Chapter 14 Acid-Base Equilibria Study Guide

This chapter will illustrate the chemistry of acid-base reactions and equilibria, and provide you with tools for quantifying the concentrations of acids and bases in solutions.

Purpose: This is a guide for your as you work through the chapter. The major topics are provided so that you can write notes on each topic and work the corresponding problems.

This should serve a s a study guide as you go on to do the problems in Sapling and take the quizzes and exams.

The Problems are embedded in the Topics and Space for Notes

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pH = -log[H_3O^+] **pOH = [OH^-]** **pH + pOH = 14** **[H_3O^+][OH^-]= 1.0x10^{-14}** **[H_3O^+]= 10^{-pH},** **$K_aK_b=K_w$** **pH = pKa + log($\frac{[base]}{[acid]}$)**

The reactions in Sections 1 – 6 are hydrolysis reactions (reaction with water) except for strong bases which just dissociate.
strong means strong electrolyte which means →and no ICE
weak means weak electrolyte which means equilibria (⇄ and ICE)
acid means hydrolysis to make H_3O^+ and pH < 7.
Base means hydrolysis to make OH^- and pH > 7.
Memorize the strong acids and strong bases. HCl, HBr, HI, HNO_3, $HClO_3$, $HClO_4$ NaOH, KOH, $Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$

Section 1 14.1 Bronsted-Lowry Acids and Bases:
By the end of this section, you will be able to:

- Identify acids, bases, and conjugate acid-base pairs according to the Brønsted-Lowry definition
- Write equations for acid and base ionization reactions (hydrolysis)
- Use the ion-product constant for water to calculate hydronium and hydroxide ion concentrations
- Describe the acid-base behavior of amphiprotic substances

Define Arrhenius acid and base.

Define Bronsted-Lowry Acid and Base

Identify acids and bases and write their hydrolysis reactions:

Identify acid/base conjugate pairs.

Define amphoteric. Describe the amphoteric behavior of water.

Use K_w to calculate the concentration of $[H_3O^+]$ and $[OH^-]$. $[H_3O^+][OH^-] = 1.0 \times 10^{-14}$

How is HCO_3^- amphoteric? Write its reactions.

Sample Question 14.1

(A) Label each as an acid or base.
Write the conjugate pair for each and then write the hydrolysis reactions for:

1. C_5H_5N

2. HClO

3. CH_3COO^-

4. NaOH (remember… no hydrolysis reaction)

5. C_6H_5COOH

(B) Fill in the table: $[H_3O^+][OH^-]= 1.0x10^{-14}$

Recognize the strong acids. (HCl, HBr, HI, HNO_3, $HClO_3HClO_4$ and H_2SO_4 first H only)

With strong acids K_a is undefined and the concentration of the acid = concentration of the H_3O^+.

Strong bases (NaOH, KOH, $Ca(OH)_2Sr(OH)_2$, $Ba(OH)_2$) .. calculate the $[OH^-]$. Take pOH and then pH.

$[H_3O^+]$	$[OH^-]$	Acidic/ basic or neutral
1.0×10^{-4} M		
	1.0×10^{-9} M	
0.050 M HCl		
	.0015 M $Ca(OH)_2$	

Section 14.2 pH and pOH

By the end of this section, you will be able to:

- Explain the characterization of aqueous solutions as acidic, basic, or neutral
- Express hydronium and hydroxide ion concentrations on the pH and pOH scales
- Perform calculations relating pH and pOH
- Autoionization of H_2O and pH. $pH = -\log[H_3O^+]$; $pOH = [OH^-]$, $pH + pOH = 14$;

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

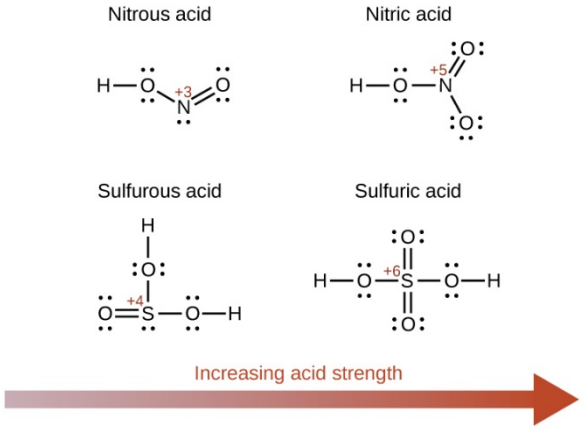
  If $[H_3O^+] > [OH^-]$, then $pH < 7$ and solution is acidic. If $[OH^-] > [H_3O^+]$, then $pH > 7$ and the solution is basic.

