Strong and Weak Acids and Bases and pH

Get your thinking caps strapped on!

pH Scale

- pH Scale: a relative scale showing the strength of acids or bases, based on the concentration of the H^+ ions in solution
- pH = -log[H+]
- For Water:
  - pH = -log(1.011 x 10^{-7})
  - pH = 7
- High [H^+] = low pH = acidic
- Low [H^+] = high pH = basic
- The further away from 7, the stronger the A/B

For Strong Acids

- Strong = complete dissociation
- This means that $[acid]=[H^+]$
- For bases, $[base] = [OH^-]$
- Whatever the concentration of the A/B is, it can be used in the pH equation

Examples

What is the pH of the following strong solutions?

1. .10M HCl
2. .10M H_2SO_4
3. .10M H_3PO_4
4. .10M NaOH
5. .10M $Ca(OH)_2$

Example Answers

1. .10M HCl
2. .10M H_2SO_4
3. .10M H_3PO_4
4. .10M NaOH
5. .10M $Ca(OH)_2$

Answers

1. pH = -log(.10) = 1.0
2. pH = -log(2*.10) = .7
3. pH = -log(3*.10) = .5
4. pOH = -log(.10) = 1
   - pH + 1 = 14 pH = 13
5. pOH = -log(2*.10) = .7
   - pH + .7 = 14 pH = 13.3

What is a Weak Acid?

- Acids can be classified into two types:
  - Strong Acids: completely dissociate
  - Weak Acids: partially dissociate, which means not all the molecules separate into their ions

Why are some weak?

- It depends on the structure of the acid
- Acids that have strong ionic character (big electronegativity differences) will be strong
- Acids that have weaker ionic character will be weak
  - These are most often the acids with polyatomic ions

Examples

<u>Strong Acids</u>

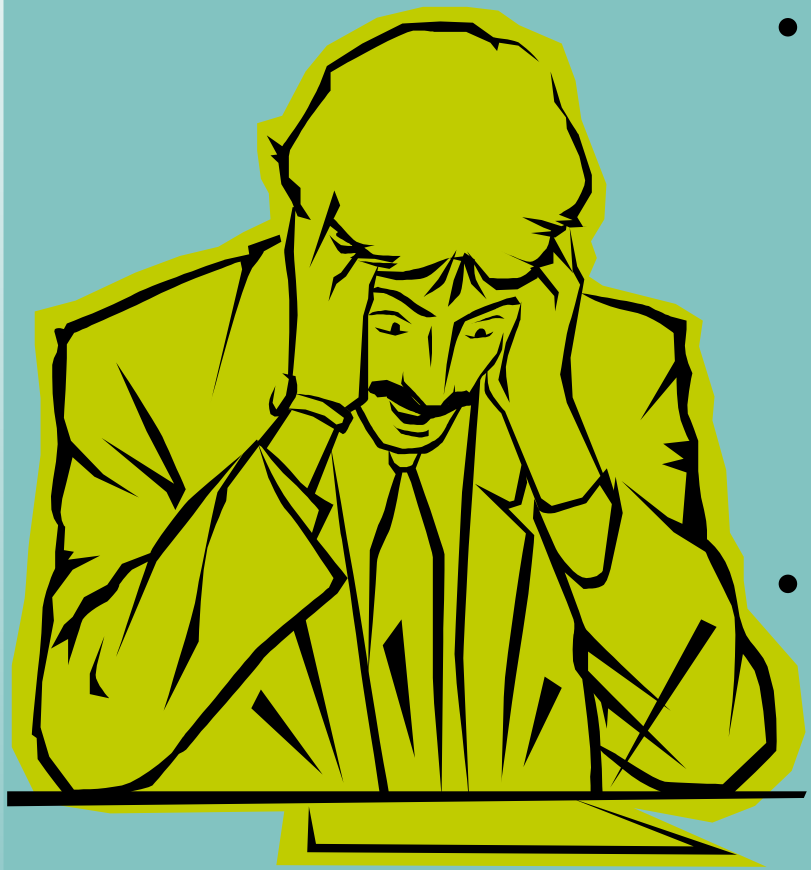
- HCl
- HF
- H_2SO_4

<u>Weak Acids</u>

- $H(C_2H_3O_2)$
- H_3PO_4

You do not need to know whether an acid is strong or weak…I will provide that information for you!!

So what is different about Weak Acids?



