Unit 5: ACID/BASE 2

Chemistry 12 Acid/Base II -Helpful Equations

1) Strong acid in water: 100% dissociation

$$\underset{.10M}{HNO_3} + H_2O \longrightarrow \underset{.10M}{H_3O^+} + \underset{.10M}{NO_3^-}$$

pH = -log(.10) = 1.00

H_2SO_4 in water:
first proton is strong:

$$\underset{.10M}{H_2SO_4} + H_2O \longrightarrow \underset{.10M}{H_3O^+} + \underset{.10M}{HSO_4^-}$$

second proton is weak:

$$\underset{.10M}{HSO_4^-} + H_2O \rightleftharpoons \underset{<.10M}{H_3O^+} + \underset{<.10M}{SO_4^{2-}}$$

2) Weak acid in water: not 100% dissociation

$$\underset{.10M}{H_3PO_4} + H_2O \rightleftharpoons \underset{0.0279M}{H_3O^+} + \underset{0.0279M}{H_2PO_4^-}$$

pH = -log(0.0279) = 1.56

3) Strong base (hydroxide base) in water: 100% dissociation (water not in reaction; just a dissociation)

$$\underset{.10M}{Sr(OH)_2} \longrightarrow \underset{.10M}{Sr^{2+}} + \underset{.20M}{2OH^-}$$

pOH = -log(.20) = 0.70
pH = 13.30

4) Weak base in water: not 100% dissociation

$$\underset{.10M}{NH_3} + H_2O \rightleftharpoons \underset{1.34 \times 10^{-3}M}{NH_4^+} + \underset{1.34 \times 10^{-3}M}{OH^-}$$

pOH = -log(1.34 x 10^{-3}) = 2.87
pH = 11.13

5) ANY reaction that involves a STRONG acid or base goes to 100% completion. So a weak acid with a strong base is 100% due to the strong base. A strong acid with a weak base is 100% due to the strong acid.

Ex: $\underset{\text{weak}}{H_3PO_4} + \underset{\text{strong}}{3NaOH} \longrightarrow 3H_2O + Na_3PO_4$

The OH^- ions take all three protons off of each H_3PO_4 molecule, such as in a titration. If H_3PO_4 was merely in water, only one proton would come off at less than 100% like #2 earlier.

6) Weak acid and weak base:

$$NH_4^+ + SO_4^{2-} \rightleftharpoons NH_3 + HSO_4$$

-side with weaker acid is favoured

http://www.media.pearson.com.au/schools/cw/au_sch_derry_ibcsl_1/int/aqueous/tutor/f5/1501.html

Acid/Base 2 Notes Key

I) Weak Acid Equilibrium and K_a

How do strong acids behave in water?

Write an equation for, and give the $[H_3O^+]$ and pH for a 0.0010M solution of HCl.

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How do strong acids behave in water?

They react 100% with H_2O to create H_3O^+ - no equilibrium exists

Write an equation for, and give the $[H_3O^+]$ and pH for a 0.0010M solution of HCl.

$$\underset{0.0010M}{HCl} + H_2O \longrightarrow \underset{0.0010M}{H_3O^+} + Cl^- \qquad [H_3O^+] = 0.0010M$$

pH = -log(0.0010) = 3.00

How do weak acids behave in water?

http://www.absorblearning.com/media/attachment.action?quick=126&att=2737

Write an equation for a 0.010M HF solution.

http://www.chembio.uoguelph.ca/educmat/chm19104/chemtoons/chemtoons4.htm

How do weak acids behave in water?

They react with H_2O (much less than 100%) to create H_3O^+, and an equilibrium results where reactants are favoured

http://www.absorblearning.com/media/attachment.action?quick=126&att=2737

Write an equation for a 0.010M HF solution.

$$\underset{0.010M}{HF} + H_2O \rightleftharpoons \underset{<<0.010M}{H_3O^+} + F^-$$

http://www.chembio.uoguelph.ca/educmat/chm19104/chemtoons/chemtoons4.htm

It is not as simple to find the pH of a 0.010M solution of a weak acid as you must first have information on the extent of dissociation for the weak acid in question. K_a, the weak acid equilibrium constant, helps to determine this information and provides a means to solve **weak acid** problems. K_a is a K_{eq} for weak acids, so it is simply an equilibrium constant, and therefore all K_{eq} rules apply.

$HF_{(aq)} + H_2O_{(l)}$

$K_a =$

It is not as simple to find the pH of a 0.010M solution of a weak acid as you must first have information on the extent of dissociation for the weak acid in question. K_a, the weak acid equilibrium constant, helps to determine this information and provides a means to solve **weak acid** problems. K_a is a K_{eq} for weak acids, so it is simply an equilibrium constant, and therefore all K_{eq} rules apply.

$$HF_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + F^-_{(aq)}$$

$$K_a = \frac{[H_3O^+][F^-]}{[HF]} = \frac{[H_3O^+]^2}{[HF]}$$

Acid/Base 2 Notes Key

Notice the K_a values ____________ as you go down the table because the acids are getting progressively ______________ (creating ______ H_3O^+ in solution).

In Chemistry 12, we work on quantitative problems that involve weak acids and weak bases.

Notice the K_a values decrease as you go down the table because the acids are getting progressively weaker (creating less H_3O^+ in solution).

In Chemistry 12, we work on quantitative problems that involve weak acids and weak bases.

Weak acid quantitative problems can be broken into three types.
Type 1 Problems: Finding pH of a weak acid solution

Example: For the reaction:

$$H_2S_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + HS^-_{(aq)}$$

1) Calculate the $[H_3O^+]$ and pH if the $[H_2S]_i$ = 0.050M

Example: For the reaction:

$$H_2S_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + HS^-_{(aq)}$$

1) Calculate the $[H_3O^+]$ and pH if the $[H_2S]_i$ = 0.050M

$$H_2S_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + HS^-_{(aq)}$$

I	0.050M	0M	0M
C	- x	+ x	+ x
E	0.050 - x	x	x

Let $x = \Delta[H_3O^+]$

$K_a = \frac{[H_3O^+]^2}{[H_2S]}$

K_a from table = 9.1×10^{-8}

Since the % dissociation of H_2S is so small (less than 5%), we can assume that x is very small compared to 0.050M

assume $0.050 - x \cong 0.050$

thus, $9.1 \times 10^{-8} = \frac{x^2}{0.050}$

so, $x = \sqrt{(9.1 \times 10^{-8})(0.050)}$

$x = \underline{6.7}45 \times 10^{-5}$

$[H_3O^+] = 6.7 \times 10^{-5}$M, pH = 4.17

Example:

Calculate the pH of a 0.450M solution of $H_2PO_4^-$

*Consider only the first proton dissociating from a polyprotic weak acid, as the dissociation of the second proton is negligible compared to the first

Acid/Base 2 Notes Key

Example:
Calculate the pH of a 0.450M solution of $H_2PO_4^-$

*Consider only the first proton dissociating from a polyprotic weak acid, as the dissociation of the second proton is negligible compared to the first

	$H_2PO_4^-$ + H_2O $\rightleftharpoons$	H_3O^+ +	HPO_4^{2-}
I	0.450M	0M	0M
C	- x	+ x	+ x
E	0.450 - x	x	x

Let x = $\Delta[H_3O^+]$

assume 0.450 - x $\cong$ 0.450

$K_a = \frac{[H_3O^+]^2}{[H_2PO_4^-]}$

thus, $6.2 \times 10^{-8} = \frac{x^2}{0.450}$

K_a from table = 6.2×10^{-8}

so, $x = \sqrt{(6.2 \times 10^{-8})(0.450)}$

$x = \underline{1.67} \times 10^{-4}$

pH = $-\log(\underline{1.67} \times 10^{-4}) = 3.78$

Sometimes when salts dissolve in water, one of the ions can act as a weak acid in solution.
Example: What is the pH of a 0.100M NH_4Cl solution?

Sometimes when salts dissolve in water, one of the ions can act as a weak acid in solution.
Example: What is the pH of a 0.100M NH_4Cl solution?

$NH_4Cl \longrightarrow NH_4^+ + Cl^-$

0.100M 0.100M $\longleftarrow$ weak acid in water

	NH_4^+ + H_2O $\rightleftharpoons$	H_3O^+ +	NH_3
	0.100M	0M	0M
	- x	+ x	+ x
	0.100 - x	x	x

Let x = $\Delta[H_3O^+]$

assume 0.100 - x $\cong$ 0.100

$K_a = \frac{[H_3O^+]^2}{[H_2S]}$

thus, $5.6 \times 10^{-10} = \frac{x^2}{0.100}$

K_a from table = 5.6×10^{-10}

so, $x = \sqrt{(5.6 \times 10^{-10})(0.100)}$

$x = \underline{7.4}83 \times 10^{-6}$

pH = 5.13

Type 2 Problems: Calculating the Initial Concentration of a Weak Acid

Example: What [HCN] would be required to produce a pH of 4.120?

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	HCN + H_2O $\rightleftharpoons$	CN^- +	H_3O^+
I	x	0M	0M
C	$-\underline{7.58}6 \times 10^{-5}$	$+\underline{7.58}6 \times 10^{-5}$	$+\underline{7.58}6 \times 10^{-5}$
E	$x - \underline{7.58}6 \times 10^{-5}$	$\underline{7.58}6 \times 10^{-5}$	$\underline{7.58}6 \times 10^{-5}$

pH = 4.120

$[H_3O^+]$ = 2ndlog(-4.120)

= $\underline{7.58}6 \times 10^{-5}$M

$4.9 \times 10^{-10} = \frac{(\underline{7.58}6 \times 10^{-5})^2}{x - \underline{7.58}6 \times 10^{-5}}$

$4.9 \times 10^{-10}(x - \underline{7.58}6 \times 10^{-5}) = \underline{5.75}5 \times 10^{-9}$

Let x = $[HCN]_i$

$x - \underline{7.58}6 \times 10^{-5} = \frac{\underline{5.75}5 \times 10^{-9}}{4.9 \times 10^{-10}}$

$K_a = \frac{[H_3O^+]^2}{[HCN]}$

$x = \underline{11}.7 + \underline{7.58}6 \times 10^{-5} = 12$ M

Assignment 1

1) Calculate the pH of a 0.50M solution of H_3BO_3.

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	$H_3BO_3 + H_2O \rightleftharpoons$	H_3O^+	$+ H_2BO_3^-$
I	0.50M	0M	0M
C	- x	+ x	+ x
E	0.50 - x	x	x

Let x = $\Delta[H_3O^+]$

$K_a = \frac{[H_3O^+]^2}{[H_3BO_3]}$

K_a from table = 7.3×10^{-10}

assume $0.50 - x \cong 0.50$

thus, $7.3 \times 10^{-10} = \frac{x^2}{0.50}$

so, $x = \sqrt{(7.3 \times 10^{-10})(0.50)}$

$x = \underline{1.91} \times 10^{-5}$

pH = 4.72

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2) Calculate the pH of a 0.235M solution of NaH_2PO_4.

