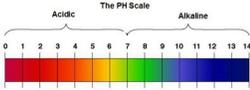


ACIDS AND BASES

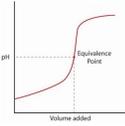




The pH Scale







Acids / Bases

ACID

- Sour taste, turn litmus paper red, low pH





BASE

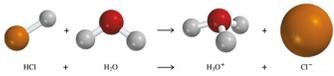
- Bitter taste, turn litmus paper blue, high pH




Arrhenius Definition

ACID

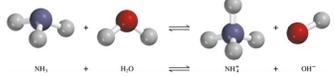
- Donates a hydrogen ion (H^+) in water



$$HCl + H_2O \rightarrow H_3O^+ + Cl^-$$

BASE

- Donates a hydroxide ion (OH^-) in water



$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

Limited to substances with those "parts"... NH_3 exception!

Bronsted-Lowry Definition

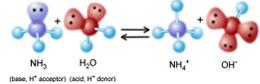
ACID

- Donates a proton in water



BASE

- Accepts a proton in water

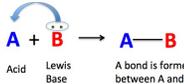


Better definition that explains NH_3 as a base!

Lewis Definition

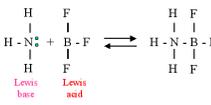
ACID

- Accepts an electron pair



BASE

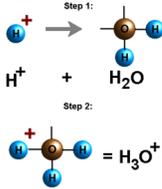
- Donates an electron pair



Explains all traditional acids and bases!

H^+ Ion

- Just like a proton in the nucleus with no surrounding valence e^-
- H^+ reacts strongly with the lone pairs of e^- around the oxygen atom of water
- H_3O^+ (hydronium ion) forms from H^+ reacting with H_2O



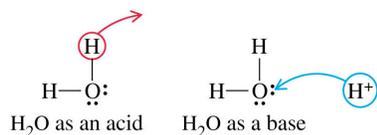
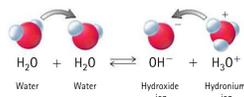
H^+ and H_3O^+ are interchangeable!

Amphiprotic / Amphoteric

Molecules or ions that can act as **EITHER** acids or bases

• Amphiprotic... capable of accepting or donating a proton

• Ex: H₂O



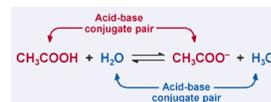
Conjugate Acid-Base Pairs

Pair of compounds that differ by the presence of one H⁺ unit



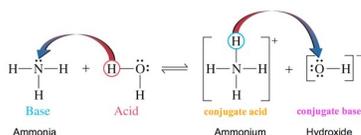
Forward: **Acid** **Conjugate Base**

Reverse: **Conjugate Acid** **Base**



Conjugate Acid-Base Pairs

Pair of compounds that differ by the presence of one H⁺ unit



• To get the **CONJUGATE BASE** of an **ACID**...
Remove H⁺ (don't forget the "-" sign!)

• To get the **CONJUGATE ACID** of a **BASE**...
Add H⁺ (don't forget the "+" sign!)

Practice

• **EXAMPLE:**

Write the conjugate base for each of the following acids.



Practice

• **EXAMPLE:**

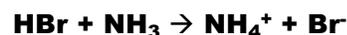
Write the conjugate acid for each of the following bases.



Practice

• **EXAMPLE:**

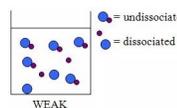
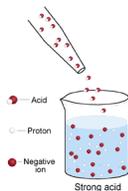
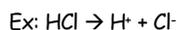
For the following reaction, identify the **acid** on the left and its **conjugate base** on the right. Similarly, identify the **base** on the left and its **conjugate acid** on the right.



Strengths of Acids and Bases

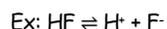
STRONG Acids / Bases

- Dissociate **COMPLETELY** in water
- Very large K values (no equilibrium though)



WEAK Acids / Bases

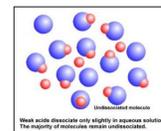
- Dissociate only **SLIGHTLY** in water
- Very small K values



Strengths of Acids and Bases

*NOTE ABOUT WEAK ACIDS / BASES AND NET-IONIC EQUATIONS:

When writing net-ionic equations on the AP Exam, the WEAK acids / bases do NOT rip apart into ions since they are not 100% dissociated!



Ex: Write net-ionic for $\text{HC}_2\text{H}_3\text{O}_2(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow$

Strengths of Acids and Bases

6 STRONG ACIDS:

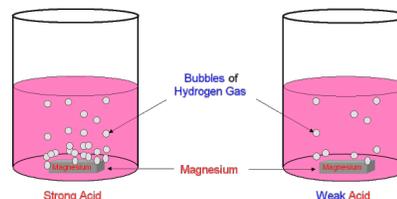
HCl
HBr
HI
HNO₃
H₂SO₄
HClO₄

5 STRONG BASES:

LiOH
NaOH
KOH
Ba(OH)₂
Sr(OH)₂

DO NOT CONFUSE STRENGTH (HOW MUCH DISSOCIATION TAKES PLACE) WITH CONCENTRATION (mol/L)!!

Strengths of Acids and Bases



If same amount of Mg added to each AND the concentration of each are the same (ex: 1.0 M), both produce the **SAME AMOUNT** of H₂ gas... but the **STRONG** acid does it faster and with more bubbles because it is more ionized!!

Strengths of Acids and Bases

EXAMPLES:

If we have 1.25 moles of HCl in water, how many moles of H⁺ and Cl⁻ will there be?

If we have 0.236 M HNO₃, what is the [H⁺] and [NO₃⁻]

Strengths of Acids and Bases

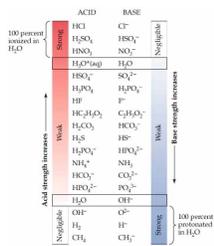
STRONGER the acid, the **WEAKER** the conjugate base!

STRONGER the base, the **WEAKER** the conjugate acid!



Conjugate base (Cl⁻) of the strong acid does **NOT** want to grab the H⁺

Strengths of Acids and Bases



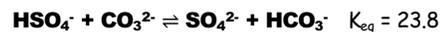
• In every acid-base reaction, the position of the equilibrium favors the transfer of the proton from the **STRONGER** acid to the **STRONGER** base to form the weaker acid and the weaker base...

WEAKER ACIDS / BASES ARE FAVORED AT EQUILIBRIUM!!

Strengths of Acids and Bases

• EXAMPLE:

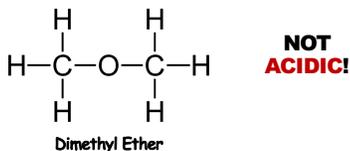
Based on the reaction given below, which substance is the strongest base in this system?



- a) SO_4^{2-}
- b) CO_3^{2-}
- c) H_2O
- d) HCO_3^-
- e) HSO_4^-

Factors Affecting Strength

- Not every molecule with an H is an acid or every molecule with OH a base...
- Strong bonds (X—H) don't want to break to give off H⁺
- High bond **POLARITY** (H bonded to electronegative atom) tends to **INCREASE** the acidity of the H atom because the H wants "get away"
- Example:



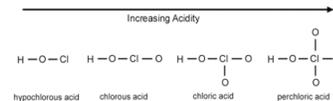
Factors Affecting Strength

- Molecules of the form H—O—X:

-If X has high electronegativity... H tends to be **ACIDIC**

-If X has low electronegativity... OH tends to come off and the compound is **BASIC**

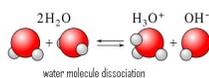
-More oxygens around X... the more **ACIDIC**



Autoionization of Water



- Very important because it allows water to act as an acid or base



$$K_w = [\text{H}^+][\text{OH}^-]$$

$$K_w = 1 \times 10^{-14} \text{ at } 25^\circ \text{ C}$$

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

pH Scale

How acidic or basic a SOLUTION is

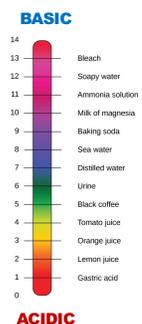
- pH is a logarithmic scale which is an exponential scale

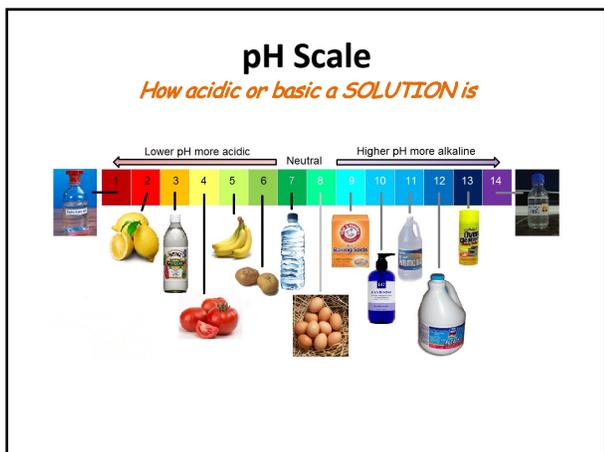
• Ex: pH of 3 is **10 times** stronger than a pH of 4 and **100 times** stronger than a pH of 5

$$[\text{H}^+] = [\text{OH}^-] \rightarrow \text{NEUTRAL (pH = 7)}$$

$$[\text{H}^+] > [\text{OH}^-] \rightarrow \text{ACIDIC (pH < 7)}$$

$$[\text{H}^+] < [\text{OH}^-] \rightarrow \text{BASIC (pH > 7)}$$





pH Scale

pH = -log [H⁺]

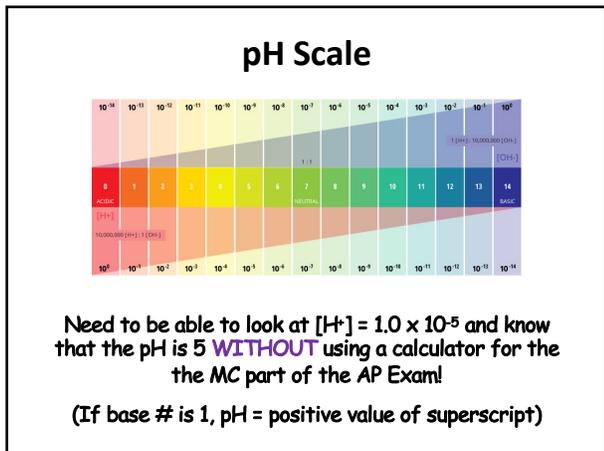
pOH = -log [OH⁻]

pH + pOH = 14

[H⁺] = 10^{-pH}

[OH⁻] = 10^{-pOH}

SIG FIGS with pH: Use as many **DECIMAL PLACES** on pH as there are SF in the **LEAST** accurate measurement... (Ex: 1.15 M has 3 SF, so pH with 3 decimal places)



pH Scale

• **EXAMPLES:**

In a sample of lemon juice, [H⁺] is 3.8 × 10⁻⁴ M.
What is the pH?

