K_a and Weak Acid Problems

Do your best on these questions and then go to the solutions at the end to see how the problem can be solved.
I also have videos for each problem the are posted at ChemAdvantage AP14.

1. Calculate the K_a and pK_a values of a 0.10 M solution of a weak acid with an [H^+] of 1.30 ×$10^{-3}M$.

Don’t skip steps. Write out the ICE. You will have to explain yourself on the FRQ, so get used to doing this!

$$HA(aq) \rightleftarrows H^+(aq) + A^-(aq)$$

Initial

Change

Equilibrium

Now you can easily fill out the entire chart.

	$HA(aq)$	$\rightleftarrows$	$H^+(aq)$	+	$A^-(aq)$
Initial					
Change					+ 0.00130 M
Equilibrium					

$$K_a = \frac{[H^+][OH^-]}{[HA]}$$ @ Equilibrium

K_a =

pK_a =

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2. Calculate the K_a and pK_a values of a 0.20 M weak acid solution with a pH of 5.03.

[H^+] =

	$HA(aq)$	$\rightleftarrows$	$H^+(aq)$	+	$A^-(aq)$
Initial					0
Change					
Equilibrium					

$$K_a = \frac{[H^+][OH^-]}{[HA]} \text{ @ Equilibrium}$$

K_a =

pK_a =

Now for two problems that have you predict the pH of weak acid solutions based on the strength of the acid.

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3. If the K_a of formic acid is 1.77×10^{-4}, what is the pH of a 1.0 *M* solution of formic acid?

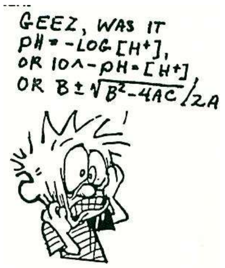
$$\text{HFormic}(aq) + H_2O(l) \rightleftarrows H_3O^+(aq) + \text{Formic}^-(aq)$$

Initial

Change

Equilibrium

K_a =



	$[H_3O^+]$ =	pH =

$[H^+] = \sqrt{K_a \times [HA]}$

$[H^+] = \sqrt{\quad ?}$

$[H^+] =$

This equation could have been used to quickly solve the problem.

$[H^+] = \sqrt{K_a \times [HA]}$ can solve these problems quickly

However, in the FRQ section, points are awarded for showing your step-by-step method of solving the problem.

So be sure to do your ICE Chart even though this approximation equation can solve the question quickly.

4. If the K_a of acetic acid is 1.8×10^{-5}, what is the pH of a 0.010 *M* solution of acetic acid?

$$\text{HAcetic}(aq) \quad + \quad H_2O(l) \quad \rightleftarrows \quad H_3O^+(aq) \quad + \quad \text{Acetic}^-(aq)$$

Initial

Change

Equilibrium

	$[H^+] = \sqrt{K_a \times [HA]}$ $[H^+] = \sqrt{\quad ? \quad}$ $[H^+] =$		**- pH**

K_b and Weak Bases

Weak bases require a slightly more complex ICE Chart since water must be included to show the ionization producing hydroxide ions.

5. Calculate the K_b and pK_b of a 0.250 *M* solution of a weak base, A^-, with a pH of 9.68.

pOH=14 - 9.68 = 4.32 $[OH^-]=10^{-4.32}\,M = 4.8\times10^{-5}M$

	$A^-(aq)$	+	H_2O	$\rightleftarrows$	$HA(aq)$	+	$OH^-(aq)$
Initial							
Change							
Equilibrium							

$$K_b = \frac{[HA][OH^-]}{[A^-]} \text{ @ Equilibrium}$$

$K_b =$

pK_b =

The pK_a for the conjugate acid in this problem, HA, will be

The K_a for the conjugate acid in this problem, HA, will be

Solutions to the Problems… don’t look at these till you have tried them already.

1. Calculate the K_a and pK_a values of a 0.10 M solution of a weak acid with an $[H^+]$ of $1.30 \times 10^{-3} M$.

Don’t get confused; write down what you know.

0.10 M solution … the given concentration of solution is always the initial concentration.
$[H^+]$ of $1.30 \times 10^{-3}\,M$ … this or the pH is always the **equilibrium** concentration.

Don’t skip steps. Write out the ICE. You will have to explain yourself on the FRQ, so get used to doing this!

	HA(*aq*)	$\rightleftarrows$	H^+(*aq*)	+	A^-(*aq*)
Initial	0.10		0		0
Change					
Equilibrium			0.00130M		

Now you can easily fill out the entire chart.

	HA(*aq*)	$\rightleftarrows$	H^+(*aq*)	+	A^-(*aq*)
Initial	0.10 M		0		0
Change	- 0.00130 M		+ 0.00130 M		+ 0.00130 M
Equilibrium	0.10 M		0.00130 M		0.00130 M

$$K_a = \frac{[H^+][OH^-]}{[HA]} \text{ @ Equilibrium}$$

$$K_a = \frac{[0.00130][0.00130]}{[0.10]} = 1.7 \times 10^{-5}$$

$$pK_a = -\log(1.7 \times 10^{-5}) = 4.77$$

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2. Calculate the K_a and pK_a values of a 0.20 M weak acid solution with a pH of 5.03.

$$[H^+] = 10^{-5.03}\,M = 9.3 \times 10^{-6}\,M$$

	HA(*aq*)	$\rightleftarrows$	H^+(*aq*)	+	A^-(*aq*)
Initial	0.20		0		0
Change	- 0.0000093 M		+ 0.0000093 M		+ 0.0000093 M
Equilibrium	0.20 M		$9.3 \times 10^{-6}\,M$		$9.3 \times 10^{-6}\,M$

$$K_a = \frac{[H^+][OH^-]}{[HA]} \text{ @ Equilibrium}$$

$$K_a = \frac{[9.3 \times 10^{-6}][9.3 \times 10^{-6}]}{[0.20]} = 4.3 \times 10^{-10}$$

$$pK_a = -\log(4.3 \times 10^{-10}) = 9.37$$

Now for two problems that have you predict the pH of weak acid solutions based on the strength of the acid.

