the middle, with the equivalence point at 50.0 mL, corresponding to a pH of 5.37.