

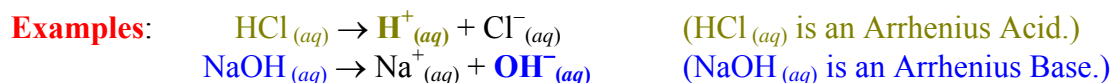
Unit 6: ACIDS AND BASESChapter 16: Acids and Bases16.1: Brønsted Acids and BasesPhysical and Chemical Properties of Acid and Base

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



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.

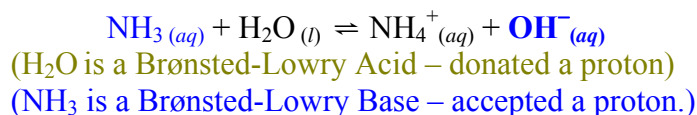
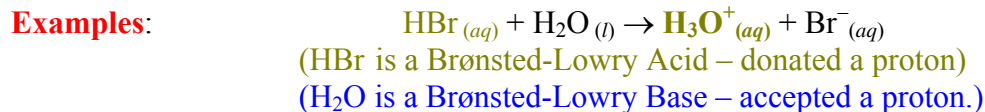


Brønsted-Lowry Model: - acids and bases **react with water** to **dissociate** where **acids are H^+ (proton) donors** and **bases are H^+ (proton) acceptors**.

- 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)**.

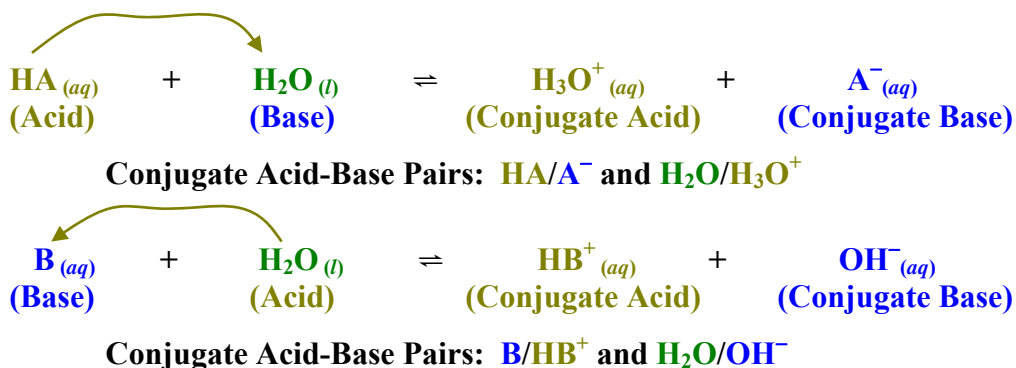
- essentially has the same function as a H^+ ion, but H_3O^+ denotes that we are using the Brønsted-Lowry model.



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

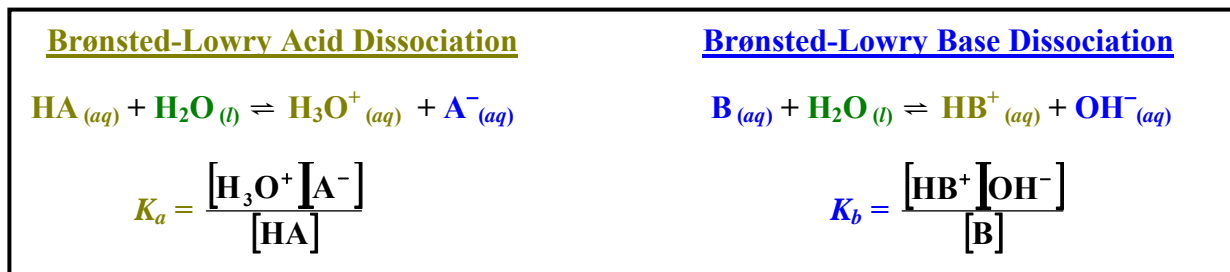
Conjugate Acid: - the product formed after the Base accepted a H^+ . (Base \rightarrow Conjugate Acid)
 - behaves like an acid when the reaction is looking from reverse.

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



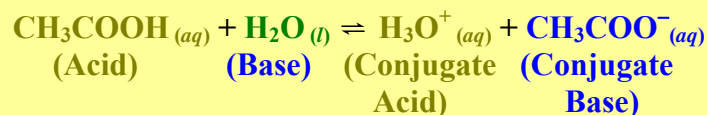
Acid Dissociation Constant (K_a): - the equilibrium constant of a Brønsted-Lowry Acid Dissociation.

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



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.

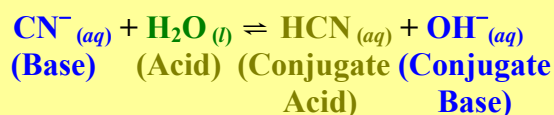
a. Acetic acid ($\text{CH}_3\text{COOH}_{(aq)}$)



Conjugate Acid-Base Pairs:
 CH₃COOH/CH₃COO⁻ and H₂O/H₃O⁺

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]}$$

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



Conjugate Acid-Base Pairs:
 CN⁻/HCN and H₂O/OH⁻

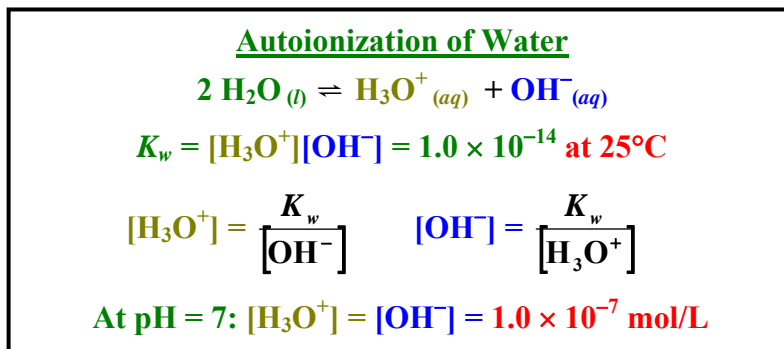
$$K_b = \frac{[\text{HCN}][\text{OH}^-]}{[\text{CN}^-]}$$

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)**, $[\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \text{ mol/L}$



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

- a. Using the ICE box, show that $[\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \text{ mol/L}$ for a neutral solution.
 b. At 100°C , $K_w = 8.19 \times 10^{-13}$. What is the $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ for a neutral solution at 100°C ?

a.

	$2 \text{H}_2\text{O} (l) \rightleftharpoons \text{H}_3\text{O}^+ (aq) + \text{OH}^- (aq)$		
	$\text{H}_2\text{O} (l)$	$[\text{H}_3\text{O}^+]$	$[\text{OH}^-]$
Initial	----	0	0
Change	----	+x	+x
Equilibrium	-----	x	x

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

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

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

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

$x = [\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \text{ M}$

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

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

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

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

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

$x = [\text{H}_3\text{O}^+] = [\text{OH}^-] = 2.86 \times 10^{-7} \text{ M}$

Example 2: Determine the $[\text{H}_3\text{O}^+]$ and/or $[\text{OH}^-]$ concentrations of the following solutions at 25°C .

