

## Unit 8: Chemical Equilibrium and Acid-Base Chemistry

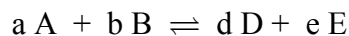
### Chemical Equilibrium

In this unit, you will learn about two major topics:

- 1) the quantitative and descriptive behavior of reversible reactions
- 2) acid-base chemistry and the pH scale.

### Equilibrium Expressions and Equilibrium Constants ( $K_{eq}$ )

A **reversible reaction** with the general form:



is quantitatively governed the following algebraic formula **at equilibrium**:

$$K_{eq} = \frac{[D]_{eq}^d [E]_{eq}^e}{[A]_{eq}^a [B]_{eq}^b}$$

This formula is called the **equilibrium expression** or equilibrium law (Historically it has also been called the mass action expression.)

Where:

1. [ ] = **Molar concentration** (moles/liter) of a dissolved (**aq**) species or gaseous (**g**) species **at equilibrium**.
2. Since the concentration of a pure solid or a pure liquid is a constant, **pure solids and liquids are omitted** from the equilibrium expression.
3.  $K_{eq}$ , the equilibrium constant is:
  - a. unique for each chemical reaction
  - b. independent of the initial concentrations of the reactants and products
  - c. dependent on (that is, it changes with) temperature
4. Since the equilibrium expression is really a ratio of  $\frac{[\text{products}]}{[\text{reactants}]}$ ,
  - if  $K_{eq}$  is a **large** number, then at equilibrium **[products] > [reactants]**
  - if  $K_{eq}$  is a **small** number, then at equilibrium **[reactants] > [products]**

**Problem Solving Applications***Problems Involving a System at Equilibrium*

1. Given all the quantities in the equilibrium expression except one, find the unknown quantity.

Solution: Substitute and solve.

*Problems Involving a System Reacting from Initial Concentrations to Equilibrium Concentrations*

2. Given the value of  $K_{eq}$  of particular reaction and the concentration of each of the reactants and products in a certain mixture, determine whether the mixture (a.k.a. the “system”) has reached equilibrium.

Solution: Consider again the generic reaction on the previous page and a mixture containing some A, B, D, and E.

Assign the letter "Q" to the ratio: 
$$Q = \frac{[D]^d [E]^c}{[A]^a [B]^b}$$

Substitute the concentrations into formula and solve for Q.

Interpretation:

- If  $Q < K_{eq}$  the reaction will proceed to increase the [D] & [E] and decrease the [A] & [B] i.e. in the **forward** direction until  $Q = K_{eq}$ .
- If  $Q > K_{eq}$  the reaction will proceed to decrease the [D] & [E] and increase the [A] & [B] i.e. in the **reverse** direction until  $Q = K_{eq}$ .
- If  $Q = K_{eq}$  the reaction is **at equilibrium** and no further observable reaction will occur.

3. Given the initial concentrations of all reactants and products and the equilibrium concentration of one of them, find  $K_{eq}$ .

4. Given the initial concentrations of all reactants and products and the  $K_{eq}$ , find the equilibrium concentration of all reactants and products.

Solutions: Strategies to solve type 3 and type 4 are difficult to describe and are better understood through the examples you will see in class.

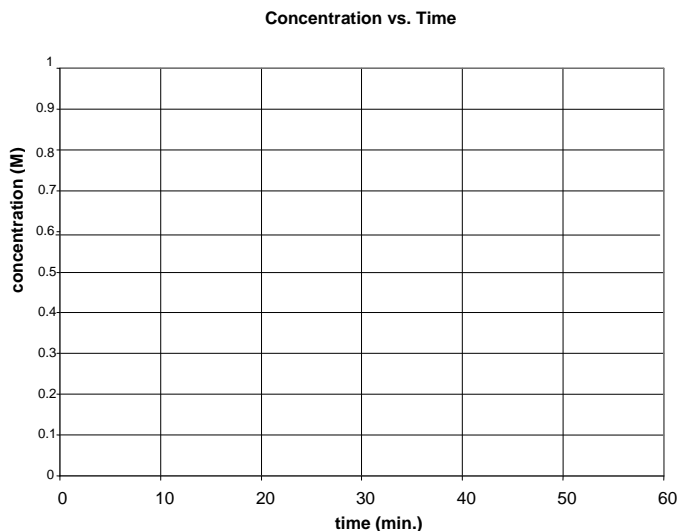
**Equilibrium Expression Problems**

1. The decomposition of phosphorus pentachloride,  $\text{PCl}_5$ , into phosphorus trichloride and chlorine gas is a reversible reaction that reaches a state of chemical equilibrium



One mole of  $\text{PCl}_5$  is placed in a sealed container and heated to  $250^\circ\text{C}$ . The concentrations of reactant and products are measured at ten-minute intervals. The following data are collected:

Time (min)	Concentrations (moles per liter)		
	$[\text{PCl}_5]$	$[\text{PCl}_3]$	$[\text{Cl}_2]$
0	1.00	0.00	0.00
10	0.90	0.10	0.10
20	0.85	0.15	0.15
30	0.82	0.18	0.18
40	0.80	0.20	0.20
50	0.80	0.20	0.20
60	0.80	0.20	0.20



Plot this data on the grid provided, and answer the questions below.

- a. How long did it take for the reaction to reach equilibrium? \_\_\_\_\_
- b. What is the equilibrium concentration of each reaction component?

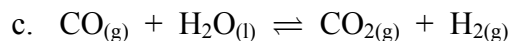
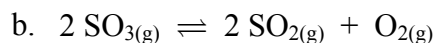
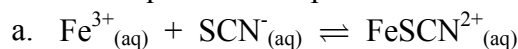
$[\text{PCl}_5]$  \_\_\_\_\_  $[\text{PCl}_3]$  \_\_\_\_\_  $[\text{Cl}_2]$  \_\_\_\_\_

- c. Why is the concentration of  $\text{PCl}_3$  always equal to the concentration of  $\text{Cl}_2$ ?
- 

- d. Predict the concentration of each component at 70 minutes.

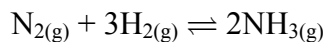
$[\text{PCl}_5]$  \_\_\_\_\_  $[\text{PCl}_3]$  \_\_\_\_\_  $[\text{Cl}_2]$  \_\_\_\_\_

2. Write the equilibrium expression for each of the following reactions:



(Answers on p.47)

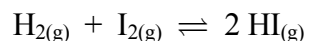
3.a. Given the following equation for the reversible reaction:



At 350 °C the equilibrium concentrations are determined to be:  $[\text{NH}_3] = 0.100 \text{ M}$ ,  $[\text{N}_2] = 0.100 \text{ M}$ , and  $[\text{H}_2] = 0.200 \text{ M}$ . Calculate  $K_{\text{eq}}$  at this temperature.

b. Some hydrogen is added to the same reaction system at the same temperature (i.e. the same  $K_{\text{eq}}$ ). After the system reaches equilibrium the  $[\text{N}_2]$  was 0.0800 M and  $[\text{H}_2]$  was 0.250 M, what is the concentration of  $\text{NH}_3$  in the system?

4. The equilibrium constant for the reaction below is 60.0 at 125°C.



a. If a mixture of 0.50 moles of HI and 0.10 moles of  $\text{H}_2$  and  $\text{I}_2$  are introduced into a 1.0L flask, is this system at equilibrium?

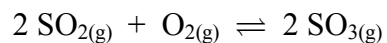
b. If it is not, will the reaction move in the forward or reverse direction? \_\_\_\_\_  
(Answers on p.47)

5. An experiment is designed such that 2.00 moles of  $\text{HI}_{(g)}$  are placed in a one liter flask at  $425^\circ\text{C}$ . It reacts to form  $\text{H}_{2(g)}$  and  $\text{I}_{2(g)}$  according to the following equation:



When equilibrium is reached, the concentration of  $\text{H}_2$  is determined to be 0.214 M. Calculate the equilibrium constant for this reaction at  $425^\circ\text{C}$ .

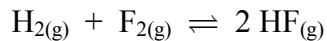
6. In an experiment 2.00 moles of  $\text{SO}_2$  and 2.00 moles of  $\text{O}_2$  are placed in a 2.00L flask at  $845^\circ\text{C}$ . The following reaction takes place in which some  $\text{SO}_3$  is formed:



When equilibrium is established, the concentration of  $\text{SO}_2$  is 0.15 M. Calculate  $K_{\text{eq}}$  for the reaction at  $845^\circ\text{C}$ .

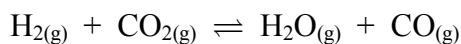
(Answers on p.47)

7. Suppose that one wanted to synthesize hydrogen fluoride from hydrogen and fluorine using the reaction below:



Initially, 3.00 moles of  $\text{H}_2$  and 3.00 moles of  $\text{F}_2$  are mixed in a 2.00 liter flask. If the equilibrium constant for this synthesis at  $25^\circ\text{C}$  is 115.0. Calculate the concentrations of all three substances when equilibrium is established.

8. Consider the following reversible reaction that has a  $K_{\text{eq}} = 4.40$  at  $1730^\circ\text{C}$ :



If 1.00 mole of  $\text{H}_2$  and 1.00 mole of  $\text{CO}_2$  are placed in a 4.50 L container at  $1730^\circ\text{C}$ , determine the equilibrium concentration of all four substances.

(Answers on p.47)

## Le Chatelier's Principle

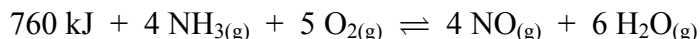
Le Chatelier's Principle states:

**When conditions are changed on a system at equilibrium, the system will react (forward or reverse) in order to partially counteract the change.**

- The forward or reverse reaction of the system is sometimes called a “**shift**”
- The system will then come to a “new” equilibrium where the substances will have a different concentration compared to their original equilibrium concentration.
- This is a general statement that has several applications. A change in conditions is sometimes referred to as a “**stress**” placed on the system and may be one of several types:
  - 1) Addition (or increase in concentration) of a gaseous or aqueous substance
  - 2) Removal (or decrease in concentration) of a gaseous or aqueous substance
  - 3) Increase the temperature of the system
  - 4) Decrease the temperature of the system
  - 5) Increase the pressure on a gaseous system by decreasing the volume
  - 6) Decrease the pressure on a gaseous system by increasing the volume
  - 7) Addition of a catalyst to the system
  - 8) Addition of a substance that is in the solid or liquid phase

### Le CHATELIER'S PRINCIPLE PROBLEMS

9. The oxidation of ammonia is a reversible reaction that proceeds as follows:



For each change (stress) described in the table:

- First use a  $\rightarrow$  or a  $\leftarrow$  to indicate the direction the system will shift.
- Then use a  $\uparrow$  to show an increase, a  $\downarrow$  to show a decrease or an = to show no change in the quantity in each column.

