Chem 1515 Section 2
Problem Set #10
Spring 1998

Name____________________________

TA Name ________________________

Lab Section #______

ALL work must be shown to receive full credit. **Due Friday, March 27th**

PS10.1. For aqueous solutions of the following substances, write the dissociation reaction and indicate whether the substance behaves as an Arrhenius acid or base.

a) $HF_{(aq)} \rightleftarrows$ $\mathbf{H^+_{(aq)} + F^-_{(aq)}}$ **acid**

b) $HC_6H_5O_{(aq)} \rightleftarrows$ $\mathbf{H^+_{(aq)} + C_6H_5O^-_{(aq)}}$ **acid**

c) $Ba(OH)_{2(aq)} \rightarrow$ $\mathbf{Ba^{2+}_{(aq)} + 2OH^-_{(aq)}}$ **base**

d) $LiOH_{(aq)} \rightarrow$ $\mathbf{Li^+_{(aq)} + OH^-_{(aq)}}$ **base**

e) $H_2O_{(aq)} \rightleftarrows$ $\mathbf{H^+_{(aq)} + OH^-_{(aq)}}$ **neutral**

f) $H_2CO_{3(aq)} \rightleftarrows$ $\mathbf{H^+_{(aq)} + HCO_3^-_{(aq)}}$ **acid**

PS10.2. Calculate the pH and pOH in each of the following aqueous solutions. In each case, indicate whether the solution is acidic or basic.

a) $[H^+] = 3.89\text{x } 10^{-5}$ M
pH = 4.41
pOH = 9.59
acidic

b) $[OH^-] = 8.34 \text{ x } 10^{-2}$ M
pH = 12.92
pOH = 1.08 basic

c) $[OH^-] = = 1.50 \text{ x } 10^{-7}$ M ($[OH^-]$ in milk)
pH = 7.18
pOH = 6.82 basic

d) $[H^+] = 9.39 \text{ x } 10^{-10}$ M
pH = 9.02
pOH = 4.97 basic

e) $[H^+] = 4.0$ M
pH = -0.60
pOH = 14.6 acidic

f) $[OH^-] = 10.1$ M
pH = 15.0
pOH = -1.00
basic

PS10.3. Calculate the $[H^+]$ and $[OH^-]$ in each of the following aqueous solutions.

a) pH = 3.40 (pH of orange juice)

$\mathbf{pH = -log[H^+]}$

$\mathbf{-3.40 = log[H^+]}$

$\mathbf{10^{-3.40} = 10^{log[H^+]}}$

$\mathbf{3.98 \text{ x } 10^{-4} \text{ M} = [H^+]}$

$\mathbf{K_w = 1 \text{ x } 10^{-14} = [H^+][OH^-]}$

$$\mathbf{[OH^-] = \frac{1 \text{ x } 10^{-14}}{[H^+]}}$$

$$\mathbf{[OH^-] = \frac{1 \text{ x } 10^{-14}}{3.98 \text{ x } 10^{-4} \text{ M}}}$$

$\mathbf{[OH^-] = 2.51 \text{ x } 10^{-11} \text{ M}}$

b) pH = 6.7 (pH of silva)
$\mathbf{[H^+] = 1.99 \text{ x } 10^{-7} \text{ M}}$
$\mathbf{[OH^-] = 5.01 \text{ x } 10^{-8} \text{ M}}$

c) pH = 4.4 (pH of beer)
$\mathbf{[H^+] = 3.98 \text{ x } 10^{-5} \text{ M}}$
$\mathbf{[OH^-] = 2.51 \text{ x } 10^{-10} \text{ M}}$

d) pOH = 2.15
$\mathbf{[H^+] = 1.41 \text{ x } 10^{-12} \text{ M}}$
$\mathbf{[OH^-] = 7.07 \text{ x } 10^{-3} \text{ M}}$

e) pOH = 12.4
$\mathbf{[H^+] = 2.51 \text{ x } 10^{-2} \text{ M}}$
$\mathbf{[OH^-] = 3.98 \text{ x } 10^{-13} \text{ M}}$

f) pH = -0.650
$\mathbf{[H^+] = 4.47 \text{ M}}$
$\mathbf{[OH^-] = 2.24 \text{ x } 10^{-15} \text{ M}}$

