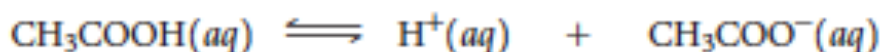


Weak Acid

Calculate the pH of a 0.30 M solution of acetic acid ($K_a = 1.8 \times 10^{-5}$)



Initial	0.30 M	0	0
Change	-x M	+x M	+x M
Equilibrium	(0.30 - x) M	x M	x M

$$K_a = \frac{[\text{H}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} = \frac{(x)(x)}{0.30 - x} = 1.8 \times 10^{-5}$$

$$K_a = \frac{x^2}{0.30} = 1.8 \times 10^{-5}$$

$$x^2 = (0.30)(1.8 \times 10^{-5}) = 5.4 \times 10^{-6}$$

$$x = \sqrt{5.4 \times 10^{-6}} = 2.3 \times 10^{-3}$$

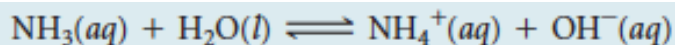
$$[\text{H}^+] = x = 2.3 \times 10^{-3} \text{ M}$$

$$\text{pH} = -\log(2.3 \times 10^{-3}) = 2.64$$

$$\text{Percent ionization of CH}_3\text{COOH} = \frac{0.0023 \text{ M}}{0.30 \text{ M}} \times 100\% = 0.77\%$$

Weak Base

Calculate the pH of a 0.15 M solution of NH_3 ($K_b = 1.8 \times 10^{-5}$).



$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = 1.8 \times 10^{-5}$$

	$\text{NH}_3(aq)$	+	$\text{H}_2\text{O}(l)$	\rightleftharpoons	$\text{NH}_4^+(aq)$	+	OH^-
Initial	0.15 M		—		0		0
Change	-x M		—		+x M		+x M
Equilibrium	(0.15 - x) M		—		x M		x M

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = \frac{(x)(x)}{0.15 - x} = 1.8 \times 10^{-5}$$

$$\frac{x^2}{0.15} = 1.8 \times 10^{-5}$$

$$x^2 = (0.15)(1.8 \times 10^{-5}) = 2.7 \times 10^{-6}$$

$$x = [\text{NH}_4^+] = [\text{OH}^-] = \sqrt{2.7 \times 10^{-6}} = 1.6 \times 10^{-3} \text{ M}$$

$[\text{OH}^-] = 1.6 \times 10^{-3} \text{ M}$. Thus, $\text{pOH} = -\log(1.6 \times 10^{-3}) = 2.80$, and $\text{pH} = 14.00 - 2.80 = 11.20$. The pH of the solution is above 7 because we are dealing with a solution of a base.

Buffer

What is the pH of a solution made by adding 0.30 mol of acetic acid and 0.30 mol of sodium acetate to enough water to make 1.0 L of solution? ($K_a = 1.8 \times 10^{-5}$)

$$\text{pH} = \text{p}K_a + \log \frac{[\text{base}]}{[\text{acid}]}$$

	$\text{CH}_3\text{COOH}(aq)$	\rightleftharpoons	$\text{H}^+(aq)$	+	$\text{CH}_3\text{COO}^-(aq)$
Initial	0.30 M		0		0.30 M
Change	$-x\text{ M}$		$+x\text{ M}$		$+x\text{ M}$
Equilibrium	$(0.30 - x)\text{ M}$		$x\text{ M}$		$(0.30 + x)\text{ M}$

$$\text{pH} = \text{p}K_a + \log (0.3/0.3) = \text{p}K_a = -\log K_a = -\log 1.8 \times 10^{-5}$$

$$\text{pH} = 4.74$$

or

$$K_a = 1.8 \times 10^{-5} = \frac{[\text{H}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]}$$

$$K_a = 1.8 \times 10^{-5} = \frac{x(0.30 + x)}{0.30 - x}$$

$$K_a = 1.8 \times 10^{-5} = \frac{x(0.30)}{0.30}$$

$$x = 1.8 \times 10^{-5}\text{ M} = [\text{H}^+]$$

$$\text{pH} = -\log(1.8 \times 10^{-5}) = 4.74$$

Buffer

What is the pH of a buffer that is 0.12 M in lactic acid [CH₃CH(OH)COOH, or HC₃H₅O₃] and 0.10 M in sodium lactate [CH₃CH(OH)COONa, or NaC₃H₅O₃]? For lactic acid, $K_a = 1.4 \times 10^{-4}$

