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AP Chemistry

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Period ____

Acids & Bases Worksheet #3

Hydrolysis Reactions:

I. Weak Acids "Dump a proton onto water" **$HA + H_2O \leftrightarrows H_3O^+ + A^{1-}$**

HCN + HOH $\leftrightarrows$ $H_3O^+ + CN^-$

HF + HOH $\leftrightarrows$ $H_3O^+ + F^-$

$HC_2H_3O_2$ + HOH $\leftrightarrows$ $H_3O^+ + C_2H_3O_2^-$

II. Weak Bases "Take a proton from water to make a weak acid" **$A^{1-} + H_2O \leftrightarrows HA + OH^{1-}$**

NH_3 + HOH $\leftrightarrows$ $NH_4^+ + OH^-$

$C_6H_5NH_2$ + HOH $\leftrightarrows$ $HC_6H_5NH_2^+ + OH^-$

CH_3NH_2 + HOH $\leftrightarrows$ $HCH_3NH_2^+ + OH^-$

III. Conjugate Base of a Weak Acid "One proton away from weak acid"

F^{1-} + HOH $\leftrightarrows$ $HF + OH^-$

CN^{1-} + HOH $\leftrightarrows$ $HCN + OH^-$

$C_2H_3O_2^{1-}$ + HOH $\leftrightarrows$ $HC_2H_3O_2 + OH^-$

IV. Conjugate Acid of a Weak Base "Act like weak acids"

NH_4^{1+} + HOH $\leftrightarrows$ $NH_3 + H_3O^+$

1. What is the pH of a 0.13 M HF solution?

$HF + H_2O \rightleftharpoons H_3O^+ + F^-$ $\quad K_a = 6.8 \times 10^{-4}$

.13 M	0	0
$-x$	$+x$	$+x$
$.13 - x$	x	x

$[H_3O^+] = .0090\ M$

$pH = 2.04$

$6.8 \times 10^{-4} = \dfrac{[x][x]}{[.13-x]}$ $\quad x = .0090$

pH = 2.04

2. What is the pH of a 1.4 M NH_3 solution?

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$ $\quad K_b = 1.8 \times 10^{-5}$

1.4 M	0	0
$-x$	$+x$	$+x$
$1.4 - x$	x	x

$[OH^-] = .0050\ M$

$pOH = 2.30$

$pH = 11.70$

$1.8 \times 10^{-5} = \dfrac{[x][x]}{[1.4-x]}$ $\quad x = .0050$

pH = 11.70

3. Calculate the pH of a 0.89 M $HC_2H_3O_2$ solution.

$HC_2H_3O_2 + H_2O \rightleftharpoons H_3O^+ + C_2H_3O_2^-$ $\quad K_a = 1.8 \times 10^{-5}$

.89 M	0	0
$-x$	$+x$	$+x$
$.89 - x$	x	x

$[H_3O^+] = .0040\ M$

$pH = 2.40$

$1.8 \times 10^{-5} = \dfrac{[x][x]}{[.89-x]}$ $\quad x = .0040$

pH = 2.40

4. What is the pH of a 0.50 M KF solution?

$F^- + H_2O \rightleftharpoons HF + OH^-$ $\quad K_b = \dfrac{1.0 \times 10^{-14}}{6.8 \times 10^{-4}}$

.50 M	0	0
$-x$	$+x$	$+x$
$.50 - x$	x	x

$K_b = 1.5 \times 10^{-11}$

$[OH^-] = 2.7 \times 10^{-6}\ M$

$pOH = 5.56$

$pH = 8.44$

$1.5 \times 10^{-11} = \dfrac{[x][x]}{[.50-x]}$ $\quad x = 2.7 \times 10^{-6}$

pH = 8.44

5. What is the pH of a 0.22 M solution of NaClO ?

$ClO^- + H_2O \rightleftharpoons HClO + OH^-$ $\quad K_b = \frac{1.0 \times 10^{-14}}{3.0 \times 10^{-8}}$ **pH = 10.43**

