

Titration Calculations

Strong Acid/Strong Base Calculations

- (1) Use balanced equation to do stoichiometric calculation.
- (2) Determine pH from amount of strong acid/base that is in excess.

Note: At stoichiometry point of equal acid and base, pH = 7.

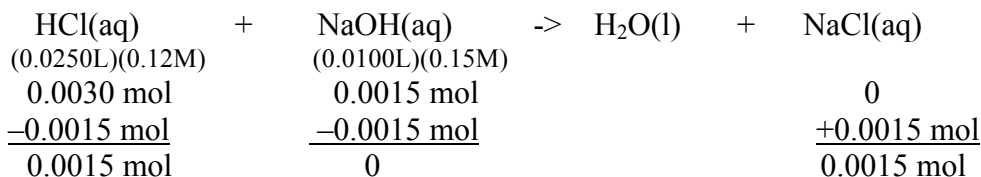
Example:

What is pH after 0.0 mL, 10.0 mL, at equivalence point, and 50.0 mL of base has been added during a titration to 25.0 mL of a 0.12M HCl solution with 0.15M NaOH solution?

For strong acid/base titration, perform stoichiometry calculation first; then calculation resulting concentration with total volume; finally, calculate pH directly.

(A) 0.0 mL base: Solution is 0.12M HCl $\text{pH} = -\log[\text{H}^+] = -\log(0.12) = 0.92$

(B) 10.0 mL added base:



$$[\text{HCl}] = 0.0015 \text{ mol} / 0.0350 \text{ L} = 0.043 \text{ M}$$

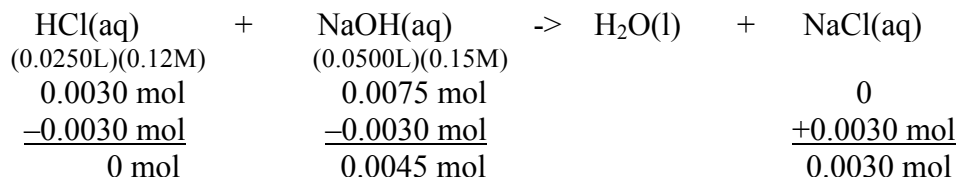
Therefore **since strong acid**: $[\text{H}^+] = 0.043 \text{ M}$ so $\text{pH} = -\log(0.043) = 1.37$

(C) At Equivalence Point:

$$\begin{aligned} \text{Volume of base added} &= (0.0030 \text{ mol HCl})(1 \text{ mol NaOH} / 1 \text{ mol HCl})(1 \text{ L} / 0.15 \text{ mol NaOH}) \\ &= 0.020 \text{ L} = 20. \text{ mL added base} \end{aligned}$$

Since NaCl does not hydrolyze water, pH is neutral 7.00.

(D) 50.0 mL added base:



$$[\text{NaOH}] = 0.0045 \text{ mol} / 0.0750 \text{ L} = 0.060 \text{ M}$$

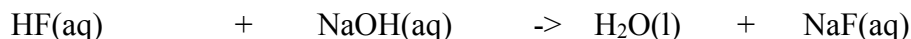
Therefore **since strong base left**: $[\text{OH}^-] = 0.060 \text{ M}$ so $\text{pOH} = -\log(0.060) = 1.22$
 $\text{pH} = 12.78$

Weak Acid/Strong Base Calculations

What is pH after 0.0 mL, 10.0mL, at equivalence point, and 50.0 mL of base has been added during a titration to 25.0 mL of a 0.12M HF solution with 0.15M NaOH solution?
 $K_a = 6.8 \times 10^{-4}$

- (1) Use balanced equation to do **stoichiometric** calculation.
- (2) Determine new concentrations by dividing by total volume.
- (3) Use appropriate **equilibrium** reaction and ICE chart to determine pH.

Stoichiometric Reaction:



Equilibrium Reaction:



(A) Addition of 0.0 mL of base:
Only weak acid present.

