

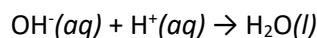
Acid-Base Equilibria

Section 98 – Acid-Base Titrations

- 98-1 A 245.0 mL sample of 0.15 M Ba(OH)₂ is added to 438.0 mL of 0.200 M HNO₃. What is the pH of the resulting solution?

Solution:

Approach: note that Ba(OH)₂ is a strong base, which dissociates in solution to form two moles of OH⁻ for every mole of Ba(OH)₂ that was dissolved in the water. The HNO₃ is a strong acid, which also fully dissociates to H⁺ and NO₃⁻. The Ba²⁺ and NO₃⁻ are spectator ions; however the OH⁻ and H⁺ will react in a neutralization reaction as follows:



Thus the key to solving this problem is to determine the number of moles of OH⁻ and H⁺ that were originally added, in order to determine what the concentrations are after the neutralization reaction. An ICE table is useful for this purpose.

Thus:

$$\text{mol OH}^- = 245.0 \text{ mL Ba(OH)}_2 \times \frac{0.15 \text{ mol Ba(OH)}_2}{1000 \text{ mL}} \times \frac{2 \text{ mol OH}^-}{1 \text{ mol Ba(OH)}_2} = 0.0735 \text{ mol OH}^-$$

$$\text{mol H}^+ = 438.0 \text{ mL HNO}_3 \times \frac{0.200 \text{ mol HNO}_3}{1000 \text{ mL}} \times \frac{1 \text{ mol H}^+}{1 \text{ mol HNO}_3} = 0.0876 \text{ mol H}^+$$

	OH ⁻ (aq) +	H ⁺ (aq)	→ H ₂ O(l)
I	0.0735 mol	.0876 mol	
C	-0.0735 mol	-0.0735 mol	
E	~0	0.0141 mol	

Thus we know that we will be left with 0.0141 mol H⁺ in 683 mL (245+438) or 0.683 L of solution. The pH can be determined as follows:

$$\text{pH} = -\log[\text{H}^+] = -\log \frac{0.0141 \text{ mol H}^+}{0.683 \text{ L}} = -\log(0.02064) = -(-1.685) = 1.685$$

- 98-2 A 50.0 mL sample of 0.100 M NaOH is titrated with 45.0 mL of 0.100 M HCl. What is the pH of the resulting solution?

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Solution:

This is another situation where we have the addition of a strong acid (H^+) to a strong base (OH^-), where we need to consider the stoichiometry of the neutralization reaction before determining the pH.

$$\text{mol } OH^- = 0.050 \text{ L} \times 0.100 \text{ mol } OH^- / \text{L} = 0.0050 \text{ mol } OH^-$$

$$\text{mol } H^+ = 0.045 \text{ L} \times 0.100 \text{ mol } H^+ / \text{L} = 0.0045 \text{ mol } H^+$$

	$OH^-(aq) +$	$H^+(aq)$	$\rightarrow H_2O(l)$
I	0.0050 mol	0.0045 mol	
C	-0.0045 mol	-0.0045 mol	
E	0.00050 mol	~ 0 mol	

$$pOH = -\log[OH^-] = -\log \frac{0.00050 \text{ mol } OH^-}{(0.050+0.045)L} = -\log 0.00526 = -(-2.279) = 2.28$$

Since we know that $pOH + pH = pK_w$, we can easily determine the pH as

$$pH = pK_w - pOH = 14.00 - 2.28 = 11.72$$

- 98-3 A 45.0 mL sample of 0.150 M HCl is added to 38.0 mL of 0.200 M NaOH. What is the $[H^+]$ of the resulting solution? What is the pH? What is the pOH?

Solution:

This is another example of a strong acid being added to a strong base, so that a neutralization reaction will occur.

$$\text{mol } OH^- = 0.0380 \text{ L} \times 0.200 \text{ mol } OH^- / \text{L} = 0.00760 \text{ mol } OH^-$$

$$\text{mol } H^+ = 0.0450 \text{ L} \times 0.150 \text{ mol } H^+ / \text{L} = 0.00675 \text{ mol } H^+$$

	$OH^-(aq) +$	$H^+(aq)$	$\rightarrow H_2O(l)$
I	0.00760 mol	0.00675 mol	
C	-0.00675 mol	-0.00675 mol	
E	0.00085 mol	~ 0 mol	

$$pOH = -\log[OH^-] = -\log \frac{0.00085 \text{ mol } OH^-}{(0.0380 + 0.0450)L} = -\log 0.0102 = -(-1.990) = 1.990$$