STRESS	SYSTEM SHIFTS	EQ. CONCENTRATIONS				$K_{\text{eq}}$
		$[\text{NH}_3]$	$[\text{O}_2]$	$[\text{NO}]$	$[\text{H}_2\text{O}]$	
NH <sub>3</sub> addition						
O <sub>2</sub> removal						
NO addition						
H <sub>2</sub> O removal						
increase temperature						
decrease temperature						
addition of catalyst						
		#MOLES PRESENT AT EQUILIBRIUM				
		NH <sub>3</sub>	O <sub>2</sub>	NO	H <sub>2</sub> O	
increase pressure (Vol. dec.)*						
decrease pressure (Vol. inc.)**						

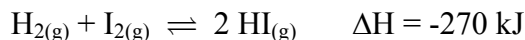
\*all equilibrium [ ]'s increase because of volume decrease

\*\*all equilibrium [ ]'s decrease because of volume increase

(Answers on p.48)

10-14. The information below applies to the five questions that follow.

The synthesis of hydrogen iodide is a reversible reaction that proceeds according to the equation:



In a laboratory experiment, hydrogen gas and iodine gas are heated to 125 °C in a sealed flask and hydrogen iodide is produced. At equilibrium the concentrations of reactants and products are measured and  $K_{\text{eq}}$  is calculated to be 60.0

Which answer best describes how the equilibrium concentrations would change when the following changes are introduced at equilibrium:

10. When the concentration of  $\text{H}_2$  is increased,
- The  $[\text{HI}]$  increases, the  $[\text{H}_2]$  and  $[\text{I}_2]$  remain the same.
  - The  $[\text{HI}]$  decreases, the  $[\text{H}_2]$  increases, and  $[\text{I}_2]$  decreases.
  - The  $[\text{HI}]$  increases, the  $[\text{H}_2]$  increases, and  $[\text{I}_2]$  decreases.
11. When the temperature of the system is decreased,
- The concentration all three components increases.
  - The  $[\text{H}_2]$  and  $[\text{I}_2]$  increases, the  $[\text{HI}]$  decreases.
  - The  $[\text{H}_2]$  and  $[\text{I}_2]$  decreases, the  $[\text{HI}]$  increases.
12. What happens to the value of  $K_{\text{eq}}$  when the temperature is decreased? \_\_\_\_\_
13. Describe what would happen to the number of moles of each gas (increase, decrease or remain the same) if the pressure were increased by decreasing the volume.
- $[\text{H}_2]$  \_\_\_\_\_  $[\text{I}_2]$  \_\_\_\_\_  $[\text{HI}]$  \_\_\_\_\_
14. If a catalyst is introduced at equilibrium, how would the equilibrium concentrations be affected?
- $[\text{H}_2]$  \_\_\_\_\_  $[\text{I}_2]$  \_\_\_\_\_  $[\text{HI}]$  \_\_\_\_\_
15. Suppose the reversible reaction system below has reached equilibrium.



How will the  $[\text{H}_2]$  be changed by imposing the following changes on the system?

- Adding  $\text{H}_2\text{O}$  \_\_\_\_\_
- Removing  $\text{C}_{(s)}$  \_\_\_\_\_
- Adding  $\text{CO}$  \_\_\_\_\_
- Removing  $\text{H}_2$  \_\_\_\_\_

Fill in the blanks for the following statements:

- Heating the reaction mixture will \_\_\_\_\_ the  $[\text{H}_2]$  and \_\_\_\_\_  $K_{\text{eq}}$ .
- Increasing pressure by decreasing volume will \_\_\_\_\_ number of moles of  $\text{H}_2$ .

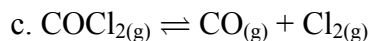
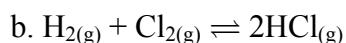
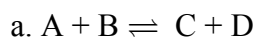
(Answers on p.48)

**UNIT 10 EQUILIBRIUM GRADED HOMEWORK ASSIGNMENT #2**

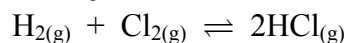
DO THIS ASSIGNMENT ON A SEPARATE SHEET OF PAPER

G1. Why are the concentrations of solids and liquids omitted from the equilibrium constant expression?

G2. Write the equilibrium constant expression (aka the equilibrium law) for the following reactions:



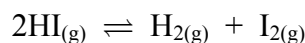
G3. At room temperature, the value of the  $K_{eq}$  for the following reaction is  $2.43 \times 10^{33}$ .



a. When this reaction reaches equilibrium, do you expect more products or more reactants to be present?

b. Explain your answer in terms of the equilibrium law for this reaction.

G4. Consider the following reversible reaction:



If the equilibrium concentrations at 448 °C are:

$$[HI]=0.0040M$$

$$[H_2]=0.0075M$$

$$[I_2]=0.00043M$$

calculate the equilibrium constant at this temperature.

G5. When ammonia dissolves in water, it reaches the equilibrium described by this equation:



If  $[NH_3]=0.25M$ ,  $[NH_4^+]=0.075M$ , what is the equilibrium concentration of  $[OH^-]$ ?

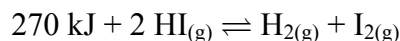
**UNIT 10 LeChatelier GRADED HOMEWORK ASSIGNMENT #5**

DO THIS ASSIGNMENT ON A SEPARATE SHEET OF PAPER

G1. When an equilibrium is shifted to the right:

- Do the concentrations of the reactants increase, decrease or remain the same?
- Do the concentrations of the products increase, decrease or remain the same?

G2. Use Le Chaterlier's principle and the following chemical equation to answer the questions below.

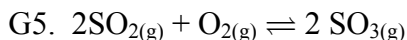
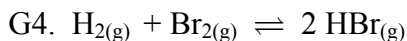
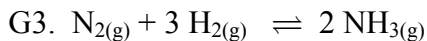


If the following conditions are changed, what will be the effect (if any) on the equilibrium concentration of hydrogen iodide?

- hydrogen gas is added
- the pressure of the system is increased by decreasing the volume
- the temperature of the system is increased
- the volume of the system is increased

For each of the following gas phase equilibrium systems:

Choose the statement below (A, B, or C) which best describes how the amount of product will be affected by a change in the pressure (due to a change volume).



- The maximum yield of products will occur if the pressure is increased.
- The maximum yield of products will occur if the pressure is decreased.
- The yield of products is unaffected by changes in pressure.

## Acid-Base Chemistry

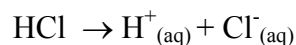
### Operational Definitions: Characteristics Physical and Chemical Properties

Acids	Bases
<ul style="list-style-type: none"> <li>taste sour</li> <li>conduct electricity when dissolved in water</li> <li>cause certain dyes (indicators) to change color (e.g. Litmus → Red)</li> <li>react with metals higher in activity series to produce H<sub>2</sub></li> <li>are <i>neutralized</i> during reaction with a base</li> <li>common examples: vinegar, lemon juice, saliva, urine, battery acid</li> </ul>	<ul style="list-style-type: none"> <li>taste bitter and feel slippery</li> <li>conduct electricity when dissolved in water</li> <li>cause certain dyes (indicators) to change color (e.g. Litmus → Blue)</li> <li>do not react with metals</li> <li>are <i>neutralized</i> during reaction with an acid</li> <li>common examples: soap, antacids, ammonia solutions</li> </ul>

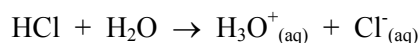
### Arrhenius Acid-Base Theory

Definition of acids and bases

An ACID is a substance that ionizes in water to produce H<sup>+</sup>:

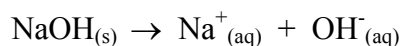


NOTE: The text represents this ionization as:



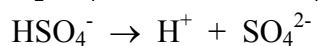
- The H<sub>3</sub>O<sup>+</sup> ion is called the *hydronium ion*. It is a combination of the H<sup>+</sup> ion and an H<sub>2</sub>O molecule.
- Don't be confused when you see it in the text; for all Arrhenius Theory purposes [H<sub>3</sub>O<sup>+</sup>] = [H<sup>+</sup>].

A BASE is a substance that ionizes in water to produce OH<sup>-</sup>:



### Polyprotic Acids

HCl is called a **monoprotic** acid because it produces one H<sup>+</sup> ions per molecule. When some compounds dissolve in water they ionize to produce two or more H<sup>+</sup> ions per molecule. These are called polyprotic acids. Sulfuric acid (H<sub>2</sub>SO<sub>4</sub>) is a **diprotic** acid because it undergoes the following two reactions in water:



Sometimes these ionizations must be considered separately, and other times they can be considered in combination as the net effect.

Two other examples of polyprotic acids are H<sub>2</sub>CO<sub>3</sub> (carbonic acid) which is present in all carbonated beverages and H<sub>3</sub>PO<sub>4</sub> (phosphoric acid) which is an example of a **triprotic** acid and is present in Coca-Cola.



**Strong Acids and Bases**

When an acid or base is described as being “strong”, it does NOT necessarily imply that it has high  $[H^+]$  or  $[OH^-]$ . Strong acids and bases **ionize completely** when dissolved in water. For example, a 1.0 M HCl solution contains virtually no HCl molecules. It does contain  $H^+$  and  $Cl^-$  ions both having a concentration of 1.0 M. When the compounds listed below dissolve in water they ionize completely to form either  $H^+$  ions (strong acids) or  $OH^-$  ions (strong bases).