A window-cleaning solution has [OH⁻] = 1.9 × 10⁻⁶ M.
What is the pH?

pH Scale

• **EXAMPLES:**

A solution formed by dissolving an antacid tablet has a pH of 9.18. Calculate the [H⁺].

The pH of a sample of human blood was measured to be 7.41 at 25° C. Calculate the pOH, [H⁺], and [OH⁻] for the sample.

Strong Acids and Bases

- Exist in solution **ENTIRELY** as ions
- Do **NOT** exist in equilibrium... reaction lies entirely to the right (side with the ions)
- Determine the [H⁺] or [OH⁻] to find the pH or pOH

Strong Acids and Bases

• EXAMPLES:

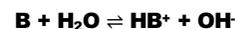
Calculate the pH of 0.050 M HNO₃.

What is the pH of 5.0 × 10⁻² M NaOH?

What is the concentration of a solution of KOH for which the pH is 11.89?

Weak Acids and Bases

- Only **PARTIALLY** ionized in solution
- Use equilibrium rules with ICE to determine how much of the acid or base dissociates and find the []
- Use K_a (**Acid Dissociation Constant**) for weak acids and K_b (**Base Dissociation Constant**) for weak bases



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

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

Weak Acids and Bases

- Only **PARTIALLY** ionized in solution
- Use equilibrium rules with ICE to determine how much of the acid or base dissociates and find the []
- Use K_a (**Acid Dissociation Constant**) for weak acids and K_b (**Base Dissociation Constant**) for weak bases

GREATER the K_a... **STRONGER** the acid!
(greater dissociation of the acid)

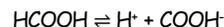
GREATER the K_b... **STRONGER** the base!
(greater protonation of the base)

Use K_{eq} knowledge to help remember this!

Weak Acids and Bases

• EXAMPLE:

What is the K_a of 0.10 M of formic acid, HCOOH, if its pH is 2.38?



	HCOOH	H ⁺	COOH ⁻
Initial			
Change			
Equilibrium			

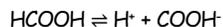
$$\text{pH} = 2.38$$

$$[\text{H}^+] = 4.2 \times 10^{-3} \text{ M}$$

Weak Acids and Bases

• EXAMPLE:

What is the K_a of 0.10 M of formic acid, HCOOH, if its pH is 2.38?



	HCOOH	H ⁺	COOH ⁻
Initial	0.10 M	0	0
Change			
Equilibrium		4.2 × 10 ⁻³ M	

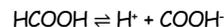
$$\text{pH} = 2.38$$

$$[\text{H}^+] = 4.2 \times 10^{-3} \text{ M}$$

Weak Acids and Bases

• EXAMPLE:

What is the K_a of 0.10 M of formic acid, HCOOH, if its pH is 2.38?



	HCOOH	H ⁺	COOH ⁻
Initial	0.10 M	0	0
Change	-4.2 × 10 ⁻³ M	+4.2 × 10 ⁻³ M	+4.2 × 10 ⁻³ M
Equilibrium		4.2 × 10 ⁻³ M	

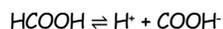
$$\text{pH} = 2.38$$

$$[\text{H}^+] = 4.2 \times 10^{-3} \text{ M}$$

Weak Acids and Bases

• **EXAMPLE:**

What is the K_a of 0.10 M of formic acid, HCOOH, if its pH is 2.38?



	HCOOH	H ⁺	COOH ⁻
Initial	0.10 M	0	0
Change	-4.2 x 10 ⁻³ M	+4.2 x 10 ⁻³ M	+4.2 x 10 ⁻³ M
Equilibrium	9.6 x 10 ⁻² M	4.2 x 10 ⁻³ M	4.2 x 10 ⁻³ M

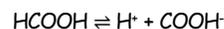
pH = 2.38

[H⁺] = 4.2 x 10⁻³ M

Weak Acids and Bases

• **EXAMPLE:**

What is the K_a of 0.10 M of formic acid, HCOOH, if its pH is 2.38?



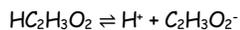
	HCOOH	H ⁺	COOH ⁻
Initial	0.10 M	0	0
Change	-4.2 x 10 ⁻³ M	+4.2 x 10 ⁻³ M	+4.2 x 10 ⁻³ M
Equilibrium	9.6 x 10 ⁻² M	4.2 x 10 ⁻³ M	4.2 x 10 ⁻³ M

$$K_a = \frac{[\text{H}^+][\text{COOH}^-]}{[\text{HCOOH}]} = \frac{[4.2 \times 10^{-3}]^2}{[9.6 \times 10^{-2}]} = 1.8 \times 10^{-4}$$

Weak Acids and Bases

• **EXAMPLE:**

What is the pH of 0.20 M HC₂H₃O₂ with a $K_a = 1.8 \times 10^{-5}$?



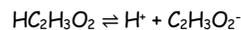
	HC ₂ H ₃ O ₂	H ⁺	C ₂ H ₃ O ₂ ⁻
Initial			
Change			
Equilibrium			

$$K_a = \frac{[\text{H}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]}$$

Weak Acids and Bases

• **EXAMPLE:**

What is the pH of 0.20 M HC₂H₃O₂ with a $K_a = 1.8 \times 10^{-5}$?



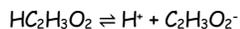
	HC ₂ H ₃ O ₂	H ⁺	C ₂ H ₃ O ₂ ⁻
Initial	0.20 M	0	0
Change			
Equilibrium			

$$K_a = \frac{[\text{H}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]}$$

Weak Acids and Bases

• **EXAMPLE:**

What is the pH of 0.20 M HC₂H₃O₂ with a $K_a = 1.8 \times 10^{-5}$?



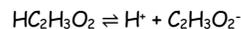
	HC ₂ H ₃ O ₂	H ⁺	C ₂ H ₃ O ₂ ⁻
Initial	0.20 M	0	0
Change	-x	+x	+x
Equilibrium			

$$K_a = \frac{[\text{H}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]}$$

Weak Acids and Bases

• **EXAMPLE:**

What is the pH of 0.20 M HC₂H₃O₂ with a $K_a = 1.8 \times 10^{-5}$?



	HC ₂ H ₃ O ₂	H ⁺	C ₂ H ₃ O ₂ ⁻
Initial	0.20 M	0	0
Change	-x	+x	+x
Equilibrium	0.20 - x M	x M	x M

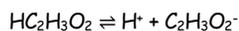
$$K_a = \frac{[\text{H}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]}$$

For these problems, we ALWAYS ignore what is subtracted or added because it is NEGLIGIBLE!

Weak Acids and Bases

• EXAMPLE:

What is the pH of 0.20 M $\text{HC}_2\text{H}_3\text{O}_2$ with a $K_a = 1.8 \times 10^{-5}$?



	$\text{HC}_2\text{H}_3\text{O}_2$	H^+	$\text{C}_2\text{H}_3\text{O}_2^-$
Initial	0.20 M	0	0
Change	-x	+x	+x
Equilibrium	$0.20 - x$ M	x M	x M

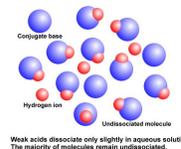
$$1.8 \times 10^{-5} = \frac{x^2}{0.20 - x} \quad x = [\text{H}^+] = 1.9 \times 10^{-3}$$

$$\text{pH} = -\log [\text{H}^+] = 2.72$$

Weak Acids and Bases

• No need to check assumptions using the 5% rule for weak acids and bases

• Determination of pH of a weak base is very similar to weak acid problems **EXCEPT** that x will be $[\text{OH}^-]$, giving the pOH not the pH!

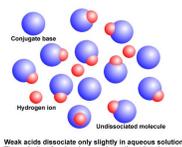


**PAY CLOSE ATTENTION
TO WHAT YOU ARE
SOLVING FOR!!**

Weak Acids and Bases

• In mixtures of **WEAK** acids, only the acid with the **LARGEST** K_a value will contribute an appreciable $[\text{H}^+]$... use this acid to determine the pH (same is true for bases too)

• In mixtures of **STRONG** acids, determine the moles of H^+ from both and **ADD** together to give total moles... determine $[\text{H}^+]$ from the new total volume to find the pH



Weak Acids and Bases

• EXAMPLES:

The hypochlorite ion (OCl^-) is a strong oxidizing agent often found in household bleaches. Calculate the pH of a 0.10 M solution of hypochlorous acid ($K_a = 3.5 \times 10^{-8}$)?