PS10.4. For each of the following acids, write the formula for the conjugate base.

a)	$H_2PO_4^-$	c)	H_2O	e)	OH^-
	$\mathbf{HPO_4^{2-}}$		$\mathbf{OH^-}$		$\mathbf{O^{2-}}$
b)	$HClO_3$	d)	$CH_3CH_2NH_3^+$	f)	NH_4^+
	$\mathbf{ClO_3^-}$		$\mathbf{CH_3CH_2NH_2}$		$\mathbf{NH_3}$

PS10.5. For each of the following bases, write the formula for the conjugate acid.

a)	OH^-	c)	HCO_3^-	e)	CH_3NH_2
	$\mathbf{H_2O}$		$\mathbf{H_2CO_3}$		$\mathbf{CH_3NH_3^+}$
b)	Cl^-	d)	H_2O	f)	$(CH_3)_3N$
	$\mathbf{HCl}$		$\mathbf{H_3O^+}$		$\mathbf{(CH_3)_3NH^+}$

PS10.6. For the following compounds, write the reaction with water and indicate the Brø nsted acids, bases, the conjugate acid and conjugate base.

a) $HBr_{(g)} + H_2O_{(l)} \rightleftarrows H_3O^+_{(aq)} + Br^-_{(aq)}$

A B CA CB

b) $NH_{3(g)} + H_2O_{(l)} \rightleftarrows NH_4^+_{(aq)} + OH^-_{(aq)}$

B A CA CB

c) $HCN_{(g)} + H_2O_{(l)} \rightleftarrows H_3O^+_{(aq)} + CN^-_{(aq)}$

A B CA CB

d) $HC_7H_5O_{2(s)} + H_2O_{(l)} \rightleftarrows H_3O^+_{(aq)} + C_7H_5O_2^-_{(aq)}$

A B CA CB

e) $CH_3NH_{2(l)} + H_2O_{(l)} \rightleftarrows CH_3NH_3^+_{(aq)} + OH^-_{(aq)}$

B A CA CB

PS10.7. For each of the following compounds, write two Brønsted-Lowry equations, one showing how the substance behaves as an acid, the second showing how the substance behaves as a base.

a) $HCO_3^-{}_{(aq)}$

$HCO_3^-{}_{(aq)} + H_2O_{(l)} \rightleftarrows H_3O^+{}_{(aq)} + CO_3^{2-}{}_{(aq)}$ acid

$HCO_3^-{}_{(aq)} + H_2O_{(l)} \rightleftarrows OH^-{}_{(aq)} + H_2CO_3{}_{(aq)}$ base

b) $NH_3{}_{(aq)}$

$NH_3{}_{(g)} + H_2O_{(l)} \rightleftarrows NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$ base

$NH_3{}_{(g)} + H_2O_{(l)} \rightleftarrows NH_2^-{}_{(aq)} + H_3O^+{}_{(aq)}$ acid

c) $HPO_4^{2-}{}_{(aq)}$

$HPO_4^{2-}{}_{(aq)} + H_2O_{(l)} \rightleftarrows H_3O^+{}_{(aq)} + PO_4^{3-}{}_{(aq)}$ acid

$HPO_4^{2-}{}_{(aq)} + H_2O_{(l)} \rightleftarrows OH^-{}_{(aq)} + H_2PO_4^-{}_{(aq)}$ base

d) $HSO_4^-{}_{(aq)}$

$HSO_4^-{}_{(aq)} + H_2O_{(l)} \rightleftarrows H_3O^+{}_{(aq)} + SO_4^{2-}{}_{(aq)}$ acid

$HSO_4^-{}_{(aq)} + H_2O_{(l)} \rightleftarrows OH^-{}_{(aq)} + H_2SO_4{}_{(aq)}$ base

PS10.7. Determine the equilibrium constant for the following solutions. (Show your work clearly!)

a) 0.250 M HF whose pH = 1.89.

