Announcements – Lecture XVI – Tuesday, Nov 13th

1. Lab 5 … Saturday, November 17th, 1:00-4:00 pm ISB 155/160 A-E

3. Final Exam … Wednesday, December 12th, ISB 135, 8:00-10:00am
   Final Review … Sunday, December 9th, ISB 135, 1:00-3:00pm

8.2 How Do We Define the Strength of Acids and Bases?

Acids:

Strong: $HA(aq) + H_2O(l) \longrightarrow H_3O^+ + A^-$ ~100% ... only 6

$HCl, HBr, HI, HNO_3, H_2SO_4, HClO_4$

Weak: $HA(aq) + H_2O(l) \Leftrightarrow H_3O^+ + A^-$ $K_a = \frac{[H_3O^+][A^-]}{[HA]}$

Bases:

Strong: "$B(aq) + H_2O(l) \longrightarrow BH^+ + OH^-$" ~100% .... only 4

$NaOH(aq) \longrightarrow Na^+ + OH^-$ $LiOH, NaOH, KOH, Ba(OH)_2$

Weak: $B(aq) + H_2O(l) \Leftrightarrow BH^+ + OH^-$ $K_b = \frac{[BH^+][OH^-]}{[B]}$

$NH_3(aq) + H_2O(l) \Leftrightarrow NH_4^+ + OH^-$

8.2 How Do We Define the Strength of Acids and Bases?

K_a Values			K_a Values		
Name of Acid	**Acid**	**K_a**	**Name of Acid**	**Acid**	**K_a**
Sulfuric acid	H_2SO_4	large	Hexaaquaaluminum ion	$Al(H_2O)_6^{3+}$	7.9×10^{-6}
Hydrochloric acid	HCl	large	Carbonic acid	H_2CO_3	4.2×10^{-7}
Nitric acid	HNO_3	large	Hydrogen sulfide	H_2S	1×10^{-7}
Hydronium ion	H_3O^+	1.0	Dihydrogen phosphate ion	$H_2PO_4^-$	6.2×10^{-8}
Hydrogen sulfate ion	HSO_4^-	1.2×10^{-2}	Hypochlorous acid	$HClO$	3.5×10^{-8}
Phosphoric acid	H_3PO_4	7.5×10^{-3}	Ammonium ion	NH_4^+	5.6×10^{-10}
Hexaaquairon(III) ion	$Fe(H_2O)_6^{3+}$	6.3×10^{-3}	Hydrocyanic acid	HCN	4.0×10^{-10}
Hydrofluoric acid	HF	7.4×10^{-4}	Hexaaquairon(II) ion	$Fe(H_2O)_6^{2+}$	3.2×10^{-10}
Formic acid	HCO_2H	1.8×10^{-4}	Hydrogen carbonate ion	HCO_3^-	4.8×10^{-11}
Benzoic acid	$C_6H_5CO_2H$	6.3×10^{-5}	Hydrogen phosphate ion	HPO_4^{2-}	3.6×10^{-13}
Acetic acid	CH_3CO_2H	1.8×10^{-5}	Water	H_2O	1.0×10^{-14}
			Hydrogen sulfide ion	HS^-	1×10^{-19}

For weak acids ... the greater the Ka ... the stronger the acid.

8.5 **How Do We Use Acid Ionization Constants?**

pKa Versus Ka

$$pK_a = -\log_{10} K_a$$

HF : $K_a = 7.4 \times 10^{-4}$ $\quad pK_a = -\log_{10}(7.4 \times 10^{-4}) = 3.13$

HCN : $K_a = 4.0 \times 10^{-10}$ $\quad pK_a = -\log_{10}(4.0 \times 10^{-10}) = 9.38$