This means: for monoprotic strong acids  $[acid] = [H^+]$   
and for Group 1 strong bases  $[base] = [OH^-]$

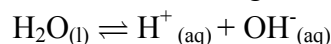
MEMORIZE THE FOLLOWING:

Strong Acids	Strong Bases
HCl (hydrochloric acid)	Group 1 hydroxides
$H_2SO_4$ (sulfuric acid)	LiOH
$HNO_3$ (nitric acid)	NaOH
	KOH
	RbOH

(other metal hydroxides, particularly those of Group 2, are strong bases but most are not very soluble in water)

**Dissociation of Water and  $K_w$** 

In any sample of pure water, a FEW molecules undergo the following reversible reaction:



The equilibrium expression for this reaction is:

$$K_{eq} = [H^+][OH^-]$$

This equilibrium expression must be satisfied in EVERY AQUEOUS SOLUTION INCLUDING ACIDS AND BASES. It such an important relationship that the equilibrium constant is given a special subscript “w”. so the equilibrium constant for the ionization of water is “ $K_w$ ” and  $K_w = 1.00 \times 10^{-14}$ . Therefore,

$$K_w = [H^+][OH^-] = 1.00 \times 10^{-14} \quad (\text{at } 25^\circ\text{C})$$

In pure water the concentrations of hydrogen ions and hydroxide ions are equal. Therefore, **in a neutral solution:**

$$[H^+] = [OH^-] = 1.00 \times 10^{-7} \text{ M}$$

When an acid or base is added to pure water, both the hydrogen ion and hydroxide ion are present. Both concentrations change but the PRODUCT OF THESE CONCENTRATIONS REMAINS A CONSTANT,  $1.00 \times 10^{-14}$ .

Acid added to water:  $[H^+] \uparrow$  and  $[OH^-] \downarrow$

Therefore, **in an acid**  $[H^+] > [OH^-]$

More specifically:  $K_w = [H^+][OH^-] = 1.00 \times 10^{-14}$

So,  $[H^+] > 1.00 \times 10^{-7} \text{ M}$

and  $[OH^-] < 1.00 \times 10^{-7} \text{ M}$

Base added to water:  $[OH^-] \uparrow$  and  $[H^+] \downarrow$

Therefore, **in a base**  $[H^+] < [OH^-]$

but  $K_w = [H^+][OH^-] = 1.00 \times 10^{-14}$

So,  $[OH^-] > 1.00 \times 10^{-7} \text{ M}$

and  $[H^+] < 1.00 \times 10^{-7} \text{ M}$

**pH and pOH**

The pH of several common substances is given at the right. You've no doubt heard the term pH many times ("pH balanced shampoo", etc.), but probably don't have a handle on just what pH means.

The pH is just a mathematical treatment of the  $[H^+]$  adopted to make it easier to express these small concentrations. The pH of a solution is the negative logarithm of the hydrogen ion concentration of that solution.

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

Similarly, the pOH of a solution is the negative log of the hydroxide ion concentration of that solution.

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

Don't worry if you haven't learned about logarithms yet. You will learn how to do the necessary calculations on your calculator.

A solution is:

**neutral if the pH = 7**

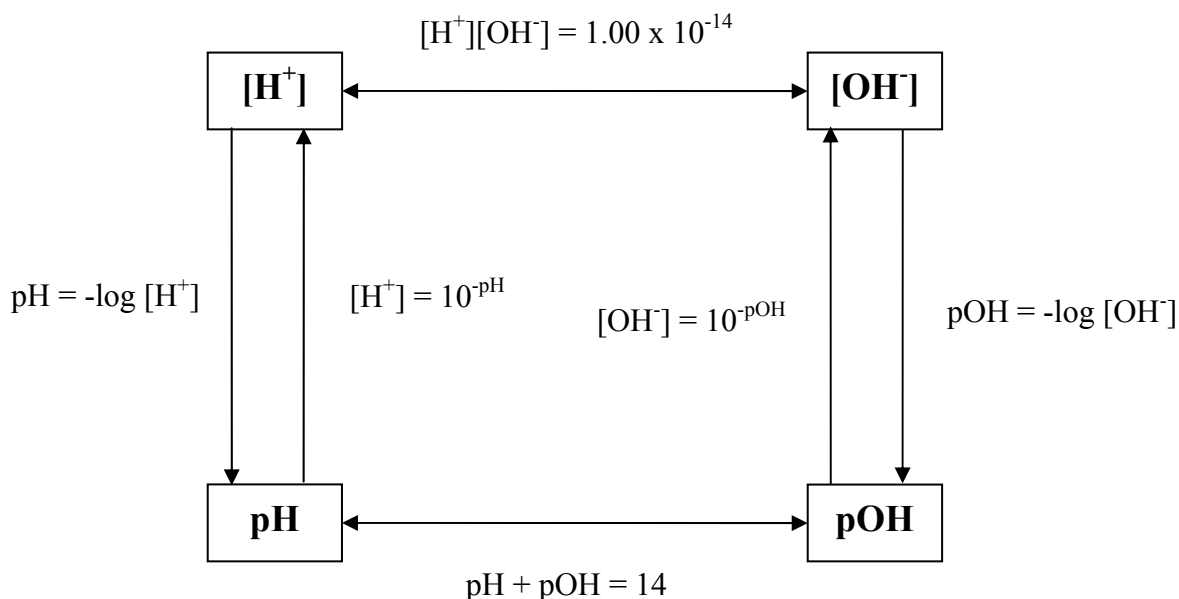
**acidic if the pH < 7**

**basic if the pH > 7**

Concentration of Hydrogen ions compared to distilled water		Examples of solutions at this pH
10,000,000	pH = 0	Battery acid, Strong Hydrofluoric Acid
1,000,000	pH = 1	Hydrochloric acid secreted by stomach lining
100,000	pH = 2	Lemon Juice, Gastric Acid Vinegar
10,000	pH = 3	Grapefruit, Orange Juice, Soda
1,000	pH = 4	Tomato Juice      Acid rain
100	pH = 5	Soft drinking water Black Coffee
10	pH = 6	Urine      Saliva
1	pH = 7	"Pure" water
1/10	pH = 8	Sea water
1/100	pH = 9	Baking soda
1/1,000	pH = 10	Great Salt Lake Milk of Magnesia
1/10,000	pH = 11	Ammonia solution
1/100,000	pH = 12	Soapy water
1/1,000,000	pH = 13	Bleaches Oven cleaner
1/10,000,000	pH = 14	Liquid drain cleaner

**Summary Diagram for  $[H^+]$ ,  $[OH^-]$ , pH, pOH Calculations – (a.k.a. “the Box”)**

The quantities  $[H^+]$ ,  $[OH^-]$ , pH and pOH are all related. If you know one, you can calculate the other three, but you can't always do it in one step. The diagram below shows you how to calculate one quantity from the others.



**Introductory Acid-Base Problems and Questions**

3. Oxalic acid ( $\text{H}_2\text{C}_2\text{O}_4$ ) is a diprotic acid. Write the two equations for its ionization(s) in water.

---



---

4.a. What is the concentration of hydrogen ions in a sample of pure water? \_\_\_\_\_

b. What is the concentration of hydroxide ions in a sample of pure water? \_\_\_\_\_

c. What is the pH of pure water? \_\_\_\_\_

5.a. What is the  $[\text{H}^+]$  in a 0.100M solution of HCl?

b. What is the  $[\text{OH}^-]$  in a 0.100M solution of HCl?

c. What is the pH of this solution?

6. What are the  $[\text{H}^+]$ ,  $[\text{OH}^-]$ , and pH of a 0.0100M NaOH solution?

a.  $[\text{H}^+] =$  \_\_\_\_\_      b.  $[\text{OH}^-] =$  \_\_\_\_\_      c. pH = \_\_\_\_\_

7. Complete the following table for these solutions:

pH	$[\text{H}^+]$	Solution is (acidic, basic, neutral)
7.00		
	$3.0 \times 10^{-12} \text{ M}$	
4.30		

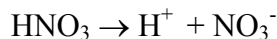
(Answers on p.49)

8. Calculate the pH of each of the solutions in parts a-c:
- $5.2 \times 10^{-2}$  M HNO<sub>3</sub>
  - 0.76 M KOH
  - a solution whose pOH is 9.40.
  - Calculate the hydrogen ion concentration of the solution in part c.
9. What is the hydroxide ion concentration of a HNO<sub>3</sub> solution whose pH is 1.20?
10. What mass of NaOH is needed to prepare 546 mL of a solution with a pH of 10.00?

(Answers on p.)

## Weak Acids

We have said before that strong acids ionize completely in water to produce hydrogen ions and anions. For example nitric acid when dissolved in water undergoes this reaction:

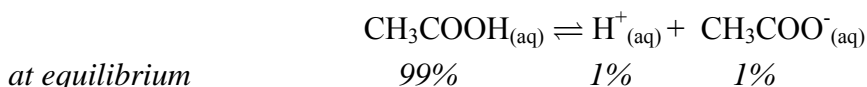


Weak acids, on the other hand, dissolve in water but only ionize to a small extent i.e. partially. (The terms ionize and dissociate are often used interchangeably, but be sure that you don't confuse the terms dissolve and dissociate.)

When added to water, weak acids form an **equilibrium** mixture. (note the  $\rightleftharpoons$  indicating reversibility) This system consists mostly of intact molecules, mixed with a few  $\text{H}^+$  ions and anions.

For example, when  $\text{CH}_3\text{COOH}$  (acetic acid) dissolves in water and when equilibrium has been established, only about 1% of the molecules dissociate to produce hydrogen ions and acetate ions. Most (99%) of the acetic acid molecules remain intact.

NOTE: The  $-\text{COOH}$  group at the end of an acid formula is characteristic of organic acids. The  $\text{H}^+$  is removed from the end of the formula when writing ionization equations.



Following the rules we've learned previously, we can write an equilibrium expression for this ionization:

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

The symbol  $K_a$  is used instead of  $K_{\text{eq}}$  to indicate that this is a weak acid equilibrium and is called the acid dissociation constant. In this case,  $K_a = 1.8 \times 10^{-5}$  at  $25^\circ\text{C}$ .

