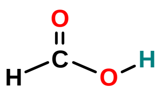
CHEM 1202 - Homework # 6 ***ANSWER KEY***

Acids & Bases # 1

Due Thursday, March 19, 2009

These compounds are the bases – you need to be able to tell acids from bases

1. (3 pts) Consider the following weak acids and bases and their pK_a values:

A)  (pK_a = 3.8) **B)** HCN (pK_a = 9.3) **C)** H_2CO_3 (pK_a = 6.4)

D) $N(CH_3)_3$ (pK_a = 12) **E)** H_2SO_3 (pK_a = 1.8) **F)** NH_3 (pK_a = 9)

Which compound is the strongest acid (use letter) ? **E**

The strongest acid has the lowest pK_a value

Which compound is the strongest base (use letter) ? **D**

The strongest base has the highest pK_a value. Note that the pK_a actually refers to the strength of the conjugate acid derived from the base

Which compound has the strongest conjugate base (use letter) ? **B**

The weakest acid has the strongest conjugate base. The weakest acid has the largest pK_a value. You need to be able to tell an acid from a base.

2. (5 pts) What are the pH's for the following solutions?

a) 0.1 *M* HBr = **1** b) 10 *M* H_2SO_4 = **–1** c) 1×10^{-10} *M* HNO_3 = **7**

d) 0.1 *M* NaOH = **13** e) 10 *M* CsOH = **15**

3. (4 pts) What is the pK_a value of a 0.1 *M* solution of palmetic acid (HA) that has a pH of 5? Clearly show all your work and put a box around your answer.

	HA(*aq*)	⇌	H^+(*aq*)	+	A^-(*aq*)
Initial:	**0.1 *M***		**0 *M***		**0 *M***
@ Equib:	**0.1 – x**		**x**		**x**

BUT, before you go any further, I've given you the pH of the solution in the problem!! So you don't have to solve for x, you already know it!! So all we have to do is take the anti-log of pH 5 to get the concentration of the H^+ (and A^-):

$[H^+]$ = antilog(–5) = 1×10^{-5} *M* Substituting this into the equlibrium expression gives us:

$$K_a = \frac{(1\times10^{-5})^2}{0.1-1\times10^{-5}} = \frac{1\times10^{-10}}{0.1} = 1\times10^{-9}$$

so, $K_a = 1 \times 10^{-9}$ **pK_a = –log(K_a) = 9**

Note that it is fine to drop the x in the 0.1 – x expression because x is much smaller than 0.1. So the answer is:

4. (3 pts) Calculate the pH of a 0.01 *M* solution of acid that has a pK_a of 6.0. Clearly show all your work and put a box around your answer.

	HA(*aq*)	⇌	H^+(*aq*)	+	A^-(*aq*)
Initial:	**0.01 *M***		**0 *M***		**0 *M***
@ Equib:	**0.01 – x**		**x**		**x**

First, convert the pK_a of 6 into a K_a value: K_a = antilog($-pK_a$) = 1×10^{-6} Now, we can solve for the $[H^+]$ and the pH.

$K_a = \frac{(x)(x)}{(0.01-x)} = 1\times10^{-6}$ assume that x << 1, $\frac{x^2}{(0.01)} = 1\times10^{-6}$, or $x^2 = 1 \times 10^{-8}$,

or $x = [H^+] = 1 \times 10^{-4}$ So the pH = $-\log(1 \times 10^{-4})$ = 4 **pH = 4**

5. (5 pts) What is the pH of a 0.1 *M* solution of the base ethyl amine ($CH_3CH_2NH_2$). $K_a = 1 \times 10^{-11}$ Clearly show all your work and put a box around your answer.

