

[TP] At 25 °C, 0.10 mol each of a strong acid and OH^- are combined in 1.0 L of water. The pH of the solution is ...

25% 1. < 7
25% 2. $= 7$
25% 3. > 7
25% 4. Further information needed

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Lecture 20 CH102 A2 (MWF 11:15 am)

Monday, March 18, 2019

- Review: [$H3O^+$] when "just enough" base added
- [H_3O^+] when "too much" base added;
- How much strong acid remains unreacted?

Next: Titration: Adding successive amounts of base; Practice: Too little, just enough, too much?; [H_3O^+] when different amounts of "not enough" base added

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Review: [H_3O^+] when "just enough" base added

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[Quiz] At 25 °C, $V_b = 200.$ mL of $c_b = 0.30$ M of OH^- is combined with $V_a = 100.$ mL of $c_a = 0.60$ M of HA, $K_a = 4.0 \times 10^{-6}$ and $K_b = K_w / K_a = 2.5 \times 10^{-9}$. The equilibrium value of [OH^-] is ...

17% 1. 2.0×10^{-4}
17% 2. 2.2×10^{-5}
17% 3. 4.0×10^{-6}
17% 4. 1.0×10^{-7}
17% 5. 2.5×10^{-9}
17% 6. Something else

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Step 2: $[H_3O^+]$ when "just enough" base added

At 25 °C, $V_b = 200.$ mL of $c_b = 0.30$ M of OH^- is combined with $V_a = 100.$ mL of $c_a = 0.60$ M of HA, $K_a = 4.0 \times 10^{-6}$ and $K_b = K_w \,/\, K_a = 2.5 \times 10^{-9}$.

	$A^-(aq)$	$HA(aq)$	$OH^-(aq)$	Q
Initial	0.20	0	10^{-7}	$0 < K_b$
Change				
Equilibrium				
Approximate				

Step 2: $[H_3O^+]$ when "just enough" base added

At 25 °C, $V_b = 200.$ mL of $c_b = 0.30$ M of OH^- is combined with $V_a = 100.$ mL of $c_a = 0.60$ M of HA, $K_a = 4.0 \times 10^{-6}$ and $K_b = K_w \,/\, K_a = 2.5 \times 10^{-9}$.

	$A^-(aq)$	$HA(aq)$	$OH^-(aq)$	Q
Initial	0.20	0	10^{-7}	$0 < K_b$
Change	$-x$	$+x$	$+x$	
Equilibrium	$0.20 - x$	x	$10^{-7} + x$	K_b
Approximate	≈ 0.20	x	$\approx x$	K_b

Step 2: $[H_3O^+]$ when "just enough" base added

At 25 °C, $V_b = 200.$ mL of $c_b = 0.30$ M of OH^- is combined with $V_a = 100.$ mL of $c_a = 0.60$ M of HA, $K_a = 4.0 \times 10^{-6}$ and $K_b = K_w \,/\, K_a = 2.5 \times 10^{-9}$.

	$A^-(aq)$	$HA(aq)$	$OH^-(aq)$	Q
Initial	0.20	0	10^{-7}	$0 < K_b$
Change	$-x$	$+x$	$+x$	
Equilibrium	$0.20 - x$	x	$10^{-7} + x$	K_b
Approximate	≈ 0.20	x	$\approx x$	K_b

$$[OH^-] = x = \sqrt{K_b \times [A^-]} = \sqrt{2.5 \times 10^{-9} \times 0.20} = 2.2 \times 10^{-5}$$

$$[H_3O^+] = K_w \,/\, [OH^-] = \frac{1.0 \times 10^{-14}}{2.2 \times 10^{-5}} = 4.5 \times 10^{-10}$$

$[H_3O^+]$ when "too much" base added

Step 2: $[H_3O^+]$ when "too much" base added

$V_b = 200.$ mL of $c_b = 0.30$ M of OH^- is combined with $V_a = 200.$ mL of $c_a = 0.20$ M of HA, $K_a = 1.0 \times 10^{-6}$ and $K_b = K_w / K_a = 1.0 \times 10^{-8}$.

Step 2: $[H_3O^+]$ when "too much" base added

$V_b = 200.$ mL of $c_b = 0.30$ M of OH^- is combined with $V_a = 200.$ mL of $c_a = 0.20$ M of HA, $K_a = 1.0 \times 10^{-6}$ and $K_b = K_w / K_a = 1.0 \times 10^{-8}$.