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

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-4}}$$

$[\text{H}_3\text{O}^+] = 1.0 \times 10^{-10} \text{ mol/L}$

b. $[\text{H}_3\text{O}^+] = 5.0 \times 10^{-5} \text{ M}$

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

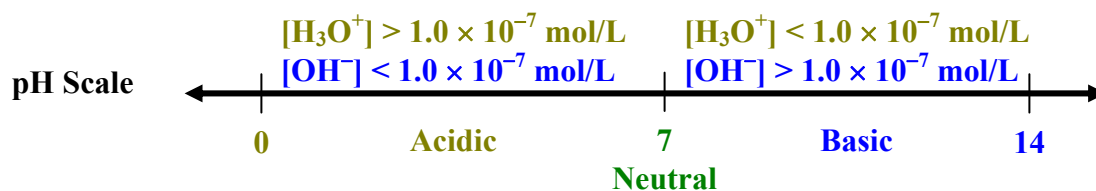
$$[\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} = \frac{1.0 \times 10^{-14}}{5.0 \times 10^{-5}}$$

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

16.3: pH – A Measure of Acidity

pH Scale: - a logarithmic scale to **measure the acidity (relative $[\text{H}_3\text{O}^+]$)** of a solution.

- the **lower the pH, the more acidic** (less basic) is the solution (**more $[\text{H}_3\text{O}^+]$ and less $[\text{OH}^-]$**).
- the **higher the pH, the more basic** (less acidic) is the solution (**less $[\text{H}_3\text{O}^+]$ and more $[\text{OH}^-]$**).
- **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 \neq 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 $[\text{H}_3\text{O}^+]$ by a factor of 10; an increase of 2 on a pH scale means a decrease of $[\text{H}_3\text{O}^+]$ by a factor of 100.**



Example: pH of Some Common Substances

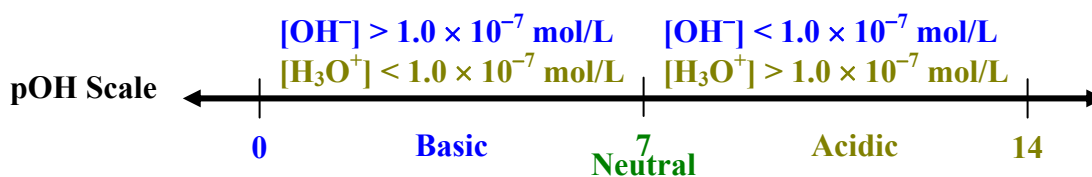
Substance	pH
1 M of HCl	0.00
Stomach Acid	2.00
Lemon Juice	2.50
Vinegar	3.00

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 $[\text{OH}^-]$)** of a solution.

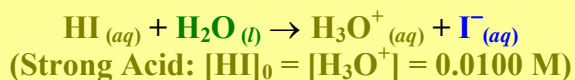
- the **lower the pOH, the more basic** (less acidic) is the solution (**more $[\text{OH}^-]$ and less $[\text{H}_3\text{O}^+]$**).
- the **higher the pOH, the less basic** (more acidic) is the solution (**less $[\text{OH}^-]$ and more $[\text{H}_3\text{O}^+]$**).
- **basicity is NOT the same as the base strength**. Just because a solution has a low pOH, it does not mean that it is a strong base. (**Highly Basic \neq 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 $[\text{OH}^-]$ by a factor of 10; an increase of 2 on a pOH scale means a decrease of $[\text{OH}^-]$ by a factor of 100.**



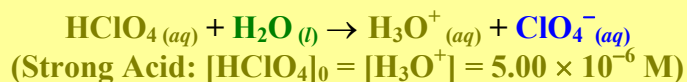
pH and pOH Scales	
$\text{pH} = -\log [\text{H}_3\text{O}^+]$	$\text{pOH} = -\log [\text{OH}^-]$
$\text{pH} + \text{pOH} = 14.00$	

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

- a. 0.0100 mol/L of HI_(aq) (completely dissociates) b. 5.00×10^{-6} M of HClO_{4(aq)} (completely dissociates)



$$\begin{aligned} \text{pH} &= -\log [\text{H}_3\text{O}^+] & \text{pOH} &= 14 - \text{pH} \\ \text{pH} &= -\log(0.0100) & \text{pOH} &= 14 - 2.00 \\ \text{pH} &= 2.00 & \text{pOH} &= 12.00 \end{aligned}$$



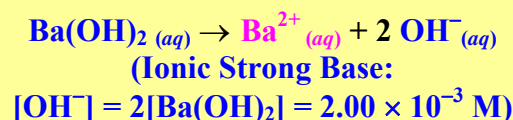
$$\begin{aligned} \text{pH} &= -\log [\text{H}_3\text{O}^+] & \text{pOH} &= 14 - \text{pH} \\ \text{pH} &= -\log(5.00 \times 10^{-6}) & \text{pOH} &= 14 - 5.30 \\ \text{pH} &= 5.30 & \text{pOH} &= 8.70 \end{aligned}$$

- c. 1.50 M of KOH_(aq)



$$\begin{aligned} \text{pOH} &= -\log [\text{OH}^-] & \text{pH} &= 14 - \text{pOH} \\ \text{pOH} &= -\log(1.50) & \text{pH} &= 14 - (-0.176) \\ \text{pOH} &= -0.176 & \text{pH} &= 14.176 \end{aligned}$$

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



$$\begin{aligned} \text{pOH} &= -\log [\text{OH}^-] & \text{pH} &= 14 - \text{pOH} \\ \text{pOH} &= -\log(2.00 \times 10^{-3}) & \text{pH} &= 14 - (2.70) \\ \text{pOH} &= 2.70 & \text{pH} &= 11.30 \end{aligned}$$

Example 2: Calculate the $[\text{H}_3\text{O}^+]$ and the $[\text{OH}^-]$ for the following solutions.

- a. pH = 4.00

$$\begin{aligned} \text{pH} &= -\log [\text{H}_3\text{O}^+] \\ [\text{H}_3\text{O}^+] &= 10^{-\text{pH}} = 10^{-4.00} \\ [\text{H}_3\text{O}^+] &= 1.00 \times 10^{-4} \text{ M} \end{aligned}$$

$$\begin{aligned} \text{pOH} &= 14 - \text{pH} \\ \text{pOH} &= 14 - 4.00 = 10.00 \end{aligned}$$

$$\begin{aligned} \text{pOH} &= -\log [\text{OH}^-] \\ [\text{OH}^-] &= 10^{-\text{pOH}} = 10^{-10.00} \\ [\text{OH}^-] &= 1.00 \times 10^{-10} \text{ M} \end{aligned}$$

- b. pOH = 3.00

$$\begin{aligned} \text{pOH} &= -\log [\text{OH}^-] \\ [\text{OH}^-] &= 10^{-\text{pOH}} = 10^{-3.00} \\ [\text{OH}^-] &= 1.00 \times 10^{-3} \text{ M} \end{aligned}$$

$$\begin{aligned} \text{pH} &= 14 - \text{pOH} \\ \text{pH} &= 14 - 3 = 11.00 \end{aligned}$$

$$\begin{aligned} \text{pH} &= -\log [\text{H}_3\text{O}^+] \\ [\text{H}_3\text{O}^+] &= 10^{-\text{pH}} = 10^{-11.00} \\ [\text{H}_3\text{O}^+] &= 1.00 \times 10^{-11} \text{ M} \end{aligned}$$