	HF (aq)	+	H ₂ O	⇌	H ₃ O ⁺ (aq)	+	F ⁻ (aq)
I	0.12 M				0		0
C	- x				+ x		+ x
E	0.12 - x				x		x

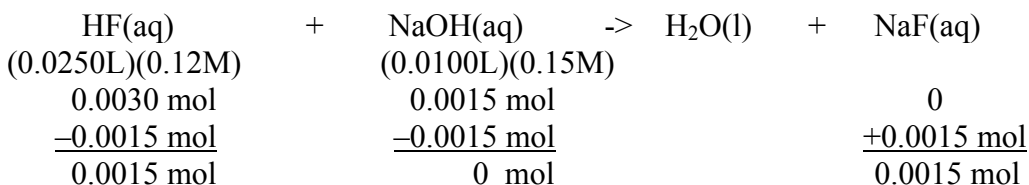
$$K_a = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]} \quad 6.8 \times 10^{-4} = \frac{x^2}{(0.12 - x)}$$

$$x = 8.7 \times 10^{-3} \text{ M} \quad \text{pH} = -\log(8.7 \times 10^{-3}) = 2.06$$

(B) What is pH after 10.0mL of 0.15M NaOH solution has been added to 25.0 mL of 0.12M HF solution? $K_a = 6.8 \times 10^{-4}$

- (1) Use balanced equation to do stoichiometric calculation.
- (2) Determine new concentrations by dividing by total volume.
- (3) Use appropriate equilibrium reaction and ICE chart to determine pH.

(1) **Stoichiometric** Reaction:



(2) New concentrations:

$$[\text{HF}] = 0.0015 \text{ mol} / 0.0350\text{L} = 0.043\text{M}$$

$$[\text{F}^-] = 0.0015 \text{ mol} / 0.0350\text{L} = 0.043\text{M}$$

(3) **Equilibrium** Reaction:

	HF (aq)	+	H ₂ O	⇌	H ₃ O ⁺ (aq)	+	F ⁻ (aq)
I	0.043M				0		0.043M
C	- x				+ x		+ x
E	0.043 - x				x		0.043 + x

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]} \quad 6.8 \times 10^{-4} = \frac{(x)(0.43 + x)}{(0.43 - x)}$$

$$x = 6.8 \times 10^{-4} \text{ M} \quad \text{pH} = -\log(6.8 \times 10^{-4}) = 3.17$$

Note: Could also use Henderson-Hasselbalch equation since this is buffer region of titration curve.

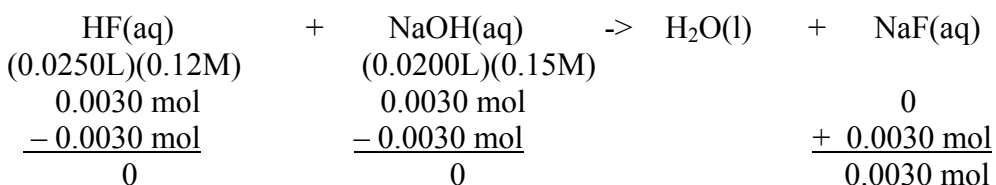
(C) What is pH at equivalence point?

First need to determine volume at equivalence point.

$$(0.0250 \text{ L}) \left(\frac{0.12 \text{ mol HF}}{1 \text{ L}} \right) \left(\frac{1 \text{ mol NaOH}}{1 \text{ mol HF}} \right) \left(\frac{1 \text{ L}}{0.15 \text{ mol NaOH}} \right) = 0.0200 \text{ L or } 20.0 \text{ mL}$$

- (1) Use balanced equation to do stoichiometric calculation.
- (2) Determine new concentrations by dividing by total volume.
- (3) Use appropriate equilibrium reaction and ICE chart to determine pH.