- Because weak acids do not completely dissociate, we cannot use the concentration directly to figure out pH
- We need to know how much of the acid splits apart vs. how much stays together

Dissocation Constants

- Dissociation Constant: a constant that tells how much of a chemical dissociates in water (like a percentage)
- Abbreviated "K" (notice it is capital K, not lowercase k…)
- In general, the constant can be calculated using this equation:

$$K = \frac{[product1] * [product\ 2]* \ldots}{[reactant1] * [reactant2]* \ldots}$$

Remember that brackets means "concentration of" something

Examples of K values

- For H_2SO_4,

K= 1.20×10^{-2}

- For $H(C_2H_3O_2)$,

K= 1.75×10^{-5}

- The smaller the K value, less of the chemical dissociates

- $K < 1x10^{-3}$ to be considered "weak"

Weak Acids and pH

- We cannot find the pH of a weak acid by using the concentration of the acid directly
- For example:

$$.10M\ H(C_2H_3O_2)$$

$$pH \neq -\log(.10)$$

- We need to figure out how much splits vs. stays together

Are we having fun yet?

How do we figure that?

- Let's look at acetic acid, $H(C_2H_3O_2)$, for an example:
- The equation for the dissociation of that acid looks like this:

$$H(C_2H_3O_2) \rightarrow H^+ + (C_2H_3O_2)^-$$

- Let's look at the solution before and after the dissolving…

How much of each chemical do I have during the dissolving?

	$H(C_2H_3O_2) \rightarrow$	H^+	$+ (C_2H_3O_2)^-$
Amount Before Dissociation	M	0	0
As it dissociates	Losing “x”	Gaining “x”	Gaining “x”
Amount After Dissociation	M-x	x	x

Using K

In general (slide #10)

$$K = \frac{[product1] * [product\ 2]*\ldots}{[reactant1] * [reactant2]*\ldots}$$

For acetic acid (slide #13)

$$K = \frac{[H^+] * [C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

Plugging in values (from last slide)

$$K = \frac{[x] * [x]}{[M-x]}$$

So what does that mean?

- In general, the equation for the dissociation of a weak acid is:

$$K = \frac{[x] * [x]}{[M\text{-}x]}$$

, which is often written like this:

$$K = \frac{[x]^2}{[M\text{-}x]}$$

K = dissociation constant

X = $[H^+]$

M= [acid]

•Know this equation!

This is how we find $[H^+]$ of

a weak acid!

How do we do that?

- If we know K and the concentration of the acid, we can solve for x, which is [H+]
- Once we know [H+], we can then use our pH equation like usual
- Should we do an example?

What is the pH of .10M $H(C_2H_3O_2)$?

- What do we know?
  - K = 1.75×10^{-5} (this is a constant I will provide to you)
  - M = .10 M
- What are we looking for/What do we need?
  - We are looking for pH
  - We need [H+]

How do we find [H+]?

$$K = \frac{[x]^2}{[M-x]}$$

Move to other side

$$1.75 \times 10^{-5} = \frac{[x]^2}{[.10-x]}$$

distribute

$$(.10-x)(1.75 \times 10^{-5}) = (x)^2$$

Move to other side

$$(1.75 \times 10^{-6})-(1.75 \times 10^{-5})(x) = (x)^2$$

$$0 = (x)^2 + 1.75 \times 10^{-5}(x) - (1.75 \times 10^{-6})$$

How do we solve this?

Quadratic Equation!! Yippee!!

$$0 = (x)^2 + 1.75 \times 10^{-5}(x) - (1.75 \times 10^{-6})$$

$X = \dfrac{-b \pm \sqrt{(b^2-4ac)}}{2a}$, where a = number in front of x^2

b = number in front of x

c = last number

The quadratic can be programmed into your calculator if you don't have it already! If your calculator is non-programmable come see me!

How do we solve this?

$$0 = (x)^2 + 1.75 \times 10^{-5}(x) - (1.75 \times 10^{-6})$$

$$X = \frac{-b \pm \sqrt{(b^2 - 4ac)}}{2a}$$

$a = 1$

$b = 1.75 \times 10^{-5}$

$c = -1.75 \times 10^{-6}$

Two Answers:

- .00131 M
- ~~-.00133 M~~

Since we are dealing with concentration, we will always use the positive answer!

Now to find pH

- Now we know [H+] = .00131 M
- Finding pH is just like strong acids
  - pH = -log[H+]
  - pH = -log (.00131) = 2.9
- Notice how much smaller the pH is compared to a strong acid of the same concentration
  - .10 M HCl: pH = - log(.10) = 1.0

What about Weak Bases?

- Weak bases are treated the exact same way
- The difference is that the “x” in the equation is the $[OH^-]$
- So when you solve the “K” equation and perform the “-log”, you get pOH
- Simply subtract your answer from 14 to find pH!