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$NaH_2PO_4 \longrightarrow Na^+ + H_2PO_4^-$

0.235M 0.235M

	$H_2PO_4^- + H_2O \rightleftharpoons$	H_3O^+	$+ HPO_4^{2-}$
I	0.235M	0M	0M
C	- x	+ x	+ x
E	0.235 - x	x	x

Let x = $\Delta[H_3O^+]$

$K_a = \frac{[H_3O^+]^2}{[H_2PO_4^-]}$

K_a from table = 6.2×10^{-8}

assume $0.235 - x \cong 0.235$

thus, $6.2 \times 10^{-8} = \frac{x^2}{0.235}$

so, $x = \sqrt{(6.2 \times 10^{-8})(0.235)}$

$x = \underline{1.21} \times 10^{-4}M$

pH = 3.92

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3) Hebden p.128 #33 & p.152 #79, 76, 78

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3) Hebden p.128 #33 & p.152 #79, 76, 78

answers in the back of Hebden

Type 3 Problems: Finding the K_a of an Unknown Weak Acid

Example: A 0.20M solution of the weak acid, HA, has a pH of 1.32. Calculate the K_a of the weak acid and use this to identify it.

Acid/Base 2 Notes Key

Type 3 Problems: Finding the K_a of an Unknown Weak Acid

Example: A 0.20M solution of the weak acid, HA, has a pH of 1.32. Calculate the K_a of the weak acid and use this to identify it.

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

	HA	H_2O	H_3O^+	A^-
I	0.20M		0	0
C	$-0.0\underline{47}86$		$+0.0\underline{47}86$	$+0.0\underline{47}86$
E	$0.\underline{15}2$		$0.0\underline{47}86$	$0.0\underline{47}86$

pH = 1.32
$[H_3O^+]$ = 2ndlog(-1.32)
= $0.0\underline{47}86$M

$$K_a = \frac{[H_3O^+]^2}{[HA]} = \frac{(0.0\underline{47}86)^2}{(0.\underline{15}2)} = 1.5 \text{ x } 10^{-2}$$ the acid is H_2SO_3

Example: A 2.00M diprotic acid has a pH of 0.50. Calculate the K_a value.

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$$H_2A + H_2O \rightleftharpoons H_3O^+ + HA^-$$

	H_2A	H_2O	H_3O^+	HA^-
I	2.00M		0	0
C	$-0.\underline{31}62$		$+0.\underline{31}62$	$+0.\underline{31}62$
E	$\underline{1.68}38$		$0.\underline{31}62$	$0.\underline{31}62$

pH = 0.50
$[H_3O^+]$ = 2ndlog(-0.50)
= $0.\underline{31}62$M

$$K_a = \frac{[H_3O^+]^2}{[H_2A]} = \frac{(0.\underline{31}62)^2}{(\underline{1.68}38)} = 5.9 \text{ x } 10^{-2}$$

II) Weak Base Equilibrium and K_b

Write a reaction for and find the pH for 0.10M NaOH. Why is it considered a 'strong' base?

Write a reaction for 0.10M NH_3 solution. Why is it considered a 'weak' base? How would you find the pH of a weak base solution?

II) Weak Base Equilibrium and K_b

Write a reaction for and find the pH for 0.10M NaOH. Why is it considered a 'strong' base?

$NaOH \longrightarrow Na^+ + OH$ pOH = -log(0.10) = 1.00, pH = 13.00
0.10M 0.10M

Write a reaction for 0.10M NH_3 solution. Why is it considered a 'weak' base? How would you find the pH of a weak base solution?

$NH_3 + H_2O \quad NH_4^+ + OH^-$
0.10M <<0.10M

NH_3 is a weak base because it only reacts to a small extent in H_2O to create OH^-, thereby forming an equilibrium. Finding the pH of a weak base solution is a similar process to finding the pH for a weak acid solution.

Weak base problems such as the one previous can be solved using the K_b constant, a K_{eq} for weak bases (similar to K_a for weak acids). The larger the K_b, the stronger the base (the more H^+ it will accept). Write a K_b expression:

$$NO_2^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HNO_{2(aq)} + OH^-{}_{(aq)}$$

K_b = but since $[OH^-] = [HNO_2]$... K_b =

Acid/Base 2 Notes Key

Weak base problems such as the one previous can be solved using the K_b constant, a K_{eq} for weak bases (similar to K_a for weak acids). The larger the K_b, the stronger the base (the more H^+ it will accept). Write a K_b expression:

$$NO_2^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HNO_{2(aq)} + OH^-{}_{(aq)}$$

$K_b = \frac{[HNO_2][OH^-]}{[NO_2^-]}$

but since $[OH^-] = [HNO_2]$...

$K_b = \frac{[OH^-]^2}{[NO_2^-]}$

Relationship of K_w, K_a, and K_b for a Conjugate Acid-Base Pair

Write the K_a for HF:

Write the K_b for F^-:

multiply the K_a of HF by the K_b of F^-:

What results?

Conclusion?

Relationship of K_w, K_a, and K_b for a Conjugate Acid-Base Pair

Write the K_a for HF: $\frac{[H_3O^+][F^-]}{[HF]}$

Write the K_b for F^-: $\frac{[OH^-][HF]}{[F^-]}$

multiply the K_a of HF by the K_b of F^-: $\frac{[H_3O^+]\cancel{[F^-]}}{\cancel{[HF]}} \times \frac{[OH^-]\cancel{[HF]}}{\cancel{[F^-]}} = [H_3O^+][OH^-]$

What results? The K_w! $K_w = [H_3O^+][OH^-]$

Conclusion? $K_w = K_a \times K_b$

The acid-base table only lists acid K_a values. Using what you learned above, how would you get the K_b for the corresponding conjugate base?

Example:
Determine K_b for the weak base SO_4^{2-}:

The acid-base table only lists acid K_a values. Using what you learned above, how would you get the K_b for the corresponding conjugate base?

get the K_a for the conjugate acid, then divide K_w by K_a to get the K_b for the corresponding conjugate base

Example:
Determine K_b for the weak base SO_4^{2-}:

conj acid is HSO_4^-

K_a of $HSO_4^- = 1.2 \times 10^{-2}$

K_b of $SO_4^{2-} = \frac{1.0 \times 10^{-14}}{1.2 \times 10^{-2}} = 8.3 \times 10^{-13}$

Example: Determine K_b for HCO_3^-:

Practice Questions: Determine the K_b values for the following:

a) HPO_4^{2-}

b) $H_2PO_4^-$

Acid/Base 2 Notes Key

Example: Determine K_b for HCO_3^-:

conj acid is HCO_3^-
K_a of HCO_3^- = 4.3 x 10^{-7}

K_b of $CO_3^{2-} = \frac{1.0 \times 10^{-14}}{4.3 \times 10^{-7}} = 2.3 \times 10^{-8}$

Practice Questions: Determine the K_b values for the following:

a) HPO_4^{2-}

conj acid is $H_2PO_4^-$
K_a of $H_2PO_4^-$ = 6.2 x 10^{-8}

K_b of $HPO_4^- = \frac{1.0 \times 10^{-14}}{6.2 \times 10^{-8}}$

$= 1.6 \times 10^{-7}$

b) $H_2PO_4^-$

conj acid is H_3PO_4
K_a of H_3PO_4 = 7.5 x 10^{-3}

K_b of $H_2PO_4^- = \frac{1.0 \times 10^{-14}}{7.5 \times 10^{-3}}$

$= 1.3 \times 10^{-12}$

Assignment 2

Hebden p.152 #77, 80, 82 & p.130 #35bce, 36

Assignment 2

Hebden p.152 #77, 80, 82 & p.130 #35bce, 36

answers in the back of Hebden

Weak base quantitative problems can be broken into three types.

Type 1 Problems: Finding the pH of a weak base solution

Example: Calculate the [OH⁻] and pH for a 0.25M solution of the weak base HS

Example: Calculate the [OH⁻] and pH for a 0.25M solution of the weak base HS

$$HS^- + H_2O \rightleftharpoons OH + H_2S$$

	HS^-	H_2O	OH	H_2S
I	0.25M		0M	0M
C	- x		+ x	+ x
E	0.25 - x		x	x

Let x = Δ[OH⁻]

$K_b = \frac{1.0 \times 10^{-14}}{9.1 \times 10^{-8}} = 1.1 \times 10^{-7}$

$K_b = \frac{[OH^-]^2}{[HS^-]}$

assume 0.25 - x $\cong$ 0.25

thus, $1.1 \times 10^{-7} = \frac{x^2}{0.25}$

so, $x = \sqrt{(1.1 \times 10^{-7})(0.25)}$

x = $\underline{1.6}6 \times 10^{-4}$, [OH⁻] = 1.7 x 10^{-4}M

pOH = -log($\underline{1.6}6 \times 10^{-4}$) = 3.78

pH = 10.22

Acid/Base 2 Notes Key

Sometimes, when salts dissolve in water, one of the ions can act as a weak base in solution, such as the salt in the next example.

Example: Calculate the pH of a 0.100M solution of $K_2C_2O_4$.

Example: Calculate the pH of a 0.100M solution of $K_2C_2O_4$. $K_2C_2O_4 \longrightarrow 2K^+ + C_2O_4^{2-}$

0.100M 0.100M

	$C_2O_4^{2-}$ + H_2O	$\rightleftharpoons$	OH	+ $HC_2O_4^-$
I	0.100M		0M	0M
C	- x		+ x	+ x
E	0.100 - x		x	x

Let x = $\Delta[OH^-]$

$K_b = \frac{1.0 \times 10^{-14}}{6.4 \times 10^{-5}} = \underline{1.5}6 \times 10^{-10}$

$K_b = \frac{[OH^-]^2}{[C_2O_4^{2-}]}$

assume $0.100 - x \approx 0.100$

thus, $\underline{1.5}6 \times 10^{-10} = \frac{x^2}{0.100}$

so, $x = \sqrt{(\underline{1.5}6 \times 10^{-10})(0.100)}$

$x = \underline{3.9}5 \times 10^{-6}$

pOH = $-\log(\underline{3.9}5 \times 10^{-6}) = 5.40$

pH = 8.60

Type 2 Problems: Calculating the Initial Concentration of a Weak Base

Example: A solution of NO_2 has a pH of 8.900. Calculate the $[NO_2]$ that would have been required to make this solution.

Example: A solution of NO_2 has a pH of 8.900. Calculate the $[NO_2]$ that would have been required to make this solution.