- The fact that  $K_a$  is such a small number implies that at equilibrium the concentration of reactants (in this case the weak acid) is much greater than the concentration of products.
- **The larger the  $K_a$ , the stronger the acid - the smaller  $K_a$ , the weaker the acid.** When comparing weak acid solutions of the same concentration, the acid with the larger  $K_a$  will ionize to greater extent and have the greater  $[\text{H}^+]$  at equilibrium.
- The coefficients will always be 1 in the ionization equation, so the exponents will always be 1 in the equilibrium expression. (Polyprotic acids have different  $K_a$  values for each successive ionization)
- Solving a problem involving weak acid equilibrium is much like solving any other equilibrium problem with one exception. If you need to, you may ASSUME THAT THE #MOLES OF ACID THAT IONIZED (per L) IS NEGLIGIBLE COMPARED TO THE INITIAL CONCENTRATION OF THE MOLECULAR FORM. (You will learn more about this assumption in class)

There are literally thousands of weak acids, most of them organic in nature (containing C, H, and O). There are many that you are probably familiar with, or should be familiar with (see the list on the next page). Every carbonated drink contains carbonic acid, aspirin is really acetylsalicylic acid, etc.

**Some Common Weak Acids**

A list of weak acids and their corresponding ionization constants is given below. These constants are a measure of the extent to which the acid dissociates in water solution.

<i>Name</i>	<i>Formula</i>	<i>K<sub>a</sub> at 25 °C</i>
<u>Phosphoric acid</u> the "tart" taste in cola drinks	H <sub>3</sub> PO <sub>4</sub>	7.3x10 <sup>-3</sup>
<u>Acetylsalicylic acid</u> aspirin	HC <sub>9</sub> H <sub>7</sub> O <sub>4</sub>	3.3x10 <sup>-4</sup>
<u>Formic acid</u> the stinging ingredient in the saliva of a red ant	HCHO <sub>2</sub>	1.8x10 <sup>-4</sup>
<u>Lactic acid</u> produced in muscles when they're overworked making them sore	H <sub>2</sub> C <sub>3</sub> H <sub>4</sub> O <sub>3</sub>	1.4x10 <sup>-4</sup>
<u>Ascorbic acid</u> Vitamin C	H <sub>2</sub> C <sub>6</sub> H <sub>4</sub> O <sub>6</sub>	7.8x10 <sup>-5</sup>
<u>Acetic acid</u> 3% of vinegar is this acid	CH <sub>3</sub> COOH	1.8x10 <sup>-5</sup>
<u>Carbonic acid</u> in all carbonated beverages and in acid rain (both beverages and raindrops contain dissolved CO <sub>2</sub> which forms carbonic acid)	H <sub>2</sub> CO <sub>3</sub>	4.2x10 <sup>-7</sup>
<u>Hydrogen sulfide</u> the "rotten egg" smell	H <sub>2</sub> S	1.0x10 <sup>-7</sup>

**Weak Acid Problems**

11.a. Write an equation that represents the dissociation of ascorbic acid in water.

---

b. Which is the stronger acid, ascorbic acid or lactic acid? Why?  
(Hint: See the table of weak acids above)

(Answers on p.50)

12. A solution is made by mixing 0.100 moles of  $\text{HCHO}_2$  (formic acid) in one liter of water. Calculate the concentrations of both ions and the formic acid molecules when the solution reaches equilibrium.  $K_a$  of formic acid is  $1.8 \times 10^{-4}$ .

13.a. Calculate the pH of a 0.500M nitrous acid ( $\text{HNO}_2$ ) solution.  $K_a$  of  $\text{HNO}_2$  is  $4.5 \times 10^{-4}$ .

b. What is the hydroxide ion concentration of this solution?

(Answers on p.50)

14. The pH of a 0.100M solution of a weak monoprotic acid, HA, is 2.85. What is the  $K_a$  of the acid?

(Answers on p.50)

**UNIT 10 ACID-BASE GRADED HOMEWORK ASSIGNMENT #8**

DO THIS ASSIGNMENT ON A SEPARATE SHEET OF PAPER

G1. Determine the following quantities:

- a. pH of 0.50 M HCl
- b. pH of 0.025 M KOH
- c. pOH of 0.00158 M HCl

G2 & G3. For each of the solutions:

- a. Find the hydrogen ion concentration.
- b. Find the hydroxide ion concentration.
- c. Determine if the solution is acidic or basic.

G2. Solution A has pH of 2.26

G3. Solution B has a pOH of 6.31

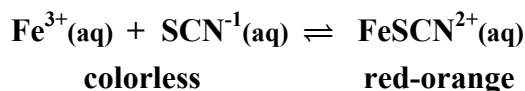
G4. Solution X has a pH of 5.0. Solution Y has a hydrogen ion concentration that is 1000 times greater than that of solution X. What is the pH of solution Y?

G5. A solution is prepared by dissolving 1.25g of NaOH in enough water to make 250.0mL. What is the pH of the prepared solution?

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BACKGROUND

In this experiment you will take a quantitative look at the reversible reaction:



The source of the  $\text{Fe}^{3+}$  ions will be a 0.00200M stock solution of  $\text{Fe}(\text{NO}_3)_3$  and the source of the  $\text{SCN}^{-}$  (thiocyanate) ions will be a 0.00200M stock solution of  $\text{KSCN}$ . You will prepare four different combinations of the reactant solutions and allow them to come to equilibrium. You will then determine the concentration of all three ions in each of these equilibrium solutions and calculate the equilibrium constant ( $K_{\text{eq}}$ ) for this reaction.

The determination of the concentration of  $\text{FeSCN}^{2+}$  will be done colorimetrically. For colored solutions containing this ion the amount of light that the solution absorbs is directly proportional to the concentration of the ion which produces the color (i.e. it follows Beer's Law). So when comparing a "solution x" that has is known absorbance but an unknown  $[\text{FeSCN}^{2+}]$  to a **standard** solution with a known absorbance and  $[\text{FeSCN}^{2+}]$ , the proportion:

$$\frac{[\text{FeSCN}^{2+}]_{\text{x}}}{\text{Abs}_{\text{x}}} = \frac{[\text{FeSCN}^{2+}]_{\text{std}}}{\text{Abs}_{\text{std}}}$$

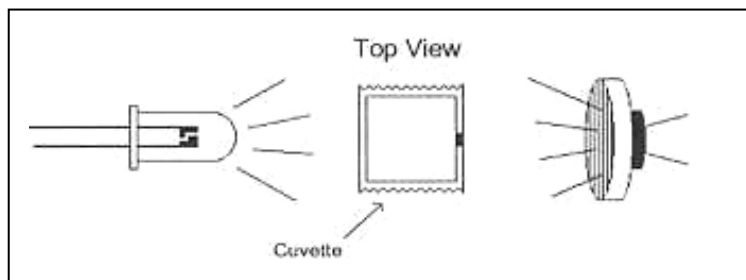
can be used to find the concentration of  $\text{FeSCN}^{2+}$  in "solution x".

The concentration of  $\text{FeSCN}^{2+}$  in our standard solution will be:  $[\text{FeSCN}^{2+}]_{\text{std}} = 2.00 \times 10^{-4} \text{ M}$

The method you will use to determine the concentration of  $\text{FeSCN}^{2+}$  ions in a solution relies on the fact that solutions containing  **$\text{FeSCN}^{2+}$  ions are red-orange**. Solutions that appear colored to the human eye absorb light in the visible range of the electromagnetic spectrum. Recall that white light is a combination of the colors red, orange, yellow, green, blue and violet. When white light passes through a red-orange solution, the yellow, green, blue and violet components of the light are absorbed and the remaining red and orange colors are transmitted. The result is that the solution appears red-orange to our eyes.

The absorbance of a solution can be measured using a **colorimeter**. The colorimeter is a device that measures the amount of light that is absorbed by a solution that is colored. As shown in the figure below, the colorimeter consists of three main parts, a source of monochromatic light (one color), a light detector, and a readout device.

The solution to be tested is poured into a **cuvette** that is inserted between the light source and the light detector. Some of the light is absorbed by the solution and some of it is transmitted through the solution. the fraction of light that is absorbed by the solution is detected, processed electronically, and displayed on a readout device as the **absorbance**. For this class, the control of the colorimeter and the readout of results will be done with the LabQuest system.



PRELAB ASSIGNMENT1) In your DATA section:

- Look in the background section and record the concentration of  $\text{FeSCN}^{2+}$  in the standard solution, and designate a place to record the absorbance of the standard solution.
- **COPY THIS TABLE OF "RECIPES" INTO YOUR DATA SECTION AND ADD A COLUMN** for recording the absorbance of each solution.

Solution #	0.00200M $\text{Fe}(\text{NO}_3)_3$	0.00200M KSCN	0.100M $\text{HNO}_3$
1	1.25 mL	0.50 mL	0.75 mL
2	1.25 mL	0.75 mL	0.50 mL
3	1.25 mL	1.00 mL	0.25 mL
4	1.25 mL	1.25 mL	0.00 mL

2) In your ANALYSIS section:

- **MAKE FOUR COPIES** THE "ICE" TABLE SHOWN BELOW AS PART OF YOUR **ANALYSIS SECTION**. Label one for each solution. The numbers get small so leave enough room to write values in scientific notation.

Solution # _____	$\text{Fe}^{3+}$	$\text{SCN}^-$	$\text{FeSCN}^{2+}$
initial [ ]			
$\Delta$ [ ]			
equilibrium [ ]			

- **IN EACH "ICE" TABLE**, fill in the initial concentration of  $\text{FeSCN}^{2+}$ , and calculate the initial concentrations of  $\text{Fe}^{3+}$  and  $\text{SCN}^-$ .

Some helpful reminders:

- Initial** means *after* the solutions have been combined but *before* a reaction occurs.
- The  $[\text{Fe}^{3+}]$  in 0.00200M  $\text{Fe}(\text{NO}_3)_3$  is also **0.00200M**
- The  $[\text{SCN}^-]$  in 0.00200M KSCN is **0.00200M**.
- The dilution formula  $M_c V_c = M_d V_d$  will be useful here for calculating the concentration of each ion after they have been diluted from their stock concentrations due to the addition of the other solutions.

## PROCEDURE

### Calibration of the Colorimeter

Carefully follow the instructions below for set up and calibration of the colorimeter. Two students will be sharing a colorimeter. Once the colorimeter has been calibrated, each student will prepare a solution in their own cuvette, take a reading and then mix the next solution.