.22M	0	0
−x	+x	+x
.22 − x	x	x

$K_b = 3.3 \times 10^{-7}$

$3.3 \times 10^{-7} = \frac{[x][x]}{[.22-x]}$ $\quad x = 2.7 \times 10^{-4}$

$[OH^-] = 2.7 \times 10^{-4}\ M$

$pOH = 3.57$

$pH = 10.43$

6. What is the pH of a 0.034 M solution of KCN ?

$CN^- + H_2O \rightleftharpoons HCN + OH^-$ $\quad K_b = \frac{1.0 \times 10^{-14}}{4.9 \times 10^{-10}}$ **pH = 10.91**

.034M	0	0
−x	+x	+x
.034 − x	x	x

$K_b = 2.0 \times 10^{-5}$

$2.0 \times 10^{-5} = \frac{[x][x]}{[.034-x]}$ $\quad x = 8.1 \times 10^{-4}$

$[OH^-] = 8.1 \times 10^{-4}\ M$

$pOH = 3.09$

$pH = 10.91$

7. What is the pH of a 0.65 M pyridine (C_5H_5N) solution?

$C_5H_5N + H_2O \rightleftharpoons HC_5H_5N^+ + OH^-$ $\quad K_b = 1.7 \times 10^{-9}$ **pH = 9.52**

.65M	0	0
−x	+x	+x
.65 − x	x	x

$1.7 \times 10^{-9} = \frac{[x][x]}{[.65-x]}$ $\quad x = 3.3 \times 10^{-5}$

$[OH^-] = 3.3 \times 10^{-5}\ M$

$pOH = 4.48$

$pH = 9.52$

8. What is the pH of a 0.56 M NH_4Br solution?

$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$ $\quad K_a = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}}$ **pH = 4.75**

.56M	0	0
−x	+x	+x
.56 − x	x	x

$K_a = 5.6 \times 10^{-10}$

$5.6 \times 10^{-10} = \frac{[x][x]}{[.56-x]}$ $\quad x = 1.8 \times 10^{-5}$

$[H_3O^+] = 1.8 \times 10^{-5}\ M$

$pH = 4.74$

9. Calculate the pH of a 0.79 M $NaC_2H_3O_2$ solution.

$C_2H_3O_2^- + H_2O \rightleftharpoons HC_2H_3O_2 + OH^-$

.79 M		0	0
−x		+x	+x
.79 − x		x	x

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

$K_b = 5.6 \times 10^{-10}$

$5.6 \times 10^{-10} = \frac{[x][x]}{[.79 - x]}$

$x = 2.1 \times 10^{-5}$

$[OH^-] = 2.1 \times 10^{-5}$ M

pOH = 4.68

pH = 9.32

pH = 9.32

10. What is the pH of a 0.16 M NH_3 solution?

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$ $\quad K_b = 1.8 \times 10^{-5}$

.16 M		0	0
−x		+x	+x
.16 − x		x	x

$1.8 \times 10^{-5} = \frac{[x][x]}{[.16 - x]}$

$x = .0017$

$[OH^-] = .0017$

pOH = 2.77

pH = 11.23

pH = 11.23

Slight Thinker!

11. A 0.020 M solution of a weak acid (HA) has a pH of 3.26.

from pH, $[H_3O^+] = 5.5 \times 10^{-4}$ M

thus, $x = 5.5 \times 10^{-4}$

(a) What is the Ka for this weak acid?

$HA + H_2O \rightleftharpoons H_3O^+ + A^-$

.020 M		0	0
− x		+x	+x
.020 − x		x	x

$K_a = \frac{[5.5 \times 10^{-4}]^2}{[.020 - 5.5 \times 10^{-4}]}$

$K_a = 1.6 \times 10^{-5}$

Ka = 1.6 x 10⁻⁵

(b) What percentage of this weak acid HA is ionized in this 0.020 M solution?

$\frac{5.5 \times 10^{-4}}{.020} \times 100 = 2.75\%$

2.7%