Calculate the pH for a 15.0 M solution of NH_3 ($K_b = 1.8 \times 10^{-5}$).

Weak Acids and Bases

• EXAMPLES:

Calculate the pH of a solution that contains 1.00 M HCN ($K_a = 6.2 \times 10^{-10}$) and 5.00 M HNO_2 ($K_a = 4.0 \times 10^{-4}$).

Determine the pH of a 1.0 M solution of methylamine (CH_3N , $K_b = 4.38 \times 10^{-4}$).

% Dissociation / Ionization

- Describes the strength of an acid just like K_a
- **GREATER** % Dissociation... **STRONGER** the acid (remember strong acids are 100% dissociated)

$$\% \text{ Dissociation / Ionization} = \frac{[\text{H}^+] \text{ or } [\text{OH}^-] \text{ Ionized}}{[\text{Acid}] \text{ or } [\text{Base}]} \times 100$$

• EXAMPLE:

What is the % Dissociation for 0.10 M HCOOH which contains $[\text{H}^+] = 4.2 \times 10^{-3}$ M?

4.2%

% Dissociation / Ionization

• EXAMPLES:

Calculate the % ionization of 0.15 M acetic acid ($K_a = 1.8 \times 10^{-5}$).

Lactic acid ($\text{HC}_3\text{H}_5\text{O}_3$) is a waste product that accumulates in muscle tissue during exertion. In a 0.10 M solution, lactic acid is 3.7% dissociated. Calculate the value of K_a for this acid.

Polyprotic Acids

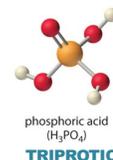
Acids that can donate MORE than one H⁺

• **MONOPROTIC:** acids donating ONE H⁺
(Ex: HNO₃)

• **DIPROTIC:** acids donating TWO H⁺
(Ex: H₂SO₄)

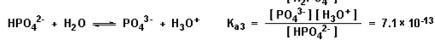
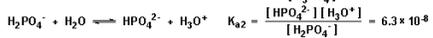
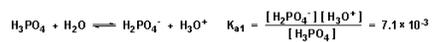
• **POLYPROTIC BASES:** accept MORE than one H⁺
(Ex: PO₄³⁻)

• Dissociate in steps one proton at a time with each step having its own K_a value



Polyprotic Acids

Acids that can donate MORE than one H⁺



• First dissociation (K_{a1}) is always **GREATEST** and those that follow have much smaller K_a values (acid gets weaker with each H⁺ removed but harder to remove)

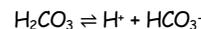
• Usually only the **FIRST** dissociation (K_{a1}) will be important in determining pH for polyprotic acids

Polyprotic Acids

Acids that can donate MORE than one H⁺

• EXAMPLE:

If $[\text{H}_2\text{CO}_3] = 0.0037 \text{ M}$, what is the pH and $[\text{CO}_3^{2-}]$? ($K_{a1} = 4.3 \times 10^{-7}$, $K_{a2} = 5.6 \times 10^{-11}$)



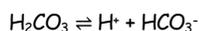
	H ₂ CO ₃	H ⁺	HCO ₃ ⁻
Initial			
Change			
Equilibrium			

Polyprotic Acids

Acids that can donate MORE than one H⁺

• EXAMPLE:

If $[\text{H}_2\text{CO}_3] = 0.0037 \text{ M}$, what is the pH and $[\text{CO}_3^{2-}]$? ($K_{a1} = 4.3 \times 10^{-7}$, $K_{a2} = 5.6 \times 10^{-11}$)



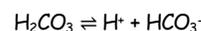
	H ₂ CO ₃	H ⁺	HCO ₃ ⁻
Initial	0.0037 M	0	0
Change			
Equilibrium			

Polyprotic Acids

Acids that can donate MORE than one H⁺

• EXAMPLE:

If $[\text{H}_2\text{CO}_3] = 0.0037 \text{ M}$, what is the pH and $[\text{CO}_3^{2-}]$? ($K_{a1} = 4.3 \times 10^{-7}$, $K_{a2} = 5.6 \times 10^{-11}$)



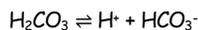
	H ₂ CO ₃	H ⁺	HCO ₃ ⁻
Initial	0.0037 M	0	0
Change	-x	+x	+x
Equilibrium			

Polyprotic Acids

Acids that can donate MORE than one H⁺

• **EXAMPLE:**

If $[H_2CO_3] = 0.0037\text{ M}$, what is the pH and $[CO_3^{2-}]$? ($K_{a1} = 4.3 \times 10^{-7}$, $K_{a2} = 5.6 \times 10^{-11}$)



	H_2CO_3	H^+	HCO_3^-
Initial	0.0037 M	0	0
Change	-x	+x	+x
Equilibrium	0.0037 - x M	x M	x M

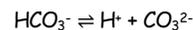
$x = [H^+] = 4.0 \times 10^{-5}$ $pH = -\log [H^+] = 4.40$

Polyprotic Acids

Acids that can donate MORE than one H⁺

• **EXAMPLE:**

If $[H_2CO_3] = 0.0037\text{ M}$, what is the pH and $[CO_3^{2-}]$? ($K_{a1} = 4.3 \times 10^{-7}$, $K_{a2} = 5.6 \times 10^{-11}$)



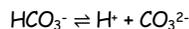
	HCO_3^-	H^+	CO_3^{2-}
Initial	$4.0 \times 10^{-5}\text{ M}$	$4.0 \times 10^{-5}\text{ M}$	0
Change			
Equilibrium			

Polyprotic Acids

Acids that can donate MORE than one H⁺

• **EXAMPLE:**

If $[H_2CO_3] = 0.0037\text{ M}$, what is the pH and $[CO_3^{2-}]$? ($K_{a1} = 4.3 \times 10^{-7}$, $K_{a2} = 5.6 \times 10^{-11}$)



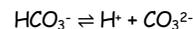
	HCO_3^-	H^+	CO_3^{2-}
Initial	$4.0 \times 10^{-5}\text{ M}$	$4.0 \times 10^{-5}\text{ M}$	0
Change	-x	+x	+x
Equilibrium			

Polyprotic Acids

Acids that can donate MORE than one H⁺

• **EXAMPLE:**

If $[H_2CO_3] = 0.0037\text{ M}$, what is the pH and $[CO_3^{2-}]$? ($K_{a1} = 4.3 \times 10^{-7}$, $K_{a2} = 5.6 \times 10^{-11}$)



	HCO_3^-	H^+	CO_3^{2-}
Initial	$4.0 \times 10^{-5}\text{ M}$	$4.0 \times 10^{-5}\text{ M}$	0
Change	-x	+x	+x
Equilibrium	$4.0 \times 10^{-5} - x\text{ M}$	$4.0 \times 10^{-5} + x\text{ M}$	x

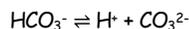
Ignore the + and - x!!

Polyprotic Acids

Acids that can donate MORE than one H⁺

• **EXAMPLE:**

If $[H_2CO_3] = 0.0037\text{ M}$, what is the pH and $[CO_3^{2-}]$? ($K_{a1} = 4.3 \times 10^{-7}$, $K_{a2} = 5.6 \times 10^{-11}$)



	HCO_3^-	H^+	CO_3^{2-}
Initial	$4.0 \times 10^{-5}\text{ M}$	$4.0 \times 10^{-5}\text{ M}$	0
Change	-x	+x	+x
Equilibrium	$4.0 \times 10^{-5} - x\text{ M}$	$4.0 \times 10^{-5} + x\text{ M}$	x

$x = [CO_3^{2-}] = 5.6 \times 10^{-11}$ $K_{a2} = [\text{conjugate}]!!$

Common Ion Effect

REDUCTION in the solubility when a soluble compound containing one of the SAME ions is added

- Common ion will affect the pH
- Weak acid or base will dissociate / ionize **LESS** than if it were alone

• **EXAMPLE:**

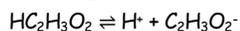
Determine the pH of 0.30 moles of $HC_2H_3O_2$ ($K_a = 1.8 \times 10^{-5}$) and 0.30 moles of $NaC_2H_3O_2$ in 1.0 L of water.

$C_2H_3O_2^-$ is the **COMMON ION!**

Common Ion Effect

• **EXAMPLE:**

Determine the pH of 0.30 moles of $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$) and 0.30 moles of $\text{NaC}_2\text{H}_3\text{O}_2$ in 1.0 L of water.

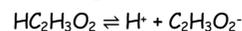


	$\text{HC}_2\text{H}_3\text{O}_2$	H^+	$\text{C}_2\text{H}_3\text{O}_2^-$
Initial	0.30 M	0	0.30 M
Change			
Equilibrium			

Common Ion Effect

• **EXAMPLE:**

Determine the pH of 0.30 moles of $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$) and 0.30 moles of $\text{NaC}_2\text{H}_3\text{O}_2$ in 1.0 L of water.

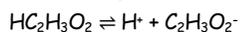


	$\text{HC}_2\text{H}_3\text{O}_2$	H^+	$\text{C}_2\text{H}_3\text{O}_2^-$
Initial	0.30 M	0	0.30 M
Change	-x	+x	+x
Equilibrium			

Common Ion Effect

• **EXAMPLE:**

Determine the pH of 0.30 moles of $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$) and 0.30 moles of $\text{NaC}_2\text{H}_3\text{O}_2$ in 1.0 L of water.



