



Titration 3: Weak Acid with Strong Base

This worksheet simulates titrating a weak acid with a strong base.

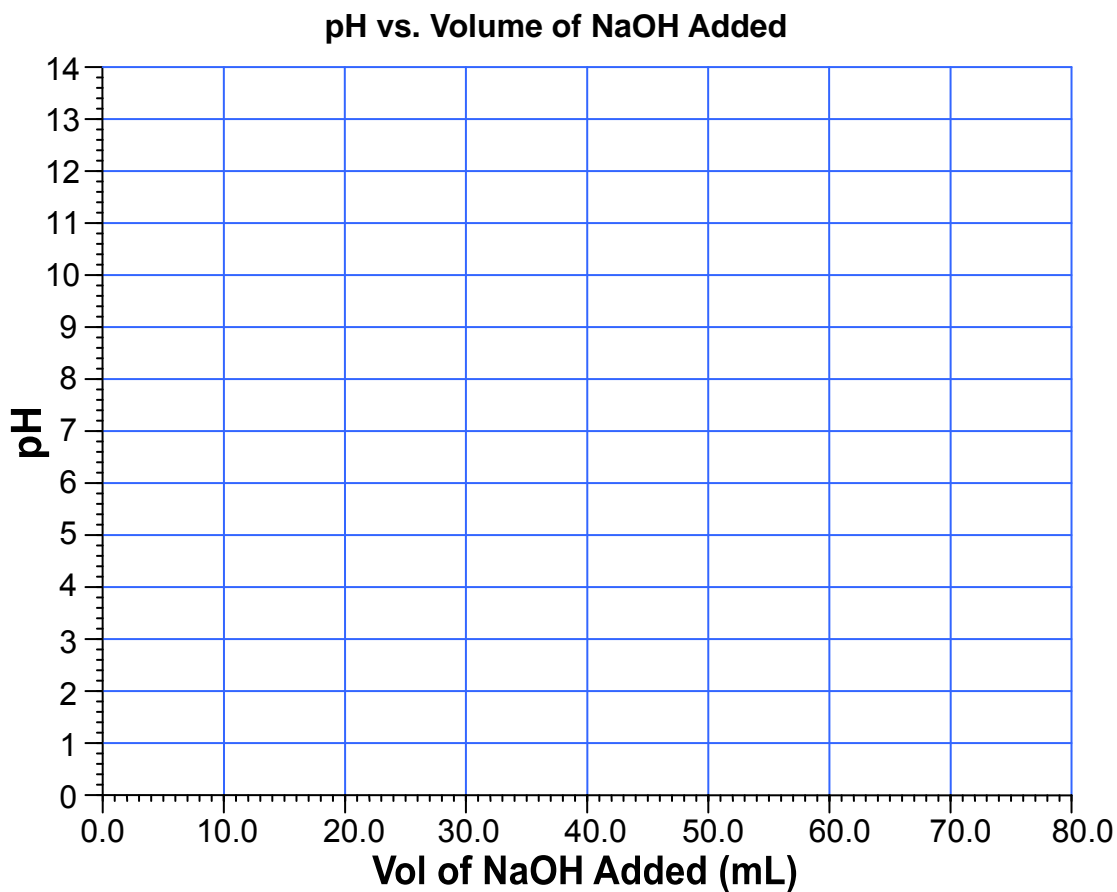
A weak acid, acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$, 0.100 M, 50.0 mL, $K_a = 1.8 \times 10^{-5}$), will be titrated with a strong base (NaOH, 0.100 M). The strong base is dispensed from a buret.

