

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) $\text{HF (aq)} \rightleftharpoons \text{H}^+ \text{ (aq)} + \text{F}^- \text{ (aq)}$ | acid |
| b) $\text{HC}_6\text{H}_5\text{O (aq)} \rightleftharpoons \text{H}^+ \text{ (aq)} + \text{C}_6\text{H}_5\text{O}^- \text{ (aq)}$ | acid |
| c) $\text{Ba(OH)}_2 \text{ (aq)} \rightarrow \text{Ba}^{2+} \text{ (aq)} + 2\text{OH}^- \text{ (aq)}$ | base |
| d) $\text{LiOH (aq)} \rightarrow \text{Li}^+ \text{ (aq)} + \text{OH}^- \text{ (aq)}$ | base |
| e) $\text{H}_2\text{O (aq)} \rightleftharpoons \text{H}^+ \text{ (aq)} + \text{OH}^- \text{ (aq)}$ | neutral |
| f) $\text{H}_2\text{CO}_3 \text{ (aq)} \rightleftharpoons \text{H}^+ \text{ (aq)} + \text{HCO}_3^- \text{ (aq)}$ | 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) $[\text{H}^+] = 3.89 \times 10^{-5} \text{ M}$
pH = 4.41
pOH = 9.59
acidic | d) $[\text{H}^+] = 9.39 \times 10^{-10} \text{ M}$
pH = 9.02
pOH = 4.97 basic |
| b) $[\text{OH}^-] = 8.34 \times 10^{-2} \text{ M}$
pH = 12.92
pOH = 1.08 basic | e) $[\text{H}^+] = 4.0 \text{ M}$
pH = -0.60
pOH = 14.6 acidic |
| c) $[\text{OH}^-] = 1.50 \times 10^{-7} \text{ M}$ ($[\text{OH}^-]$ in milk)
pH = 7.18
pOH = 6.82 basic | f) $[\text{OH}^-] = 10.1 \text{ M}$
pH = 15.0
pOH = -1.00 basic |

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

- | | |
|---|--|
| a) pH = 3.40 (pH of orange juice)
pH = $-\log[\text{H}^+]$
$-3.40 = \log[\text{H}^+]$
$10^{-3.40} = 10^{\log[\text{H}^+]}$
$3.98 \times 10^{-4} \text{ M} = [\text{H}^+]$
$K_w = 1 \times 10^{-14} = [\text{H}^+][\text{OH}^-]$
$[\text{OH}^-] = \frac{1 \times 10^{-14}}{[\text{H}^+]}$
$[\text{OH}^-] = \frac{1 \times 10^{-14}}{3.98 \times 10^{-4} \text{ M}}$
$[\text{OH}^-] = 2.51 \times 10^{-11} \text{ M}$ | c) pH = 4.4 (pH of beer)
$[\text{H}^+] = 3.98 \times 10^{-5} \text{ M}$
$[\text{OH}^-] = 2.51 \times 10^{-10} \text{ M}$ |
| b) pH = 6.7 (pH of saliva)
$[\text{H}^+] = 1.99 \times 10^{-7} \text{ M}$
$[\text{OH}^-] = 5.01 \times 10^{-8} \text{ M}$ | d) pOH = 2.15
$[\text{H}^+] = 1.41 \times 10^{-12} \text{ M}$
$[\text{OH}^-] = 7.07 \times 10^{-3} \text{ M}$ |
| | e) pOH = 12.4
$[\text{H}^+] = 2.51 \times 10^{-2} \text{ M}$
$[\text{OH}^-] = 3.98 \times 10^{-13} \text{ M}$ |
| | f) pH = -0.650
$[\text{H}^+] = 4.47 \text{ M}$
$[\text{OH}^-] = 2.24 \times 10^{-15} \text{ M}$ |

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

- | | | |
|---------------------------------------|---|-----------------------------------|
| a) H_2PO_4^- | c) H_2O | e) OH^- |
| HPO_4^{2-} | OH^- | O^{2-} |
| b) HClO_3 | d) $\text{CH}_3\text{CH}_2\text{NH}_3^+$ | f) NH_4^+ |
| ClO_3^- | $\text{CH}_3\text{CH}_2\text{NH}_2$ | NH_3 |

