Unit 6: ACIDS AND BASES

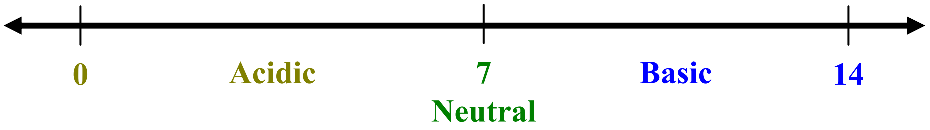
Chapter 16: Acids and Bases

16.1: Brønsted Acids and Bases

Physical and Chemical Properties of Acid and Base

Acids	Bases
Taste Sour (Citric Acids).	Taste Bitter.
Burning Sensation (Stomach Acid).	Feels Slippery (Detergent, Degreaser).
Corrosive with Metals (reacts to give off $H_{2\,(g)}$).	Alkaline in Nature (NaOH, Baking Soda).
Red litmus remains Red; Blue litmus turns Red.	Red litmus turns Blue; Blue litmus remains Blue.
Bromothymol Blue turns Yellow	Bromothymol Blue turns Blue.
Phenolphthalein turns Colourless.	Phenolphthalein turns Pink.
$pH < 7$	$pH > 7$

pH Scale



Conceptual Definition: - an explanation that attempts to describe why things are the way they are.

Arrhenius Concept: - **acids are H^+ (proton) producers** and **bases are OH^- producers**.

Examples: $HCl_{(aq)} \rightarrow \mathbf{H^+_{(aq)}} + Cl^-_{(aq)}$ ($HCl_{(aq)}$ is an Arrhenius Acid.)
$NaOH_{(aq)} \rightarrow Na^+_{(aq)} + \mathbf{OH^-_{(aq)}}$ ($NaOH_{(aq)}$ is an Arrhenius Base.)

Brønsted-Lowry Model: - acids and bases **react with water** to **dissociate** where **acids are H^+ (proton) donors** and **bases are H^+ (proton) acceptors**.
- first proposed by Johannes Brønsted and Thomas Lowry.

Hydronium Ion: - an ion formed when an **acid "donated" H^+ ion combined with a H_2O molecule** to form a **H_3O^+ ion (hydronium ion)**.
- essentially has the same function as a H^+ ion, but H_3O^+ denotes that we are using the Brønsted-Lowry model.

Examples: $HBr_{(aq)} + H_2O_{(l)} \rightarrow \mathbf{H_3O^+_{(aq)}} + Br^-_{(aq)}$
(HBr is a Brønsted-Lowry Acid – donated a proton)
(H_2O is a Brønsted-Lowry Base – accepted a proton.)

$NH_{3\,(aq)} + H_2O_{(l)} \rightleftharpoons NH_4^+{}_{(aq)} + \mathbf{OH^-_{(aq)}}$
(H_2O is a Brønsted-Lowry Acid – donated a proton)
(NH_3 is a Brønsted-Lowry Base – accepted a proton.)

Conjugate Base: - the product formed **after** the Acid donated a H^+. (**Acid → Conjugate Base**)
- behaves like a base when the reaction is looking from reverse.

Conjugate Acid: - the product formed **after** the Base accepted a H^+. (**Base → Conjugate Acid**)
- behaves like a acid when the reaction is looking from reverse.

Conjugate Acid-Base Pair: - the (**acid**/**conjugate base**) or (**base**/**conjugate acid**) pairs.

$$\underset{\text{(Acid)}}{HA_{(aq)}} + \underset{\text{(Base)}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{(Conjugate Acid)}}{H_3O^+_{(aq)}} + \underset{\text{(Conjugate Base)}}{A^-_{(aq)}}$$

Conjugate Acid-Base Pairs: HA/A^- and H_2O/H_3O^+

$$\underset{\text{(Base)}}{B_{(aq)}} + \underset{\text{(Acid)}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{(Conjugate Acid)}}{HB^+_{(aq)}} + \underset{\text{(Conjugate Base)}}{OH^-_{(aq)}}$$

Conjugate Acid-Base Pairs: B/HB^+ and H_2O/OH^-

Acid Dissociation Constant (**K_a**): - the equilibrium constant of a Brønsted-Lowry Acid Dissociation.

Base Dissociation Constant (**K_b**): - the equilibrium constant of a Brønsted-Lowry Base Dissociation.

Brønsted-Lowry Acid Dissociation	**Brønsted-Lowry Base Dissociation**
$HA_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + A^-_{(aq)}$	$B_{(aq)} + H_2O_{(l)} \rightleftharpoons HB^+_{(aq)} + OH^-_{(aq)}$
$K_a = \dfrac{[H_3O^+][A^-]}{[HA]}$	$K_b = \dfrac{[HB^+][OH^-]}{[B]}$

Example 1: Write the Brønsted-Lowry dissociation reaction of the following. Identify the Brønsted-Lowry acid and base, along with the conjugate acid and base. Determine the conjugate acid-base pairs. State the equilibrium expression of the dissociation reaction.

a. Acetic acid ($CH_3COOH_{(aq)}$)

$$\underset{\text{(Acid)}}{CH_3COOH_{(aq)}} + \underset{\text{(Base)}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{(Conjugate Acid)}}{H_3O^+_{(aq)}} + \underset{\text{(Conjugate Base)}}{CH_3COO^-_{(aq)}}$$

Conjugate Acid-Base Pairs:
CH_3COOH/CH_3COO^- and H_2O/H_3O^+

$$K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]}$$

b. Cyanide Ion ($CN^-_{(aq)}$)

$$\underset{\text{(Base)}}{CN^-_{(aq)}} + \underset{\text{(Acid)}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{(Conjugate Acid)}}{HCN_{(aq)}} + \underset{\text{(Conjugate Base)}}{OH^-_{(aq)}}$$

Conjugate Acid-Base Pairs:
CN^-/HCN and H_2O/OH^-

$$K_b = \frac{[HCN][OH^-]}{[CN^-]}$$

Assignment
16.1 pg. 568 #1 to 8

16.2: The Acid-Base Properties of Water

Autoionization of Water: - the process where water dissociates into hydronium and hydroxide ions.
- water's dissociation constant is called **ion-product constant** (**$K_w = 1.0 \times 10^{-14}$**)
- when the solution is **neutral (pH = 7)**, **$[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$ mol/L**

Autoionization of Water

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

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14} \text{ at } 25°\text{C}$$

$$[H_3O^+] = \frac{K_w}{[OH^-]} \qquad [OH^-] = \frac{K_w}{[H_3O^+]}$$

At pH = 7: $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$ mol/L

Example 1: At 25°C, $K_w = 1.0 \times 10^{-14}$.

a. Using the ICE box, show that $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$ mol/L for a neutral solution.

b. At 100°C, $K_w = 8.19 \times 10^{-13}$. What is the $[H_3O^+]$ and $[OH^-]$ for a neutral solution at 100°C?

a.

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

	$H_2O_{(l)}$		$[H_3O^+]$	$[OH^-]$
Initial	----		0	0
Change	----		$+x$	$+x$
Equilibrium	-----		x	x

$$K_w = [H_3O^+][OH^-]$$

$$1.0 \times 10^{-14} = (x)(x)$$

$$x^2 = \sqrt{1.0 \times 10^{-14}}$$

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

$$\boxed{x = [H_3O^+] = [OH^-] = 1.0 \times 10^{-7}\ \text{M}}$$

b. At 100°C, using $K_w = 8.19 \times 10^{-13}$

$$K_w = [H_3O^+][OH^-]$$

$$8.19 \times 10^{-13} = (x)(x)$$

$$x^2 = \sqrt{8.19 \times 10^{-13}}$$

$$x = 2.86 \times 10^{-7}$$

$$\boxed{x = [H_3O^+] = [OH^-] = 2.86 \times 10^{-7}\ \text{M}}$$

Example 2: Determine the $[H_3O^+]$ and/or $[OH^-]$ concentrations of the following solutions at 25°C.

a. $[OH^-] = 1.0 \times 10^{-4}$ mol/L

$$K_w = [H_3O^+][OH^-]$$

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-4}}$$

$$\boxed{[H_3O^+] = 1.0 \times 10^{-10}\ \text{mol/L}}$$

b. $[H_3O^+] = 5.0 \times 10^{-5}$ M

$$K_w = [H_3O^+][OH^-]$$

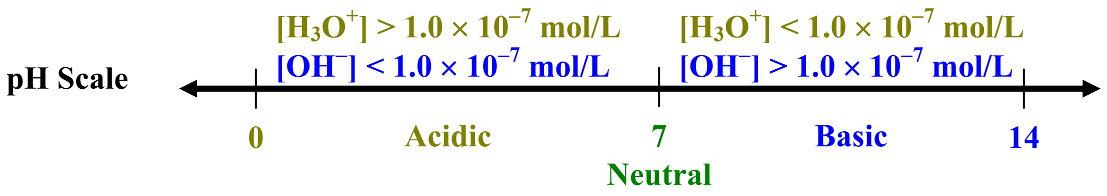
$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0 \times 10^{-14}}{5.0 \times 10^{-5}}$$

$$\boxed{[OH^-] = 2.0 \times 10^{-10}\ \text{M}}$$

16.3: pH – A Measure of Acidity

pH Scale: - a logarithmic scale to **measure the acidity (relative $[H_3O^+]$)** of a solution.

- the **lower the pH, the more acidic** (less basic) is the solution **(more $[H_3O^+]$ and less $[OH^-]$).**
- the **higher the pH, the more basic** (less acidic) is the solution **(less $[H_3O^+]$ and more $[OH^-]$).**
- **acidity is NOT the same as the acid strength**. Just because a solution has a low pH, it does not mean that it is a strong acid. **(Highly Acidic ≠ Strong Acid)**
- it is normally reported between 0 to 14 (**with 7 as neutral**), but it **can be above 14 (very basic)** or **below 0 (very acidic).**
- **an increase of 1 on a pH scale means a decrease of $[H_3O^+]$ by a factor of 10; an increase of 2 on a pH scale means a decrease of $[H_3O^+]$ by a factor of 100.**



Example: pH of Some Common Substances

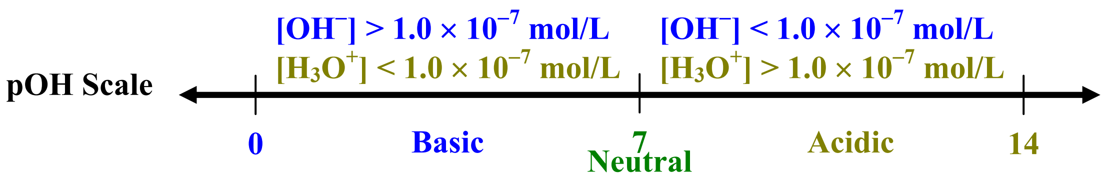
Substance	pH
1 M of HCl	0.00
Stomach Acid	2.00
Lemon Juice	2.50
Vinegar	3.00

Substance	pH
Milk	6.30
Rain Water	6.70
Pure Water	7.00
Blood	7.50

Substance	pH
1 M of Baking Soda ($NaHCO_3$)	9.68
Ammonia as Household Cleaner	12.00
1 M of NaOH	14.00

pOH Scale: - a logarithmic scale to **measure the basicity (relative $[OH^-]$)** of a solution.

- the **lower the pOH, the more basic** (less acidic) is the solution **(more $[OH^-]$ and less $[H_3O^+]$).**
- the **higher the pOH, the less basic** (more acidic) is the solution **(less $[OH^-]$ and more $[H_3O^+]$).**
- **basicity is NOT the same as the base strength**. Just because a solution has a low pOH, it does not mean that it is a strong base. **(Highly Basic ≠ Strong Base)**
- it is normally reported between 0 to 14 (**with 7 as neutral**), but it **can be above 14 (very acidic)** or **below 0 (very basic).**
- **an increase of 1 on a pOH scale means a decrease of $[OH^-]$ by a factor of 10; an increase of 2 on a pOH scale means a decrease of $[OH^-]$ by a factor of 100.**



pH and pOH Scales

$$pH = -\log [H_3O^+] \qquad pOH = -\log [OH^-]$$

$$pH + pOH = 14.00$$

Example 1: Calculate the pH and the pOH for the following solutions.

a. 0.0100 mol/L of $HI_{(aq)}$ (completely dissociates)

$$HI_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + I^-_{(aq)}$$

(Strong Acid: $[HI]_0 = [H_3O^+] = 0.0100$ M)

$pH = -\log [H_3O^+]$

$pH = -\log(0.0100)$

$\boxed{pH = 2.00}$

$pOH = 14 - pH$

$pOH = 14 - 2.00$

$\boxed{pOH = 12.00}$

b. 5.00×10^{-6} M of $HClO_{4\,(aq)}$ (completely dissociates)

$$HClO_{4\,(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + ClO_4^-{}_{(aq)}$$

(Strong Acid: $[HClO_4]_0 = [H_3O^+] = 5.00 \times 10^{-6}$ M)

$pH = -\log [H_3O^+]$

$pH = -\log(5.00 \times 10^{-6})$

$\boxed{pH = 5.30}$

$pOH = 14 - pH$

$pOH = 14 - 5.30$

$\boxed{pOH = 8.70}$

c. 1.50 M of $KOH_{(aq)}$

$$KOH_{(aq)} \rightarrow K^+_{(aq)} + OH^-_{(aq)}$$

(Ionic Strong Base: $[KOH]_0 = [OH^-] = 1.50$ M)

$pOH = -\log [OH^-]$

$pOH = -\log(1.50)$

$\boxed{pOH = -0.176}$

$pH = 14 - pOH$

$pH = 14 - (-0.176)$

$\boxed{pH = 14.176}$

d. 1.00×10^{-3} mol/L of $Ba(OH)_{2\,(aq)}$

$$Ba(OH)_{2\,(aq)} \rightarrow Ba^{2+}_{(aq)} + 2\,OH^-_{(aq)}$$

(Ionic Strong Base:

$[OH^-] = 2[Ba(OH)_2] = 2.00 \times 10^{-3}$ M)

$pOH = -\log [OH^-]$

$pOH = -\log(2.00 \times 10^{-3})$

$\boxed{pOH = 2.70}$

$pH = 14 - pOH$

$pH = 14 - (2.70)$

$\boxed{pH = 11.30}$

Example 2: Calculate the $[H_3O^+]$ and the $[OH^-]$ for the following solutions.

a. pH = 4.00

$pH = -\log [H_3O^+]$

$[H_3O^+] = 10^{-pH} = 10^{-4.00}$

$\boxed{[H_3O^+] = 1.00 \times 10^{-4}\text{ M}}$

$pOH = 14 - pH$

$pOH = 14 - 4.00 = 10.00$

$pOH = -\log [OH^-]$

$[OH^-] = 10^{-pOH} = 10^{-10.00}$

$\boxed{[OH^-] = 1.00 \times 10^{-10}\text{ M}}$

b. pOH = 3.00

$pOH = -\log [OH^-]$

$[OH^-] = 10^{-pOH} = 10^{-3.00}$

$\boxed{[OH^-] = 1.00 \times 10^{-3}\text{ M}}$

$pH = 14 - pOH$

$pH = 14 - 3 = 11.00$

$pH = -\log [H_3O^+]$

$[H_3O^+] = 10^{-pH} = 10^{-11.00}$

$\boxed{[H_3O^+] = 1.00 \times 10^{-11}\text{ M}}$

c. pH = 12.83

$pH = -\log [H_3O^+]$

$[H_3O^+] = 10^{-pH} = 10^{-12.83}$

$\boxed{[H_3O^+] = 1.48 \times 10^{-13}\text{ M}}$

$K_w = [H_3O^+][OH^-]$

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0\times10^{-14}}{1.48\times10^{-13}}$$

$\boxed{[OH^-] = 0.0676\text{ M}}$

d. pOH = 9.67

$pOH = -\log [OH^-]$

$[OH^-] = 10^{-pOH} = 10^{-9.67}$

$\boxed{[OH^-] = 2.14 \times 10^{-10}\text{ M}}$

$K_w = [H_3O^+][OH^-]$

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0\times10^{-14}}{2.14\times10^{-10}}$$

$\boxed{[H_3O^+] = 4.68 \times 10^{-5}\text{ mol/L}}$

Assignment

16.2 & 16.3 pg. 568–569 #9 to 24

16.4: Strength of Acids and Bases

Strong Acids: - acids that **dissociate completely (100%) in water**.
- when the **acid dissociation constant (K_a) is much greater than 1.** **($K_a >> 1$)**
- **the equilibrium position strongly favours the products.**
- at equilibrium, the original acid concentration, $[HA]_{eq} \approx 0$; $[H_3O^+]_{eq} = [A^-]_{eq} = [HA]_0$.
- the **conjugate base, A^-**, of a strong acid, HA **is itself a weak base** (cannot easily accept protons to do the reverse reaction).
- ***Note:* Strong Acids DO NOT MEAN that they are VERY CORROSIVE. It is the $[H_3O^+]$ that defines acidity.**

Strong Acid ($K_a >> 1$)

$$K_a = \frac{[H_3O^+][A^-]}{\cancel{[HA]}\rightarrow} \; 0\text{ M}$$

$$\underset{\text{(Strong Acid)}}{HA_{(aq)}} + \underset{\text{(Base)}}{H_2O_{(l)}} \overset{\longrightarrow}{\rightleftharpoons} \underset{\text{(Conjugate Acid)}}{H_3O^+_{(aq)}} + \underset{\text{(Conjugate Base-Weak)}}{A^-_{(aq)}}$$

	$[HA]$		$[H_3O^+]$	$[A^-]$
Initial	x		0	0
Change	$-x$		$+x$	$+x$
Equilibrium	0		x	x

Examples: Strong Acids: $HClO_{4\,(aq)}$, $HI_{(aq)}$, $HBr_{(aq)}$, $HCl_{(aq)}$, $H_2SO_{4\,(aq)}$ and $HNO_{3\,(aq)}$

Example 2: Write the dissociation reaction of 0.250 M of $HBr_{(aq)}$ and determine its $[H_3O^+]$.