Step 1 results: Ignoring acid-base equilibrium, after the limiting reagent reaction is complete, the resulting concentrations of HA, A^-, and OH^- are ...

$[HA] =$

$[A^-] =$

$[OH^-] =$

Step 2: $[H_3O^+]$ when "too much" base added

$V_b = 200.$ mL of $c_b = 0.30$ M of OH^- is combined with $V_a = 200.$ mL of $c_a = 0.20$ M of HA, $K_a = 1.0 \times 10^{-6}$ and $K_b = K_w / K_a = 1.0 \times 10^{-8}$.

Step 1 results: Ignoring acid-base equilibrium, after the limiting reagent reaction is complete, the resulting concentrations of HA, A^-, and OH^- are ...

$[HA] = 0$

$[A^-] = \frac{c_a V_a}{V_a + V_b} = \frac{0.040 \text{ mol}}{0.400 \text{ L}} = 0.10$ M

$[OH^-] = \frac{c_b V_b - c_a V_a}{V_a + V_b} = \frac{(0.060 - 0.040) \text{ mol}}{0.400 \text{ L}} = 0.050$ M ("too much" base added)

Step 2: $[H_3O^+]$ when "too much" base added

If conjugate base and excess $[OH^-]$ are both present, then always work with

$A^-(aq) + H_2O(l) \rightleftarrows HA(aq) + OH^-(aq)$, $K_b = K_w/K_a$

Step 2: $[H_3O^+]$ when "too much" base added

$V_b = 200.$ mL of $c_b = 0.30$ M of OH^- is combined with $V_a = 200.$ mL of $c_a = 0.20$ M of HA, $K_a = 1.0 \times 10^{-6}$ and $K_b = K_w / K_a = 1.0 \times 10^{-8}$.

	$A^-(aq)$	$HA(aq)$	$OH^-(aq)$	Q
Initial				
Change				
Equilibrium				
Approximate				

$[OH^-] = ?$

$[H_3O^+] = ?$

$[HA] = ?$

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Step 2: $[H_3O^+]$ when "too much" base added

$V_b = 200.$ mL of $c_b = 0.30$ M of OH^- is combined with $V_a = 200.$ mL of $c_a = 0.20$ M of HA, $K_a = 1.0 \times 10^{-6}$ and $K_b = K_w / K_a = 1.0 \times 10^{-8}$.

	$A^-(aq)$	$HA(aq)$	$OH^-(aq)$	Q
Initial	0.10	0	0.050	0
Change	$-x$	$+x$	$+x$	
Equilibrium	$0.10 - x$	x	$0.050 + x$	K_b
Approximate	≈ 0.10	x	≈ 0.050	K_b

$[OH^-] = 0.050$ (easy!)

$[H_3O^+] = K_w / [OH^-] = \frac{1.0\times10^{-14}}{0.050} = 2.0 \times 10^{-13}$

$[HA] = x = \frac{K_b[A^-]}{[OH^-]} = \frac{1.0\times10^{-8}\times0.10}{0.050} = 2.0 \times 10^{-8}$ (tiny!)

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Unreacted strong acid

Strong acids have $K_a \gg 1$ and so react with water with nearly 100% yield.

However, there must always be some unreacted acid.

Let's see how to calculate the molarity of unreacted acid.

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Revised initial method for strong acid

An acid, HA, has $K_a = 1$ x 10^5. Calculate the molarity of unreacted $HA(aq)$ in a 0.02 M solution of HA.

$[H_3O^+] = ?$

$[A^-] = ?$

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Revised initial method for strong acid

An acid, HA, has $K_a = 1 \times 10^5$. Calculate the molarity of unreacted HA(aq) in a 0.02 M solution of HA.

$[H_3O^+] = 0.02$

$[A^-] = 0.02$

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Revised initial method for strong acid

An acid, HA, has $K_a = 1 \times 10^5$. Calculate the molarity of unreacted HA(aq) in a 0.02 M solution of HA.

$[H_3O^+] = 0.02$

$[A^-] = 0.02$

$[HA] = ?$

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Revised initial method for strong acid

An acid, HA, has $K_a = 1 \times 10^5$. Calculate the molarity of unreacted HA(aq) in a 0.02 M solution of HA.

$[H_3O^+] = 0.02$

$[A^-] = 0.02$

$[HA] = 4 \times 10^{-9}$

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