	$\text{HC}_2\text{H}_3\text{O}_2$	H^+	$\text{C}_2\text{H}_3\text{O}_2^-$
Initial	0.30 M	0	0.30 M
Change	-x	+x	+x
Equilibrium	0.30 - x M	x M	0.30 + x M

$$x = [\text{H}^+] = 1.8 \times 10^{-5} \quad \text{pH} = -\log [\text{H}^+] = 4.74$$

K_a and K_b Relationship

$$K_a \times K_b = K_w = 1 \times 10^{-14}$$

$$\text{p}K_a = -\log K_a$$

$$\text{p}K_b = -\log K_b$$

$$\text{p}K_a + \text{p}K_b = \text{p}K_w = 14$$

• $\text{p}K_a$ / $\text{p}K_b$: how acidic or basic a hydrogen or hydroxide is in a molecule

K_a and K_b Relationship

STRONG Acids / Bases

- Larger K_a or K_b = **STRONGER**
- Lower $\text{p}K_a$ or $\text{p}K_b$ = **STRONGER**



WEAK Acids / Bases

- Smaller K_a or K_b = **WEAKER**
- Higher $\text{p}K_a$ or $\text{p}K_b$ = **WEAKER**

K_a and K_b Relationship

• **EXAMPLES:**

Calculate the pH of 0.20 M lactic acid, $\text{HC}_3\text{H}_5\text{O}_3$, with a $\text{p}K_a$ of 3.08.

What is the pH of a 0.045 M solution of KOB^- ?
The $\text{p}K_a$ of HOBr is 8.63. (*Hint: write the weak base reaction of OBr^- with water first)

Neutralization Reactions

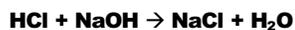
When an acid and base mix, the acid will donate protons to the base (usually get a salt and water)

FOUR TYPES:

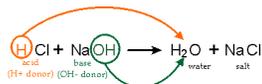
1) Strong Acid and Strong Base

• Products are always a salt and water

• Example:



• Try it: $\text{HClO}_4 + \text{NaOH}$



Acid + Base →
Salt + Water

Neutralization Reactions

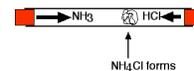
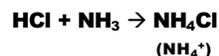
When an acid and base mix, the acid will donate protons to the base (usually get a salt and water)

FOUR TYPES:

2) Strong Acid and Weak Base

• Salt forms but water may not be present if no "OH" in the base

• Example:



Neutralization Reactions

When an acid and base mix, the acid will donate protons to the base (usually get a salt and water)

FOUR TYPES:

3) Weak Acid and Strong Base

• Salt is the conjugate base of the weak acid and water are products

• Example:



Neutralization Reactions

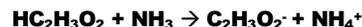
When an acid and base mix, the acid will donate protons to the base (usually get a salt and water)

FOUR TYPES:

4) Weak Acid and Weak Base

• Products are the conjugate base and conjugate acid (or salts of them)

• Example:



Practice

• EXAMPLES:

Write the products for the following acid-base reactions.



Acid-Base Properties of Salts



• Salts produced in neutralization reactions are **NOT** always neutral... might contribute to pH!!

• Most salts **FULLY** dissociate into their ions

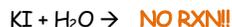
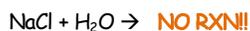
• Some can **hydrolyze** (react with water) to produce acidic or basic solutions

Acid-Base Properties of Salts

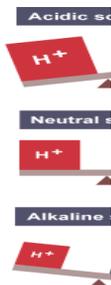
GENERAL STATEMENTS:

• Salts of strong acids and strong bases (ex: NaCl) do NOT hydrolyze and solution is **NEUTRAL**... pH = 7

• Examples:



You MUST look at which acid and base made the salt!!



Acid-Base Properties of Salts

GENERAL STATEMENTS:

• Salts of weak acids and strong bases (ex: NaC₂H₃O₂) do hydrolyze and solution is **BASIC**... pH > 7 (anion acts as a base)

Remember... the salt FULLY dissociates into its ions which might be able to hydrolyze with water!

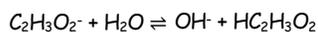


Acid-Base Properties of Salts

GENERAL STATEMENTS:

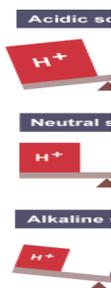
• Salts of weak acids and strong bases (ex: NaC₂H₃O₂) do hydrolyze and solution is **BASIC**... pH > 7 (anion acts as a base)

• Example:



BASE

***NOTE: Hydrolysis reactions are equilibrium reactions!**

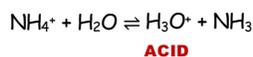


Acid-Base Properties of Salts

GENERAL STATEMENTS:

• Salts of strong acids and weak bases (ex: NH₄Cl) do hydrolyze and solution is **ACIDIC**... pH < 7 (cation acts as an acid)

• Example:



ACID



Acid-Base Properties of Salts

GENERAL STATEMENTS:

• Salts of weak acids and weak bases (Ex: NH₄C₂H₃O₂) do hydrolyze but pH depends on **K_a** and **K_b** values (both components hydrolyze)

-If **K_a** of acidic ion is greater than **K_b** of basic ion... **ACIDIC pH**

-If **K_b** of basic ion is greater than **K_a** of acidic ion... **BASIC pH**

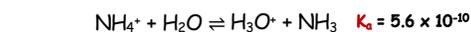
-If **K_a** of acidic ion and **K_b** of basic ion are equal... **NEUTRAL pH**

Acid-Base Properties of Salts

GENERAL STATEMENTS:

• Salts of weak acids and weak bases (Ex: NH₄C₂H₃O₂) do hydrolyze but pH depends on **K_a** and **K_b** values (both components hydrolyze)

• Example:



NEUTRAL!

Practice

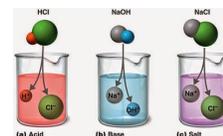
• EXAMPLES:

Predict whether an aqueous solution of the following salts will be acidic, basic, or neutral. Prove with appropriate equations.



Acid-Base Properties of Salts

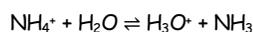
- When calculating the pH of salt solutions, pay attention to whether or not the salt will hydrolyze to form **acid**, **basic**, or **neutral** pH
- Write the appropriate equilibrium reaction for the hydrolysis
- If H^+ is formed, K_a will be needed to solve the equilibrium... if OH^- is formed, K_b will be needed! ($K_a \times K_b = K_w$)



Acid-Base Properties of Salts

• EXAMPLE:

Calculate the pH of a 0.10 M NH_4Cl solution.
The K_b value for NH_3 is 1.8×10^{-5} .



Because an **ACID** is forming,
we need the **K_a** value!!

$$K_a \times K_b = K_w$$

$$K_a \times 1.8 \times 10^{-5} = 1 \times 10^{-14}$$

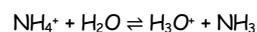
$$K_a = 5.6 \times 10^{-10}$$

From strong acid and weak base... so forms **ACIDIC!**

Acid-Base Properties of Salts

• EXAMPLE:

Calculate the pH of a 0.10 M NH_4Cl solution.
The K_b value for NH_3 is 1.8×10^{-5} .

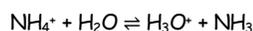


	NH_4^+	H^+	NH_3
Initial			
Change			
Equilibrium			

Acid-Base Properties of Salts

• EXAMPLE:

Calculate the pH of a 0.10 M NH_4Cl solution.
The K_b value for NH_3 is 1.8×10^{-5} .

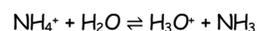


	NH_4^+	H^+	NH_3
Initial	0.10 M	0	0
Change			
Equilibrium			

Acid-Base Properties of Salts

• EXAMPLE:

Calculate the pH of a 0.10 M NH_4Cl solution.
The K_b value for NH_3 is 1.8×10^{-5} .

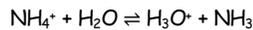


	NH_4^+	H^+	NH_3
Initial	0.10 M	0	0
Change	-x	+x	+x
Equilibrium			

Acid-Base Properties of Salts

• **EXAMPLE:**

Calculate the pH of a 0.10 M NH_4Cl solution.
The K_b value for NH_3 is 1.8×10^{-5} .



	NH_4^+	H^+	NH_3
Initial	0.10 M	0	0
Change	-x	+x	+x
Equilibrium	0.10 - x M	x M	x M

$x = [\text{H}^+] = 7.5 \times 10^{-6}$ $\text{pH} = -\log [\text{H}^+] = 5.12$

Acid-Base Properties of Salts

• **EXAMPLE:**

Calculate the pH of a 0.30 M NaF solution.
The K_a value for HF is 7.2×10^{-4} .

