

Experiment 9 Acid-Base Equilibria

CH204
Spring 2006
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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

This week

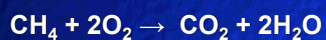
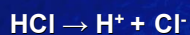
Weak acid titration.

Determine K_a of acetic acid by a couple
different methods.

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

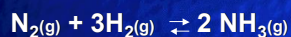
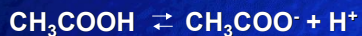
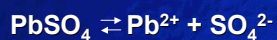
Non-Equilibrium Reaction

Reaction goes to completion.



Equilibrium Reaction

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



Equilibrium Expression

For any equilibrium reaction,



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

For a weak acid dissociation,



Two Ecksepshins

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

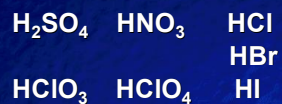


$$K_{\text{sp}} = [\text{Pb}^{2+}][\text{SO}_4^{2-}]$$



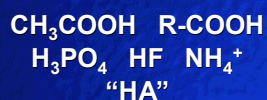
$$K_{\text{w}} = [\text{H}^+][\text{OH}^-]$$

Strong Acids



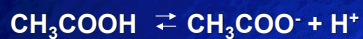
Weak Acids

All the rest!



pH of weak acids

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



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

Dissociation of acetic acid

	CH ₃ COOH	CH ₃ COO ⁻	H ⁺
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

$$x^2 = 0.1K_a$$

$$x^2/0.1 = K_a$$

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⁻
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

	CH ₃ COOH	CH ₃ COO ⁻	H ⁺
initial	1.5	0	0
equilibrium	1.5 - x	x	x

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

Assume $x \ll 1.5$ 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