

[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

Response Counter

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Lecture 21 CH102 A1 (MWF 9:05 am)

Friday, March 17, 2017

See Overview of acid-base calculations, https://goo.gl/GEPDqo

- $[H_3O^+]$ when "too much" base added
- Practice: Too little, just enough, too much?
- More practice
- $[H_3O^+]$ when different amounts of "not enough" base added

Next lecture: Complete ch14; begin ch15: Solubility, precipitation, and complexation

$[H_3O^+]$ 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 …

$[HA] = 0$

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

$[OH^-] = (c_b V_b - c_a V_a) / (V_a + V_b) = (0.060 - 0.040)\text{ mol} / (0.400\text{ L})$

$= 0.020\text{ mol} / (0.400\text{ L}) = 0.050\text{ M}$

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

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

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

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

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	$A^-(aq)$	$HA(aq)$	$OH^-(aq)$	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	$0.050 + x$	1.0×10^{-8}
Approximate	≈ 0.10	x	≈ 0.050	1.0×10^{-8}

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

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

$[HA] = x = K_b [A^-] / [OH^-] = 1.0 \times 10^{-8} \times 0.10 / 0.050 = 2.0 \times 10^{-8}$ (tiny!)

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Too little, just enough, too much?

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25% 1. < 7
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25% 4. Further information needed

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[TP] At 25 ºC, 0.10 mol each of a weak acid HA, its conjugate base A^-, 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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[Quiz] At 25 ºC, 0.10 mol each of a weak acid HA and a strong acid HB, and 0.20 mol of 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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25% 1. < 7
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Response Counter

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[Quiz] An acid HA has $K_a = 1.0 \times 10^{-4}$ at 25 °C. An aqueous solution is made by combining of 0.137 moles each of HA and A^- in a total volume of 250 mL at 25 °C. The pH of the solution is …

17% 1. 1
17% 2. 4
17% 3. 7
17% 4. 10
17% 5. 13
17% 6. More information needed

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Different amounts of "not enough" base

At 25 °C, the pH of a 1.0 L solution of $c_a = c_b = 1.00$ M, $K_a = 1 \times 10^{-5}$ is …
pH = 5.00

Add 100. mL of 0.100 M NaOH …
$\downarrow HA(aq) + OH^-(aq) \rightarrow H_2O(l) + \uparrow A^-(aq)$
HA → 1.00 mol – 0.010 mol = 0.99 mol
A^- → 1.00 mol + 0.010 mol = 1.01 mol

The pH of a 1.0 L solution of $c_a = 0.99$ mol/1.10 L, $c_b = 1.01$ mol/1.10 L is …
$c_a/c_b = 1.00 \rightarrow 0.99/1.01$, pH → 5.01 (tiny change!)

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Different amounts of "not enough" base

At 25 °C, the pH of a 1.0 L solution pure water, $K_a = 1 \times 10^{-14}$ is …
pH = 7.00

Add 100. mL of 0.100 M NaOH …
$[OH^-] = 0.010$ mol/1.10 L = 0.0091
pOH = 2.04, pH = 11.96 (huge change!)

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Different amounts of "not enough" base

At 25 °C, the pH of a 1.0 L solution of $c_a = c_b = 1.00$ M, $K_a = 1 \times 10^{-5}$ is ...

pH = 5.00

Add 100. mL of 0.100 M HCl ...

$HCl(aq) + \downarrow A^-(aq) \rightarrow \uparrow HA(aq) + Cl^-(aq)$

HA → 1.00 mol + 0.010 mol = 1.01 mol

A^- → 1.00 mol – 0.010 mol = 0.99 mol

The pH of a 1.0 L solution of $c_a = 1.01$ mol/1.10 L, $c_b = 0.99$ mol/1.10 L ...

$c_a/c_b = 1.00 \rightarrow 1.01/0.99$, pH → 4.99 (tiny change!)

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Different amounts of "not enough" base

At 25 °C, the pH of a 1.0 L solution pure water, $K_a = 1 \times 10^{-14}$ is ...

pH = 7.00

Add 100. mL of 0.100 M HCl ...

$[H_3O^+] = 0.010$ mol/1.10 L = 0.0091

pH = 2.04 (huge change!)

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Buffers

Mixtures of a weak acid and its conjugate base (or weak base and its conjugate acid) exhibit the special property that they resist changes to pH.

For this reason such mixtures are known as buffers.

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Buffers resist change in pH

Added strong base (say, OH^-) is gobbled up ...

$\downarrow HA(aq) + OH^-(aq) \rightarrow H_2O(l) + \uparrow A^-(aq)$

c_a lowered, c_b raised, c_a/c_b lowered

Added strong acid (say, HCl) is gobbled up ...

$HCl(aq) + \downarrow A^-(aq) \rightarrow \uparrow HA(aq) + Cl^-(aq)$

c_b lowered, c_a raised, c_a/c_b raised

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