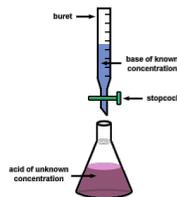
Titration

Solution with a known [] is added to a solution with an unknown [] in order to determine the []

• Neutralization reactions are generally performed this way (acid added to base or vice versa)

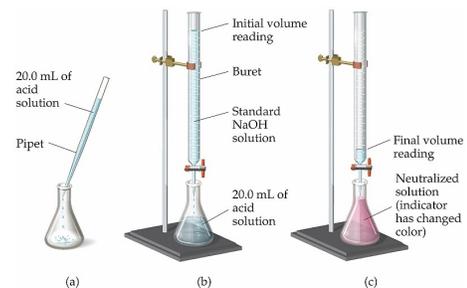
• Uses a burette filled with **STANDARDIZED** (know the exact []) solution

• Exact amount of solution added is important... Proper techniques **MUST** be followed (you will learn these during lab)



Titration

Solution with a known [] is added to a solution with an unknown [] in order to determine the []

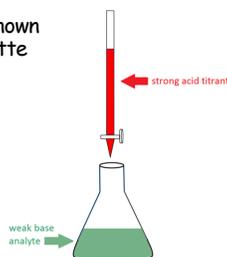


Titration

Solution with a known [] is added to a solution with an unknown [] in order to determine the []

• **TITRANT:** substance with a known [] that is added from the burette to substance with unknown []

• **ANALYTE:** substance being analyzed with unknown []



Titration

Solution with a known [] is added to a solution with an unknown [] in order to determine the []

****Watch equations for ratio... is it 1:1, etc? →**

• **EQUIVALENCE POINT:** point at which equal stoichiometric amounts of acid and base have been added (**moles acid = moles base**)

• **HALF-EQUIVALENCE POINT:** enough substance added to convert exactly half of the acid or base into its conjugate (**pH = pK_a**)



• **INDICATORS** (change color at specific pH) are added to identify the equivalence point

Titrations

Solution with a known [] is added to a solution with an unknown [] in order to determine the []



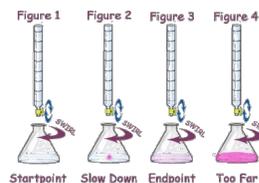
Good Endpoint

Bad Endpoint (Overly Titrated)

Be careful when adding the titrant to the **ENDPOINT (Point in titration when the indicator undergoes a color change)... important not to go too far over!**

Titrations

Solution with a known [] is added to a solution with an unknown [] in order to determine the []



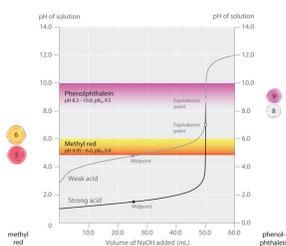
Startpoint

Slow Down

Endpoint

Too Far

Selecting Indicators



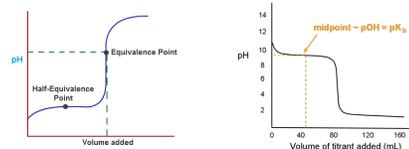
- Indicator color changes can be sharp and occur with the addition of a single extra drop of titrant
- Choose a suitable indicator based on the **EQUIVALENCE POINT** of the reaction

Indicators with a $pK_a = pH$ at the equivalence point work best

Titration Curves

Graphs of pH (y-axis) as a function of titrant added (x-axis)

- Can be generated from burette and pH readings (paper or meter)
- pH increases **SLOWLY** at first and then **JUMPS** just before the equivalence point



Base added to Acid

Acid added to Base

Titration Curves

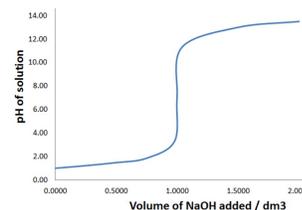
Graphs of pH (y-axis) as a function of titrant added (x-axis)

- When acid and base are both **STRONG**... equivalence point $pH = 7$
- When using **STRONG** acid and **WEAK** base... equivalence point $pH < 7$
- When using **WEAK** acid and **STRONG** base... equivalence point $pH > 7$

pH can be found for titrations at different points...

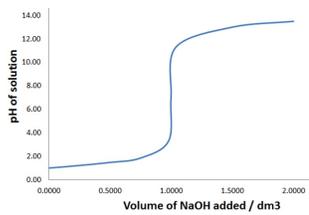


Shape of the Curve



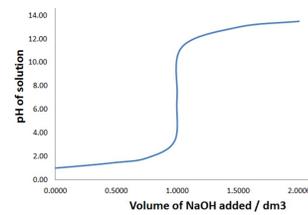
INITIALLY: pH does NOT change much because of the "buffering" action of the analyte (large amount present versus only small amount of titrant... Analyte destroys what is added)

Shape of the Curve



AT EQUIVALENCE POINT: pH dramatically changes and increases quickly

Shape of the Curve

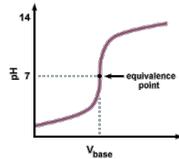


EXCESS ADDED: pH levels off and does not change much (as long as the titrant remains the same)

Strong Acid / Strong Base

• Four Regions of the Curve:

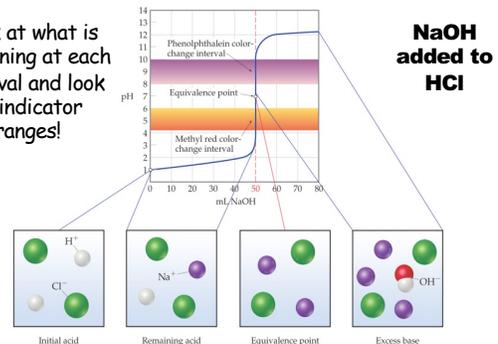
- 1) Initial pH (initial [H^+])
- 2) Between initial pH and equivalence point
- 3) Equivalence point
- 4) After equivalence point



• **ENDPOINT:** point where indicator changes color (usually close to the equivalence point)

Strong Acid / Strong Base

Look at what is happening at each interval and look at indicator ranges!

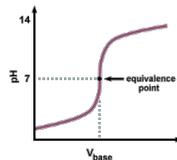


Strong Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl.

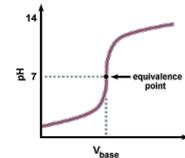
Since the titration involves a **STRONG** acid and **STRONG** base, we know the equivalence point pH = 7!



Strong Acid / Strong Base

• Four Regions of the Curve:

- 1) Initial pH (initial [H^+])
- 2) Between initial pH and equivalence point
- 3) Equivalence point
- 4) After equivalence point



• **ENDPOINT:** point where indicator changes color (usually close to the equivalence point)

Strong Acid / Strong Base

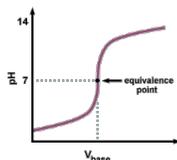
• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl.

To find the pH **BEFORE** any NaOH is added...

Just use $\text{pH} = -\log [\text{H}^+]$
(strong acid so completely dissociated)

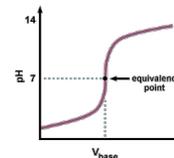
Ans: **1.00**



Strong Acid / Strong Base

• Four Regions of the Curve:

- 1) Initial pH (initial [])
- 2) Between initial pH and equivalence point
- 3) Equivalence point
- 4) After equivalence point



• **ENDPOINT:** point where indicator changes color (usually close to the equivalence point)

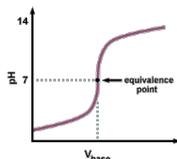
Strong Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl.

To find the pH **AFTER** the addition of 10.0 mL of NaOH...

- 1) Find the moles of NaOH added
- 2) Find amount of moles of HCl started
- 3) Determine moles of HCl left
- 4) Use **TOTAL** volume to determine $[\text{H}^+]$ and then find the pH



Strong Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl.

To find the pH **AFTER** the addition of 10.0 mL of NaOH...

$$10.0 \text{ mL} \times \frac{\text{L}}{1000 \text{ mL}} \times \frac{0.10 \text{ mol}}{\text{L}} = 0.0010 \text{ mol NaOH added}$$

$$25.0 \text{ mL} \times \frac{\text{L}}{1000 \text{ mL}} \times \frac{0.10 \text{ mol}}{\text{L}} = 0.0025 \text{ mol HCl started}$$

Strong Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl.

To find the pH **AFTER** the addition of 10.0 mL of NaOH...

If **0.0010 mol of NaOH** are added, that means **0.0010 mol of HCl** are now gone (**neutralized**) because the reaction is a **1:1 ratio!**

Strong Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl.

To find the pH **AFTER** the addition of 10.0 mL of NaOH...

$$0.0025 \text{ mol} - 0.0010 \text{ mol} = 0.0015 \text{ mol HCl left}$$

$$\frac{0.0015 \text{ mol}}{35.0 \text{ mL}} \times \frac{1000 \text{ mL}}{\text{L}} = 0.043 \text{ M } [\text{H}^+]$$

$$\text{pH} = -\log [\text{H}^+] \\ \text{Don't forget!} \quad -\log [0.043] = \mathbf{1.37}$$

Strong Acid / Strong Base

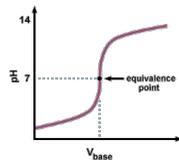
• Four Regions of the Curve:

1) Initial pH (initial [H^+])

2) Between initial pH and equivalence point

3) Equivalence point

4) After equivalence point



• **ENDPOINT:** point where indicator changes color (usually close to the equivalence point)

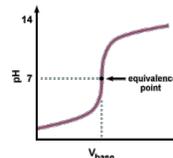
Strong Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl.

To find the pH at the **EQUIVALENCE POINT**...

We know the pH is 7 because both are strong... BUT since both have the same [] we need equal amounts of each to complete neutralization!



25 mL of HCl and 25 mL of NaOH added = pH 7
($[H^+] = [OH^-]$)

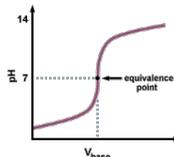
Strong Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl.

To find the pH at the **EQUIVALENCE POINT**...

MOLES ACID = MOLES BASE
AT THE EQUIVALENCE
POINT!!!!!!!!!!!!