### SETUP AND CALIBRATION OF COLORIMETER WITH LABQUEST

#### Setup and Calibration

1. Turn on the LabQuest.
2. Connect your colorimeter to Channel 1. The LabQuest will automatically identify the colorimeter sensor
3. From the discussion in the introduction, it should be obvious that the blue light is the best choice for the red-orange solutions you will be measuring. Use the navigation keys on the colorimeter to select the wavelength to be used for the experiment (**470 nm**).
4. Fill a cuvette  $\frac{3}{4}$  full with distilled water.
5. Clean the clear sides of the cuvette with a **KimWipe**.
6. Solutions are placed into the colorimeter by opening the lid, inserting the cuvette, and closing the lid. Put the cuvette in the colorimeter with the **clear sides perpendicular to the white mark** (direction of the light beam) and close the lid.
7. Push the calibration button "C" in the center of the navigation keys.
8. The absorbance being measured by the colorimeter will be displayed on the screen.
9. Take out the first cuvette and put in the cuvette that will be used by the other person. The absorbance value should be about 0.000 for each cuvette.

You have now set up the colorimeter so that the absorbance of the water and cuvette is being read as zero and you are ready to begin your experiment.

**DO NOT PUSH ANY MORE BUTTONS!** The colorimeter can now be used to measure the absorbance of the standard solution and the four equilibrium system solutions.

#### Determination of the Absorbance of the Standard Solution

10. Discard the water used for the calibration from the cuvette.
11. Rinse the cuvette with a little of the standard solution from one of the burettes which have been placed around the lab.
12. Fill the cuvette about  $\frac{3}{4}$  full with the standard solution and return to your station.
13. Clean the clear sides of the cuvette with a KimWipe and place it in the colorimeter with the clear sides facing the proper direction.
14. Close the lid and record the absorbance of the standard solution.

Determination of the Absorbance of the Four Equilibrium System Solutions

**When you are preparing your four solutions, be sure to use the proper pipette for each stock solution or you will contaminate the stock solution!**

**Solutions 1-4 will be prepared directly in the cuvette. Be sure to mix thoroughly by inverting several times after capping the cuvette.**

15. Discard the standard solution. Be careful to ensure that none of the standard solution remains in the cuvette.
16. For the first solution, use the Beral pipette labeled  $\text{Fe}(\text{NO}_3)_3$  to dispense the proper volume of  $\text{Fe}(\text{NO}_3)_3$  stock solution in the cuvette.
17. Use the  $\text{H}_2\text{O}$  Beral pipette to dispense the proper volume in the cuvette.
18. Use the  $\text{KSCN}$  Beral pipette to dispense the proper volume of  $\text{KSCN}$  stock solution into each cuvette.
19. Place cap on cuvette, mix by inverting cuvette and clean the clear sides with a KimWipe.
20. Place the cuvette properly in the colorimeter, close the lid and record the absorbance.
21. Discard the solution and shake the cuvette to ensure that none of the solution remains in the cuvette or on the cap before refilling it.

**22. Repeat steps 16-21 for equilibrium solutions 2-4.**

23. After the last measurement is made, **REMOVE THE CUVETTE FROM THE COLORIMETER** and turn the colorimeter off.
24. On the LabQuest screen touch **File**, then touch **Quit** and hold the power button down until the LabQuest shuts down.
25. Rinse the cuvette and cap with distilled water and invert the cuvette in the rack to dry.
26. Clean your counter top with a damp sponge

ANALYSIS AND CONCLUSION:

Show a complete sample calculation for one solution. The answers to for the other three solutions should be entered into the designated “ice” table.

1. Use the proportion on the first page with the concentration and absorbance of the standard solution and the absorbance of each of your prepared solution to calculate the equilibrium concentration of  $\text{FeSCN}^{2+}$ . (This will be a different value for each of your solutions.)

**THE EXAMPLE CALCULATION BELOW CAN BE USED AS A MODEL FOR ANALYSIS QUESTIONS 2 AND 3. (x = your answer to question 1)**

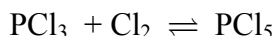
2. Use stoichiometry and subtraction to find the change in concentration of  $\text{Fe}^{3+}$  and  $\text{SCN}^-$  and the equilibrium concentrations of  $\text{Fe}^{3+}$  and  $\text{SCN}^-$ .
3. Use the equilibrium concentrations to calculate a value of  $K_{\text{eq}}$  for each of your solutions, and put them in a results table like the one shown below
4. This value should be characteristic of this reaction at room temperature. Average your results for  $K_{\text{eq}}$ . Use your average as an “accepted value” and calculate a percent difference for each  $K_{\text{eq}}$ . Describe how “constant” your  $K_{\text{eq}}$  turned out to be and what errors may have led to inconsistencies.

RESULTS

Solution #	$K_{\text{eq}}$	% difference
1		
2		
3		
4		
average		

Example Calculation:

$\text{PCl}_3$  and  $\text{Cl}_2$  react to form  $\text{PCl}_5$  by the equation:



The initial concentration of  $\text{PCl}_3$  is 0.298M and the initial concentration of  $\text{Cl}_2$  is  $2.00 \times 10^{-3}\text{M}$ . If the concentration of the  $\text{PCl}_5$  at equilibrium is  $1.79 \times 10^{-3}\text{M}$ , calculate the value of  $K_{\text{eq}}$ .

“ice” table

	$\text{PCl}_3$	$\text{Cl}_2$	$\text{PCl}_5$
initial [ ]	0.298M	$2.00 \times 10^{-3}\text{M}$	0
$\Delta$ [ ]	-x	-x	+x
equilibrium [ ]	$0.298-x$	$2.00 \times 10^{-3} -x$	x ( $=1.79 \times 10^{-3}\text{M}$ )

- The value of x represents the equilibrium  $[\text{PCl}_5]$ . If this has been experimentally determined to be  $1.79 \times 10^{-3}\text{M}$ , the change in concentration of  $\text{PCl}_5$  can be determined.
- The change in concentration of  $\text{PCl}_3$  and  $\text{Cl}_2$  can be determined by stoichiometry.
- The equilibrium concentrations of  $\text{PCl}_3$  and  $\text{Cl}_2$  can then be calculated by subtraction.
- Therefore, at equilibrium:

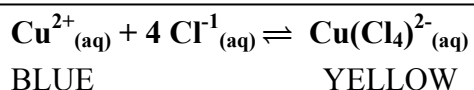
$$[\text{PCl}_3] = 0.296\text{M} \quad [\text{Cl}_2] = 2.1 \times 10^{-4}\text{M} \quad [\text{PCl}_5] = 1.79 \times 10^{-3}\text{M}$$

- Substituting into the equilibrium expression gives:  $K_{\text{eq}} = \frac{[\text{PCl}_5]}{[\text{PCl}_3][\text{Cl}_2]} = 28.8$

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NAME \_\_\_\_\_

**PART 1****DEMONSTRATION PROCEDURE**

1. Progressively add solid NaCl to 100 mL of a .03M solution of  $\text{Cu}(\text{NO}_3)_2$ . (7g + 3g + 5g)
2. Record your observations using the solutions in the display tubes as a color reference.

Solution # →	1	2	3	4
total mass NaCl	0 g	7 g	10 g (total)	15 g (total)
Color	Blue			

**ANALYSIS**

Look at the equation for the equilibrium system above to review how the color change indicates a shift of the system and subsequent changes in the concentrations of the reactants and product.

1. As the successive equilibrium systems 2 – 4 are prepared, what factor is changing in the equilibrium system in comparison to solution 1? (i.e. What is the stress on the system?)

- 2.a. According to Le Châtelier's Principle, in which direction should the system shift in order to counteract the change (i.e. relieve the stress)?

(circle one)      RIGHT                      LEFT

- b. Explain specifically how the shift you chose in 2a counteracts the change.

- 3.a. Is your answer to 2a consistent with the color change you observed?      YES              NO

- b. Write a brief explanation for your answer to 3a.

4. As a result of relieving the stress on the equilibrium, what is the effect on the concentration of the  $\text{Cu}^{2+}$  ion and  $\text{Cu}(\text{Cl}_4)^{2-}$  complex?

$[\text{Cu}^{2+}]$  \_\_\_\_\_                       $[\text{Cu}(\text{Cl}_4)^{2-}]$  \_\_\_\_\_

**PART 2****DEMONSTRATION PROCEDURE**

A solution of  $\text{AgNO}_3$  will be used. When  $\text{AgNO}_3$  dissolves it dissociates into  $\text{Ag}^+$  and  $\text{NO}_3^{-1}$  ions.

1. Add 10 mL of the equilibrium system solution to two test tubes.
2. Add 3 mL of the  $\text{AgNO}_3$  solution to the equilibrium system solution in one of the tubes.
3. Add 6 mL of the  $\text{AgNO}_3$  solution to the other test tube.
4. Tap the tube gently on the counter to help the precipitate settle.
5. Record the color of the solution in each test tube after the precipitate has settled.

NOTE: In making your observations, bear in mind that the white precipitate will affect the clarity which may affect the overall brightness of the solution. You are interested in the color change, not the change in brightness.

	Original	after 3 mL $\text{AgNO}_3$	after 6 mL $\text{AgNO}_3$
color	green		

**ANALYSIS**

The precipitate is silver chloride ( $\text{AgCl}$ )

5. What is the net ionic equation for the precipitation reaction?

6. What is the stress (disturbance) to the system at equilibrium? (HINT: What species from the original equilibrium system is consumed in the formation of this precipitate?)

7.a. According to Le Châtelier's Principle, in which direction should the system shift in order to counteract the change (i.e. relieve the stress)?

(circle one)      RIGHT                      LEFT

b. Explain specifically how the shift you chose in 7a counteracts the change.

8.a. Is your answer to 7a consistent with the color change you observed?      YES              NO

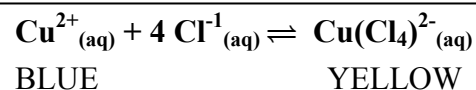
b. Write a brief explanation for your answer to 8a.

9. As a result of relieving the stress on the equilibrium, what is the effect on the concentration of the  $\text{Cu}^{2+}$  ion and  $\text{Cu}(\text{Cl}_4)^{2-}$  complex?

$[\text{Cu}^{2+}]$  \_\_\_\_\_                       $[\text{Cu}(\text{Cl}_4)^{2-}]$  \_\_\_\_\_

**PART 3****DEMONSTRATION PROCEDURE**

1. Add about 15 mL of the equilibrium system to each of two test tubes.
2. Put one in a hot water bath and another in an ice water bath and wait a few minutes for the solutions to come to thermal equilibrium with the bath.
3. Record your observations.



	room temp	Hot	Cold
color	green		

**ANALYSIS**

For the systems in the hot and cold water baths the change in conditions is obviously an increase and decrease in temperature respectively.

