Chapter 20: Acids and Bases
Part 1 – Notes: Introduction to Acids and Bases

Electrolytes

An **electrolyte** is . . .

Question: How does a strong electrolyte differ from a weak electrolyte?

Properties of Acids and Bases

Acid Ionization:

***Acids** are **molecular compounds** that **WILL** form ions when dissolved in water.* An acid will donate a proton (a hydrogen ion) to the water forming the *hydronium ion* and a negative ion.

Question: Ions are found in water solution as a result of two different processes, ionization and dissociation. Describe the differences between ionization and dissociation.

Arrhenius Definitions of Acids and Bases

Arrhenius Acid:

Ex.

Arrhenius Base:

Ex.

Bronsted-Lowry Definitions of Acids and Bases

Bronsted Acid:

Ex.

Bronsted Base:

Ex.

Lewis Definitions of Acids and Bases

Lewis Acid:

Ex.

Lewis Base:

Ex.

Salt:

Oxyacids: acids that contain hydrogen, oxygen, and at least one additional element

Ex.

Binary Acid: acids that contain only two elements

Ex.

Monoprotic Acid: any acid that contains *ONE* ionizable hydrogen/proton

Ex.

Diprotic Acid: any acid that contains *TWO* ionizable hydrogens/protons

Ex.

Triprotic Acid: any acid that contains *THREE* ionizable hydrogens/protons

Ex.

Amphoteric:

Example:

The pH Scale

The pH scale shows the relationship between pH and the hydrogen-ion concentration. Notice that acids have low pHs and bases have high pHs. A pH of 7 means a substance is neither acidic or basic – it is neutral.

pH	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
$[H^+]$	10^0	10^{-1}	10^{-2}	10^{-3}	10^{-4}	10^{-5}	10^{-6}	10^{-7}	10^{-8}	10^{-9}	10^{-10}	10^{-11}	10^{-12}	10^{-13}	10^{-14}

<<<<<<<Stronger acid Stronger base >>>>>>>

	$[H_3O^+]$	***$[OH^-]$***	***pH***	***aqueous system***
Acidic	1×10^0	1×10^{-14}	0	1*M* HCl
Acidic	1×10^{-1}	1×10^{-13}	1	0.10*M* HCl (1.0)
Acidic	1×10^{-2}	1×10^{-12}	2	gastric juice (1.6-1.8)
Acidic	1×10^{-3}	1×10^{-11}	3	lemon juice (2.3), vinegar (2.3-2.4)
Acidic	1×10^{-4}	1×10^{-10}	4	soda water (3.8), tomato juice (4.2)
Acidic	1×10^{-5}	1×10^{-9}	5	black coffee (5.0)
Acidic	1×10^{-6}	1×10^{-8}	6	milk (6.3-6.6), urine (5.5-5.7)
Neutral	1×10^{-7}	1×10^{-7}	7	pure water (7.0), saliva (6.2-7.4)
Basic	1×10^{-8}	1×10^{-6}	8	blood (7.35-7.45), bile (7.8-8.6)
Basic	1×10^{-9}	1×10^{-5}	9	sodium bicarbonate (8.4), sea water (8.4)
Basic	1×10^{-10}	1×10^{-4}	10	milk of magnesia (10.5)
Basic	1×10^{-11}	1×10^{-3}	11	household ammonia (11.5)
Basic	1×10^{-12}	1×10^{-2}	12	washing soda (12.0)
Basic	1×10^{-13}	1×10^{-1}	13	0.10*M* NaOH (13.0)
Basic	1×10^{-14}	1×10^0	14	1*M* NaOH

Chapter 20: Acids and Bases
Part 1 – Assignment: Introduction to Acids and Bases

Determine whether the following acids are monoprotic, diprotic, or triprotic. (M, D, or T???)

1.	HCl	4.	H_2SO_4	7.	H_3PO_4
2.	$HClO_4$	5.	$HClO_3$	8.	HNO_3
3.	H_3X	6.	H_2X	9.	HX

Are the following substances amphiprotic (amphoteric)? Write YES or NO.