- c. pH = 12.83

$$\begin{aligned} \text{pH} &= -\log [\text{H}_3\text{O}^+] \\ [\text{H}_3\text{O}^+] &= 10^{-\text{pH}} = 10^{-12.83} \\ [\text{H}_3\text{O}^+] &= 1.48 \times 10^{-13} \text{ M} \end{aligned}$$

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

$$[\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} = \frac{1.0 \times 10^{-14}}{1.48 \times 10^{-13}}$$

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

- d. pOH = 9.67

$$\begin{aligned} \text{pOH} &= -\log [\text{OH}^-] \\ [\text{OH}^-] &= 10^{-\text{pOH}} = 10^{-9.67} \\ [\text{OH}^-] &= 2.14 \times 10^{-10} \text{ M} \end{aligned}$$

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{2.14 \times 10^{-10}}$$

$$[\text{H}_3\text{O}^+] = 4.68 \times 10^{-5} \text{ mol/L}$$

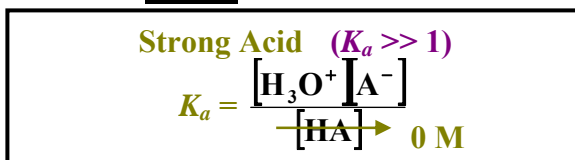
Assignment

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.**

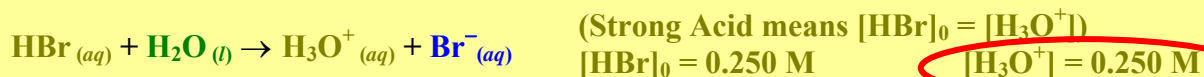
- when the **acid dissociation constant (K_a) is much greater than 1.** ($K_a \gg 1$)
- **the equilibrium position strongly favours the products.**
- at equilibrium, the original acid concentration, $[HA]_{eq} \approx 0$; $[H_3O^+]_{eq} = [A^-]_{eq} = [HA]_0$.
- the **conjugate base, A^-** , of a strong 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.**



	[HA]		[H ₃ O ⁺]	[A ⁻]
Initial	x		0	0
Change	$-x$		$+x$	$+x$
Equilibrium	0		x	x

Examples: Strong Acids: HClO₄ (aq), HI (aq), HBr (aq), HCl (aq), H₂SO₄ (aq) and HNO₃ (aq)

Example 2: Write the dissociation reaction of 0.250 M of HBr (aq) and determine its [H₃O⁺].



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.**



	[HA]		[H ₃ O ⁺]	[A ⁻]
Initial	x		0	0
Change	$-y$ (where $y \ll x$)		$+y$	$+y$
Equilibrium	$(x - y) \approx x$		y	y

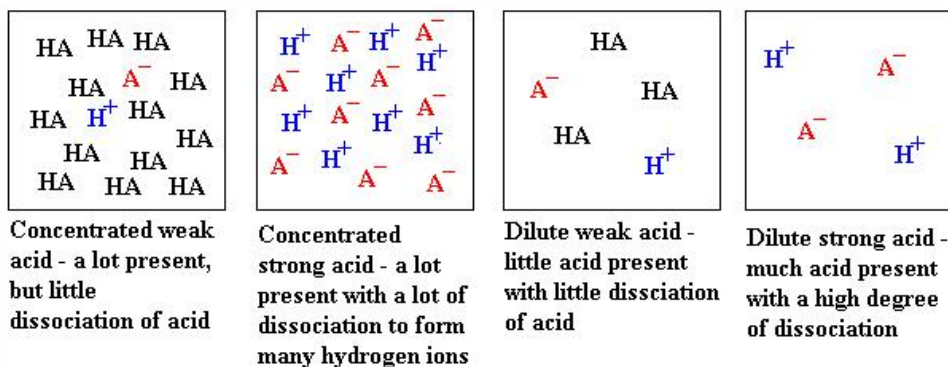
Examples: Some Weak Acids: HOCCOOH (aq), H₂SO₃ (aq), HSO₄⁻ (aq), H₃PO₄ (aq), HNO₂ (aq), H₃C₆H₅O₇ (aq), HF (aq), HCOOH (aq), C₆H₈O₆ (aq), C₆H₅COOH (aq), CH₃COOH (aq), H₂CO₃ (aq), H₂S (aq), HOCl (aq), HCN (aq), NH₄⁺ (aq), and H₃BO₃ (aq)