10. According to your observations, in which direction did the system shift in order to counteract the change (i.e. relieve the stress)?

Temperature increase	Temperature decrease

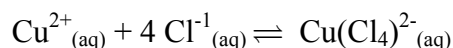
11. As a result of relieving the stress, what is the effect on the concentration of each of the reactants and the product?

	Temperature increase	Temperature decrease
$\text{Cu}^{2+}$		
$\text{Cl}^{-1}$		
$\text{Cu}(\text{Cl}_4)^{2-}$		

12. How does a change in temperature affect the value of  $K_{eq}$  for this system?

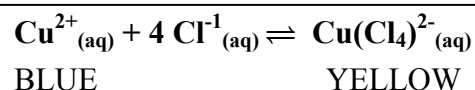
Temperature increase	Temperature decrease

- 13.a. Write the “energy” or “heat” term on the proper side of the equation below.



- b. Using **either** the increase **or** decrease in temperature, write a complete explanation for your answer to 13a. Your answer should be based on Le Châtelier's Principle.

**PART 4** (OPTIONAL FOR INSTRUCTOR)  
**DEMONSTRATION PROCEDURE**



1. Add about 10 mL of the equilibrium system solution to a test tube.
2. Add distilled water until there is a detectable color change and record your results.

	Original	after water is added
color	green	

**ANALYSIS**

**NOTE: ADDING WATER TO AN AQUEOUS EQUILIBRIUM SYSTEM INCREASES THE VOLUME OF THE ENTIRE SYSTEM AND IS ANALOGOUS TO DECREASING THE PRESSURE ON A SYSTEM OF GASES BY VOLUME INCREASE.**

When the system is diluted:

14. What is the stress (disturbance) to the system at equilibrium?

- 15.a. According to Le Châtelier's Principle, in which direction should the system shift in order to counteract the change (i.e. relieve the stress)?

(circle one)      RIGHT                      LEFT

- b. Explain specifically how the shift you chose in 15a counteracts the change.

- 16.a. Is your answer to 15a consistent with the color change you observed?    YES            NO

- b. Write a brief explanation for your answer to 16a.

17. As a result of relieving the stress, what is the effect on the number of moles of each of the reactants and the product?

moles  $\text{Cu}^{2+}$  \_\_\_\_\_      moles  $\text{Cl}^{-1}$  \_\_\_\_\_      moles  $\text{Cu}(\text{Cl}_4)^{2-}$  \_\_\_\_\_

18. After the system has come to equilibrium, how does the concentration of each of the reactants and the product compare to their concentration before the dilution?

$[\text{Cu}^{2+}]$  \_\_\_\_\_       $[\text{Cl}^{-1}]$  \_\_\_\_\_       $[\text{Cu}(\text{Cl}_4)^{2-}]$  \_\_\_\_\_

## LAB 10-3: ANTACIDS AND THE pH BALANCE IN YOUR STOMACH

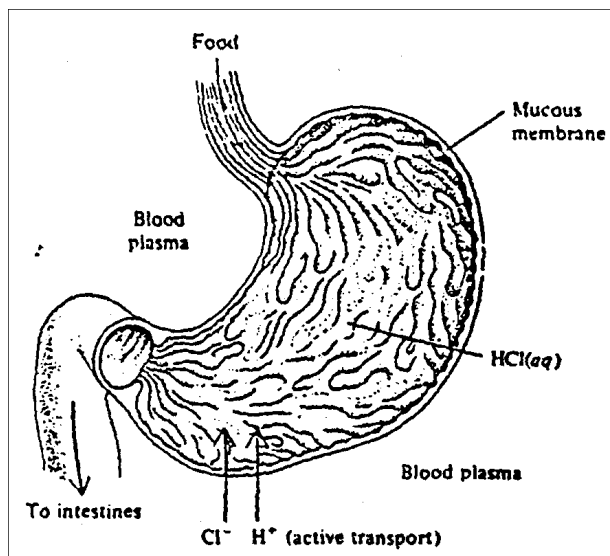
### BACKGROUND

An average adult produces between 2 and 3 liters of gastric juice daily. Gastric juice is a thin, acidic digestive fluid secreted by glands in the mucous membrane lining the stomach. It contains among other substances, hydrochloric acid. The pH of the gastric juice is about 1.5. This corresponds to a hydrogen ion concentration strong enough to react readily with zinc metal! Where do the  $H^+$  and  $Cl^-$  ions come from? What is the purpose of this highly acidic medium? What happens when there is an excess of  $H^+$  ions present in the stomach?

Both  $H^+$  and  $Cl^-$  ions move into the stomach from cells lining the stomach. Once in the stomach, most of these ions are prevented from diffusing back into the blood plasma by cell membranes.

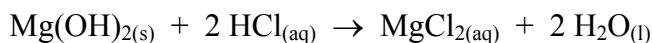
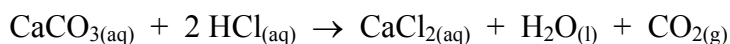
The purpose of the highly acidic medium within the stomach is to digest food and to activate certain digestive enzymes. Just the act of eating stimulates  $H^+$  secretion. However, if the acid content is excessively high, the constant influx of  $H^+$  ions can cause muscular discomfort, pain, swelling and bleeding.

One way to temporarily reduce the  $H^+$  ion concentration in the stomach is to take an antacid. The major function of antacids is to neutralize excess HCl in gastric juice. The table below lists the active ingredients of some popular antacids.



<b>Some Common Commercial Antacid Preparations</b>	
Alka-2	calcium carbonate
Alka-Seltzer	aspirin, sodium bicarbonate, citric acid
Buffered aspirin	aspirin, magnesium carbonate, aluminum hydroxide
Milk of magnesia	Magnesium hydroxide
Rolaids	Dihydroxy aluminum sodium carbonate
Tums	calcium carbonate

Three examples of how these antacids neutralize stomach acid are given in the reactions below:



One additional benefit of the acid/antacid reaction is that the  $CO_2$  produced by these reactions increases the gas pressure in the stomach, causing the person to belch and (hopefully) feel better.

The purpose of this exercise is to determine which brand of antacid is most effective in neutralizing stomach acid. First, you will use a titration technique to standardize the hydrochloric acid. Second, you will neutralize the hydrochloric acid with several different commercial antacids in order to calculate the effectiveness in **moles of HCl neutralized per gram of antacid**.

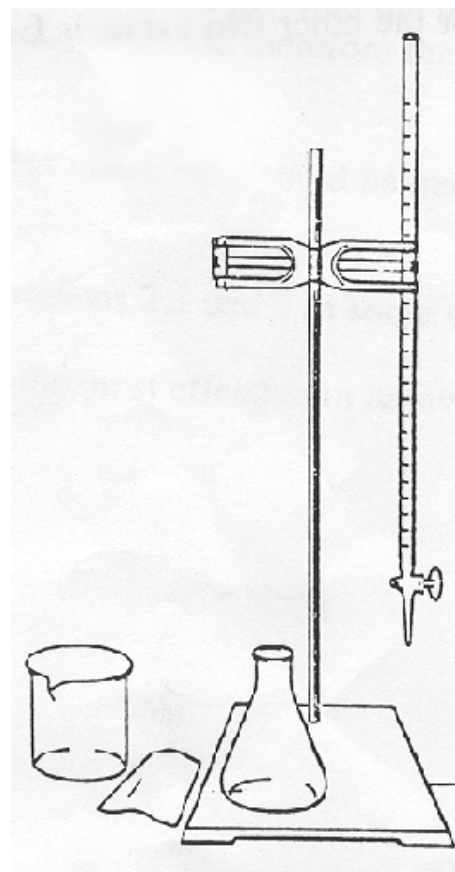
## Part 1: Standardizing the HCl solution.

### METHOD

Before carrying out the reactions between antacid and stomach acid (HCl), it is necessary to determine the actual concentration of HCl. You will do this by titrating it with a standardized NaOH solution, that is, a NaOH solution whose concentration is known to be **0.400 M**. Record this concentration.

### PROCEDURE

1. Pipette 10.0mL of the HCl solution into a clean 250 mL beaker. Record this volume.
2. Add 2 drops of the **methyl orange** indicator to the beaker and put in the magnetic stirring bar. This indicator should be pink in the presence of this excess acid.
- You have done a titration, but REMEMBER to close the valve on the buret before filling and avoid overfilling.
3. Fill the 100 mL beaker about half full of the standard NaOH solution.
4. Fill the buret with the standard NaOH solution
5. Drain some solution through the tip.
5. Record the initial reading on the buret.
6. Turn the magnetic stirrer to a medium setting and **carefully** titrate the HCl solution with the base. When you are near the endpoint there will be a yellow color which persists a bit longer after each drop. You must then add the base a drop at a time so that the **exact endpoint** (when the acid has been exactly neutralized) can be accurately determined.
7. Record the buret reading.
8. Discard the solution in the beaker and perform another trial of this same titration (HCl with NaOH) to confirm the results.



## Part 2: Titrating the Antacids.

### METHOD

Determining the quantity of acid that can be neutralized by an antacid tablet is not as straightforward as it might at first appear. The difficulty is that these tablets do not dissolve easily in water. If you simply added acid to them, they would dissolve slowly and the titration would take a very long time. The resolution of this problem is to dissolve the tablet in an excess of the acid which you have been using. The antacid will dissolve nicely in the excess and the solution is distinctly pink when you add the methyl orange indicator. You then titrate this solution with the standard base to determine how much acid was NOT neutralized by the tablet. The calculations will be outlined later.

THERE ARE FOUR DIFFERENT ANTACIDS TO BE TESTED. YOU SHOULD DO TWO OF THEM, AND YOUR PARTNER SHOULD DO THE OTHER TWO. YOU CAN THEN EXCHANGE DATA.

