

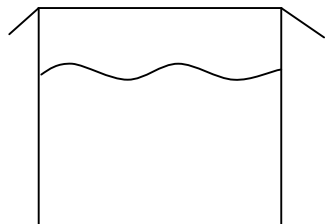
CHEMISTRY 111 LECTURE  
EXAM IV Material

PART 1 CHEMICAL EQUILIBRIUM

Chapter 14

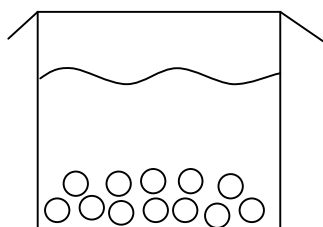
I Dynamic Equilibrium

I. In a closed system a liquid obtains a dynamic equilibrium with its vapor state



*Dynamic equilibrium:* rate of evaporation = rate of condensation

II. In a closed system a solid obtains a dynamic equilibrium with its dissolved state



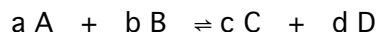
*Dynamic equilibrium:* rate of dissolving = rate of crystallization

II Chemical Equilibrium

I. EQUILIBRIUM

A. BACKGROUND

Consider the following reversible reaction:



1. The forward reaction (→) and reverse (←) reactions are occurring simultaneously.
2. The rate for the forward reaction is **equal** to the rate of the reverse reaction and a dynamic equilibrium is achieved.
3. The ratio of the concentrations of the products to reactants is constant.

B. THE EQUILIBRIUM CONSTANT - Types of K's

Solutions	$K_c$
Gases	$K_c$ & $K_p$
Acids	$K_a$
Bases	$K_b$
Solubility	$K_{sp}$
Ionization of water	$K_w$
Hydrolysis	$K_h$
Complex ions	$\beta_n$

General

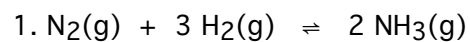
Keq

### C. EQUILIBRIUM CONSTANT

For the reaction,  $aA + bB \rightleftharpoons cC + dD$  The equilibrium constant,  $K_c$ , has the form:

$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

### D. WRITING K's



### E. MEANING OF K

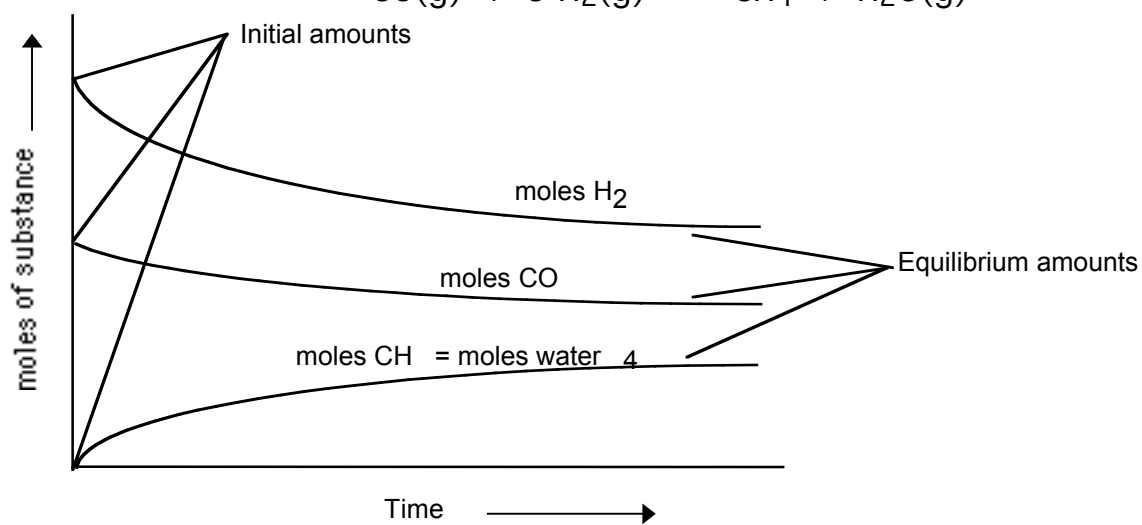
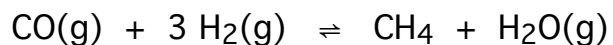
1. If  $K > 1$ , equilibrium favors the products

