N37 – Acid Base

Weak Problems

N37 – Acid Base

Yay, ICE Tables are back!

A Weak Acid Equilibrium Problem

What is the pH of a 0.50 M solution of acetic acid, $HC_2H_3O_2$, $K_a = 1.8 \times 10^{-5}$?

Step #1: Write the dissociation equation

$$HC_2H_3O_2 \leftrightarrows C_2H_3O_2^- + H^+$$

A Weak Acid Equilibrium Problem

What is the pH of a 0.50 M solution of acetic acid, $HC_2H_3O_2$, $K_a = 1.8 \times 10^{-5}$?

Step #2: ICE it!

	$HC_2H_3O_2 \leftrightharpoons$	$C_2H_3O_2^-$	$+ \; H^+$
I	0.50	0	0
C	- x	+x	+x
E	0.50 - x	x	x

A Weak Acid Equilibrium Problem

What is the pH of a 0.50 M solution of acetic acid, $HC_2H_3O_2$, K_a = 1.8 x 10^{-5} ?

Step #3: Set up the law of mass action

	$HC_2H_3O_2$ ⇆	$C_2H_3O_2^-$	+ H^+
E	0.50 - x	x	x

$$1.8x10^{-5} = \frac{(x)(x)}{(0.50 - x)} \cong \frac{x^2}{(0.50)}$$

Can use the 5% rule because K < 1 and K at least 1000 time smaller than [initial]

A Weak Acid Equilibrium Problem

What is the pH of a 0.50 M solution of acetic acid, $HC_2H_3O_2$, $K_a = 1.8 \times 10^{-5}$?

Step #4: **Solve for x, which is also $[H^+]$**

$$HC_2H_3O_2 \leftrightarrows C_2H_3O_2^- + H^+$$

E	0.50 - x	x	x

$$1.8 x 10^{-5} = \frac{x^2}{(0.50)}$$

$[H^+] = 3.0 \times 10^{-3}$ M

A Weak Acid Equilibrium Problem

What is the pH of a 0.50 M solution of acetic acid, $HC_2H_3O_2$, $K_a = 1.8 \times 10^{-5}$?

Step #5: Convert $[H^+]$ to pH

	$HC_2H_3O_2 \leftrightarrows$	$C_2H_3O_2^-$	$+ \quad H^+$
E	$0.50 - x$	x	x

$$pH = -\log(3.0\,x10^{-3}) = 4.52$$

Reaction of Weak Bases with Water

The base reacts with water, producing its conjugate acid and hydroxide ion:

$$CH_3NH_2 + H_2O \leftrightarrows CH_3NH_3^+ + OH^- \quad K_b = 4.38 \times 10^{-4}$$

$$K_b = 4.38\,x10^{-4} = \frac{[CH_3NH_3{}^+][OH^-]}{[CH_3NH_2]}$$

K_b for Some Common Weak Bases

Many students struggle with identifying weak bases and their conjugate acids. What patterns do you see that may help you?

Base	Formula	Conjugate Acid	K_b
Ammonia	NH_3	NH_4^+	1.8×10^{-5}
Methylamine	CH_3NH_2	$CH_3NH_3^+$	4.38×10^{-4}
Ethylamine	$C_2H_5NH_2$	$C_2H_5NH_3^+$	5.6×10^{-4}
Diethylamine	$(C_2H_5)_2NH$	$(C_2H_5)_2NH_2^+$	1.3×10^{-3}
Triethylamine	$(C_2H_5)_3N$	$(C_2H_5)_3NH^+$	4.0×10^{-4}
Hydroxylamine	$HONH_2$	$HONH_3^+$	1.1×10^{-8}
Hydrazine	H_2NNH_2	$H_2NNH_3^+$	3.0×10^{-6}
Aniline	$C_6H_5NH_2$	$C_6H_5NH_3^+$	3.8×10^{-10}
Pyridine	C_5H_5N	$C_5H_5NH^+$	1.7×10^{-9}

Reaction of Weak Bases with Water

The generic reaction for a base reacting with water, producing its conjugate acid and hydroxide ion:

$$B + H_2O \leftrightarrows BH^+ + OH^-$$

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

(Yes, all weak bases do this – DO NOT make this more complicated then it needs to be.)

A Weak Base Equilibrium Problem

What is the pH of a 0.50 M solution of ammonia, NH_3, $K_b = 1.8 \times 10^{-5}$?

Step #1: **Write the equation for the reaction**

$$NH_3 + H_2O \leftrightarrows NH_4^+ + OH^-$$

A Weak Base Equilibrium Problem

What is the pH of a 0.50 M solution of ammonia, NH_3, $K_b = 1.8 \times 10^{-5}$?

Step #2: ICE it!

	NH_3	+	H_2O	$\leftrightarrows$	NH_4^+	+	OH^-
I	0.50				0		0
C	- x				+x		+x
E	0.50 - x				x		x

A Weak Base Equilibrium Problem

What is the pH of a 0.50 M solution of ammonia, NH_3, $K_b = 1.8 \times 10^{-5}$?

Step #3: **Set up the law of mass action**

	NH_3	+	H_2O	$\leftrightarrows$	NH_4^+	+	OH^-
E	0.50 - x				x		x

$$1.8 x 10^{-5} = \frac{(x)(x)}{(0.50 - x)} \cong \frac{x^2}{(0.50)}$$

Can use the 5% rule because K < 1 and K at least 1000 time smaller than [initial]

A Weak Base Equilibrium Problem

What is the pH of a 0.50 M solution of ammonia, NH_3, $K_b = 1.8 \times 10^{-5}$?

Step #4: **Solve for x, which is also $[OH^-]$**

	NH_3	+	H_2O	$\leftrightarrows$	NH_4^+	+	OH^-
E	0.50 - x				x		x

$$1.8x10^{-5} = \frac{x^2}{(0.50)}$$

$[OH^-] = 3.0 \times 10^{-3}$ M

A Weak Base Equilibrium Problem

What is the pH of a 0.50 M solution of ammonia, NH_3, $K_b = 1.8 \times 10^{-5}$?

Step #5: **Convert $[OH^-]$ to pH**

$$NH_3 + H_2O \leftrightarrows NH_4^+ + OH^-$$

E $\quad 0.50 - x \qquad x \qquad x$

$$pOH = -\log(3.0\,x10^{-3}) = 4.52$$

$$pH = 14.00 - pOH = 9.48$$

Remember…

You can convert back and forth from Ka to Kb and vice versa. If you are given Ka for an acid but are doing problems with the acid's conjugate base you can use that Ka to find the Kb that you need.

$$Ka \times Kb = Kw$$

$$Ka \times Kb = (1 \times 10^{-14})$$