

# **Experiment 9**

## **Acid-Base Equilibria**

**CH204**

**Spring 2006**

**Dr. Brian Anderson**

# Last Week

## Heat in chemical reactions:

heat is a measurable quantity

produced and consumed in stoichiometric amounts

## Heat Capacity:

how much heat is required to raise the temperature of something by one degree Celsius (or 1 Kelvin)

## Specific Heat Capacities (J/gK):

Lead 0.128

Iron 0.449

Water 4.184

# **This week**

**Weak acid titration.**

**Determine  $K_a$  of acetic acid by a couple different methods.**

**Witness the awesome power of a buffer solution to resist changes in pH.**



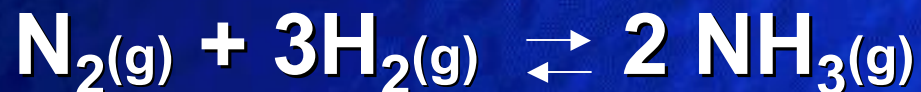
# Non-Equilibrium Reaction

Reaction goes to completion.



# Equilibrium Reaction

Products react with each other to re-form the reactants.



# Equilibrium Eckspresshin

For any equilibrium reaction,



products  
reactants

$$K_{\text{eq}} = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

For a weak acid dissociation,



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



# Two Ecksepshins

Don't include *liquids* or *solids* in equilibrium expressions.

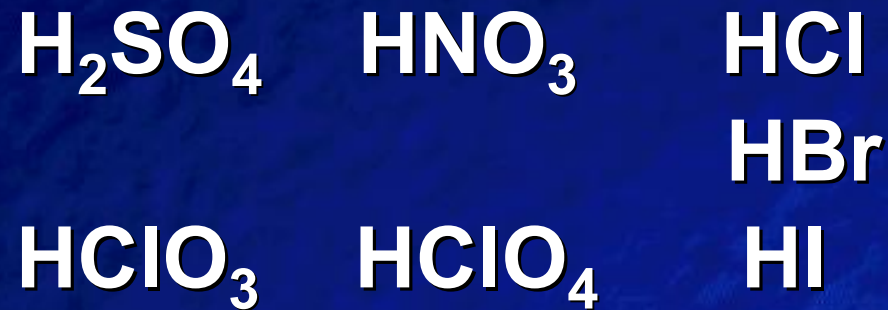


$$K_{\text{sp}} = [\text{Pb}^{2+}][\text{SO}_4^{2-}]$$



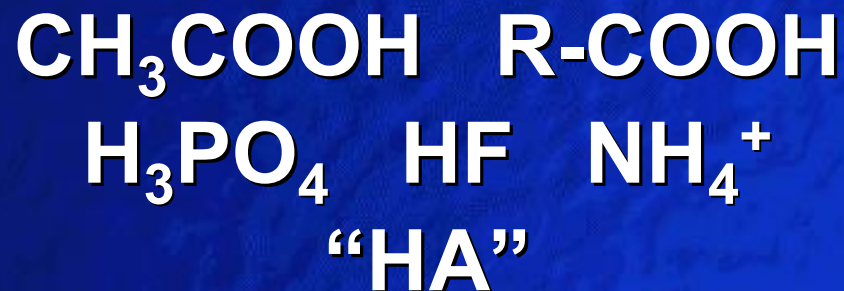
$$K_{\text{w}} = [\text{H}^+][\text{OH}^-]$$

# Strong Acids



# Weak Acids

*All the rest!*





# pH of weak acids

What is the pH of a 0.1M solution of acetic acid?



	$\text{CH}_3\text{COOH}$	$\text{CH}_3\text{COO}^-$	$\text{H}^+$
initial	0.1	0	0
equilibrium	$0.1 - x$	$x$	$x$

# Dissociation of acetic acid

	$\text{CH}_3\text{COOH}$	$\text{CH}_3\text{COO}^-$	$\text{H}^+$
initial	0.1	0	0
equilibrium	$0.1 - x$	$x$	$x$

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} = \frac{[x][x]}{[0.1 - x]}$$

Assume  $x \ll 0.1 \text{ M}$

$$x^2 = 0.1K_a$$

$$x^2/0.1 = K_a$$

# Four-Part Lab

## 1) Calibrate pH meter

Make sure you're in CALIBRATION mode.

Calibrate the pH meter in the order in the lab manual: **pH 7** first, then **pH 4**, then **pH 10**.

Press ENTER or CON to confirm calibration.



# Last Two Titrations of Your Life

## 2) Titrate 25 ml of 0.1 M acetic acid using 0.1 M NaOH

DO NOT add water! No indicator this time. Use a beaker, not a flask, because you need room for the pH electrode.

Record pH after the addition of every 1.0 ml of NaOH at first, and as the pH begins to change more quickly, record smaller volume increments, down to 0.2 or 0.1 ml. Try to catch points on the vertical portion of the graph.

Switch roles with your lab partner and repeat the titration a second time.

Graph pH versus ml added in Excel.

# At the Equivalence Point

All of the HA has been reacted away.

If the solution was initially 0.1M acetic acid,  
it is now 0.05 M acetate

At the half-equivalence point, half of the HA  
has been reacted away, and the HA and A<sup>-</sup>  
concentrations are equal.

# Half-Equivalence Point

At the half-equivalence point,  $[HA] = [A^-]$ .

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

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

$$pK_a = pH - \log \frac{[A^-]}{[HA]}$$

$\log(1) = 0$ , so when  $[A^-] = [HA]$ ,  $pH = pK_a$



# A Short-Cut to $K_a$

- 3) Measure the pH of 1.5 M acetic acid and two buffer solutions

Use measured  $[H^+]$  and known acetic acid and acetate concentrations to calculate  $K_a$

# Deja Vu

	$\text{CH}_3\text{COOH}$	$\text{CH}_3\text{COO}^-$	$\text{H}^+$
initial	1.5	0	0
equilibrium	$1.5 - x$	$x$	$x$

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} = \frac{[x][x]}{[1.5 - x]}$$

Assume  $x \ll 1.5 \text{ M}$

$$x^2 = 1.5K_a$$

$$x^2/1.5 = K_a$$

# And Finally...

- 4) Add strong acid & base to buffers and to water and compare the changes in pH.



# pH meters need love, too

**Glass bulb is very thin**

**Remove carefully from storage bottle – turn the bottle, not the cap**

**Rinse well between samples**

**Dab, don't wipe**

**Swish samples to get better reading**