2. If  $K < 1$ , equilibrium favors the reactants

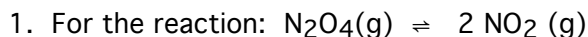
3. If  $K = 1$ , neither is favored

### F. ACHIEVEMENT OF EQUILIBRIUM

Chemical equilibrium is established when the rates of the forward and reverse reactions are equal.



### G. OBTAINING EQUILIBRIUM CONSTANTS FOR REACTIONS



the concentrations of the substances present in an equilibrium mixture at 25°C are:

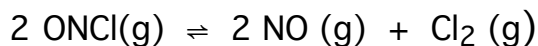
$$[\text{N}_2\text{O}_4] = 4.27 \times 10^{-2} \text{ mol/L}$$

$$[\text{NO}_2] = 1.41 \times 10^{-2} \text{ mol/L}$$

What is the value of  $K_C$  for this temperature?



1.00 mol of ONCl (g) at 500°C is introduced into a one-liter container. At equilibrium, the ONCl is 9.0% dissociated. What is the value of  $K_C$  at 500°C?



*Initial*

*Change*

*at Equilibrium*

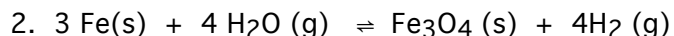
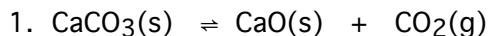
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### H. HETEROGENEOUS EQUILIBRIA

When equilibrium between substances involve two or more phases it is called **Heterogeneous Equilibria**. The concentration of a pure solid or a pure liquid in their standard states is constant (at constant  $T^\circ$  and  $P$ ).

Therefore, the concentrations of solids or liquids involved in a heterogeneous equilibrium are included in the value of  $K_C$  and do not appear in the equilibrium constant expression.

Write the  $K_C$  expression for:



**Solids** do not appear in the equilibrium constant.

**Pure liquids** do not appear in the equilibrium constant.

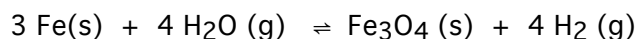
$\text{H}_2\text{O}(\text{l})$  for example.

## I. THE EQUILIBRIUM CONSTANT, $K_p$ FOR GASES

Amounts of gases are often measure by pressure. The partial pressure of a gas is a measure if its concentration. The equilibrium constant,  $K_p$ , is written in terms of the partial pressures of gases.



The  $K_p$  expression is:



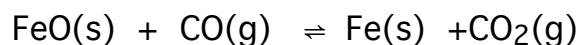
The  $K_p$  expression is:

Example : For the reaction:  $\text{FeO}(\text{s}) + \text{CO}(\text{g}) \rightleftharpoons \text{Fe}(\text{s}) + \text{CO}_2(\text{g})$

at 1000. °C,  $K_p$  is 0.403. If  $\text{CO}(\text{g})$ , at a pressure of 1.000 atm, and excess  $\text{FeO}(\text{s})$  are placed

in

a container at 1000°C, what are the pressures of  $\text{CO}(\text{g})$  and  $\text{CO}_2(\text{g})$  when equilibrium is attained?