Since we know that $pOH + pH = pK_w$, we can easily determine the pH as

$$pH = pK_w - pOH = 14.00 - 1.990 = 12.010$$

The $[H^+]$ can be determined from the pH as follows:

$$pH = -\log[H^+]$$

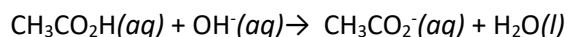
$$12.010 = -\log[\text{H}^+]$$

$$[\text{H}^+] = 10^{-12.010} = 9.80 \times 10^{-13} \text{ M}$$

- 98-4 A 45.0 mL sample of 0.15 M NaOH is added to 38.0 mL of 0.20 M acetic acid. What is the $[\text{H}^+]$ of the resulting solution? What is the pH? What is the pOH?

Solution:

This is another example of a neutralization reaction, but this situation could be a little different since a strong base is added to a weak acid. We start off the same way (looking at the stoichiometry of the neutralization reaction), but the method might be a little different afterwards, depending on the answer that we get.



$$\text{mol OH}^- = 0.0450 \text{ L} \times 0.150 \text{ mol OH}^- / \text{L} = 0.00675 \text{ mol OH}^-$$

$$\text{mol CH}_3\text{CO}_2\text{H} = 0.0380 \text{ L} \times 0.200 \text{ mol CH}_3\text{CO}_2\text{H} / \text{L} = 0.00760 \text{ mol CH}_3\text{CO}_2\text{H}$$

	$\text{CH}_3\text{CO}_2\text{H}(aq)$	$+ \text{OH}^-(aq) \rightarrow$	$\text{CH}_3\text{CO}_2^-(aq)$	$+ \text{H}_2\text{O}(l)$
I	0.00760 mol	0.00675 mol		
C	-0.00675 mol	-0.00675 mol	+0.00675 mol	
E	0.00085 mol	~ 0	0.00675 mol	

We can see that the resulting solution has appreciable quantities of both a weak acid ($\text{CH}_3\text{CO}_2\text{H}$) and its conjugate base (CH_3CO_2^-), so we have created a buffer system. The pH can be determined using the Henderson-Hasselbach equation. Note that the equation technically requires the concentration of the acid and base, but since they are both in the same solution the volumes will cancel and we can just use moles. (You COULD include the volume – (45 mL + 38 mL = 83 mL or 0.083 L, but it is unnecessary to work it out since it will cancel).

$$\text{pH} = \text{pK}_a + \log \frac{[\text{base}]}{[\text{acid}]}$$

$$\text{pH} = 4.76 + \log \frac{0.00675 \text{ mol/volume}}{0.00085 \text{ mol/volume}} = 4.76 + \log 7.941 = 4.76 + 0.8999 = 5.66$$

Since we know that $\text{pOH} + \text{pH} = \text{pK}_w$, we can easily determine the pH as

$$\text{pOH} = \text{pK}_w - \text{pH} = 14.00 - 5.66 = 8.34$$

The $[H^+]$ can be determined from the pH as follows:

$$pH = -\log[H^+]$$

$$5.66 = -\log[H^+]$$

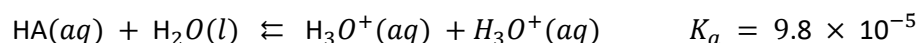
$$[H^+] = 10^{-5.66} = 2.190 \times 10^{-6} \text{ M}$$

98-5 Calculate the pH at the following points in a titration of 40.0 mL of 0.100 M barbituric acid ($K_a = 9.8 \times 10^{-5}$) with 0.100 M KOH.

- (a) no KOH added
- (b) 20 mL of KOH solution added
- (c) 39 mL of KOH solution added
- (d) 40 mL of KOH solution added
- (e) 41 mL of KOH solution added

Solution

(a) Let HA represent barbituric acid and A^- represent the conjugate base. The reaction and equilibrium constant are:



$$K_a = \frac{[H_3O^+][A^-]}{[HA]} = 9.8 \times 10^{-5}$$

The initial and equilibrium concentrations for this system can be written as follows:

	[HA]	$[H_3O^+]$	$[A^-]$
Initial concentration (M)	0.100	0	0
Change (M)	-x	+x	+x
Equilibrium (M)	$0.100 - x$	x	x