	HC ₃ H ₅ O ₃ (aq)	⇌	H ⁺ (aq)	+	C ₃ H ₅ O ₃ ⁻ (aq)
Initial	0.12 M		0		0.10 M
Change	-x M		+x M		+x M
Equilibrium	(0.12 - x) M		x M		(0.10 + x) M

$$\begin{aligned} \text{pH} &= \text{p}K_a + \log \frac{[\text{base}]}{[\text{acid}]} = 3.85 + \log \left(\frac{0.10}{0.12} \right) \\ &= 3.85 + (-0.08) = 3.77 \end{aligned}$$

or

$$K_a = 1.4 \times 10^{-4} = \frac{[\text{H}^+][\text{C}_3\text{H}_5\text{O}_3^-]}{[\text{HC}_3\text{H}_5\text{O}_3]} = \frac{x(0.10 + x)}{(0.12 - x)}$$

$$K_a = 1.4 \times 10^{-4} = \frac{x(0.10)}{0.12}$$

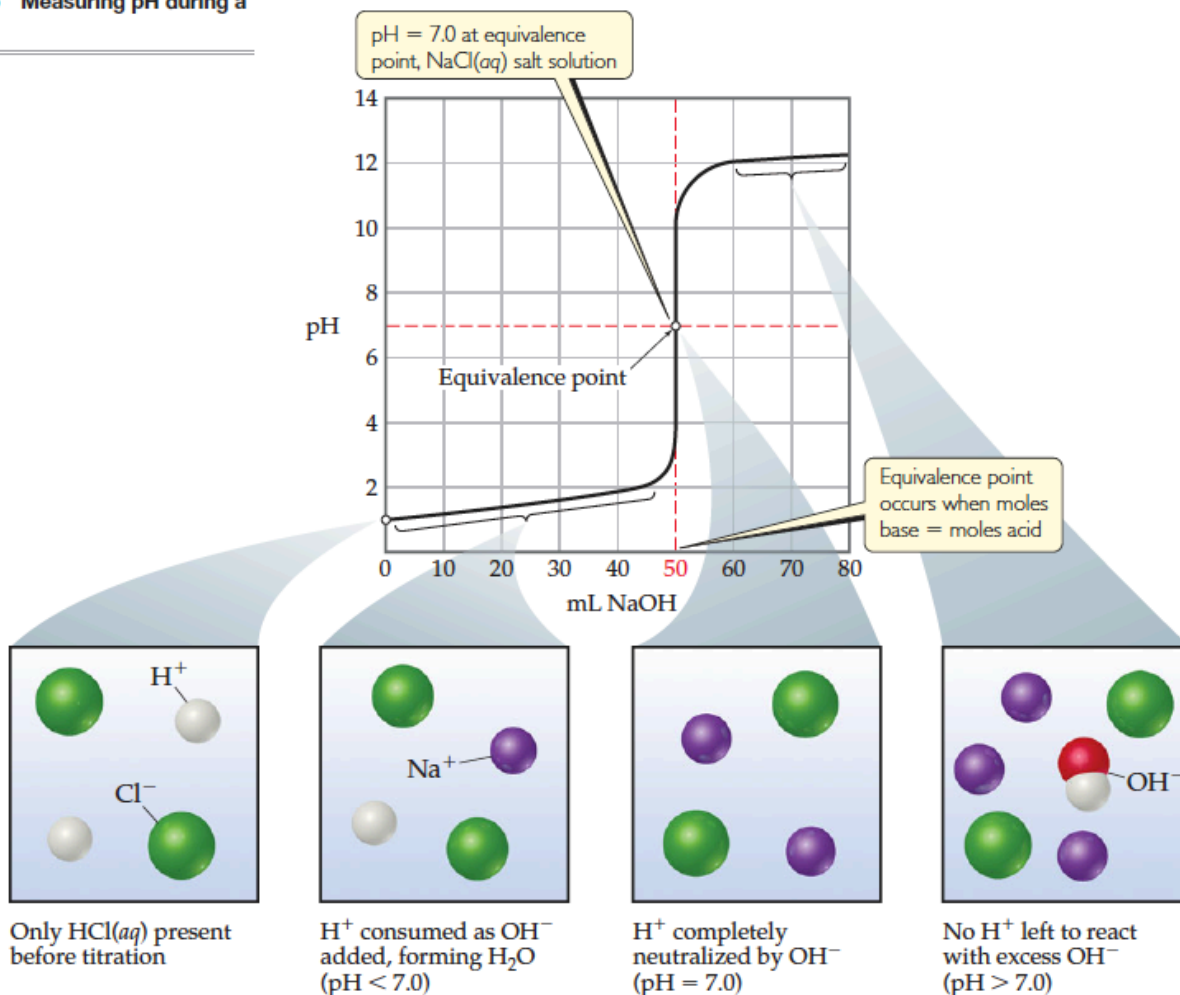
$$[\text{H}^+] = x = \left(\frac{0.12}{0.10} \right) (1.4 \times 10^{-4}) = 1.7 \times 10^{-4} \text{ M}$$