*Initial*

*Change*

*at Equilibrium*

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## J. THE RELATIONSHIP BETWEEN $K_C$ AND $K_p$

$$K_p = K_C (RT)^{\Delta n}$$

Where: **R** is the gas constant

**T** is the temperature of the system

**$\Delta n$**  is the change in the number of moles of gas

$\{ \sum \text{moles of products gases} - \sum \text{moles of reactants gases} \}$

Example 1: For the reaction:  $2 \text{SO}_3(\text{g}) \rightleftharpoons 2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g})$

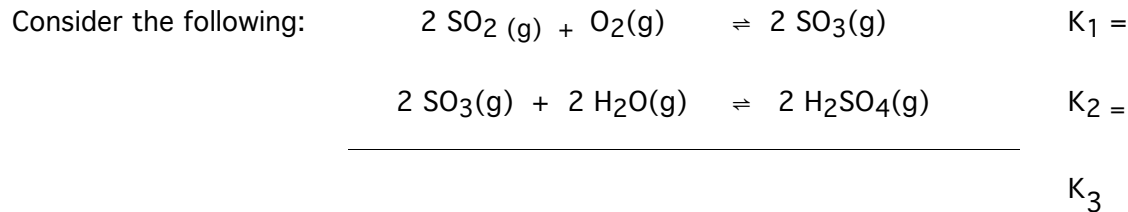
at 1100. K,  $K_C$  is 0.0271 mol/L. What is  $K_p$  at this temperature?

Example 2: For the reaction:  $\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightleftharpoons 2 \text{NH}_3(\text{g})$

at 500. °C,  $K_p$  is  $1.50 \times 10^{-5}/\text{atm}^2$ . What is  $K_C$  at this temperature?

### K. THE EQUILIBRIUM CONSTANT FOR THE SUM OF REACTIONS

When a new chemical equation is obtained by summing individual equations, the equilibrium constant,  $K_{\text{Overall}}$ , is the product of the individual equations.



$$K_3 = K_{\text{Overall}} =$$

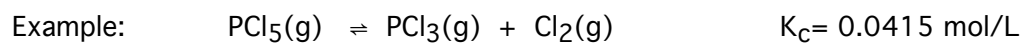
### L. PREDICTING THE DIRECTION OF REACTION

Given the initial concentrations of reactants and products, how would you predict which direction the reaction will "go" to obtain equilibrium? Left ? or Right?

1. Calculate Reaction Quotient,  $Q_C$

$$Q_C = \frac{[\text{C}]_{\text{ini}}^c [\text{D}]_{\text{ini}}^d}{[\text{A}]_{\text{ini}}^a [\text{B}]_{\text{ini}}^b}$$

2. Compare  $Q_C$  with  $K_C$
3. If  $Q_C > K_C$ , the reaction will go to the left to obtain equilibrium.  
If  $Q_C < K_C$ , the reaction will go to the right to obtain equilibrium.  
If  $Q_C = K_C$ , the reaction is at equilibrium.

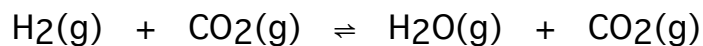


Initially, 0.100 mol of  $\text{PCl}_5(\text{g})$ , 0.0500 mol of  $\text{PCl}_3(\text{g})$ , and 0.0300 mol of  $\text{Cl}_2(\text{g})$  is placed in a 1.00 L container. Is this system at equilibrium? If not, which way will the reaction go to obtain equilibrium?

### M. CALCULATING EQUILIBRIUM CONCENTRATIONS

For the Reaction:  $\text{H}_2(\text{g}) + \text{CO}_2(\text{g}) \rightleftharpoons \text{H}_2\text{O}(\text{g}) + \text{CO}(\text{g})$   $K_C = 0.771$  at  $75^\circ\text{C}$

If 0.0100 mol of  $\text{H}_2$  and 0.0100 mol of  $\text{CO}_2$  are mixed in a one-liter container at  $750^\circ\text{C}$ , what are the concentrations of all substances present at equilibrium?



*Initial*

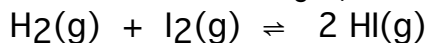
*Change*

*at Equilibrium*

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### N. LE CHATELIER'S PRINCIPLE

Le Chatelier's principle states: Any change to a system in equilibrium causes a shift in the system to reduce the effect of the change (until equilibrium is again reached)



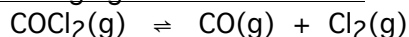
If more  $\text{H}_2$  were added to a mixture at equilibrium, the system is no longer at equilibrium.