10.	H_2O	14.	$H_2PO_4^{-1}$	18.	HCO_3^{-1}
11.	HCl	15.	HPO_4^{-2}	19.	H_2SO_4
12.	H I	16.	H_3PO_4	20.	HSO_4^{-1}
13.	NaOH	17.	H_2CO_3	21.	SO_4^{-2}

22-24 Write an equation for the ***ionization*** of each of the following acids in water solution.

22. HCl

23. H_2CO_3

24. H_3PO_4

25-27 Write an equation for the ***dissociation*** of each of the following in water solution.

25. NaOH

26. $Ca(OH)_2$

27. $Sc(OH)_3$

28. A Bronsted acid is a substance that donates a proton to another substance. Is a proton removed from the nucleus of one of the atoms in an acid? If yes, how does this happen since it is not a nuclear reaction? If no, how does the substance donate a proton?

29. What is the range for the pH scale?

Chapter 20: Acids and Bases
Part 2 – Notes: Conjugate Acid-Base Pairs and the Water Constant

Proton-Transfer Reactions and Conjugate Acid-Base Pairs
The Bronsted definition of acids and bases provides a basis for the study of **PROTOLYSIS**, or *proton-transfer reactions*.

Recall: An Bronsted acid is a species capable of **donating** a proton. A Bronsted base is a species that can **accept** a proton

Let's suppose an acid gives up a proton. What is true of the piece that remains after the proton is transferred?

Conjugate base:

What is the conjugate base of HCl?

Conjugate acid:

What is the conjugate acid of NH_3?

You need to be able to identify conjugate acid-base pairs.

Example 1: $HClO_4 + H_2O \rightarrow H_3O^+ + ClO_4^-$

Example 2: $NH_3 + H_2O \rightarrow OH^- + NH_4^+$

Key point:

In a Bronsted acid-base reaction, the **stronger acid** has the **weaker conjugate base**. In Example 1, ClO_4^- ions are the **conjugate base** of $HClO_4$. Since the base, H_2O, was able to win the competition for the protons, it is stronger than ClO_4^- ions. Thus, H_3O^+, the **conjugate acid** of H_2O, is weaker than $HClO_4$, the **conjugate acid** of ClO_4 ions. So the stronger a Bronsted acid, the weaker its **conjugate base**.

Also, Example 2 shows that **the conjugate base of a very weak acid is *always* a strong base**. Water is a very weak acid; hence OH^- ions, the **conjugate base** of H_2O, is a very strong base and it will remove protons from any acid stronger than water.

Auto-Dissociation of Water

Pure water dissociates (or ionizes) very slightly into ions. Perhaps collisions between water molecules may result in a transfer of proton from one molecule to another. The equation for the dissociation (or ionization) of water may be written as:

It has been found by experimentation that 1 L of pure water contains only one ten-millionth of a mole of hydronium ions (1×10^{-7} mol) and one ten-millionth of a mole of hydroxide ions (1×10^{-7} mol). If a substance contains more hydronium ions that hydroxide, it is acidic. If a substance contains more hydroxide ions than hydronium, it is basic. *Water is considered a neutral substance because it contains equal numbers of hydronium and hydroxide ions.*

The ions are in ***equilibrium*** in pure water; the concentration of the hydronium ions is $[H_3O^+] = 1.00 \times 10^{-7}$ M and the concentration of the hydroxide ion is $[OH^-] = 1.00 \times 10^{-7}$ M.

Since we know the equation, we can write the equilibrium expression for the dissociation of water:

The concentration, in molarity of water is:

Since the concentration of the ions (1.00×10^{-7} M) is so small in comparison with the total concentration of water (55.5 M), ***the concentration of water can be assumed to remain constant***. If the concentration of water is constant, we can multiply both sides of the equation by this constant without destroying the relationship.

Since both $[H_2O]$ and K_{eq} are constants, their product will also be a constant. We call this constant the ***dissociation constant of water***, and it is given the symbol $\boldsymbol{K_w}$.

Chapter 20: Acids and Bases
Part 2 – Assignment: Conjugate Acid-Base Pairs and the Water Constant

Show the conjugate acid base pairs for the following reactions. Write A and B on the reactant side ***above*** the appropriate substance. Then write, CA and CB on the product side ***above*** the appropriate substance.

