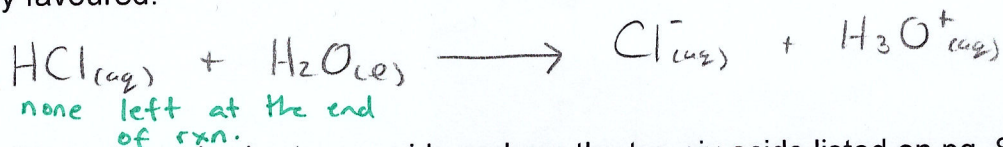


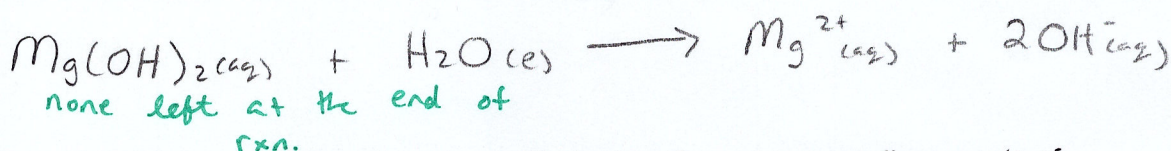
Acid-Base Equilibrium Constants & Expressions

- **Strong acids** are defined as proton donors that will completely ionize in water. This means that all reactants get consumed to produce lots of products. Therefore a strong acid cannot form equilibrium because the products are so heavily favoured.



- There are only six strong acids and are the top six acids listed on pg. 8-9 in data booklet (very large K values)
- **Strong bases** are defined as proton acceptors that will completely dissociate in water. This means that all reactants get consumed to produce products. Therefore a strong base cannot form equilibrium because the products are so heavily favoured.

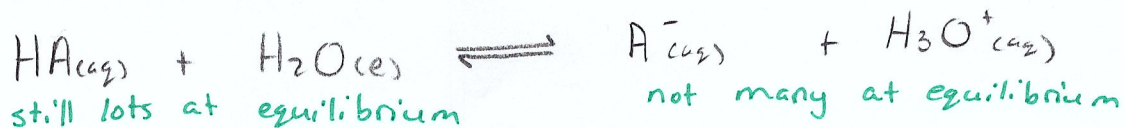
- A strong base is an ionic compound that contains the hydroxide ion (OH^-)



- **Weak acids** in water will establish equilibrium and produce small amounts of $\text{H}^+/\text{H}_3\text{O}^+$. This means not much of the weak acid actually ionizes (ie. not many products form).
- **Weak bases** in water will establish equilibrium and produce small amounts of OH^- ions. This means not much of the weak base actually dissociates (ie. not many products form).
- If very few products are produced with weak acids and weak bases, the equilibrium constant must be small for all weak acids and weak bases.
 - The equilibrium constant for weak acids and weak bases is used to determine exactly how many $\text{H}^+/\text{H}_3\text{O}^+$ ions will be produced when a weak acid breaks down or how many OH^- ions will be produced when a weak base breaks down

ACID IONIZATION EXPRESSIONS

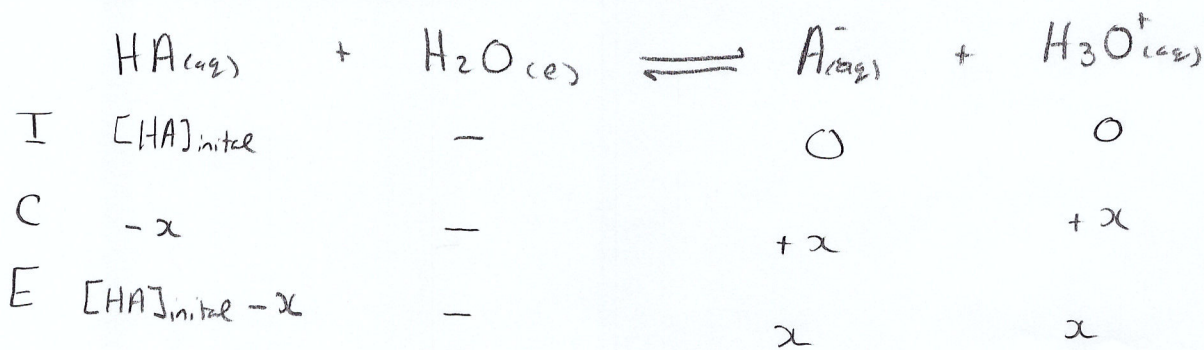
- The general reaction written below represents the equilibrium for a weak acid