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

Acid Name	Acid Formula	Conjugate Base Formula	K_a
perchloric acid	$\text{HClO}_4 (aq)$	$\text{ClO}_4^- (aq)$	Very Large
hydroiodic acid	$\text{HI} (aq)$	$\text{I}^- (aq)$	Very Large
hydrobromic acid	$\text{HBr} (aq)$	$\text{Br}^- (aq)$	Very Large
hydrochloric acid	$\text{HCl} (aq)$	$\text{Cl}^- (aq)$	Very Large
sulfuric acid	$\text{H}_2\text{SO}_4 (aq)$	$\text{HSO}_4^- (aq)$	Very Large
nitric acid	$\text{HNO}_3 (aq)$	$\text{NO}_3^- (aq)$	Very Large
hydronium ion	$\text{H}_3\text{O}^+ (aq)$	$\text{H}_2\text{O} (l)$	1
oxalic acid	$\text{HOOC}(\text{COOH}) (aq)$	$\text{HOOC}(\text{COO})^- (aq)$	6.5×10^{-2}
sulfurous acid	$\text{H}_2\text{SO}_3 (aq)$	HSO_3^-	1.5×10^{-2}
hydrogen sulfate ion	$\text{HSO}_4^- (aq)$	SO_4^{2-}	1.2×10^{-2}
chlorous acid	$\text{HClO}_2 (aq)$	$\text{ClO}_2^- (aq)$	1.2×10^{-2}
phosphoric acid	$\text{H}_3\text{PO}_4 (aq)$	$\text{H}_2\text{PO}_4^- (aq)$	7.5×10^{-3}
arsenic acid	$\text{H}_3\text{AsO}_4 (aq)$	$\text{H}_2\text{AsO}_4^- (aq)$	5×10^{-3}
monochloroacetic acid	$\text{HC}_2\text{H}_2\text{ClO}_2$	$\text{C}_2\text{H}_2\text{ClO}_2^- (aq)$	1.35×10^{-3}
citric acid	$\text{H}_3\text{C}_6\text{H}_5\text{O}_7 (aq)$	$\text{H}_2\text{C}_6\text{H}_5\text{O}_7^- (aq)$	8.4×10^{-4}
hydrofluoric acid	$\text{HF} (aq)$	$\text{F}^- (aq)$	7.2×10^{-4}
nitrous acid	$\text{HNO}_2 (aq)$	$\text{NO}_2^- (aq)$	4.0×10^{-4}
methanoic (formic) acid	$\text{HCOOH} (aq)$	$\text{HCOO}^- (aq)$	1.8×10^{-4}
lactic acid	$\text{HC}_3\text{H}_5\text{O}_3 (aq)$	$\text{C}_3\text{H}_5\text{O}_3^- (aq)$	1.38×10^{-4}
ascorbic acid (vitamin C)	$\text{H}_2\text{C}_6\text{H}_6\text{O}_6 (aq)$	$\text{HC}_6\text{H}_6\text{O}_6^- (aq)$	7.9×10^{-5}
benzoic acid	$\text{C}_6\text{H}_5\text{COOH} (aq)$	$\text{C}_6\text{H}_5\text{COO}^- (aq)$	6.4×10^{-5}
hydrogen oxalate ion	$\text{HOOC}(\text{COO})^- (aq)$	$\text{OOC}(\text{COO})^{2-} (aq)$	6.1×10^{-5}
ethanoic (acetic) acid	$\text{CH}_3\text{COOH} (aq)$	$\text{CH}_3\text{COO}^- (aq)$	1.8×10^{-5}
dihydrogen citrate ion	$\text{H}_2\text{C}_6\text{H}_5\text{O}_7^- (aq)$	$\text{HC}_6\text{H}_5\text{O}_7^{2-} (aq)$	1.8×10^{-5}
hydrated aluminum (III) ion	$[\text{Al}(\text{H}_2\text{O})_6]^{3+}$	$[\text{AlOH}(\text{H}_2\text{O})_5]^{2+}$	1.4×10^{-5}
propanoic acid	$\text{C}_2\text{H}_5\text{COOH} (aq)$	$\text{C}_2\text{H}_5\text{COO}^- (aq)$	1.3×10^{-5}
hydrogen citrate ion	$\text{HC}_6\text{H}_5\text{O}_7^{2-} (aq)$	$\text{C}_6\text{H}_5\text{O}_7^{3-} (aq)$	4.0×10^{-6}
carbonic acid	$\text{H}_2\text{CO}_3 (aq)$	$\text{HCO}_3^- (aq)$	4.3×10^{-7}
hydrosulfuric acid	$\text{H}_2\text{S} (aq)$	$\text{HS}^- (aq)$	1.0×10^{-7}
hydrogen sulfite ion	$\text{HSO}_3^- (aq)$	$\text{SO}_3^{2-} (aq)$	1.0×10^{-7}
dihydrogen arsenate ion	$\text{H}_2\text{AsO}_4^- (aq)$	$\text{HAsO}_4^{2-} (aq)$	8×10^{-8}
dihydrogen phosphate ion	$\text{H}_2\text{PO}_4^- (aq)$	$\text{HPO}_4^{2-} (aq)$	6.2×10^{-8}
hypochlorous acid	$\text{HOCl} (aq)$	$\text{OCl}^- (aq)$	3.5×10^{-8}
hypobromous acid	$\text{HOBr} (aq)$	$\text{OBr}^- (aq)$	2×10^{-9}
hydrocyanic acid	$\text{HCN} (aq)$	$\text{CN}^- (aq)$	6.2×10^{-10}
hydrogen arsenate ion	$\text{HAsO}_4^{2-} (aq)$	$\text{AsO}_4^{3-} (aq)$	6×10^{-10}
boric acid	$\text{H}_3\text{BO}_3 (aq)$	$\text{H}_2\text{BO}_3^- (aq)$	5.8×10^{-10}
ammonium ion	$\text{NH}_4^+ (aq)$	$\text{NH}_3 (aq)$	5.6×10^{-10}
Phenol	$\text{C}_6\text{H}_5\text{OH} (aq)$	$\text{C}_6\text{H}_5\text{O}^- (aq)$	1.6×10^{-10}
hydrogen carbonate ion	$\text{HCO}_3^- (aq)$	$\text{CO}_3^{2-} (aq)$	5.6×10^{-11}
hypoiodous acid	$\text{HOI} (aq)$	$\text{OI}^- (aq)$	2×10^{-11}
hydrogen ascorbate ion	$\text{HC}_6\text{H}_6\text{O}_6^- (aq)$	$\text{C}_6\text{H}_6\text{O}_6^{2-} (aq)$	1.6×10^{-12}
hydrogen phosphate ion	$\text{HPO}_4^{2-} (aq)$	$\text{PO}_4^{3-} (aq)$	4.8×10^{-13}
water (55.49 mol/L)	$\text{H}_2\text{O} (l)$	$\text{OH}^- (aq)$	$1.0 \times 10^{-14} = K_w$
hydrogen sulfide ion	$\text{HS}^- (aq)$	$\text{S}^{2-} (aq)$	$\sim 10^{-19}$

Increasing Acid Strength

Increasing Base Strength

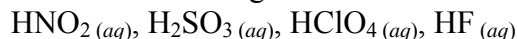


Check out Strong and Weak Acids Animations at
<http://www.sgc.peachnet.edu/users/larnold/WWW/courses/1212/rev1212.html>

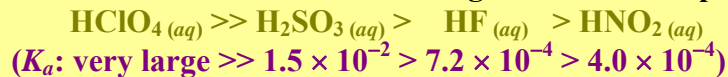
Relative Strength of Acids and Conjugate Bases:

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.



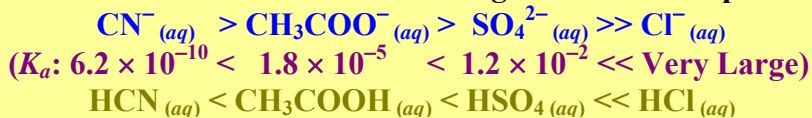
According to the Acid and Bases Relative Strength Table on the previous page:



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

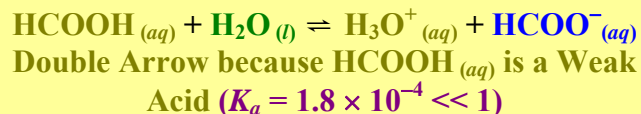
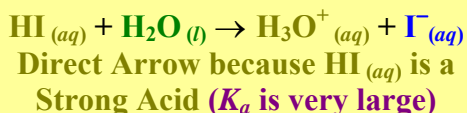


According to the Acids and Bases Relative Strength Table on the previous page:

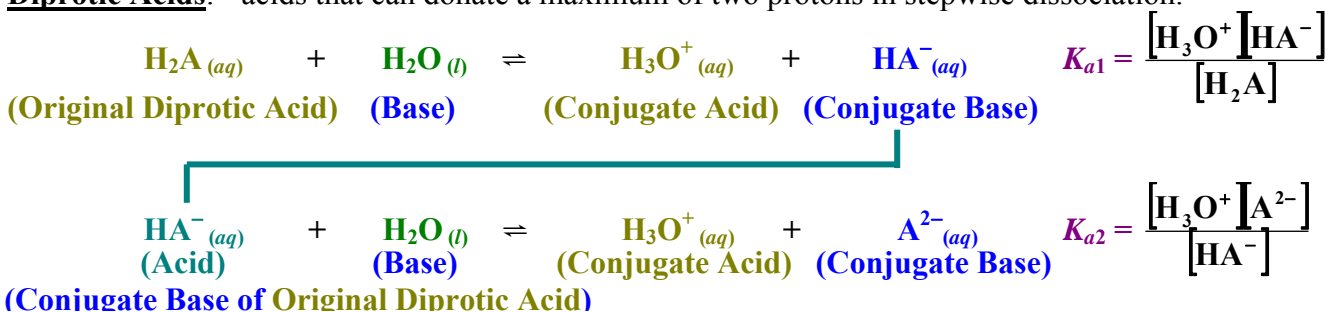


Monoprotic Acids: - acids that can donate a maximum of one proton.