1. $CO_3^{-2} + H_2O \rightarrow HCO_3^{-1} + OH^{-1}$

2. $HCN + H_2O \rightarrow H_3O^+ + CN^{-1}$

3. $H_2S + (CH_3)_2NH \rightarrow (CH_3)_2NH_2^+ + HS^{-1}$

4. $HF + H_2O \rightarrow H_3O^+ + F^{-1}$

5. $HC_2H_3O_2 + H_2O \rightarrow H_3O^+ + C_2H_3O_2^{-1}$

6. $HSO_4^{-1} + HPO_4^{-2} \rightarrow SO_4^{-2} + H_2PO_4^{-1}$

7. $N_2H_4 + H_2O \rightarrow N_2H_5^+ + OH^{-1}$

8. $NH_2OH + HCl \rightarrow NH_3OH^+ + Cl^{-1}$

9. $NH_2^{-1} + N_2H_4 \rightarrow NH_3 + N_2H_3^{-1}$

10. $HNO_3 + H_2SO_4 \rightarrow H_3SO_4^+ + NO_3^{-1}$

Chapter 20: Acids and Bases
Part 3 – Notes: pH, pOH, and K_w – Acid Concentrations

Recall: The equation for the ionization of water is: $2\ H_2O \rightarrow H_3O^+ + OH^-$

$$K_w = [H_3O^+][OH^-] = [1.00 \times 10^{-7}][1.00 \times 10^{-7}] = 1.00 \times 10^{-14} \text{ (at 25°C)}$$

The Concentration of Hydronium and Hydroxide

The simple relationship above has been used to construct the pH scale, used to indicate how acidic or basic a substance is.

What happens when an acid is dissolved in water? An acid is a substance whose water solution contains a large concentration of hydronium ions and at 25°C, the above equation must always be true.

So, what if the concentration of H_3O^+ goes up?

Similarly, a water solution of a base has a large concentration of OH^- and a small concentration of H_3O^+.

The above formula may be rearranged to determine the concentrations of both:

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

The pH Scale

The pH of a solution is the *negative logarithm, to base 10, of the hydronium ion concentration*. It is expressed by the equation:

In general, the pH scale is a numerical scale that normally extends from 0 through 14. The numbers on the scale represent the acidity of solutions and can be converted to hydronium concentrations. The midpoint of the scale is taken as 7.

At 25°C, a solution with a pH of 7.0 . . .

A solution with a pH ***less*** than 7.0 . . .

A solution with a pH ***greater*** than 7.0 . . .

The actual relationship between the pH and the molar concentration of the H_3O^+ is:

The pOH Scale

The pOH scale is used to indicate the hydroxide concentration of a solution. Note that as pH increases, the pOH decreases, and vice versa, because the product of the ion concentrations must always be 1.00×10^{-14}.

The pOH of a solution is the *negative logarithm, to base 10, of the hydroxide ion concentration*. It is expressed by the equation:

The actual relationship between the pOH and the molar concentration of the OH^- is:

The Relationship Between pH and pOH

We know that as pH increases the pOH decreases and that as pH decreases the pOH increases, we are able to relate the two scales. In any aqueous solution at 25°C:

Question: What is the $[H_3O^+]$ if the acid is 0.75*M* HCl?

What is the $[H_3O^+]$ if the acid is 0.75*M* H_2SO_4?

Example 1: What is the pH of a solution that contains 1.00×10^{-4} M hydronium ion?

Example 2: Calculate the $[H_3O^+]$ of a solution that has a pH of 3.70.

Example 3: Find the pH and pOH of a solution that contains 0.00350 M H_3O^+ ion.

Example 4: What is the $[OH^-]$ and the $[H_3O^+]$ of a solution if the pOH is 4.40?

Example 5: What is the pH of 0.250 M H_3PO_4? This is a strong acid and dissociates 100%.

Chapter 20: Acids and Bases
Part 3 – Notes: pH, pOH, and K_w – Acid Concentrations

Supply answers for the following questions. Show all work, set-ups, etc. Use additional paper if necessary to show all work.

