

[TP] At 25 °C the $pK_a = -\log(K_a)$ of a certain acid is 4.17. A solution is made by combining 0.314 mol of HA and 0.314 mol of NaA in a total volume of 716 mL. What is the pH of the resulting solution?

14%	1.	1.78
14%	2.	2.19
14%	3.	3.14
14%	4.	4.17
14%	5.	5.78
14%	6.	7.00
14%	7.	Something else

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Lecture 17 CH102 A2 (MWF 11:15 am)

Monday, March 4, 2019

- $pK_a = -\log(K_a)$
- Titration: What happens when some OH^- is added to an acid

Next: $[H3O^+]$ when "too little" base added; $[H_3O^+]$ when "too much" base added; Practice: Too little, just enough, too much?; $[H_3O^+]$ when different amounts of "not enough" base added

Review: Logarithm Tutorial, https://goo.gl/RtLg9X

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$pK_a = -\log(K_a)$

Since $K_a = [H_3O^+]\,[A^-] / [HA]$, we can express the hydronium ion concentration as $[H_3O^+] = K_a\ [HA] / [A^-]$

From this we can express pH as

$$pH = -\log(K_a) + \log(\,[A^-] / [HA]\,)$$

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$pK_a = -\log(K_a)$

Since $K_a = [H_3O^+]\,[A^-] / [HA]$, we can express the hydronium ion concentration as $[H_3O^+] = K_a\ [HA] / [A^-]$

From this we can express pH as

$$pH = -\log(K_a) + \log(\,[A^-] / [HA]\,)$$

$$= pK_a + \log(\,[A^-] / [HA]\,)$$

where pK_a is defined as $pK_a = -\log(K_a)$

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$pK_a = -\log(K_a)$

An important application of

$pH = pK_a + \log([A^-] / [HA])$

is that …

if equal amounts of weak acid and its conjugate base are combined, then

$pH = pK_a + \log(1) = pK_a$

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[TP] At 25 °C the $pK_a = -\log(K_a)$ of a certain acid is 4.17. A solution is made by combining 0.314 mol of HA and 0.314 mol of NaA in a total volume of 716 mL. What is the pH of the resulting solution?

0% 1. 1.78
0% 2. 2.19
0% 3. 3.14
0% 4. 4.17
0% 5. 5.78
0% 6. 7.00
0% 7. Something else

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$pK_a = -\log(K_a)$

Since $pH = pK_a + \log([A^-] / [HA])$, when equal amounts of acid and its conjugate based are combine in solution (an equimolar solution), the pH will be pK_a.

This means that a way to measure K_a is to make an equimolar solution and measure its pH, for then

$K_a = 10^{-pH}$

Very useful to remember!

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Titration: What happens when some OH^- is added to an acid?

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[Quiz] An acid has $K_a = 1.0 \times 10^{-7}$ at 25 °C. The hydronium concentration of a $c_a = 0.40$ M solution of this acid solution is $[H_3O^+] = 2.0 \times 10^{-4}$. After adding some 0.20 M OH^-, $[H_3O^+]$ must be ...

0% 1. smaller than 2.0×10^{-4}
0% 2. less than 1.0×10^{-7} (basic)
0% 3. equal to 1.0×10^{-7} (neutral)
0% 4. greater than 1.0×10^{-7} (acidic)
0% 5. More information needed

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Add "some" OH^- to an acid

When some OH^- solution is added to an HA solution, some HA will be converted to conjugate base, A^-.

Therefore, [HA] will be less, the equilibrium

$$HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$$

will shift toward reactants, and so, by Le Chatelier, $[H_3O^+]$ must decrease.

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Add "some" OH^- to an acid

When some OH^- solution is added to an HA solution, some HA will be converted to conjugate base A^-, and so $[H_3O^+]$ must decrease.

There are two steps to find the new value of $[H_3O^+]$.

Step 1: Assume HA and the added OH^- react 100%, as a limiting reagent problem.

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Add "some" OH^- to an acid

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There are two steps to find the new value of $[H_3O^+]$.

Step 1: Assume HA and the added OH^- react 100%, as a limiting reagent problem.

Step 2: Use the results of step 1 to re-solve the equilibrium.

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Add "some" OH^- to an acid

When some OH^- solution is added to an HA solution, some HA will be converted to conjugate base A^-, and so $[H_3O^+]$ must decrease.

There are two steps to find the new value of $[H_3O^+]$.

Step 1: Assume HA and the added OH^- react 100%, as a limiting reagent problem.

Step 2: Use the results of step 1 to re-solve the equilibrium.

To begin, let's learn about the results of step 1.

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Step 1: HA + OH^- limiting reagent reaction

There are three possible results of step 1.

"Too little" base:

If fewer moles of OH^- are added than the moles of acid present, then only some of the acid HA will be converted to conjugate base A^-, consuming all of the added OH^-.

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Step 1: HA + OH^- limiting reagent reaction

There are three possible results of step 1.

"Just enough" base:

If the moles of OH^- added is the same as the moles of acid present, then all of the acid HA will be converted to conjugate base A^-, consuming all of the added OH^-.

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Step 1: HA + OH^- limiting reagent reaction

There are three possible results of step 1.

"Too much" base:

If the moles of OH^- added is more than the moles of acid present, then all of the acid HA will be converted to conjugate base A^- and unused OH^- will be left over.

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Step 1: HA + OH^- limiting reagent reaction

There are three possible results of step 1.

	$HA(aq)$	$OH^-(aq)$	$A^-(aq)$
"Too little"			
"Just enough"			
"Too much"			

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Step 1: HA + OH^- limiting reagent reaction

There are three possible results of step 1.

	$HA(aq)$	$OH^-(aq)$	$A^-(aq)$
"Too little"	Some	None	Some
"Just enough"	None	None	Some
"Too much"	None	Some	Some

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[TP] V_b = 100. mL of c_b = 0.40 M of OH^- is combined with V_a = 100. mL of c_a = 0.40 M of HA. The result is …

0% 1. "too little" OH^-
0% 2. "just enough" OH^-
0% 3. "too much" OH^-
0% 4. Further information needed

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[Quiz] V_b = 100. mL of c_b = 0.20 M of OH^- is combined with V_a = 100. mL of c_a = 0.40 M of HA. The result is …

0% 1. "too little" OH^-
0% 2. "just enough" OH^-
0% 3. "too much" OH^-
0% 4. Further information needed

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Add “some” OH^- to an acid

There are two steps to find the new value of $[H_3O^+]$.

Step 1: Assume the acid HA and the added OH^- react 100%, as a limiting reagent problem (ignoring equilibrium!).

Step 2: Based on whether we have added “too little”, “just enough”, or “too much” OH^-, we use the results of step 1 to re-calculate the equilibrium concentration of $[H_3O^+]$.