	HF	$\rightleftarrows$	H^+	+	F^-	
I	**.250**		**~0**		**0**	
C	**-x**		**+x**		**+x**	**x = $[HF]_D$**
E	**.250 - x**		**x**		**x**	

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

Since the pH = 1.89, the $[H^+] = 1.29 \times 10^{-2}$ M which is equal to x as shown in the ICE table above.

$$K_a = \frac{(x)(x)}{(0.250 - x)}$$

Substituting for x,

$$K_a = \frac{(1.29 \times 10^{-2})^2}{.250 - 0.0129} = 7.02 \times 10^{-4}$$

b) 0.235 M NH_3 whose pH = 11.31.

	NH_3	$+ H_2O$	$\rightleftarrows$	$NH_4^+(aq)$ +	$OH^-(aq)$	
I	**0.235 M**	-		**0**	**0**	
C	**-x**	-		**+x**	**+x**	**let x = $[NH_3]_R$**
E	**0.235 - x**	-		**0 + x**	**0 + x**	

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

Since the pH = 11.31, the pOH = 2.69 and the $[OH^-] = 2.0 \times 10^{-3}$ M which is also equal to x as shown in the ICE table above.

$$K_b = \frac{(x)(x)}{(0.1 - x)} = \frac{(2.0 \times 10^{-3})(2.0 \times 10^{-3})}{(0.235 - 2.0 \times 10^{-3})} = 1.79 \times 10^{-5}$$

c) 0.500 M B whose pH = 10.67.

	B	+ H_2O	⇄	$BH^+(aq)$ +	$OH^-(aq)$	
I	0.500 M	-		0	0	
C	-x	-		+x	+x	let $x = [B]_R$
E	0.500 - x	-		0 + x	0 + x	

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

Since the pH = 10.67, the pOH = 3.33 and the $[OH^-] = 4.68 \times 10^{-4}$ M which is also equal to x as shown in the ICE table above.

$$K_b = \frac{(x)(x)}{(0.1 - x)}$$

Substituting for x,

$$= \frac{(4.68 \times 10^{-4})(4.68 \times 10^{-4})}{(0.500 - 4.68 \times 10^{-4})}$$

$$K_b = \frac{(4.68 \times 10^{-4})^2}{0.500} = 4.38 \times 10^{-7}$$

d) 0.0751 M HA whose pH = 4.00.

	HA	⇄	H^+	+ A^-	
I	0.0751		~0	0	
C	-x		+x	+x	$x = [HA]_D$
E	0.0751 - x		x	x	

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

Since the pH = 4.00, the $[H^+] = 1.00 \times 10^{-4}$ M which is also equal to x as shown in the ICE table above.

$$K_a = \frac{(x)(x)}{(0.0751 - x)}$$

Substituting for x,

$$K_a = \frac{(1.00 \times 10^{-4})^2}{.0751 - 1.00 \times 10^{-4}}$$

$$= \frac{(1.00 \times 10^{-4})^2}{.0750} = 1.33 \times 10^{-7}$$

PS10.8. Given the following substances and their initial concentration:

a)	0.100 M HNO_3	e)	0.100 M HF	i)	0.100 M HC_6H_5O
b)	55.5 M H_2O	f)	0.100 M HNO_2	j)	0.100 M $Ba(OH)_2$
c)	0.100 M NaOH	g)	0.100 M CH_3NH_2	k)	0.00491 M HF
d)	0.100 M $C_2H_5NH_2$	h)	0.100 M C_5H_5N	l)	0.100 M HOCl

Answer the following,

i) identify each as an acid (**A**), base (**B**) or neutral (**N**) substance.