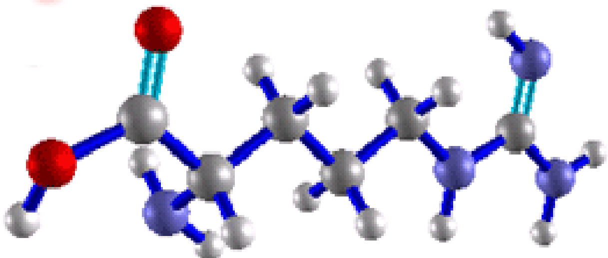
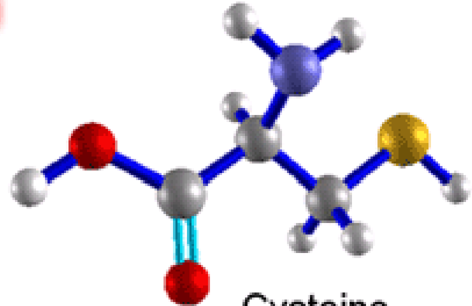
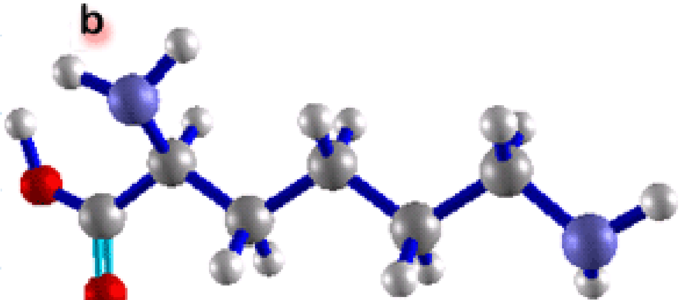
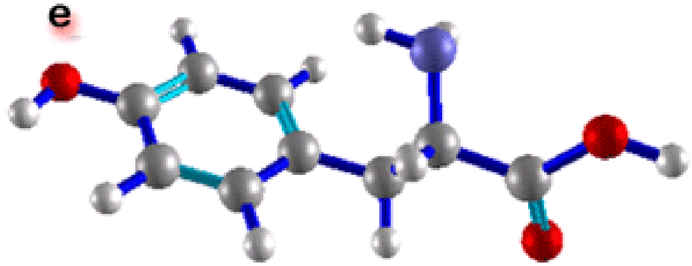
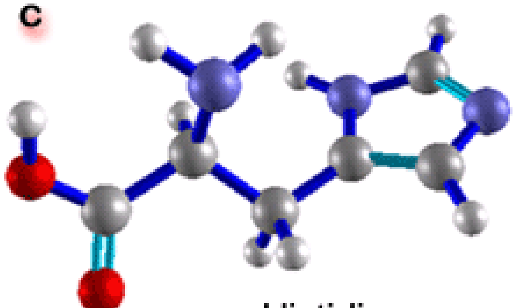
Which is the stronger acid?

a) The one with the largest Ka ... HF

b) The one with the smallest pKa ... HF

8.5 How Do We Use Acid Ionization Constants?

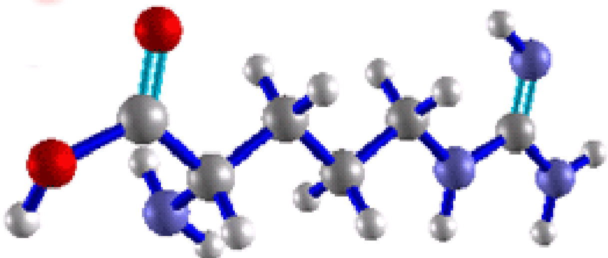
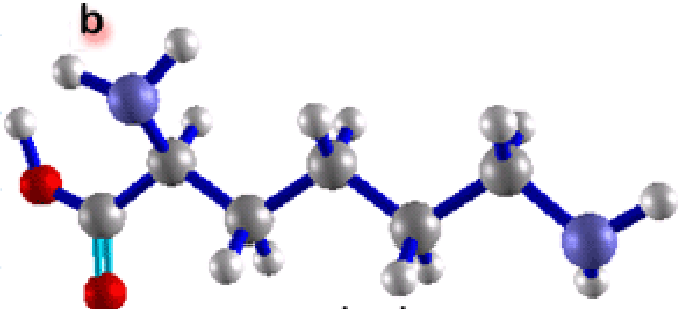
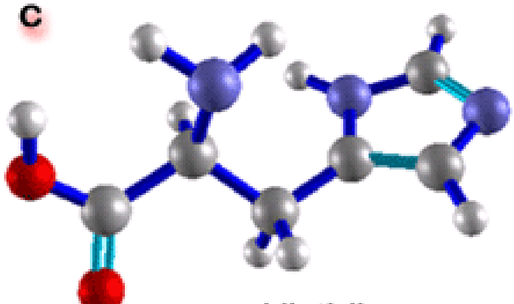
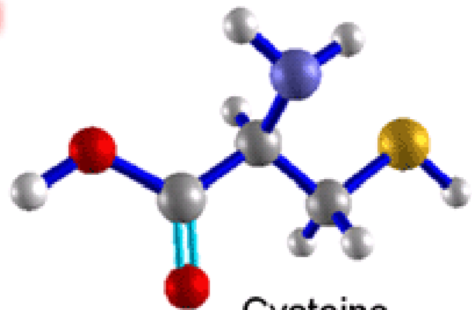
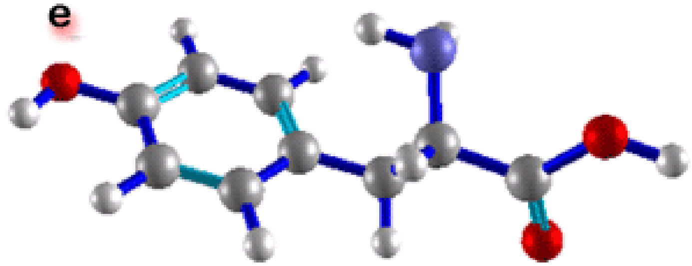
pKa Versus Ka