$$\text{pH} = -\log(1.7 \times 10^{-4}) = 3.77$$

Strong acid- Strong base titration

Equivalence point: $\text{pH} = 7$

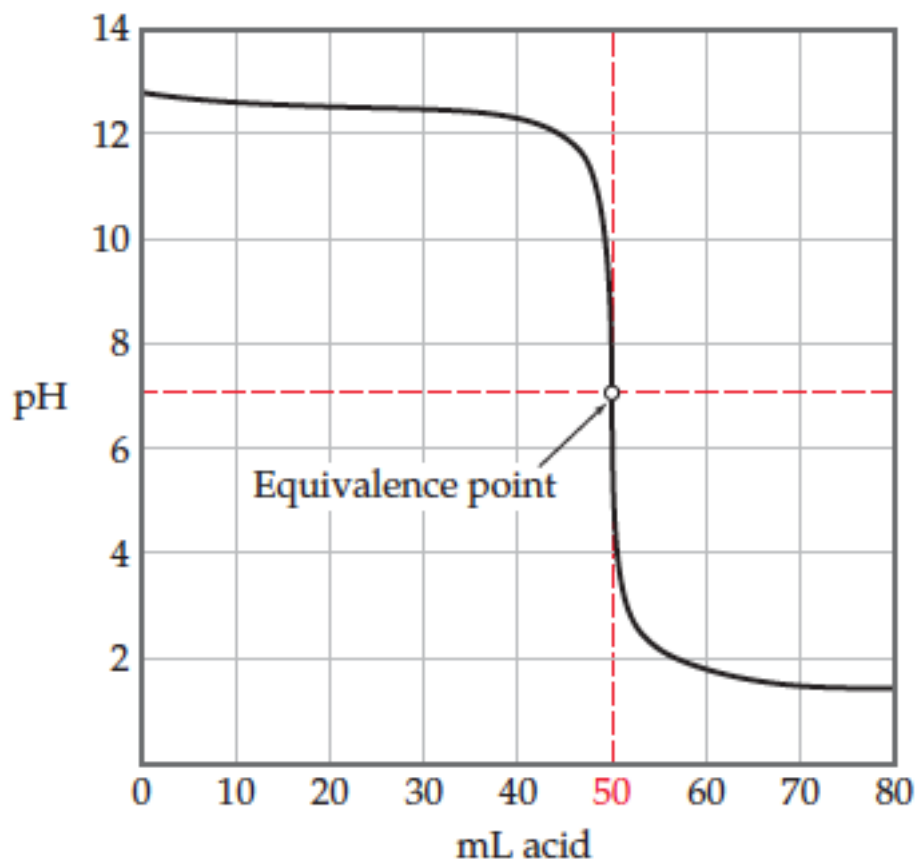
17.6 Measuring pH during a



▲ **FIGURE 17.7** Titration of a strong acid with a strong base. The pH curve for titration of 50.0 mL of a 0.100 M solution of hydrochloric acid with a 0.100 M solution of $\text{NaOH}(aq)$. For clarity, water molecules have been omitted from the molecular art.