Initial:	**0.1 M**	**pure liq**		**0 M**	**0 M**
	Base(*aq*)	+ $H_2O(l)$	⇌	$BaseH^+(aq)$	+ $OH^-(aq)$
@ Equib:	**0.1 – x**	**pure liq**		**x**	**x**

You first need to convert the K_a into a K_b for this ***basic*** equilibrium:

$$K_b = \frac{K_w}{K_a} = \frac{1\times10^{-14}}{1\times10^{-11}} = 1\times10^{-3}$$ -- now you can setup your equilib: $$K_b = \frac{(x)(x)}{(0.1-x)} = 1\times10^{-3}$$

assume that x << 1, $\frac{(x)(x)}{(0.1)} = 1\times10^{-3}$ or $x^2 = 1 \times 10^{-4}$, or $x = [OH^-] = 1 \times 10^{-2}$ pOH = 2. ***BUT THIS IS NOT YOUR ANSWER***, since I asked for the pH!! **pH = 14 – 2 = 12** **pH = 12**

6. (5 pts) What is the K_b of a 0.01 *M* solution of a base that has a pH of 10? Clearly show all your work and put a box around your answer.

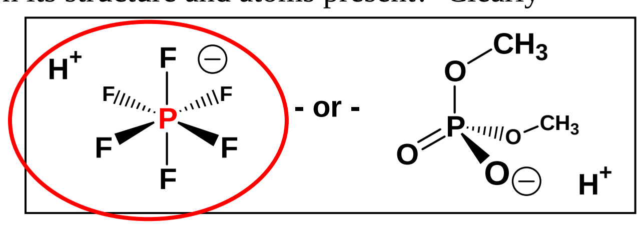
Initial:	**0.01 M**	**pure liq**		**0 M**	**0 M**
	$Base^-(aq)$	**+ $H_2O(l)$**	⇌	**BaseH (*aq*)**	**+ $OH^-(aq)$**
@ Equib:	**1 – x**	**pure liq**		**x**	**x**
@ Equib:	**$0.01 – 1 \times 10^{-4}$**			**1×10^{-4}**	**1×10^{-4}**

A pH of 10 means that the pOH is 4. We do NOT use the pH directly in this problem because we are dealing with a basic equilibrium!! Thus, we know that the $[OH^-] = 1 \times 10^{-4}$ M.

Plug into the K_b expression and solve: $$K_b = \frac{(1\times10^{-4})^2}{0.01-1\times10^{-4}} = 1\times10^{-6}$$ **$K_b = 1 \times 10^{-6}$**

You can drop this since it is small compared to 0.01 M

7. (5 pts) Which of the following acids is the strongest based on its structure and atoms present? Clearly discuss your reasoning.



HPF_6 is an extremely strong acid, while $HPO_2(OCH_3)_2$ is a rather weak acid. The two main reasons is that the negative charge for the $[PF_6]^-$ anion is spread out evenly over the 6 flourine atoms making for a very "dilute" charge. In contrast, the negative charge in the $[PO_2(OCH_3)_2]^-$ anion is localized (concentrated) on one (or at most two) oxygen atoms. More localized negative charges will more strongly attract the H^+ cation. This is why F^- (small concentrated charge) by itself is a moderately good conjugate base for the weak acid HF. But the $[PF_6]^-$ anion is not a simple F^- anion, its charge is spread out over 6 flourine atoms.

The other significant factor is that fluorine has a valence of 1, that is, it only wants to make one chemical bond to another atom. In the $[PF_6]^-$ anion each fluorine is already coordinated to the central phosphorus atom and won't want to make another covalent bond to a H^+. The $[PO_2(OCH_3)_2]^-$ anion, on the other hand, has one oxygen atom with only a single bond to the central phosphorus atom (and a negative charge). Oxygen atoms like to have two chemical bonds in general (H_2O !!) so it will have a fairly strong tendency to want to make a chemical bond to the H+. Remember that OH- loves to coordinate to H^+. The PO_3 group attached to the O^- is electron-withdrawing, so the P-O^- group will not be as electron-rich as OH^- and will not want to bond to the H^+ nearly as strongly.