		pKa			pKa
a	 Arginine	12.0	d	 Cysteine	8.3
b	 Lysine	9.0	e	 Tyrosine	9.8
c	 Histidine	6.1			

The strongest acid depicted?

6.1 → Smallest pKa

pKa Versus Ka

		pKa
a	 Arginine	12.0
b	 Lysine	9.0
c	 Histidine	6.1
d	 Cysteine	8.3
e	 Tyrosine	9.8

The one with an acid strength closest to that of NH_4^+, $K_a = 5.6\times10^{-10}$ @ 25°C?

$$pK_a = -\log_{10}(5.6\times10^{-10}) = \boxed{9.25}$$

Acid Base Properties of Pure Water

Autoionization of Water

$$H_2O(l) + H_2O(l) \Leftrightarrow H_3O^+ + OH^-$$

$$K = [H_3O^+][OH^-]$$

↳ K_W

@ 25°C , $K_W = 1 \times 10^{-14}$

$$[H_3O^+][OH^-] = 1 \times 10^{-14}$$

$$[H_3O^+] = 1 \times 10^{-7}$$

$$[OH^-] = 1 \times 10^{-7}$$

NEUTRAL: $[H_3O^+] = [OH^-]$

Acidic: $[H_3O^+] > [OH^-]$

Basic: $[OH^-] > [H_3O^+]$

8.7 Acid Base Properties of Pure Water

Curiosity!

The autoionization of water is an endothermic process.
$H_2O(l) + H_2O(l) \Leftrightarrow H_3O^+ + OH^-$

Thus as the temperature increases
then – the $[H_3O^+]$ should – 

a) Decrease b) Increase c) Remain the same

shift →

$$H_2O(l) + H_2O(l) + heat \Leftrightarrow H_3O^+ + OH^-$$

↑ Increase T

8.7 Acid Base Properties of Pure Water

Curiosity!

With the $[H_3O^+]$ increasing with increasing temperature this must mean that as the temperature of water increases the water –

a) becomes acidic

b) becomes basic

c) remain neutral

shift →

$$H_2O(l) + H_2O(l) + heat \rightleftharpoons H_3O^+ + OH^-$$

↑ Increase T

but

$[H_3O^+]$ still equals the $[OH^-]$

8.7 Acid Base Properties of Pure Water

Example I

An aqueous solution has a hydronium ion, H_3O^+, concentration of $1x10^{-11}M$ @ 25ºC. This solution is –

a) acidic
b) basic
c) neutral

$K_w = 1 \times 10^{-14}$ @ 25°C

$K_w = [H_3O^+][OH^-]$

$[H_3O^+][OH^-] = 1 \times 10^{-14}$

$(1 \times 10^{-11})[OH^-] = 1 \times 10^{-14}$

$[OH^-] = \frac{1 \times 10^{-14}}{1 \times 10^{-11}} = 1 \times 10^{-3}$

$[OH^-] > [H_3O^+]$

8.8 What are pH and pOH?

$$pH = -\log_{10}[H_3O^+] \qquad pOH = -\log_{10}[OH^-]$$

@ 25°C $\quad [H_3O^+][OH^-] = 1\times10^{-14}$

$\log_{10}$ of both sides: $\quad \log_{10}\{[H_3O^+][OH^-]\} = \log_{10}(1\times10^{-14})$

tidy this up: $\quad \log_{10}[H_3O^+] + \log_{10}[OH^-] = -14$

multiply both sides by (-1): $\quad \underbrace{-\log_{10}[H_3O^+]}_{pH} \underbrace{- \log_{10}[OH^-]}_{pOH} = 14$

$$\boxed{pH + pOH = 14 \quad @\ 25°C}$$