$[H_3O^+]$	$[OH^-]$	pH	pOH	A/B or neutral
1.0 M				
		4.5		
	2.5×10^{-4} M			
			3.5	
	1.0×10^{-7} M			
		10.8		
0.00025 M				

Section 14.3 Relative Strength's of Acids and Bases

By the end of this section, you will be able to:

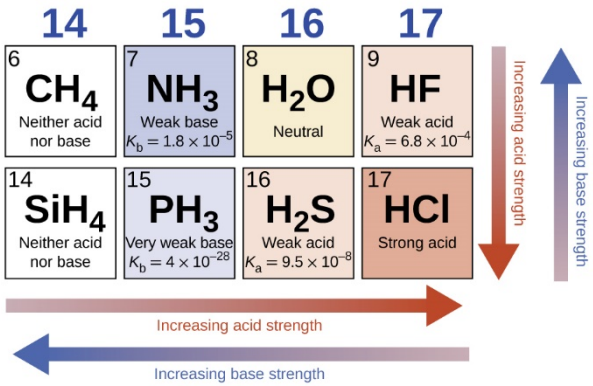
- Assess the relative strengths of acids and bases according to their ionization constants
- Rationalize trends in acid–base strength in relation to molecular structure
- Carry out equilibrium calculations for weak acid–base systems



14.3.1
Know K_a, K_b for strong acids and strong bases

Be able to find and use K_a, K_b for weak acids and weak bases

What is the trend of acid strength as a function of K_a? (Bases?)



14.3.2 Acid base strength as a function of molecular structure. (Binary acids)

For oxyacids:
As the net electronegativity increases, the acid strength increases.. (as the number of oxygens increases, the acid strength increases.)

14.4 Equilibrium calculations of weak acid and weak base solutions.

For weak acids (rest of inorganic acids and organic acids with the formula: : R-OH and R-COOH…
look up K_a, write your balanced reaction, write the equation, Do the ICE chart, solve for something.
For Weak bases. (Most weak bases contain an N with 3 bonds and a lone pair. NH_3 is the prototype. For weak bases, look up K_b, write the balanced reaction, write the equation, do the ICE chart and solve for something.

The charts of K_a's and K_b's are at the back of the text.

1. Calculate the pH of 0.25 M HClO(aq).

2. The pH of a 0.50 M solution of a weak acid is 1.37. What is the K_a for the acid?

3. What is the pH of 0.035 M pyridine (aq)?

14.5 By the end of this section, you will be able to:

- **Predict whether a salt solution will be acidic, basic, or neutral**
- **Calculate the concentrations of the various species in a salt solution**
- **Describe the process that causes solutions of certain metal ions to be acidic**

Thoughts: Acid/base properties of salts. Cations will be acidic and Anions will be basic. The conjugate acids of strong bases will inert. (Na^+, K^+, and the rest of group 1, Ca^{2+}, Sr^{2+}, Ba^{2+}.)
The conjugate bases of the strong acids are inert. (Cl^-, Br^-, I^-, NO_3^-, ClO_3^-, ClO_4^-).
Otherwise conjugate bases of weak acids (A^-) are basic. And conjugate acids of weak bases are acidic.
Remember that K_a(acid)K_b(conjugate base)=K_w. KaKb=Kw. So.. as the acid strength (Ka) goes up, the strength of the conjugate base (Kb) goes down. The conjugates of weaks will be weak. So.. if you have the conjugate of a weak acid or base (example NaF). Separate into ions.. Na^+ and F^-. recognize that Na^+ will be inert. Calculate Kb for F^- from Ka of HF (KaKb=Kw). Hydrolysis, ICE, solve for something.
Remember… the stronger the acid the weaker the conjugate base and the stronger the base the weaker the conjugate acid.

1. Calculate the pH of 0.25 NaBrO(aq).

2. Calculate the pH of 0.15 M $C_5H_5NHNO_3$(aq).

3. Rank from the most acidic to the most basic 0.10 M solutions of:
$(CH_3)_3NHBr$, KNO_2, HNO_2, $(CH_3)_3N$, KBr, KOH

(Hint: look up the K's) calculate the relevant K!