Substituting the equilibrium concentrations into the equilibrium expression, and making the assumption that $(0.100 - x) \approx 0.100$, gives:

$$\frac{[H_3O^+][A^-]}{[HA]} = \frac{(x)(x)}{(0.100 - x)} \approx \frac{(x)(x)}{0.100} = 9.8 \times 10^{-5}$$

Solving for x gives $3.13 \times 10^{-3} \text{ M}$. Because this value is less than 5% of 0.100, our assumption is correct. Therefore, $[H_3O^+] = 3.13 \times 10^{-3} \text{ M}$:

$$pH = -\log(3.13 \times 10^{-3}) = 2.504 = 2.50;$$

(b) The moles of OH^- added are given by:

$$\text{mol } OH^- = M \times V = (0.100 \text{ M}) \times (0.020 \text{ L}) = 0.00200 \text{ mol}$$

The initial moles of barbituric acid are given by:

$$\text{mol HA} = M \times V = (0.100 \text{ M}) \times (0.040 \text{ L}) = 0.00400 \text{ mol}$$

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Assume that the added hydroxide ion reacts completely with an equal number of moles of HA, forming an equal number of moles of A⁻ in the process. Thus, the moles of the ions are given by:

$$\begin{aligned}\text{mol HA} &= 0.00400 - 0.00200 = 0.00200 \text{ mol} \\ \text{mol A}^- &= 0.00200 \text{ mol}\end{aligned}$$

The total volume is:

$$40.0 \text{ mL} + 20.0 \text{ mL} = 60.0 \text{ mL} = 0.0600 \text{ L}$$

The initial concentrations of the ions are given by:

$$\begin{aligned}[\text{HA}] &= \frac{0.00200 \text{ mol}}{0.0600 \text{ L}} = 0.0333 \text{ M} \\ [\text{A}^-] &= \frac{0.00200 \text{ mol}}{0.0600 \text{ L}} = 0.0333 \text{ M}\end{aligned}$$

The initial and equilibrium concentrations for this system can be written as follows:

	[HA]	[H ₃ O ⁺]	[A ⁻]
Initial concentration (M)	0.0333	0	0.0333
Change (M)	-x	+x	+x
Equilibrium (M)	0.0333 - x	x	0.0333 + x

Substituting the equilibrium concentrations into the equilibrium expression, and making the assumptions that $(0.0333 - x) \approx 0.0333$ and $(0.0333 + x) \approx 0.0333$, gives:

$$\frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} = \frac{(x)(0.0333 + x)}{(0.0333 - x)} \approx \frac{(x)(0.0333)}{0.0333} = 9.8 \times 10^{-5}$$

Solving for x gives $9.8 \times 10^{-5} \text{ M}$. Because this value is less than 5% of 0.0333, our assumptions are correct. Therefore, $[\text{H}_3\text{O}^+] = 9.8 \times 10^{-5} \text{ M}$:

$$\text{pH} = -\log(9.8 \times 10^{-5}) = 4.009 = 4.01;$$

(c) The moles of OH⁻ added are given by:

$$\text{mol OH}^- = M \times V = (0.100 \text{ M}) \times (0.039 \text{ L}) = 0.00390 \text{ mol}$$

The initial moles of barbituric acid are given by:

$$\text{mol HA} = M \times V = (0.100 \text{ M}) \times (0.040 \text{ L}) = 0.00400 \text{ mol}$$

Assume that the added hydroxide ion reacts completely with an equal number of moles of HA, forming an equal number of moles of A⁻ in the process. Thus, the moles of the ions are given by:

$$\text{mol HA} = 0.00400 - 0.00390 = 0.00010 \text{ mol}$$

$$\text{mol A}^- = 0.00390 \text{ mol}$$

The total volume is:

$$40.0 \text{ mL} + 39.0 \text{ mL} = 79.0 \text{ mL} = 0.0790 \text{ L}$$

The initial concentrations of the ions are given by:

$$\begin{aligned}[\text{HA}] &= \frac{0.00010 \text{ mol}}{0.0790 \text{ L}} = 0.00127 \text{ M} \\ [\text{A}^-] &= \frac{0.00390 \text{ mol}}{0.0790 \text{ L}} = 0.0494 \text{ M}\end{aligned}$$

The initial and equilibrium concentrations for this system can be written as follows:

	[HA]	[H ₃ O ⁺]	[A ⁻]
Initial concentration (M)	0.00127	0	0.0494
Change (M)	-x	+x	+x
Equilibrium (M)	0.00127 - x	x	0.0494 + x