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

(Strong Acid means $[HBr]_0 = [H_3O^+]$)

$[HBr]_0 = 0.250$ M $[H_3O^+] = 0.250$ M

Weak Acids: - acids that **dissociate LESS than 100% in water**.
- when the **acid dissociation constant (K_a) is less than 1.** **($K_a < 1$)**
- **the equilibrium position strongly favours the reactants.**
- at equilibrium, the hydronium concentration is much less than the original acid concentration, $[HA]_{eq} > [H_3O^+]_{eq}$ or $[HA]_0 \approx [HA]_{eq}$.
- the **conjugate base, A^-**, of a weak acid **is itself a stronger weak base** (can easily accept protons to do the reverse reaction).
- ***Note:* Weak Acids DO NOT MEAN that they are NOT CORROSIVE. It is the $[H_3O^+]$ that defines acidity. At a high enough concentration, a weak acid can be corrosive.**

$$\underset{\text{(Weak Acid)}}{HA_{(aq)}} + \underset{\text{(Base)}}{H_2O_{(l)}} \overset{\longleftarrow}{\rightleftharpoons} \underset{\text{(Conjugate Acid)}}{H_3O^+_{(aq)}} + \underset{\text{(Conjugate Base-Stronger)}}{A^-_{(aq)}}$$

	$[HA]$		$[H_3O^+]$	$[A^-]$
Initial	x		0	0
Change	$-y$ (where $y << x$)		$+y$	$+y$
Equilibrium	$(x - y) \approx x$		y	y

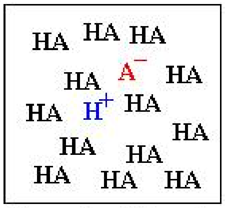
Examples: Some Weak Acids: $HOOCCOOH_{(aq)}$, $H_2SO_{3\,(aq)}$, $HSO_4^-{}_{(aq)}$, $H_3PO_{4\,(aq)}$, $HNO_{2\,(aq)}$, $H_3C_6H_5O_{7\,(aq)}$, $HF_{(aq)}$, $HCOOH_{(aq)}$, $C_6H_8O_{6\,(aq)}$, $C_6H_5COOH_{(aq)}$, $CH_3COOH_{(aq)}$, $H_2CO_{3\,(aq)}$, $H_2S_{(aq)}$, $HOCl_{(aq)}$, $HCN_{(aq)}$, $NH_4^+{}_{(aq)}$, and $H_3BO_{3\,(aq)}$

K_a and Relative Strength of Some Common Acids and Bases at 25°C

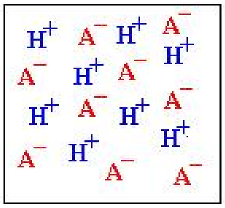
Increasing Acid Strength

Increasing Base Strength

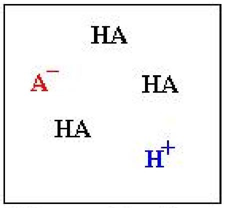
Acid Name	Acid Formula	Conjugate Base Formula	K_a
perchloric acid	$\mathrm{HClO_4}_{(aq)}$	$\mathrm{ClO_4^{-}}_{(aq)}$	Very Large
hydroiodic acid	$\mathrm{HI}_{(aq)}$	$\mathrm{I^{-}}_{(aq)}$	Very Large
hydrobromic acid	$\mathrm{HBr}_{(aq)}$	$\mathrm{Br^{-}}_{(aq)}$	Very Large
hydrochloric acid	$\mathrm{HCl}_{(aq)}$	$\mathrm{Cl^{-}}_{(aq)}$	Very Large
sulfuric acid	$\mathrm{H_2SO_4}_{(aq)}$	$\mathrm{HSO_4^{-}}_{(aq)}$	Very Large
nitric acid	$\mathrm{HNO_3}_{(aq)}$	$\mathrm{NO_3^{-}}_{(aq)}$	Very Large
hydronium ion	$\mathrm{H_3O^{+}}_{(aq)}$	$\mathrm{H_2O}_{(l)}$	1
oxalic acid	$\mathrm{HOOCCOOH}_{(aq)}$	$\mathrm{HOOCCOO^{-}}_{(aq)}$	6.5×10^{-2}
sulfurous acid	$\mathrm{H_2SO_3}_{(aq)}$	$\mathrm{HSO_3^{-}}$	1.5×10^{-2}
hydrogen sulfate ion	$\mathrm{HSO_4^{-}}_{(aq)}$	$\mathrm{SO_4^{2-}}$	1.2×10^{-2}
chlorous acid	$\mathrm{HClO_2}_{(aq)}$	$\mathrm{ClO_2^{-}}_{(aq)}$	1.2×10^{-2}
phosphoric acid	$\mathrm{H_3PO_4}_{(aq)}$	$\mathrm{H_2PO_4^{-}}_{(aq)}$	7.5×10^{-3}
arsenic acid	$\mathrm{H_3AsO_4}_{(aq)}$	$\mathrm{H_2AsO_4^{-}}_{(aq)}$	5×10^{-3}
monochloracetic acid	$\mathrm{HC_2H_2ClO_2}$	$\mathrm{C_2H_2ClO_2^{-}}_{(aq)}$	1.35×10^{-3}
citric acid	$\mathrm{H_3C_6H_5O_7}_{(aq)}$	$\mathrm{H_2C_6H_5O_7^{-}}_{(aq)}$	8.4×10^{-4}
hydrofluoric acid	$\mathrm{HF}_{(aq)}$	$\mathrm{F^{-}}_{(aq)}$	7.2×10^{-4}
nitrous acid	$\mathrm{HNO_2}_{(aq)}$	$\mathrm{NO_2^{-}}_{(aq)}$	4.0×10^{-4}
methanoic (formic) acid	$\mathrm{HCOOH}_{(aq)}$	$\mathrm{HCOO^{-}}_{(aq)}$	1.8×10^{-4}
lactic acid	$\mathrm{HC_3H_5O_3}_{(aq)}$	$\mathrm{C_3H_5O_3^{-}}_{(aq)}$	1.38×10^{-4}
ascorbic acid (vitamin C)	$\mathrm{H_2C_6H_6O_6}_{(aq)}$	$\mathrm{HC_6H_6O_6^{-}}_{(aq)}$	7.9×10^{-5}
benzoic acid	$\mathrm{C_6H_5COOH}_{(aq)}$	$\mathrm{C_6H_5COO^{-}}_{(aq)}$	6.4×10^{-5}
hydrogen oxalate ion	$\mathrm{HOOCCOO^{-}}_{(aq)}$	$\mathrm{OOCCOO^{2-}}_{(aq)}$	6.1×10^{-5}
ethanoic (acetic) acid	$\mathrm{CH_3COOH}_{(aq)}$	$\mathrm{CH_3COO^{-}}_{(aq)}$	1.8×10^{-5}
dihydrogen citrate ion	$\mathrm{H_2C_6H_5O_7^{-}}_{(aq)}$	$\mathrm{HC_6H_5O_7^{2-}}_{(aq)}$	1.8×10^{-5}
hydrated aluminum (III) ion	$\mathrm{[Al(H_2O)_6]^{3+}}$	$\mathrm{[AlOH(H_2O)_5]^{2+}}$	1.4×10^{-5}
propanoic acid	$\mathrm{C_2H_5COOH}_{(aq)}$	$\mathrm{C_2H_5COO^{-}}_{(aq)}$	1.3×10^{-5}
hydrogen citrate ion	$\mathrm{HC_6H_5O_7^{2-}}_{(aq)}$	$\mathrm{C_6H_5O_7^{3-}}_{(aq)}$	4.0×10^{-6}
carbonic acid	$\mathrm{H_2CO_3}_{(aq)}$	$\mathrm{HCO_3^{-}}_{(aq)}$	4.3×10^{-7}
hydrosulfuric acid	$\mathrm{H_2S}_{(aq)}$	$\mathrm{HS^{-}}_{(aq)}$	1.0×10^{-7}
hydrogen sulfite ion	$\mathrm{HSO_3^{-}}_{(aq)}$	$\mathrm{SO_3^{2-}}_{(aq)}$	1.0×10^{-7}
dihydrogen arsenate ion	$\mathrm{H_2AsO_4^{-}}_{(aq)}$	$\mathrm{HAsO_4^{2-}}_{(aq)}$	8×10^{-8}
dihydrogen phosphate ion	$\mathrm{H_2PO_4^{-}}_{(aq)}$	$\mathrm{HPO_4^{2-}}_{(aq)}$	6.2×10^{-8}
hypochlorous acid	$\mathrm{HOCl}_{(aq)}$	$\mathrm{OCl^{-}}_{(aq)}$	3.5×10^{-8}
hypobromous acid	$\mathrm{HOBr}_{(aq)}$	$\mathrm{OBr^{-}}_{(aq)}$	2×10^{-9}
hydrocyanic acid	$\mathrm{HCN}_{(aq)}$	$\mathrm{CN^{-}}_{(aq)}$	6.2×10^{-10}
hydrogen arsenate ion	$\mathrm{HAsO_4^{2-}}_{(aq)}$	$\mathrm{AsO_4^{3-}}_{(aq)}$	6×10^{-10}
boric acid	$\mathrm{H_3BO_3}_{(aq)}$	$\mathrm{H_2BO_3^{-}}_{(aq)}$	5.8×10^{-10}
ammonium ion	$\mathrm{NH_4^{+}}_{(aq)}$	$\mathrm{NH_3}_{(aq)}$	5.6×10^{-10}
Phenol	$\mathrm{C_6H_5OH}_{(aq)}$	$\mathrm{C_6H_5O^{-}}_{(aq)}$	1.6×10^{-10}
hydrogen carbonate ion	$\mathrm{HCO_3^{-}}_{(aq)}$	$\mathrm{CO_3^{2-}}_{(aq)}$	5.6×10^{-11}
hypoiodous acid	$\mathrm{HOI}_{(aq)}$	$\mathrm{OI^{-}}_{(aq)}$	2×10^{-11}
hydrogen ascorbate ion	$\mathrm{HC_6H_6O_6^{-}}_{(aq)}$	$\mathrm{C_6H_6O_6^{2-}}_{(aq)}$	1.6×10^{-12}
hydrogen phosphate ion	$\mathrm{HPO_4^{2-}}_{(aq)}$	$\mathrm{PO_4^{3-}}_{(aq)}$	4.8×10^{-13}
water (55.49 mol/L)	$\mathrm{H_2O}_{(l)}$	$\mathrm{OH^{-}}_{(aq)}$	$1.0 \times 10^{-14} = K_w$
hydrogen sulfide ion	$\mathrm{HS^{-}}_{(aq)}$	$\mathrm{S^{2-}}_{(aq)}$	$\sim 10^{-19}$



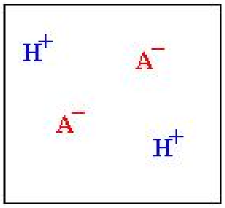
Concentrated weak acid - a lot present, but little dissociation of acid



Concentrated strong acid - a lot present with a lot of dissociation to form many hydrogen ions



Dilute weak acid - little acid present with little dissciation of acid



Dilute strong acid - much acid present with a high degree of dissociation

Check out Strong and Weak Acids Animations at
(http://www.sgc.peachnet.edu/users/larnold/WWW/courses/1212/rev1212.html)

Relative Strength of Acids and Conjugate Bases:

1. **The stronger the acid (the bigger the value of K_a), the weaker its conjugate base.**
2. **The weaker the acid (the smaller the value of K_a), the stronger its conjugate base.**

Example 2: Order the following acids from the strongest to the weakest.

$HNO_{2\,(aq)}$, $H_2SO_{3\,(aq)}$, $HClO_{4\,(aq)}$, $HF_{(aq)}$

According to the Acid and Bases Relative Strength Table on the previous page:

$HClO_{4\,(aq)} >> H_2SO_{3\,(aq)} > HF_{(aq)} > HNO_{2\,(aq)}$

$(K_a\text{: very large} >> 1.5 \times 10^{-2} > 7.2 \times 10^{-4} > 4.0 \times 10^{-4})$

Example 3: Order the following conjugate base from the strongest to the weakest.

$CN^-_{(aq)}$, $SO_4^{2-}{}_{(aq)}$, $Cl^-_{(aq)}$, $CH_3COO^-_{(aq)}$

According to the Acids and Bases Relative Strength Table on the previous page:

$CN^-_{(aq)} > CH_3COO^-_{(aq)} > SO_4^{2-}{}_{(aq)} >> Cl^-_{(aq)}$

$(K_a\text{: } 6.2 \times 10^{-10} < 1.8 \times 10^{-5} < 1.2 \times 10^{-2} << \text{Very Large})$

$HCN_{(aq)} < CH_3COOH_{(aq)} < HSO_{4\,(aq)} << HCl_{(aq)}$

Monoprotic Acids: - acids that can donate a maximum of one proton.

Example 3: Write the dissociation reaction for the following monoprotic acids.

a. $HI_{(aq)}$

$HI_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + I^-_{(aq)}$

Direct Arrow because $HI_{(aq)}$ is a Strong Acid (K_a is very large)

b. $HCOOH_{(aq)}$

$HCOOH_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + HCOO^-_{(aq)}$

Double Arrow because $HCOOH_{(aq)}$ is a Weak Acid ($K_a = 1.8 \times 10^{-4} << 1$)

Diprotic Acids: - acids that can donate a maximum of two protons in stepwise dissociation.

$$H_2A_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + HA^-_{(aq)} \qquad K_{a1} = \frac{[H_3O^+][HA^-]}{[H_2A]}$$

(Original Diprotic Acid) (Base) (Conjugate Acid) (Conjugate Base)

$$HA^-_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + A^{2-}_{(aq)} \qquad K_{a2} = \frac{[H_3O^+][A^{2-}]}{[HA^-]}$$

(Acid) (Base) (Conjugate Acid) (Conjugate Base)

(Conjugate Base of Original Diprotic Acid)

Example 4: Write the stepwise dissociation reaction for the following diprotic acids.

a. $H_2SO_{4\,(aq)}$

$$H_2SO_{4\,(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + HSO_4^-{}_{(aq)}$$

Direct Arrow because $H_2SO_{4\,(aq)}$ is a Strong Acid (K_{a1} is very large)

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

Double Arrow because $HSO_4^-{}_{(aq)}$ is a Weak Acid ($K_{a2} = 1.2 \times 10^{-2} << 1$)

b. $HOOCCOOH_{(aq)}$

$$HOOCCOOH_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + HOOCCOO^-{}_{(aq)}$$

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

Double Arrow because both $HOOCCOOH_{(aq)}$ and $HOOCCOO^-$ are Weak Acids (K_{a1} and $K_{a2} << 1$)

Amphoteric Substances: - chemical **species that can be an acid or a base**.
- all **intermediate species of a diprotic acid** is an amphoteric substance.

Examples: Some Amphoteric Substances: $HOOCCOO^-{}_{(aq)}$, $HSO_4^-{}_{(aq)}$, $HSO_3^-{}_{(aq)}$, $HCO_3^-{}_{(aq)}$, $HS^-{}_{(aq)}$, $HC_6H_6O_6^-{}_{(aq)}$, and **$H_2O_{(l)}$**

Calculating the pH of Strong Acid Solutions

Major Species: - the predominant species of an acid or a base after dissociation.
- in acid and base dissociation, because they are aqueous, **water ($H_2O_{(l)}$) is always listed as a major species**.

Major Species of a Strong Acid: - as strong acid dissociates completely ($K_a > 1$) in water, **the major species of all strong acids are H_3O^+ and their conjugate bases.**

$$\underset{\text{(Strong Acid)}}{HA_{(aq)}} + \underset{\text{(Base)}}{H_2O_{(l)}} \longrightarrow \underset{\text{(Conjugate Acid)}}{H_3O^+_{(aq)}} + \underset{\text{(Conjugate Base-Weak)}}{A^-_{(aq)}}$$

Example 5: List the major species $H_2SO_{4\,(aq)}$, and calculate its pH if it has a concentration of 1.00×10^{-5} M.

$$H_2SO_{4\,(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + HSO_4^-{}_{(aq)}$$

Since H_2SO_4 is a strong acid, the major species are: **$H_3O^+_{(aq)}$, $HSO_4^-{}_{(aq)}$ and $H_2O_{(l)}$**

(Even if H_2SO_4 is a **diprotic acid**, it **dissociates one proton at a time**. The conjugate base, HSO_4^-, is a weak acid. Weak acids dissociate differently than strong acids – next section.)

$[H_3O^+] = [H_2SO_4]_0 = 1.00 \times 10^{-5}$ M

$pH = -\log [H_3O^+]$

$pH = -\log (1.00 \times 10^{-5})$ **$pH = 5.00$**

Example 6: Determine the $[H_3O^+]$ and/or $[OH^-]$ concentrations of the following solutions at 25°C.

a. [HCl] = 0.350 mol/L

$HCl_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$

(Strong Acid: $[HCl]_0 = [H_3O^+] = 0.350$ M)

$$K_w = [H_3O^+][OH^-]$$

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0\times10^{-14}}{0.350}$$

$$\boxed{[OH^-] = 2.86 \times 10^{-14} \text{ mol/L}}$$

b. [NaOH] = 0.0750 M

$NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$

(Ionic Strong Base: $[NaOH]_0 = [OH^-] = 0.0750$ M)

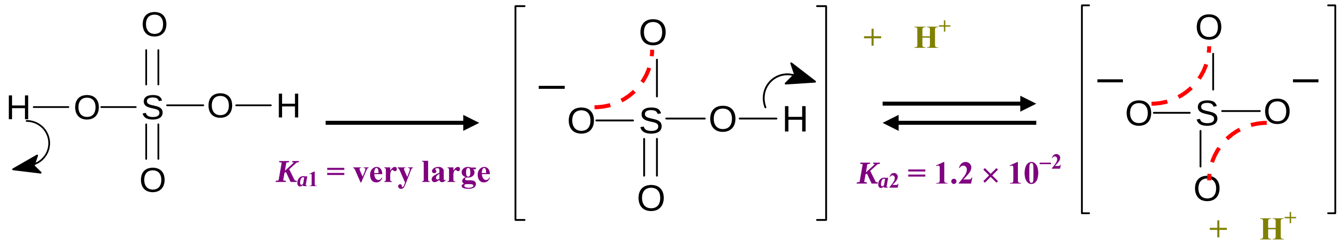
$$K_w = [H_3O^+][OH^-]$$

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0\times10^{-14}}{0.0750}$$

$$\boxed{[H_3O^+] = 1.33 \times 10^{-13} \text{ mol/L}}$$

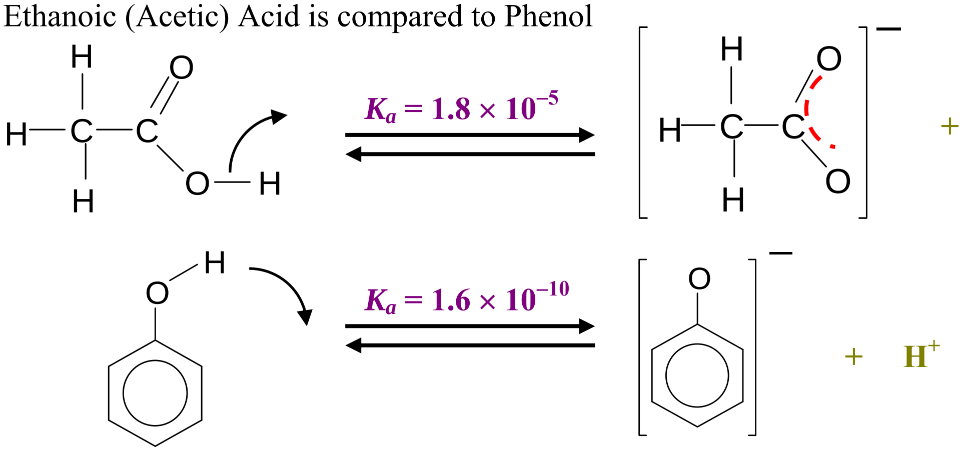
Oxoacids: - acids where the donating proton is attached to an oxygen atom.
- most acids are oxoacids because of the strong electronegativity of the oxygen atom, the hydrogen atom is more readily to leave as H^+ ion.

Example: Sulfuric Acid ($H_2SO_{4\,(aq)}$), successively donate protons to finally reach $SO_4^{2-}{}_{(aq)}$



Organic Acids: - acids of organic compounds, commonly known as carboxylic acids.
- alcohols can also be considered as organic acids, but they are really weak. This is because carboxylic acids have double-bonded oxygen for resonance to occur, whereas alcohols only contain single-bonded oxygen.