	NO_2^- + H_2O	$\rightleftharpoons$	OH	+ HNO_2
I	x		0M	0M
C	$-\underline{7.94}3 \times 10^{-6}$		$+\underline{7.94}3 \times 10^{-6}$	$+\underline{7.94}3 \times 10^{-6}$
E	$x - \underline{7.94}3 \times 10^{-6}$		$\underline{7.94}3 \times 10^{-6}$	$\underline{7.94}3 \times 10^{-6}$

pH = 8.900, pOH = 5.100

$[OH^-]$ = 2ndlog(-5.100)

$= \underline{7.94}3 \times 10^{-6}$M

Let $x = [NO_2^-]_i$

$K_b = \frac{1.0 \times 10^{-14}}{4.6 \times 10^{-4}} = \underline{2.1}7 \times 10^{-11}$

$K_b = \frac{[OH^-]^2}{[NO_2^-]}$

$\underline{2.1}7 \times 10^{-11} = \frac{(\underline{7.94}3 \times 10^{-6})^2}{x - \underline{7.94}3 \times 10^{-6}}$

$\underline{2.1}7 \times 10^{-11}x - \underline{1.7}27 \times 10^{-16} = \underline{6.30}9 \times 10^{-11}$

x = 2.9, $[NO_2^-]_i$ = 2.9M

Acid/Base 2 Notes Key

Example: A solution of ammonia, NH_3, has a pH of 10.50. Calculate the $[NH_3]$ used to make the solution.

Example: A solution of ammonia, NH_3, has a pH of 10.50. Calculate the $[NH_3]$ used to make the solution.

	NH_3	$+ H_2O \rightleftharpoons$	OH	$+$	NH_4^+
I	x		0M		0M
C	$-\underline{3.1}62 \times 10^{-4}$		$+\underline{3.1}62 \times 10^{-4}$		$+\underline{3.1}62 \times 10^{-4}$
E	$x - \underline{3.1}62 \times 10^{-4}$		$\underline{3.1}62 \times 10^{-4}$		$\underline{3.1}62 \times 10^{-4}$

pH = 10.50, pOH = 3.50

$[OH^-]$ = 2ndlog(-3.50)

$= \underline{3.1}62 \times 10^{-4}M$

$\underline{1.7}86 \times 10^{-5} = \dfrac{(\underline{3.1}62 \times 10^{-4})^2}{x - \underline{3.1}62 \times 10^{-}}$

Let $x = [NH_3]_i$

$x = 5.9 \times 10^{-3}$, $[NH_3]_i = 5.9 \times 10^{-3}M$

$K_b = \dfrac{1.0 \times 10^{-14}}{5.6 \times 10^{-10}} = \underline{1.7}86 \times 10^{-5}$

$K_b = \dfrac{[OH^-]^2}{[NH_3]}$

Type 3 Problems: Finding the K_b of an Unknown Weak Base

Example: A 0.44M solution of the weak base B^- has a pH of 11.12. Calculate the K_b for this base, and the K_a for the conjugate acid, HB at 25°C.

Example: A 0.44M solution of the weak base B^- has a pH of 11.12. Calculate the K_b for this base, and the K_a for the conjugate acid, HB at 25°C.

	B^-	$+ H_2O \rightleftharpoons$	OH^{--}	$+$	HB
I	0.44M		0		0
C	$-0.00\underline{13}2$		$+0.00\underline{13}2$		$+0.00\underline{13}2$
E	$0.\underline{43}87$		0.00132		0.00132

pH = 11.12, pOH = 2.88

$[OH^-]$ = 2ndlog(-2.88)

$= \underline{1.3}2 \times 10^{-3}M$

$K_b = \dfrac{[OH^-]^2}{[B^-]} = \dfrac{(0.00132)^2}{(0.\underline{43}87)} = \underline{3.9}7 \times 10^{-6} = 4.0 \times 10^{-6}$

K_a of HB $= \dfrac{1.0 \times 10^{-14}}{\underline{3.9}7 \times 10^{-6}} = 2.5 \times 10^{-9}$

Assignment 3

Hebden p.153 #84-89

Assignment 3

Hebden p.153 #84-89

answers in the back of Hebden

III) Writing Formula (Molecular), Complete Ionic, and Net Ionic Equations for Acid/Base Reactions

1. Strong Acid / Strong Base (Neutralization)

F: $HCl_{(aq)} + NaOH_{(aq)} \rightarrow$ _______ + _______

C:

N:

III) Writing Formula (Molecular), Complete Ionic, and Net Ionic Equations for Acid/Base Reactions

1. Strong Acid / Strong Base (Neutralization)

F: $HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

C: $H^+_{(aq)} + Cl^-_{(aq)} + Na^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)} + Na^+_{(aq)} + Cl^-_{(aq)}$

N: $H^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)}$

Since 100% of strong acids and bases dissociate, they should be written as ions in the complete ionic and net ionic equations.

http://preparatorychemistry.com/Bishop_Solubility_frames.htm

If the resulting salt is low solubility and precipitates, it is included in the net ionic equation.

2) Weak Acid / Strong Base:

F: $HF_{(aq)} + KOH_{(aq)} \longrightarrow$ _______ + _______

C:

N:

Since less than 5% of weak acids and bases dissociate, don't split them into ions for the complete ionic and net ionic equations, since the majority of weak acid molecules stay intact.

2) Weak Acid / Strong Base:

F: $HF_{(aq)} + KOH_{(aq)} \longrightarrow H_2O_{(l)} + KF_{(aq)}$

C: $HF_{(aq)} + K^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)} + K^+_{(aq)} + F^-_{(aq)}$

N: $HF_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)} + F^-_{(aq)}$

Since less than 5% of weak acids and bases dissociate, don't split them into ions for the complete ionic and net ionic equations, since the majority of weak acid molecules stay intact.

Acid/Base 2 Notes Key

3) Strong Acid / Weak Base

F: $HCl_{(aq)} + NaCN_{(aq)} \longrightarrow$ ________ + ________

C:

N:

Many weak bases originate as salts since weak bases often have a negative charge. The salt will dissociate 100% into ions, and the weak base component will then react 100% due to the strong acid present.

3) Strong Acid / Weak Base

F: $HCl_{(aq)} + NaCN_{(aq)} \longrightarrow NaCl_{(aq)} + HCN_{(aq)}$

C: $H^{+}_{(aq)} + Cl^{-}_{(aq)} + Na^{+}_{(aq)} + CN^{-}_{(aq)} \longrightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)} + HCN_{(aq)}$

N: $H^{+}_{(aq)} + CN^{-}_{(aq)} \longrightarrow HCN_{(aq)}$

Many weak bases originate as salts since weak bases often have a negative charge. The salt will dissociate 100% into ions, and the weak base component will then react 100% due to the strong acid present.

Here is an example when the weak base does not originate as a salt:

F: $HCl_{(aq)} + NH_{3(aq)} \longrightarrow$

C:

N:

Sometimes when an acid and base react, only a salt is produced as the base does not contain OH , so no water can form.

Here is an example when the weak base does not originate as a salt:

F: $HCl_{(aq)} + NH_{3(aq)} \longrightarrow NH_4Cl_{(aq)}$

C: $H^{+}_{(aq)} + Cl^{-}_{(aq)} + NH_{3(aq)} \longrightarrow NH_4{}^{+}_{(aq)} + Cl^{-}_{(aq)}$

N: $H^{+}_{(aq)} + NH_{3(aq)} \longrightarrow NH_4{}^{+}_{(aq)}$

Sometimes when an acid and base react, only a salt is produced as the base does not contain OH , so no water can form.

4) Weak Acid / Weak Base

F: $NH_{3(aq)} + HF_{(aq)}$

C:

N:

4) Weak Acid / Weak Base

F: $NH_{3(aq)} + HF_{(aq)} \rightleftharpoons NH_4F_{(aq)}$

C: $NH_{3(aq)} + HF_{(aq)} \rightleftharpoons NH_4{}^{+}_{(aq)} + F^{-}_{(aq)}$

N: $NH_{3(aq)} + HF_{(aq)} \rightleftharpoons NH_4{}^{+}_{(aq)} + F^{-}_{(aq)}$

Acid/Base 2 Notes Key

Assignment 4: Write Formula (Molecular), Complete Ionic, and Net Ionic Equations for the following Acid/Base reactions:

1) $HClO_{4(aq)} + KOH_{(aq)} \longrightarrow$

Assignment 4: Write Formula (Molecular), Complete Ionic, and Net Ionic Equations for the following Acid/Base reactions:

1) $HClO_{4(aq)} + KOH_{(aq)} \longrightarrow$

F: $HClO_{4(aq)} + KOH_{(aq)} \longrightarrow KClO_{4(aq)} + H_2O_{(l)}$

C: $H^+_{(aq)} + ClO_4^-{}_{(aq)} + K^+_{(aq)} + OH^-_{(aq)} \longrightarrow K^+_{(aq)} + ClO_4^-{}_{(aq)} + H_2O_{(l)}$

N: $H^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)}$

2) $HBr_{(aq)} + NaCH_3COO_{(aq)} \longrightarrow$

2) $HBr_{(aq)} + NaCH_3COO_{(aq)} \longrightarrow$

F: $HBr_{(aq)} + NaCH_3COO_{(aq)} \longrightarrow NaBr_{(aq)} + CH_3COOH_{(aq)}$

C: $H^+_{(aq)} + Br^-_{(aq)} + Na^+_{(aq)} + CH_3COO^-_{(aq)} \longrightarrow Na^+_{(aq)} + Br^-_{(aq)} + CH_3COOH_{(aq)}$

N: $H^+_{(aq)} + CH_3COO^-_{(aq)} \longrightarrow CH_3COOH_{(aq)}$

3) $HCOOH_{(aq)} + LiOH_{(aq)} \longrightarrow$

3) $HCOOH_{(aq)} + LiOH_{(aq)} \longrightarrow$

F: $HCOOH_{(aq)} + LiOH_{(aq)} \longrightarrow H_2O_{(l)} + LiCOOH_{(aq)}$

C: $HCOOH_{(aq)} + Li^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)} + Li^+_{(aq)} + HCOO^-_{(aq)}$

N: $HCOOH_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)} + HCOO^-_{(aq)}$

Acid/Base 2 Notes Key

4) $HI_{(aq)} + NH_{3(aq)} \longrightarrow$

4) $HI_{(aq)} + NH_{3(aq)} \longrightarrow$

F: $HI_{(aq)} + NH_{3(aq)} \longrightarrow NH_4I_{(aq)}$

C: $H^+_{(aq)} + I^-_{(aq)} + NH_{3(aq)} \longrightarrow NH_4^+{}_{(aq)} + I^-_{(aq)}$

N: $H^+_{(aq)} + NH_{3(aq)} \longrightarrow NH_4^+{}_{(aq)}$

5) $Sr(OH)_{2(aq)} + HNO_{3(aq)} \longrightarrow$

5) $Sr(OH)_{2(aq)} + HNO_{3(aq)} \longrightarrow$

F: $Sr(OH)_{2(aq)} + 2HNO_{3(aq)} \rightarrow 2H_2O_{(l)} + Sr(NO_3)_{2(aq)}$

C: $Sr^{2+}_{(aq)} + 2OH^-_{(aq)} + 2H^+_{(aq)} + 2NO_3^-{}_{(aq)} \longrightarrow H_2O_{(l)} + Sr^{2+}_{(aq)} + 2NO_3^-{}_{(aq)}$

N: $H^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)}$

6) $NaCN_{(aq)} + NH_4Cl_{(aq)} \longrightarrow$

6) $NaCN_{(aq)} + NH_4Cl_{(aq)} \rightleftharpoons$

F: $NaCN_{(aq)} + NH_4Cl_{(aq)} \rightleftharpoons NaCl_{(aq)} + NH_{3(aq)} + HCN_{(aq)}$

C: $Na^+_{(aq)} + CN^-_{(aq)} + NH_4^+{}_{(aq)} + Cl^-_{(aq)} \rightleftharpoons Na^+_{(aq)} + Cl^-_{(aq)} + NH_{3(aq)} + HCN_{(aq)}$

N: $NH_4^+{}_{(aq)} + CN^-_{(aq)} \rightleftharpoons HCN_{(aq)} + NH_{3(aq)}$

Acid/Base 2 Notes Key

IV) Hydrolysis

When any acid (strong or weak) reacts with a strong hydroxide base, the reaction is 100% (due to the strong base):

Examples

$HBr_{(aq)} + Ca(OH)_{2(aq)} \longrightarrow$

$CH_3COOH_{(aq)} + KOH_{(aq)} \longrightarrow$

In general, the products are __________ and ________. These reactions are called _____________________ reactions.

