

N-48

# Weak Acid/Base Equilibria

## Target

I can use equilibrium expressions and ICE tables to calculate the concentrations of reactants and products at equilibrium for acids and bases, and use that information to find the pH or other related values.



# Weak Acids and Bases

## What do chemists mean by WEAK?

They do not completely ionize in water.

Only a LITTLE BIT will be dissociated.



# Connection back to...Equilibrium!

Dissociation is a reversible reaction right?

So...

We can use equilibrium constants, expressions, ice tables to determine [ ]'s which let us find...

*pH values!*

## K<sub>a</sub> and K<sub>b</sub>

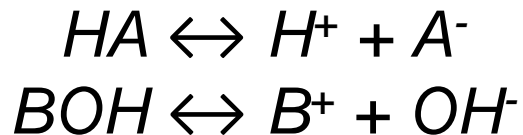
Remember that K<sub>eq</sub> is just generic.

Could be K<sub>c</sub>, K<sub>p</sub>, K<sub>sp</sub> if you are trying to be specific. So for acid bases use:

- $K_a$  (for acids)
- $K_b$  (for bases)

Still  $\frac{\text{Products}}{\text{Reactants}}$  which will be

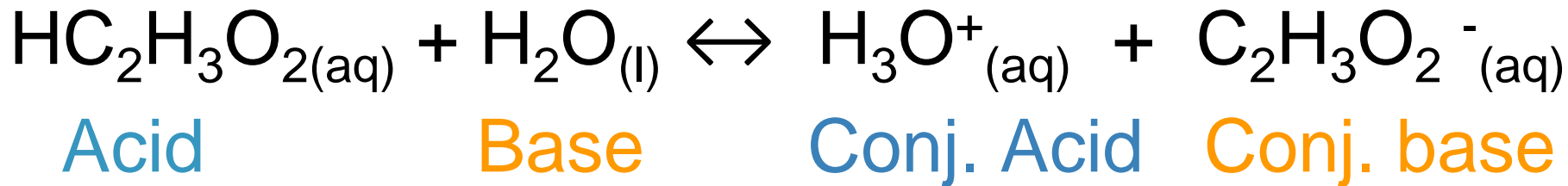
$$\frac{[\text{Dissociated Ions}]}{[\text{Undissociated Molecule}]}$$





## Practice Problem

**Identify Acid/Conj Base/Base/Conj Acid for Acetic Acid  
 $\text{HC}_2\text{H}_3\text{O}_2$  (HOAc). Then write the  $K_a$  Expression**



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{OAc}^-]}{[\text{HOAc}]}$$

# Size of Ka for Weak Acids

$$K_a = \frac{[H_3O^+][OAc^-]}{[HOAc]} = 1.8 \times 10^{-5}$$

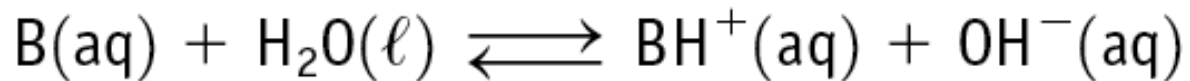
## **Why is the Ka so small for acetic acid???**

- Small Ka means equilibrium lies to the LEFT
- Reactant Favored – not much dissociated
- It is a WEAK acid!
- Weak acids have  $K_a < 1$ 
  - *Leads to low  $[H^+] \rightarrow pH$  from 2 - 6.9*



# Size of $K_b$ for Weak Bases

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



## Weak Bases have small $K_b$ values

- Small  $K_b$  means equilibrium lies to the LEFT
- Reactant Favored – not much dissociated
- It is a WEAK base!
- Weak bases have  $K_b < 1$ 
  - *Leads to low  $[\text{OH}^-] \rightarrow \text{pH from } 12 - 7.1$*



# Relationship between $K_a$ , $[H^+]$ , pH

Increase  
in ACID  
strength

$K_a$  and  $[H_3O^+]$   
increase



pH  
decreases



Increase  
in BASE  
strength

$K_b$  and pH  
increase



$[H_3O^+]$   
decreases



Glue Into Your Notes!



# Practice with Weak Acid/Bases

You have 1.00 M HOAc. Calculate the equilibrium concentrations of HOAc,  $\text{H}_3\text{O}^+$ ,  $\text{OAc}^-$ , and the pH if the  $K_a = 1.8 \times 10^{-5}$ .