Example: Ethanoic (Acetic) Acid is compared to Phenol



Assignment

16.4 pg. 569 #25 to 36; pg. 571 #86

16.5: Weak Acids and Acid Ionization Constants

Major Species of a Weak Acid: - since weak acid do not dissociate completely ($K_a < 1$) in water, **the major species of all weak acids are their original form and water.**

Procedure to calculate pH of Weak Acid Solutions:

1. **List all the major species from all weak acids, including water.**
2. **Determine which species has the highest K_a.** This is the **Strongest Acid (SA)** of the list.
3. **Write the Brønsted-Lowry dissociation of this Strongest Acid.**
4. Set up the **ICE Box** and the **equilibrium expression**.
5. If the **original weak acid concentration, $[HA]_0$, is much larger than K_a**, we can **approximate by assuming $[HA]_{eq} = ([HA]_0 - x) \approx [HA]_0$**. Thereby, simplifying the calculation. **(A general rule of thumb: if $[HA]_0 \geq 1000 \times K_a$, we can use the approximation.)**
6. **Find the concentration of H_3O^+.**
7. **Verify** any **approximation** made by using the **5% rule**. $$\left(\frac{[H_3O^+]}{[HA]_0} \times 100\% \leq 5\%\right)$$

Example 1: Determine the $[H_3O^+]$, and pH of 0.200 mol/L of $HF_{(aq)}$ ($K_a = 7.2 \times 10^{-4}$).

Major Species: **$HF_{(aq)}$, $H_2O_{(l)}$**
Strongest Acid: **HF ($K_a = 7.2 \times 10^{-4}$)** (K_a for H_2O is 1.0×10^{-14})
HF is a weak acid and undergoes Brønsted-Lowry Dissociation.

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

	[HF]		$[H_3O^+]$	$[F^-]$
Initial	0.200 M		0	0
Change	$-x$		$+x$	$+x$
Equilibrium	$(0.2 - x)$		x	x

$$K_a = \frac{[H_3O^+][F^-]}{[HF]} \qquad 7.2 \times 10^{-4} = \frac{(x)(x)}{(0.2-x)} = \frac{x^2}{(0.2-x)}$$

$$7.2 \times 10^{-4}\,(0.2 - x) = x^2$$

$$(1.44 \times 10^{-4}) - (7.2 \times 10^{-4})\,x = x^2$$

$$x^2 + (7.2 \times 10^{-4})\,x - (1.44 \times 10^{-4}) = 0$$

(Quadratic Equation: Apply the Quadratic Formula!)

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} \qquad a = 1 \qquad b = 7.2 \times 10^{-4} \qquad c = -1.44 \times 10^{-4}$$

$$x = \frac{-(7.2\times10^{-4}) \pm \sqrt{(7.2\times10^{-4})^2 - 4(1)(-1.44\times10^{-4})}}{2(1)}$$

$x = 0.0116453988$ $x = -0.0247307976$ (omit negative x)

$[H_3O^+] = 0.0116$ mol/L

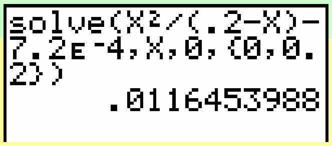
pH = – log $[H_3O^+]$
pH = – log(0.0116)

pH = 1.93

CANNOT use Approximation:

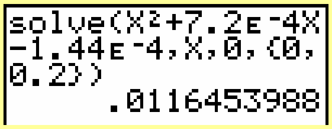
$$\frac{[HF]_0}{K_a} = \frac{0.200\text{ M}}{7.2\times10^{-4}} = \mathbf{277.8 < 1000}$$

Have to use $(0.2 - x)$ in the denominator



```
solve(X²/(.2-X)-
7.2ᴇ-4,X,0,{0,0.
2})
        .0116453988
```

OR



```
solve(X²+7.2ᴇ-4X
-1.44ᴇ-4,X,0,{0,
0.2})
        .0116453988
```

Verify that we could NOT use Approximation:

$$\frac{[H_3O^+]}{[HF]_0} \times 100\% = \frac{0.0116\text{ M}}{0.200\text{ M}} \times 100\% = \mathbf{5.8\% > 5\%}$$

Therefore, approximation would NOT be appropriate.

Percent Dissociation: - the amount of $[H_3O^+]$ dissociated from the original $[HA]_0$ expressed in percentage.
- strong acids will have % dissociation = 100%.
- weak acids will have % Dissociation < 100%

Percent Dissociation of Acids

$$\textbf{\% Dissociation} = \frac{[H_3O^+]}{[HA]_0} \times 100\%$$

Example 2: Determine the $[H_3O^+]$ and pH of an acid mixture of 0.500 mol/L $HOBr_{(aq)}$ ($K_a = 2 \times 10^{-9}$) of 0.200 mol/L of $CH_3COOH_{(aq)}$ ($K_a = 1.8 \times 10^{-5}$). Calculate the % dissociation of this acid mixture.

Major Species: $HOBr_{(aq)}$, $CH_3COOH_{(aq)}$, $H_2O_{(l)}$

Strongest Acid: CH_3COOH ($K_a = 1.8 \times 10^{-5}$) (K_a for HOBr and H_2O are 2×10^{-9} and 1.0×10^{-14})

CH_3COOH is a weak acid and undergoes Brønsted-Lowry Dissociation.

$$CH_3COOH_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + CH_3COO^-_{(aq)}$$

	$[CH_3COOH]$		$[H_3O^+]$	$[CH_3COO^-]$
Initial	0.200 M		0	0
Change	$-x$		$+x$	$+x$
Equilibrium	$(0.2 - x)$		x	x

CAN use Approximation:

$$\frac{[CH_3COOH]_0}{K_a} = \frac{0.200\text{ M}}{1.8 \times 10^{-5}}$$

$$= 11111 \geq 1000$$

Use 0.2 in the denominator, because $(0.2 - x) \approx 0.2$ [x is so small compared to 0.2 M]

$$K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]} \qquad 1.8 \times 10^{-5} = \frac{(x)(x)}{(0.2 - x)} \approx \frac{x^2}{(0.2)}$$

$$1.8 \times 10^{-5}\,(0.2) \approx x^2$$

$$3.6 \times 10^{-6} \approx x^2$$

$$x \approx \sqrt{3.6 \times 10^{-6}}$$

$$x \approx 0.00190$$

$$\boxed{[H_3O^+] = 0.00190\text{ mol/L}}$$

$$\text{pH} = -\log[H_3O^+]$$

$$\text{pH} = -\log(0.00190)$$

$$\boxed{\text{pH} = 2.72}$$

Verify that we could use Approximation:

$$\frac{[H_3O^+]}{[CH_3COOH]_0} \times 100\% = \frac{0.00190\text{ M}}{0.200\text{ M}} \times 100\%$$

$$= 0.95\% \leq 5\%$$

Therefore, approximation would be appropriate.

$$\text{\% Dissociation} = \frac{[H_3O^+]}{[CH_3COOH]_0} \times 100\%$$

$$\text{\% Dissociation} = \frac{0.00190\text{ M}}{0.200\text{ M}} \times 100\%$$

$$\boxed{\text{\% Dissociation} = 0.95\%}$$

Example 3: A 0.0500 mol/L of an unknown acid, HA, has a percent dissociation of 0.38%. What is the acid dissociation constant of this acid?

Major Species: $HA_{(aq)}$, $H_2O_{(l)}$

Strongest Acid: HA (K_a = ?) (K_a for H_2O is 1.0×10^{-14} and has a 0.000 01 % dissociation)

HA is a weak acid and undergoes Brønsted-Lowry Dissociation.

$$HA_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + A^-_{(aq)}$$

	$[HA]$		$[H_3O^+]$	$[A^-]$
Initial	0.0500 M		0	0
Change	-1.9×10^{-4}		$+1.9 \times 10^{-4}$	1.9×10^{-4}
Equilibrium	0.04981 M		1.9×10^{-4} M	1.9×10^{-4} M

$$\text{\% Dissociation} = \frac{[H_3O^+]}{[HA]_0} \times 100\%$$

$$[H_3O^+] = (\text{\%Dissociation})[HA]_0 / 100\%$$

$$[H_3O^+] = (0.38\%)(0.0500\text{ M}) / 100\%$$

$$[H_3O^+] = 1.9 \times 10^{-4}\text{ M} = [A^-]$$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} = \frac{(1.9 \times 10^{-4})(1.9 \times 10^{-4})}{(0.04981)}$$

$$\boxed{K_a = 7.2 \times 10^{-7}}$$

Polyprotic Acids: - acids that can donate more than one protons.
- this includes all diprotic and **triprotic acids** (acids that can **donate three protons**).
- **polyprotic acids dissociate one proton at a time**. Each successive proton donation has its own K_a, which gets smaller until the last proton is donated. ($K_{a1} > K_{a2} > K_{a3} > \ldots$).
- the **intermediates (conjugate bases of each dissociation except the last one) are themselves acids**. Thus, *diprotic acid can generate one amphoteric species*, and *triprotic acid can generate two amphoteric species*.
- **except for sulfuric acid (because K_{a1} is very large and K_{a2} is small), all polyprotic acids' $[H_3O^+]$ are calculated from their K_{a1}** because successive proton donation from smaller subsequent K_a do not amount to any significant increase in the $[H_3O^+]$.

Examples: Some Polyprotic Acids:

a. Diprotic Acids: $H_2SO_{4\,(aq)}$, $HOOCCOOH_{(aq)}$, $H_2SO_{3\,(aq)}$, $H_2C_6H_6O_{6\,(aq)}$, $H_2CO_{3\,(aq)}$, $H_2S_{(aq)}$

Diprotic Amphoteric Intermediates: $HSO_4^-{}_{(aq)}$, $HOOCCOO^-{}_{(aq)}$, $HSO_3^-{}_{(aq)}$, $HC_6H_6O_6^-{}_{(aq)}$, $HCO_3^-{}_{(aq)}$, $HS^-{}_{(aq)}$

b. Triprotic Acids: $H_3PO_{4\,(aq)}$, $H_3AsO_{4\,(aq)}$, $H_3C_6H_5O_{7\,(aq)}$, $H_3BO_{3\,(aq)}$

Triprotic Amphoteric Intermediates: $H_2PO_4^-{}_{(aq)}$, $HPO_4^{2-}{}_{(aq)}$, $H_2AsO_4^-{}_{(aq)}$, $HAsO_4^{2-}{}_{(aq)}$, $H_2C_6H_5O_7^-{}_{(aq)}$, $HC_6H_5O_7^{2-}{}_{(aq)}$, $H_2BO_3^-{}_{(aq)}$, $HBO_3^{2-}{}_{(aq)}$

Example 1: Determine the pH of 0.0500 M of sulfuric acid ($H_2SO_{4\,(aq)}$) and the concentrations of $HSO_4^-{}_{(aq)}$, and $SO_4^{2-}{}_{(aq)}$. The acid dissociation constant is $K_{a2} = 1.2 \times 10^{-2}$.

$$H_2SO_{4\,(aq)} + H_2O_{(l)} \rightarrow H_3O^+{}_{(aq)} + HSO_4^-{}_{(aq)}$$

Since H_2SO_4 is a strong acid, the major species are: **$H_3O^+{}_{(aq)}$, $HSO_4^-{}_{(aq)}$ and $H_2O_{(l)}$.**

$[H_3O^+] = [H_2SO_4]_0 = 0.0500$ M

$\text{pH} = -\log\,[H_3O^+]$

$\text{pH} = -\log\,(0.0500)$

pH = 1.30 (after K_{a1})

Next, we have to calculate $[HSO_4^-]$, $[SO_4^{2-}]$ and the final pH after the last proton donation.

Major Species: $HSO_4^-{}_{(aq)}$, $H_2O_{(l)}$ **Strongest Acid: $HSO_4^-{}_{(aq)}$ ($K_{a2} = 1.2 \times 10^{-2}$)**

HSO_4^- is a weak acid (Brønsted-Lowry Dissociation).

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

	$[HSO_4^-]$		$[H_3O^+]$	$[SO_4^{2-}]$
Initial	**0.0500 M**		**0.0500 M**	**0**
Change	$-x$		$+x$	$+x$
Equilibrium	$(0.05 - x)$		$(0.05 + x)$	x

CANNOT use Approximation:

$$\frac{[HSO_4]_0}{K_{a2}} = \frac{0.0500\text{ M}}{1.2\times10^{-2}}$$

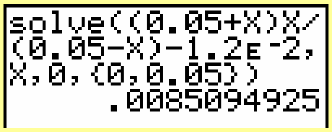
$$= 4.167 < 1000$$

$$K_a = \frac{[H_3O^+][SO_4^{2-}]}{[HSO_4^-]}$$

$$1.2 \times 10^{-2} = \frac{(0.05+x)(x)}{(0.05-x)}$$

$$0 = \frac{(0.05+x)(x)}{(0.05-x)} - 1.2 \times 10^{-2}$$

$$x = 0.00851$$



```
solve((0.05+X)X/
(0.05-X)-1.2ᴇ-2,
X,0,{0,0.05})
      .0085094925
```

Final $[H_3O^+]$ = 0.0500 M (from K_{a1}) + 0.00851 M (from K_{a2})

Final $[H_3O^+]$ = 0.0585 M
(a significant change from before)

$[HSO_4^-]_{eq}$ = 0.0500 M − 0.00851 M

$[HSO_4^-]_{eq}$ = 0.0415 M
$[SO_4^{2-}]_{eq}$ = 0.00585 M

Final pH = – log $[H_3O^+]$ = –log (0.0585)

Final pH = 1.23
(different from 1.30 of the pH at K_{a1})

From the last two examples, we can see that sulfuric acid ($H_2SO_{4\,(aq)}$) requires the calculation of $[H_3O^+]$ in both steps of the dissociation. Other polyprotic acids, such as $H_3C_6H_5O_{7\,(aq)}$ only require the first dissociation step to calculate the $[H_3O^+]$.

Assignment
16.5 pg. 569−570 #37 to 52; pg. 571 #93

16.6 & 16.7: Weak Bases and Base Ionization Constants & The Relationship Between the Ionization and their Conjugate Bases

Strong Bases: - bases that **dissociate completely (100%) in water**.
- all **alkali bases (Group IA cations with OH^-)** and some **alkaline bases (Group IIA cations with OH^-)** are considered as strong bases because they are ionic compound that dissociates completely.

Examples: Strong Ionic Bases:

a. Alkali Bases: $LiOH_{(aq)}$, $NaOH_{(aq)}$, $KOH_{(aq)}$, $RbOH_{(aq)}$, and $CsOH_{(aq)}$
(gives off 1 mole of OH^- when 1 mole of alkali base is dissolved)

$$NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)} \qquad KOH_{(aq)} \rightarrow K^+_{(aq)} + OH^-_{(aq)}$$

b. Alkaline Bases: $Ca(OH)_{2\,(aq)}$, $Ba(OH)_{2\,(aq)}$, $Sr(OH)_{2\,(aq)}$
(gives off 2 moles of OH^- when 1 mole of alkaline base is dissolved)

$$Ba(OH)_{2\,(aq)} \rightarrow Ba^{2+}_{(aq)} + 2\ OH^-_{(aq)} \qquad Sr(OH)_{2\,(aq)} \rightarrow Sr^{2+}_{(aq)} + 2\ OH^-_{(aq)}$$

Major Species of a Strong Base: - as alkali- and alkaline bases dissociate completely in water, **the major species of all strong bases are OH^- and water**.
- special care must be taken with **alkaline bases** as they **generate 2 moles of OH^- per 1 mole of solid dissolved.**

$$\underset{\text{(Strong Base)}}{MOH_{(aq)}} \longrightarrow \underset{\text{(Spectator Metal Cation)}}{M^+_{(aq)}} + \underset{\text{(Major Species of Strong Base)}}{OH^-_{(aq)}}$$

Example 1: List the major species NaOH $_{(aq)}$, and calculate its pH, pOH, $[H_3O^+]$ and $[OH^-]$ if it has a concentration of 1.00×10^{-4} M.

$$NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

Since NaOH is a strong base, the major species are: $OH^-_{(aq)}$ **and** $H_2O_{(l)}$.

$[OH^-] = [NaOH]_0 = 1.00 \times 10^{-4}$ M

$pOH = -\log[OH^-]$

$pOH = -\log(1.00 \times 10^{-4})$

pOH = 4.00

$pH = 14 - pOH$

$pH = 14 - (4.00)$

pH = 10.00

$pH = -\log[H_3O^+]$

$[H_3O^+] = 10^{-pH} = 10^{-10.00}$

$[H_3O^+] = 1.00 \times 10^{-10}$ M

Example 2: List the major species $Sr(OH)_{2\,(aq)}$, and calculate its pH, pOH, $[H_3O^+]$ and $[OH^-]$ if it has a concentration of 1.00×10^{-4} M.

$$Sr(OH)_{2\,(aq)} \rightarrow Sr^{2+}_{(aq)} + 2\ OH^-_{(aq)}$$

Since $Sr(OH)_2$ is a strong base, the major species are: $OH^-_{(aq)}$ **and** $H_2O_{(l)}$.

$[OH^-] = 2 \times [Sr(OH)_2]_0 = 2.00 \times 10^{-4}$ M

$pOH = -\log[OH^-]$

$pOH = -\log(2.00 \times 10^{-4})$

pOH = 3.70

$K_w = [H_3O^+][OH^-]$

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{2.00 \times 10^{-4}}$$

$[H_3O^+] = 5.00 \times 10^{-11}$ mol/L

$pH = -\log[H_3O^+]$

$pH = -\log(5.00 \times 10^{-11})$

pH = 10.3

Weak Bases: - bases that **dissociate LESS than 100% in water**.

- when the **base dissociation constant (K_b) is less than 1.** **($K_b < 1$)**
- **the equilibrium position strongly favours the reactants.**
- at equilibrium, the hydroxide concentration is much less than the original base concentration, $[OH]_{eq} < [B]_{eq}$ or $[B]_0 \approx [B]_{eq}$.
- the **conjugate acid, HA^-**, of a weak base **is itself a stronger weak acid** (can easily donate protons to do the reverse reaction).

$$\underset{\text{(Weak Base)}}{B_{(aq)}} + \underset{\text{(Acid)}}{H_2O_{(l)}} \overset{\longleftarrow}{\rightleftharpoons} \underset{\text{(Conjugate Base)}}{OH^-_{(aq)}} + \underset{\text{(Conjugate Acid-Stronger)}}{HB^+_{(aq)}}$$

	$[B]$		$[OH^-]$	$[HB^+]$
Initial **Change**	x $-y$ (where $y << x$)		0 $+y$	0 $+y$
Equilibrium	$(x - y) \approx x$		y	y

Examples: Some Weak Bases: $HOOCCOO^-_{(aq)}$, $HSO_3{}^-_{(aq)}$, $HSO_4{}^-_{(aq)}$, $H_2PO_4{}^-_{(aq)}$, $NO_2{}^-_{(aq)}$, $HCO_3{}^-_{(aq)}$, $H_2C_6H_5O_7{}^-_{(aq)}$, $F^-_{(aq)}$, $HCOO^-_{(aq)}$, $C_6H_7O_6{}^-_{(aq)}$, $C_6H_5COO^-_{(aq)}$, $CO_3{}^{2-}_{(aq)}$ $CH_3COO^-_{(aq)}$, $HS^-_{(aq)}$, $OCl^-_{(aq)}$, $CN^-_{(aq)}$, $NH_{3\,(aq)}$, and $NO_3{}^-_{(aq)}$

Base Dissociation Constant (K_b): - the equilibrium constant of a dissociation of a weak base in water.
- **K_b of a weak base can be calculated from K_a of its conjugate acid and K_w.**

$$\underset{\text{(Weak Base)}}{B_{(aq)}} + \underset{\text{(Acid)}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{(Conjugate Base)}}{OH^-_{(aq)}} + \underset{\text{(Conjugate Acid)}}{HB^+_{(aq)}} \qquad K_b = \frac{[OH^-][HB^+]}{[B]}$$

$$\underset{\text{(Weak Acid)}}{HB^+_{(aq)}} + \underset{\text{(Base)}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{(Conjugate Acid)}}{H_3O^+_{(aq)}} + \underset{\text{(Conjugate Base)}}{B_{(aq)}} \qquad K_a = \frac{[H_3O^+][B]}{[HB^+]}$$

$$K_a \times K_b = \frac{[H_3O^+]\cancel{[B]}}{\cancel{[HB^+]}} \times \frac{[OH^-]\cancel{[HB^+]}}{\cancel{[B]}} = [H_3O^+][OH^-] = K_w$$

Relationship between Conjugate Acid-Base Pair Dissociation Constants

$$K_w = K_a \times K_b$$

Major Species of a Weak Base: - since weak acid do not dissociate completely (K_b < 1) in water, **the major species of all weak bases are their original form and water.**

Procedure to calculate pH of Weak Base Solutions:

1. **List all the major species from all weak bases, including water.**
2. **Calculate K_b from all weak bases using $K_w = K_a \times K_b$**
3. **Determine which species has the highest K_b.** This is the **Strongest Base (SB)** of the list.
4. **Write the Brønsted-Lowry dissociation of this Strongest Base**
5. Set up the **ICE Box** and the **equilibrium expression**.
6. If the **original weak base concentration, $[B]_0$, is much larger than K_b**, we can **approximate by assuming $[B]_{eq} = ([B]_0 - x) \approx [B]_0$**. Thereby, simplifying the calculation. **(A general rule of thumb: if $[B]_0 \ge 1000 \times K_b$, we can use the approximation.)**
7. **Find the concentration of OH^-.**
8. **Verify** any **approximation** made by using the **5% rule**. $\left(\frac{[OH^-]}{[B]_0} \times 100\% \le 5\%\right)$

Example 3: Determine the $[OH^-]$, $[H_3O^+]$, pOH, pH and % dissociation of 0.200 mol/L of $NaCN_{(aq)}$.
(K_a of HCN = 6.2×10^{-10}).

