

## Experiment 9 Acid-Base Equilibria

CH 204  
Fall 2007  
Dr. Brian Anderson

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### 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

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### 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.

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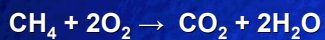
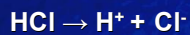
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## Non-Equilibrium Reaction

Reaction goes to completion.



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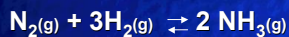
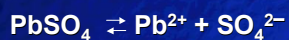
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## Equilibrium Reaction

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



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## Equilibrium Expression

For any equilibrium reaction,



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

For a weak acid dissociation,



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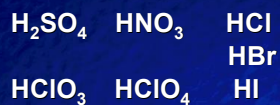
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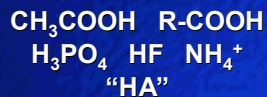
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### Strong Acids



### Weak Acids

*All the rest!*



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### Review of pH

Neutral                     $pH = 7$   
Acidic                      $pH < 7$   
Basic (alkaline)         $pH > 7$

pH is a negative log scale for acid concentrations. Increasing pH by one decreases  $[H^+]$  by a factor of 10.

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### $pH = \log_{10}[H^+]$

$[H^+]$	$\log[H^+]$	pH
$10^{-1}$	- 1	1
$10^{-2}$	- 2	2
$10^{-3}$	- 3	3
$10^{-4}$	- 4	4
$1.78 \times 10^{-5}$	- 4.75	4.75
$4.68 \times 10^{-7}$	- 6.33	6.33
$8.13 \times 10^{-12}$	- 11.90	11.90

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### Calculating pH

Given the  $[H^+]$ ,  $pH = -\log[H^+]$

What is the pH of a 0.025 M HCl solution?

$$pH = -\log[.025] = -(-1.6) = 1.6$$

Given the pH,  $[H^+] = 10^{-pH}$

What is the  $[H^+]$  concentration in human arterial blood, which has a pH of 7.40?

$$[H^+] = 10^{-pH} = 10^{-7.40} = 3.98 \times 10^{-8} \text{ M}$$

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### Review

If you are given the  $[H^+]$   
 $pH = -\log[H^+]$

If you are given the pH  
 $[H^+] = 10^{-pH}$

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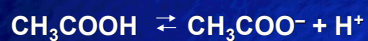
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### pH of weak acids

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



	$CH_3COOH$	$CH_3COO^-$	$H^+$
initial	0.1	0	0
equilibrium	$0.1 - x$	$x$	$x$

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## Dissociation of acetic acid

	CH <sub>3</sub> COOH	CH <sub>3</sub> COO <sup>-</sup>	H <sup>+</sup>
initial	0.1	0	0
equilibrium	0.1 - x	x	x

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

Assume  $x \ll 0.1$  M

$$K_a = x^2 / 0.1$$

$$K_a = (10^{-pH})^2 / 0.1$$

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## 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.

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## 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. Titrate in 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 (y-axis) versus ml added (x-axis) in Excel.

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## 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  
 $A^-$  concentrations are equal.

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## Half-Equivalence Point

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

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

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

$$-\log K_a = -\log [H^+]$$

so when  $[A^-] = [HA]$ ,  $pH = pK_a$

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## 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  $K_a$

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

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## Deja Vu

	CH <sub>3</sub> COOH	CH <sub>3</sub> COO <sup>-</sup>	H <sup>+</sup>
initial	1.4	0	0
equilibrium	1.4 - x	x	x

$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{[x][x]}{[1.4 - x]}$$

Assume  $x \ll 1.4$  M

$$K_a = x^2/1.4$$

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## And Finally...

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

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## 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



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## AWOL

I will be out of town Thursday and Friday of this week.

Post-Lab problem 4 is a little tricky.

There is a hint sheet for Post-Lab 9 on the [Freebies](#) page of the course web site.

The Cheat Sheet for Experiment 9 is also online.

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## Next week

Final Class Meeting

Course/Instructor Surveys

TA Evaluations

Kinetics lab

Lab check-out

If you have missed more than one lab, e-mail me to discuss make-up week.

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## Next week's quiz

LAST QUIZ!

Given  $[H^+]$  calculate pH

Given pH, calculate  $[H^+]$

Know how to recognize a buffer solution

Know how to make up a buffer solution

Given three variables in an equilibrium expression, calculate the fourth.

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