IV) Hydrolysis

When any acid (strong or weak) reacts with a strong hydroxide base, the reaction is 100% (due to the strong base):

Examples

$2HBr_{(aq)} + Ca(OH)_{2(aq)} \longrightarrow CaBr_{2(aq)} + 2H_2O_{(l)}$

$CH_3COOH_{(aq)} + KOH_{(aq)} \longrightarrow H_2O_{(l)} + KCH_3COO_{(aq)}$

In general, the products are __water_____ and __salt_____. These reactions are called ______neutralization________ reactions.

When a strong acid reacts with a base that does not contain hydroxide, it is still a neutralization reaction, however the only product is a salt. The reaction is 100% because a strong acid is reacting.

Example:

$HCl_{(aq)} + NH_{3(aq)} \longrightarrow$

When a strong acid reacts with a base that does not contain hydroxide, it is still a neutralization reaction, however the only product is a salt. The reaction is 100% because a strong acid is reacting.

Example:

$HCl_{(aq)} + NH_{3(aq)} \longrightarrow NH_4Cl_{(aq)}$

The salts that are produced can be soluble or insoluble (use your solubility table). The insoluble salts will form a solid and precipitate out of solution. The soluble salts will stay in solution as cations and anions and may act as weak acids or bases. For example, the salt produced above was $NH_4Cl_{(aq)}$. Is it soluble (use your table)?

Therefore, how will it actually exist in solution?

Is the cation or anion a weak acid or base?

The salts that are produced can be soluble or insoluble (use your solubility table). The insoluble salts will form a solid and precipitate out of solution. The soluble salts will stay in solution as cations and anions and may act as weak acids or bases. For example, the salt produced above was $NH_4Cl_{(aq)}$. Is it soluble (use your table)? Yes

Therefore, how will it actually exist in solution?

as ions $NH_4^+{}_{(aq)}$ and $Cl^-{}_{(aq)}$

Is the cation or anion a weak acid or base?

NH_4^+ is a weak acid

Acid/Base 2 Notes Key

So what will that ion do in solution?

The reaction you just wrote above is a **hydrolysis** reaction, and it can cause salt solutions to be acidic or basic (acidic in the example above due to H_3O^+ formation).
So, whenever a salt is dissolved in solution OR whenever a salt is formed due to an acid reacting with a base, the resulting salt solution may be acidic or basic if a **hydrolysis** reaction occurs. If not, the solution will be neutral.

So what will that ion do in solution?

react with water and form an equilibrium:

$$NH_4^+{}_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+{}_{(aq)} + NH_{3(aq)}$$

The reaction you just wrote above is a **hydrolysis** reaction, and it can cause salt solutions to be acidic or basic (acidic in the example above due to H_3O^+ formation).
So, whenever a salt is dissolved in solution OR whenever a salt is formed due to an acid reacting with a base, the resulting salt solution may be acidic or basic if a **hydrolysis** reaction occurs. If not, the solution will be neutral.

The ions that make up the salts produced from the neutralization reactions may or may not undergo hydrolysis. Here are the guidelines:

1) Ions that will not undergo hydrolysis are...

2) Ions that will not undergo hydrolysis are...

The ions that make up the salts produced from the neutralization reactions may or may not undergo hydrolysis. Here are the guidelines:

1) Ions that will not undergo hydrolysis are...

conjugate bases of strong acids, because they are not bases at all. These include: ClO_4^-, I^-, Br^-, Cl^-, NO_3^-

HSO_4^- is not a base, but is actually a weak acid so will hydrolyze acidically

2) Ions that will not undergo hydrolysis are...

conjugate acids of strong bases, because they are not acids at all. These include: Li^+, Na^+, K^+, Rb^+, Cs^+, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}
(group 1 & 2 cations)

OH^- & NH_3 though not acids, are both bases

3) Ions that will undergo hydrolysis are...

3) Ions that will undergo hydrolysis are...

weaks acids and bases, found on the middle (unshaded) portion of the acid/base table

Acid/Base 2 Notes Key

Write a dissociation equation for each salt, and then predict whether the resulting salt solutions will be acidic, basic, or neutral, and write any hydrolysis equations as support.

1) $KCl_{(aq)}$

2) $NH\text{-}_4NO_{3(aq)}$

http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/animations/Nh4ClsaltV8.html

Write a dissociation equation for each salt, and then predict whether the resulting salt solutions will be acidic, basic, or neutral, and write any hydrolysis equations as support.

1) $KCl_{(aq)} \longrightarrow K^+_{(aq)} + Cl^-_{(aq)}$ NEUTRAL

doesn't hydrolyze (conj acid of strong base)

doesn't hydrolyze (conj base of strong acid)

2) $NH\text{-}_4NO_{3(aq)} \longrightarrow NH_4^+{}_{(aq)} + NO_3^-{}_{(aq)}$

weak acid, therefore hydrolyzes

doesn't hydrolyze (conj base of str acid)

http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/animations/Nh4ClsaltV8.html

$NH_4^+{}_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+{}_{(aq)} + NH_{3(aq)}$

ACIDIC (due to formation of H_3O^+)

3) $Na_2CO_{3(aq)}$

3) $Na_2CO_{3(aq)} \longrightarrow 2Na^+_{(aq)} + CO_3^{2-}{}_{(aq)}$

doesn't hydrolyze

weak base so it hydrolyzes

$CO_3^{2-}{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HCO_3^-{}_{(aq)} + OH^-_{(aq)}$

BASIC (due to OH^- formation)

Sometimes, an ion is amphiprotic, so will it act more as an acid or a base?

4) $NaHSO_{3(aq)}$

Sometimes, an ion is amphiprotic, so will it act more as an acid or a base?

4) $NaHSO_{3(aq)} \longrightarrow Na^+_{(aq)} + HSO_3{}_{(aq)}$

$HSO_3^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+{}_{(aq)} + SO_3^{2-}{}_{(aq)}$

$K_a = 1.0 \times 10^{-7}$

$HSO_3^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons OH_{(aq)} + H_2SO_{3(aq)}$

$K_b = \frac{1.0 \times 10^{-14}}{1.5 \times 10^{-2}} = 6.7 \times 10^{-13}$

both reactions occur, but since $K_a > K_b$, the acidic hydrolysis (top reaction) occurs to a greater extent, therefore there is more H_3O^+ created compared to OH , therefore an ACIDIC solution

5) $KHCO_{3(aq)}$

---

5) $KHCO_{3(aq)} \longrightarrow K^+_{(aq)} + HCO_{3(aq)}$

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

$K_a = 5.6 \times 10^{-11}$

$HCO_3^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons OH^-_{(aq)} + H_2CO_{3(aq)}$

$$K_b = \frac{1.0 \times 10^{-14}}{4.3 \times 10^{-7}} = 2.3 \times 10^{-8}$$

$K_b > K_a$, therefore BASIC

If an ion is amphiprotic, write an acidic hydrolysis with a K_a value, and a basic hydrolysis with a K_b value. Whichever K value is greater, that reaction will occur to a greater extent.

How do you predict if the solution is acidic or basic when both ions in the salt hydrolyze?

6) $NH_4NO_{2(aq)}$

How do you predict if the solution is acidic or basic when both ions in the salt hydrolyze?

6) $NH_4NO_{2(aq)} \longrightarrow NH_4^+{}_{(aq)} + NO_2^-{}_{(aq)}$

$NH_4^+{}_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + NH_{3(aq)}$

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

$NO_2^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons OH^-_{(aq)} + HNO_{2(aq)}$

$$K_b = \frac{1.0 \times 10^{-14}}{4.6 \times 10^{-4}} = 2.2 \times 10^{-11}$$

$K_a > K_b$, therefore the solution is ACIDIC.

---

7) $Al(NO_2)_{3(aq)}$

*when Al^{3+}, Cr^{3+}, or Fe^{3+} exist in solution, they will gain six water molecules around them and then act as a weak acid (see table).

Acid/Base 2 Notes Key

7) $Al(NO_2)_{3(aq)} \longrightarrow Al^{3+} + 3NO_2^-$

*when Al^{3+}, Cr^{3+}, or Fe^{3+} exist in solution, they will gain six water molecules around them and then act as a weak acid (see table).