NaCN dissociates completely in water: $NaCN_{(aq)} \rightarrow Na^+_{(aq)} + CN^-_{(aq)}$ **(Weak Base)**

Major Species: $CN^-_{(aq)}$, $H_2O_{(l)}$

$$K_b = \frac{K_w}{K_a} = \frac{1.00\times10^{-14}}{6.2\times10^{-10}} \qquad K_b = 1.613\times10^{-5}$$

(taking a few more decimal places to avoid round off errors)

Strongest Base: CN^- **($K_b = 1.613 \times 10^{-5}$)** ($K_b$ for H_2O is $K_w = 1.0 \times 10^{-14}$)

CN^- is a weak base and undergoes Brønsted-Lowry Dissociation.

$$CN^-_{(aq)} + H_2O_{(l)} \rightleftharpoons HCN_{(aq)} + OH^-_{(aq)}$$

	$[CN^-]$		$[H_3O^+]$	$[OH^-]$
Initial	**0.200 M**		0	**0**
Change	$-x$		$+x$	$+x$
Equilibrium	$(0.2 - x)$		x	x

$$K_b = \frac{[HCN][OH^-]}{[CN^-]} \qquad 1.613\times10^{-5} = \frac{(x)(x)}{(0.2-x)} \approx \frac{x^2}{(0.2)}$$

$$1.613\times10^{-5}\,(0.2) \approx x^2$$

$$3.226\times10^{-6} \approx x^2$$

$$x \approx \sqrt{3.226\times10^{-6}}$$

$$x \approx 0.0018$$

CAN use Approximation:

$$\frac{[CN^-]_0}{K_b} = \frac{0.200\text{ M}}{1.613\times10^{-5}}$$

$$= 12399 \geq 1000$$

Use 0.2 in the denominator, because $(0.2 - x) \approx 0.2$ [x is so small compared to 0.2 M]

$$\boxed{[OH^-] = 0.0018\text{ mol/L} = 1.8\text{ mmol/L}}$$

$$pOH = -\log[OH^-]$$

$$pOH = -\log(0.0018)$$

$$\boxed{pOH = 2.75}$$

<u>Verify that we could use Approximation:</u>

$$\frac{[OH^-]}{[CN^-]_0}\times100\% = \frac{0.0018\text{ M}}{0.200\text{ M}}\times100\%$$

$$= 0.90\% \leq 5\%$$

Therefore, approximation would be appropriate.

$$\%\ \text{Dissociation} = \frac{[OH^-]}{[CN^-]_0}\times100\%$$

$$\%\ \text{Dissociation} = \frac{0.0018\text{ M}}{0.200\text{ M}}\times100\%$$

$$\boxed{\%\ \text{Dissociation} = 0.90\%}$$

$$pH = 14 - pOH$$

$$pH = 14 - (2.75)$$

$$\boxed{pH = 11.25}$$

$$pH = -\log[H_3O^+]$$

$$[H_3O^+] = 10^{-pH} = 10^{-11.25}$$

$$\boxed{[H_3O^+] = 5.6\times10^{-12}\text{ M}}$$

<u>Assignment</u>

16.6 & 16.7 pg. 570 #53 to 58; pg. 571 #94

16.9: Acid-Base Properties of Salts

Salts: - ionic compounds that might dissociate in water.

1. **Neutral Salts**: - when the **Cation comes from a Strong Base** and the **Anion is the Conjugate-Base of Strong Acid**.
   - no effect on pH; if dissolve in pure water, pH will remain at 7.

Examples: Some Neutral Salts:

a. $KNO_{3\,(aq)}$ (K^+ can be from a strong base – $KOH_{(aq)}$; NO_3^- is the conjugate-base of a strong acid – $HNO_{3\,(aq)}$)

b. $NaCl_{(aq)}$ (Na^+ can be from a strong base – $NaOH_{(aq)}$; Cl^- is the conjugate-base a strong acid – $HCl_{(aq)}$)

2. **Basic Salts**: - when the **Cation comes from a Strong Base** and the **Anion is the Conjugate-Base of a Weak Acid**.
   - pH will increase; if dissolve in pure water, pH > 7 (Basic).

Examples: Some Basic Salts:

a. $NaCH_3COO_{(aq)}$ (Na^+ can be from a strong base – $NaOH_{(aq)}$; CH_3COO^- is the conjugate-base of a weak acid – $CH_3COOH_{(aq)}$)

b. $KF_{(aq)}$ (K^+ can be from a strong base – $KOH_{(aq)}$; F^- is the conjugate base of a weak acid – $HF_{(aq)}$)

Hydrolysis: - the reaction of a base and a water to form a conjugate acid and OH^-.
- it is basically the Brønsted-Lowry Dissociation of a weak base.

$$\underset{\text{(Weak Base)}}{B_{(aq)}} + \underset{\text{(Acid)}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{(Conjugate Base)}}{OH^-_{(aq)}} + \underset{\text{(Conjugate Acid)}}{HB^+_{(aq)}} \qquad K_b = \frac{[OH^-][HB^+]}{[B]}$$

Percent Hydrolysis: - the amount of $[OH^-]$ dissociated from the original $[B]_0$ expressed in percentage.
- strong bases will have % Hydrolysis = 100%.
- weak bases will have % Hydrolysis < 100%

Percent Hydrolysis of Bases

$$\%\ \text{Hydrolysis} = \frac{[OH^-]}{[B]_0} \times 100\%$$

3. **Acidic Salts**: - when the **Cation is the Conjugate-Acid of a Weak Base** and the **Anion is the Conjugate-Base of a Strong Acid**.

Example: $NH_4ClO_{4\,(aq)}$ is an acidic Salts: (NH_4^+ is the conjugate-acid of a weak base – $NH_{3\,(aq)}$; ClO_4^- is the conjugate-base of a strong acid – $HClO_{4\,(aq)}$)

Example 1: Determine the pH and the percent hydrolysis of 0.235 M of sodium nitrite. The acid dissociation constant for nitrous acid is $K_a = 4.0 \times 10^{-4}$.

Sodium nitrite ($NaNO_2$) dissociates completely in H_2O: $NaNO_{2\,(s)} \rightarrow Na^+_{(aq)} + NO_2^-{}_{(aq)}$ **(Weak Base)**

Major Species: $NO_2^-{}_{(aq)}$, $H_2O_{(l)}$

$$K_b = \frac{K_w}{K_a} = \frac{1.00\times10^{-14}}{4.0\times10^{-4}} \qquad K_b = 2.5\times10^{-11}$$

Strongest Base: NO_2^- ($K_b = 2.5 \times 10^{-11}$) ($K_b$ for H_2O is $K_w = 1.0 \times 10^{-14}$)

NO_2^- is a weak base and undergoes Brønsted-Lowry Dissociation.

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

	$[NO_2^-]$		$[HNO_2]$	$[OH^-]$
Initial	**0.235 M**		**0**	**0**
Change	$-x$		$+x$	$+x$
Equilibrium	$(0.235 - x)$		x	x

CAN use Approximation:

$$\frac{[NO_2^-]_0}{K_b} = \frac{0.235\text{ M}}{2.5\times10^{-11}} = 9.4\times10^{9} \geq 1000$$

Use 0.235 in the denominator, because $(0.235 - x) \approx 0.235$ [x is so small compared to 0.235 M]

$$K_b = \frac{[HNO_2][OH^-]}{[NO_2^-]} \qquad 2.5\times10^{-11} = \frac{(x)(x)}{(0.235-x)} \approx \frac{x^2}{(0.235)}$$

$$2.5\times10^{-11}\,(0.235) \approx x^2$$

$$5.875\times10^{-12} \approx x^2$$

$$x \approx \sqrt{5.875\times10^{-12}}$$

$$x = [OH^-] \approx 2.4\times10^{-6}\text{ M}$$

$$\text{pOH} = -\log\,[OH^-] \qquad \text{pH} = 14 - \text{pOH}$$

$$\text{pOH} = -\log(2.4\times10^{-6}) \qquad \text{pH} = 14 - (5.62)$$

$$\text{pOH} = 5.62 \qquad \boxed{\text{pH} = 8.38}$$

$$\%\text{ Hydrolysis} = \frac{[OH^-]}{[NO_2^-]_0}\times100\% = \frac{2.4\times10^{-6}\text{ M}}{0.235\text{ M}}\times100\% \qquad \boxed{\%\text{ Hydrolysis} = 0.00102\%}$$

As the small K_b suggests, the $[OH^-]$ is very small compared to $[NO_2^-]_0$. Therefore, % hydrolysis is also verv small as a result.

Example 2: Determine the pH of 0.0750 M of ammonium nitrate. The base dissociation constant for $NH_{3\,(aq)}$ is $K_b = 1.8 \times 10^{-5}$.

Ammonium nitrate, NH_4NO_3, dissociates completely in water:

$$\underset{0.0750\text{ M}}{NH_4NO_{3\,(s)}} \rightarrow \underset{0.0750\text{ M}}{NH_4^+{}_{(aq)}\text{ (Weak Acid)}} + NO_3^-{}_{(aq)}$$

Major Species: $NH_4^+{}_{(aq)}$, $H_2O_{(l)}$

$$K_a = \frac{K_w}{K_b} = \frac{1.00 \times 10^{-14}}{1.8 \times 10^{-5}} \qquad K_a = 5.556 \times 10^{-10}$$

Strongest Acid: NH_4^+ **($K_a = 5.556 \times 10^{-10}$)** ($K_a$ for H_2O is $K_w = 1.0 \times 10^{-14}$)

NH_4^+ is a weak acid and undergoes Brønsted-Lowry Dissociation.

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

	$[NH_4^+]$		$[H_3O^+]$	$[NH_3]$
Initial	0.0750 M		0	0
Change	$-x$		$+x$	$+x$
Equilibrium	$(0.075 - x)$		x	x

CAN use Approximation:

$$\frac{[NH_4^+]_0}{K_a} = \frac{0.0750\text{ M}}{5.556 \times 10^{-10}}$$

$$= 1.35 \times 10^8 \geq 1000$$

Use 0.075 in the denominator, because $(0.075 - x) \approx 0.075$ [x is so small compared to 0.075 M]

$$K_a = \frac{[H_3O^+][NH_3]}{[NH_4^+]} \qquad 5.556 \times 10^{-10} = \frac{(x)(x)}{(0.075 - x)} \approx \frac{x^2}{(0.075)}$$

$$5.556 \times 10^{-10}\,(0.075) \approx x^2$$

$$4.167 \times 10^{-11} \approx x^2$$

$$x \approx \sqrt{4.167 \times 10^{-11}}$$

$$x = [H_3O^+] \approx 6.455 \times 10^{-6}\text{ M}$$

$$pH = -\log[H_3O^+]$$

$$pH = -\log(6.455 \times 10^{-6})$$

$$\boxed{pH = 5.19}$$

Verify for Using Approximation:

$$\frac{[H_3O^+]}{[NH_4^+]_0} \times 100\% = \frac{6.455 \times 10^{-6}\text{ M}}{0.0750\text{ M}} \times 100\%$$

$= 0.0086\% \leq 5\%$ **Therefore, approximation would be appropriate.**

Assignment

16.9 pg. 570−571 #65, 68, 70 to 74

Chapter 17: Acid-Base Equilibria

17.2: The Buffer Solutions

Buffered Solution: - a solution that **resists a change in pH** when a small amount of H_3O^+ or OH^- is added.
- consists of a pair of **weak acid**/**conjugate base common ion** or a pair of **weak base**/**conjugate acid common ion**.

Acidic Buffered Solution: $HA_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + A^-_{(aq)}$

a. Small Amounts of H_3O^+ is Added: $A^-_{(aq)} + H_3O^+_{(aq)} \rightarrow H_2O_{(l)} + HA_{(aq)}$

(H^+ - Strong Acid reacts completely with A^-)

(More HA – Weak Acid: pH will only be lowered SLIGHTLY!)

b. Small Amounts of OH^- is Added: $HA_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)} + A^-_{(aq)}$

(OH^- - Strong Base reacts completely with HA)

(More A^- – Weak Base: pH will only be raised SLIGHTLY!)

Basic Buffered Solution: $B_{(aq)} + H_2O_{(l)} \rightleftharpoons OH^-_{(aq)} + HB^+_{(aq)}$

a. Small Amounts of H_3O^+ is Added: $B_{(aq)} + H_3O^+_{(aq)} \rightarrow H_2O_{(l)} + HB^+_{(aq)}$

(H^+ - Strong Acid reacts completely with B)

(More HB^+– Weak Acid: pH will only be lowered SLIGHTLY!)

b. Small Amounts of OH^- is Added: $B_{(aq)} + H_2O_{(l)} \rightleftharpoons OH^-_{(aq)} + HB^+_{(aq)}$

(OH^- - Strong Base increases $[OH^-]$; eq shifts left)

(More B – Weak Base: pH will only be raised SLIGHTLY!)

Example: $HCO_3^-{}_{(aq)}$ / CO_3^{2-} is a buffer because HCO_3^- acts as an acid and CO_3^{2-} acts as a base whenever a small amount of strong base or acid is added to it. Thus, pH will remain somewhat constant.

When small amounts of strong base is added: $HCO_3^-{}_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)} + CO_3^{2-}{}_{(aq)}$

When small amounts of strong acid is added: $CO_3^{2-}{}_{(aq)} + H_3O^+_{(aq)} \rightarrow H_2O_{(l)} + HCO_3^-{}_{(aq)}$

Example 1: Specify which of these systems can be classified as a buffer system.

a. HNO_3 / KNO_3	b. CH_3COOH / $NaCH_3COO$	c. H_2SO_4 / $NaHSO_4$	d. NH_3 / NH_4Cl
HNO_3 is a strong acid. Hence, NO_3^- is not a strong enough base to resist the change in pH. **HNO_3 / KNO_3 is *NOT* a buffer system.**	**CH_3COOH is a weak acid.** Therefore, CH_3COO^- is a ***strong enough weak base*** to resist the change in pH. $NaCH_3COO$ provides this base. **CH_3COOH / CH_3COO^- is a buffer system.**	**H_2SO_4 is a strong acid.** Hence, HSO_4^- is not a strong enough base to resist the change in pH. **H_2SO_4 / $NaHSO_4$ is *NOT* a buffer**	**NH_3 is a weak base.** Therefore, NH_4^+ is a ***strong enough weak acid*** to resist the change in pH. NH_4Cl provides this acid. **NH_3 / NH_4Cl is a buffer system.**

Example 2: Calculate the pH of a 1.00 L buffered solution consisting of 0.500 M of methylamine ($K_b = 4.38 \times 10^{-4}$) with 0.450 M of CH_3NH_3Cl when:

a. 0.0200 mol of KOH is added to it.
b. 3.00 mL of 0.750 M of $HNO_{3\,(aq)}$ is added to it.
c. Contrast the pH's of the above two additions against the same additions to 1.00 L of water.

First, we have to figure out the [OH⁻] and pH for the buffered solution system.

Methylamine (Base) Dissociation: $CH_2NH_{2\,(aq)} + H_2O_{(l)} \rightleftharpoons CH_2NH_3^+{}_{(aq)} + OH^-{}_{(aq)}$

CH₃NH₃Cl: $CH_3NH_3Cl_{(aq)} \rightarrow CH_2NH_3^+{}_{(aq)} + Cl^-{}_{(aq)}$

0.450 M **0.450 M**

$$CH_2NH_{2\,(aq)} + H_2O_{(l)} \rightleftharpoons CH_2NH_3^+{}_{(aq)} + OH^-{}_{(aq)}$$

	$[CH_2NH_2]$		$[CH_2NH_3^+]$	$[OH^-]$
Initial	**0.500 M**		**0.450 M**	**0**
Change	$-x$		$+x$	$+x$
Equilibrium	$(0.5 - x)$		$(0.45 + x)$	x

$$K_b = \frac{[OH^-][CH_2NH_3^+]}{[CH_2NH_2]} \qquad 4.38\times10^{-4} = \frac{(x)(0.45+x)}{(0.5-x)} \approx \frac{x(0.45)}{(0.5)}$$

$$4.38\times10^{-4}\,\frac{(0.5)}{(0.45)} \approx x$$

$$x \approx 4.87\times10^{-4}$$

CAN use Approximation:

$$\frac{[CH_2NH_2]_0}{K_b} = \frac{0.500\text{ M}}{4.38\times10^{-4}}$$

$$= 1142 \geq 1000$$

Use 0.45 in the numerator, because $(0.45 + x) \approx 0.45$ [x is so small compared to 0.45 M].

Use 0.5 in the denominator, because $(0.5 - x) \approx 0.5$ [x is so small compared to 0.5 M].