- The equilibrium expression can be written for any weak acid equilibrium based on the general reaction above and is known as the acid ionization expression or K_a expression

$$K_a = \frac{[\text{A}^{-}][\text{H}_3\text{O}^{+}]}{[\text{HA}]}$$

- ICE tables can be applied to weak acids in equilibrium, but the general outcome is the same for all weak acids



$$\therefore K_a = \frac{[\text{A}^{-}][\text{H}_3\text{O}^{+}]}{[\text{HA}]_{\text{equil.}}} = \frac{(x)(x)}{([\text{HA}]_{\text{initial}} - x)}$$

but for weak acids, K_a is small, so use approximation rule $[\text{HA}]_{\text{initial}} - x \approx [\text{HA}]_{\text{initial}}$

$$K_a = \frac{(x)(x)}{[\text{HA}]_{\text{initial}}} = \frac{x^2}{[\text{HA}]_{\text{initial}}} \quad \text{where } x = [\text{H}_3\text{O}^{+}] / [\text{H}^{+}] \text{ or } [\text{A}^{-}]$$

$$\therefore K_a = \frac{[\text{H}_3\text{O}^{+}]^2}{[\text{HA}]_{\text{initial}}}$$

- Recall other acid-base equations that will also be used:

$$[\text{H}_3\text{O}^{+}_{(aq)}] = 10^{-\text{pH}} \quad \text{and} \quad \text{pH} = -\log[\text{H}_3\text{O}^{+}_{(aq)}] \quad \text{*memorize!}$$

EXAMPLES:

1. Methanoic acid, $\text{CHOOH}_{(\text{aq})}$, is present in the sting of certain ants. What is the pH of a 0.25 mol/L solution of methanoic acid?

$$\text{pH} = ?$$

$$[\text{HA}] = 0.25 \text{ mol/L}$$

$$K_a = 1.8 \times 10^{-4}$$

↑
from data book

$$\textcircled{2} \quad \text{pH} = -\log [\text{H}_3\text{O}^+]$$

↓

$$\textcircled{1} \quad K_a = \frac{[\text{H}_3\text{O}^+]^2}{[\text{HA}]}$$

approximation
rule good!

$$\textcircled{1} \quad K_a = \frac{[\text{H}_3\text{O}^+]^2}{[\text{HA}]} \Rightarrow [\text{H}_3\text{O}^+] = \sqrt{K_a [\text{HA}]}$$

$$[\text{H}_3\text{O}^+] = \sqrt{(1.8 \times 10^{-4})(0.25)} = 6.708 \dots \times 10^{-3} \text{ mol/L}$$

$$\textcircled{2} \quad \text{pH} = -\log [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log(6.708 \dots \times 10^{-3}) = 2.17339 \dots$$

$$\boxed{\text{pH} = 2.17}$$

2. Propanoic acid, $\text{CH}_3\text{CH}_2\text{COOH}_{(\text{aq})}$, is a weak monoprotic acid that is used to inhibit mould formation in bread. A student prepared a 0.10 mol/L solution of propanoic acid and found the pH was 2.96. What is the acid ionization constant for propanoic acid?

$$[\text{HA}] = 0.10 \text{ mol/L}$$

$$\text{pH} = 2.96$$

$$K_a = ?$$

(assume approximation rule is good)

$$\textcircled{2} \quad K_a = \frac{[\text{H}_3\text{O}^+]^2}{[\text{HA}]}$$

$$\textcircled{1} \quad [\text{H}_3\text{O}^+] = 10^{-\text{pH}}$$

$$\textcircled{1} \quad [\text{H}_3\text{O}^+] = 10^{-\text{pH}}$$

$$[\text{H}_3\text{O}^+] = 10^{-2.96} = 0.00109... \text{ mol/L}$$

$$\textcircled{2} \quad K_a = \frac{[\text{H}_3\text{O}^+]^2}{[\text{HA}]} = \frac{(0.00109... \text{ mol/L})^2}{0.10}$$

