CHM 1046 - JOENS WORKSHEET #8 Due date: Wednesday, October 17th

WORKSHEETS ARE DUE AT THE BEGINNING OF CLASS ON THE DATE GIVEN ON THE WORKSHEET. LATE WORKSHEETS WILL NOT BE ACCEPTED.

NAME ______________________________ Panther ID ____________________________

For problems involving calculations you must show your work for credit.

1) Find the pH for a 0.200 M aqueous solution of phenol (C_6H_5OH). The value for K_a for phenol at this temperature is $K_a = 1.3 \times 10^{-10}$. Also find the percent dissociation of phenol for these conditions.

The reaction is $C_6H_5OH(aq) + H_2O(\ell) \leftrightarrows H_3O^+(aq) + C_6H_5O^-(aq)$

$$K_a = \frac{[H_3O^+]\,[C_6H_5O^-]}{[C_6H_5OH]} = 1.3 \times 10^{-10}$$

	Initial	Change	Equilibrium
H_3O^+	0	x	x
$C_6H_5O^-$	0	x	x
C_6H_5OH	0.200	- x	0.200 - x

So $$\frac{(x)\,(x)}{(0.200 - x)} = 1.3 \times 10^{-10}$$

If we assume x << 0.200, then we get $$\frac{x^2}{0.200} = 1.3 \times 10^{-10}$$

$x^2 = (1.3 \times 10^{-10})\,(0.200) = 2.6 \times 10^{-11}$

$x = (2.6 \times 10^{-11})^{1/2} = 5.1 \times 10^{-6}$ Since x is small compared to 0.200,our assumption was good

$[H_3O^+] = 5.1 \times 10^{-6}$ M pH = $-\log_{10}(5.1 \times 10^{-6}) = 5.29$

$$\text{percent dissociation} = \frac{[C_6H_5O^-]_{eq}}{[C_6H_5OH]_{initial}} \times 100\ \% = \frac{5.1 \times 10^{-6}}{0.200} \times 100\ \% = 2.6 \times 10^{-3}\ \%$$

2) A 0.100 M aqueous solution of which of the following ionic compounds would be expected to have a pH much less than 7.0?

a) $NaNO_3$
b) KF
c) KOH
d) Both a and b
e) None of the above _____E______

$NaNO_3$ is the salt of a strong acid and a strong base, and so when added to water will give pH = 7.0. KF is the salt of a weak acid and a strong base, and so it contains a weak base (F^-). When added to water it will give pH > 7.0 KOH is a strong soluble base, and so it will also give a pH > 7.0 So the correct answer is E. None of the solutions will give pH < 7.0

3) The pH of a 0.0400 M aqueous solution of a weak monoprotic acid HA, measured at T = 25.0 °C, is pH = 3.71.

a) Find the value for $[H_3O^+]$ and $[OH^-]$ present in the above solution.

pOH = 14.00 - pH = 14.00 - 3.71 = 10.29

$[H_3O^+] = 10^{-3.71} = 1.95 \times 10^{-4}$ M $\qquad$ $[OH^-] = 10^{-10.29} = 5.1 \times 10^{-11}$ M

b) Find the value for K_a for the above acid, at T = 25.0 °C.

The reaction is $\quad HA(aq) + H_2O(\ell) \leftrightarrows H_3O^+(aq) + A^-(aq)$

$$K_a = \frac{[H_3O^+]\,[A^-]}{[HA]}$$

If we set up an ICE table, we get

	Initial	Change	Equilibrium
H_3O^+	0	x	x
A^-	0	x	x
HA	0.0400	- x	0.0400 - x

But x = $[H_3O^+] = 1.95 \times 10^{-4}$ M, so $[A^-] = x = 1.95 \times 10^{-4}$ M
$[HA] = 0.0400 - x = 0.0400 - 1.95 \times 10^{-4} = 3.98 \times 10^{-2}$ M

So $\quad K_a = \dfrac{(1.95 \times 10^{-4})\,(1.95 \times 10^{-4})}{(3.98 \times 10^{-2})} = 9.6 \times 10^{-7}$

4) The value for the acid ionization constant for hypochlorous acid (HOCl) is $K_a = 3.5 \times 10^{-8}$ at T = 25.°C. What is the value for K_b for the hypochlorite ion (OCl^-) at this temperature?

For an acid/conjugate base pair we know (at T = 25. °C) that $K_a K_b = 1.0 \times 10^{-14}$

So $K_b = \dfrac{1.0 \times 10^{-14}}{K_a} = \dfrac{1.0 \times 10^{-14}}{3.5 \times 10^{-8}} = 2.9 \times 10^{-7}$