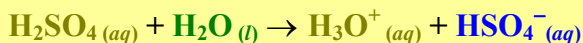
Example 3: Write the dissociation reaction for the following monoprotic acids.



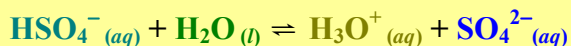
Diprotic Acids: - acids that can donate a maximum of two protons in stepwise dissociation.



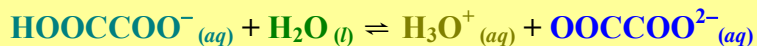
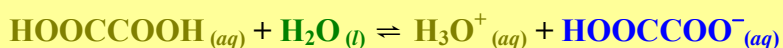
Example 4: Write the stepwise dissociation reaction for the following diprotic acids.



Direct Arrow because $\text{H}_2\text{SO}_4(aq)$ is a Strong Acid (K_{a1} is very large)



Double Arrow because $\text{HSO}_4^-(aq)$ is a Weak Acid ($K_{a2} = 1.2 \times 10^{-2} \ll 1$)



Double Arrow because both $\text{HOOC}\text{COOH}(aq)$ and HOOCCOO^- are Weak Acids (K_{a1} and $K_{a2} \ll 1$)

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: $\text{HOOC}\text{COO}^-(aq)$, $\text{HSO}_4^-(aq)$, $\text{HSO}_3^-(aq)$, $\text{HCO}_3^-(aq)$, $\text{HS}^-(aq)$, $\text{HC}_6\text{H}_6\text{O}_6^-(aq)$, and $\text{H}_2\text{O}(l)$

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 ($\text{H}_2\text{O}(l)$) 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.



Example 5: List the major species $\text{H}_2\text{SO}_4(aq)$, and calculate its pH if it has a concentration of 1.00×10^{-5} M.



Since H_2SO_4 is a strong acid, the major species are: $\text{H}_3\text{O}^+(aq)$, $\text{HSO}_4^-(aq)$ and $\text{H}_2\text{O}(l)$.

(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.)

$$[\text{H}_3\text{O}^+] = [\text{H}_2\text{SO}_4]_0 = 1.00 \times 10^{-5} \text{ M}$$

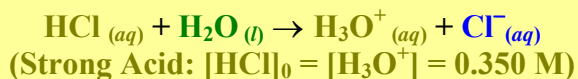
$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

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

$$\text{pH} = 5.00$$

Example 6: Determine the $[\text{H}_3\text{O}^+]$ and/or $[\text{OH}^-]$ concentrations of the following solutions at 25°C .

a. $[\text{HCl}] = 0.350 \text{ mol/L}$

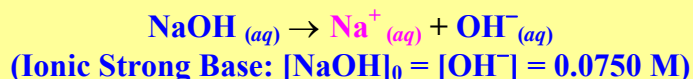


$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

$$[\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} = \frac{1.0 \times 10^{-14}}{0.350}$$

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

b. $[\text{NaOH}] = 0.0750 \text{ M}$



$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

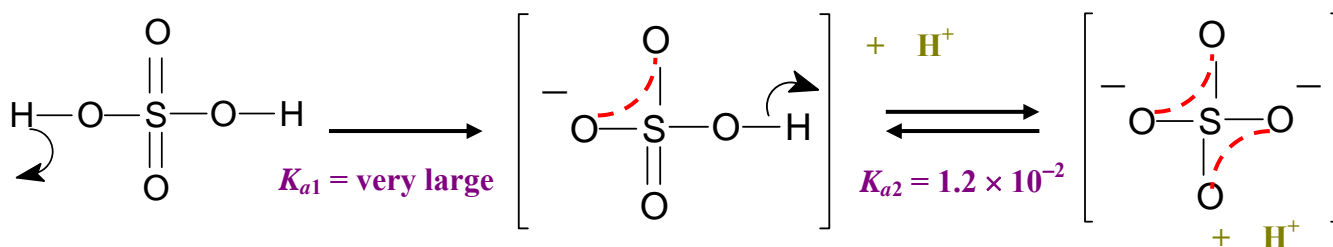
$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{0.0750}$$

$$[\text{H}_3\text{O}^+] = 1.33 \times 10^{-13} \text{ mol/L}$$

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.

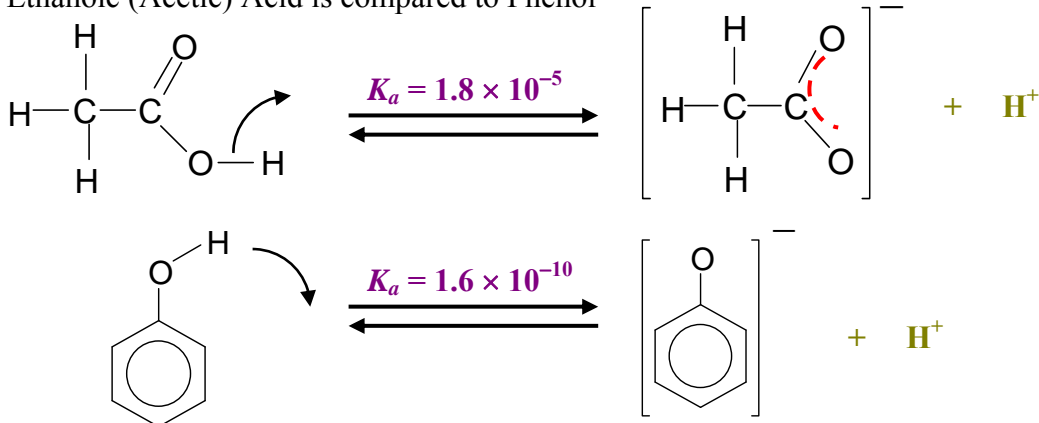
Example: Sulfuric Acid ($\text{H}_2\text{SO}_4_{(aq)}$), successively donate protons to finally reach $\text{SO}_4^{2-}_{(aq)}$



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), H₂O (l)

Strongest Acid: HF ($K_a = 7.2 \times 10^{-4}$) (K_a for H₂O is 1.0×10^{-14})

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

	HF (aq)	H ₂ O (l)	H ₃ O ⁺ (aq)	F ⁻ (aq)
Initial	0.200 M		0	0
Change	-x		+x	+x
Equilibrium	(0.2 - x)		x	x

CANNOT use Approximation:

$$\frac{[HF]_0}{K_a} = \frac{0.200 \text{ M}}{7.2 \times 10^{-4}} = 277.8 < 1000$$

Have to use (0.2 - x) in the denominator

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

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

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

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

(Quadratic Equation: Apply the Quadratic Formula!)