$$K_a = 1.20226... \times 10^{-5}$$

$$K_a = 1.2 \times 10^{-5}$$

Now try the Practice Problems

Practice Problems

- Write the equilibrium expression for each acid ionizing in an aqueous solution.
 - nitrous acid, $\text{HNO}_{2(\text{aq})}$
 - benzoic acid, $\text{C}_6\text{H}_5\text{COOH}_{(\text{aq})}$
 - carbonic acid, $\text{H}_2\text{CO}_{3(\text{aq})}$
- Phenol, $\text{C}_6\text{H}_6\text{O}_{(\text{aq})}$, is a weak monoprotic acid used as a disinfectant. The K_a expression is $K_a = \frac{[\text{C}_6\text{H}_5\text{O}^-][\text{H}_3\text{O}^+]}{[\text{C}_6\text{H}_6\text{O}]}$. Write the ionization reaction for phenol in an aqueous solution.
- Calculate the pH of a sample of vinegar that contains 0.83mol/L ethanoic acid.
- Hypochlorous acid, $\text{HOCl}_{(\text{aq})}$ is used to make bleach. A chemist finds that a 0.40mol/L solution of $\text{HOCl}_{(\text{aq})}$ will ionize and produce 1.08×10^{-4} mol/L of $\text{OCl}^-_{(\text{aq})}$. What is the K_a for the acid?
- Hexanoic acid, commonly known as caproic acid, $\text{C}_5\text{H}_{11}\text{COOH}_{(\text{s})}$, occurs naturally in coconut and palm oil. It is a weak monoprotic acid, with $K_a = 1.3 \times 10^{-5}$. A certain aqueous solution of hexanoic acid has a pH of 2.94. What was the initial concentration of the hexanoic acid solution?
- Lactic acid is a monoprotic acid produced by muscle activity. It is also produced in milk by the action of bacteria. What is the pH of a 0.16mol/L solution of lactic acid?
- What is the acid ionization constant for gallic acid if a 0.68mol/L solution has a pH of 2.29?

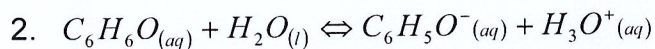
Practice Problem: Answers

1.

a. $K_a = \frac{[NO_2^-][H_3O^+]}{[HNO_2]}$

b. $K_a = \frac{[C_6H_5COO^-][H_3O^+]}{[C_6H_5COOH]}$

c. $K_{a1} = \frac{[HCO_3^-][H_3O^+]}{[H_2CO_3]}$ and $K_{a2} = \frac{[CO_3^{2-}][H_3O^+]}{[HCO_3^-]}$



3. 2.41

4. 2.9×10^{-8}

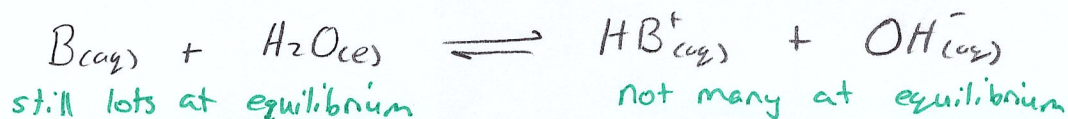
5. 0.10 mol/L

6. 2.32

7. 3.87×10^{-5}

BASE EQUILIBRIUM EXPRESSIONS

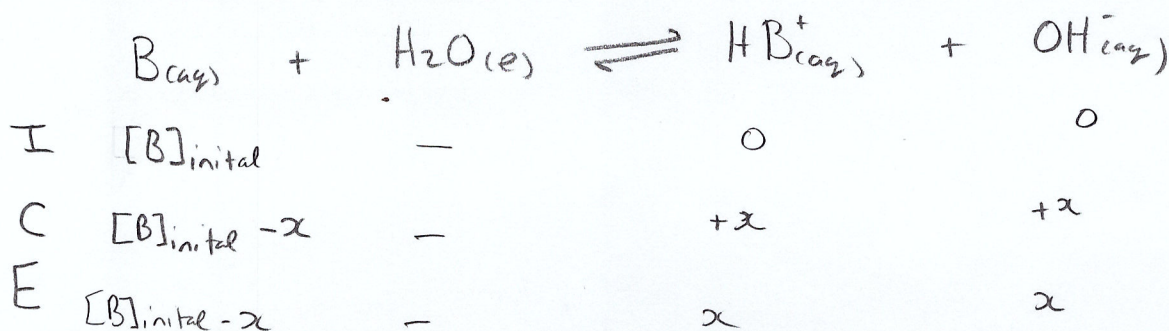
- The general reaction written below represents the equilibrium for a weak base



- The equilibrium expression can be written for any weak base equilibrium based on the general reaction above and is known as the base equilibrium expression or K_b expression

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

- ICE tables can be applied to weak bases in equilibrium, but the general outcome is the same for all weak bases



$$K_b = \frac{[\text{HB}^+][\text{OH}^-]}{[\text{B}]_{\text{equil.}}} = \frac{(x)(x)}{[\text{B}]_{\text{initial}} - x}$$

but for weak bases, K_b will be small, \therefore use approximation rule $[\text{B}]_{\text{initial}} - x \simeq [\text{B}]_{\text{initial}}$

$$K_b = \frac{(x)(x)}{[\text{B}]_{\text{initial}}} \quad \text{where } x = [\text{OH}^-] = [\text{HB}^+]$$

$$K_b = \frac{[\text{OH}^-]^2}{[\text{B}]_{\text{initial}}}$$

- Calculations are needed to relate acidic information to basic information (because as acidic properties increase, basic properties decrease and vice versa)
 - The relationship between pH and pOH is

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

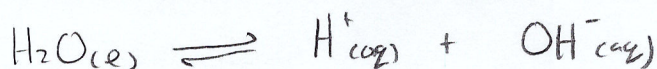
and recall $\text{pOH} = -\log[\text{OH}^-_{(\text{aq})}]$

&

$$[\text{OH}^-_{(\text{aq})}] = 10^{-\text{pOH}}$$

} memorize!