1. Calculate the pH of the solutions with these concentrations:
   (a) 0.00010 mol H_3O^+/L
   (b) 0.0018 mol OH^-/L
   (c) 1.62×10^{-5} mol OH^-/L
   (d) 4.09×10^{-2} mol H_3O^+/L

2. Calculate the pOH of the solutions with these concentrations:
   (a) 4.57×10^{-6} mol OH^-/L
   (b) 5.75×10^{-2} mol H_3O^+/L

3. Calculate the $[H_3O^+]$ of the following solutions.
   (a) pH = 3.72
   (b) pOH = 12.0

4. Calculate the $[OH^-]$ of the following solutions.
   (a) pOH = 11.0
   (b) pH = 4.25

5. Calculate the (i) pH and the (ii) pOH of solutions having the following concentrations. Assume 100% ionization. **Remember:** 1 mole of H_2SO_4 produces TWO moles of H_3O^+ ions, etc.

(a) 0.0025 M NaOH (b) 0.0025 M H_2SO_4

(c) 0.075 M H_2SO_4 (d) 0.048 M HCl

(e) 0.032 M KOH (f) 0.00017 M $Ca(OH)_2$

6. The pH of human blood is 7.4. What is the $[H_3O^+]$?

7. What is the $[H_3O^+]$ in a red ripe tomato that has a pH of 4.2?
What is the pOH? What is the $[OH^-]$?

8. Find the $[H_3O^+]$ and the pH of the following solutions:

(a) 0.325 M H_2SO_4 (b) 0.0375 M $Ca(OH)_2$

9. Find the $[OH^-]$ and the pH of the following solutions:

(a) 0.00862 M $Mg(OH)_2$ (b) 1.25 M H_3PO_4

Chapter 20: Acids and Bases
Part 4 – Notes: Titration and Neutralization

Titration:

In acid-base titration, a base is added to an acid (or vice versa) until . . .

This point of **neutralization** is called the ***end point*** or ***equivalence point***. These *ions react to form water*. An indicator tells when the right amount of acid and base have been mixed.

When a hydronium ion combines with a hydroxide ion, the result is the neutral compound water: $H_3O^+ + OH^- \rightarrow HOH$

Strong and Weak Acids and Bases
A strong acid or a strong base will completely ionize or dissociate. Weak acids and weak bases will not completely ionize.

The ***strong acids*** you will commonly encounter are:

The ***strong bases*** you will commonly encounter are:

The ***moderately strong acids*** you will encounter are:

The ***weak acids*** you will commonly encounter are:

Examples:

	$HCl \rightarrow H_3O^+ + Cl^-$	one mole of HCl produces one mole of H_3O^+ ions
	$NaOH \rightarrow Na^+ + OH^-$	one mole of NaOH produces one mole of OH^- ions
	$H_2SO_4 \rightarrow 2\ H_3O^+ + SO_4^{-2}$	one mole of sulfuric acid produces 2 moles of H_3O^+ ions

Steps for Solving Titration Problems
1. Write a balanced equation for the neutralization of the acid and base.
2. Use the information given in the problem to find the moles of either acid or base.
3. Use the mole ratio from the balanced equation to calculate the moles of the unknown acid or base.
4. Convert the moles of unknown into whatever units the problem requires.

Example 1:
45.00 mL of a 0.350 M H_2SO_4 solution is neutralized by a 0.425 M NaOH solution. Calculate the volume of NaOH required.

Example 2:
50.0 mL NaOH solution is used to exactly neutralize 2.00 g of a solid acid with a molar mass of 180. g/mol. This acid has a single ionizable hydrogen. Find the concentration of the sodium hydroxide solution.

Example 3:
2.50 g of H_2X (a strong acid with two ionizable hydrogen) is exactly neutralized by 57.8 mL of 1.25 M NaOH solution. What is the molar mass of the acid?

Chapter 20: Acids and Bases
Part 4 – Assignment: Titration and Neutralization

Answer the following questions; show all work, set-ups, units, etc. Use additional paper if necessary.

1. 30.0 mL of 0.200 M HCl is exactly neutralized by 50.0 mL of $Ca(OH)_2$ solution. Calculate the concentration of the $Ca(OH)_2$ solution.