Substituting the equilibrium concentrations into the equilibrium expression, and making the assumption that $(0.00127 - x) \approx 0.00127$ and $(0.0494 + x) \approx 0.0494$, gives:

$$\frac{[H_3O^+][A^-]}{[HA]} = \frac{(x)(0.0494 + x)}{(0.00127 - x)} \approx \frac{(x)(0.0494)}{0.00127} = 9.8 \times 10^{-5}$$

Solving for x gives $2.52 \times 10^{-6} M$. Because this value is less than 5% of 0.00127 and 0.0494, our assumptions are correct. Therefore, $[H_3O^+] = 2.52 \times 10^{-6} M$:

$$pH = -\log(2.52 \times 10^{-6}) = 5.599 = 5.60;$$

(d) The moles of OH⁻ added are given by:

$$\text{mol OH}^- = M \times V = (0.100 M) \times (0.040 L) = 0.00400 \text{ mol}$$

The initial moles of barbituric acid are given by:

$$\text{mol HA} = M \times V = (0.100 M) \times (0.040 L) = 0.00400 \text{ mol}$$

This is the equivalence point, where the moles of base added equal the moles of acid present initially. At the equivalence point:

$$\text{mol A}^- = 0.00400 \text{ mol}$$

The total volume is:

$$40.0 \text{ mL} + 40.0 \text{ mL} = 80.0 \text{ mL} = 0.0800 L$$

The initial concentration of the conjugate base is:

$$[A^-] = \frac{0.00400 \text{ mol}}{0.0800 L} = 0.0500 M$$

The reaction and equilibrium constant are:

$$A^-(aq) + H_2O(l) \rightleftharpoons HA(aq) + OH^-(aq)$$

$$K_b = \frac{[HA][OH^-]}{[A^-]} = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{9.8 \times 10^{-4}} = 1.02 \times 10^{-10}$$

The initial and equilibrium concentrations for this system can be written as follows:

	[A ⁻]	[HA]	[OH ⁻]
Initial concentration (M)	0.0500	0	0
Change (M)	-x	+x	+x
Equilibrium (M)	0.0500 - x	x	x

Substituting the equilibrium concentrations into the equilibrium expression, and making the assumption that $(0.0500 - x) \approx 0.0500$, gives:

$$\frac{[HA][OH^-]}{[A^-]} = \frac{(x)(x)}{(0.0500 - x)} \approx \frac{(x)(x)}{0.0500} = 1.02 \times 10^{-10}$$

Solving for x gives $2.26 \times 10^{-6} M$. Because this value is less than 5% of 0.0500, our assumption is correct. Therefore, $[OH^-] = 2.26 \times 10^{-6} M$:

$$pOH = -\log(2.26 \times 10^{-6}) = 5.646$$

$$pH = 14.000 - pOH = 14.000 - 5.646 = 8.354 = 8.35;$$

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(e) The moles of OH^- added are given by:

$$\text{mol OH}^- = M \times V = (0.100 \text{ M}) \times (0.041 \text{ L}) = 0.00410 \text{ mol}$$

The initial moles of barbituric acid are given by:

$$\text{mol HA} = M \times V = (0.100 \text{ M}) \times (0.040 \text{ L}) = 0.00400 \text{ mol}$$

This is past the equivalence point, where the moles of base added exceed the moles of acid present initially. The excess moles of hydroxide ion are given by:

$$\text{mol OH}^- = 0.00410 - 0.00400 = 0.00010 \text{ mol}$$

The total volume is:

$$40.0 \text{ mL} + 41.0 \text{ mL} = 81.0 \text{ mL} = 0.0810 \text{ L}$$

The concentration of OH^- is:

$$\begin{aligned} [\text{OH}^-] &= \frac{0.00010 \text{ mol}}{0.0810 \text{ L}} = 0.0012 \text{ M} \\ \text{pOH} &= -\log(0.0012) = 2.921 \\ \text{pH} &= 14.000 - \text{pOH} = 14.000 - 2.921 = 11.079 = 11.08 \end{aligned}$$

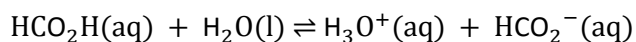
98-6 Consider the titration of 100.0 mL of 0.150 M formic acid (HCOOH , $K_a = 1.8 \times 10^{-4}$) with 0.100 M NaOH. What is the pH at the following points?