3. If the K_a of formic acid is 1.77×10^{-4}, what is the pH of a 1.0 *M* solution of formic acid?

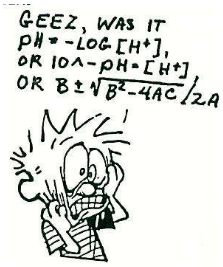
	HFormic(*aq*)	+	$H_2O(l)$	⇄	$H_3O^+(aq)$	+	$Formic^-(aq)$
Initial	1.0				0		0
Change	$-x$				$+x$		$+x$
Equilibrium	$1.0 - x$				x		x

$$K_a = \frac{[x][x]}{[1.0 - x]} \text{ @ Equilbrium}$$

The acid is weak and its original concentration is high, so this approximation will work.

$$K_a = \frac{[x][x]}{[1.0]}$$

$[1.0] - [x] \approx 1.0]$[1]



$1.77 \times 10^{-4} = \frac{[x]^2}{[1.0]}$	$0.013\ M = x$	$-\log[0.013] = 1.88$ pH

$H_3O^+(aq) = 0.013\ M$ (simplifying assumption is correct since $1.0 - x$, $1.0 - 0.013 = 1.0$)

Solving using Solver[2] or the quadratic equation has $x = 0.0132\ M$, with a pH of 1.88.

$$[H^+] = \sqrt{K_a \times [HA]}$$
$$[H^+] = \sqrt{1.77 \times 10^{-4} \times 1.0}$$
$$[H^+] = 0.013M$$

This equation could have been used to quickly solve the problem.

$[H^+] = \sqrt{K_a \times [HA]}$ can solve these problems quickly and is useful in the MC section.

However, in the FRQ section, points are awarded for showing your step-by-step method of solving the problem.

So be sure to do your ICE Chart even though this approximation equation can solve the question quickly.

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[1] You can solve the more complex equation using Solver on the Ti. However only do this if you are good with Solver and methodical enough not to make errors writing the complete equation.

[2] SOLVER eqn:0=X²/(1-X)-1.77E-4

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4. If the K_a of acetic acid is 1.8×10^{-5}, what is the pH of a 0.010 M solution of acetic acid?

	HAcetic(aq)	+	$H_2O(l)$	⇄	$H_3O^+(aq)$	+	Acetic$^-$(aq)
Initial	1.0				0		0
Change	$-x$				$+x$		$+x$
Equilibrium	$0.010 - x$				$0 + x$		$0 + x$

$K_a = \dfrac{[x][x]}{[0.010 - x]} = 1.8\times10^{-5}$	$[H^+] = \sqrt{K_a \times [HA]}$ $[H^+] = \sqrt{1.8 \times 10^{-5} \times 0.010}$	$x = 4.2 \times 10^{-4}\ M$ Quadratic-Solver answer[3]	$-\log(4.2 \times 10^{-4}] = $ **3.38 pH**

K_b and Weak Bases

Weak bases require a slightly more complex ICE Chart since water must be included to show the ionization producing hydroxide ions.

5. Calculate the K_b and pK_b of a 0.250 M solution of a weak base, A^-, with a pH of 9.68.

pOH=14 - 9.68 = 4.32 $[OH^-]=10^{-4.32}\ M = 4.8\times10^{-5} M$

	$A^-(aq)$	+	H_2O	⇄	HA(aq)	+	$OH^-(aq)$
Initial	0.250				0		0
Change							
Equilibrium							$4.8\times10^{-5} M$

With these values in the table, the rest of the equilibrium can be found

	$A^-(aq)$	+	H_2O	⇄	HA(aq)	+	$OH^-(aq)$
Initial	$0.250 M$				0		0
Change	$-4.8\times10^{-5} M$				$+4.8\times10^{-5} M$		$+4.8\times10^{-5} M$
Equilibrium	$0.250\ M - 4.8\times10^{-5}\ M$				$4.8\times10^{-5}\ M$		$4.8\times10^{-5}\ M$

$$K_b = \frac{[HA][OH^-]}{[A^-]} \text{ @ Equilibrium}$$

$$K_b = \frac{[4.8\times10^{-5}][4.8\times10^{-5}]}{[0.250 - 4.8\times10^{-5}]} = \frac{[4.8\times10^{-5}]^2}{[0.250]} = 9.2\times10^{-9}$$

$pK_b = -\log(9.2 \times 10^{-9}) = $ **8.04**

The pK_a for the conjugate acid in this problem, HA, will be 14.00-8.04 = 5.96
The K_a for the conjugate acid in this problem, HA, will be $1.0\times10^{-14} / 9.2 \times 10^{-9} = 1.1\times 10^{-6}$

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[3] Solving this problem using Solver would produce a concentration of $4.15 \times 10^{-4}\ M$. One advantage of Solver is that if you set x as your variable, you can use that same x variable when finding the pH, by -log(x). However, you must be good at using Solver since if you do not set your x in Solver to start at 0, you could get the negative root of this problem, -4.3×10^{-4}.