2. 30.0 mL of 0.200 M H_2SO_4 is exactly neutralized by 50.0 mL NaOH solution. Calculate the NaOH concentration.

3. How many mL of 0.400 M HCl are required to react with 10.0 g of solid calcium hydroxide?

4. 32.50 mL of NaOH solution are required to exactly neutralize 1.50 g of a solid acid having a molar mass of 180. g/mol. The acid has two ionizable hydrogen. Find the concentration of the NaOH solution.

5. 2.25 g of an unknown acid, H_2X, having two ionizable hydrogens, is exactly neutralized by 42.5 mL of 0.800 M NaOH solution. Calculate the molar mass of the acid, H_2X.

6. How many grams of barium hydroxide would be needed to completely neutralize 30.0 mL of 0.500 M phosphoric acid?

7. 40.0 mL of an acetic acid solution is exactly neutralized by 30.0 mL of 0.250 M barium hydroxide solution. Calculate the concentration of the acetic acid solution.

8. What mass, in grams, of calcium hydroxide would be required to react completely with 25.0 mL of 0.300 M sulfuric acid solution?

Chapter 20: Acids and Bases
Part 5 – Notes: Hydrolysis of Salts

Hydrolysis:

Hydrolysis occurs when certain salts dissolve in water to form solutions that have acidic or basic properties. A rule for predicting the properties of solutions of salts is based upon the concept of strong acids and strong bases in the Arrhenius sense. This rule applies to salts from different combinations of strong and weak acids and bases *except* for those salts formed from a weak acid and a weak base. Below is a summary of the rules for the formation of acidic or basic solutions when salts are dissolved in water.

Salt formed from . . .	***Water solution exhibits***	***Example***
Strong acid + strong base		
Strong acid + weak base		
Weak acid + strong base		
Weak acid + weak base		

Recall: The strong acids commonly used are:

The strong bases commonly used are:

What does an acid generally supply during the formation of a salt?

What does a base generally supply during the formation of a salt?

Example 1:

Does hydrolysis occur when KCl is dissolved in water? Why? If yes, what type of solution is produced?

Example 2:

Does hydrolysis occur when $Al_2(SO_4)_3$ is dissolved in water? Why? If yes, what type of solution is produced?

Example 3:

Does hydrolysis occur when K_2CO_3 is dissolved in water? Why? If yes, what type of solution is produced?

Chapter 20: Acids and Bases
Part 5 – Assignment: Hydrolysis of Salts

Supply answers to the following questions.

Predict the hydrolysis effect in a solution of each of the following salts. For answer, write *acidic*, *basic*, or *neutral*. Also indicate the pH (<7, >7, or =7) of the resulting solution and list the acid and base that react to produce the result indicated by the problem.

		A/B/N	***pH***	***Acid + Base***
1.	Na_3PO_4			
2.	Na_2SO_4			
3.	$K_2C_2O_4$			
4.	NH_4Cl			
5.	$FeCl_3$			
6.	NH_4NO_3			
7.	$Ca(NO_3)_2$			
8.	$KC_2H_3O_2$			
9.	$NaC_2H_3O_2$			
10.	K_2CO_3			

Chapter 20: Acids and Bases
Part 6 – Notes & Assignment: Preparation of Acids, Bases, and Salts

A ***binary acid*** can be prepared by a direct combination (synthesis) reaction. When the product is dissolved in water, the acid solution is formed.

The two reactions that take place in the preparation of hydrogen bromide are:

$$H_{2\,(g)} + Br_{2\,(g)} \rightarrow 2\,HBr_{(g)}$$
$$HBr_{(g)} + H_2O \rightarrow H_3O^+_{(aq)} + Br^-_{(aq)}$$

Question 1:
A 40.0 g sample of $Br_{2\,(g)}$ reacts with excess hydrogen. The $HBr_{(g)}$ is then dissolved in 1.75 L of solution.

a. What is the number of moles of HBr produced?

b. What is the molarity of the resulting solution?

c. What is the number of H_3O^+ ions in solution?

d. What is the pH of the solution?