Verify that we could use Approximation:

$$\frac{[OH^-]}{[CH_2NH_2]_0}\times 100\% = \frac{4.87\times10^{-4}\text{ M}}{0.500\text{ M}}\times 100\%$$

$$= 0.0973\% \leq 5\%$$

Therefore, approximation would be appropriate.

$$\boxed{[OH^-] = 4.87\times10^{-4}\text{ mol/L}}$$

pOH = – log [OH⁻]
pOH = – log(4.87×10^{-4})
$\boxed{pOH = 3.31}$

pH = 14 – pOH
pH = 14 – 3.31
$\boxed{pH = 10.69}$

a. 0.0200 mol of KOH is added to the buffered solution **(Complete Rxn between OH⁻ & $CH_2NH_3^+$).**

$$OH^-{}_{(aq)} + CH_2NH_3^+{}_{(aq)} \rightarrow H_2O_{(l)} + CH_2NH_{2\,(aq)}$$

	n of OH^-	n of $CH_2NH_3^+$		n of CH_2NH_2
Before	**0.0200 mol (Limiting)**	**0.450 mol**		**0.500 mol**
Change	**– 0.0200 mol**	**– 0.0200 mol**		**+ 0.0200 mol**
After	**0**	**0.430 mol**		**0.520 mol**

Recalculate equilibrium concentrations using **NEW $[CH_2NH_3^+]_0$** and **$[CH_2NH_2]_0$**

$$CH_2NH_{2\,(aq)} + H_2O_{(l)} \rightleftharpoons CH_2NH_3^+{}_{(aq)} + OH^-{}_{(aq)}$$

	$[CH_2NH_2]$		$[CH_2NH_3^+]$	$[OH^-]$
Initial	**0.520 M**		**0.430 M**	**0**
Change	$-y$		$+y$	$+y$
Equilibrium	$(0.52 - y)$		$(0.43 + y)$	y

$$K_b = \frac{[OH^-][CH_2NH_3^+]}{[CH_2NH_2]} \qquad 4.38\times10^{-4} \approx \frac{y(0.43)}{(0.52)}$$

$$4.38\times10^{-4} = \frac{(y)(0.43+y)}{(0.52-y)} \qquad 4.38\times10^{-4}\,\frac{(0.52)}{(0.43)} \approx y$$

$$y \approx 5.30\times10^{-4}$$

CAN use Approximation:

$$\frac{[CH_2NH_2]_0}{K_b} = \frac{0.520\text{ M}}{4.38\times10^{-4}} = 1187 \geq 1000$$

Use 0.43 in the numerator, because $(0.43 + y) \approx 0.43$ [y is so small compared to 0.43 M].

Use 0.52 in the denominator, because $(0.52 - y) \approx 0.52$ [y is so small compared to 0.52 M].

Verify that we could use Approximation:

$$\frac{[OH^-]}{[CH_2NH_2]_0}\times 100\% = \frac{5.30\times10^{-4}\text{ M}}{0.520\text{ M}}\times 100\%$$

= 0.102% ≤ 5% (Appropriate Approximation)

$$\boxed{[OH^-] = 5.30\times10^{-4}\text{ mol/L}}$$

pOH = – log [OH⁻]
pOH = – log(5.30×10^{-4})
$\boxed{pOH = 3.28}$

pH = 14 – pOH
pH = 14 – 3.28
$\boxed{pH = 10.72}$

b. 3.00 mL of 0.750 M (0.00225 mol) of $HNO_{3\ (aq)}$ is added to the buffered solution

(Complete Reaction between H_3O^+ & CH_2NH_2).

$$H_3O^+_{(aq)} + CH_2NH_{2\ (aq)} \rightarrow H_2O_{(l)} + CH_2NH_3^+_{(aq)}$$

	n of H_3O^+	*n* of CH_2NH_2		*n* of $CH_2NH_3^+$
Before	**0.00225 mol (Limiting)**	**0.500 mol**		**0.450 mol**
Change	**– 0.00225 mol**	**– 0.00225 mol**		**+ 0.00225 mol**
After	**0**	**0.49775 mol**		**0.45225 mol**

Recalculate equilibrium concentrations using **NEW $[CH_2NH_2]_0$** and **$[CH_2NH_3^+]_0$ (neglect Δvolume)**

$$CH_2NH_{2\ (aq)} + H_2O_{(l)} \rightleftharpoons CH_2NH_3^+_{(aq)} + OH^-_{(aq)}$$

	$[CH_2NH_2]$		$[CH_2NH_3^+]$	$[OH^-]$
Initial	**0.49775 M**		**0.45225 M**	**0**
Change	**–z**		**+z**	**+z**
Equilibrium	**(0.49775 – z)**		**(0.45225 + z)**	**z**

$$K_b = \frac{[OH^-][CH_2NH_3^+]}{[CH_2NH_2]} \qquad 4.38 \times 10^{-4} \approx \frac{z(0.45225)}{(0.49775)}$$

$$4.38 \times 10^{-4} = \frac{(y)(0.45225 + z)}{(0.49775 - z)} \qquad 4.38 \times 10^{-4}\ \frac{(0.49775)}{(0.45225)} \approx z$$

CAN use Approximation:

$$\frac{[CH_2NH_2]_0}{K_b} = \frac{0.520\ \text{M}}{4.38 \times 10^{-4}} = 1187 \geq 1000$$

Use 0.45225 in the numerator, because (0.45225 + z) ≈ 0.45225 [z is so small compared to 0.45225 M].

Use 0.49775 in the denominator, because (0.49775 – z) ≈ 0.49775 [z is so small compared to 0.49775 M].

Verify that we could use Approximation:

$$\frac{[OH^-]}{[CH_2NH_2]_0} \times 100\% = \frac{4.82 \times 10^{-4}\ \text{M}}{0.49775\ \text{M}} \times 100\%$$

= 0.0968% ≤ 5% (Appropriate Approximation)

$$\boxed{[OH^-] = 4.82 \times 10^{-4}\ \text{mol/L}}$$

$pOH = -\log [OH^-]$

$pOH = -\log(4.82 \times 10^{-4})$

$$\boxed{pOH = 3.32}$$

$pH = 14 - pOH$

$pH = 14 - 3.28$

$$\boxed{pH = 10.68}$$

c. Contrast the pH's of the above two additions against the same additions to 1.00 L of water.

i. 0.0200 mol of KOH is added to 1.00 L of water

$pOH = -\log [OH^-]$

$pOH = -\log(0.0200)$

$pOH = 1.70$

$pH = 14 - pOH$

$pH = 14 - 1.70$

$$\boxed{pH = 12.30}$$

Adding the 0.0200 mol of KOH to buffered solution.

pH changes from 10.69 to 10.72

ii. 3.00 mL of 0.750 M (0.00225 mol) of $HNO_{3\ (aq)}$ is added to 1.00 L of water (neglect Δvolume).

$pH = -\log [H_3O^+]$

$pH = -\log(0.00225)$

$$\boxed{pH = 2.65}$$

Adding the 0.00225 mol of HNO_3 to buffered solution.

pH changes from 10.69 to 10.68

Assignment

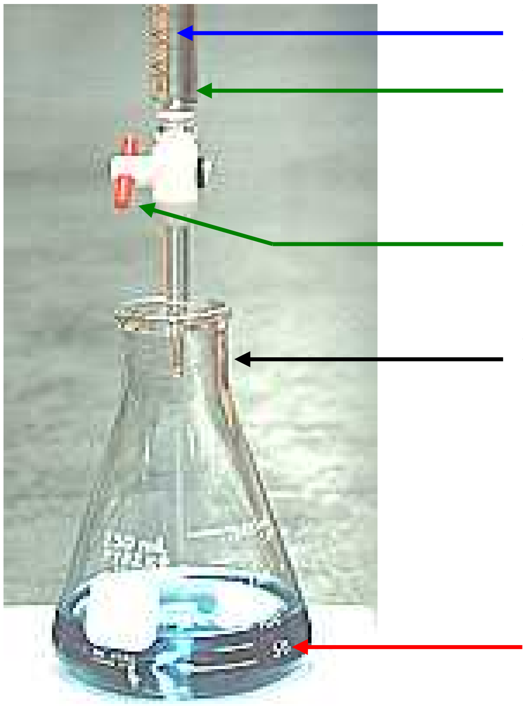
17.2 pg. 604 #1, 5 and 6

17.3: A Closer Look at Acid-Base Titrations

pH Curve: - a graph that shows the how the pH changes as the titration proceeds (as titrant is to the analyzed).

Titration: - a volumetric analysis that involves measuring the volume of known concentration solution to measure a given volume of an unknown concentration solution.

Titration Set-up



Titrant: - the solution of known concentration.

Buret: - a precise apparatus to deliver the titrant.
- the volume of the titrant added is read by subtracting the final volume and the initial volume.

Buret Valve: - can be adjusted to let one drop out at a time.

Erlenmeyer Flask: - a container commonly uses to hold the analyte. (Narrow mouth prevents splash and spillage.)

Analyte: - the solution of an unknown concentration.
- the exact volume is usually delivered by a pipet.

Acid-Base Titration: - volumetric analysis that assist in determining the unknown concentration in an acid and base neutralization.

Equivalent Point (Stoichiometric Point): - a point where the number of moles of H^+ is equivalent to the number of moles of OH^-. ($n_{H^+} = n_{OH^-}$)

Endpoint: - a point where the indicator actually changes colour to indicate neutralization is completed.

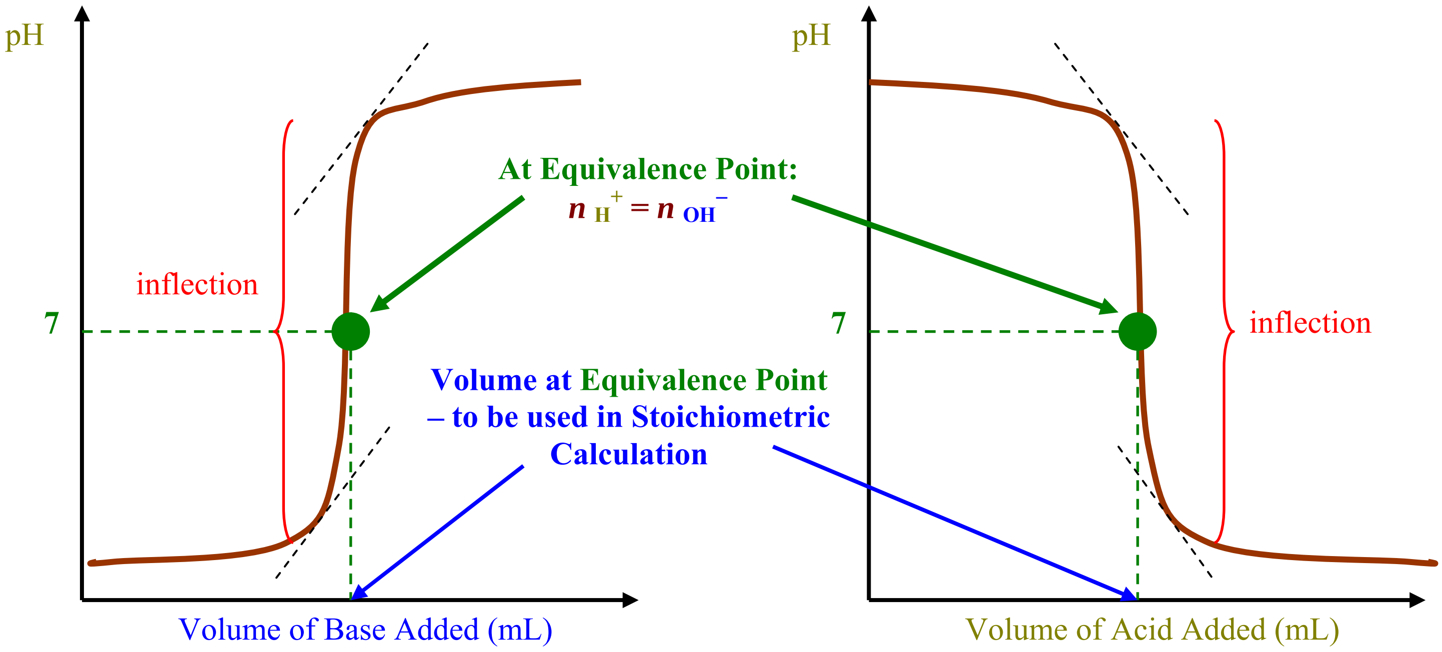
Indicator: - a chemical that changes colour due to the pH of the solution (more in the next section).

Inflection: - the part of the curve where there is a sudden rise or drop in pH.
- the midpoint of the inflection marks the equivalence point.

1. **Titration Between Strong Acids and Strong Base**: - **Equivalence Point** always occur at **pH = 7**.

Strong Acid (unknown concentration with known volume - analyte) titrated with Strong Base (known concentration - titrant)

Strong Base (unknown concentration and known volume - analyte) titrated with Strong Acid (known concentration - titrant)



Example 1: Calculate the pH when 30.0 mL of 0.100 M of $HCl_{(aq)}$ is titrated with 0.200 M of $KOH_{(aq)}$ at:

a. 0 mL of $KOH_{(aq)}$ added.
b. 5.00 mL of $KOH_{(aq)}$ added.
c. 15.00 mL of $KOH_{(aq)}$ added.
d. 20.00 mL of $KOH_{(aq)}$ added.

a. At 0 mL of $KOH_{(aq)}$ added, the pH will be solely based on the [HCl] in the flask:

$$\mathbf{HCl_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}}$$

$pH = -\log [H_3O^+] = -\log(0.100)$ **pH = 1.00**

b. At 5.00 mL of $KOH_{(aq)}$ added: **(Before Stoichiometric Point)**

Chemical Species Present: $\overset{SA}{H_3O^+}$, $\underset{B}{Cl^-}$, K^+, $\underset{SB}{OH^-}$, $\overset{A}{\underset{B}{H_2O}}$

Net-Ionic Equation: $H_3O^+_{(aq)} + OH^-_{(aq)} \rightarrow 2\ H_2O_{(l)}$

	n of H_3O^+	*n* of OH^-
Before	(0.100 mol/L)(30.00 mL) = 3.00 mmol	(0.200 mol/L)(5.00 mL) = 1.00 mmol (LR)
Change	**– 1.00 mmol**	**– 1.00 mmol**
After	**2.00 mmol**	**0.00 mmol**

$$[H_3O^+] = \frac{n_{H_3O^+}}{\text{Total Volume}} = \frac{2.00\ \cancel{m}mol}{(30.00\ \cancel{m}L + 5.00\ \cancel{m}L)} = 0.05714\ mol/L$$

$pH = -\log [H_3O^+] = -\log(0.05714)$ **pH = 1.24**

(pH's results from similar Calculations with other volumes before the stoichiometric point is listed on the next page.)

c. At 15.00 mL of KOH $_{(aq)}$ added: **(Stoichiometric Point)**

SA A
Chemical Species Present: H_3O^+, Cl^-, K^+, OH^-, H_2O
B **SB** B

Net-Ionic Equation: $H_3O^+_{(aq)} + OH^-_{(aq)} \rightarrow 2\ H_2O_{(l)}$

	***n* of H_3O^+**	***n* of OH^-**
Before	(0.100 mol/L)(30.00 mL) = 3.00 mmol	(0.200 mol/L)(15.00 mL) = 3.00 mmol
Change	**– 3.00 mmol**	**– 3.00 mmol**
After	**0.00 mmol**	**0.00 mmol**

$[H_3O^+] = [OH^-] = 1.00 \times 10^{-7}$ mol/L

$pH = -\log [H_3O^+] = -\log(1.00 \times 10^{-7})$ **pH = 7.00**

d. At 20.00 mL of KOH $_{(aq)}$ added: **(After Stoichiometric Point)**

SA A
Chemical Species Present: H_3O^+, Cl^-, K^+, OH^-, H_2O
B **SB** B

Net-Ionic Equation: $H_3O^+_{(aq)} + OH^-_{(aq)} \rightarrow 2\ H_2O_{(l)}$

	***n* of H_3O^+**	***n* of OH^-**
Before	(0.100 mol/L)(30.00 mL) = 3.00 mmol (Limiting Reagent)	(0.200 mol/L)(20.00 mL) = 4.00 mmol
Change	**– 3.00 mmol**	**– 3.00 mmol**
After	**0.00 mmol**	**1.00 mmol**

$$[OH^-] = \frac{n_{OH^-}}{\text{Total Volume}} = \frac{1.00\ \cancel{m}\text{mol}}{(30.00\ \cancel{m}\text{L} + 20.00\ \cancel{m}\text{L})} = 0.0200 \text{ mol/L}$$

$pOH = -\log [OH^-] = -\log(0.0200)$ pOH = 1.70

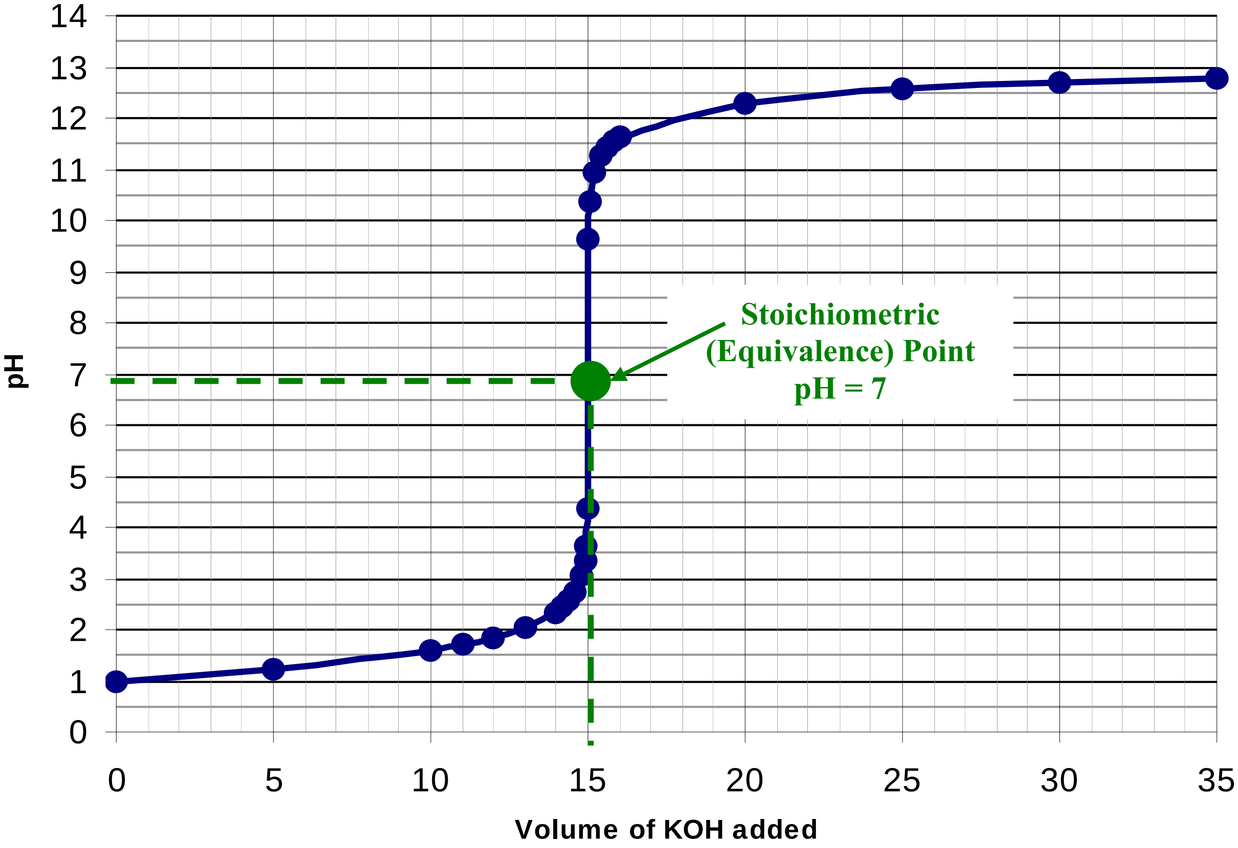
pH = 14 – pOH **pH = 12.30**

(pH's results from similar Calculations with other volumes after the stoichiometric point is listed below.)