Strong base- Strong acid titration

Equivalence point: $\text{pH} = 7$

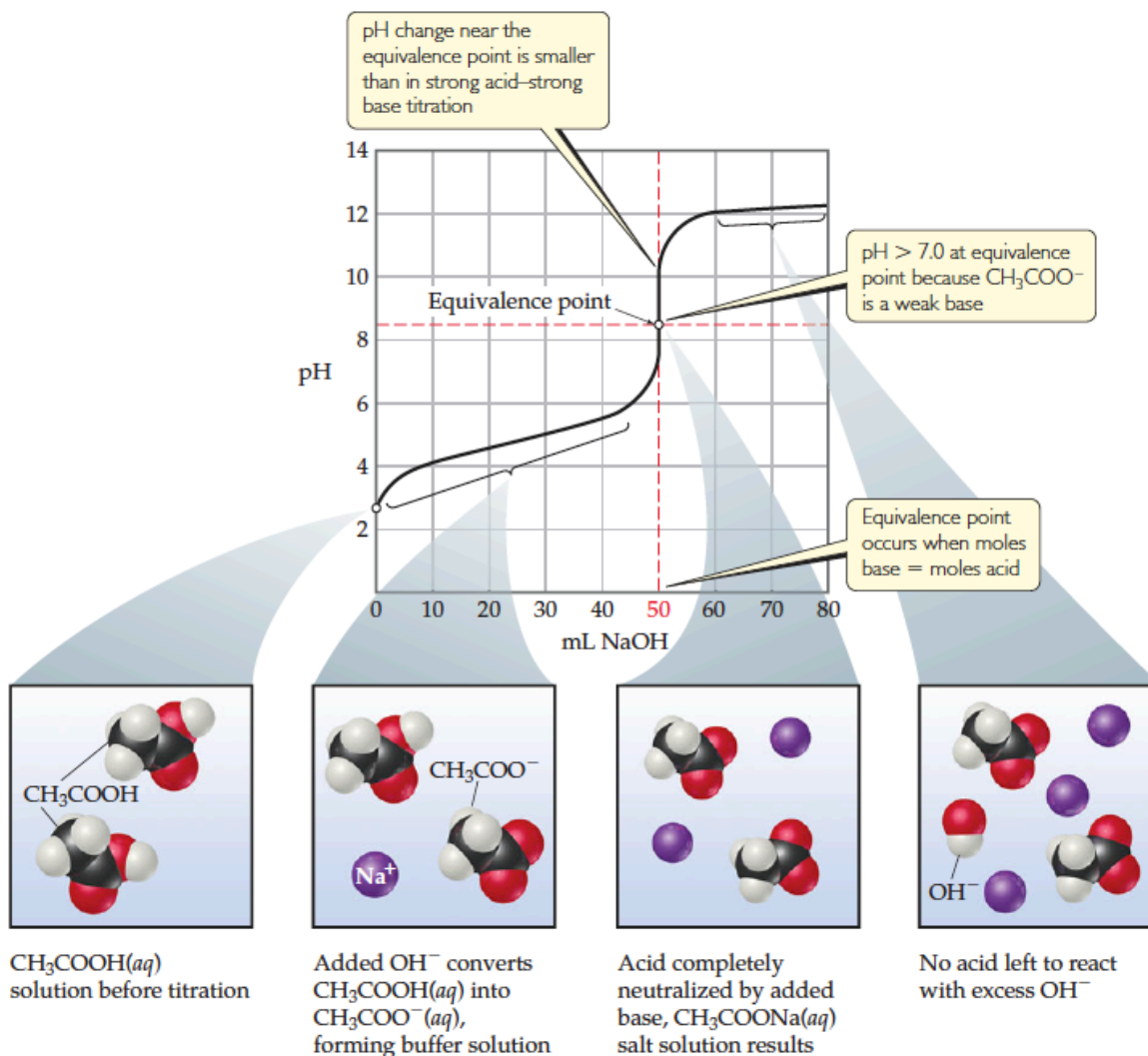


▲ **FIGURE 17.8** Titration of a strong base with a strong acid. The pH curve for titration of 50.0 mL of a 0.100 M solution of a strong base with a 0.100 M solution of a strong acid.