A ***ternary acid*** may be prepared by adding a salt of the desired acid to concentrated sulfuric acid. The mixture is then heated. Sulfuric acid is used as a reactant (source of H^+) because it has a high boiling point. Since the desired acid has a relatively low boiling point, it boils out of the reacting mixture and is collected as a gas. The gas is then dissolved in water to form the acid solution. When nitric acid is prepared this way, the two reactions that take place are:

$$NaNO_{3\,(s)} + H_2SO_{4\,(l)} \rightarrow NaHSO_{4\,(s)} + HNO_{3\,(g)}$$
$$HNO_{3\,(g)} + H_2O \rightarrow H_3O^+_{(aq)} + NO_3^-{}_{(aq)}$$

Question 2:
A 75.0 g sample of $NaNO_3$ is treated with excess H_2SO_4, forming $HNO_{3\,(g)}$. The HNO_3 gas produced is then dissolved in 0.500 kg water?

a. What mass of H_2SO_4 reacts?

b. What is the number of moles of HNO_3 produced?

c. What is the molality of the solution?

d. What additional information would be required to determine the molarity of the sample?

An ***acid anhydride*** is an oxide of a nonmetal that will dissolved in water to form an acid as the only product. The reaction that takes place when P_4O_{10} reacts in this was is:

$$P_4O_{10\,(s)} + 6\ H_2O_{(l)} \rightarrow 4\ H_3PO_{4\,(aq)}$$

Question 3:
An 85.0 g sample of solid P_4O_{10} is dissolved to form 1.50 L of aqueous acid solution.

a. What is the molarity of H_3PO_4 in the resulting solution?

b. What mass of water reacts with the solid P_4O_{10}?

c. What is the number of PO_4^{-3} ions available in the aqueous solution?

d. What is the number of H_3O^+ ions available in the aqueous solution?

e. What is the pH of the acidic solution?

Preparation of Bases

When an active metal is added to water, a solution of a hydroxide compound is produced. When potassium hydroxide is prepared this way, the reaction is:

$$2\ K_{(s)} + 2\ HOH_{(l)} \rightarrow 2\ KOH_{(aq)} + H_{2\,(g)}$$

(Note: Using HOH to represent water in acid/base systems sometimes makes equations easier to understand and balance.)

Question 4:
A 5.00 g sample of solid potassium is added to enough water to prepare 5.00 L of solution.
(Caution: Addition of an active metal to water is very hazardous and should be done only under supervision.)

a. What volume of $H_{2\,(g)}$, measured at STP, is produced?

b. What is the concentration, in molarity, of KOH in the resulting solution?

c. What is the pOH of the solution?

A ***basic anhydride*** is a metal oxide that reacts with water to form a base as the only product. When CaO reacts in this way, the reaction is:

$$CaO_{(s)} + HOH_{(l)} \rightarrow Ca(OH)_{2\ (aq)}$$

Question 5:

The solubility of $Ca(OH)_2$ in water is 0.509 g/L at room temperature.

a. What is the minimum mass of solid CaO that must be added to water to produce 2.50 L of a saturated solution of $Ca(OH)_{2\ (aq)}$?

b. What is the concentration, in molarity, of $OH^-_{(aq)}$ in a saturated solution of $Ca(OH)_{2\ (aq)}$?

Ionization and Dissociation

In water solutions, acids dissolve and ionize to form $H_3O^+_{(aq)}$ ions and corresponding anions. In water solutions, bases (generally, hydroxide compounds) dissolve and dissociate into $OH^-_{(aq)}$ ions and corresponding cations.

Question 6:

What is the molarity of H_3O^+ when 45.0 g $HCl_{(g)}$ is dissolved into 400. mL of solution.

Question 7:

A solution is prepared by adding 14.0 g $NaOH_{(s)}$ to 100. mL of 1.50 M $KOH_{(aq)}$. (Assume no change in volume.) What is the total molarity of the OH^- in solution?

Chapter 20: Acids and Bases
Part 7 – Notes: Acids, Bases and Ionization Equilibria

A ***strong acid*** is one that ionizes completely (or almost completely) in an aqueous system.

What does this mean in terms of the equilibrium? How is it shifted?

What kind of value does the K_{eq} have for this system?

When a ***weak acid*** dissolve in water, only partial ionization occurs.

