Mr. Bracken
AP Chemistry

Name BRACKEN KEY
Period ____

Acids & Bases Worksheet #3.5

1. The pH of an acetic acid solution is measured to be 2.25. What is the molarity of this solution?

$HC_2H_3O_2 + H_2O \rightleftarrows H_3O^+ + C_2H_3O_2^-$ $K_a = 1.8 \times 10^{-5}$ *1.75 M*

$HC_2H_3O_2$	H_3O^+	$C_2H_3O_2^-$
y	0	0
−x	+x	+x
y−x	x	x

from pH

$[H_3O^+] = .0056$ M

$x = .0056$

$K_a = 1.8 \times 10^{-5} = \frac{x^2}{y-x}$ $1.8 \times 10^{-5} = \frac{[.0056]^2}{y-.0056}$ $y = 1.75$

Bracken Guess .1

2. A student adds 21 grams of solid NaF to 450 mL of distilled water. Assuming the volume of the water does not change, what is the expected pH of the resulting solution?

$\frac{21}{42} = .50$ moles NaF $\frac{.50}{.45} = 1.1$ M NaF

pH = 8.61

$F^- + HOH \rightleftarrows HF + OH^-$ $K_b = \frac{1.0 \times 10^{-14}}{6.8 \times 10^{-4}} = 1.5 \times 10^{-11}$

F^-	HF	OH^-
1.1 M	0	0
−x	+x	+x
1.1−x	x	x

$K_b = 1.5 \times 10^{-11} = \frac{x^2}{1.1-x}$ $x = 4.1 \times 10^{-6}$

$-\log(4.1 \times 10^{-6}) = pOH$

$pOH = 5.39$

$pH = 8.61$

3. A student adds 9.7 grams of sulfamic acid (HNH_2SO_3) crystals to 2.0 Liters of water. If the pH is 1.44, what is the acid dissociation constant, Ka, for sulfamic acid?

$\frac{9.7}{97} = .10$ moles HNH_2SO_3 $\frac{.10}{2.0} = .050$M HNH_2SO_3

Ka= 0.093
(Rounding differences)

$HNH_2SO_3 + H_2O \rightleftarrows H_3O^+ + NH_2SO_3^-$

HNH_2SO_3	H_3O^+	$NH_2SO_3^-$
.050M	0	0
−x	+x	+x
.050−x	x	x

from pH,

$[H^+] = .036$ M

$x = .036$

$K_a = \frac{x^2}{.050-x}$ $K_a = \frac{[.036]^2}{[.050-.036]}$ $K_a = .093$

4. What is the expected pH of a solution that was prepared by adding 53 grams of solid NH_4Cl to 0.50 Liters of distilled water? Assume the volume of the water does not change.

NH_4Cl moles $= \frac{53}{53} = 1.0$ mole $\qquad \frac{1.0}{.50} = 2.0\,M\ NH_4Cl$

NH_4^+	+ H_2O ⇌	H_3O^+	+	NH_3
2.0 M		0		0
−x		+x		+x
2.0 − x		x		x

$K_a = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$

$K_a = 5.6 \times 10^{-10} = \frac{x^2}{[2.0 - x]}$ $\qquad x = 3.3 \times 10^{-5}$

$x = [H_3O^+]$

$pH = -\log[3.3 \times 10^{-5}]$

$pH = 4.48$

pH = 4.48

5. A solution of NH_3 was measured to have a pH of 10.4. What is the molarity of this solution?

NH_3	+ H_2O ⇌	NH_4^+	+	OH^-
y		0		0
−x		+x		+x
y − x		x		x

$K_b = 1.8 \times 10^{-5}$

0.0037 M

from pH, get

pOH = 3.6

$[OH^-] = 2.5 \times 10^{-4}$ M

This is my x-value

$K_b = \frac{[2.5 \times 10^{-4}][2.5 \times 10^{-4}]}{[y - 2.5 \times 10^{-4}]} = 1.8 \times 10^{-5}$

Bracken Guess = .1 $\qquad y = .0037\ M$

6. What percent of a 0.23 M solution of HF is ionized?

HF	+ H_2O ⇌	H_3O^+	+	F^-
.23 M		0		0
−x		+x		+x
.23 − x		x		x

5.2%

$K_a = 6.8 \times 10^{-4} = \frac{x^2}{[.23 - x]}$ $\qquad x = .012$

% ionization $= \frac{.012}{.23} \times 100 = 5.2\%$