Chemistry 1105 Lab: pH and Buffers

Goals:

1. To become familiar with the reactions of acids and bases. Strong and weak.
2. Calculate the $[H^+(aq)]$, $[OH^-(aq)]$, and pH.
3. Calculating pH of acids and bases.
4. Determine K_A or K_B and pH of a weak acid or base.
5. Buffers.

Proton Concentration:

Proton is a hydrogen atom that has lost its lone electron.

$$H^+$$

H^+ very reactive. Attaches to water a molecule.

$$H_2O + H^+ \rightarrow H_3O^+$$

H_3O^+: hydronium ion

H_3O^+ same as $H^+(aq)$.

Important to know concentration, $[H^+(aq)]$.

Importance of $H^+(aq)$ Concentration:

Knowing the concentration of H^+ very important.

Ex: aquarium, pools, etc.

Fishtank

Aquarium in Toilet

Acids and Bases:

Bronsted/Lowry Theory

acid- Proton donor. Increase [H^+(aq)].

base- Proton acceptor. Decrease [H^+(aq)].

The strength of an acid or base depends on their ability to donate or accept protons.

K_w: Ion product of water:

Water also contains OH^- ions. At 25 °C

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$$

$$K_w = [H^+(aq)][OH^-(aq)] = 1.0\times10^{-14}$$

If $[H^+(aq)]$ increases, $[OH^-(aq)]$ decreases.

If you know one. Can calculate the other.

Quantifying Acidity(pH):

Could quantify acidity by concentration of aqueous H^+. Not practicle.

pH is a logarithmic scale of H^+ concentration.

$$pH = -Log[H^+(aq)] \quad \text{or} \quad pH = -Log[H_3O^+(aq)]$$

pH Scale:

The pH scale ranges from 0(very acidic) to 14(very basic). Pure water has a pH of 7.

0 - 7 - 14

Acidic Neutral Basic

Likewise,

$pOH = -Log[OH^{-}(aq)]$

$$pH + pOH = 14$$

Questions:

1. Determine $[H_3O^+]$ for the following:

a) $[OH^-] = 1.5\times10^{-2}$ M

$$K_w = [H^+(aq)][OH^-(aq)] = 1.0\times10^{-14}$$

$$[H^+(aq)] = \frac{1\times10^{-14}}{1.5\times10^{-2}\text{M}} = 6.7\times10^{-13}\text{M}$$

$$pH = -Log[H^+(aq)]$$

$$pH = -Log(6.7\times10^{-13}\text{M}) = 12.18$$

BASIC

b) pH = 7(Neutral); c) pH = 4.91(Acidic)

Questions cont:

2. Determine $[H_3O^+]$ and $[OH^-]$ for the following:

a) pH = 4.32

pH = -Log$[H^+(aq)]$ or

$[H^+(aq)]$ = antilog(-pH)

$[H^+(aq)]$ = antilog(-4.32) = 4.8×10^{-5}M

$[OH^-] = 2.1\times10^{-10}$ M

b) $[H^+] = 2.3\times10^{-10}$ M, $[OH^-] = 4.3\times10^{-5}$ M

c) $[H^+] = 1.3\times10^{-11}$ M, $[OH^-] = 7.8\times10^{-4}$ M

pH of a Weak Acid:

Consider a 1.0 M weak acid HA.

$$HA(aq) \rightleftharpoons H^+(aq) + A^-(aq)$$

	HA	H^+	A^-
I:	1.0 M	0	0
C:	-X	+X	+X
E:	1.0 - X	X	X

If X is small. 1.0 – X ≈ 1.0.

X = [H^+(aq)].

By measuring pH. Can calculate [H^+(aq)].

Questions cont:

3. Find the $[H_3O^+]$ of a 0.250 M solution of HF acid? Weak acid. Has a $K_A = 3.53\times10^{-4}$

$$HF(aq) \leftrightharpoons H^+(aq) + F^-(aq)$$

	HF(aq)	H⁺(aq)	F⁻(aq)
I:	0.250 M	0	0
C:	-X	+X	+X
E:	0.250 - X	X	X

$$K_A = \frac{[H^+(aq)][F^-(aq)]}{[HF(aq)]}$$

Questions cont:

$$K_A = \frac{(X)(X)}{(0.250 - X)} = \frac{X^2}{(0.250 - X)}$$

Assume X is small.

0.250 - X ~ 0.250

$$3.53\times10^{-4} = \frac{X^2}{(0.250)}$$

X = [H$^+$] = 0.00939 M

Questions cont:

4. Find the pH of a 0.250 M solution of acetic acid? Weak acid. $K_A = 1.76\times10^{-5}$

$$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$$

	$CH_3COOH(aq)$	$H^+(aq)$	$CH_3COO^-(aq)$
I:	0.250 M	0	0
C:	-X	+X	+X
E:	0.250 - X	X	X

$$K_A = \frac{[H^+(aq)][CH_3COO^-(aq)]}{[CH_3COOH(aq)]}$$

Questions cont:

$$K_A = \frac{(X)(X)}{(0.250 - X)} = \frac{X^2}{(0.250 - X)}$$

Assume *X* is small.