Vol NaOH added (mL)	Total Volume (mL)	Calculation Workspace	pH																								
0	50.0	<p>$\text{HC}_2\text{H}_3\text{O}_2$ is a weak acid and will dissociate:</p> $\text{HC}_2\text{H}_3\text{O}_2 \leftrightarrow \text{H}^+ + \text{C}_2\text{H}_3\text{O}_2^-$ <table style="width: 100%; border-collapse: collapse;"> <tr> <td style="text-align: center;">0.100</td> <td style="text-align: center;">0</td> <td style="text-align: center;">0</td> <td style="text-align: right;">initial molarity</td> </tr> <tr> <td style="text-align: center;">- x</td> <td style="text-align: center;">+ x</td> <td style="text-align: center;">+ x</td> <td style="text-align: right;">change in molarity</td> </tr> <tr> <td colspan="3"></td> <td style="text-align: right;">final molarity</td> </tr> </table> <p>Use the K_a expression for $\text{HC}_2\text{H}_3\text{O}_2$ to solve for $[\text{H}^+]$, then calculate pH.</p>	0.100	0	0	initial molarity	- x	+ x	+ x	change in molarity				final molarity	2.87												
0.100	0	0	initial molarity																								
- x	+ x	+ x	change in molarity																								
			final molarity																								
10.0	60.0	<p>When NaOH is added, the OH^- completely reacts with the acetic acid. Find the remaining moles after the reaction is complete:</p> $\text{OH}^- + \text{HC}_2\text{H}_3\text{O}_2 \rightarrow \text{C}_2\text{H}_3\text{O}_2^- + \text{H}_2\text{O}$ <table style="width: 100%; border-collapse: collapse;"> <tr> <td style="text-align: center;">1.00×10^{-3}</td> <td style="text-align: center;">5.00×10^{-3}</td> <td style="text-align: center;">0</td> <td style="text-align: right;">— initial moles</td> </tr> <tr> <td style="text-align: center;">- 1.00×10^{-3}</td> <td style="text-align: center;">- 1.00×10^{-3}</td> <td style="text-align: center;">+ 1.00×10^{-3}</td> <td style="text-align: right;">— change in moles</td> </tr> <tr> <td colspan="3"></td> <td style="text-align: right;">— final moles</td> </tr> </table> <p>Convert the final moles to molarities using 60.0 mL as the volume:</p> <p>Now you can calculate the pH two ways: Method 1: Do the weak acid dissociation calculation for acetic acid using the new molarities from above:</p> $\text{HC}_2\text{H}_3\text{O}_2 \leftrightarrow \text{H}^+ + \text{C}_2\text{H}_3\text{O}_2^-$ <table style="width: 100%; border-collapse: collapse;"> <tr> <td style="text-align: center;">6.67×10^{-2}</td> <td style="text-align: center;">0</td> <td style="text-align: center;">1.67×10^{-2}</td> <td style="text-align: right;">initial molarity</td> </tr> <tr> <td style="text-align: center;">- x</td> <td style="text-align: center;">+ x</td> <td style="text-align: center;">+ x</td> <td style="text-align: right;">change in molarity</td> </tr> <tr> <td colspan="3"></td> <td style="text-align: right;">final molarity</td> </tr> </table> <p>Use the K_a expression to solve for $[\text{H}^+]$, then calculate pH.</p> <p>Method 2: Using the molarities calculated after the addition of NaOH, plug into the Henderson-Hasselbalch equation to find the pH (since you have in essence created a buffer solution).</p>	1.00×10^{-3}	5.00×10^{-3}	0	— initial moles	- 1.00×10^{-3}	- 1.00×10^{-3}	+ 1.00×10^{-3}	— change in moles				— final moles	6.67×10^{-2}	0	1.67×10^{-2}	initial molarity	- x	+ x	+ x	change in molarity				final molarity	4.14
1.00×10^{-3}	5.00×10^{-3}	0	— initial moles																								
- 1.00×10^{-3}	- 1.00×10^{-3}	+ 1.00×10^{-3}	— change in moles																								
			— final moles																								
6.67×10^{-2}	0	1.67×10^{-2}	initial molarity																								
- x	+ x	+ x	change in molarity																								
			final molarity																								
20.0		You try this one:	4.56																								

25.0			4.74
50.0		<p>By now, 5.00×10^{-3} moles OH^- reacts completely with the 5.00×10^{-3} mol $\text{HC}_2\text{H}_3\text{O}_2$ originally present. This is the equivalence point. But because $\text{C}_2\text{H}_3\text{O}_2^-$ is produced in the neutralization, and it is a weak base, there will be a small equilibrium reaction after.</p> $ \begin{array}{ccccccc} \text{OH}^- & + & \text{HC}_2\text{H}_3\text{O}_2 & \rightarrow & \text{C}_2\text{H}_3\text{O}_2^- & + & \text{H}_2\text{O} \\ 5.00 \times 10^{-3} & & 5.00 \times 10^{-3} & & 0 & & \text{---} & \text{initial moles} \\ -5.00 \times 10^{-3} & & -5.00 \times 10^{-3} & & +5.00 \times 10^{-3} & & \text{---} & \text{change in moles} \\ & & & & & & \text{---} & \text{final moles} \end{array} $ <p>Convert the final moles of acetate to molarity using the total volume.</p> <p>$\text{C}_2\text{H}_3\text{O}_2^-$ is a weak base and will react with water:</p> $ \begin{array}{ccccccc} \text{C}_2\text{H}_3\text{O}_2^- & + & \text{H}_2\text{O} & \leftrightarrow & \text{HC}_2\text{H}_3\text{O}_2 & + & \text{OH}^- \\ 5.00 \times 10^{-2} & & \text{---} & & 0 & & 0 & \text{initial molarity} \\ -x & & \text{---} & & +x & & +x & \text{change in molarity} \\ & & \text{---} & & & & & \text{final molarity} \end{array} $ <p>Use K_a and K_w to find K_b and solve for the concentration of OH^- produced in equilibrium.</p> <p>Then find the H^+ concentration and the pH.</p> <p>Note that for a weak acid and strong base, the equivalence point occurs at $\text{pH} > 7$.</p>	8.72

60.0	110.0	<p>Do the stoichiometric calculation:</p> $ \begin{array}{ccccccc} \text{OH}^- & + & \text{HC}_2\text{H}_3\text{O}_2 & \rightarrow & \text{C}_2\text{H}_3\text{O}_2^- & + & \text{H}_2\text{O} \\ 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} & & \\ \hline & & & & & & \\ & & & & & & \text{initial moles} \\ & & & & & & \text{change in moles} \\ & & & & & & \text{final moles} \end{array} $ <p>$\text{C}_2\text{H}_3\text{O}_2^-$ is a weak base but contributes so little OH^- in comparison to the excess OH^- in solution. Therefore, the pH is mainly determined by the excess OH^-.</p> <p>Convert the final moles of OH^- to molarity using 110.0 mL as the volume:</p> <p>$[\text{OH}^-] =$</p> <p>From here, you can calculate pOH and then subtract from 14.00 to find pH. Or you can use $[\text{H}^+] = K_w \div [\text{OH}^-]$ to calculate hydrogen concentration and then find pH.</p>	11.96
70.0			12.22
80.0			12.36

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



When completed, the curve should look like a horizontally-stretched S-shaped graph, with the equivalence point at 50.0 mL, corresponding to a pH of 8.72.