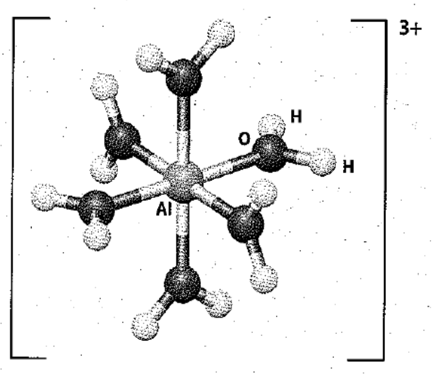
$Al(H_2O)_6^{3+} + H_2O \rightleftharpoons Al(H_2O)_5(OH)^{2+} + H_3O^+$

$K_a = 1.4 \times 10^{-5}$

$NO_2^- + H_2O \rightleftharpoons HNO_2 + OH^-$

$K_b = 2.2 \times 10^{-11}$

$K_a > K_b$, therefore the solution is ACIDIC



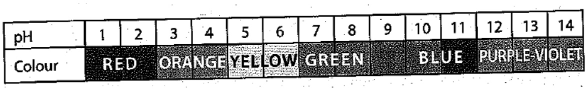
Summarize how you can predict whether a salt solution with two ions that hydrolyze will be acidic, basic, or neutral:

Summarize how you can predict whether a salt solution with two ions that hydrolyze will be acidic, basic, or neutral:

write the hydrolysis equation for each and find the K_a and K_b

- if $K_a > K_b$, solution is acidic
- if $K_b > K_a$, solution is basic
- if $K_a = K_b$, solution is neutral

Universal Indicator



pH	1	2	3	4	5	6	7	8	9	10	11	12	13	14
Colour	RED		ORANGE		YELLOW		GREEN			BLUE		PURPLE-VIOLET		

Assignment 5

1) Hydrolysis Mini-Lab: answer the questions from the lab for #1 of this assignment

Acid/Base 2 Notes Key

2) Write dissociation equations, and any hydrolysis equation(s) occurring when the following salts are added to water and predict whether the resulting solution will be acidic, basic, or neutral.

a) Na_2HPO_4

2) Write dissociation equations, and any hydrolysis equation(s) occurring when the following salts are added to water and predict whether the resulting solution will be acidic, basic, or neutral.

a) $Na_2HPO_4 \longrightarrow 2Na^+ + HPO_4^{2-}$

$HPO_4^{2-} + H_2O \rightleftharpoons H_3O^+ + PO_4^{3-}$

$K_a = 2.2 \times 10^{-13}$

$HPO_4^{2-} + H_2O \rightleftharpoons OH + H_2PO_4^-$

$K_b = \frac{1.0 \times 10^{-14}}{6.2 \times 10^{-8}} = 1.6 \times 10^{-7}$

$K_b > K_a$, so the solution is BASIC

b) $Cr_2(SO_4)_3$

b) $Cr_2(SO_4)_3 \longrightarrow 2Cr^{3+} + 3SO_4^{2-}$

$Cr(H_2O)_6^{3+} + H_2O \rightleftharpoons H_3O^+ + Cr(H_2O)_5(OH)^{2+}$

$K_a = 1.5 \times 10^{-4}$

$SO_4^{2-} + H_2O \rightleftharpoons HSO_4^- + OH^-$

$K_b = \frac{1.0 \times 10^{-14}}{1.2 \times 10^{-2}} = 8.3 \times 10^{-13}$

$K_a > K_b$, therefore the solution is ACIDIC

3) NH_3 is titrated with HI. When the two react in the titration, what salt is formed? Does the salt undergo hydrolysis? If so, what is the hydrolysis equation and will the resulting pH be above or below 7?

3) NH_3 is titrated with HI. When the two react in the titration, what salt is formed? Does the salt undergo hydrolysis? If so, what is the hydrolysis equation and will the resulting pH be above or below 7?

NH_4I
Yes
$NH_4^+ + H_2O \rightleftharpoons H_3O^+ + NH_3$

Below 7 (acidic)

Acid/Base 2 Notes Key

4) In a titration, which of the following combinations would result in an equivalence point with pH greater than 7.0? *HINT: find the resulting salt from each reaction and see if and how it undergoes hydrolysis

A. HCl and NaOH
B. HNO_3 and NH_3
C. HBr and $NaCH_3COO$
D. CH_3COOH and NaOH

4) In a titration, which of the following combinations would result in an equivalence point with pH greater than 7.0? *HINT: find the resulting salt from each reaction and see if and how it undergoes hydrolysis

A. HCl and NaOH
B. HNO_3 and NH_3
C. HBr and $NaCH_3COO$
[✓ D.] CH_3COOH and NaOH

5) Calculate the pH of a 0.20M KCN solution (type 1 K_b problem after considering hydrolysis)

5) Calculate the pH of a 0.20M KCN solution (type 1 K_b problem after considering hydrolysis)

$$\underset{0.20M}{KCN} \longrightarrow K^+ + \underset{0.20M}{CN^-} \longleftarrow \text{undergoes hydrolysis}$$

	CN^-	+ H_2O $\rightleftharpoons$	OH	+ HCN
I	0.20M		0M	0M
C	- x		+ x	+ x
E	0.20 - x		x	x

Let x = $\Delta[OH^-]$

$K_b = \dfrac{1.0 \times 10^{-14}}{4.9 \times 10^{-10}} = \underline{2.0}4 \times 10^{-5}$

$K_b = \dfrac{[OH^-]^2}{[CN^-]}$

assume $0.20 - x \overset{\sim}{=} 0.20$

thus, $\underline{2.0}4 \times 10^{-5} = \dfrac{x^2}{0.20}$

so, $x = \sqrt{(\underline{2.0}4 \times 10^{-5})(0.20)}$

$x = \underline{2.0}2 \times 10^{-3}$

$pOH = -\log(\underline{2.0}2 \times 10^{-3}) = 2.69$

pH = 11.31

V) Indicators

Indicators are used to signal the equivalence point (when ______________________________) of an acid-base titration using a colour change. An indicator is a solution of a weak organic acid (an acid that contains ____________), **HIn**, and its conjugate base, **In**$^-$, at equilibrium. The acid form of the indicator, HIn, is a different colour than the conjugate base form, In . The following is the general equilibrium for any acid-base indicator:

$$HIn + H_2O \rightleftharpoons In^- + H_3O^+$$

V) Indicators

Indicators are used to signal the equivalence point (when moles of H_3O^+ = moles of OH^-) of an acid-base titration using a colour change. An indicator is a solution of a weak organic acid (an acid that contains carbon), **HIn**, and its conjugate base, **In**$^-$, at equilibrium. The acid form of the indicator, HIn, is a different colour than the conjugate base form, In . The following is the general equilibrium for any acid-base indicator:

$$HIn + H_2O \rightleftharpoons In^- + H_3O^+$$

Acid/Base 2 Notes Key

Let's look at how an indicator equilibrium works in solution using the indicator bromthymol blue:

$$HIn + H_2O \rightleftharpoons In^- + H_3O^+$$

for bromthymol blue yellow blue

If $[HIn] > [In^-]$, the system favours the ________ side and the solution will be a __________ colour.
If $[In^-] > [HIn]$, the system favours the __________ side and the solution will be a __________ colour.

Let's look at how an indicator equilibrium works in solution using the indicator bromthymol blue:

$$HIn + H_2O \rightleftharpoons In^- + H_3O^+$$

for bromthymol blue yellow blue

If $[HIn] > [In^-]$, the system favours the ___left___ side and the solution will be a ___yellow___ colour.
If $[In^-] > [HIn]$, the system favours the ___right___ side and the solution will be a ___blue___ colour.

What happens to the equilibrium if bromthymol blue is put into an acidic solution, and what is the resulting solution colour?

$$HIn + H_2O \rightleftharpoons In^- + H_3O^+$$

yellow blue

What happens to the equilibrium if bromthymol blue is put into an acidic solution, and what is the resulting solution colour?

$$HIn + H_2O \rightleftharpoons In^- + H_3O^+$$

yellow blue

in acid, $[H_3O^+]$ is high, therefore a shift left occurs, causing $[HIn]$ to be higher than $[In^-]$, therefore the solution is yellow

What will happen if bromthymol blue is put into basic solution, and what is the resulting solution colour?

$$HIn + H_2O \rightleftharpoons In^- + H_3O^+$$

yellow blue

What will happen if bromthymol blue is put into basic solution, and what is the resulting solution colour?

$$HIn + H_2O \rightleftharpoons In^- + H_3O^+$$

yellow blue

in basic solution, [OH] is high, so $[H_3O^+]$ is low, causing a shift right, so $[In^-]$ is high and $[HIn]$ is low, therefore the solution is blue in colour

Acid/Base 2 Notes Key

During a titration, pH is constantly changing as base is being added to acid (or *visa versa*). If an indicator such as bromthymol blue is present, it will eventually undergo a colour change due to the continual change in $[H_3O^+]$ and resulting shift of the indicator equilibrium.
If there is acid in a flask with some bromthymol blue, what colour will it be?
If base is continually added from the buret, what shift results in the equilibrium?
How does this affect [HIn] and [In]?

During a titration, pH is constantly changing as base is being added to acid (or *visa versa*). If an indicator such as bromthymol blue is present, it will eventually undergo a colour change due to the continual change in $[H_3O^+]$ and resulting shift of the indicator equilibrium.
If there is acid in a flask with some bromthymol blue, what colour will it be? yellow
If base is continually added from the buret, what shift results in the equilibrium? right
How does this affect [HIn] and [In]?

$[In^-]$ increases and [HIn] decreases

What result will this have on the colour of the solution?

$$\underset{\text{yellow}}{HIn} + H_2O \rightleftharpoons \underset{\text{blue}}{In^-} + H_3O^+$$

What result will this have on the colour of the solution?

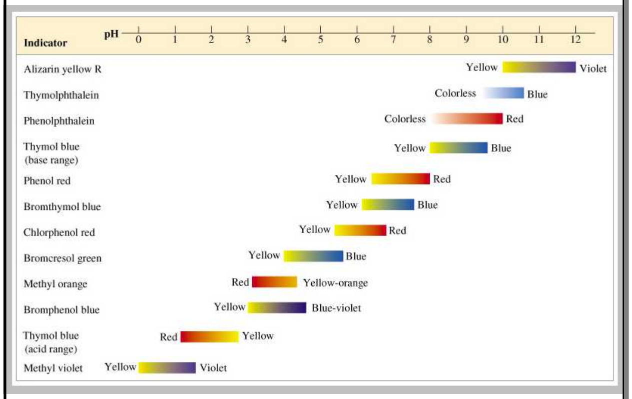
$$\underset{\text{yellow}}{HIn} + H_2O \rightleftharpoons \underset{\text{blue}}{In^-} + H_3O^+$$

At first, [HIn] > $[In^-]$, thus the solution is yellow. As $[In^-]$ increases and [HIn] decreases due to addition of base, eventually [HIn] = [In], and the solution is green (the transition point - a mixture of yellow and blue). Then, as more base is added, $[In^-]$ > [HIn], so the solution turns blue.

The point at which the colour is an equal mixture of the [HIn] colour and the [In] colour is called the **transition point** for the indicator. Another name is the **endpoint**, as this is when a titration would come to an end as the endpoint signals that the equivalence point has been reached.

The endpoint occurs at different pHs for different indicators, as each indicator has its own unique equilibrium. The acid-base indicator table in the data booklet shows different indicators and the pH range of their colour changes. Most indicators change colour over a range of about 2 pH units. For example, bromthymol blue is yellow at pH 6.0 and below and blue at pH 7.6 and above. From 6.0 to 6.8, it's yellow-green, at 6.8 it's perfect green, and from 6.8-7.6 it's blue-green.

Acid/Base 2 Notes Key



It is very important to be able to distinguish between the two terms **equivalence point** and **endpoint**. The equivalence point is the point in the titration where moles of H_3O^+ = moles of OH^-. The endpoint is the point in the titration where the colour of the indicator changes. If the indicator is chosen correctly, it will change the colour of the solution at or very near the equivalence point.

Practice Questions:

1) Which of the following indicators is red at pH 13?

A. Orange IV
B. Alizarin Yellow
C. Indigo Carmine
D. Thymol Blue

Practice Questions:

1) Which of the following indicators is red at pH 13?