**Step 1:** Create an ICE Table

**Step 2:** Write  $K_a$  expression (or  $K_b$  expression depending on Q)

**Step 3:** Solve for x using quadratic or 5% rule if valid.

**Step 4:** Solve for pH (or pOH depending on the Q)

# Practice with Weak Acid/Bases

You have 1.00 M HOAc. Calculate the equilibrium concentrations of HOAc,  $H_3O^+$ ,  $OAc^-$ , and the pH if the  $K_a = 1.8 \times 10^{-5}$ .  $HOAc + H_2O \leftrightarrow H_3O^+ + OAc^-$

Reaction	[HOAc]	[ $H_3O^+$ ]	[ $OAc^-$ ]
I	1.00	0.00	0.00
C	-x	+x	+x
E	1.00-x	x	x
5%	1.00	x	x
Answer			

# Practice with Weak Acid/Bases

You have 1.00 M HOAc. Calculate the equilibrium concentrations of HOAc,  $H_3O^+$ ,  $OAc^-$ , and the pH if the  $K_a = 1.8 \times 10^{-5}$ .  $HOAc + H_2O \leftrightarrow H_3O^+ + OAc^-$

$$K_a = \frac{[H_3O^+][OAc^-]}{[HOAc]} \qquad 1.8 \times 10^{-5} = \frac{x^2}{1.00}$$

$$x = 4.2 \times 10^{-3}$$

# Practice with Weak Acid/Bases

You have 1.00 M HOAc. Calc. the equilibrium concentrations of HOAc,  $H_3O^+$ ,  $OAc^-$ , and the pH if  $K_a = 1.8 \times 10^{-5}$ .  $HOAc + H_2O \leftrightarrow H_3O + OAc^-$

Reaction	[HOAc]	[ $H_3O^+$ ]	[ $OAc^-$ ]
I	1.00	0.00	0.00
C	-x	+x	+x
E	1.00-x	x	x
5%	1.00	x	x
Answer	1.00	$4.2 \times 10^{-3}$	$4.2 \times 10^{-3}$

# Practice with Weak Acid/Bases

You have 1.00 M HOAc. Calc. the equilibrium concentrations of HOAc,  $H_3O^+$ ,  $OAc^-$ , and the pH if  $K_a = 1.8 \times 10^{-5}$ .  $HOAc + H_2O \leftrightarrow H_3O^+ + OAc^-$

**Check that the 5% Rule is Valid!**

$$x = \frac{x}{[initial]} \times 100 \leq 5\%$$

$$x = \frac{4.2 \times 10^{-3}}{1.00} \times 100 = 0.42\% \quad \checkmark$$

# Practice with Weak Acid/Bases

*You have 1.00 M HOAc. Calc. the equilibrium concentrations of HOAc,  $H_3O^+$ ,  $OAc^-$ , and the pH if  $K_a = 1.8 \times 10^{-5}$ .  $HOAc + H_2O \leftrightarrow H_3O + OAc^-$*

**Now Solve for pH – Don't forget!!!**

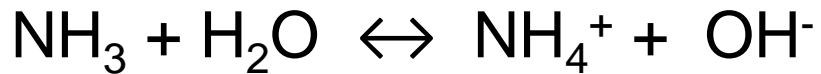
$$pH = -\log [H_3O^+]$$

$$pH = -\log (4.2 \times 10^{-3})$$

$$pH = 2.37$$

# Practice with Weak Acid/Bases

You have 0.010 M  $\text{NH}_3$ . Calculate the pH.  $K_b = 1.8 \times 10^{-5}$



Reaction	$[\text{NH}_3]$	$[\text{NH}_4]$	$[\text{OH}^-]$
I	0.010	0.00	0.00
C	-x	+x	+x
E	0.010-x	x	x
5%	0.010	x	x
Answer			

# Practice with Weak Acid/Bases

You have 0.010 M  $\text{NH}_3$ . Calculate the pH.  $K_b = 1.8 \times 10^{-5}$



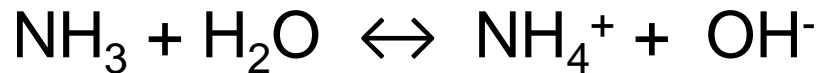
$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} \qquad 1.8 \times 10^{-5} = \frac{[x][x]}{0.010}$$

$$x = 4.2 \times 10^{-4}$$



# Practice with Weak Acid/Bases

You have 0.010 M  $\text{NH}_3$ . Calculate the pH.  $K_b = 1.8 \times 10^{-5}$



Reaction	$[\text{NH}_3]$	$[\text{NH}_4^+]$	$[\text{OH}^-]$
I	0.010	0.00	0.00
C	-x	+x	+x
E	0.010-x	x	x
5%	1.00	x	x
Answer	1.00	$4.2 \times 10^{-4}$	$4.2 \times 10^{-4}$

# Practice with Weak Acid/Bases

You have 0.010 M  $\text{NH}_3$ . Calculate the pH.  $K_b = 1.8 \times 10^{-5}$



**Now Calculate pH!**

$$[\text{OH}^-] = 4.2 \times 10^{-4} \text{ M}$$

$$\text{pOH} = -\log [\text{OH}^-] = 3.37$$

$$\text{pH} + \text{pOH} = 14 \rightarrow 14 = \text{pH} + 3.37$$

$$\text{pH} = 10.63$$