0.250 - *X* ~ 0.250

$$1.76\times10^{-5} = \frac{X^2}{(0.250)}$$

$X = [H^+] = 6.6\times10^{-4}$ M

pH = 3.18

Questions:

5. A 0.175 M weak acid solution has a pH of 3.25. Find the K_A of the acid.

pH of 3.25 corresponds to
$[H^+(aq)] = 5.6\times10^{-4}$ M

	$HW(aq)$	$\leftrightharpoons$	$H^+(aq)$	+	$W^-(aq)$
I:	0.175 M		0		0
C:	-5.6×10^{-4} M		$+5.6\times10^{-4}$ M		$+5.6\times10^{-4}$ M
E:	(0.175 M -5.6×10^{-4} M)		5.6×10^{-4} M		5.6×10^{-4} M

Questions cont:

	$HW(aq)$	$\rightleftharpoons$	$H^{+}(aq)$	+	$W^{-}(aq)$
I:	0.175 M		0		0
C:	-5.6×10^{-4} M		$+5.6\times10^{-4}$ M		$+5.6\times10^{-4}$ M
E:	(0.174 M)		5.6×10^{-4} M		5.6×10^{-4} M

$$K_A = \frac{(5.6\times10^{-4}\ \text{M})(5.6\times10^{-4}\ \text{M})}{(0.174\ \text{M})}$$

$$K_A = 1.8\times10^{-6}$$

Questions:

6. Find the $[OH^-(aq)]$ and pH of a 0.33 M methylamine solution. $K_B = 4.4\times10^{-4}$

$CH_3NH_2(aq) + H_2O(l) \rightleftharpoons OH^-(aq) + CH_3NH_3^+(aq)$

	$CH_3NH_2(aq)$	$OH^-(aq)$	$CH_3NH_3^+(aq)$
I:	0.33 M	0	0
C:	-X	+X	+X
E:	0.33 - X	X	X

$$K_B = \frac{[OH^-(aq)][CH_3NH_3^+(aq)]}{[CH_3NH_2(aq)]}$$

Questions cont:

$$K_B = \frac{(X)(X)}{(0.33 - X)} = \frac{X^2}{(0.33 - X)}$$

Assume X is small.

0.33 - X ~ 0.33

$$4.4\times10^{-4} = \frac{X^2}{(0.33)}$$

X = [OH$^-$] = 0.012 M

pH = 12.08

pH of a Weak Base:

Consider a 1.0 M weak base B.