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

$$x = \frac{-(7.2 \times 10^{-4}) \pm \sqrt{(7.2 \times 10^{-4})^2 - 4(1)(-1.44 \times 10^{-4})}}{2(1)}$$

$$x = 0.0116453988$$

$$x = -0.0247307976 \text{ (omit negative } x)$$

$$[H_3O^+] = 0.0116 \text{ mol/L}$$

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

$$\text{pH} = -\log(0.0116)$$

$$\text{pH} = 1.93$$

Verify that we could NOT use Approximation:

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

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

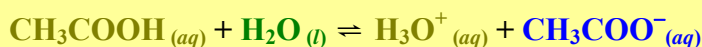
$$\% \text{ Dissociation} = \frac{[H_3O^+]}{[HA]_0} \times 100\%$$

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 CH₃COOH_(aq) ($K_a = 1.8 \times 10^{-5}$). Calculate the % dissociation of this acid mixture.

Major Species: HOBr_(aq), CH₃COOH_(aq), H₂O_(l)

Strongest Acid: CH₃COOH ($K_a = 1.8 \times 10^{-5}$) (K_a for HOBr and H₂O are 2×10^{-9} and 1.0×10^{-14})

CH₃COOH is a weak acid and undergoes Brønsted-Lowry Dissociation.



	[CH ₃ COOH]	[H ₃ O ⁺]	[CH ₃ COO ⁻]
Initial	0.200 M	0	0
Change	-x	+x	+x
Equilibrium	(0.2 - x)	x	x

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

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

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

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

$$x \approx 0.00190$$

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.

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]

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

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

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

$$\text{pH} = 2.72$$

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

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

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

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), H₂O_(l)

Strongest Acid: HA ($K_a = ?$) (K_a for H₂O is 1.0×10^{-14} and has a 0.000 01 % dissociation)

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



	[HA]	[H ₃ O ⁺]	[A ⁻]
Initial	0.0500 M	0	0
Change	-1.9 × 10 ⁻⁴	+1.9 × 10 ⁻⁴	1.9 × 10 ⁻⁴
Equilibrium	0.04981 M	1.9 × 10 ⁻⁴ M	1.9 × 10 ⁻⁴ M

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

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

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

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

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

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

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. ($K_{a1} > K_{a2} > K_{a3} > \dots$).
- 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' $[\text{H}_3\text{O}^+]$ are calculated from their K_{a1}** because successive proton donation from smaller subsequent K_a do not amount to any significant increase in the $[\text{H}_3\text{O}^+]$.

Examples: Some Polyprotic Acids:

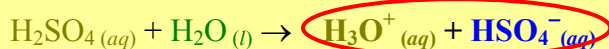
a. Diprotic Acids: $\text{H}_2\text{SO}_4(aq)$, $\text{HOCCOOH}(aq)$, $\text{H}_2\text{SO}_3(aq)$, $\text{H}_2\text{C}_6\text{H}_6\text{O}_6(aq)$, $\text{H}_2\text{CO}_3(aq)$, $\text{H}_2\text{S}(aq)$

Diprotic Amphoteric Intermediates: $\text{HSO}_4^-(aq)$, $\text{HOCCOO}^-(aq)$, $\text{HSO}_3^-(aq)$, $\text{HC}_6\text{H}_6\text{O}_6^-(aq)$,
 $\text{HCO}_3^-(aq)$, $\text{HS}^-(aq)$

b. Triprotic Acids: $\text{H}_3\text{PO}_4(aq)$, $\text{H}_3\text{AsO}_4(aq)$, $\text{H}_3\text{C}_6\text{H}_5\text{O}_7(aq)$, $\text{H}_3\text{BO}_3(aq)$

Triprotic Amphoteric Intermediates: $\text{H}_2\text{PO}_4^-(aq)$, $\text{HPO}_4^{2-}(aq)$, $\text{H}_2\text{AsO}_4^-(aq)$, $\text{HASO}_4^{2-}(aq)$,
 $\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-(aq)$, $\text{HC}_6\text{H}_5\text{O}_7^{2-}(aq)$, $\text{H}_2\text{BO}_3^-(aq)$, $\text{HBO}_3^{2-}(aq)$

Example 1: Determine the pH of 0.0500 M of sulfuric acid ($\text{H}_2\text{SO}_4(aq)$) and the concentrations of $\text{HSO}_4^-(aq)$, and $\text{SO}_4^{2-}(aq)$. The acid dissociation constant is $K_{a2} = 1.2 \times 10^{-2}$.



Since H_2SO_4 is a strong acid, the major species are: $\text{H}_3\text{O}^+(aq)$, $\text{HSO}_4^-(aq)$ and $\text{H}_2\text{O}(l)$.

$$[\text{H}_3\text{O}^+] = [\text{H}_2\text{SO}_4]_0 = 0.0500 \text{ M}$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

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

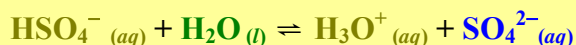
$$\text{pH} = 1.30 \text{ (after } K_{a1}\text{)}$$

Next, we have to calculate $[\text{HSO}_4^-]$, $[\text{SO}_4^{2-}]$ and the final pH after the last proton donation.

Major Species: $\text{HSO}_4^-(aq)$, $\text{H}_2\text{O}(l)$

Strongest Acid: $\text{HSO}_4^-(aq)$ ($K_{a2} = 1.2 \times 10^{-2}$)

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



	$[\text{HSO}_4^-]$	$[\text{H}_3\text{O}^+]$	$[\text{SO}_4^{2-}]$
Initial	0.0500 M	0.0500 M	0
Change	-x	+x	+x
Equilibrium	(0.05 - x)	(0.05 + x)	x

CANNOT use Approximation:

$$\frac{[\text{HSO}_4^-]_0}{K_{a2}} = \frac{0.0500 \text{ M}}{1.2 \times 10^{-2}} = 4.167 < 1000$$

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{SO}_4^{2-}]}{[\text{HSO}_4^-]}$$

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

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

$$x = 0.00851$$

```
solve((0.05+x)*x/
(0.05-x)-1.2E-2,
x, 0, (0, 0.05))
.0085094925
```

Final $[\text{H}_3\text{O}^+] = 0.0500 \text{ M}$ (from K_{a1}) + 0.00851 M (from K_{a2})

Final $[\text{H}_3\text{O}^+] = 0.0585 \text{ M}$
(a significant change from before)

$[\text{HSO}_4^-]_{eq} = 0.0500 \text{ M} - 0.00851 \text{ M}$

$[\text{HSO}_4^-]_{eq} = 0.0415 \text{ M}$
 $[\text{SO}_4^{2-}]_{eq} = 0.00585 \text{ M}$

Final pH = $-\log [\text{H}_3\text{O}^+] = -\log (0.0585)$

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 ($\text{H}_2\text{SO}_4(aq)$) requires the calculation of $[\text{H}_3\text{O}^+]$ in both steps of the dissociation. Other polyprotic acids, such as $\text{H}_3\text{C}_6\text{H}_5\text{O}_7(aq)$ only require the first dissociation step to calculate the $[\text{H}_3\text{O}^+]$.

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.

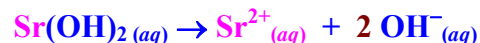
- all alkali bases (Group IA cations with OH^-) and some alkaline bases (Group IIA cations with OH^-) are considered as strong bases because they are ionic compound that dissociates completely.

