

Lecture 18 CH102 A1 (MWF 9 am) Spring 2016 Copyright © 2016 Dan Dill dan@bu.edu

[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 18 CH102 A1 (MWF 9:05 am)
 Friday, March 3, 2017

- Complete: Getting weak acid K_a values
- $pK_a = -\log(K_a)$

Next lecture: Using K_a to get $[H_3O^+]$. Practice: $[H_3O^+]$ of a weak acid. What happens when some OH^- is added to an acid?

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Weak acid K_a values

We can get the numerical value of K_a by analyzing the acid equilibrium.

	HA(aq)	H ₃ O ⁺ (aq)	A ⁻ (aq)	Q
Initial	c_a	10^{-7}	0	0
Change	$-x$	$+x$	$+x$	
Equilibrium	$c_a - x$	$10^{-7} + x$	x	K_a

The result is a quadratic equation in x ,

$$K_a = (10^{-7} + x)(x) / (c_a - x)$$

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Approximate	?	?	x	K_a

Let's see how we can approximate the equilibrium concentrations, to simplify equation we need to solve for x .

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Approximate	?	?	x	K_a

Since $K_a \ll 1$, what is a good approximation to $c_a - x$?



9

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Approximate	$c_a - x \approx c_a$?	x	K_a

Since $K_a \ll 1$, the acid reacts with water **only a little bit**.

Therefore $c_a - x \approx c_a$



10

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Approximate	$c_a - x \approx c_a$?	x	K_a

Since $K_a \gg K_w$, what is a good approximation to $10^{-7} + x$?



11

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Weak acid K_a values

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Change	$-x$	$+x$	$+x$	
Equilibrium	$c_a - x$	$10^{-7} + x$	x	K_a
Approximate	$c_a - x \approx c_a$	$10^{-7} + x \approx x$	x	K_a

Since $K_a \gg K_w$, HA is a **much greater source of H₃O⁺** than water itself.

Therefore $10^{-7} + x \approx x$



12

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With these simplifications, the acid equilibrium constant is ...

$$K_a = [\text{H}_3\text{O}^+][\text{A}^-] / [\text{HA}] = x^2 / c_a = (10^{-\text{pH}})^2 / c_a$$



13

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Approximate	$c_a - x \approx c_a$	$10^{-7} + x \approx x$	x	K_a

For example, if $c_a = 0.10$ and $\text{pH} = 4$, then ...

$$K_a = (10^{-\text{pH}})^2 / c_a = (10^{-4.0})^2 / 0.10 = 1.0 \times 10^{-7}$$



14

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$\text{p}K_a = -\log(K_a)$

Since $K_a = [\text{H}_3\text{O}^+][\text{A}^-] / [\text{HA}]$ we can express the hydronium ion concentration as

$$[\text{H}_3\text{O}^+] = K_a [\text{HA}] / [\text{A}^-]$$

From this we can express pH as

$$\begin{aligned} \text{pH} &= -\log(K_a) + \log([\text{A}^-] / [\text{HA}]) \\ &= \text{p}K_a + \log([\text{A}^-] / [\text{HA}]) \end{aligned}$$

where $\text{p}K_a$ is defined as

$$\text{p}K_a = -\log(K_a)$$



15

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10

16

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$$\text{p}K_a = -\log(K_a)$$

Since

$$\text{pH} = \text{p}K_a + \log([A^-] / [HA])$$

when equal amounts of acid and its conjugate base are combined in solution (an **equimolar solution**), the pH will be $\text{p}K_a$!

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

$$K_a = 10^{-\text{pH}}$$

Very useful to remember!



17