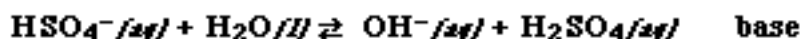
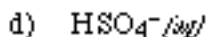
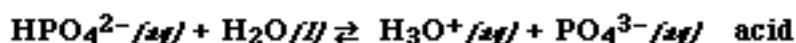
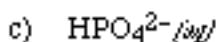
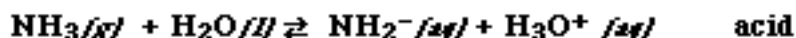
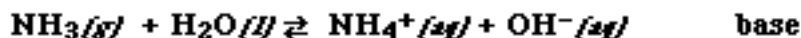
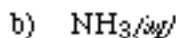
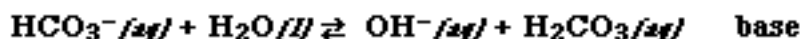
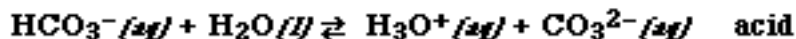
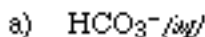
PS10.5. For each of the following bases, write the formula for the conjugate acid.

- | | | |
|--|---|--|
| a) OH^- | c) HCO_3^- | e) CH_3NH_2 |
| H_2O | H_2CO_3 | CH_3NH_3^+ |
| b) Cl^- | d) H_2O | f) $(\text{CH}_3)_3\text{N}$ |
| HCl | H_3O^+ | $(\text{CH}_3)_3\text{NH}^+$ |

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

- | |
|---|
| a) $\text{HBr}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{Br}^-(\text{aq})$ |
| A B CA CB |
| b) $\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$ |
| B A CA CB |
| c) $\text{HCN}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{CN}^-(\text{aq})$ |
| A B CA CB |
| d) $\text{HC}_7\text{H}_5\text{O}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{C}_7\text{H}_5\text{O}_2^-(\text{aq})$ |
| A B CA CB |
| e) $\text{CH}_3\text{NH}_2(\text{l}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{CH}_3\text{NH}_3^+(\text{aq}) + \text{OH}^-(\text{aq})$ |
| 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.



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

a) 0.250 M HF whose pH = 1.89.

	HF	\rightleftharpoons	H⁺	+	F⁻	
I	.250		~0		0	
C	-x		+x		+x	$x = [\text{HF}]_D$
E	.250 - x		x		x	

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

Since the pH = 1.89, the $[\text{H}^+] = 1.29 \times 10^{-2} \text{ M}$ 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.

	NH₃	+	H₂O	\rightleftharpoons	NH₄⁺(aq)	+	OH⁻(aq)	
I	0.235 M		-		0		0	
C	-x		-		+x		+x	let x = $[\text{NH}_3]_R$
E	0.235 - x		-		0 + x		0 + x	

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$

Since the pH = 11.31, the pOH = 2.69 and the $[\text{OH}^-] = 2.0 \times 10^{-3} \text{ M}$ 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	+ H₂O	⇌	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	

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

Since the pH = 10.67, the pOH = 3.33 and the $[\text{OH}^-] = 4.68 \times 10^{-4} \text{ M}$ 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	

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

Since the pH = 4.00, the $[\text{H}^+] = 1.00 \times 10^{-4} \text{ M}$ 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 ₆ H ₅ O |
| b) 55.5 M H ₂ O | f) 0.100 M HNO ₂ | j) 0.100 M Ba(OH) ₂ |
| c) 0.100 M NaOH | g) 0.100 M CH ₃ NH ₂ | k) 0.00491 M HF |
| d) 0.100 M C ₂ H ₅ NH ₂ | h) 0.100 M C ₅ H ₅ N | 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 ₆ H ₅ O | A |
| b) 55.5 M H ₂ O | 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 ₃ NH ₂ | B | k) 0.00491 M HF | A |
| d) 0.100 M C ₂ H ₅ NH ₂ | B | h) 0.100 M C ₅ H ₅ N | B | l) 0.100 M HOCl | A |

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

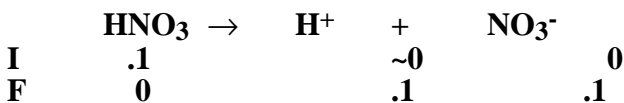
- | | |
|--|---|
| a) 0.100 M HNO ₃ | K_a is very large |
| b) 55.5 M H ₂ O | K_a = 1 x 10⁻¹⁴ |
| c) 0.100 M NaOH | K_b is very large |
| d) 0.100 M C ₂ H ₅ NH ₂ | K_b = 6.4 x 10⁻⁴ |
| e) 0.100 M HF | K_a = 6.8 x 10⁻⁴ |
| f) 0.100 M HNO ₂ | K_a = 4.5 x 10⁻⁴ |
| g) 0.100 M CH ₃ NH ₂ | K_b = 4.4 x 10⁻⁴ |
| h) 0.100 M C ₅ H ₅ N | K_b = 1.7 x 10⁻⁹ |
| i) 0.100 M HC ₆ H ₅ O | K_a = 1.3 x 10⁻¹⁰ |
| j) 0.100 M Ba(OH) ₂ | K_b is very large |
| k) 0.00491 M HF | K_a = 6.8 x 10⁻⁴ |
| l) 0.100 M HOCl | K_b = 3.0 x 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 ₆ H ₅ O | WA |
| b) 55.5 M H ₂ O | 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 ₃ NH ₂ | WB | k) 0.00491 M HF | WA |
| d) 0.100 M C ₂ H ₅ NH ₂ | WB | h) 0.100 M C ₅ H ₅ N | WB | l) 0.100 M HOCl | WA |