- (a) The start
- (b) At the midpoint
- (c) After 25.0 mL of the NaOH has been added
- (d) After 90.0 mL of the NaOH has been added.
- (e) At the equivalence point.

Solution:

(a) At the start of the titration, we have a solution that only contains a weak acid. We can determine the pH using the K_a expression:

The reaction is:



The equilibrium expression is:

$$K_a = \frac{[\text{HCO}_2^-][\text{H}_3\text{O}^+]}{[\text{HCO}_2\text{H}]} = 1.8 \times 10^{-4}$$

The initial and equilibrium concentrations for this system can be written as follows:

	$\text{HCO}_2\text{H} \rightleftharpoons$	H_3O^+	HCO_2^-
I	0.150 M	~0	0
C	-x	+x	+x
E	0.150 - x	x	x

$$K_a = \frac{[\text{HCO}_2^-][\text{H}_3\text{O}^+]}{[\text{HCO}_2\text{H}]}$$

$$1.8 \times 10^{-4} = \frac{(x)(x)}{(0.150 - x)}$$

Since the K_a value is so small, we can assume that x would be negligible compared to 0.150 so that we can assume that we can neglect it. The equation thus becomes:

$$1.8 \times 10^{-4} = \frac{(x)(x)}{(0.150)}$$

$$2.7 \times 10^{-5} = x^2$$

$$\sqrt{2.7 \times 10^{-5}} = \sqrt{x^2}$$

$$x = 5.2 \times 10^{-3} \text{ M}$$

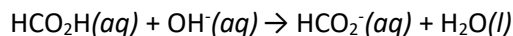
Thus x is very small compared to the [] of the acid, so our assumption that we could neglect x was valid.

Since $x = [\text{H}^+] = 5.20 \times 10^{-3}$. The $\text{pH} = -\log[\text{H}^+] = -\log(5.20 \times 10^{-3}) = -(-2.28) = 2.28$

- (b) At the midpoint, the acid would be “half-titrated”, so that the [acid] = [conjugate base]. When this occurs, the $\text{pH} = \text{p}K_a$.

$$\text{p}K_a = -\log(K_a) = -\log(1.8 \times 10^{-4}) = 3.74.$$

- (c) The addition of 25.0 mL of 0.100 M NaOH to 100.0 mL of 0.150 M formic acid will cause a neutralization reaction.



Depending on the stoichiometry of reaction, both formic acid and the conjugate base will be present in appreciable quantities, and so a buffer could be formed (depending on the ratio).

$$\text{mol OH}^- = 0.0250 \text{ L} \times 0.100 \text{ mol OH}^- / \text{L} = 0.00250 \text{ mol OH}^-$$

$$\text{mol HCO}_2\text{H} = 0.1000 \text{ L} \times 0.150 \text{ mol HCO}_2\text{H} / \text{L} = 0.0150 \text{ mol HCO}_2\text{H}$$

	$\text{HCO}_2\text{H}(aq)$	$+ \text{OH}^-(aq) \rightarrow$	$\text{HCO}_2^-(aq)$	$+ \text{H}_2\text{O}(l)$
I	0.0150 mol	0.00250 mol		
C	-0.00250 mol	-0.00250 mol	+0.00250 mol	
E	0.0125 mol	~0	0.00250 mol	

We can use the H-H equation to determine the pH. Since the volumes cancel out, we need only concern ourselves with the ratio of moles:

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

$$\text{pH} = 3.74 + \log \frac{[0.00250 \text{ mol/volume}]}{[0.0125 \text{ mol/volume}]}$$

$$\text{pH} = 3.74 + \log 0.200 = 3.74 + (-0.699) = 3.04$$

Note that this pH is less than the pH at the midpoint (part b), which makes sense since we are not yet at the midpoint.