A. Orange IV
[✓ B.] Alizarin Yellow
C. Indigo Carmine
D. Thymol Blue

use your indicator table in your data booklet

2) What colour is a 1 x 10^{-3}M NaOH solution containing the indicator Neutral Red?

2) What colour is a 1 x 10^{-3}M NaOH solution containing the indicator Neutral Red?

$[OH^-]$ = 1 x 10^{-3}, so pOH = 3.0, so pH = 11.0

Therefore, the colour is AMBER.

Acid/Base 2 Notes Key

Recall that the general equilibrium equation for an indicator is a follows:

$HIn_{(aq)} + H_2O_{(l)} \rightleftharpoons In^-_{(aq)} + H_3O^+_{(aq)}$

Write the K_a equation for the above:

K_a =

At the endpoint, what is true about [HIn] and $[In^-]$?

Therefore, what will the K_a reduce to?

K_a =

Recall that the general equilibrium equation for an indicator is a follows:

$HIn_{(aq)} + H_2O_{(l)} \rightleftharpoons In^-_{(aq)} + H_3O^+_{(aq)}$

Write the K_a equation for the above:

$$K_a = \frac{[In^-][H_3O^+]}{[HIn]}$$

At the endpoint, what is true about [HIn] and $[In^-]$? at the endpt, [HIn] = $[In^-]$

Therefore, what will the K_a reduce to?

$$K_a = \frac{\cancel{[In^-]}[H_3O^+]}{\cancel{[HIn]}}$$ so $K_a = [H_3O^+]$ at the endpt.

So, **at the endpoint** (point of colour change), the $[H_3O^+]$ equals the value of the K_a for the indicator.

It is easy to find the K_a of each indicator (remember, indicators are weak organic acids) using the indicator data table and some simple calculations.

i) Find the pH of the endpoint of the indicator using the table

ii) Use the endpoint pH to find the $[H_3O^+]$ at this point (2ndlog(-pH))

iii) At the endpoint, the $[H_3O^+]$ is equal to the K_a

Example: Find the K_a of Orange IV

Example: Find the K_a of Orange IV

Orange IV: 1.4 - 2.8
so endpt pH is 2.1

$[H_3O^+]$ = 2ndlog(-2.1)
= $\underline{7}.943 \times 10^{-3}$M

$K_a = 8 \times 10^{3}$ for Orange IV

Assignment 6

1) Which of the following chemical indicators has a $K_a = 2.5 \times 10^{5}$?

A. methyl orange
B. phenolphthalein
C. thymolphthalein
D. bromcresol green

2) Find the K_a of Alizarin Yellow.

Acid/Base 2 Notes Key

Assignment 6

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A. methyl orange
B. phenolphthalein
C. thymolphthalein
[✓ D.] bromcresol green

2) Find the K_a of Alizarin Yellow.

Alizarin Yellow: 10.1 - 12.0
endpt pH = $11.\underline{0}5$

$[H_3O^+] = K_a$ = 2ndlog(-11.05) = 9×10^{-12}

3) A weak acid is titrated with a strong base using the indicator phenolphthalein to detect the endpoint. What is the approximate pH at the transition point?

A. 7.0
B. 8.0
C. 9.0
D. 10.0

3) A weak acid is titrated with a strong base using the indicator phenolphthalein to detect the endpoint. What is the approximate pH at the transition point?

A. 7.0
B. 8.0
[✓ C.] 9.0
D. 10.0

4) Read Hebden p.161 (bottom) & 162 on Universal Indicators. Do Hebden p.162 #108-112 and p.163 #116-119

5) The indicator thymol blue has two colour changes (see the acid-base indicator table). How is this possible?

4) Read Hebden p.161 (bottom) & 162 on Universal Indicators. Do Hebden p.162 #108-112 and p.163 #116-119

answers in the back of Hebden

5) The indicator thymol blue has two colour changes (see the acid-base indicator table). How is this possible?

It is a diprotic weak acid:

$$\underset{\text{red}}{H_2In} + H_2O \rightleftharpoons \underset{\text{yellow}}{HIn^-} + H_3O^+$$

$$\underset{\text{yellow}}{HIn^-} + H_2O \rightleftharpoons \underset{\text{blue}}{In^-} + H_3O^+$$

VI) Buffers

What is a buffer?

What composes a buffer?

Acid/Base 2 Notes Key

VI) Buffers

What is a buffer?

A weak acid / weak conjugate base equilibrium solution that keeps pH steady even when acid or base is added.

What composes a buffer?

Large, equal concentrations of a weak acid and its conjugate base.

How would you make an HSO_3^- / SO_3^{2-} 2.0M buffer solution?

http://www.chembio.uoguelph.ca/educmat/chm19104/chemtoons/chemtoons5.htm

Would you have a buffer if you simply add 2.0M $NaHSO_3$ to water? Why or why not?

How would you make an HSO_3^- / SO_3^{2-} 2.0M buffer solution?

1) Add 2.0M $NaHSO_3$ to water to create an equilibrium.

$NaHSO_3 \qquad Na^+ + HSO_3$ then $\underset{2.0M}{HSO_3} + H_2O \rightleftharpoons \underset{\text{v. small}}{SO_3^{2-}} + \underset{\text{v. small}}{H_3O^+}$

2) Add Na_2SO_3 to increase $[SO_3^{2-}]$ to 2.0M

$\underset{2.0M}{HSO_3^-} + H_2O \rightleftharpoons \underset{2.0M}{SO_3^{2-}} + \underset{\text{v. small}}{H_3O^+}$

http://www.chembio.uoguelph.ca/educmat/chm19104/chemtoons/chemtoons5.htm

Would you have a buffer if you simply add 2.0M $NaHSO_3$ to water? Why or why not?

No, because the conjugate base SO_3^{2-} would have a very small molarity, not even close to 2.0M (like step 1 above).

Let`s suppose you`ve made a HSO_3 / SO_3^{2-} 2.0M buffer: $\underset{2.0M}{HSO_3^-} + H_2O \rightleftharpoons \underset{2.0M}{SO_3^{2-}} + \underset{\text{v. small}}{H_3O^+}$

How does a buffer work?

What determines pH in any solution?

Therefore, if $[H_3O^+]$ and [OH] can somehow be kept constant in the buffer solution, the pH will not drastically change.

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How does a buffer work?

What determines pH in any solution?

$[H_3O^+]$ and/or $[OH^-]$

Therefore, if $[H_3O^+]$ and [OH] can somehow be kept constant in the buffer solution, the pH will not drastically change.

What would happen if a small amount of HCl was added to the buffer solution?

$\underset{2.0M}{HSO_3^-} + H_2O \rightleftharpoons \underset{2.0M}{SO_3^{2-}} + \underset{\text{v. small}}{H_3O^+}$

http://www.chembio.uoguelph.ca/educmat/chm19104/chemtoons/chemtoons8.htm

Acid/Base 2 Notes Key

What would happen if a small amount of HCl was added to the buffer solution?

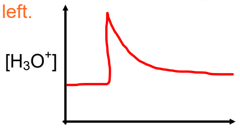
$$HSO_3^- + H_2O \rightleftharpoons SO_3^{2-} + H_3O^+$$

2.0M 2.0M v. small

$HCl + H_2O \longrightarrow H_3O^+ + Cl^-$ adding HCl is adding H_3O^+

$[H_3O^+]$ will initially increase, then a shift left will occur and $[H_3O^+]$ will decrease. Thus, $[HSO_3^-]$ will increase, $[SO_3^{2-}]$ will decrease, and $[H_3O^+]$ will initially increase, then decrease due to the shift left, ut overall will **slightly** increase. This cause pH to only **slightly** decrease.

The large $[SO_3^{2-}]$ allow for much H_3O^+ to be added with only a slight increase in $[H_3O^+]$ due to the shift left.



http://www.chembio.uoguelph.ca/educmat/chm19104/chemtoons/chemtoons8.htm

You could keep adding HCl (which immediately becomes H_3O^+) and as long as there is sufficient SO_3^{2-} available in the solution, the H_3O^+ that forms due to HCl addition will react with SO_3^{2-} to form H_2O and HSO_3 (a shift left in the equilibrium). Since most of the H_3O^+ reacts and is no longer present, the pH will not drastically change.

If HCl continues to be added, eventually SO_3^{2-} will be depleted enough so that H_3O^+ will no longer have any base to react with. What happens in this situation?

If HCl continues to be added, eventually SO_3^{2-} will be depleted enough so that H_3O^+ will no longer have any base to react with. What happens in this situation?

The buffer equilibrium will 'break down' if SO_3^{2-} is used up, as there is no longer any base present to reacth with H_3O^+, which causes the pH to drastically decrease.

If OH is added to our buffer solution, a small amount of it will react with the small amount of H_3O^+ present, but he majority will react with the weak acid HSO_3^- in the following reaction:

$$HSO_3 + OH \longrightarrow$$

This is just like the original buffer equilibrium shifting to the right since HSO_3^- is turning into SO_3^{2-}:

$$HSO_3^- + H_2O \rightleftharpoons SO_3^{2-} + H_3O^+$$

2.0M 2.0M v. small

http://www.chembio.uoguelph.ca/educmat/chm19104/chemtoons/chemtoons7.htm

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$$HSO_3 + OH \longrightarrow H_2O + SO_3^{2-}$$

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$$\longrightarrow$$

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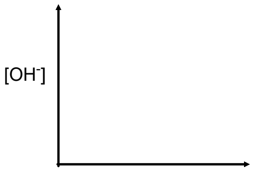
2.0M 2.0M v. small

http://www.chembio.uoguelph.ca/educmat/chm19104/chemtoons/chemtoons7.htm

Acid/Base 2 Notes Key

The critical idea here is that almost all of the OH being added reacts with HSO_3^- to make water and SO_3^{2-}, thereby 'getting rid' of the OH , so the pH of the solution will not rise drastically. It will however, rise slightly. Show why graphically:



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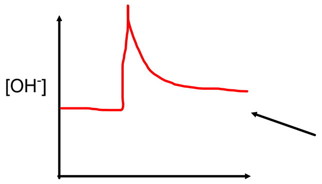


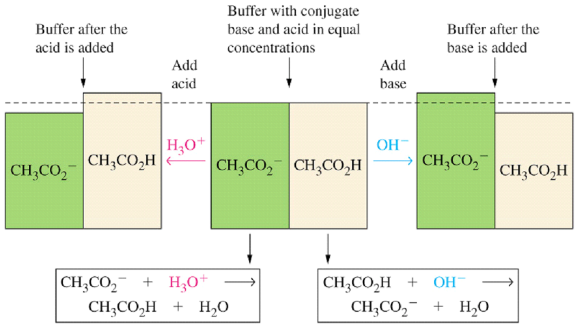
a slight increase in $[OH^-]$ results in a slight increase in pH

Not quite all of the OH^- added will react, thus the pH will rise slightly. If too much OH^- is added, all of the HSO_3 will eventually be used up and the buffer will collapse.