### PROCEDURE

9. Record the type of antacid that you are testing.
10. Place the tablet in a mortar and grind it with a pestle.
11. Take your mortar and a 250mL beaker to the balance.
12. Tare the beaker.
13. Mass between 0.50 and 1.00 grams of antacid into the beaker. **Do not use more than 1.00 gram of antacid.**
14. Record the exact mass.
15. Dissolve/neutralize the antacid in exactly 20.0 mL of your HCl solution. This will require **two** 10 mL portions from the pipet. Record this volume
16. Add 3 drops of the methyl orange indicator, and put in the magnetic stirring bar.
17. If necessary, add more NaOH to your buret and record the initial reading on the buret.
18. Titrate as before until you notice a **permanent** color change and record the final reading on the buret.
- BE PATIENT. Some of the active ingredients in the antacid are not very soluble.
19. Record the final reading on the buret.
20. Discard the solution and rinse the beaker thoroughly with tap and then distilled water.
21. **DO NOT RINSE THE MORTAR.** Wipe the mortar and pestle as clean as possible with a **DRY** paper towel.
22. Repeat the process with a second antacid tablet.

### CLEAN UP

23. Drain any remaining NaOH back into the **proper** stock bottle.
24. Wipe the mortar and pestle clean with a dry paper towel.
25. Wash the 250 mL beaker with a brush and then rinse it thoroughly; first with tap and then distilled water. Place it in the rack - inverted.
26. Rinse the 100 mL beaker; first with tap and then distilled water. Place it in the rack - inverted.
27. Rinse the buret and place it back into the clamp -inverted with the stopcock open.
28. Rinse the pipette and place it in the clamp - inverted.
29. Wipe down your counter top with a damp sponge.
- Remember to obtain the **data** (NOT CALCULATED RESULTS) for the other two antacids from your partner.

ANALYSIS AND CONCLUSION:***Part 1: Standardizing the HCl solution.***

1. Write a balanced molecular equation for the neutralization reaction between HCl and NaOH.
2. Calculate the volume of NaOH added during each trial and average the two volumes.
3. Use your balanced equation and knowledge of solution stoichiometry to calculate the concentration (molarity) of the HCl solution. Put this answer in your results table.

***Part 2: Titrating the Antacids.***

In questions 4-6 you will calculate the number of moles of acid neutralized by each of your antacids. Recall that you used an **excess of HCl**; therefore, only some of the acid was neutralized by the antacid and the remainder was neutralized by NaOH during the titration. The underlying principle for the calculation is that:

$$\underline{\text{total moles HCl added}} - \underline{\text{moles HCl neutralized by NaOH}} = \underline{\text{moles HCl neutralized by antacid}}$$

Using the mass of antacid, the moles of HCl neutralized by each antacid can then be reported on a “per gram” basis.

*Perform calculations 4-7 for each of the four antacids. Show work for one antacid.*

4. Calculate the total number of moles of HCl you initially added to the antacid sample.
5. Keeping in mind the balanced equation you wrote for question (1), calculate the moles of HCl neutralized by the NaOH (i.e. acid that was NOT neutralized by the antacid sample).
6. To find the number of moles of HCl neutralized by your antacid, subtract the moles neutralized by the NaOH from the total number of moles of HCl initially added to the antacid.
7. Finally, use the recorded mass of antacid in each experiment, to proportionately calculate the number of moles of acid which would be neutralized by 1.00g of each antacid. Put these answers in your results table.
8. Rank the four antacids from most effective to least effective in terms of moles of HCl neutralized per gram.

**UNIT 10 PRACTICE TEST**

**Part 1: Multiple Choice:** Choose the correct answer for each of the following questions.

1. What is the equilibrium expression for the following reaction:  $\text{H}_2\text{S}_{(g)} \rightleftharpoons \text{H}_{2(g)} + \text{S}_{(s)}$
- a.  $K_{\text{eq}} = \frac{[\text{H}_2][\text{S}]}{[\text{H}_2\text{S}]}$       b.  $K_{\text{eq}} = \frac{[\text{H}_2]}{[\text{H}_2\text{S}]}$       c.  $K_{\text{eq}} = \frac{[\text{H}_2\text{S}]}{[\text{H}_2]}$       d.  $K_{\text{eq}} = \frac{[\text{H}_2\text{S}]}{[\text{H}_2][\text{S}]}$

The **next four questions** refer to the following system at equilibrium:

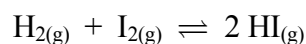


2. An increase in temperature will favor which reaction?  
a. forward reaction    b. reverse reaction    c. neither reaction
3. How will the number of moles of  $\text{UF}_4$  be changed if the pressure on the system is increased by decreasing volume?  
a. It will decrease    b. It will increase    c. It will remain the same
4. Which reaction will be favored by the addition of more  $\text{HF}_{(g)}$ ?  
a. forward reaction    b. reverse reaction    c. neither reaction
5. How will the concentration of  $\text{H}_2\text{O}_{(g)}$  be changed by the addition of some  $\text{UF}_{4(g)}$ ?  
a. It will decrease    b. It will increase    c. It will remain the same
6. The value for the equilibrium constant ( $K_{\text{eq}}$ ) for a system always changes when:  
a. the pressure is changed  
b. the concentration of the reactants is changed  
c. a catalyst is added  
d. the temperature is changed
7. For a given chemical reaction, if the value of  $K_{\text{eq}}$  is much greater than one, the reaction system at equilibrium consists of:  
a. mainly reactants    b. mainly products    c. equal amounts of reactants and products
8. When a catalyst is added to a system at equilibrium, the equilibrium concentrations of the reactants:  
a. decreases    b. increases    c. remains the same
9. A solution has a pH of 8. How is this solution best described?  
a. very basic    b. slightly basic    c. slightly acidic    d. very acidic
10. Which of the following solutions is the poorest conductor of electricity? (i.e. lowest [ions])  
a. 0.01M HCl    b. 0.01M  $\text{H}_2\text{SO}_4$     c. 0.01M  $\text{CH}_3\text{COOH}$     d. 0.01M  $\text{HNO}_3$
11. Which of the following properties describes a common acid? There may be more than one correct answer.  
a. turns litmus blue    b. tastes bitter    c. feels slippery    d. reacts with metals to produce  $\text{H}_2$  gas
12. When hydrochloric acid reacts neutralizes a base, the compounds formed are always:  
a. a hydroxide and water    c. a sodium compound and a chloride  
b. a chloride and water    d. a chloride and a hydroxide

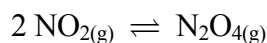
13. A hydronium ion is a water molecule:
- a. minus the oxygen
  - b. plus a neutron
  - c. plus a proton
  - d. minus an electron
14. When an acid is added to pure water:
- a. the  $[H^+]$  increases and the  $[OH^-]$  decreases
  - b. the  $[H^+]$  increases and the  $[OH^-]$  stays the same
  - c. both the  $[H^+]$  and the  $[OH^-]$  increase
  - d. both the  $[H^+]$  and the  $[OH^-]$  decrease

**Part 2: Problems and Questions**

15. The equilibrium constant at  $448^\circ\text{C}$  for the reaction below is 50.2. When a 2.00L flask at  $448^\circ\text{C}$  containing these three species at equilibrium was analyzed, it was found to contain 1.20 moles of  $\text{H}_2$  and 0.400 moles of  $\text{I}_2$ . What was the concentration of HI in this flask?



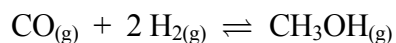
16. At  $25^\circ\text{C}$ ,  $K_{\text{eq}}$  is 8.80 for the equilibrium system:



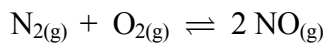
Analysis of an experimental system shows that  $2.00 \times 10^{-4}$  moles of  $\text{NO}_2$  are present in a 1.00L flask along with  $1.50 \times 10^{-3}$  moles of  $\text{N}_2\text{O}_4$ .

- a. Is the system at equilibrium? \_\_\_\_\_
- b. If it is not at equilibrium, which reaction (forward or reverse) must occur for the system to come to equilibrium? Explain carefully.

17. When 1.360 moles of  $\text{H}_2$  and 0.780 moles  $\text{CO}$  are put in a 1.00L sealed container at  $160^\circ\text{C}$ , the equilibrium system below is established. Measurements indicate that the equilibrium concentration of  $\text{H}_2$  is 0.120 M. Calculate  $[\text{CO}]$  and  $[\text{CH}_3\text{OH}]$  at equilibrium and  $K_{\text{eq}}$ .



18. At  $950^\circ\text{C}$ , elemental nitrogen and oxygen react with each other to form nitrogen monoxide according to the equation below. The  $K_{\text{eq}} = 0.690$



Determine what the equilibrium concentrations of all three gases will be if the reaction is initiated with 0.600 M  $\text{NO}_{(\text{g})}$ .

$$[\text{N}_2] = \underline{\hspace{2cm}} \quad [\text{O}_2] = \underline{\hspace{2cm}} \quad [\text{NO}] = \underline{\hspace{2cm}}$$

19. Write balanced equations for the following:

a. Sulfuric acid is neutralized by titration with potassium hydroxide.

---

b. Write equations for the three ionizations of  $\text{H}_3\text{PO}_4$  (phosphoric acid) in water.

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20. A student added 40.0mL of 0.287M  $\text{HNO}_3$  and three drops of phenolphthalein to a flask and titrated it with an  $\text{Ba}(\text{OH})_2$  solution. The solution turned pink when 10.7mL of the base had been added.

a. Write the balanced molecular equation for the reaction between  $\text{HNO}_3$  and  $\text{Ba}(\text{OH})_2$ .

---

b. Calculate the concentration of the  $\text{Ba}(\text{OH})_2$  solution.

21. Determine the pH of the following solutions:

a. 0.1M HCl      pH = \_\_\_\_\_

b. 0.0500 M HCl    pH = \_\_\_\_\_

c. 0.01M NaOH    pH = \_\_\_\_\_

d. Complete this statement:

*When the pH of a solution changes by a factor of "1", the  $[\text{H}^+]$  changes by a factor of \_\_\_\_\_.*

22. Calculate the pH of a solution that is prepared by mixing 0.122 grams of KOH in 555mL of water.
23. Calculate the pH of a 0.200M solution of hydrofluoric acid ( $K_a$  of HF is  $6.60 \times 10^{-4}$ )

24. The pH of a 0.200M solution of a weak monoprotic acid (HA) is 3.100. Calculate  $K_a$  for this acid.

**Part 3: Essays**

*There may be one or more short essay questions on the test. The question(s) will be taken from the following topics unless you are notified otherwise.*

**E1. Le Châtelier's Principle.**

- What happens to an equilibrium system when substance is added or removed or the temperature or pressure is changed?
- Why does  $K_{eq}$  vary with temperature?
- Why does the variation of  $K_{eq}$  with temperature depend upon whether a reaction is exothermic or endothermic?