Titration of weak acid with a strong base

Equivalence point: $\text{pH} > 7$

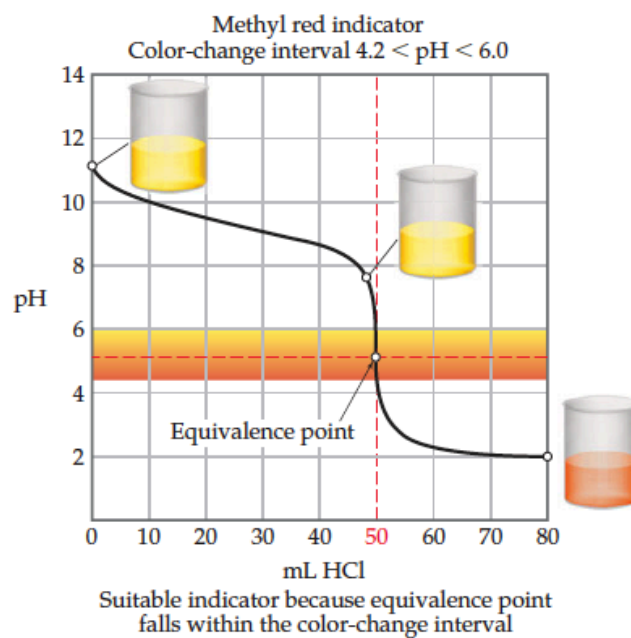
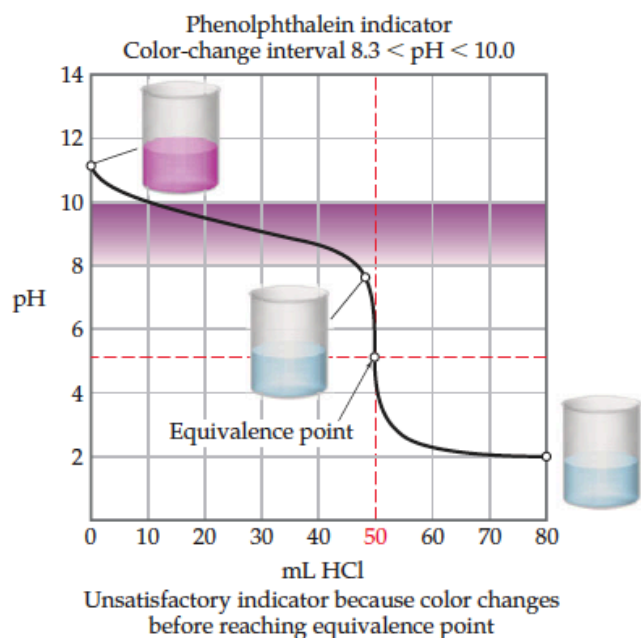
If the acetic acid being titrated here were replaced by hydrochloric acid, would the amount of base needed to reach the equivalence point change? Would the pH at the equivalence point change?



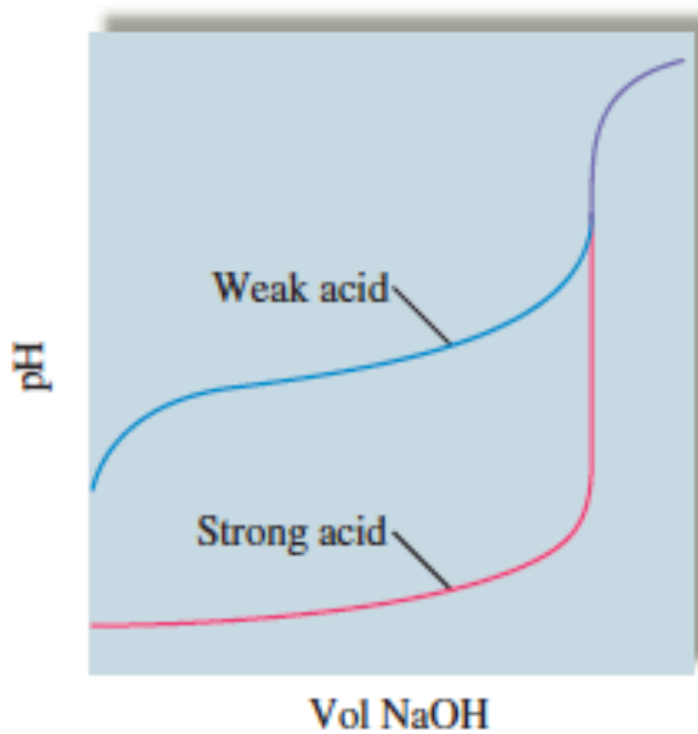
▲ **FIGURE 17.9** Titration of a weak acid with a strong base. The pH curve for titration of 50.0 mL of a 0.100 M solution of acetic acid with a 0.100 M solution of NaOH(aq) . For clarity, water molecules have been omitted from the molecular art.