pH of 30.0 mL of 0.100 M HCl $_{(aq)}$ titrated by 0.200 M of KOH $_{(aq)}$

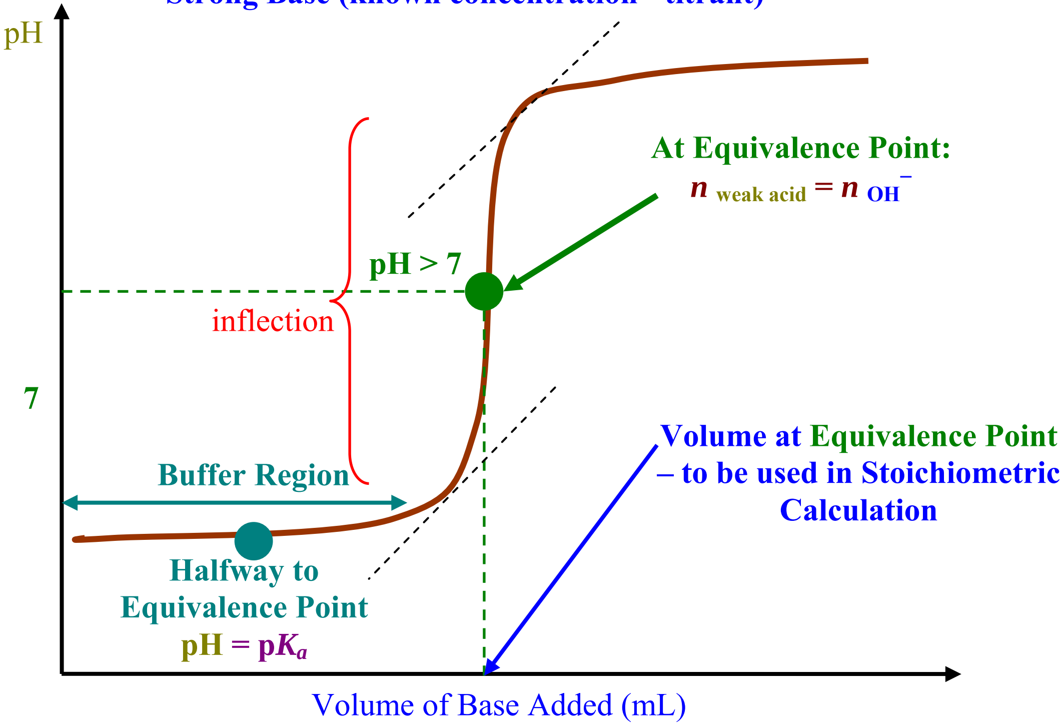
Volume of KOH added	pH	Volume of KOH added	pH	Volume of KOH added	pH
0.00 mL	1.00	14.60 mL	2.75	15.40 mL	11.25
5.00 mL	1.24	14.80 mL	3.05	15.60 mL	11.42
10.00 mL	1.60	14.90 mL	3.35	15.80 mL	11.54
11.00 mL	1.71	14.95 mL	3.65	16.00 mL	11.64
12.00 mL	1.85	14.99 mL	4.35	20.00 mL	12.30
13.00 mL	2.03	**15.00 mL (Equivalence)**	**7.00**	25.00 mL	12.56
14.00 mL	2.34	15.01 mL	9.65	30.00 mL	12.70
14.20 mL	2.44	15.05 mL	10.35	35.00 mL	12.79
14.40 mL	2.57	15.20 mL	10.95		

pH of 30.0 mL of 0.100 M HCl titrated by 0.200 M of KOH



2. **Titration Between Weak Acids and Strong Base**: - **Equivalence Point** always occur at **pH > 7**.

Weak Acid (unknown concentration with known volume - analyte) titrated with Strong Base (known concentration - titrant)



Example 2: Calculate the pH when 30.0 mL of 0.100 M of HCOOH $_{(aq)}$ ($K_a = 1.8 \times 10^{-4}$) is titrated with 0.200 mol/L of KOH $_{(aq)}$ at:

a. 0 mL of KOH $_{(aq)}$ added.
b. 7.50 mL of KOH $_{(aq)}$ added.
c. 15.00 mL of KOH $_{(aq)}$ added.
d. 20.00 mL of KOH $_{(aq)}$ added.

a. At 0 mL of KOH $_{(aq)}$ added, the pH will be calculated the weak acid dissociation in the flask:

$$\mathbf{HCOOH}_{(aq)} + \mathbf{H_2O}_{(l)} \rightleftharpoons \mathbf{H_3O^+}_{(aq)} + \mathbf{HCOO^-}_{(aq)}$$

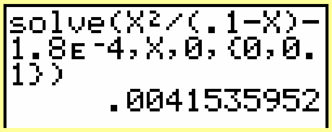
	[HCOOH]		**[H_3O^+]**	**[$HCOO^-$]**
Initial	**0.100 M**		**0**	**0 M**
Change	**$-x$**		**$+x$**	**$+x$**
Equilibrium	**$(0.1 - x)$**		**x**	**x**

CANNOT use Approximation:

$$\frac{[\mathrm{HCOOH}]_0}{K_a} = \frac{0.100\text{ M}}{1.8\times10^{-4}}$$

= 555.6 < 1000

$$K_a = \frac{[\mathrm{H_3O^+}][\mathrm{HCOO^-}]}{[\mathrm{HCOOH}]} \qquad 1.8\times10^{-4} = \frac{(x)(x)}{(0.1-x)} = \frac{x^2}{(0.1-x)}$$



```
solve(X²/(.1-X)-
1.8ᴇ-4,X,0,{0,0.
1})
      .0041535952
```

[H_3O^+] = x = 0.0042 mol/L

pH = −log [H_3O^+] = −log(0.042)

pH = 2.38

b. At 7.50 mL of KOH $_{(aq)}$ added: **(Before Stoichiometric Point: at Halfway to Equivalence Point)**

Chemical Species Present: **HCOOH,** K^+, **OH^-,** H_2O
SA **SB** A/B

Net-Ionic Equation: $\mathbf{HCOOH}_{(aq)} + \mathbf{OH^-}_{(aq)} \rightarrow \mathbf{H_2O}_{(l)} + \mathbf{HCOO^-}_{(aq)}$

	***n* of HCOOH**	***n* of OH^-**		***n* of $HCOO^-$**
Before	(0.100 mol/L)(30.00 mL) = 3.00 mmol	(0.200 mol/L)(7.50 mL) = 1.50 mmol (LR)		0 mmol
Change	**− 1.50 mmol**	**− 1.50 mmol**		**+ 1.50 mmol**
After	**1.50 mmol**	**0 mmol**		**1.50 mmol**

$$[\mathrm{HCOOH}] = \frac{n_{\mathrm{HCOOH}}}{\text{Total Volume}} = \frac{1.50\ \cancel{\mathrm{m}}\mathrm{mol}}{(30.00\ \cancel{\mathrm{m}}\mathrm{L} + 7.50\ \cancel{\mathrm{m}}\mathrm{L})} = 0.0400\text{ mol/L}$$

$$[\mathrm{HCOO^-}] = \frac{n_{\mathrm{HCOO^-}}}{\text{Total Volume}} = \frac{1.50\ \cancel{\mathrm{m}}\mathrm{mol}}{(30.00\ \cancel{\mathrm{m}}\mathrm{L} + 7.50\ \cancel{\mathrm{m}}\mathrm{L})} = 0.0400\text{ mol/L}$$

[HCOOH] = [$HCOO^-$]
pH ≈ pK_a
(Max Buffer Capacity)

After all the OH^- is used up, the resulting solution follows the regular weak acid dissociation.

$$\mathbf{HCOOH}_{(aq)} + \mathbf{H_2O}_{(l)} \rightleftharpoons \mathbf{H_3O^+}_{(aq)} + \mathbf{HCOO^-}_{(aq)}$$

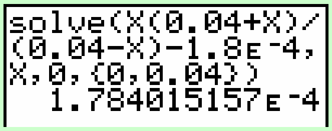
	[HCOOH]		**[H_3O^+]**	**[$HCOO^-$]**
Initial	**0.0400 M**		**0**	**0.0400 M**
Change	**$-x$**		**$+x$**	**$+x$**
Equilibrium	**$(0.04 - x)$**		**x**	**$(0.04 + x)$**

CANNOT use Approximation:

$$\frac{[\mathrm{HCOOH}]_0}{K_a} = \frac{0.0400\text{ M}}{1.8\times10^{-4}}$$

= 222.22 < 1000

$$K_a = \frac{[\mathrm{H_3O^+}][\mathrm{HCOO^-}]}{[\mathrm{HCOOH}]} \qquad 1.8\times10^{-4} = \frac{(x)(0.04+x)}{(0.04-x)}$$



```
solve(X(0.04+X)/
(0.04-X)-1.8ᴇ-4,
X,0,{0,0.04})
   1.784015157ᴇ-4
```

[H_3O^+] = x = 1.8 × 10^{-4} mol/L pH = pK_a = −log(1.8 × 10^{-4})

(pH's results from similar Calculations with other volumes before the stoichiometric point is listed on page 306.)

pH = 3.75

c. At 15.00 mL of KOH $_{(aq)}$ added: **(Stoichiometric Point)**

Chemical Species Present: **HCOOH,** K^+, **OH^-,** H_2O

SA SB A/B

Net-Ionic Equation: $HCOOH_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)} + HCOO^-_{(aq)}$

	***n* of HCOOH**	***n* of OH⁻**		***n* of HCOO⁻**
Before	(0.100 mol/L)(30.00 mL) = 3.00 mmol	(0.200 mol/L)(15.00 mL) = 3.00 mmol		0 mmol
Change	**– 3.00 mmol**	**– 3.00 mmol**		**+ 3.00 mmol**
After	**0 mmol**	**0 mmol**		**3.00 mmol**

$$[HCOO^-] = \frac{n_{HCOO^-}}{\text{Total Volume}} = \frac{3.00\ \cancel{m}\text{mol}}{(30.00\ \cancel{m}\text{L} + 15.00\ \cancel{m}\text{L})} = 0.0667 \text{ mol/L}$$

At Stoichiometric Point, after all the OH^- and HCOOH are used up, the resulting solution follows the regular weak base dissociation.

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

	[HCOO⁻]		**[HCOOH]**	**[OH⁻]**
Initial	**0.0667 M**		**0**	**0 M**
Change	**–*x***		+*x*	**+*x***
Equilibrium	**(0.0667 – *x*)**		*x*	***x***

CAN use Approximation:

$$\frac{[HCOO^-]_0}{K_b} = \frac{0.0667\text{ M}}{5.556 \times 10^{-11}}$$

$$= \mathbf{1.2 \times 10^9 \geq 1000}$$

Use 0.0667 in the denominator, because (0.0667 – *x*) ≈ 0.03667 [*x* is so small compared to 0.0667 M].

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-4}} \qquad K_b = 5.556 \times 10^{-11} = \frac{[HCOOH][OH^-]}{[HCOO^-]}$$

$$5.556 \times 10^{-11} = \frac{(x)(x)}{(0.0667 - x)} \approx \frac{x^2}{0.0667}$$

$$\mathbf{[OH^-] = x = 1.92 \times 10^{-6} \text{ mol/L}} \qquad pOH = -\log(1.92 \times 10^{-6}) = 5.72$$

$$pH = 14 - pOH = 14 - 5.72 \qquad \mathbf{pH = 8.28}$$

d. At 20.00 mL of KOH $_{(aq)}$ added: **(After Stoichiometric Point)**

Chemical Species Present: **HCOOH,** K^+, **OH^-,** H_2O

SA SB A/B

Net-Ionic Equation: $HCOOH_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)} + HCOO^-_{(aq)}$

	***n* of HCOOH**	***n* of OH⁻**
Before	(0.100 mol/L)(30.00 mL) = 3.00 mmol (LR)	(0.200 mol/L)(20.00 mL) = 4.00 mmol
Change	**– 3.00 mmol**	**– 3.00 mmol**
After	**0 mmol**	**1.00 mmol**

After all the HCOOH is used up, the resulting solution has the following major species.

Chemical Species Present: K^+, **OH^-,** H_2O

SB A/B

$$[OH^-] = \frac{n_{OH^-}}{\text{Total Volume}} = \frac{1.00\ \cancel{m}\text{mol}}{(30.00\ \cancel{m}\text{L} + 20.00\ \cancel{m}\text{L})} = 0.0200 \text{ mol/L} \qquad pOH = -\log(0.0200) = 1.70$$

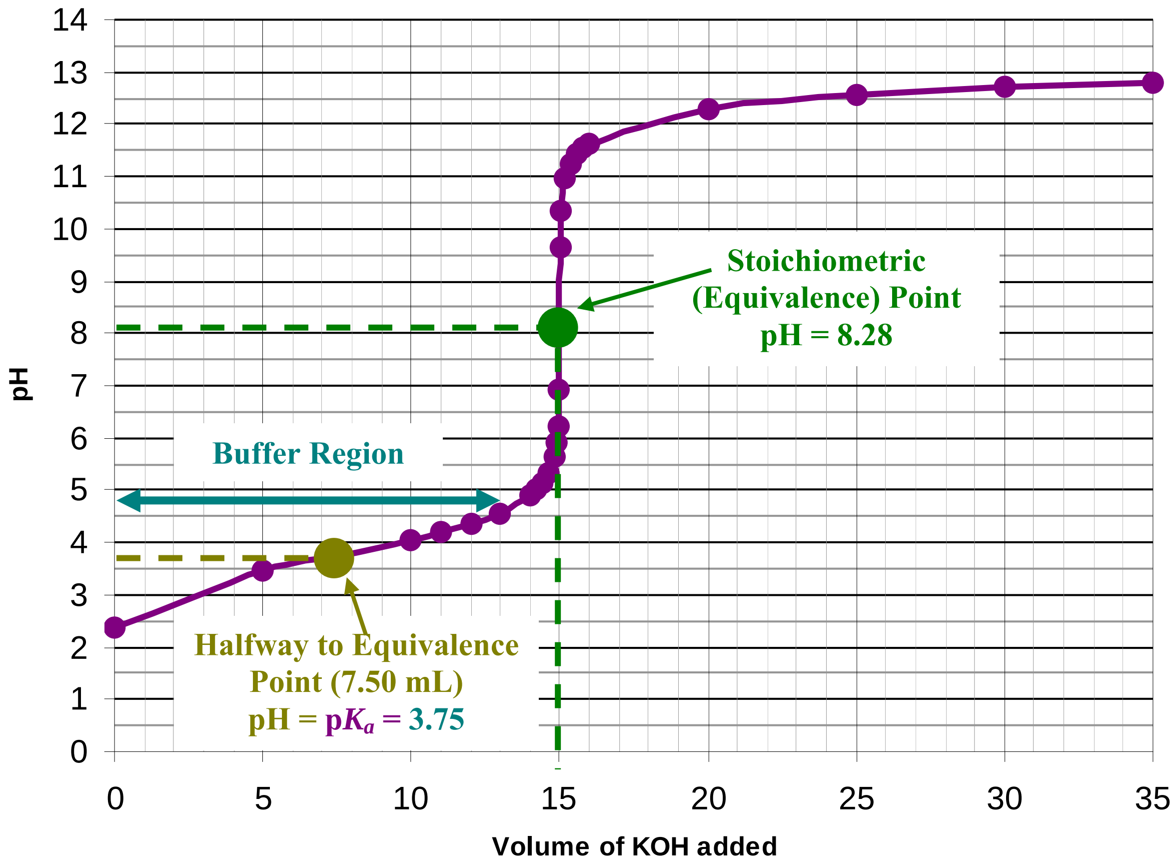
$$pH = 14 - pOH = 14 - 1.70 \qquad \mathbf{pH = 12.30}$$

(pH's results from similar Calculations with other volumes after the stoichiometric point is listed on the next page.)

pH of 30.0 mL of 0.100 M HCOOH $_{(aq)}$ titrated by 0.200 M of KOH $_{(aq)}$

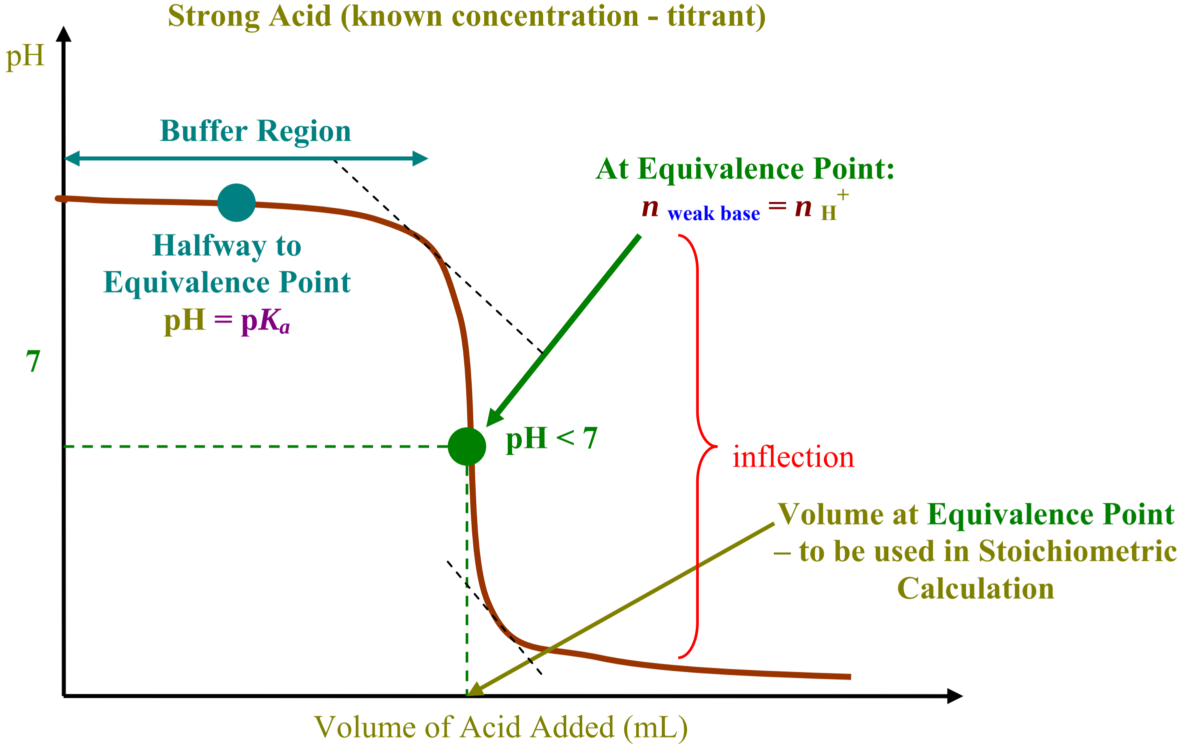
Volume of KOH added	pH	Volume of KOH added	pH	Volume of KOH added	pH
0.00 mL	2.38	14.40 mL	5.13	15.20 mL	10.95
5.00 mL	3.45	14.60 mL	5.31	15.40 mL	11.25
7.50 mL (Halfway to Equivalence)	3.75	14.80 mL	5.62	15.60 mL	11.42
10.00 mL	4.05	14.90 mL	5.91	15.80 mL	11.54
11.00 mL	4.19	14.95 mL	6.22	16.00 mL	11.64
12.00 mL	4.35	14.99 mL	6.92	20.00 mL	12.30
13.00 mL	4.56	**15.00 mL (Equivalence)**	**8.28**	25.00 mL	12.56
14.00 mL	4.89	15.01 mL	9.65	30.00 mL	12.70
14.20 mL	5.00	15.05 mL	10.35	35.00 mL	12.79

pH of 30.0 mL of 0.100 M HCOOH titrated by 0.200 M of KOH



3. **Titration Between Weak Bases and Strong Acid**: - **Equivalence Point** always occur at **$pH < 7$**.