_____ _____ _____ _____ _____

Most acidic most basic

4. Circle the correct answer.

(A) Strongest acid: HNO_2 or HNO_3

(B) Strongest acid: Co^{2+} or Co^{3+}

(C) Weakest base: Cl^- or F^-

(D) Strongest base: ClO_3^- or ClO^-

14.6 By the end of this section, you will be able to:

- **Extend previously introduced equilibrium concepts to acids and bases that may donate or accept more than one proton**
- **Polyprotic acids. More than 1 H^+, If $K_{a1}>>K_{a2}$.. ignore K_{a2} and just do an ICE chart for the first H^+. If K_{a2} is not small, then you will have to use the quadratic.**

1. What is the pH of 0.10 M HCl(aq) with 0.10 M HF(aq)?

2. What is the pH of 0.500 M H_2CO_3(aq)?

3. What is the pH of 0.75 M H_2SO_4(aq)?

14.7 By the end of this section, you will be able to:

- **Describe the composition and function of acid–base buffers**
- **Calculate the pH of a buffer before and after the addition of added acid or base**

Thoughts: Use the Henderson-Hasselbach**.** pH = pKa + log([base]/[acid])
Add a strong acid or base to a buffer (write your mole chart):
Understand Blood Buffers:
Creating a Buffer (get a pKa near desired pH)

1. What is a buffer? Give an example.

2. A solution is make of 1.00 M of 0.45 M NH_3(aq) with 0.65 M NH_4Cl(aq). What is the pH of the solution?

What is the pH after 0.10 moles of NaOH is added to the solution?

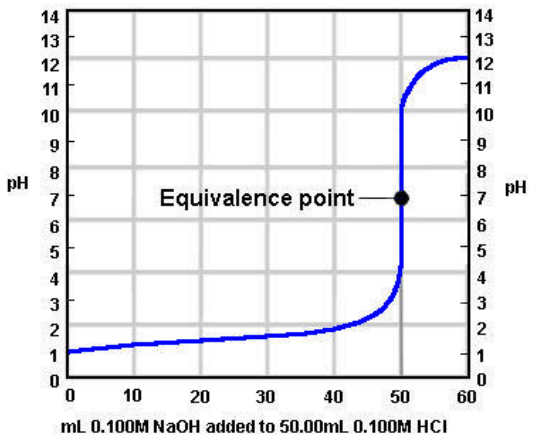
3. Which of these solutions will make a buffer?

(A) HCl with NaCl (B) NH_3 with NH_4NO_3 (C) C_5H_5COOH with NaC_5H_5COO

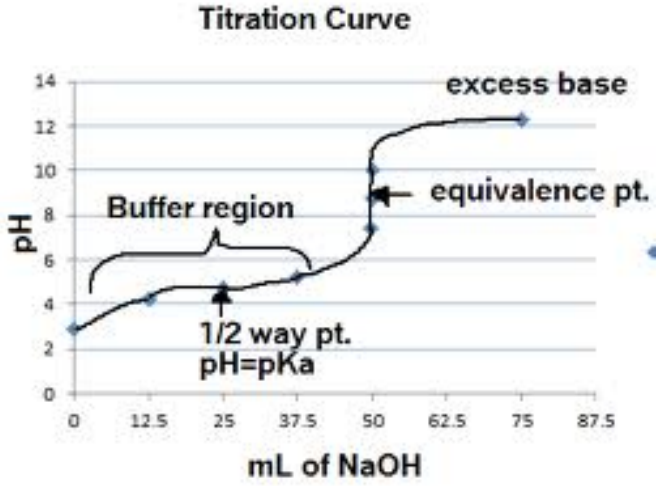
4. Write the reactions that will occur when strong acid and strong base is added to a solution that contains H_2CO_3 and $NaHCO_3$.

14.8 By the end of this section, you will be able to:

- **Interpret titration curves for strong and weak acid-base systems**
- **Compute sample pH at important stages of a titration**
- **Explain the function of acid-base indicators**



Thoughts: Strong/ Strong titrations
Mole chart needed
pH at equivalence point is 7



Thoughts: strong/weak titrations
mole chart needed
pH at equivalence point ≠ 7
½ way to equivalence point.. pH= pKa

1. 50.0 mL of 0.100 M KOH (aq) is titrated with 0.100 M HNO_3(aq)

What is the pH..

Initially

½ way to the equivalence point

At the equivalence point

1 mL past the equivalence point.

Graph the titration curve!

2. 40.0 mL of 0.200 M benzoic acid is titrated with 0.200 M NaOH (aq).

Question	What is the solution	V of NaOH added	V total?	Re-evaluate and what do we have?	pH
(A) Initial	Weak acid, need….?	0 mL	NA	still weak acid	
(B) ½ way to eq pt.	Adding strong base				
(C) 5.0 mL before Eq Pt.					
(D) At eq pt					
(E) 2.0 mL past eq pt					

(A) What is the pH of the acid solution initially?

(B) What is the pH of the acid solution ½ way to the equivalence point?

(C) What is the pH 5.0 mL before the equivalence point?

(D) What is the pH at the equivalence point?

(E) What is the pH 2.0 mL past the equivalence point?

(F) Graph this titration curve. What would be a good indicator for this titration?