



N-48

Weak Acid/Base Equilibria



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Target

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

Link to YouTube Presentation: <https://youtu.be/fkc4USA25l8>



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_a and K_b

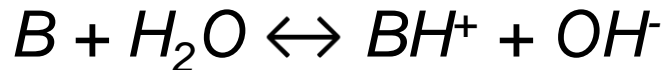
Remember that K_{eq} is just generic.

Could be K_c, K_p, K_{sp} if you are trying to be specific. So for acid bases use:

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

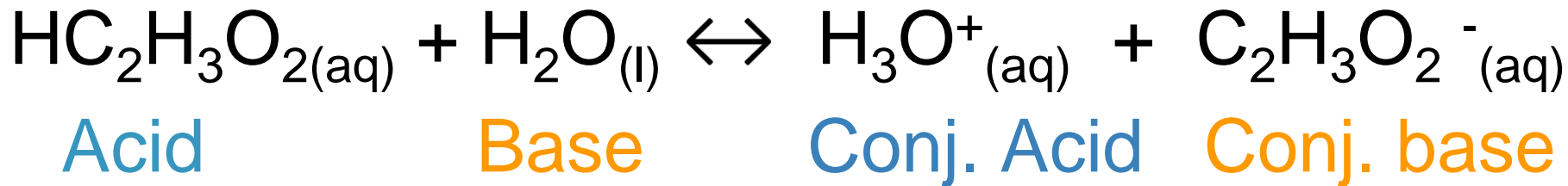
Still $\frac{\text{Products}}{\text{Reactants}}$ which will be

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



Practice Problem

Identify Acid/Conj Base/Base/Conj Acid for $\text{HC}_2\text{H}_3\text{O}_2$ (abbreviated as 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

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

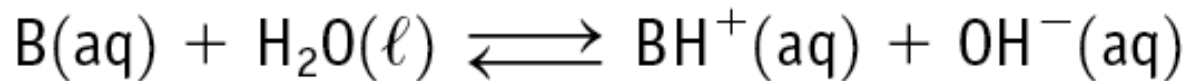
Why is the Ka so small for acetic acid???

- 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 Kb for Weak Bases

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



Weak Bases have small Kb values

- Small Kb 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



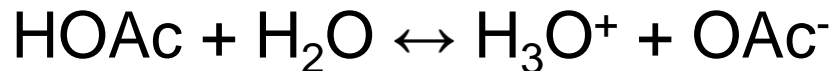
$[H_3O^+]$
decreases



Glue Into Your Notes!

Practice with Weak Acid/Bases

You have 1.00 M HOAc. Calculate the equilibrium concentrations of HOAc, H_3O^+ , 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	4.2×10^{-3}	4.2×10^{-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!!!

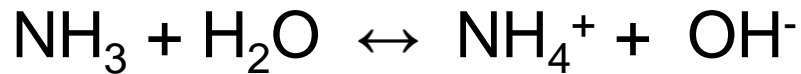
$$pH = -\log [H_3O^+]$$

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

$$pH = 2.37$$

Practice with Weak Acid/Bases

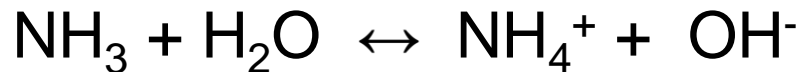
You have 0.010 M NH_3 . Calculate the pH. $K_b = 1.8 \times 10^{-5}$



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 NH_3 . Calculate the pH. $K_b = 1.8 \times 10^{-5}$

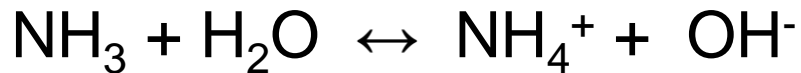


$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} \qquad 1.8 \times 10^{-5} = \frac{[x][x]}{0.010}$$

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

Practice with Weak Acid/Bases

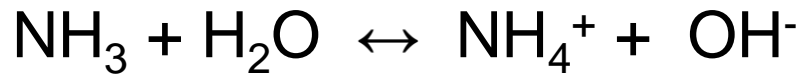
You have 0.010 M NH_3 . Calculate the pH. $K_b = 1.8 \times 10^{-5}$



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	0.010	4.2×10^{-4}	4.2×10^{-4}

Practice with Weak Acid/Bases

You have 0.010 M NH_3 . Calculate the pH. $K_b = 1.8 \times 10^{-5}$



Now Calculate pH!

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

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

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

$$\text{pH} = 10.63$$



Weird Fact...

The % dissociation INCREASES as the concentration of the weak acid or weak base DECREASES!



Weird Fact...



$$\% \text{ dissociation} = \frac{[\textit{ions}]}{[\textit{undissociated}]} \times 100$$

WHICH IS LIKE....

$$K_a = \frac{[H^+][A^-]}{HA}$$



Weird Fact...

Let's pretend we take our bottle of acid and double the amount of water.

The concentration of everything is cut in half correct????



Weird Fact...

So our K_a turns into...

$$Q = \frac{\left[\frac{1}{2} H^+\right] \left[\frac{1}{2} A^-\right]}{\left[\frac{1}{2} HA\right]}$$

vs. $\frac{[H^+][A^-]}{[HA]}$ like it had been.

Weird Fact...

So...

$$Q = \frac{\frac{1}{2} [H^+][A^-]}{[HA]}$$

So $Q = \frac{1}{2} K_a$

Weird Fact...

$$Q < K$$

MEANS...

Not enough products!

SHIFT RIGHT to get to equilibrium!

Make more products!



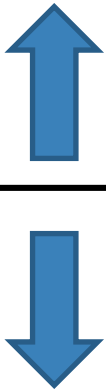
Weird Fact...

**And the products are...
[H⁺] and [A⁻] !**

**So MORE ions means...
HIGHER % dissociation!**



Weird Fact...

$$\% \text{ dissociation} = \frac{\textit{ions}}{\textit{undissociated}} \times 100$$


A bigger number!

Larger % dissociation!

YouTube Link to Presentation

<https://youtu.be/fkc4USA25l8>