$$B(aq) + H_2O(l) \rightleftharpoons HB(aq) + OH^-(aq)$$

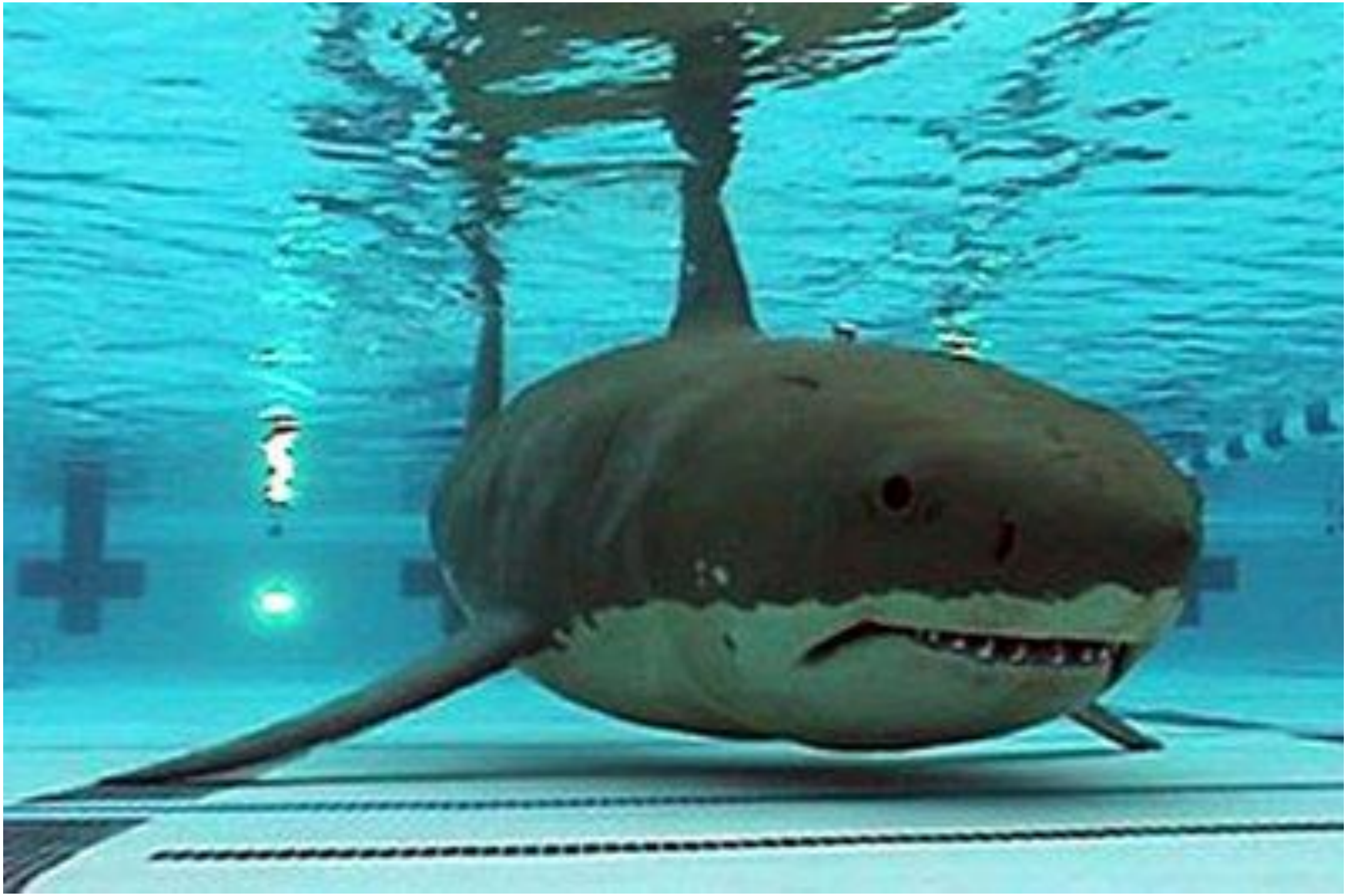
I:	1.0 M	0	0
C:	-X	+X	+X
E:	1.0 - X	X	X

If X is small. 1.0 – X ≈ 1.0.

X = [OH^-(aq)].

By measuring pH. Can calculate [OH^-(aq)].

Buffers:



Buffers:

A buffer is an aqueous system that allows only small changes in pH on the addition of acids or bases.

Consider a system consisting of equal amounts of $HC_2H_3O_2$ and $C_2H_3O_2^-$.

If strong base added:

$$HC_2H_3O_2(aq) + OH^-(aq) \rightarrow C_2H_3O_2^-(aq) + H_2O(l)$$

If strong acid added:

$$C_2H_3O_2^-(aq) + H^+(aq) \rightarrow HC_2H_3O_2(aq)$$

Identifying Buffers:

Buffers consist of a weak acid and its corresponding conjugate base.

Ex: Is the system NH_3/NH_4Cl a buffer?

Ex: Is the system HCl/Cl^- a buffer?

Calculating the pH of Buffer Systems:

Can use K_A or K_B to determine pH.

DIFFICULT!!!

Henderson-Hasselbach Equation:

$$pH = pK_A + \text{Log}\frac{[\text{base}]}{[\text{acid}]}$$

$$pH = pK_A + \text{Log}\frac{\text{moles of base}}{\text{moles of acid}}$$

Consider a buffer consisting of the weak acid HA and its conjugate base A^-.

$$pH = pK_A + \text{Log}\frac{[A^-]}{[HA]}$$

Where $pK_A = -\text{Log}(K_A)$

Questions:

7. Calculate the pH of a buffer system that is 0.250 M HCN and 0.170 M in KCN.
K_A for HCN = 4.9×10^{-10}

$$pH = pK_A + Log\frac{[KCN]}{[HCN]}$$

$$pK_A = -Log(4.9\times10^{-10}) = 9.31$$

$$pH = 9.31 + Log\frac{0.170\ M}{0.250\ M}$$

$$pH = 9.14$$

Questions:

8. What is the pH of a solution that contains 15.0 g of HF and 25.0 g of NaF in 125 mL?

$$pH = pK_A + \text{Log}\frac{[NaF]}{[HF]}$$

[HF] = 6.00 M **[NaF] = 4.76 M**

pH = 3.35

Percent Ionization of a Weak Acid:

$$\% \text{ ionization} = \frac{\text{ionized acid concentration}}{\text{initial concentration of acid}} \text{ x } 100\ \%$$

$$\% \text{ ionization} = \frac{[H^{+}(aq)]}{[\mathrm{HA(aq)}]} \text{ x } 100\ \%$$

Percent Ionization of a Weak Base:

$$\% \text{ ionization} = \frac{\text{ionized base concentration}}{\text{initial concentration of base}} \text{ x } 100\ \%$$

$$\% \text{ ionization} = \frac{[OH^{-}(aq)]}{[\text{B(aq)}]} \text{ x } 100\ \%$$