Examples: Strong Ionic Bases:

- a. Alkali Bases: $\text{LiOH}(aq)$, $\text{NaOH}(aq)$, $\text{KOH}(aq)$, $\text{RbOH}(aq)$, and $\text{CsOH}(aq)$
(gives off 1 mole of OH^- when 1 mole of alkali base is dissolved)



- b. Alkaline Bases: $\text{Ca}(\text{OH})_2(aq)$, $\text{Ba}(\text{OH})_2(aq)$, $\text{Sr}(\text{OH})_2(aq)$
(gives off 2 moles of OH^- when 1 mole of alkaline base is dissolved)



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.



Example 1: List the major species $\text{NaOH}_{(aq)}$, and calculate its pH, pOH, $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ if it has a concentration of $1.00 \times 10^{-4} \text{ M}$.

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

Since NaOH is a strong base, the major species are: $\text{OH}^-_{(aq)}$ and $\text{H}_2\text{O}_{(l)}$.

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

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

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

$$\text{pOH} = 4.00$$

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

$$\text{pH} = 14 - (4.00)$$

$$\text{pH} = 10.00$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-10.00}$$

$$[\text{H}_3\text{O}^+] = 1.00 \times 10^{-10} \text{ M}$$

Example 2: List the major species $\text{Sr}(\text{OH})_2_{(aq)}$, and calculate its pH, pOH, $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ if it has a concentration of $1.00 \times 10^{-4} \text{ M}$.

$$\text{Sr}(\text{OH})_2_{(aq)} \rightarrow \text{Sr}^{2+}_{(aq)} + 2 \text{OH}^-_{(aq)}$$

Since Sr(OH)₂ is a strong base, the major species are: $\text{OH}^-_{(aq)}$ and $\text{H}_2\text{O}_{(l)}$.

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

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

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

$$\text{pOH} = 3.70$$

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{2.00 \times 10^{-4}}$$

$$[\text{H}_3\text{O}^+] = 5.00 \times 10^{-11} \text{ mol/L}$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

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

$$\text{pH} = 10.3$$

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

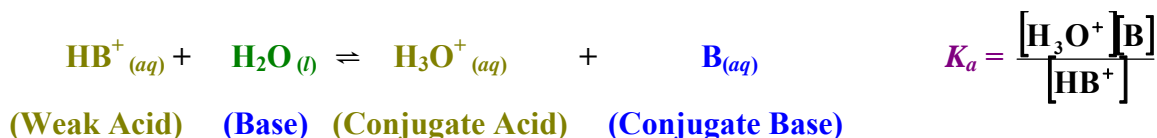
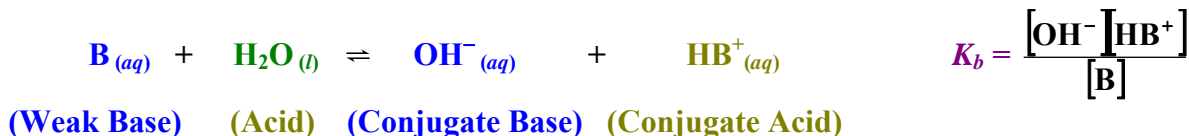
- when the base dissociation constant (K_b) is less than 1, ($K_b < 1$)
- the equilibrium position strongly favours the reactants.
- at equilibrium, the hydroxide concentration is much less than the original base concentration, $[\text{OH}]_{eq} < [\text{B}]_{eq}$ or $[\text{B}]_0 \approx [\text{B}]_{eq}$.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



	[B]		$[\text{OH}^-]$	$[\text{HB}^+]$
Initial	x		0	0
Change	$-y$ (where $y \ll x$)		$+y$	$+y$
Equilibrium	$(x - y) \approx x$		y	y

Examples: Some Weak Bases: $\text{HOCCOO}^- (aq)$, $\text{HSO}_3^- (aq)$, $\text{HSO}_4^- (aq)$, $\text{H}_2\text{PO}_4^- (aq)$, $\text{NO}_2^- (aq)$, $\text{HCO}_3^- (aq)$, $\text{H}_2\text{C}_6\text{H}_5\text{O}_7^- (aq)$, $\text{F}^- (aq)$, $\text{HCOO}^- (aq)$, $\text{C}_6\text{H}_7\text{O}_6^- (aq)$, $\text{C}_6\text{H}_5\text{COO}^- (aq)$, $\text{CO}_3^{2-} (aq)$, $\text{CH}_3\text{COO}^- (aq)$, $\text{HS}^- (aq)$, $\text{OCl}^- (aq)$, $\text{CN}^- (aq)$, $\text{NH}_3 (aq)$, and $\text{NO}_3^- (aq)$

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 .



$$K_a \times K_b = \frac{[\text{H}_3\text{O}^+][\text{B}]}{[\text{HB}^+]} \times \frac{[\text{OH}^-][\text{HB}^+]}{[\text{B}]} = [\text{H}_3\text{O}^+][\text{OH}^-] = K_w$$

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.

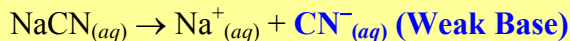
Procedure to calculate pH of Weak Base Solutions:

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

$$\left(\frac{[\text{OH}^-]}{[\text{B}]_0} \times 100\% \leq 5\% \right)$$

Example 3: Determine the $[\text{OH}^-]$, $[\text{H}_3\text{O}^+]$, pOH, pH and % dissociation of 0.200 mol/L of $\text{NaCN}_{(aq)}$. (K_a of $\text{HCN} = 6.2 \times 10^{-10}$).

NaCN dissociates completely in water:



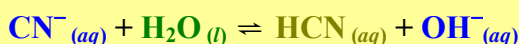
Major Species: $\text{CN}^-_{(aq)}$, $\text{H}_2\text{O}_{(l)}$

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

(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.



	$[\text{CN}^-]$	$[\text{H}_3\text{O}^+]$	$[\text{OH}^-]$
Initial	0.200 M	0	0
Change	-x	+x	+x
Equilibrium	(0.2 - x)	x	x

$$K_b = \frac{[\text{HCN}][\text{OH}^-]}{[\text{CN}^-]} \quad 1.613 \times 10^{-5} = \frac{(x)(x)}{(0.2-x)} \approx \frac{x^2}{(0.2)}$$

$$1.613 \times 10^{-5} (0.2) \approx x^2$$

$$3.226 \times 10^{-6} \approx x^2$$

$$x \approx \sqrt{3.226 \times 10^{-6}}$$

$$x \approx 0.0018$$

CAN use Approximation:

$$\frac{[\text{CN}^-]_0}{K_b} = \frac{0.200 \text{ M}}{1.613 \times 10^{-5}}$$

$$= 12399 \geq 1000$$

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

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

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

$$\text{pOH} = -\log(0.0018)$$

$$\text{pOH} = 2.75$$

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

$$\% \text{ Dissociation} = \frac{0.0018 \text{ M}}{0.200 \text{ M}} \times 100\%$$

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

Verify that we could use Approximation:

$$\frac{[\text{OH}^-]}{[\text{CN}^-]_0} \times 100\% = \frac{0.0018 \text{ M}}{0.200 \text{ M}} \times 100\%$$

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

Therefore, approximation would be appropriate.