a) 0.100 M HNO_3	**A**	e) 0.100 M HF	**A**	i) 0.100 M HC_6H_5O	**A**
b) 55.5 M H_2O	**N**	f) 0.100 M HNO_2	**A**	j) 0.100 M $Ba(OH)_2$	**B**
c) 0.100 M NaOH	**B**	g) 0.100 M CH_3NH_2	**B**	k) 0.00491 M HF	**A**
d) 0.100 M $C_2H_5NH_2$	**B**	h) 0.100 M C_5H_5N	**B**	l) 0.100 M HOCl	**A**

ii) list the K_a value for each acid and K_b value for each base.

a) 0.100 M HNO_3	**K_a is very large**
b) 55.5 M H_2O	**$K_a = 1 \times 10^{-14}$**
c) 0.100 M NaOH	**K_b is very large**
d) 0.100 M $C_2H_5NH_2$	**$K_b = 6.4 \times 10^{-4}$**
e) 0.100 M HF	**$K_a = 6.8 \times 10^{-4}$**
f) 0.100 M HNO_2	**$K_a = 4.5 \times 10^{-4}$**
g) 0.100 M CH_3NH_2	**$K_b = 4.4 \times 10^{-4}$**
h) 0.100 M C_5H_5N	**$K_b = 1.7 \times 10^{-9}$**
i) 0.100 M HC_6H_5O	**$K_a = 1.3 \times 10^{-10}$**
j) 0.100 M $Ba(OH)_2$	**K_b is very large**
k) 0.00491 M HF	**$K_a = 6.8 \times 10^{-4}$**
l) 0.100 M HOCl	**$K_b = 3.0 \times 10^{-8}$**

iii) identify each substance as strong (S) or weak (W).

a) 0.100 M HNO_3	**SA**	e) 0.100 M HF	**WA**	i) 0.100 M HC_6H_5O	**WA**
b) 55.5 M H_2O	**N**	f) 0.100 M HNO_2	**WA**	j) 0.100 M $Ba(OH)_2$	**SB**
c) 0.100 M NaOH	**SB**	g) 0.100 M CH_3NH_2	**WB**	k) 0.00.491 M HF	**WA**
d) 0.100 M C_2H_5NH2	**WB**	h) 0.100 M C_5H_5N	**WB**	l) 0.100 M HOCl	**WA**

iv) calculate the $[H^+]$ and the pH of each of the solutions.

a) 0.100 M HNO_3

	HNO_3 $\rightarrow$	H^+	+	NO_3^-
I	**.1**		**~0**	**0**
F	**0**		**.1**	**.1**

HNO_3 is a strong acid and completely dissociates in water.
Therefore,

$[H^+]$ = 0.1 M and pH = 1

b) 55.5 M H_2O **$[H^+] = 1.0 \times 10^{-7}$ M: pH = 7.00**

c) 0.100 M NaOH

	$NaOH \rightarrow$	Na^+	+	OH^-
I	**.1**		**0**	**~0**
F	**0**		**.1**	**.1**

NaOH is a strong base and completely dissociates in water. Therefore,

$[OH^-] = 0.1$ M and pOH = 1, pH = 13: $[H^+] = 1.0 \times 10^{-13}$ M

d) 0.100 M $C_2H_5NH_2$

	$C_2H_5NH_2$	$+ H_2O$	$\rightleftarrows$	$C_2H_5NH_3^+(aq)$ +	$OH^-(aq)$	
I	**0.1 M**	**-**		**0**	**0**	
C	**-x**	**-**		**+x**	**+x**	**let x = $[C_2H_5NH_2]$R**
E	**0.1 - x**	**-**		**0+x**	**0+x**	

