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!

Ka and Kb

Remember that Keq is just generic. Could be Kc, Kp, Ksp if you are trying to be specific. So for acid bases use:

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

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

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

$HA \leftrightarrow H^+ + A^-$

$BOH \leftrightarrow B^+ + OH^-$

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

Practice Problem

Identify Acid/Conj Base/Base/Conj Acid for $HC_2H_3O_2$ (abbreviated as HOAc). Then write the Ka Expression

$$HC_2H_3O_{2(aq)} + H_2O_{(l)} \leftrightarrow H_3O^+_{(aq)} + C_2H_3O_2{}^-_{(aq)}$$

Acid | Base | Conj. Acid | Conj. base

$$Ka = \frac{[H_3O^+][OAc^-]}{[HOAc]}$$

Size of Ka for Weak Acids

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

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 Ka < 1
  - *Leads to low [H^+] → pH from 2 - 6.9*

Size of Kb for Weak Bases

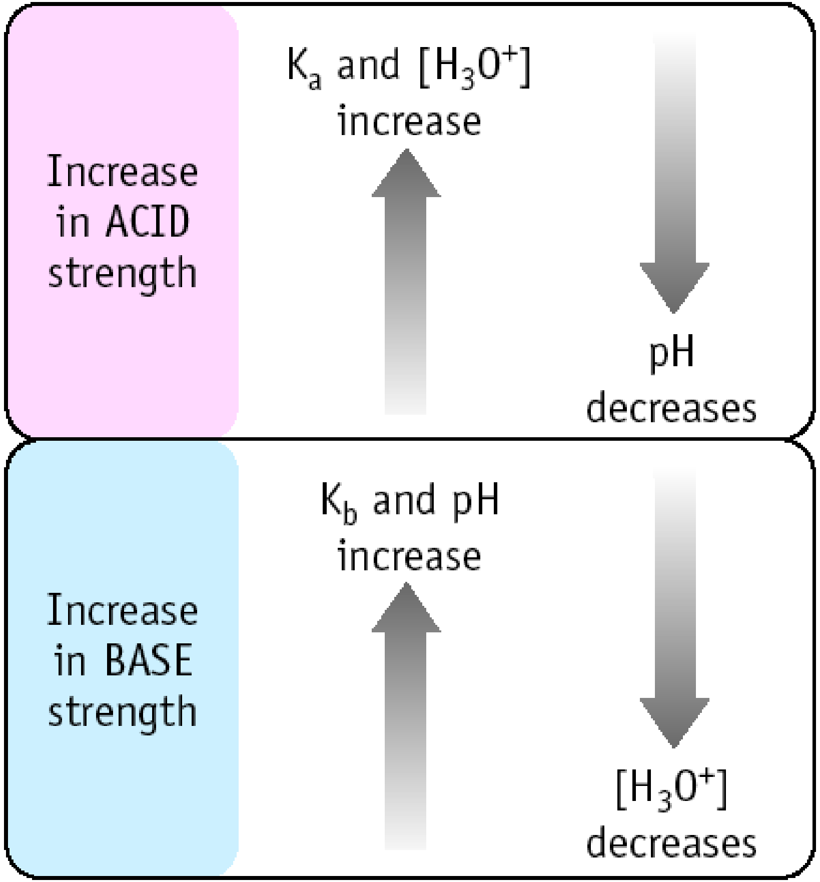
$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

$$B(aq) + H_2O(\ell) \rightleftharpoons BH^+(aq) + OH^-(aq)$$

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 Kb < 1
  - *Leads to low [OH^-] → pH from 12 – 7.1*

Relationship between Ka, [H+], pH



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 Ka = 1.8×10^{-5}.

$$HOAc + H_2O \leftrightarrow H_3O^+ + OAc^-$$

Step 1: Create an ICE Table

Step 2: Write Ka expression *(or Kb 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 Ka = 1.8 x 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 Ka = 1.8 x 10^{-5}. HOAc + $H_2O \leftrightarrow H_3O^+ + OAc^-$

$$Ka = \frac{[H_3O^+][OAc^-]}{[HOAc]}$$

$$1.8\ x\ 10^{-5} = \frac{x^2}{1.00}$$

x = 4.2 x 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 Ka = $1.8x10^{-5}$. 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 Ka = $1.8x10^{-5}$. $HOAc + H_2O \leftrightarrow H_3O + OAc^-$

Check that the 5% Rule is Valid!

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

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

Practice with Weak Acid/Bases

You have 1.00 M HOAc. Calc. the equilibrium concentrations of HOAc, H_3O^+, OAc^-, and the pH if $Ka = 1.8x10^{-5}$. $HOAc + H_2O \leftrightarrow H_3O + OAc^-$

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. $K_b = 1.8 \times 10^{-5}$

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

Reaction	$[NH_3]$	$[NH_4^+]$	$[OH^-]$
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. $K_b = 1.8 \times 10^{-5}$

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

$$\boldsymbol{Kb = \frac{[NH_4^+]\,[OH^-]}{[NH_3]}}$$

$$1.8\ x\ 10^{-5} = \frac{[x][x]}{0.010}$$

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

Practice with Weak Acid/Bases

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

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

Reaction	$[NH_3]$	$[NH_4^+]$	$[OH^-]$
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. $K_b = 1.8 \times 10^{-5}$

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

Now Calculate pH!

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

$$pOH = -\log [OH^-] = 3.37$$

$$pH + pOH = 14 \rightarrow 14 = pH + 3.37$$

$$\mathbf{pH = 10.63}$$

Weird Fact...

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

Weird Fact...

$$HA \rightarrow H^+ + A^-$$

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

WHICH IS LIKE....

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

Weird Fact…

$$HA \rightarrow H^{+} + A^{-}$$

$$\% \text{ dissociation} = \frac{[H^{+}]}{[undissociated]} x\ 100$$

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...

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

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

Weird Fact...

So...

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

So Q = ½ Ka

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...

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

A bigger number!

Larger % dissociation!

YouTube Link to Presentation

https://youtu.be/fkc4USA25l8