Experiment 9
Acid-Base Equilibria

CH 204

Fall 2007

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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.

$$HCl \rightarrow H^+ + Cl^-$$

$$KNO_3 \rightarrow K^+ + NO_3^-$$

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

Equilibrium Reaction

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

$$PbSO_4 \rightleftarrows Pb^{2+} + SO_4^{2-}$$

$$CH_3COOH \rightleftarrows CH_3COO^- + H^+$$

$$N_{2(g)} + 3H_{2(g)} \rightleftarrows 2\ NH_{3(g)}$$

Equilibrium Expression

For any equilibrium reaction,

$$aA + bB \rightleftharpoons cC + dD$$

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

For a weak acid dissociation,

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

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

Strong Acids

H_2SO_4 HNO_3 HCl
HBr
$HClO_3$ $HClO_4$ HI

Weak Acids

All the rest!

CH_3COOH R-COOH
H_3PO_4 HF NH_4^+
"HA"

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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
1.78×10^{-5}	- 4.75	4.75
4.68×10^{-7}	- 6.33	6.33
8.13×10^{-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?

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

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Review

If you are given the $[H^+]$

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

If you are given the pH

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

pH of weak acids

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

$$CH_3COOH \rightleftarrows CH_3COO^- + H^+$$

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

Dissociation of acetic acid

	CH_3COOH	CH_3COO^-	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 << 0.1 M

$$K_a = x^2 / 0.1$$

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

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. 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.

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, $[HA] = [A^-]$.

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

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

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

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

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

Deja Vu

	CH_3COOH	CH_3COO^-	H^+
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 << 1.4 M

$$K_a = x^2/1.4$$

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

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.

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.

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.