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

- a) 0.100 M HNO₃



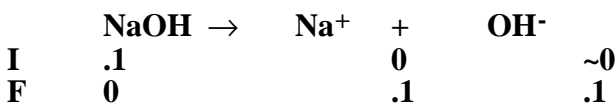
HNO₃ is a strong acid and completely dissociates in water.

Therefore,

$$[\text{H}^+] = 0.1 \text{ M and pH} = 1$$

- b) 55.5 M H₂O **[H⁺] = 1.0 x 10⁻⁷ M: pH = 7.00**

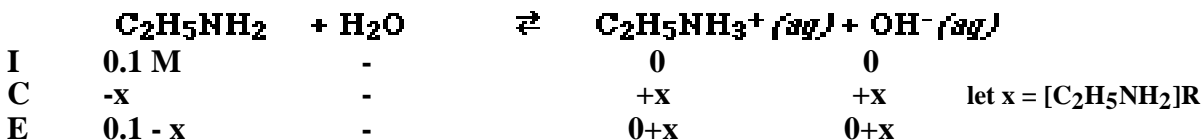
c) 0.100 M NaOH



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

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

d) 0.100 M C₂H₅NH₂



$$K_b = \frac{[\text{C}_2\text{H}_5\text{NH}_3^+][\text{OH}^-]}{[\text{C}_2\text{H}_5\text{NH}_2]}$$

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

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

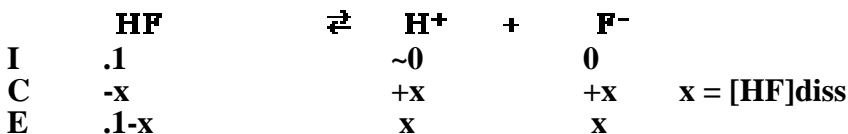
$$6.4 \times 10^{-5} = x^2$$

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

$$\text{pOH} = 2.10$$

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

e) 0.100 M HF



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

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

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

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

$$8.2 \times 10^{-3} \text{ M} = x = [\text{H}^+]$$

$$\text{pH} = 2.09$$

f) 0.100 M HNO₂ $[\text{H}^+] = 6.71 \times 10^{-3} \text{ M: pH} = 2.17$

g) 0.100 M CH₃NH₂ $[\text{H}^+] = 1.51 \times 10^{-12} \text{ M: pH} = 11.82$

h) 0.100 M C₅H₅N $[\text{H}^+] = 7.67 \times 10^{-10} \text{ M: pH} = 9.11$

i) 0.10 M HC₆H₅O $[\text{H}^+] = 3.61 \times 10^{-6} \text{ M: pH} = 5.44$

j) 0.100 M Ba(OH)₂ $[\text{H}^+] = 5.0 \times 10^{-14} \text{ M: pH} = 13.3$

(the answer is not 13.0 since Ba(OH)₂ produces 2 moles of OH⁻.)

k) 0.00491 M HF

	HF	\rightleftharpoons	H⁺	+	F⁻	
I	.00491		~0		0	
C	-x				+x	+x
E	.00491-x		x		x	x = [HF]diss

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

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

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

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

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

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

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

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

$$\text{pH} = 2.82$$

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

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

a) 0.100 M HNO₃	pH = 1.00
e) 0.100 M HF	pH = 2.09
f) 0.100 M HNO₂	pH = 2.17
k) 0.00491 M HF	pH = 2.82
l) 0.100 M HOCl	pH = 4.26
i) 0.100 M HC₆H₅O	pH = 5.44
b) 55.5 M H₂O	pH = 7.00
h) 0.100 M C₅H₅N	pH = 9.11
g) 0.100 M CH₃NH₂	pH = 11.82
d) 0.100 M C₂H₅NH₂	pH = 11.90
c) 0.100 M NaOH	pH = 13.0
j) 0.100 M Ba(OH)₂	pH = 13.3

