Acid/Base Problems

High $[H^+]$ weak acid

Neutral

weak base

Low $[H^+]$ Strong Base

#2 $N_2H_4 + H_2O \rightleftarrows N_2H_5^+ + OH^-$

$$K_b = \frac{[N_2H_5^+] \times [OH^-]}{[N_2H_4]}$$

4. $C_6H_5COOH + H_2O \rightleftarrows C_6H_5COO^- + H_3O^+$

2. $C_6H_5COOH + H_2O \rightleftarrows H_3O^+ + C_6H_5COO^-$

	C_6H_5COOH	H_2O	H_3O^+	$C_6H_5COO^-$
I	0.20		0	0
C	$-x$		$+x$	$+x$
E	$0.20-x$		x	x

$$K_a = \frac{[H_3O^+][C_6H_5COO^-]}{[C_6H_5COOH]}$$

$$6.3\times10^{-5} = \frac{x \cdot x}{0.20 - x}$$ assume small

$$6.3\times10^{-5} = \frac{x^2}{0.20}$$

$$\sqrt{1.26\times10^{-5}} = \sqrt{x^2}$$

3.6×10^{-3} mol/L $= x = [H_3O^+]$

Percent (%) Ionization Problem 2

A 1.2 mol/L solution of base ionizes 13%. Calculate K_b.

$$B + H_2O \rightleftarrows BH^+ + OH^-$$

	B	H_2O	BH^+	OH^-
I	1.2		0	0
C	-0.156		+0.156	0.156
E	1.044		0.156	0.156

13% of base

0.13 x 1.2 = 0.156

$$K_b = \frac{[BH^+][OH^-]}{[B]} = \frac{(0.156)(0.156)}{1.044}$$

$K_b = 0.0233$ (2.33×10^{-2})

% Ionization example

A 0.50 mol/L solution of HCOOH has a $K_a = 1.8 \times 10^{-4}$. Calculate % ionization

$$HCOOH + H_2O \rightleftharpoons H_3O^+ + HCOO^-$$

	$HCOOH$	H_2O	H_3O^+	$HCOO^-$
I	0.50		0	0
C	$-x$		$+x$	$+x$
E	$0.50-x$		x	x

$$K_a = \frac{[H_3O^+][HCOOH]}{[HCOOH]}$$

$$1.8 \times 10^{-4} = \frac{x^2}{0.50-x}$$

assume x is small

$$\sqrt{0.50 \times 1.8 \times 10^{-4}} = \sqrt{x^2}$$

$$9.5 \times 10^{-3} = x$$

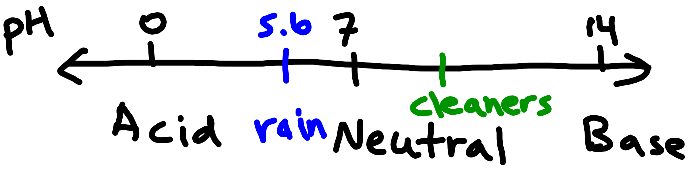
% ionization

$$= \frac{9.5 \times 10^{-3}}{0.50} \times 100\%$$

$$= 1.9\%$$

Worksheet-
#1,3,4,8

pH-Definition and Calculations



Definition- "p" means $-\log$

pH means $-\log$ (hydronium ion)

Logarithms-

$\log 100 = 2 \qquad 10^2 = 100$

$\log 10 = 1 \qquad 10^1 = 10$

$\log 1 = 0 \qquad 10^0 = 1$

$\log 0.1 = -1 \qquad 10^{-1} = 0.1$

2. $[H_3O^+] \longrightarrow pH$

a) $[H_3O^+] = 1 \times 10^{-7}$ $\quad pH = -\log(1 \times 10^{-7})$

$= -(-7) = \underline{7}$

b) $[H_3O^+] = 3.3 \times 10^{-2}$ $\quad pH = -(-1.48)$

$= 1.48$

3. $pH \longrightarrow [H_3O^+]$

ex. $\underline{8.1 \longrightarrow [H_3O^+]}$

$8.1 = -\log[H_3O^+]$

$-8.1 = \log[H_3O^+]$

$10^{-8.1} = [H_3O^+]$

Use 10^x button $\quad 7.9 \times 10^{-9} = [H_3O^+]$

pH and Ka/Kb Problems

Ex. Find the pH of a 0.45 mol/L solution of HF ($K_a = 6.6 \times 10^{-4}$)

$$HF + H_2O \rightleftarrows H_3O^+ + F^-$$

	HF	H_2O	H_3O^+	F^-
I	0.45		0	0
C	$-x$		$+x$	$+x$
E	$0.45-x$		x	x

$$K_a = \frac{[H_3O^+][F^-]}{[HF]}$$

$$6.6\times10^{-4} = \frac{x^2}{0.45-x}$$

assume....

$$1.7\times10^{-2} = x$$

$$pH = -\log(1.7\times10^{-2})$$
$$= -(-1.8)$$
$$= 1.8$$

Worksheet
#2,7,5

Text p.591 #5,7