- (d) The calculation of the pH after 90.0 mL of base has been added is very similar to the one that we just did.

$$\text{mol OH}^- = 0.0900 \text{ L} \times 0.100 \text{ mol OH}^- / \text{L} = 0.00900 \text{ mol OH}^-$$

$$\text{mol HCO}_2\text{H} = 0.1000 \text{ L} \times 0.150 \text{ mol HCO}_2\text{H} / \text{L} = 0.0150 \text{ mol HCO}_2\text{H}$$

	HCO ₂ H(aq)	+ OH ⁻ (aq) →	HCO ₂ ⁻ (aq)	+ H ₂ O(l)
I	0.0150 mol	0.00900 mol		
C	-0.00900 mol	-0.00900 mol	+0.00900 mol	
E	.00600 mol	~0	0.00900 mol	

We can use the H-H equation to determine the pH. Since the volumes cancel out, we need only concern ourselves with the ratio of moles:

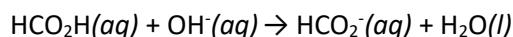
$$\text{pH} = \text{pKa} + \log \frac{[\text{base}]}{[\text{acid}]}$$

$$\text{pH} = 3.74 + \log \frac{[0.00900 \text{ mol/volume}]}{[0.00600 \text{ mol/volume}]}$$

$$\text{pH} = 3.74 + \log 1.5 = 3.74 + (0.176) = 3.92$$

Here we have a pH > pKa, which makes sense since passed the midpoint (where the pH = pKa).

- (e) At the equivalence point, enough base has been added to completely titrate the weak acid to its conjugate base.



Since the number of moles of acid are 0.0150 mol

$$\text{mol HCO}_2\text{H} = 0.1000 \text{ L} \times 0.150 \text{ mol HCO}_2\text{H} / \text{L} = 0.0150 \text{ mol HCO}_2\text{H}$$

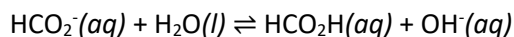
according to the stoichiometry of the neutralization reaction, it will required 0.0150 mol of hydroxide to attain the equivalence point. The volume of 0.100 M NaOH required to provide 0.0150 mol of base can be determined as follows:

$$\text{volume of OH}^- = 0.0150 \text{ mol OH}^- \times \frac{1 \text{ L}}{0.100 \text{ mol}} = 0.150 \text{ L}$$

Thus at the equivalence point, the pH-determining species will be the formate (HCO₂⁻), which will be at a concentration of 0.0150 mol in 0.250 L (0.100 L + 0.150 L), or 0.0600 M.

The formate is, itself, a weak base and so will react with water as indicated:

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We can determine the pH of the solution using the K_b expression. Note that we can determine the K_b since $K_a \times K_b = K_w$.

	HCO_2^-	$\rightleftharpoons \text{HCO}_2\text{H}$	$\text{OH}^-(aq)$
I	0.0600 M	~ 0	0
C	-x	+x	+x
E	0.0600 M - x	x	x

$$K_b = \frac{[\text{HCO}_2\text{H}][\text{OH}^-]}{[\text{HCO}_2^-]}$$

$$\frac{1 \times 10^{-14}}{1.8 \times 10^{-4}} = \frac{(x)(x)}{(0.0600 - x)}$$

Since the K_b value is so small, we can assume that x would be negligible compared to 0.0600 so that we can assume that we can neglect it. The equation thus becomes:

$$5.56 \times 10^{-11} = \frac{(x)(x)}{(0.0600)}$$

$$3.33 \times 10^{-12} = x^2$$

$$\sqrt{3.33 \times 10^{-12}} = \sqrt{x^2}$$

$$x = 1.82 \times 10^{-6}$$

Thus x is <5% of the [] of the base, so our assumption that we could neglect x was valid.

Since $x = [\text{OH}^-] = 1.82 \times 10^{-6}$; The $\text{pOH} = -\log[\text{OH}^-] = -\log(1.82 \times 10^{-6}) = -(-5.74) = 5.74$
 $\text{pH} = \text{p}K_w - \text{pOH} = 14.00 - 5.74 = 8.26$

This number looks about right, since we expect that the equivalence point of the titration of a weak acid with a strong base will be slightly basic.

98-7 Explain how to choose the appropriate acid-base indicator for the titration of a weak base with a strong acid.

Solution

At the equivalence point in the titration of a weak base with a strong acid, the resulting solution is slightly acidic due to the presence of the conjugate acid. Thus, pick an indicator that changes color in the acidic range and brackets the pH at the equivalence point. Methyl orange is a good example

98-8 Explain why an acid-base indicator changes color over a range of pH values rather than at a specific pH.

Solution

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Acid-Base Equilibria*

Section 98-Titrations-Questions and Worked Answers

The color in an indicator depends on the ratio of nonionized to ionized forms. The predominant color in acid requires at least a 1:10 ratio of forms, whereas the predominant color in base requires a 10:1 ratio of forms. The transition from one condition to the other comes about by addition of acid or base requiring, in general, a change of two pH units between colors.