- Water can ionize and create an equilibrium reaction shown below



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

where

$$K_w = 1.0 \times 10^{-14} \text{ at } 25^\circ\text{C}$$

↑ pg. 3 of data book

memorize!

- There is a relationship between the K_a and K_b for a conjugate acid-base pair

$$K_w = K_a \times K_b$$

memorize!

EXAMPLES:

- The characteristic bitter taste of tonic water is due to the addition of quinine, $\text{C}_{20}\text{H}_{24}\text{N}_2\text{O}_2(\text{s})$, a naturally occurring, white crystalline compound. It is also used to treat malaria. The base equilibrium constant, K_b , for quinine is 3.3×10^{-6} . What is the hydroxide ion concentration of a $3.6 \times 10^{-2} \text{ mol/L}$ solution of quinine?

$$K_b = 3.3 \times 10^{-6}$$

$$[\text{B}]_{\text{initial}} = 3.6 \times 10^{-2} \text{ mol/L}$$

$$[\text{OH}^-] = ?$$

$$K_b = \frac{[\text{OH}^-]^2}{[\text{B}]_{\text{initial}}}$$

approximation
rule good!

$$[\text{OH}^-] = \sqrt{K_b [\text{B}]_{\text{initial}}}$$

$$[\text{OH}^-] = \sqrt{(3.3 \times 10^{-6})(3.6 \times 10^{-2})}$$

$$[\text{OH}^-] = 1.1 \times 10^{-4} \text{ mol/L}$$

2. An aqueous solution of ammonia has a pOH of 3.15. What is the concentration of the ammonia solution? ^{NH₃} basic!

$$\text{pOH} = 3.15$$

$$[\text{B}]_{\text{initial}} = ?$$

$$K_a = 5.6 \times 10^{-10} \text{ for } \text{NH}_4^+ \Rightarrow \therefore K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{5.6 \times 10^{-10}} = 1.7857 \dots \times 10^{-5}$$

↑
from data book

$$\textcircled{2} K_b = \frac{[\text{OH}^-]}{[\text{B}]_{\text{initial}}}$$

$$\textcircled{1} [\text{OH}^-] = 10^{-\text{pOH}}$$

$$\textcircled{1} [\text{OH}^-] = 10^{-\text{pOH}} = 10^{-(3.15)} = 7.07945 \dots \times 10^{-4} \text{ mol/L}$$

$$\textcircled{2} K_b = \frac{[\text{OH}^-]^2}{[\text{B}]_{\text{initial}}} \Rightarrow [\text{B}]_{\text{initial}} = \frac{[\text{OH}^-]^2}{K_b}$$

$$[\text{B}]_{\text{initial}} = \frac{(7.07945 \dots \times 10^{-4})^2}{1.7857 \dots \times 10^{-5}} = 0.02806 \dots$$

$$[\text{B}]_{\text{initial}} = 0.028 \text{ mol/L}$$

3. One of the uses for aniline, $C_6H_5NH_2(l)$, is in the manufacture of dyes. Aniline is soluble in water and acts as a weak base. When a 0.054 mol/L solution of aniline was prepared, the pH was found to be 8.68 . Calculate the base equilibrium constant for aniline.
- ↳ basic

$$[B]_{\text{initial}} = 0.054 \text{ mol/L}$$

$$\text{pH} = 8.68$$

$$K_b = ?$$

assume approximation rule is good!

$$\textcircled{3} \quad K_b = \frac{[\text{OH}^-]^2}{[B]_{\text{initial}}}$$

$$\textcircled{2} \quad [\text{OH}^-] = 10^{-\text{pOH}} \quad \text{or}$$

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

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

$$\textcircled{1} \quad 14 = \text{pH} + \text{pOH}$$

* going to go with green path *

$$\textcircled{1} \quad \text{pOH} + \text{pH} = 14 \Rightarrow \text{pOH} = 14 - \text{pH}$$

$$\text{pOH} = 14 - 8.68 = 5.32$$

$$\textcircled{2} \quad [\text{OH}^-] = 10^{-\text{pOH}} = 10^{-5.32} = 4.7863 \dots \times 10^{-6} \text{ mol/L}$$

$$\textcircled{3} \quad K_b = \frac{[\text{OH}^-]^2}{[B]_{\text{initial}}} = \frac{(4.7863 \dots \times 10^{-6})^2}{(0.054)} = 4.2423 \dots \times 10^{-10}$$

$$K_b = 4.2 \times 10^{-10}$$