Titration of weak base with a strong acid

Equivalence point: $\text{pH} < 7$

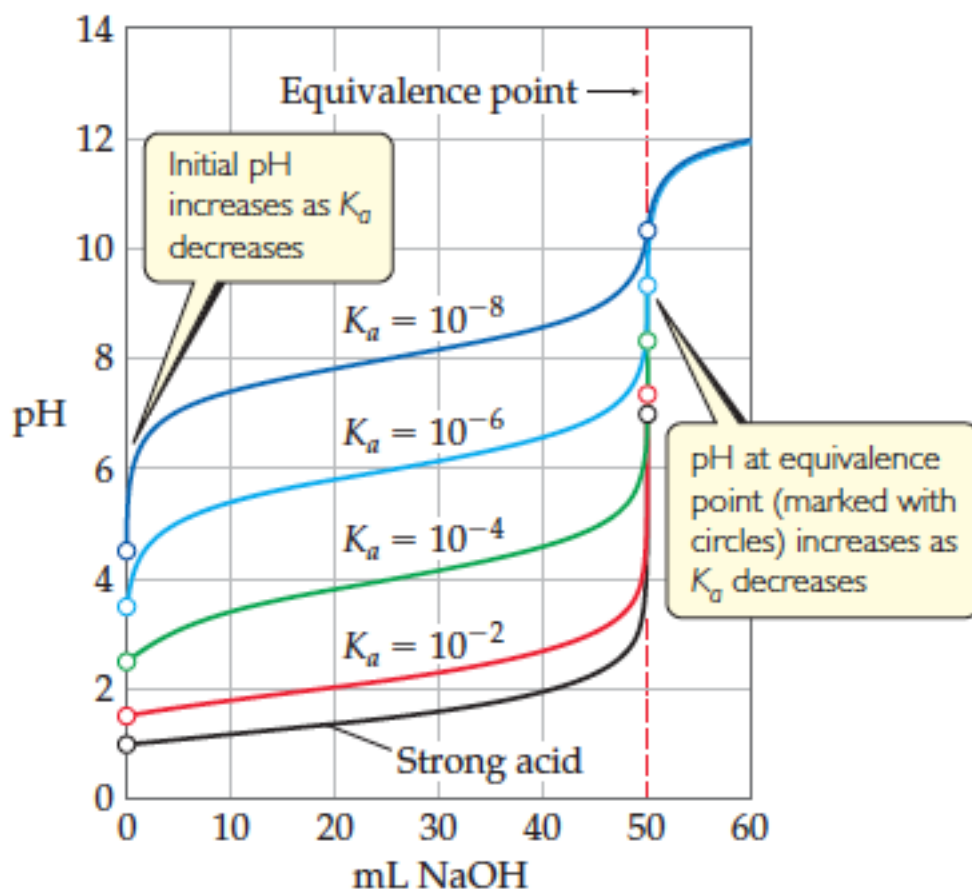


▲ FIGURE 17.15 Good and poor indicators for titration of a weak base with a strong acid.



The equivalence point is defined by the stoichiometry, not by the pH.

pH at the equivalence point increases as the K_a decreases.



▲ **FIGURE 17.11** A set of curves showing the effect of acid strength on the characteristics of the titration curve when a weak acid is titrated by a strong base. Each curve represents titration of 50.0 mL of 0.10 M acid with 0.10 M NaOH.