What does this mean in terms of the equilibrium? How is it shifted?

Consider the ionization of HNO_2:

The above equation represents the state of equilibrium in a solution of the weak acid HNO_2. Although it does not totally ionize, the concentration of the hydronium ions is sufficient enough to give the solution acidic properties.

Ionization constants:

When ***strong bases***, such as NaOH or KOH, dissolve in water they undergo the process of dissociation, which is nearly 100% complete. Like strong acids, strong bases have an *equilibrium constant that is so large that its numerical value is meaningless*.

In a solution of a ***weak base*** a state of equilibrium exists. For the weak base NH_4OH there is an equilibrium represented by the equation: $NH_3 + H_2O \longleftrightarrow NH_4^+ + OH^-$. Since the hydroxide ion is formed by a reaction in which dissociation occurs, for NH_4OH the equilibrium constant is known as an ionization constant, like that of weak acids.

Ionization Constants

For the reaction of a weak acid, such as HNO_2, with water, we can derive the equilibrium constant as follows:

$$HNO_2 + H_2O \longleftrightarrow H_3O^+ + NO_2^-$$

$$K_a \text{ or } K_I = \frac{[H_3O^+][NO_2^-]}{[HNO_2]}$$

H_2O is a constant and therefore not included in the K_a.

What does the K_a or the K_b (more generally, the K_I) indicate?

If the value of K_a or K_b is large, what does it indicate?

If the vale of K_a or K_b is small, what does it indicate?

NOTE: All values of K are positive.

Solutions of Weak Acids and Bases

General equation for the ionization of a weak acid is written: $HX + H_2O \longleftrightarrow H_3O^+_{(aq)} + X^-_{(aq)}$

In this system, the equilibrium is NOT completely displaced to the right; so the dissociation is NOT complete. **All equilibrium systems containing weak acids or weak bases favor reactants instead of products.** You need to use the dissociation constant of the acid (K_a) or the dissociation constant of the base (K_b).

Remember the following when working with problems of weak acids or weak bases:

a. The original concentration of the weak acid is **NOT** equal to the concentration of hydronium. **$[H_3O^+]$ < [weak acid]**

b. The original concentration of the weak base is **NOT** equal to the concentration of hydroxide. **$[OH^-]$ < [weak base]**

c. Equilibrium systems having weak acids or weak bases always favor the REACTANTS, and not the products, as in the case of strong acids or strong bases. So you must use K_a or K_b in order to find $[H_3O^+]$ or $[OH^-]$.

d. You can calculate the percentage dissociation of the acid or the base by using the following formula:

$$\textbf{percentage dissociation} = \frac{[H_3O^+]}{\text{original concentration of the acid}}$$

Example 1 : Find (a) the $[H_3O^+]$, (b) pH, and (c) percentage dissociation for 0.100 M acetic acid at 25°C. $K_a = 1.80 \times 10^{-5}$

Example 2: A 0.500 M solution of a weak acid, HX, is only partially ionized. The $[H_3O^+]$ was found to be 2.63×10^{-3}M. Find the dissociation constant for this acid.

Chapter 20: Acids and Bases
Part 7 – Notes: Acids, Bases and Ionization Equilibria

Answer the following questions on a **separate sheet of paper**. Show all work, set-ups, units, etc.

1. Calculate the (a) pH and (b) K_a of a 0.100 M acetic acid solution if 1.34% of the CH_3COOH molecules in this solution have ionized to form product ions.

2. Calculate (a) the pH and (b) the percent dissociation of a 0.200 M solution of HCN. $K_a = 4.90 \times 10^{-10}$

3. A 0.100 M solution of acetic acid is only partially ionized. Using a measure of pH, the $[H_3O^+]$ is determined to be $1.34\text{x}10^{-3}$ M. What is the acid dissociation constant (K_a) of acetic acid?

4. A 0.375 M solution of a weak acid, H2X, is only partially ionized. The $[H_3O^+]$ was found to be 4.58×10^{-4}M. Find the dissociation constant for this acid.

5. Calculate the acid dissociation constant of a weak monoprotic acid if a 0.500 M solution of this acid gives a hydronium ion concentration of 0.000100 M.