Strong Acid / Strong Base

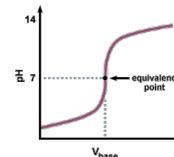
• Four Regions of the Curve:

1) Initial pH (initial [H^+])

2) Between initial pH and equivalence point

3) Equivalence point

4) After equivalence point



• **ENDPOINT:** point where indicator changes color (usually close to the equivalence point)

Strong Acid / Strong Base

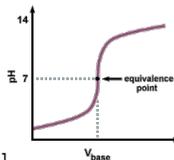
• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl.

To find the pH **AFTER** the addition of 35.0 mL NaOH...

- 1) Find the moles of NaOH added
- 2) Find amount of moles of HCl started
- 3) Determine moles of NaOH excess

4) Use **TOTAL** volume to determine $[OH^-]$ and then find the pH from pOH



Strong Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl.

To find the pH **AFTER** the addition of 35.0 mL NaOH...

$$35.0 \text{ mL} \times \frac{\text{L}}{1000 \text{ mL}} \times \frac{0.10 \text{ mol}}{\text{L}} = 0.0035 \text{ mol NaOH added}$$

$$25.0 \text{ mL} \times \frac{\text{L}}{1000 \text{ mL}} \times \frac{0.10 \text{ mol}}{\text{L}} = 0.0025 \text{ mol HCl started (SAME AS BEFORE)}$$

Strong Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl.

To find the pH **AFTER** the addition of 35.0 mL NaOH...

If 0.0035 mol of NaOH are added, that means **ALL** of HCl are now gone (**neutralized**) because the reaction is a 1:1 ratio!

Strong Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl.

To find the pH **AFTER** the addition of 35.0 mL NaOH...

$$0.0035 \text{ mol} - 0.0025 \text{ mol} = 0.0010 \text{ mol NaOH left}$$

$$\frac{0.0010 \text{ mol}}{60.0 \text{ mL}} \times \frac{1000 \text{ mL}}{\text{L}} = 0.017 \text{ M } [\text{OH}^-]$$

$$\text{pOH} = -\log [\text{OH}^-]$$

$$-\log [0.017] = 1.77$$

** (25.0 + 35.0 mL)**

Strong Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HCl.

To find the pH **AFTER** the addition of 35.0 mL NaOH...

$$0.0035 \text{ mol} - 0.0025 \text{ mol} = 0.0010 \text{ mol NaOH left}$$

$$\frac{0.0010 \text{ mol}}{(25.0 + 35.0 \text{ mL})} \times \frac{1000 \text{ mL}}{\text{L}} = 0.017 \text{ M } [\text{OH}^-]$$

$$\text{pH} + \text{pOH} = 14$$

$$\text{pH} = 14 - 1.77 = \mathbf{12.23}$$

Strong Acid / Strong Base

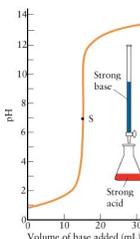
If a **STRONG** acid is added to a **STRONG** base instead, just do the **OPPOSITE!**

Curve will look different too!!

Summary

For **Strong Acid / Strong Base** Titrations:

- pH initially is just $-\log [\text{analyte}]$ because it is 100% dissociated
- pH = 7 at the equivalence point
- At any other point, find moles of acid and moles of base and subtract the two... determine [] of whichever there is more of and use that to determine pH



Practice

• EXAMPLE:

Determine the pH when the following quantities of 0.10 M HNO₃ have been added to 25.0 mL of 0.10 M KOH solution: a) 24.0 mL and b) 27.0 mL

Practice

• EXAMPLE:

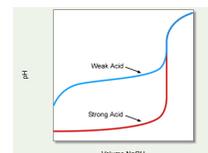
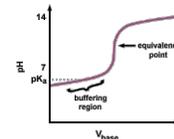
A solution of 0.10 M NaOH is added to a burette and used to titrate 50.0 mL of 0.20 M HNO₃. a) Find the initial pH of the acid solution b) after 65.0 mL of NaOH have been added and c) after 150.0 mL have been added.

Weak Acid / Strong Base

• pH at equivalence point is > 7

• THREE DIFFERENCES:

- 1) Initial acid pH is **HIGHER** (less acidic)
- 2) Smaller change of pH near the equivalence point
- 3) pH at equivalence point is higher

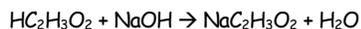


Weak Acid / Strong Base

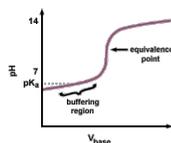
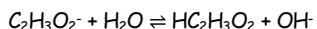
• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HC₂H₃O₂ (K_a = 1.8 × 10⁻⁵).

Get in the habit of writing the **neutralization** and **hydrolysis** reactions first!



Salt is **NOT** neutral... from strong base and weak acid so needs to form OH⁻!

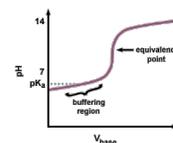


*Note the difference in the arrows!!

Weak Acid / Strong Base

• Four Regions of the Curve:

- 1) Initial pH (initial [])
- 2) Between initial pH and equivalence point
- 3) Equivalence point
- 4) After equivalence point



• **ENDPOINT:** point where indicator changes color (usually close to the equivalence point)

Weak Acid / Strong Base

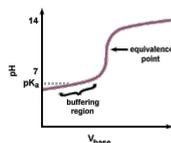
• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M HC₂H₃O₂ (K_a = 1.8 × 10⁻⁵).

To find the pH **BEFORE** any NaOH is added...

Set up the equilibrium problem with ICE table because HC₂H₃O₂ is a weak acid that does not completely dissociate (show the rxn)!

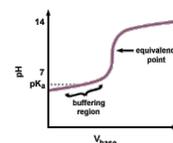
Ans: 2.89



Weak Acid / Strong Base

• Four Regions of the Curve:

- 1) Initial pH (initial [])
- 2) Between initial pH and equivalence point
- 3) Equivalence point
- 4) After equivalence point



• **ENDPOINT:** point where indicator changes color (usually close to the equivalence point)

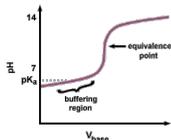
Weak Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$).

To find the pH **AFTER** the addition of 10.0 mL of NaOH...

- 1) Find the moles of NaOH added
- 2) Find amount of moles of acid started
- 3) Determine moles of acid left
- 4) Determine moles of weak base formed
- 5) Determine [] of acid and weak base then plug into HH equation to get pH



Henderson-Hasselbalch Equation

$$\text{pH} = \text{p}K_a + \log \frac{[\text{Base}]}{[\text{Acid}]}$$

OR

$$\text{pOH} = \text{p}K_b + \log \frac{[\text{Acid}]}{[\text{Base}]}$$

*Used for **BUFFERS** (weak acids / bases and their conjugate bases / acids)... More to come!

We will use it here because with only 10.0 mL of NaOH added, we have a weak acid and its conjugate base remaining!

Weak Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$).

To find the pH **AFTER** the addition of 10.0 mL of NaOH...

$$10.0 \text{ mL} \times \frac{\text{L}}{1000 \text{ mL}} \times \frac{0.10 \text{ mol}}{\text{L}} = 0.0010 \text{ mol NaOH added}$$

$$25.0 \text{ mL} \times \frac{\text{L}}{1000 \text{ mL}} \times \frac{0.10 \text{ mol}}{\text{L}} = 0.0025 \text{ mol acid started}$$

Weak Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$).

To find the pH **AFTER** the addition of 10.0 mL of NaOH...

If 0.0010 mol of NaOH are added, that means 0.0010 mol of $\text{HC}_2\text{H}_3\text{O}_2$ are neutralized because the reaction is a 1:1 ratio!

Weak Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$).

To find the pH **AFTER** the addition of 10.0 mL of NaOH...

$$0.0025 \text{ mol} - 0.0010 \text{ mol} = 0.0015 \text{ mol acid left}$$

If 0.0010 mol of NaOH are added and consumed, that means 0.0010 mol of the conjugate base ($\text{NaC}_2\text{H}_3\text{O}_2$) must have formed because the reaction is a 1:1 ratio!

Weak Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$).

To find the pH **AFTER** the addition of 10.0 mL of NaOH...

$$\frac{0.0015 \text{ mol}}{35.0 \text{ mL}} \times \frac{1000 \text{ mL}}{\text{L}} = 0.043 \text{ M [Acid]}$$

$$\frac{0.0010 \text{ mol}}{35.0 \text{ mL}} \times \frac{1000 \text{ mL}}{\text{L}} = 0.028 \text{ M } [\text{C}_2\text{H}_3\text{O}_2^-]$$

BUFFER

Weak Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$).

To find the pH **AFTER** the addition of 10.0 mL of NaOH...

$$\text{pH} = \text{p}K_a + \log \frac{[\text{Base}]}{[\text{Acid}]}$$

$$\text{p}K_a = -\log K_a$$

$$\text{p}K_a = -\log (1.8 \times 10^{-5})$$

$$\text{p}K_a = 4.74$$

Weak Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$).

To find the pH **AFTER** the addition of 10.0 mL of NaOH...

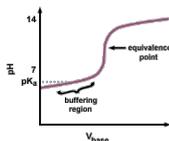
$$\text{pH} = \text{p}K_a + \log \frac{[\text{Base}]}{[\text{Acid}]}$$

$$\text{pH} = 4.74 + \log \frac{[0.028]}{[0.043]} = 4.55$$

Weak Acid / Strong Base

• Four Regions of the Curve:

- 1) Initial pH (initial [])
- 2) Between initial pH and equivalence point
- 3) Equivalence point
- 4) After equivalence point



• **ENDPOINT:** point where indicator changes color (usually close to the equivalence point)

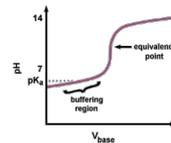
Weak Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$).