$$Kb = \frac{[C_2H_5NH_3^+][OH^-]}{[C_2H_5NH_2]}$$

$$6.4 \times 10^{-4} = \frac{(x)(x)}{0.1 - x} \quad \text{assume } x << 0.1$$

$$6.4 \times 10^{-4} = \frac{x^2}{.1}$$

$$6.4 \times 10^{-5} = x^2$$

$$8 \times 10^{-3} \text{ M} = x = [OH^-]$$

pOH = 2.10

pH = 11.90: $[H^+] = 1.26 \times 10^{-12}$ M

e) 0.100 M HF

	HF	$\rightleftarrows$	H^+	+	F^-	
I	**.1**		**~0**		**0**	
C	**-x**		**+x**		**+x**	**x = [HF]diss**
E	**.1-x**		**x**		**x**	

$$Ka = \frac{[H^+][F^-]}{[HF]}$$

$$6.8 \times 10^{-4} = \frac{(x)(x)}{0.1 - x} \qquad \text{assume } x << 0.1$$

$$6.8 \times 10^{-4} = \frac{(x)(x)}{0.1}$$

$$6.8 \times 10^{-5} = x^2$$

8.2 x 10-3 M = x = [H+]

pH = 2.09

f) 0.100 M HNO_2 **$[H^+] = 6.71 \times 10^{-3}$ M: pH = 2.17**

g) 0.100 M CH_3NH_2 **$[H^+] = 1.51 \times 10^{-12}$ M: pH = 11.82**

h) 0.100 M C_5H_5N **$[H^+] = 7.67 \times 10^{-10}$ M: pH = 9.11**

i) 0.10 M HC_6H_5O **$[H^+] = 3.61 \times 10^{-6}$ M: pH = 5.44**

j) 0.100 M $Ba(OH)_2$ **$[H^+] = 5.0 \times 10^{-14}$ M: pH = 13.3**

(the answer is not 13.0 since $Ba(OH)_2$ produces 2 moles of OH^-.)

k) 0.00491 M HF

	HF	$\rightleftarrows$	H^+	+	F^-		
I	**.00491**		**~0**		**0**		
C	**-x**				**+x**	**+x**	**x = [HF]diss**
E	**.00491-x**		**x**		**x**		

$$Ka = \frac{[H+][F^-]}{[HF]}$$

$$6.8 \times 10^{-4} = \frac{(x)(x)}{.00491-x}$$

$$6.8 \times 10^{-4} = \frac{x^2}{.00491-x}$$

$$3.39 \times 10^{-6} - 6.8 \times 10^{-4}x = x^2$$

$$3.39 \times 10^{-6} - 6.8 \times 10^{-4}x - x^2 = 0$$

$$x^2 + 6.8 \times 10^{-4}x - 3.39 \times 10^{-6} = 0$$

solving the quadratic equation $x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$

$$x = \frac{-6.8 \times 10^{-4} \pm \sqrt{(6.8 \times 10^{-4})^2 - 4(1)(-3.39 \times 10^{-6})}}{2(1)}$$

$$x = \frac{-6.8 \times 10^{-4} \pm 3.75 \times 10^{-3}}{2}$$

Use only the positive root

$x = 1.53 \times 10^{-3}$ M = $[H^+]$
pH = 2.82

1) 0.100 M HOCl **$[H^+] = 5.48 \times 10^{-5}$ M: pH = 4.26**

v) rank all substances from strongest acid...weakest acid...neutrals...
...weakest base..strongest base.

a) 0.100 M HNO_3	**pH = 1.00**
e) 0.100 M HF	**pH = 2.09**
f) 0.100 M HNO_2	**pH = 2.17**
k) 0.00491 M HF	**pH = 2.82**
l) 0.100 M HOCl	**pH = 4.26**
i) 0.100 M HC_6H_5O	**pH = 5.44**
b) 55.5 M H_2O	**pH = 7.00**
h) 0.100 M C_5H_5N	**pH = 9.11**
g) 0.100 M CH_3NH_2	**pH = 11.82**
d) 0.100 M $C_2H_5NH_2$	**pH = 11.90**
c) 0.100 M NaOH	**pH = 13.0**
j) 0.100 M $Ba(OH)_2$	**pH = 13.3**