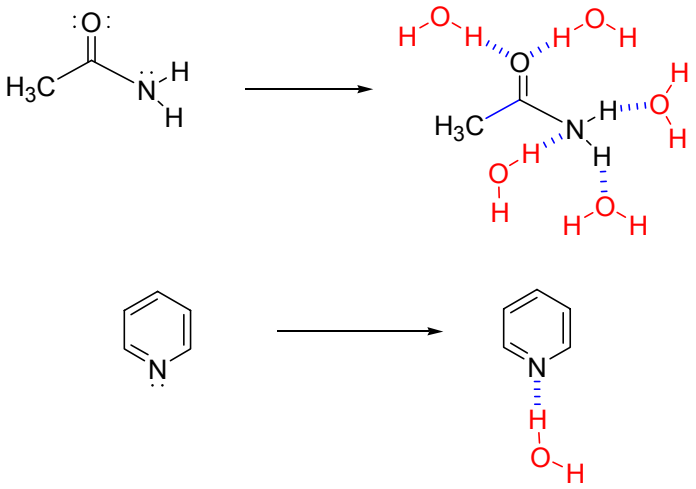
BMB 401 (Spring 2004) Answers to Problem Set #1

1. Answers to Problems from Lehninger given in a separate document.
2.



Two of the hydrogen bonds shown above for acetamide (Not formamide as stated in the problem) are not ideal. Can you pick out which two?

Also, notice that there is a numbering problem in the problem set handout. I give the number 2 to two successive problems. Sorry.

2 again! The Dielectric constant of a solvent is related to its dipole moment and hydrogen bonding ability, and generally reflects the molecule's polarity.

a) Formamide has a significantly larger dipole moment than acetone. In addition, it can form hydrogen bonds, which is not the case for acetone.

b) Methanol has a higher dipole moment than ethanol, although only slightly. It's smaller size permits more hydrogen bonding interactions per unit volume. The more non-polar nature of ethanol gives it a lower dielectric constant.

c) Carbon tetrachloride is a symmetrical molecule. It has no net dipole because the four microscopic dipoles cancel out. Chloroform, on the other hand has a net dipole.

3a) pH= -log [H+]

HCl is a strong acid therefore,

pH = -log(0.1M)

= 1

b) There are several ways to do this:

The fastest is pOH = -log[OH] for a strong base

= 1

pH + pOH = 14

pH = 13

c) Acetic Acid is a weak acid

$$CH_3COOH \rightleftharpoons CH_3COO^{\ominus} + H^+$$

$Ka = 1.74 \times 10^{-5}$

0.1-X X X

[X][X]/(0.1-X) simplify because the assumption that X is much smaller than 0.1

Therefore: $X^2/0.1 = 1.74 \times 10^{-5}$

pH= 2.88

d) HClO4 is a strong acid. It's concentration is 300 times higher than the [H+] of pure water. Therefore, H+ from water can be neglected. The pH is 4.52.

e) You're adding base to pure water. The pH cannot be less than 7. If you take the negative logarithm (7.7) and use it in the following equation, pH + 7.7 = 14 = 6.3, you will get a pH that is less than 7. You must account for the hydroxide concentration in pure water. Therefore the total [OH] = $2 \times 10^{-8} + 1 \times 10^{-7} = 1.2 \times 10^{-7}$

pOH = -log (1.2×10^{-7}) = 6.92 **pH=14-6.92 = 7.08**

4) pH = pKa + log $[A^-]/[HA]$

5 = 4.76 + log $[A^-]/[HA]$

0.24 = log $[A^-]/[HA]$

1.74 = $[A^-]/[HA]$

In a 0.2M acetate solution, the ratio of the base to the acid should be 1.74.

Let A^- = X

Let HA = 0.2-X

$$X/0.2-X = 1.74$$

$$X = 0.348 - 1.74X$$

$$2.74X = 0.348$$

$$X = 0.127$$

You need **0.127M sodium acetate** and therefore **0.073M Acetic Acid**

5) Imidazole #1

First, find the concentration of acid and conjugate base at pH = 6.7 using the H-H equation. $pH = pKa + \log [A^-]/[HA]$

$$6.7 = 6.99 + \log [A^-]/[HA]$$

$$-0.29 = \log [A^-]/[HA]$$

$$0.51 = [A^-]/[HA]$$

$$0.51\,[HA] = [A^-]$$

The total imidazole buffer concentration is 0.10M, therefore the total of the two species must equal this concentration, $[HA] + [A^-] = 0.10M$. Solve for the concentration of A-, $[A^-] = 0.1M - [HA]$. Substitute into the above equation to solve for the concentrations, $0.51([HA]) = 0.1M - [HA]$.

$$1.51[HA] = 0.1M$$

$$[HA] = 0.066M \text{ and}$$

$$[A^-] = 0.034M \text{ at pH } 6.7$$

Upon addition of 0.02M NaOH, 0.02M of the acid, HA, will be deprotonated to form 0.02M additional A^-. Therefore at the final pH, the concentration of HA will be 0.046M and the concentration of A- will be 0.054M. You can now solve for the final pH using the H-H equation.

$$pH = pKa + \log [A^-]/[HA]$$

$$pH = 6.99 + \log 0.54/0.46$$

$$pH = 6.99 + 0.070$$

pH = 7.1 after addition of 0.02M NaOH

****The following two problems should be calculated in the same manner as the above.**

Buffer Species	Buffer Conc.	Initial pH	pKa	NaOH Added
*Imidazole	0.03M	6.5	6.99	0.02M
Phosphate	1.0 M	6.35	7.21	0.10M

***Imidazole: New pH should be 7.95**
Phosphate: New pH should be 6.66