To find the pH at the **EQUIVALENCE POINT...**

When 25.0 mL of NaOH is added, all of the weak acid has been neutralized... BUT the conjugate base will hydrolyze to give a pH that is higher than 7



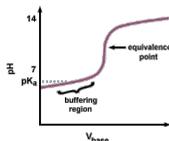
Weak Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$).

To find the pH at the **EQUIVALENCE POINT...**

- 1) Find the moles of NaOH added
- 2) Determine moles of weak base formed
- 3) Find the [] of weak base
- 4) Set up ICE table and equilibrium for the hydrolysis rxn... Need K_b since OH^- formed
- 5) Determine pH from pOH



Weak Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$).

To find the pH at the **EQUIVALENCE POINT...**

$$25.0 \text{ mL} \times \frac{\text{L}}{1000 \text{ mL}} \times \frac{0.10 \text{ mol}}{\text{L}} = 0.0025 \text{ mol NaOH added}$$

So 0.0025 mol of $\text{C}_2\text{H}_3\text{O}_2^-$ formed!

Weak Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$).

To find the pH at the **EQUIVALENCE POINT...**

$$\frac{0.0025 \text{ mol}}{50.0 \text{ mL}} \times \frac{1000 \text{ mL}}{\text{L}} = 0.050 \text{ M } [\text{C}_2\text{H}_3\text{O}_2^-]$$

Weak Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$).

To find the pH at the **EQUIVALENCE POINT...**

	$\text{C}_2\text{H}_3\text{O}_2^-$	$\text{HC}_2\text{H}_3\text{O}_2$	OH^-
Initial	0.050 M	0	0
Change			
Equilibrium			

Weak Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$).

To find the pH at the **EQUIVALENCE POINT...**

	$\text{C}_2\text{H}_3\text{O}_2^-$	$\text{HC}_2\text{H}_3\text{O}_2$	OH^-
Initial	0.050 M	0	0
Change	-x	+x	+x
Equilibrium			

Weak Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$).

To find the pH at the **EQUIVALENCE POINT...**

	$\text{C}_2\text{H}_3\text{O}_2^-$	$\text{HC}_2\text{H}_3\text{O}_2$	OH^-
Initial	0.050 M	0	0
Change	-x	+x	+x
Equilibrium	0.050 M - x M	x M	x M

$$K_a \times K_b = K_w$$

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

$$1.8 \times 10^{-5} \times K_b = 1 \times 10^{-14}$$

Weak Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$).

To find the pH at the **EQUIVALENCE POINT...**

	$\text{C}_2\text{H}_3\text{O}_2^-$	$\text{HC}_2\text{H}_3\text{O}_2$	OH^-
Initial	0.050 M	0	0
Change	-x	+x	+x
Equilibrium	0.050 M - x M	x M	x M

$$x = [\text{OH}^-] = 5.3 \times 10^{-6} \quad \text{pOH} = -\log [\text{OH}^-] = 5.28$$

Weak Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$).

To find the pH at the **EQUIVALENCE POINT...**

	$\text{C}_2\text{H}_3\text{O}_2^-$	$\text{HC}_2\text{H}_3\text{O}_2$	OH^-
Initial	0.050 M	0	0
Change	-x	+x	+x
Equilibrium	0.050 M - x M	x M	x M

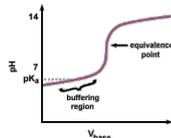
$$\text{pH} + \text{pOH} = 14$$

$$\text{pH} = 14 - 5.28 = 8.72$$

Weak Acid / Strong Base

• Four Regions of the Curve:

- 1) Initial pH (initial [])
- 2) Between initial pH and equivalence point
- 3) Equivalence point
- 4) After equivalence point



• **ENDPOINT:** point where indicator changes color (usually close to the equivalence point)

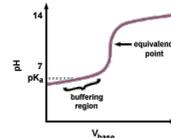
Weak Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$).

To find the pH **AFTER** the addition of 40.0 mL NaOH...

- 1) Find the moles of NaOH added
- 2) Find amount of moles of acid started
- 4) Determine the moles of excess NaOH
- 5) Get the [] of excess... pOH to get pH



Weak Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$).

To find the pH **AFTER** the addition of 40.0 mL NaOH...

$$40.0 \text{ mL} \times \frac{\text{L}}{1000 \text{ mL}} \times \frac{0.10 \text{ mol}}{\text{L}} = 0.0040 \text{ mol NaOH added}$$

$$25.0 \text{ mL} \times \frac{\text{L}}{1000 \text{ mL}} \times \frac{0.10 \text{ mol}}{\text{L}} = 0.0025 \text{ mol acid started}$$

Weak Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$).

To find the pH **AFTER** the addition of 40.0 mL NaOH...

$$0.0040 \text{ mol} - 0.0025 \text{ mol} = 0.0015 \text{ mol NaOH excess}$$

$$\frac{0.0015 \text{ mol}}{65.0 \text{ mL}} \times \frac{1000 \text{ mL}}{\text{L}} = 0.023 \text{ M } [\text{OH}^-]$$

$$\text{pOH} = -\log [\text{OH}^-]$$

$$-\log [0.023] = 1.64$$

Weak Acid / Strong Base

• TITRATION:

Addition of 0.10 M NaOH solution from a burette into a flask containing 25.0 mL of 0.10 M $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$).

To find the pH **AFTER** the addition of 40.0 mL NaOH...

$$\text{pH} + \text{pOH} = 14$$

$$\text{pH} = 14 - 1.64 = 12.36$$

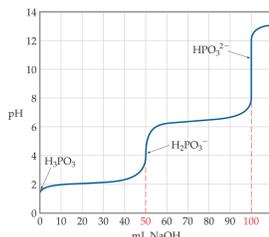
Weak Base / Strong Acid

If a **STRONG** acid is added to a **WEAK** base instead, just do the **OPPOSITE!**

Curve will look different too!!

Polyprotic Acid Titrations

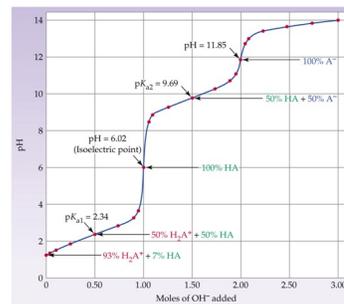
- Neutralization reaction occurs in steps due to **MULTIPLE** hydrogens
- Curve has multiple equivalence points... one for each H^+



Polyprotic Acid Titrations

**Don't forget...
pH = pKa at the
midpoint of each!**

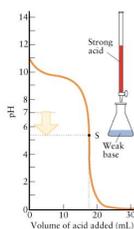
**Also, notice as K_a
gets smaller (acid
gets weaker)... pH
change at
equivalence point
is smaller!**



Summary

For Weak / Strong Titrations:

- pH initially is found by doing equilibrium problem for weak (ICE)
- If some weak still remains... determine $[]$ of weak and $[]$ of the conjugate and plug into HH
- At equivalence point, find $[]$ of conjugate salt formed and use ICE table with hydrolysis rxn
- If excess strong remains... use moles and $[]$ of excess to find pH



Practice

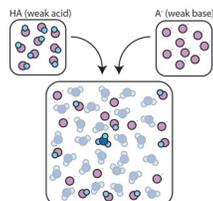
• EXAMPLE:

A solution is formed by adding 0.050 M NaOH to 40.0 mL of 0.025 M benzoic acid, $HC_7H_5O_2$, ($K_a = 6.3 \times 10^{-5}$). Calculate the pH a) of the initial solution *b) after 10.0 mL of NaOH added c) after 20.0 mL added and d) after 35.0 mL added

Buffers

Solutions of a weak acid or base and its conjugate (salt) that resist changes in pH

- As acid is added to the solution, the base component **neutralizes** it and vice versa
- The weak acid or base and its conjugate **CANNOT** consume each other
- pH remains stable even if solution is diluted with water or water is lost through evaporation (*WATCH THIS!)



Buffers

Solutions of a weak acid or base and its conjugate (salt) that resist changes in pH

• TWO WAYS TO MAKE:

1) Mix weak acid or base with salt of its conjugate

Ex: Add $HC_2H_3O_2$ to $NaC_2H_3O_2$

Add NH_3 to NH_4Cl



Buffers

Solutions of a weak acid or base and its conjugate (salt) that resist changes in pH

• TWO WAYS TO MAKE:

2) Add **SOME** strong acid or base to a weak base or acid to make the conjugate salt **OR** vice versa... **SOME** strong added to a salt to make the conjugate acid or base (just enough strong added to neutralize about half the weak)

Ex: Add some NaOH to $\text{HC}_2\text{H}_3\text{O}_2$ ($\text{C}_2\text{H}_3\text{O}_2^-$ forms and is with the remaining weak)

Add some HCl to NaF (HF forms and is with the remaining F^-)

Buffers

Solutions of a weak acid or base and its conjugate (salt) that resist changes in pH

• EXAMPLE:

Which of the following could be added to a solution of sodium acetate to produce a buffer?
Choose any that apply.