Weak Base (unknown concentration with known volume - analyte) titrated with Strong Acid (known concentration - titrant)



Example 3: Calculate the pH when 30.0 mL of 0.100 M of $CH_3NH_{2\,(aq)}$ ($K_b = 4.38 \times 10^{-4}$) is titrated with 0.200 mol/L of $HCl_{(aq)}$ at:

a. 0 mL of $HCl_{(aq)}$ added.
b. 7.50 mL of $HCl_{(aq)}$ added.
c. 15.00 mL of $HCl_{(aq)}$ added.
d. 20.00 mL of $HCl_{(aq)}$ added.

a. At 0 mL of $HCl_{(aq)}$ added, the pH will be calculated the weak base dissociation in the flask:

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

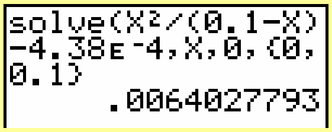
	$[CH_3NH_2]$		$[CH_3NH_3{}^+]$	$[OH^-]$
Initial **Change**	0.100 M $-x$		0 $+x$	0 M $+x$
Equilibrium	$(0.1 - x)$		x	x

CANNOT use Approximation:

$$\frac{[CH_3NH_2]_0}{K_b} = \frac{0.100\text{ M}}{4.38 \times 10^{-4}}$$

$$= 228.3 < 1000$$

$$K_b = \frac{[CH_3NH_3{}^+][OH^-]}{[CH_3NH_2]} \qquad 4.38 \times 10^{-4} = \frac{(x)(x)}{(0.1 - x)} = \frac{x^2}{(0.1 - x)}$$



```
solve(X²/(0.1-X)
-4.38E-4,X,0,{0,
0.1}
           .006402793
```

$[OH^-] = x = 0.00640$ mol/L

pOH = –log [OH⁻] = –log(0.00640) = 2.19

pH = 14 – pOH = 14 – 2.19

pH = 11.81

b. At 7.50 mL of HCl $_{(aq)}$ added: **(Before Stoichiometric Point: at Halfway to Equivalence Point)**

Chemical Species Present: CH_3NH_3, H_3O^+, Cl^-, H_2O

SB SA A/B

Net-Ionic Equation: $CH_3NH_{2\,(aq)} + H_3O^+_{(aq)} \rightarrow H_2O_{(l)} + CH_3NH_3^+{}_{(aq)}$

	***n* of CH_3NH_2**	***n* of H_3O^+**		***n* of $CH_3NH_3^+$**
Before	(0.100 mol/L)(30.00 mL) = 3.00 mmol	(0.200 mol/L)(7.50 mL) = 1.50 mmol (LR)		0 mmol
Change	**– 1.50 mmol**	**– 1.50 mmol**		**+ 1.50 mmol**
After	**1.50 mmol**	**0 mmol**		**1.50 mmol**

$$[CH_3NH_2] = \frac{n_{CH_3NH_2}}{\text{Total Volume}} = \frac{1.50\ \cancel{m}\text{mol}}{(30.00\ \cancel{m}\text{L} + 7.50\ \cancel{m}\text{L})} = 0.0400\text{ mol/L}$$

$$[CH_3NH_3^+] = \frac{n_{CH_3NH_3^+}}{\text{Total Volume}} = \frac{1.50\ \cancel{m}\text{mol}}{(30.00\ \cancel{m}\text{L} + 7.50\ \cancel{m}\text{L})} = 0.0400\text{ mol/L}$$

$[CH_3NH_3^+] = [CH_2NH_2]$

pH ≈ pK_a

(Max Buffer Capacity)

After all the H_3O^+ is used up, the resulting solution follows the regular weak base dissociation.

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

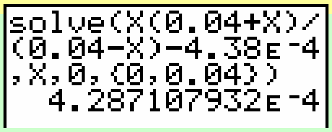
	$[CH_3NH_2]$		**$[CH_3NH_3^+]$**	**$[OH^-]$**
Initial	**0.0400 M**		**0**	**0.0400 M**
Change	**$-x$**		$+x$	**$+x$**
Equilibrium	**$(0.04 - x)$**		x	**$(0.04 + x)$**

CANNOT use Approximation:

$$\frac{[CH_3NH_2]_0}{K_b} = \frac{0.0400\text{ M}}{4.38\times10^{-4}}$$

= 91.3 < 1000

$$K_b = \frac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]} \qquad 4.38 \times 10^{-4} = \frac{(x)(0.04+x)}{(0.04-x)}$$



```
solve(X(0.04+X)/
(0.04-X)-4.38E-4
,X,0,{0,0.04})
  4.287107932E-4
```

$[OH^-] = x = 4.29 \times 10^{-4}$ mol/L

pOH = –log $[OH^-]$ = –log(4.29 × 10^{-4}) = 3.37

pH = 14 – pOH = 14 – 3.37

(pH's results from similar Calculations with other volumes before the stoichiometric point is listed on the next page.)

pH = 10.63

c. At 15.00 mL of HCl $_{(aq)}$ added: **(Stoichiometric Point)**

Chemical Species Present: CH_3NH_2, H_3O^+, Cl^-, H_2O

SB SA A/B

Net-Ionic Equation: $CH_3NH_{2\,(aq)} + H_3O^+_{(aq)} \rightarrow H_2O_{(l)} + CH_3NH_3^+{}_{(aq)}$

	***n* of CH_3NH_2**	***n* of H_3O^+**		***n* of $CH_3NH_3^+$**
Before	(0.100 mol/L)(30.00 mL) = 3.00 mmol	(0.200 mol/L)(15.00 mL) = 3.00 mmol		0 mmol
Change	**– 3.00 mmol**	**– 3.00 mmol**		**+ 3.00 mmol**
After	**0 mmol**	**0 mmol**		**3.00 mmol**

$$[CH_3NH_3^+] = \frac{n_{CH_3NH_3^+}}{\text{Total Volume}} = \frac{3.00\ \cancel{m}\text{mol}}{(30.00\ \cancel{m}\text{L} + 15.00\ \cancel{m}\text{L})} = 0.0667\text{ mol/L}$$

At Stoichiometric Point, after all the H_3O^+ and CH_3NH_2 are used up, the resulting solution follows the regular weak acid dissociation.

$$CH_3NH_3^+{}_{(aq)} + H_2O_{(l)} \rightleftharpoons CH_3NH_2{}_{(aq)} + H_3O^+{}_{(aq)}$$

	$[CH_3NH_3^+]$		$[CH_3NH_2]$	$[H_3O^+]$
Initial	**0.0667 M**		**0**	**0 M**
Change	$-x$		$+x$	$+x$
Equilibrium	$(0.0667 - x)$		x	x

CAN use Approximation:

$$\frac{[CH_3NH_3^+]_0}{K_a} = \frac{0.0667\text{ M}}{2.28\times10^{-11}}$$

$$= 2.9 \times 10^9 \geq 1000$$

Use 0.0667 in the denominator, because $(0.0667 - x) \approx 0.03667$ [x is so small compared to 0.0667 M].

$$K_a = \frac{K_w}{K_b} = \frac{1.0\times10^{-14}}{4.38\times10^{-4}} \qquad K_a = 2.28 \times 10^{-11} = \frac{[CH_2NH_2][H_3O^+]}{[CH_3NH_3]}$$

$$2.28 \times 10^{-11} = \frac{(x)(x)}{(0.0667 - x)} \approx \frac{x^2}{0.0667}$$

$[H_3O^+] = x = 1.23 \times 10^{-6}$ mol/L $\quad$ $pH = -\log[H_3O^+] = -\log(1.23 \times 10^{-6})$ $\quad$ $\boxed{pH = 5.91}$

d. At 20.00 mL of $HCl_{(aq)}$ added: **(After Stoichiometric Point)**

Chemical Species Present: $\quad$ **CH_3NH_2, H_3O^+**, Cl^-, H_2O

SB	**SA**		A/B

Net-Ionic Equation: $\quad CH_3NH_2{}_{(aq)} + H_3O^+{}_{(aq)} \rightarrow H_2O_{(l)} + CH_3NH_3^+{}_{(aq)}$

	***n* of CH_3NH_2**	***n* of H_3O^+**
Before	(0.100 mol/L)(30.00 mL) = 3.00 mmol (LR)	(0.200 mol/L)(20.00 mL) = 4.00 mmol
Change	**– 3.00 mmol**	**– 3.00 mmol**
After	**0 mmol**	**1.00 mmol**

After all the CH_3NH_2 is used up, the resulting solution has the following major species.

Chemical Species Present: $\quad Cl^-$, **H_3O^+**, H_2O

SA	A/B

$$[H_3O^+] = \frac{n_{H_3O^+}}{\text{Total Volume}} = \frac{1.00\ \cancel{m}\text{mol}}{(30.00\ \cancel{m}\text{L} + 20.00\ \cancel{m}\text{L})} = 0.0200\text{ mol/L}$$

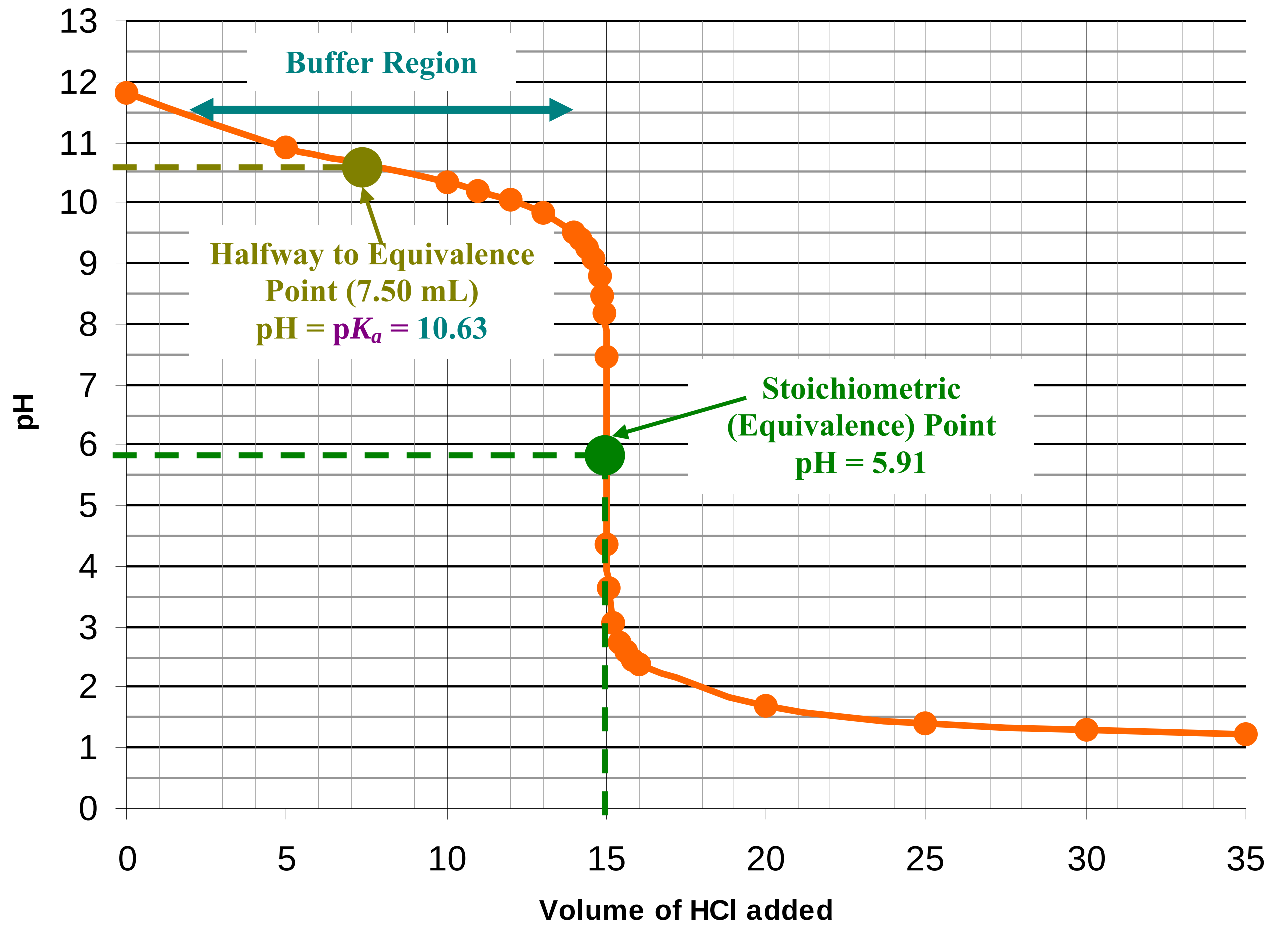
$pH = -\log[H_3O^+] = -\log(0.0200)$ $\quad$ $\boxed{pH = 1.70}$

(pH's results from similar Calculations with other volumes after the stoichiometric point is listed below.)

pH of 30.0 mL of 0.100 M $CH_3NH_2{}_{(aq)}$ titrated by 0.200 M of $HCl_{(aq)}$

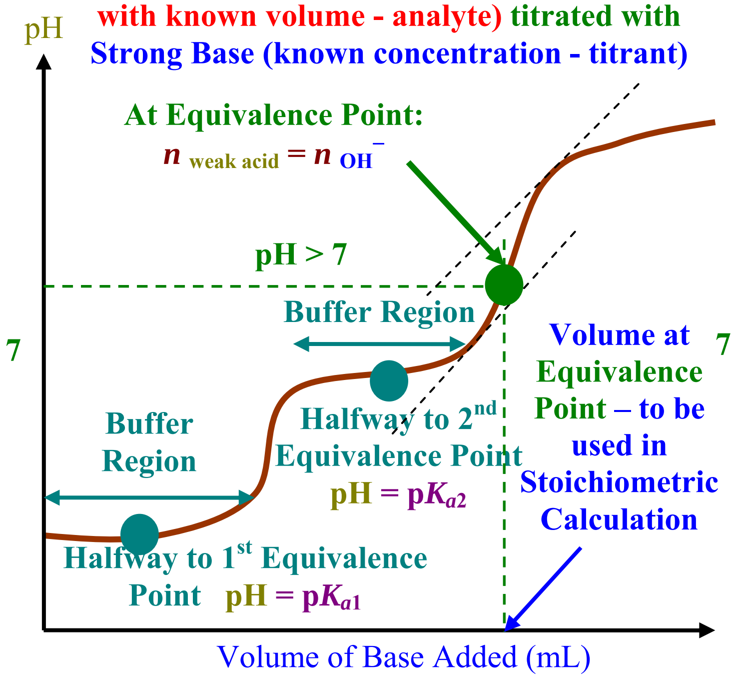
Volume of HCl added	pH	Volume of HCl added	pH	Volume of HCl added	pH
0.00 mL	11.81	14.40 mL	9.26	15.20 mL	3.05
5.00 mL	10.92	14.60 mL	9.08	15.40 mL	2.75
7.50 mL (Halfway to Equivalence)	10.63	14.80 mL	8.77	15.60 mL	2.58
10.00 mL	10.33	14.90 mL	8.47	15.80 mL	2.46
11.00 mL	10.20	14.95 mL	8.16	16.00 mL	2.36
12.00 mL	10.03	14.99 mL	7.46	20.00 mL	1.70
13.00 mL	9.82	**15.00 mL (Equivalence)**	**5.91**	25.00 mL	1.40
14.00 mL	9.49	15.01 mL	4.35	30.00 mL	1.30
14.20 mL	9.39	15.05 mL	3.65	35.00 mL	1.21

pH of 30.0 mL of 0.100 M Methylamine titrated by 0.200 M of HCl

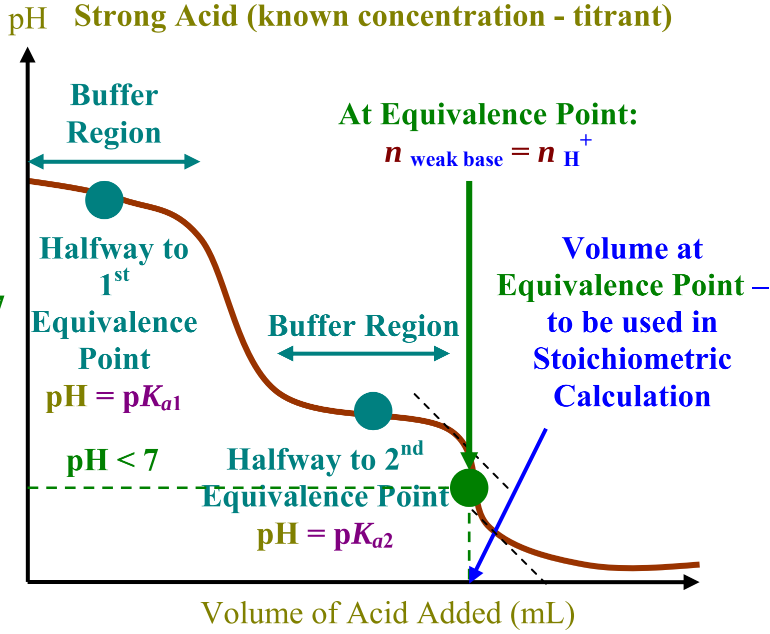


4. **Titration Between Strong Acid/Bases and Polyprotic Acid/Base**: - there are multiple equivalence points.
   - the last equivalence point indicates the stoichiometric volume.

Weak Polyprotic Acid (unknown concentration with known volume - analyte) titrated with Strong Base (known concentration - titrant)



Weak Polyprotic Base (unknown concentration with known volume - analyte) titrated with Strong Acid (known concentration - titrant)



Example 4: Calculate the concentration of 10.00 mL of $Na_2S_{(aq)}$ is titrated with 0.0350 M of $HBr_{(aq)}$ when the first and second equivalence points are at 13.34 mL and 27.88 mL respectively.

For Acid-Base Stoichiometry, do NOT write the Net-Ionic Equation. Write the Molecular Equation to do Stoichiometry. We use the second equivalence point because Na_2S can accept two protons.