The key to a functional buffer is the large, equal concentrations of a weak acid and its conjugate base. The weak conjugate base (in our example, SO_3^{2-}) is present in large amounts to react with and deplete any H_3O^+ added to the solution. The weak acid (in our example HSO_3) is present in large amounts to react with and deplete any OH added to the buffer. Thus, the pH of the solution cannot drastically change as it is solely dependent on $[H_3O^+]$ and $[OH^-]$.

http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/buffer12.swf

Summary of how a buffer operates:



http://www.wwnorton.com/college/chemistry/gilbert2/tutorials/interface.asp?chapter=chapter_16&folder=buffers

Diluting a Buffer

Does diluting a buffer with water affect its performance?
Take the example used previously:

$$HSO_3^- + H_2O \rightleftharpoons SO_3^{2-} + H_3O^+$$

2.0M 2.0M v. small

Adding water to the above equilibrium system will cause each concentration to ______________.
Will it still work as a buffer?

Acid/Base 2 Notes Key

Diluting a Buffer

Does diluting a buffer with water affect its performance?
Take the example used previously:

$$\underset{2.0M}{HSO_3^-} + H_2O \rightleftharpoons \underset{2.0M}{SO_3^{2-}} + \underset{\text{v. small}}{H_3O^+}$$

Adding water to the above equilibrium system will cause each concentration to __decrease__.
Will it still work as a buffer?

Yes, though a due to the lower molarity, an equivalent volume of the diluted buffer would not hold up as long as the original.

Acidic and Basic Buffers

Each weak conjugate acid/base buffer system has a unique pH that it buffers. The pH can be found by using the K_a of the weak acid. For example:

$$\underset{2.0M}{HSO_3^-} + H_2O \rightleftharpoons \underset{2.0M}{SO_3^{2-}} + \underset{\text{v. small}}{H_3O^+}$$

$$K_a = \qquad = \qquad =$$

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$$\underset{2.0M}{HSO_3^-} + H_2O \rightleftharpoons \underset{2.0M}{SO_3^{2-}} + \underset{\text{v. small}}{H_3O^+}$$

$$K_a = \frac{[SO_3^{2-}][H_3O^+]}{[HSO_3^-]} = \frac{(2.0)[H_3O^+]}{(2.0)} = [H_3O^+]$$

Because the concentrations of the weak acid and its conjugate base are equal, for buffers, $K_a = [H_3O^+]$

For the HSO_3 buffer: $K_a = [H_3O^+] = 1.0 \times 10^{-7}$

Therefore, the pH of the HSO_3 / SO_3^{2-} buffer system is $-\log(1.0 \times 10^{-7})$ = __________

Thus, the HSO_3 / SO_3^{2-} buffer system is called a ____________ buffer.

Weak acids that have K_a values greater than 1.0×10^{-7} will have buffering pHs less than 7, thus they are called **acidic buffers**.
Weak acids that have K_a values less than 1.0×10^{-7} will have buffering pHs greater than 7, and are deemed **basic buffers**.

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Therefore, the pH of the HSO_3 / SO_3^{2-} buffer system is $-\log(1.0 \times 10^{-7})$ = __7.00__

Thus, the HSO_3 / SO_3^{2-} buffer system is called a __neutral__ buffer.

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Weak acids that have K_a values less than 1.0×10^{-7} will have buffering pHs greater than 7, and are deemed **basic buffers**.

Example: Find the buffering pH of an acetic acid / acetate buffer.

Example: Find the buffering pH of an HCN / CN buffer.

Example: Find the buffering pH of an acetic acid / acetate buffer.

$K_a = 1.8 \times 10^{5}$
thus, $[H_3O^+] = 1.8 \times 10^{-5}$

buffering pH = $-\log(1.8 \times 10^{-5}) = 4.74$ ACIDIC buffer

Example: Find the buffering pH of an HCN / CN buffer.

$K_a = 4.9 \times 10^{-10}$
thus, $[H_3O^+] = 4.9 \times 10^{-10}$

buffering pH = $-\log(4.9 \times 10^{-10}) = 9.31$ BASIC buffer

Assignment 7

Read Hebden pages 177-181 (stop after first paragraph on p.181)

1) What composes a buffer and how would you make one?

2) Do Hebden p.181-182 #131-140

Assignment 7

Read Hebden pages 177-181 (stop after first paragraph on p.181)

1) What composes a buffer and how would you make one?

-large, equal concentrations of a weak acid and its conjugate base

-Add a large concentration of a weak acid to water. Then add an equal concentration of the conjugate base.

2) Do Hebden p.181-182 #131-140

answers in the back of Hebden

3) Read Hebden p.182-183: *Buffers in Biological Systems*
What is the buffering pH in your blood?
What two buffering systems contribute to this?

3) Read Hebden p.182-183: *Buffers in Biological Systems*
What is the buffering pH in your blood?
What two buffer systems contribute to this?

buffering pH = 7.35

two systems: see Hebden p. 182-183

VII) Acid/Base Titration Curves

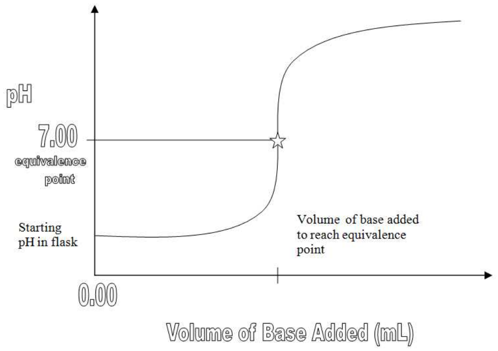
Titrations are commonly carried out to find the concentration of an acidic or basic solution.

http://www.chem-ilp.net/labTechniques/TitrationAnimation.htm

A standard titration curve has an *x* axis that is the *Volume of Base (or Acid) Added* from the buret, and the *y* axis is the *pH* in the flask.

Acid/Base 2 Notes Key

Strong Acid/Strong Base Titration Curve
(strong acid in the flask, strong base in the buret)



Notice the general shape of the titration curve. The pH rises very slowly at the start of the titration, drastically in the middle region, and then very slowly again at the end. Why is this so?
At the start, in order to make the pH change 1 unit (from say 2 to 3), you have to add a large amount of OH from the buret. pH 2 is an $[H_3O^+]$ = _________ and pH 3 is an $[H_3O^+]$ = _________.
Thus, you must add
_________ - _________ = _________ OH ,
quite a large amount.
In the middle region, to change the pH from 6 to 7, you must add
_________ - _________ = _________ OH^-,
one ten-thousandth of the OH needed to change the pH from 2 to 3!

Notice the general shape of the titration curve. The pH rises very slowly at the start of the titration, drastically in the middle region, and then very slowly again at the end. Why is this so?
At the start, in order to make the pH change 1 unit (from say 2 to 3), you have to add a large amount of OH from the buret. pH 2 is an $[H_3O^+]$ =
0.01M and pH 3 is an $[H_3O^+]$ = 0.001M.
Thus, you must add
0.01M - 0.001M = 0.009M OH ,
quite a large amount.
In the middle region, to change the pH from 6 to 7, you must add
0.000001M - 0.0000001M = 0.0000009M OH^-,
one ten-thousandth of the OH needed to change the pH from 2 to 3!

This is why, from about pH 4 to pH 10, you add very little OH and the pH changes so quickly. After pH 10, the same effect takes place as early on.
An analogy using $: If you need to pay a $10 000 loan down to $1 000, it costs $9 000. This may take a while to pay off! But suppose you had to pay a $1 000 loan down to $100, a $100 loan down to $10, a $10 loan down to $1, a $1 loan down to $0.10, and a $0.10 loan down to $0.01 (simulating the middle of the pH curve). This, in total, costs $999.99, a fraction of the first loan!

http://www.mhhe.com/physsci/chemistry/animations/chang_7e_esp/crm3s5_5.swf

Here is an example of a strong acid/strong base titration reaction:

$HCl_{(aq)} + NaOH_{(aq)} \longrightarrow$

What products result?

Water is, of course, neutral. Is the resulting salt neutral? Why or why not?

http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/animations/HClandNaOHtgV8.html

Therefore, what is the pH at the **equivalence point** of a strong acid/strong base titration?

Here is an example of a strong acid/strong base titration reaction:

$HCl_{(aq)} + NaOH_{(aq)} \longrightarrow H_2O_{(l)} + NaCl_{(aq}$

What products result?

Water is, of course, neutral. Is the resulting salt neutral? Why or why not?
Yes, because neither Na^+ or Cl^- hydrolyze.

http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/animations/HClandNaOHtgV8.html

Therefore, what is the pH at the **equivalence point** of a strong acid/strong base titration?
7, as both products are neutral

Acid/Base 2 Notes Key

Strong acid/strong base titrations result in a salt that does not hydrolyze, therefore the **equivalence point** is always 7.

An ideal indicator for a titration is one in which the colour change encompasses the equivalence point. List the ideal indicators for a strong acid/strong base titration.

Strong acid/strong base titrations result in a salt that does not hydrolyze, therefore the **equivalence point** is always 7.

An ideal indicator for a titration is one in which the colour change encompasses the equivalence point. List the ideal indicators for a strong acid/strong base titration.

bromthymol blue
phenol red
neutral red

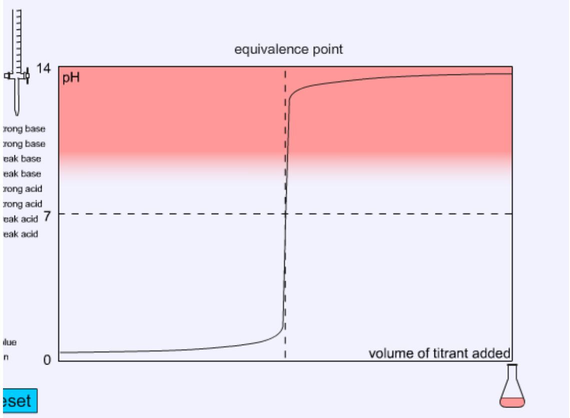
However, in our lab, we used phenolphthalein to indicate the equivalence point of a strong acid/strong base titration, even though the endpoint of phenolphthalein is 9.1. Why is this okay?

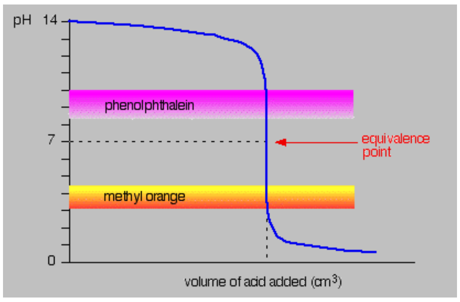
http://www.chem-ilp.net/labTechniques/AcidBaseIdicatorSimulation.htm

However, in our lab, we used phenolphthalein to indicate the equivalence point of a strong acid/strong base titration, even though the endpoint of phenolphthalein is 9.1. Why is this okay?