**Unit 10 Practice Test Answers**

- |      |       |
|------|-------|
| 1. b | 8. c  |
| 2. a | 9. b  |
| 3. b | 10. c |
| 4. a | 11. d |
| 5. a | 12. b |
| 6. d | 13. c |
| 7. b | 14. a |

15.  $[\text{HI}] = 2.45\text{M}$

16.  $Q = 37500$  which is  $\neq K_{\text{eq}}$  therefore, *NO* the system is not at equilibrium.

The fact that  $Q > K_{\text{eq}}$  indicates that the current value for  $[\text{N}_2\text{O}_4]$  is too high and the current value for  $[\text{NO}_2]$  is too low. Therefore the reverse reaction must occur for the system to come to equilibrium.

17.  $[\text{CO}] = .160\text{M}$

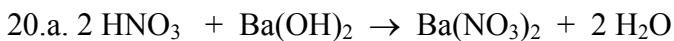
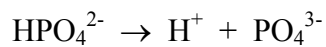
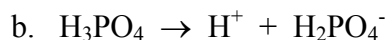
$[\text{CH}_3\text{OH}] = 0.620\text{M}$

$K_{\text{eq}} = 269$

18.  $[\text{N}_2] = 0.212\text{ M}$  .

$[\text{O}_2] = 0.212\text{ M}$

$[\text{NO}] = 0.176\text{ M}$ .



b.  $[\text{Ba}(\text{OH})_2] = .536\text{M}$

21.a.  $\text{pH} = 1.0$

b.  $\text{pH} = 1.301$

c.  $\text{pH} = 12.0$

d. ...factor of 10.

22.  $\text{pH} = 11.593$

23.  $\text{pH} = 1.94$

24.  $K_a = 3.15 \times 10^{-6}$

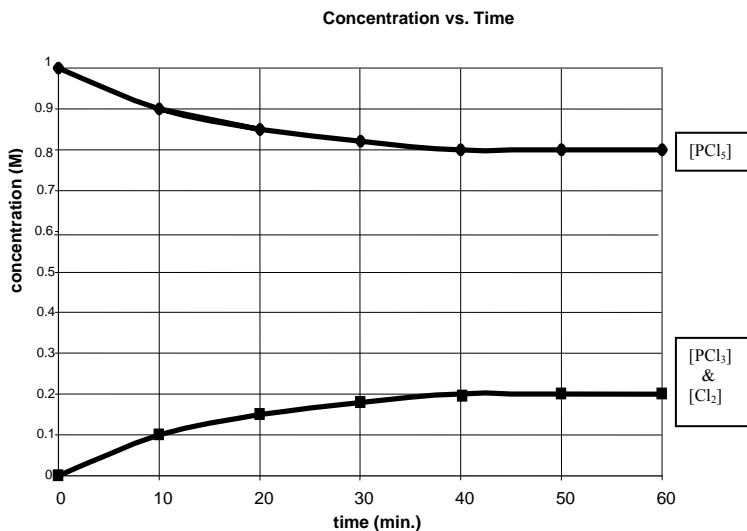
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## Answers to Unit 10 Homework Assignments

### Part 1: Chemical Equilibrium

#### Assignment 2

p.3-4 problems. 1-3



1.
  - a. 40 min.
  - b. [PCl<sub>5</sub>] = .80 M  
[PCl<sub>3</sub>] and [Cl<sub>2</sub>] = .20 M
  - c. Both substances have an initial concentration of 0, and their coefficient (or mole) ratio is 1:1 in the equation. Therefore, they will be produced in equal mole amounts.
  - d. [PCl<sub>5</sub>] = .80 M  
[PCl<sub>3</sub>] and [Cl<sub>2</sub>] = .20 M

$$2.a. K_{eq} = \frac{[FeSCN^{+2}]}{[Fe^{+3}][SCN^{-}]} \quad b. K_{eq} = \frac{[SO_2]^2[O_2]}{[SO_3]^2} \quad c. K_{eq} = \frac{[CO_2][H_2]}{[CO]}$$

- 3.a. 12.5
- b. .125 M

#### Assignment 3

p.4-6 problems 4-8

4. No,  $Q = 25$ . Therefore, the system must react to the right to achieve equilibrium.

5.  $K_{eq} = 0.0185$

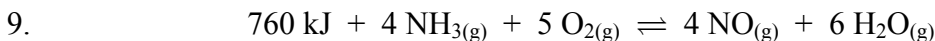
6.  $K_{eq} = 55.8$

7.  $[H_2] = 0.24 M$      $[F_2] = 0.24 M$      $[HF] = 2.52 M$

8.  $[H_2] = [CO_2] = 0.072 M$      $[H_2O] = [CO_2] = 0.150 M$

**Assignment 4**

p.7-8 problems 9-15



STRESS	SYSTEM SHIFTS	EQ. CONCENTRATIONS				K <sub>eq</sub>
		[NH <sub>3</sub> ]	[O <sub>2</sub> ]	[NO]	[H <sub>2</sub> O]	
NH <sub>3</sub> addition	→	↑	↓	↑	↑	=
O <sub>2</sub> removal	←	↑	↓	↓	↓	=
NO addition	←	↑	↑	↑	↓	=
H <sub>2</sub> O removal	→	↓	↓	↑	↓	=
increase temperature	→	↓	↓	↑	↑	↑
decrease temperature	←	↑	↑	↓	↓	↓
addition of catalyst	none	=	=	=	=	=
		#MOLES PRESENT AT EQUILIBRIUM				
		NH <sub>3</sub>	O <sub>2</sub>	NO	H <sub>2</sub> O	
increase pressure (Vol. dec.)*	←	↑	↑	↓	↓	=
decrease pressure (Vol. inc.)**	→	↓	↓	↑	↑	=

\*all equilibrium [ ]'s increase because of volume decrease

\*\*all equilibrium [ ]'s decrease because of volume increase

10. C

11. C

12. K<sub>eq</sub> increases

13. There would be no change in the number of moles of each gas in the system.

(All concentrations would increase if pressure increase was caused by a decrease in volume.)

14. The system would achieve equilibrium sooner (flat portion of each curve would start at an earlier time). However, the equilibrium concentration of each substance would be the same as it was without the catalyst (flat portion of each curve would occur at the same concentration).

15.a. [H<sub>2</sub>] increasesb. [H<sub>2</sub>] does not change (Addition or removal of solid substance does not cause system to shift.)c. [H<sub>2</sub>] decreasesd. [H<sub>2</sub>] decreases (System only *partially* counteracts change caused by addition or removal.)e. increase the [H<sub>2</sub>] and increase K<sub>eq</sub>.f. #moles H<sub>2</sub> decreases (but [H<sub>2</sub>] increases)

## Answers to Unit 8 Homework Assignments

### Part 2: Acid-Base Chemistry

#### Assignment 6

p.12 problems 1-2

- 1.a.  $\text{KOH} + \text{HNO}_3 \rightarrow \text{H}_2\text{O} + \text{KNO}_3$
- b.  $2 \text{H}_3\text{PO}_4 + 3 \text{Ba}(\text{OH})_2 \rightarrow 6 \text{H}_2\text{O} + \text{Ba}_3(\text{PO}_4)_2$
- c.  $2 \text{NaOH} + \text{H}_2\text{CO}_3 \rightarrow 2 \text{H}_2\text{O} + \text{Na}_2\text{CO}_3$
- 2.a.  $2 \text{HNO}_3 + \text{Ba}(\text{OH})_2 \rightarrow 2 \text{H}_2\text{O} + \text{Ba}(\text{NO}_3)_2$
- b. 0.369 M

#### Assignment 7

p.15-16 problems 3-10

3.  $\text{H}_2\text{C}_2\text{O}_4 \rightarrow \text{H}^+ + \text{HC}_2\text{O}_4^-$   
 $\text{HC}_2\text{O}_4^- \rightarrow \text{H}^+ + \text{C}_2\text{O}_4^{2-}$
- 4.a.  $[\text{H}^+] = 1.00 \times 10^{-7} \text{M}$
- b.  $[\text{OH}^-] = 1.00 \times 10^{-7} \text{M}$
- c. pH = 7
- 5.a.  $[\text{H}^+] = 0.100 \text{M}$
- b.  $[\text{OH}^-] = 1.00 \times 10^{-13} \text{M}$
- c. pH = 1
- 6.a.  $[\text{H}^+] = 1.00 \times 10^{-12} \text{M}$
- b.  $[\text{OH}^-] = 1.00 \times 10^{-2} \text{M}$
- c. pH = 12

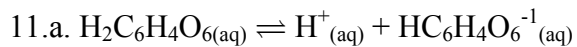
7.

pH	$[\text{H}^+]$	Solution is (acidic, basic, neutral)
(7)	$1.00 \times 10^{-7} \text{M}$	neutral
11.5	$(3.00 \times 10^{-12} \text{M})$	basic
(4.3)	$5.0 \times 10^{-5} \text{M}$	acidic

- 8.a. 1.28
- b. 13.88
- c. 4.60
- d.  $2.5 \times 10^{-5} \text{M}$
9.  $[\text{OH}^-] = 1.6 \times 10^{-13} \text{M}$
10.  $2.18 \times 10^{-3} \text{g NaOH}$

**Assignment 9**

p.18-20 Problems 11-14

b. Lactic acid is a stronger weak acid because it has a higher  $K_a$  value.

12.  $[\text{H}^+] = 4.12 \times 10^{-3}\text{M}$        $[\text{CHO}_2^{-1}] = 4.12 \times 10^{-3}\text{M}$        $[\text{HCHO}_2] = 0.096 \text{ M}$

13. a.  $\text{pH} = 1.82$

b.  $6.6 \times 10^{-13}\text{M}$

14.  $K_a = 2.00 \times 10^{-5}$  (if  $[\text{H}^+]$  is considered negligible to initial  $[\text{HA}]$ )

$K_a = 2.02 \times 10^{-5}$  (if  $[\text{H}^+]$  is not considered negligible to initial  $[\text{HA}]$  and is subtracted from initial  $[\text{HA}]$  )