6. Determine the (a) $[H_3O^+]$, (b) $[OH^-]$, (c) pH, and (d) pOH of a 1.00×10^{-4}M HCN solution that undergoes 3.5% ionization.

7. Determine the (a) $[H_3O^+]$, (b) $[OH^-]$, (c) pH, and (d) pOH of a 1.60×10^{-3} oxalic acid solution that undergoes 0.166% ionization. $H_2C_2O_4 + H_2O \Leftrightarrow H_3O^+ + HC_2O_4^-$

8. What would be the ionization constant for acetic acid at 25°C is a 0.0100 M solution of this acid is ionized to the extent of 1.38%?

9. A 0.500 M solution of a weak acid, HX, has a pH of 1.75. Calculate the K_a of HX.

10. A weak acid, HX, has $K_a = 5.00 \times 10^{-6}$. A certain solution of this acid has pH = 2.70. Calculate the concentration of the acid in this solution.

11. Nitrous acid, HNO_2, has K_a of 5.1×10^{-4}.
    a. Calculate the pH of 0.25 M HNO_2.

    b. Calculate the percent ionization of 0.25 M HNO_2.

12. A 0.010 M solution of acid HX is 7.5% ionized. Calculate the K_a and the pH of this acid.

Chapter 20: Acids and Bases
Part 8 – Notes: Buffers and Acid-Base Wrap Up Information

Buffers:

You can create a buffer by . . .

Example: If acetic acid (HAc) and sodium acetate (NaAc) were dissolved into the same solvent you would create a buffer.

a. The conjugate base of HAc is Ac^{-1}. So by dissolving NaAc into water . . .

b. The sodium ions don't play any role here so they are left out of the chemical equation.

c. Notice this is the same as the equilibrium equation for the ionization of acetic acid with a Ka of $1.8x10^{-5}$.

d. Given such a small Ka, not many acetate ions (Ac^{-1}) will form naturally. However, by adding sodium acetate, a large number of the Ac^{-1} ions will be present. So, once you set up the buffer, there will be a large number of undissociated HAc and also a large number of Ac^{-1} ions present.

e. Now consider adding a strong acid to this equilibrium system, (addition of a H_3O^+ donor). According to Le Chatelier, what will happen to this system?

f. Now consider adding a strong base to this equilibrium system (addition of a H^+ acceptor). Anything with OH^{-1} ions will dissolve to react with the H_3O^+ ions to form H_2O. According to Le Chatelier, what will happen to this system?

Without buffers, we would not be able to maintain the proper pH of our blood. Blood pH must remain between 7.35 and 7.45 or we might die.

There are several buffer systems in our bodies that help to maintain our proper pH levels. One of them involves our breathing.

- Our blood contains significant amounts of both carbonic acid (H_2CO_3) and its conjugate base, bicarbonate ion (HCO_3^{-1}). This creates the equilibrium:

- Carbonic acid also forms the equilibrium:

- By combining these two equations, you get:

- As our bodies metabolic processes produce acids, equilibrium is shifted from right to left and we exhale the CO_2, maintaining proper levels of all chemicals.

- You can raise your blood pH by hyperventilating which removes excessive amounts of carbon dioxide causing the entire system of equilibrium equations to shift left removing hydronium ions from the system. If you hyperventilate for too long you could enter *alkalosis*.

What are some common acids and bases you have encountered in and around your home?

Baking soda:

Drain cleaner and oven cleaner:

Vinegar:

Citrus fruits:

Battery acid:

Antacids like Milk of magnesia:

Household ammonia (aqueous ammonia):

What are some problems associated with "acid rain"?

Why do some shampoo bottles say "acid-balanced"?

What do antacids do? Are there any problems with taking too much?

What is the difference between acid strength and acid concentration?

Is it true that a neutral solution has a pH of 7?

How does endpoint and neutralization point differ?

MISCONCEPTION: For sulfuric acid, to say that dissociation is complete means that the acid has lost both protons.

To say that the dissociation of sulfuric acid is complete does not mean that the species in solution are two hydronium ions and one sulfate ion for each molecule dissolved. The fact is, sulfuric acid is "strong" only in regard to the removal of the first hydrogen ion.