$2\ HBr_{(aq)}$	+	$Na_2S_{(aq)}$	→	$H_2S_{(aq)}$	+	$2\ NaBr_{(aq)}$
27.88 mL		10.0 mL				
0.0350 mol/L		? mol/L				

① $n_{HBr} = CV = (0.0350\ \text{mol}/\cancel{\text{L}})\ (27.88\ \text{m}\cancel{\text{L}}) = 0.9758\ \text{mmol}$

② $n_{Na_2S} = 0.9758\ \text{m}\cancel{\text{mol HBr}} \times \dfrac{1\ \text{mol}\ Na_2S}{2\ \cancel{\text{mol HBr}}} = 0.4879\ \text{mmol}\ Na_2S$

③ $[Na_2S] = \dfrac{n}{V} = \dfrac{0.4879\ \cancel{\text{m}}\text{mol}}{10.0\ \cancel{\text{m}}\text{L}} = 0.04879\ \text{mol/L}$

$\boxed{[Na_2S] = 0.0488\ \text{mol/L}}$

Neutralization Reactions Between Weak Acids and Weak Bases

1. **The reaction is NOT one way, but in equilibrium.**
2. When writing out the equilibrium system, **list all major species first (Do NOT Break up any Weak Acids and Weak Bases)**. Then, **reacts the Strongest Acid (SA – donates a proton) and the Strongest Base (SB – accepts a proton) from the list**.
3. Using the Table of the Relative Strengths of Acids, decide on **which side of the equilibrium is the system favoured**. **(The equilibrium favours the side with the Weakest Acid (Lower K_a) – because the stronger acid of the other side dissociates more.)**

Weak Acid and Weak Base Neutralization

$$\overset{\Longrightarrow}{\underset{\text{(Stronger Acid)}}{HA_{(aq)}} + B_{(aq)} \rightleftharpoons A^{-}_{(aq)} + \underset{\text{(Weaker Acid)}}{HB^{+}_{(aq)}}}$$

$$\overset{\Longleftarrow}{\underset{\text{(Weaker Acid)}}{HX_{(aq)}} + Y_{(aq)} \rightleftharpoons X^{-}_{(aq)} + \underset{\text{(Stronger Acid)}}{HY^{+}_{(aq)}}}$$

Example 5: Write the net-ionic equation for the following neutralizations. Indicate the side of the equilibrium system that is favoured.

a. Acetic Acid reacts with Ammonia

Chemical Species Present: $\underset{}{\overset{\text{SA}}{CH_3COOH}},\ \underset{\text{SB}}{NH_3},\ \underset{\text{B}}{\overset{\text{A}}{H_2O}}$

Net-Ionic Equation:

$\Longrightarrow$ **Favour Products**

$$\underset{\substack{\text{(Stronger Acid)}\\ K_a = 1.8 \times 10^{-5}}}{CH_3COOH_{(aq)}} + NH_{3\,(aq)} \rightleftharpoons CH_3COO^{-}_{(aq)} + \underset{\substack{\text{(Weaker Acid)}\\ K_a = 5.6 \times 10^{-10}}}{NH_4^{+}{}_{(aq)}}$$

b. Sodium Hydrogen Oxalate reacts with Calcium Sulfite

Chemical Species Present: Na^+, $\underset{\text{B}}{\overset{\text{SA}}{\mathbf{HOOCCOO^-}}}$, Ca^{2+}, $\underset{\textbf{SB}}{\mathbf{SO_3^{2-}}}$, $\underset{\text{B}}{\overset{\text{A}}{H_2O}}$

Net-Ionic Equation:

$\longrightarrow$ **Favour Products**

$$\mathbf{HOOCCOO^-_{(aq)} + SO_3^{2-}{}_{(aq)} \rightleftharpoons OOCCOO^{2-}{}_{(aq)} + HSO_3^-{}_{(aq)}}$$

(Stronger Acid) — **(Weaker Acid)**

$K_a = 1.5 \times 10^{-4}$ — $K_a = 6.3 \times 10^{-8}$

c. Sodium Hydrogen Carbonate reacts with Potassium Hydrogen Sulfite

Chemical Species Present: Na^+, $\underset{\textbf{SB}}{\overset{\text{A}}{\mathbf{HCO_3^-}}}$, K^+, $\underset{\text{B}}{\overset{\text{SA}}{\mathbf{HSO_3^-}}}$, $\underset{\text{B}}{\overset{\text{A}}{H_2O}}$

Net-Ionic Equation:

Favour Reactants $\longleftarrow$

$$\mathbf{HSO_3^-{}_{(aq)} + HCO_3^-{}_{(aq)} \rightleftharpoons SO_3^{2-}{}_{(aq)} + H_2CO_3{}_{(aq)}}$$

(Weaker Acid) — **(Stronger Acid)**

$K_a = 6.3 \times 10^{-8}$ — $K_a = 4.5 \times 10^{-7}$

Example 6: Four different weak acids are reacted with various conjugate bases, and the results are recorded as follow.

$HW_{(aq)} + X^-_{(aq)} \rightleftharpoons HX_{(aq)} + W^-_{(aq)}$ [products favoured]

$HY_{(aq)} + W^-_{(aq)} \rightleftharpoons HW_{(aq)} + Y^-_{(aq)}$ [products favoured]

$HZ_{(aq)} + Y^-_{(aq)} \rightleftharpoons HY_{(aq)} + Z^-_{(aq)}$ [reactants favoured]

$HZ_{(aq)} + X^-_{(aq)} \rightleftharpoons HX_{(aq)} + Z^-_{(aq)}$ [reactants favoured]

List the acids in order of **increasing** strength.

First, we have to decide which acid is the weaker acid for each of these reactions.

$\longrightarrow$

$\mathbf{HW}_{(aq)} + X^-_{(aq)} \rightleftharpoons \mathbf{HX}_{(aq)} + W^-_{(aq)}$ [products favoured]

(Stronger Acid) **(Weaker Acid)** **HW > HX**

$\longrightarrow$

$\mathbf{HY}_{(aq)} + W^-_{(aq)} \rightleftharpoons \mathbf{HW}_{(aq)} + Y^-_{(aq)}$ [products favoured]

(Stronger Acid) **(Weaker Acid)** **HY > HW**

$\longleftarrow$

$\mathbf{HZ}_{(aq)} + Y^-_{(aq)} \rightleftharpoons \mathbf{HY}_{(aq)} + Z^-_{(aq)}$ [reactants favoured]

(Weaker Acid) **(Stronger Acid)** **HY > HZ**

$\longleftarrow$

$\mathbf{HZ}_{(aq)} + X^-_{(aq)} \rightleftharpoons \mathbf{HX}_{(aq)} + Z^-_{(aq)}$ [reactants favoured]

(Weaker Acid) **(Stronger Acid)** **HX > HZ**

Putting the inequalities together, we obtain in order of increasing acid strengths:

$$\boxed{\mathbf{HZ < HX < HW < HY}}$$

Assignment

17.3 pg. 604 #17 to 20a

17.4: Acid-Base Indicators

Acid-Base Indicators: - chemicals that change colours at a specific pH range.
- they are themselves organic acids. Since they are usually very big structurally, we usually use abbreviations to describe them in chemical equations.
- this is due to the acidic form of the indictor (HIn) has a different colour than its basic form (In^-).

Example: Phenolphthalein ($HPh_{(aq)}$ /$Ph^-_{(aq)}$) changes colours from colourless to pink at a pH range of 8.0 to 10.0. At its **mid-range (9.0 – average of 8.0 and 10.0)**, the colour would be **light pink (colourless + pink)**.

$\mathbf{HPh}_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + \mathbf{Ph}^-_{(aq)}$ $K_a = 1.0 \times 10^{-9}$ $pK_a = 9$ Colour Change = 9 ± 1
(Colourless) (Pink)

At pH ≤ 9, phenolphthalein is colourless. At pH ≥ 10, phenolphthalein is pink.

Example: Bromothymol Blue ($HBb_{(aq)}$ /$Bb^-_{(aq)}$) changes colours from yellow to blue at a pH range of 6.0 to 8.0. At its **mid-range (7.0 – average of 6.0 and 8.0)**, the colour would be **green (yellow + blue)**.

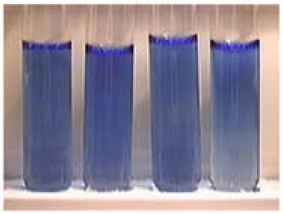
$\mathbf{HBb}_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + \mathbf{Bb}^-_{(aq)}$ $K_a = 1.0 \times 10^{-7}$ $pK_a = 7$ Colour Change = 7 ± 1
(Yellow) **(Blue)**

At pH ≤ 6, bromothymol blue is yellow. At pH ≥ 8, bromothymol blue is blue.

1 to 4

5 to 9

10 to 13



Example 1: A sample of rainwater is poured into five test tubes. A different indicator is added to each test tube. Four of the observations are recorded in the table below.

methyl red	yellow
phenol red	yellow
bromocresol green	blue
phenolphthalein	colourless
bromothymol blue	?

Determine the pH of the rainwater and the predicted colour of the sample containing bromothymol blue.

First, we need to decide on the pH range that these indicators will change colors.

Indicators	Colour Change	pH range	Colour Observed	Indicated pH
methyl red	red to yellow	4.8 to 6.0	Yellow	pH > 6.0
phenol red	yellow to red	6.6 to 8.0	Yellow	pH < 6.6
bromocresol green	yellow to blue	3.8 to 5.4	Blue	pH > 5.4
phenolphthalein	colourless to pink	8.2 to 10.0	colourless	pH < 8.2

From the indicated pH ranges, we can tell than the pH of the rain water is **between 6.0 to 6.6**. At such pH range, **bromothymol blue would appear green**. (The pH and colour range for bromothymol blue is 6.0 – 7.6; yellow to blue)

Assignment
17.4 pg. 604 #21 to 24

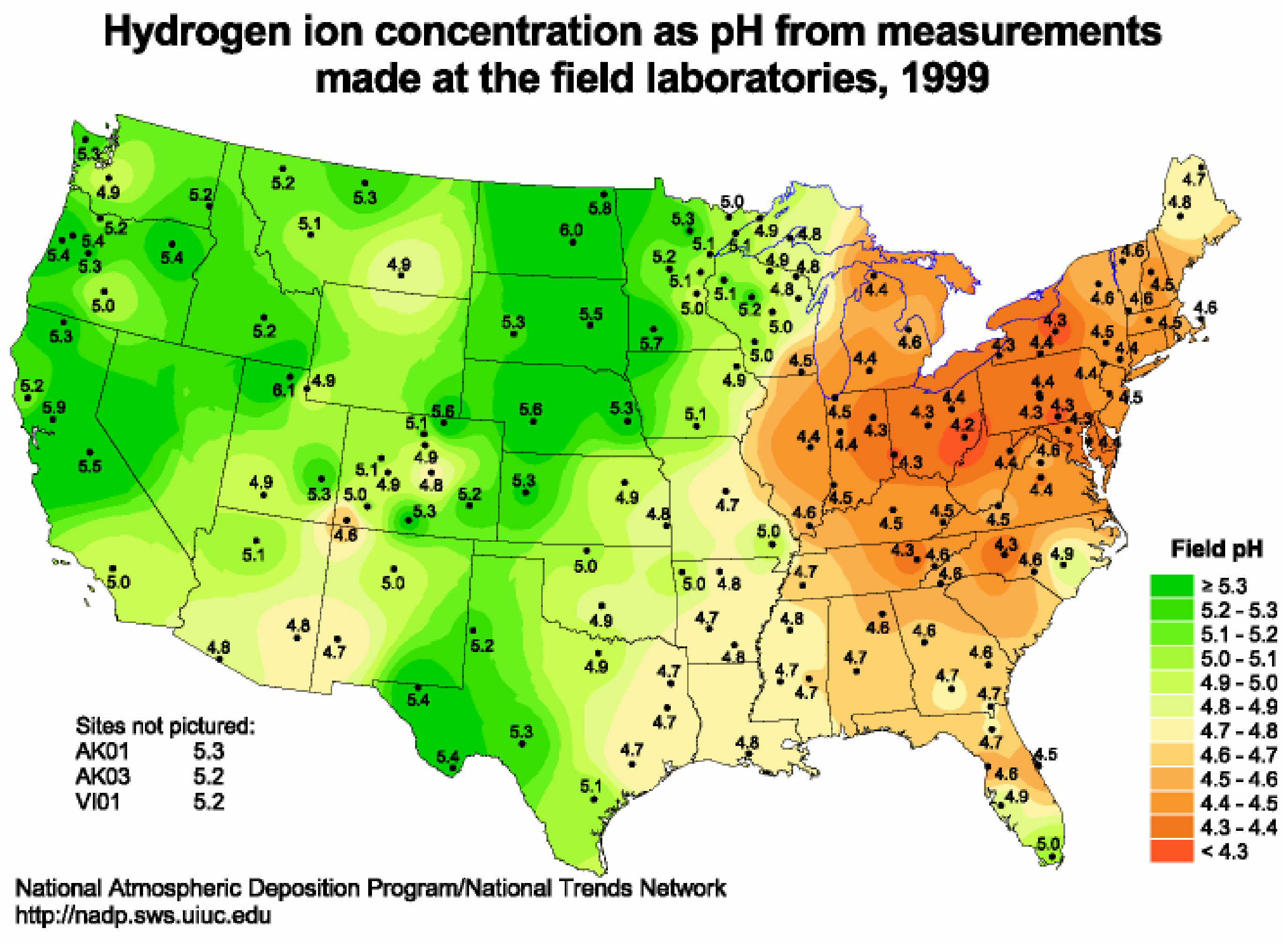
Extra Section: Environmental Issues Involving Acids and Bases

Acid Rain: - precipitation of water (rain) or (snow) that has a low pH (less than 5.0).
- natural rain is slightly acidic due to the small amount of CO_2 in the atmosphere.

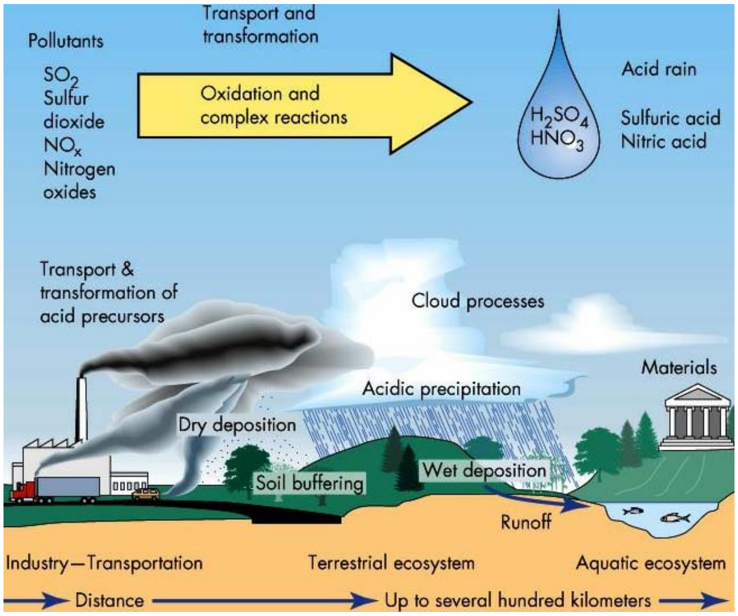
$$CO_{2\,(g)} + H_2O_{\,(l)} \rightarrow H_2CO_{3\,(aq)} \qquad \text{(Carbonic Acid} - K_a = 4.5 \times 10^{-7}\text{)}$$

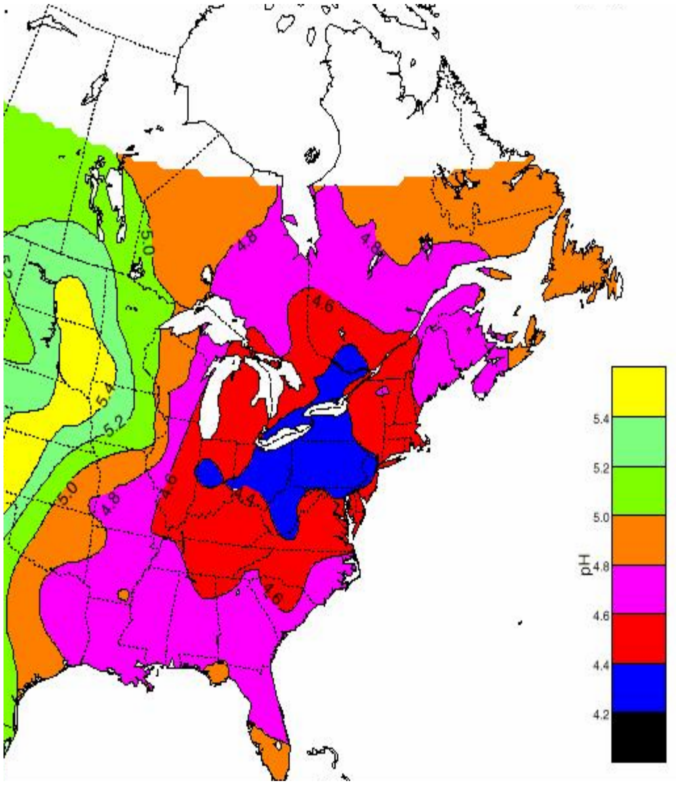
- however, the burning of fossil fuel by vehicles and industrial activities, and also emits other oxides such as $NO_{2\,(g)}$, $SO_{2\,(g)}$ and $SO_{3\,(aq)}$. These gaseous oxides (Lewis acids) combined with water in the atmosphere (Lewis base) to form various strong and weak acids. Thus, lowering the pH of rain and snow.

$$SO_{2\,(g)} + H_2O_{\,(l)} \rightarrow H_2SO_{3\,(aq)} \qquad \text{(Sulfurous Acid} - K_a = 1.4 \times 10^{-2}\text{)}$$
$$SO_{3\,(g)} + H_2O_{\,(l)} \rightarrow H_2SO_{4\,(aq)} \qquad \text{(Sulfuric Acid} - \underline{\text{Strong Acid}}\text{)}$$
$$2\ NO_{2\,(g)} + H_2O_{\,(l)} \rightarrow HNO_{3\,(aq)} + HNO_{2\,(aq)}$$
$$\text{(Nitric Acid} - \underline{\text{Strong Acid}} \text{ and Nitrous Acid} - K_a = 4.5 \times 10^{-7}\text{)}$$



(Above) Measurements of pH in the continental U.S in 1999. Note the most severely affected regions are concentrated in the East Coast as well as in the major cities.





(Above) The Formation of Acid Rain.
(Right) Measurements of pH in Eastern Canada and U.S in 1996. The lowest pHs are in the Great Lakes Region where most of the heavy industries are located.

Effects of Acid Rain:

1. **Marine Life (especially in-land lakes and rivers) is affected greatly**. Acid Rain lowers the pH that can destroy all life (plants and animals). Usually, acid lake and river appears to be crystal clear, with no sign of algae, plant and animal life.
2. **Vegetation (including forests) can be decimated**. Since the **soil acidity level are lowered**, further plant growth may not be possible and the surrounding **ecosystem is destroyed** as a result. Forests affected by acid rain appear to be charred with trees having no leaves and all branches are brittle.

(Left) an Acid Lake. Note the clear water with the absence of life. (Right) an Acid Forest. Note the lack of leaves, the brittle branches, and the charred looked of the tree trunks.

3. In areas of where there is a sustain amount of humidity, acid rain can turn into **Acid Fog or Smog (SO_2 + H_2O) – Combination of *Sm*oke from Coal Burning and F*og***. **Humans may experience severe respiratory problems as a result.** The most notable case was in London, England and surrounding areas in December 1952 when smog was so serious that there were estimated 4700 deaths.

4. **Acid deposition** (acid rain depositing onto a solid – metal or base that causes corrosion and neutralization) on city structures such as historical statues (usually made of limestone, $CaCO_3$) and buildings can cause expensive repairs and lost of priceless arts.

(Above) Smog in London during December, 1952. (Below) Smog over the Los Angeles skyline. (Right) Acid Depositions on Statues at the Schloss Sanssouci Palace in Berlin, Germany. Note the black "burn" deposits on the statutes.

Ways to Reduce Acid Rain:

1. **Burn Less Fossil Fuel**: drive more fuel efficient vehicles, develop alternate ways to generate electricity without combustion of fossil fuel, and reduce the amount of electricity used are all possible ways to cut down on gaseous oxides emissions.

2. **Enact Legislation to Regulate Emissions**: governments can pass legislation on industry to regulate the amount of gases that cause acid rain. Better filtering technology and frequent inspections of industrial facilities can hold companies accountable to these environmental laws.

3. **Adding Limestone** to acid lakes and rivers can raise the pH and reintroduce marine life that has been affected by acid rain.