Remember the sharp vertical rise in the middle region. It only takes one or two drops to the pH to go from about 4 to 10. Thus, the pH will pass through 7 and 9 either in the same drop or within one drop of base, a small error only.

http://www.chem-ilp.net/labTechniques/AcidBaseIdicatorSimulation.htm

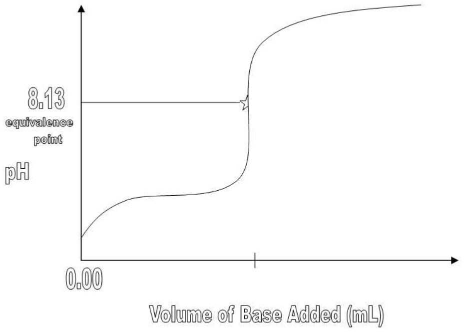




Acid/Base 2 Notes Key

Weak Acid/Strong Base Titration

(strong base is in the buret)



Though two parts of this curve are different than the strong acid/strong base curve, the vertical rise is still present.

What is different compared to the strong acid/strong base titration?

1)

2)

What is different compared to the strong acid/strong base titration?

1) Notice the small pH jump at the beginning of the curve. This is characteristic of any weak/strong titration.

2) The equivalence point for a weak acid/strong base titration is between pH 8 & 9.

Why, for a weak acid/strong base titration, is the equivalence point between pH 8 & 9?

http://www.chembio.uoguelph.ca/educmat/chm19104/chemtoons/chemtoons9.htm

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Why, for a weak acid/strong base titration, is the equivalence point between pH 8 & 9?

The salt produced hydrolyzes basically, producing extra OH^-, thereby increasing the pH.

Example:

$HF_{(aq)} + NaOH_{(aq)} \longrightarrow H_2O_{(l)} + NaF_{(aq)}$

$NaF_{(aq)} \longrightarrow Na^+_{(aq)} + F^-_{(aq)}$

Na^+ doesn't hydrolyze but F^- does:

$F^- + H_2O \rightleftharpoons HF + OH^-$

The extra OH make the equiv pt. pH higher (b/w 8 & 9)

http://www.chembio.uoguelph.ca/educmat/chm19104/chemtoons/chemtoons9.htm

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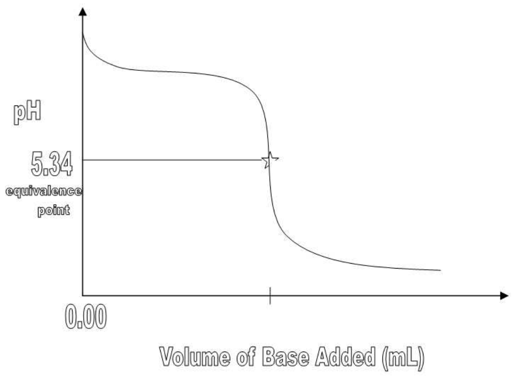
Which indicators would be **ideal** for a weak acid/strong base titration?

Acid/Base 2 Notes Key

Which indicators would be **ideal** for a weak acid/strong base titration?

thymol blue
phenolphthalein

Weak Base/Strong Acid Titration
(strong acid in the buret)



Why does this curve start at a high pH and end at a low pH?

Why does this curve start at a high pH and end at a low pH?

The curve monitors the pH in the flask. If we are starting with weak base in the flask, the pH will be high. When we add acid to the flask from the buret, the pH will decrease.

Characteristics include an initial dip in pH and an equivalence point pH of 5-6. This is because the salt produced will hydrolyze acidically:

Ex: $HCl + NH_3 \longrightarrow NH_4Cl$

$NH_4Cl \longrightarrow NH_4^+ + Cl^-$

$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$

The extra H_3O^+ produced due to the hydrolysis of the salt lowers the equivalence point pH to 5-6.

List the ideal indicators for a weak base/strong acid titration:

http://www.chem-ilp.net/labTechniques/AcidBaseIdicatorSimulation.htm

Acid/Base 2 Notes Key

List the ideal indicators for a weak base/strong acid titration:

methyl red
chlorophenol red

http://www.chem-ilp.net/labTechniques/AcidBaseIdicatorSimulation.htm

Assignment 8

1) Do Hebden p.176 #125

Assignment 8

1) Do Hebden p.176 #125

answers in the back of Hebden

2) A student titrated a 25.00mL sample of 0.20M HX acid with 0.20M NaOH. The following data was collected:

Volume of NaOH added (mL)	pH
0.00	2.72
10.00	4.57
24.90	7.14
24.99	8.14
25.00	8.88
25.01	9.60
26.00	11.59
35.00	12.52

a) What volume of NaOH must be added to reach the endpoint?

b) Is HX weak or strong? How do you know?

c) Select an indicator that would be ideal for this titration and give the colour at the equiv. pt.

2) A student titrated a 25.00mL sample of 0.20M HX acid with 0.20M NaOH. The following data was collected:

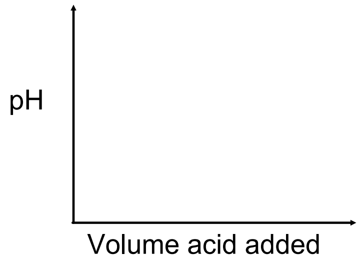
Volume of NaOH added (mL)	pH
0.00	2.72
10.00	4.57
24.90	7.14
24.99	8.14
25.00	8.88
25.01	9.60
26.00	11.59
35.00	12.52

a) What volume of NaOH must be added to reach the endpoint? 25.00mL

b) Is HX weak or strong? How do you know?
weak, as the pH at the equiv pt is 8.88

c) Select an indicator that would be ideal for this titration and give the colour at the equiv. pt.
thymol blue; green colour

3) Draw a typical curve for a weak base / strong acid titration (strong acid in the buret). Show all characteristics and estimate the endpt. pH:



Acid/Base 2 Notes Key

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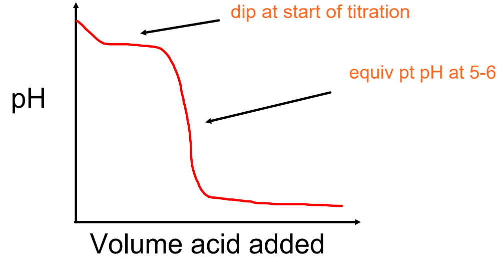

dip at start of titration

equiv pt pH at 5-6

IX) Acidic & Basic Anhydrides

How are acids and bases produced in nature?
Both are produced from **oxides.**
What are oxides?

Acids are produced from **non-metal oxides** and bases are produced from **metal oxides**.

IX) Acidic & Basic Anhydrides

How are acids and bases produced in nature?
Both are produced from **oxides.**
What are oxides?
compounds made from oxygen and one other element

Acids are produced from **non-metal oxides** and bases are produced from **metal oxides**.

When **non-metal oxides** react with water in a synthesis reaction, an acid is formed. Thus, **non-metal oxides** are called **acidic anhydrides** **(__________ ____________ __________).** These reactions often occur in the atmosphere as most non-metal oxides are gases.

Examples:

$SO_{2(g)} + H_2O \longrightarrow$

$SO_{3(g)} + H_2O \longrightarrow$

$CO_{2(g)} + H_2O \longrightarrow$

When **non-metal oxides** react with water in a synthesis reaction, an acid is formed. Thus, **non-metal oxides** are called **acidic anhydrides** (acids without water). These reactions often occur in the atmosphere as most non-metal oxides are gases.

Examples:

$SO_{2(g)} + H_2O \longrightarrow H_2SO_3$

$SO_{3(g)} + H_2O \longrightarrow H_2SO_4$

$CO_{2(g)} + H_2O \longrightarrow H_2CO_3$

When Group 1 and 2 **metal oxides** react with water in a synthesis reaction, a base is formed. **Metal oxides** are called **basic anhydrides**.

$Na_2O_{(s)} + H_2O \longrightarrow$

$MgO_{(s)} + H_2O \longrightarrow$

$CaO_{(s)} + H_2O \longrightarrow$

Acid/Base 2 Notes Key

When Group 1 and 2 **metal oxides** react with water in a synthesis reaction, a base is formed. **Metal oxides** are called **basic anhydrides**.

$$Na_2O_{(s)} + H_2O \longrightarrow 2NaOH$$

$$MgO_{(s)} + H_2O \longrightarrow Mg(OH)_2$$

$$CaO_{(s)} + H_2O \longrightarrow Ca(OH)_2$$

X) Acid Rain

Fuels that contain sulphur are combusted in an industrial setting to form sulphur dioxide gas (________). Some of that sulphur dioxide then reacts with oxygen in the air to produce sulphur trioxide gas (_________).

What do we call these compounds?

What will they do when they go up into the atmosphere?

Thus, what results?

X) Acid Rain

Fuels that contain sulphur are combusted in an industrial setting to form sulphur dioxide gas (__SO_2__). Some of that sulphur dioxide then reacts with oxygen in the air to produce sulphur trioxide gas (__SO_3__).

What do we call these compounds?

acidic anhydrides

What will they do when they go up into the atmosphere? react with water vapour to make acids

$$SO_2 + H_2O \longrightarrow H_2SO_4$$

$$SO_3 + H_2O \longrightarrow H_2SO_3$$

Thus, what results? Acid Rain

When fuel combusts in a car engine, it is so hot that N_2 from the air reacts with O_2 to form nitrogen monoxide gas (__________). Though the catalytic converter reverses this reaction, some nitrogen monoxide escapes through the exhaust and reacts with O_2 in the air to make nitrogen dioxide (_____)

What happens next?

When fuel combusts in a car engine, it is so hot that N_2 from the air reacts with O_2 to form nitrogen monoxide gas (__NO__). Though the catalytic converter reverses this reaction, some nitrogen monoxide escapes through the exhaust and reacts with O_2 in the air to make nitrogen dioxide (_NO_2_)

What happens next?

The gaseous acidic anhydride NO_2 reacts with H_2O in the atmosphere to make acid rain:

$$2NO_2 + H_2O \longrightarrow HNO_2 + HNO_3$$

Acid rain is defined as rainwater that has a pH less than **5.6**.

It is important to note that even 'normal' rainwater is acidic (between pH 5.6 & 7) due to atmospheric CO_2 (an acidic anhydride) dissolving in water to produce carbonic acid:

$$CO_{2(aq)} + H_2O \longrightarrow$$

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Assignment 9

1) There are many environmental problems associated with acid rain. Read Hebden pages 187 & 188.

2) Do Hebden p.185 #144, 145 & p.188 #147

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1) There are many environmental problems associated with acid rain. Read Hebden pages 187 & 188.

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answers in the back of Hebden

XI) Acid/Base Testing

Suppose you had a 1.0M solution of strong acid and weak acid but did not know which was which. Describe any testing you could do to identify each solution.

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XI) Acid/Base Testing

Suppose you had a 1.0M solution of strong acid and weak acid but did not know which was which. Describe any testing you could do to identify each solution.

1) Use an indicator that changes colour at low pH such as thymol blue
2) Use pH paper or universal indicator
3) Use a pH meter
4) Do a conductivity test - the stronger acid will light the bulb up brighter

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