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, 10.0mL, 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 pH = -log[H+] = -log(0.12) = 0.92

(B) 10.0mL added base:

$HCl(aq)$	+	$NaOH(aq)$	->	$H_2O(l)$	+	$NaCl(aq)$
(0.0250L)(0.12M)		(0.0100L)(0.15M)				
0.0030 mol		0.0015 mol				0
–0.0015 mol		–0.0015 mol				+0.0015 mol
0.0015 mol		0				0.0015 mol

[HCl] = 0.0015 mol/0.0350L = 0.043 M
Therefore since strong acid: $[H^+]$ = 0.043 M so pH = –log (0.043) = 1.37

(C) At Equivalence Point:
Volume of base added = (0.0030mol HCl)(1mol NaOH/1mol HCl)(1L/0.15mol NaOH)
= 0.020 L = 20. mL added base

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

(D) 50.0mL added base:

$HCl(aq)$	+	$NaOH(aq)$	->	$H_2O(l)$	+	$NaCl(aq)$
(0.0250L)(0.12M)		(0.0500L)(0.15M)				
0.0030 mol		0.0075 mol				0
–0.0030 mol		–0.0030 mol				+0.0030 mol
0 mol		0.0045 mol				0.0030 mol

[NaOH] = 0.0045 mol/0.0750L = 0.060 M
Therefore since strong base left: $[OH^-]$ = 0.060 M so pOH = –log (0.060) = 1.22
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:

$$HF(aq) + NaOH(aq) \rightarrow H_2O(l) + NaF(aq)$$

Equilibrium Reaction:

$$HF(aq) + H_2O(l) \rightarrow H_3O^+(aq) + F^-(aq)$$

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

	HF (aq)	+	H_2O	⇔	H_3O^+(aq)	+	F^- (aq)
I	0.12 M				0		0
C	- x				+ x		+ x
E	0.12 - x				x		x

$$K_a = \frac{[H_3O^+][F^-]}{[HF]} \qquad 6.8\text{x}10^{-4} = \frac{x^2}{(0.12 - x)}$$

$x = 8.7 \times 10^{-3}$ M pH = -log (8.7×10^{-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:

HF(aq)	+	NaOH(aq)	->	H_2O(l)	+	NaF(aq)
(0.0250L)(0.12M)		(0.0100L)(0.15M)				
0.0030 mol		0.0015 mol				0
–0.0015 mol		–0.0015 mol				+0.0015 mol
0.0015 mol		0 mol				0.0015 mol

(2) New concentrations:

[HF] = 0.0015 mol/ 0.0350L = 0.043M

$[F^-]$ = 0.0015 mol/ 0.0350L = 0.043M

(3) Equilibrium Reaction:

	HF (aq)	+	H_2O	⇔	H_3O^+(aq)	+	F^- (aq)
I	0.043M				0		0.043M
C	- x				+ x		+ x
E	0.043 - x				x		0.043 + x

$$K_a = \frac{[H_3O^+][F^-]}{[HF]} \qquad 6.8\text{x}10^{-4} = \frac{(x\ (0.43 + x))}{(0.43 - x)}$$

x = 6.8 x 10^{-4} M   pH = -log (6.8 x 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:

HF(aq)	+	NaOH(aq)	->	$H_2O(l)$	+	NaF(aq)
(0.0250L)(0.12M)		(0.0200L)(0.15M)				
0.0030 mol		0.0030 mol				0
– 0.0030 mol		– 0.0030 mol				+ 0.0030 mol
0		0				0.0030 mol

(2) New concentrations:

[HF] = 0 mol/0.0450L = 0 M

$[F^-]$ = 0.0030 mol/0.0450L = 0.067 M

$$K_b = \frac{1\text{x}10^{-14}}{6.8\text{x}10^{-4}} = 1.5\text{x}10^{-11}$$

(3) Equilibrium Reaction:

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

	F^- (aq)	+	H_2O	⇔	OH^-(aq)	+	HF (aq)
I	0.067 M				0		0
C	- x				+ x		+ x
E	0.067 - x				x		x

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

x = 1.0 x 10^{-6} M   pOH = -log (1.0 x 10^{-6}) = 6.00   so 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:

HF(aq)	+	NaOH(aq)	->	$H_2O(l)$	+	NaF(aq)
(0.0250L)(0.12M)		(0.0500L)(0.15M)				
0.0030 mol		0.0075 mol				0
– 0.0030 mol		– 0.0030 mol				+ 0.0030 mol
0		0.0045 mol				0.0030 mol

(2) New concentrations:

$[OH^-]$ = 0.0045 mol/0.0750L = 0.060 M

$[F^-]$ = 0.0030 mol/0.0750L = 0.040 M

$$K_b = \frac{1x10^{-14}}{6.8x10^{-4}} = 1.5x10^{-11}$$

(3) Equilibrium Reaction:

	F^- (aq) + H_2O ⇔	OH^-(aq) +	HF (aq)
I	0.040 M	0.060 M	0
C	- x	+ x	+ x
E	0.040 - x	0.060 + x	x

$$K_b = \frac{[OH^-][HF]}{[F^-]} \qquad 1.5x10^{-11} = \frac{(x\,(0.060 + x))}{(0.040 - x)}$$

x = 1.0 x 10^{-11} M ~~pOH = -log (1.0 x 10-11) = 11.00 so 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.

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

pOH = –log (0.060) = 1.22 so pH = 12.78