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

$$\text{pH} = 14 - (2.75)$$

$$\text{pH} = 11.25$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-11.25}$$

$$[\text{H}_3\text{O}^+] = 5.6 \times 10^{-12} \text{ M}$$

Assignment

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:

- $\text{KNO}_3(aq)$ (K^+ can be from a strong base – $\text{KOH}(aq)$; NO_3^- is the conjugate-base of a strong acid – $\text{HNO}_3(aq)$)
- $\text{NaCl}(aq)$ (Na^+ can be from a strong base – $\text{NaOH}(aq)$; Cl^- is the conjugate-base a strong acid – $\text{HCl}(aq)$)

2. **Basic Salts:** - when the **Cation comes from a Strong Base** and the **Anion is the Conjugate-Base of a Weak Acid**.

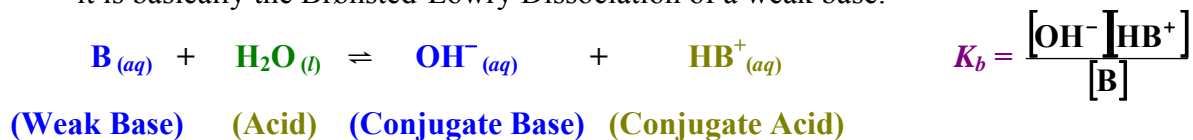
- pH will increase; if dissolve in pure water, pH > 7 (Basic).

Examples: Some Basic Salts:

- $\text{NaCH}_3\text{COO}(aq)$ (Na^+ can be from a strong base – $\text{NaOH}(aq)$; CH_3COO^- is the conjugate-base of a weak acid – $\text{CH}_3\text{COOH}(aq)$)
- $\text{KF}(aq)$ (K^+ can be from a strong base – $\text{KOH}(aq)$; F^- is the conjugate base of a weak acid – $\text{HF}(aq)$)

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.



Percent Hydrolysis: - the amount of $[\text{OH}^-]$ dissociated from the original $[\text{B}]_0$ expressed in percentage.

- strong bases will have % Hydrolysis = 100%.

- weak bases will have % Hydrolysis < 100%

<p>Percent Hydrolysis of Bases</p> $\% \text{ Hydrolysis} = \frac{[\text{OH}^-]}{[\text{B}]_0} \times 100\%$

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: $\text{NH}_4\text{ClO}_4(aq)$ is an acidic Salts: (NH_4^+ is the conjugate-acid of a weak base – $\text{NH}_3(aq)$; ClO_4^- is the conjugate-base of a strong acid – $\text{HClO}_4(aq)$)

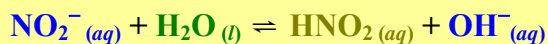
Example 1: Determine the pH and the percent hydrolysis of 0.235 M of sodium nitrite. The acid dissociation constant for nitrous acid is $K_a = 4.0 \times 10^{-4}$.

Sodium nitrite (NaNO_2) dissociates completely in H_2O : $\text{NaNO}_2(s) \rightarrow \text{Na}^+(aq) + \text{NO}_2^-(aq)$ (Weak Base)

Major Species: $\text{NO}_2^-(aq)$, $\text{H}_2\text{O}(l)$ $K_b = \frac{K_w}{K_a} = \frac{1.00 \times 10^{-14}}{4.0 \times 10^{-4}} \quad K_b = 2.5 \times 10^{-11}$

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.



	$[\text{NO}_2^-]$	$[\text{HNO}_2]$	$[\text{OH}^-]$
Initial	0.235 M	0	0
Change	-x	+x	+x
Equilibrium	(0.235 - x)	x	x

CAN use Approximation:

$$\frac{[\text{NO}_2^-]_0}{K_b} = \frac{0.235 \text{ M}}{2.5 \times 10^{-11}} = 9.4 \times 10^9 \geq 1000$$

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

$$K_b = \frac{[\text{HNO}_2][\text{OH}^-]}{[\text{NO}_2^-]} \quad 2.5 \times 10^{-11} = \frac{(x)(x)}{(0.235 - x)} \approx \frac{x^2}{(0.235)}$$

$$2.5 \times 10^{-11} (0.235) \approx x^2$$

$$5.875 \times 10^{-12} \approx x^2$$

$$x \approx \sqrt{5.875 \times 10^{-12}}$$

$$x = [\text{OH}^-] \approx 2.4 \times 10^{-6} \text{ M}$$

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

$$\text{pOH} = -\log(2.4 \times 10^{-6})$$

$$\text{pOH} = 5.62$$

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

$$\text{pH} = 14 - (5.62)$$

$$\text{pH} = 8.38$$

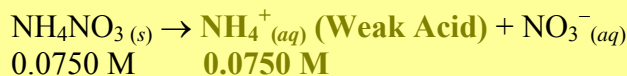
$$\% \text{ Hydrolysis} = \frac{[\text{OH}^-]}{[\text{NO}_2^-]_0} \times 100\% = \frac{2.4 \times 10^{-6} \text{ M}}{0.235 \text{ M}} \times 100\%$$

$$\% \text{ Hydrolysis} = 0.00102\%$$

As the small K_b suggests, the $[\text{OH}^-]$ is very small compared to $[\text{NO}_2^-]_0$. Therefore, % hydrolysis is also very small as a result.

Example 2: Determine the pH of 0.0750 M of ammonium nitrate. The base dissociation constant for $\text{NH}_3(aq)$ is $K_b = 1.8 \times 10^{-5}$.

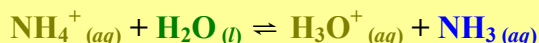
Ammonium nitrate, NH_4NO_3 , dissociates completely in water:



Major Species: $\text{NH}_4^+(aq)$, $\text{H}_2\text{O}(l)$ $K_a = \frac{K_w}{K_b} = \frac{1.00 \times 10^{-14}}{1.8 \times 10^{-5}} \quad K_a = 5.556 \times 10^{-10}$

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.



	$[\text{NH}_4^+]$	$[\text{H}_3\text{O}^+]$	$[\text{NH}_3]$
Initial	0.0750 M	0	0
Change	-x	+x	+x
Equilibrium	(0.075 - x)	x	x

CAN use Approximation:

$$\frac{[\text{NH}_4^+]_0}{K_a} = \frac{0.0750 \text{ M}}{5.556 \times 10^{-10}} = 1.35 \times 10^8 \geq 1000$$

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

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{NH}_3]}{[\text{NH}_4^+]}$$

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

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

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

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

$$x = [\text{H}_3\text{O}^+] \approx 6.455 \times 10^{-6} \text{ M}$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

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

$$\text{pH} = 5.19$$

Verify for Using Approximation:

$$\frac{[\text{H}_3\text{O}^+]}{[\text{NH}_4^+]_0} \times 100\% = \frac{6.455 \times 10^{-6} \text{ M}}{0.0750 \text{ M}} \times 100\%$$

$$= 0.0086\% \leq 5\% \quad \text{Therefore, approximation would be appropriate.}$$

Assignment

16.9 pg. 570–571 #65, 68, 70 to 74