(1) **Stoichiometric** Reaction:



(2) New concentrations:

$$[\text{HF}] = 0 \text{ mol} / 0.0450 \text{ L} = 0 \text{ M}$$

$$[\text{F}^-] = 0.0030 \text{ mol} / 0.0450 \text{ L} = 0.067 \text{ M} \qquad K_b = \frac{1 \times 10^{-14}}{6.8 \times 10^{-4}} = 1.5 \times 10^{-11}$$

(3) **Equilibrium** Reaction:

Only conjugate base now left. So must use equilibrium reaction for conjugate base and calculate K_b .

	F ⁻ (aq)	+	H ₂ O	⇌	OH ⁻ (aq)	+	HF (aq)
I	0.067 M				0		0
C	- x				+ x		+ x
E	0.067 - x				x		x

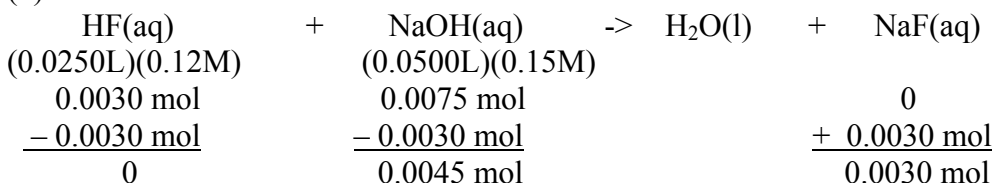
$$K_b = \frac{[\text{OH}^-][\text{HF}]}{[\text{F}^-]} \qquad 1.5 \times 10^{-11} = \frac{x^2}{(0.067 - x)}$$

$$x = 1.0 \times 10^{-6} \text{ M} \quad \text{pOH} = -\log (1.0 \times 10^{-6}) = 6.00 \quad \text{so} \quad \text{pH} = 8.00$$

(D) What is pH after 50.0mL of 0.15M NaOH solution has been added to 25.0 mL of 0.12M HF solution? $K_a = 6.8 \times 10^{-4}$

- (1) Use balanced equation to do stoichiometric calculation.
- (2) Determine new concentrations by dividing by total volume.
- (3) Use appropriate equilibrium reaction and ICE chart to determine pH.

(1) **Stoichiometric** Reaction:



(2) New concentrations:

$$[\text{OH}^-] = 0.0045 \text{ mol} / 0.0750 \text{ L} = 0.060 \text{ M}$$

$$[\text{F}^-] = 0.0030 \text{ mol} / 0.0750 \text{ L} = 0.040 \text{ M} \quad K_b = \frac{1 \times 10^{-14}}{6.8 \times 10^{-4}} = 1.5 \times 10^{-11}$$

(3) **Equilibrium** Reaction:

	F ⁻ (aq)	+	H ₂ O	⇌	OH ⁻ (aq)	+	HF (aq)
I	0.040 M				0.060 M		0
C	- x				+ x		+ x
E	0.040 - x				0.060 + x		x

$$K_b = \frac{[\text{OH}^-][\text{HF}]}{[\text{F}^-]} \quad 1.5 \times 10^{-11} = \frac{(x)(0.060 + x)}{(0.040 - x)}$$

$$x = 1.0 \times 10^{-11} \text{ M} \quad \text{pOH} = -\log(1.0 \times 10^{-11}) = 11.00 \quad \text{so} \quad \text{pH} = 3.00 ???$$

NOTE: "x" is NOT the OH⁻ concentration. The OH⁻ concentration is 0.060M + x. Since there is excess strong base in this last addition, the pH is determined by the strong base concentration. The weak conjugate base F⁻ adds an insignificant amount.

$$[\text{OH}^-] = 0.060 + x = 0.060 \text{ M} + 1.0 \times 10^{-11} \text{ M} = 0.060 \text{ M}$$

$$\text{pOH} = -\log(0.060) = 1.22 \quad \text{so} \quad \text{pH} = 12.78$$