Acetic acid

Hydrochloric acid

Potassium acetate

Sodium chloride

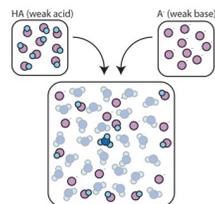
Buffers

Solutions of a weak acid or base and its conjugate (salt) that resist changes in pH

• Buffers are an example of the **common ion effect**

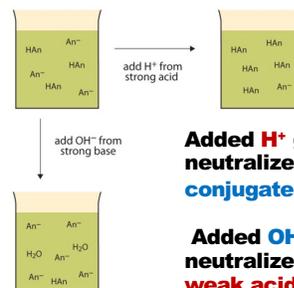
• Most effective or **IDEAL** when **EQUAL** amounts of weak acid or base and conjugate are used

• Resists change in pH until the conjugate acid or base is used up



Buffers

Solutions of a weak acid or base and its conjugate (salt) that resist changes in pH

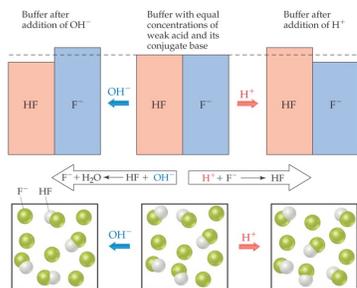


Added H^+ gets neutralized by the conjugate base!

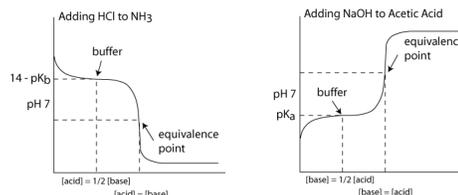
Added OH^- gets neutralized by the weak acid!

Buffers

Solutions of a weak acid or base and its conjugate (salt) that resist changes in pH



pH of a Buffer



• These are not new problems... can use **common ion effect** equilibrium or **HH** to solve

• Any weak acid or base titration is a buffer problem from the time the titrant is added until the equivalence point

Henderson-Hasselbalch Equation

$$\text{pH} = \text{pK}_a + \log \frac{[\text{Base}]}{[\text{Acid}]}$$

OR

$$\text{pOH} = \text{pK}_b + \log \frac{[\text{Acid}]}{[\text{Base}]}$$

*NOTE: Technically, HH becomes less accurate as the acid or base becomes **STRONGER** or the [] becomes **SMALLER**... This could affect some lab data in theory, so just be aware!

Henderson-Hasselbalch Equation

$$\text{pH} = \text{pK}_a + \log \frac{[\text{Base}]}{[\text{Acid}]}$$

OR

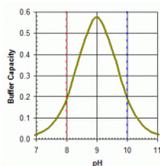
$$\text{pOH} = \text{pK}_b + \log \frac{[\text{Acid}]}{[\text{Base}]}$$

• **EXAMPLE:**

What is the pH of the buffer of a 0.12 M $\text{HC}_3\text{H}_5\text{O}_3$ and a 0.10 M $\text{NaC}_3\text{H}_5\text{O}_3$ where $K_a = 1.4 \times 10^{-4}$?

Buffer Capacity

Amount of acid or base the buffer can neutralize before the pH begins to change



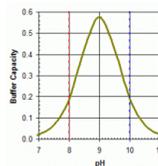
• Depends on the amount of acid and base in the buffer

• **INCREASE** the amount of acid and base... the **MORE RESISTANT** the buffer is to changes in pH

• pH range: buffers act best at the $\text{pH} = \text{pK}_a$ (or pK_b) of the weak acid / base because that is where the [] of the two components are equal... **IDEAL!**

Buffer Capacity

Amount of acid or base the buffer can neutralize before the pH begins to change



• Buffers with a **1:1** ratio of [conjugate]:[weak] have the greatest buffer capacity

• Maximum amount of acid that can be added to buffer = amount of conjugate base in the buffer and vice versa

• Buffer is still a buffer IF...

-pH within $\text{pK}_a \pm 1$

-[base]:[weak] = .10 - 10

Preparing a Buffer

• **EXAMPLE:**

How many moles of NH_4Cl must be added to 2.0 L of 0.10 M NH_3 ($K_b = 1.8 \times 10^{-5}$) to form a **buffer** whose pH is 9.00? (Assume that the addition of solid does not change the volume of the solution.)

$$\text{pOH} = 14 - \text{pH}$$

$$14 - 9.00 = 5.00$$

$$-\log K_b = \text{pK}_b$$

$$-\log 1.8 \times 10^{-5} = 4.74$$

Use HH with pOH and pK_b!!

Preparing a Buffer

• **EXAMPLE:**

How many moles of NH_4Cl must be added to 2.0 L of 0.10 M NH_3 ($K_b = 1.8 \times 10^{-5}$) to form a **buffer** whose pH is 9.00? (Assume that the addition of solid does not change the volume of the solution.)

$$\text{pOH} = \text{pK}_b + \log \frac{[\text{Acid}]}{[\text{Base}]}$$

$$5.00 = 4.74 + \log \frac{[\text{Acid}]}{[\text{0.10}]}$$

Preparing a Buffer

• **EXAMPLE:**

How many moles of NH_4Cl must be added to 2.0 L of 0.10 M NH_3 ($K_b = 1.8 \times 10^{-5}$) to form a **buffer** whose pH is 9.00? (Assume that the addition of solid does not change the volume of the solution.)

$$5.00 = 4.74 + \log \frac{[\text{Acid}]}{[0.10]}$$

$$0.26 = \log \frac{[\text{Acid}]}{[0.10]}$$

Preparing a Buffer

• **EXAMPLE:**

How many moles of NH_4Cl must be added to 2.0 L of 0.10 M NH_3 ($K_b = 1.8 \times 10^{-5}$) to form a **buffer** whose pH is 9.00? (Assume that the addition of solid does not change the volume of the solution.)

$$0.26 = \log \frac{[\text{Acid}]}{[0.10]}$$

$$10^{0.26} = \frac{[\text{Acid}]}{[0.10]}$$

Preparing a Buffer

• **EXAMPLE:**

How many moles of NH_4Cl must be added to 2.0 L of 0.10 M NH_3 ($K_b = 1.8 \times 10^{-5}$) to form a **buffer** whose pH is 9.00? (Assume that the addition of solid does not change the volume of the solution.)

$$10^{0.26} = \frac{[\text{Acid}]}{[0.10]}$$

$$1.82 = \frac{[\text{Acid}]}{[0.10]} \quad [\text{Acid}] = 0.18 \text{ M}$$

Preparing a Buffer

• **EXAMPLE:**

How many moles of NH_4Cl must be added to 2.0 L of 0.10 M NH_3 ($K_b = 1.8 \times 10^{-5}$) to form a **buffer** whose pH is 9.00? (Assume that the addition of solid does not change the volume of the solution.)

$$2.0 \text{ L} \times \frac{0.18 \text{ mol}}{\text{L}} = 0.36 \text{ mol } \text{NH}_4\text{Cl}$$

pH Changes in Buffers

• **EXAMPLE:**

A **buffer** is made by adding 0.30 mol $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$) and 0.30 mol $\text{NaC}_2\text{H}_3\text{O}_2$ to enough water to make 1.0 L of solution. The pH of the buffer is 4.74. Calculate the pH of this solution after 0.020 mol of NaOH is added.

0.30 mol acid - 0.020 mol base = 0.28 mol acid left over
 *Added **base** "attacks" **acid** part so subtracted

0.30 mol conj. + 0.020 mol base = 0.32 mol conj. left over
 ***Acid** reacting forms **more of the base** part so added

pH Changes in Buffers

• **EXAMPLE:**

A **buffer** is made by adding 0.30 mol $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$) and 0.30 mol $\text{NaC}_2\text{H}_3\text{O}_2$ to enough water to make 1.0 L of solution. The pH of the buffer is 4.74. Calculate the pH of this solution after 0.020 mol of NaOH is added.

$$\frac{0.28 \text{ mol}}{1.0 \text{ L}} = 0.28 \text{ M } [\text{HC}_2\text{H}_3\text{O}_2]$$

$$\frac{0.32 \text{ mol}}{1.0 \text{ L}} = 0.32 \text{ M } [\text{C}_2\text{H}_3\text{O}_2^-]$$

pH Changes in Buffers

• **EXAMPLE:**

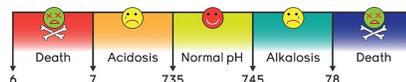
A **buffer** is made by adding 0.30 mol $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$) and 0.30 mol $\text{NaC}_2\text{H}_3\text{O}_2$ to enough water to make 1.0 L of solution. The pH of the buffer is 4.74. Calculate the pH of this solution after 0.020 mol of NaOH is added.

$$\text{pH} = \text{p}K_a + \log \frac{[\text{Base}]}{[\text{Acid}]}$$

$$\text{pH} = 4.74 + \log \frac{[0.32]}{[0.28]} = 4.80$$

Real World Buffers

Blood pH Levels



- Buffers are needed in the body to maintain pH levels in order for certain enzymes to work
- Human blood **MUST** maintain a pH between 7.35 - 7.45 in order to function normally
- pH falls below... **ACIDOSIS**
- pH rises above... **ALKALOSIS**

Practice

• **EXAMPLE:**

A chemist needs a solution buffered at pH 4.30 and can choose from the following acids and their salts. Which should be chosen?

Chloroacetic acid $K_a = 1.35 \times 10^{-3}$

Propanoic acid $K_a = 1.30 \times 10^{-5}$

Benzoic acid $K_a = 6.40 \times 10^{-5}$

Hypochlorous acid $K_a = 3.5 \times 10^{-8}$

Practice

• **EXAMPLES:**

What is the pH of a buffer solution that is composed of 0.45 M NH_4Cl and 0.15 M NH_3 ? ($K_b = 1.8 \times 10^{-5}$)

Calculate the concentration of sodium benzoate that must be present in 0.20 M $\text{HC}_7\text{H}_5\text{O}_2$ to produce a pH of 4.00. ($K_a = 6.4 \times 10^{-5}$)