$Q_C < K_C$ . The net reaction would be to the right to restore equilibrium.

If more  $\text{HI}$  were added to a mixture at equilibrium, the system is no longer at equilibrium.

$Q_C > K_C$ . The net reaction would be to the left to restore equilibrium.

#### 1. Changing the Concentration



- Adding more  $\text{CO}$
- Removing some  $\text{Cl}_2$
- Adding more  $\text{COCl}_2$
- How will any of the above effect the numerical value of  $K_C$ ?

## 2. Changing the Pressure

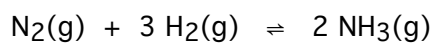
a.  $\Delta n > 0$



(1) Increase Pressure

(2) Decrease Pressure

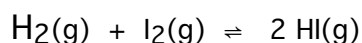
b.  $\Delta n < 0$



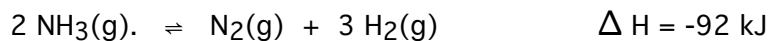
(1) Increase Pressure

(2) Decrease Pressure

c.  $\Delta n = 0$



## 3. Changing the Temperature



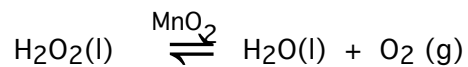
a. Increase Temperature

b. Decrease Temperature

## 4. Adding a catalyst

A catalyst increases the rate of reaction and is not consumed in a reaction:

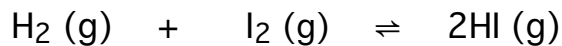
The catalyst increases the rate by lowering the energy of activation,  $E_a$



Adding a catalyst to a system has no effect on the equilibrium.



1. A mixture of 0.500 mole of hydrogen and 0.500 mole of iodine gas was placed in a one-liter flask at 430°C. Calculate the concentration of the hydrogen gas, iodine gas and hydrogen iodide gas at equilibrium. The Kc for the reaction is 54.3 at 430 °C



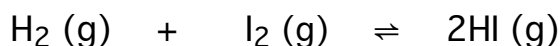
*Initial*

*Change*

*at Equilibrium*

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2. Starting with 4.20 moles of HI (g) in a 9.60 liter reaction vessel, calculate the concentration of the hydrogen gas, iodine gas and hydrogen iodide gas at equilibrium. at 430°C. The Kc for the reaction is 54.3 at 430 °C



*Initial*

*Change*

*at Equilibrium*

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3. For the reaction:  $\text{C} (\text{s}) + \text{CO}_2 (\text{g}) \rightleftharpoons 2 \text{CO}(\text{g})$

Kp = 167.5 atm at 1000°C. What is the partial pressure of CO(g) in an equilibrium system in which the partial pressure of carbon dioxide gas is 0.100 atm?

## CHAPTER 15

### I. ACIDS AND BASES

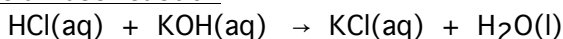
#### A. Acidic Characteristics

1. Tart/Sour taste
2. Produces color changes with indicators
3. Will react with and neutralize a base to form water
4. Will react with certain metals with  $H_2$  as a product

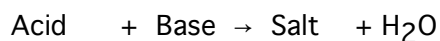
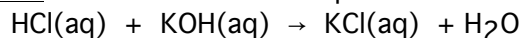
#### B. Basic Characteristics → Ionic Compounds that contains $OH^-$

1. Bitter taste
2. Slippery feeling
3. Produces color changes with indicators
4. Will neutralize an acid to form water
5. Will form a precipitate (ppt) with certain cations

#### D. Acid-Base reaction

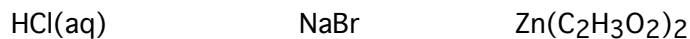


#### E. Salts- A salt is an ionic compound which is produced in an acid-base reaction.



How to recognize:

A salt is an ionic compound that does not contain  $OH^-$  and is not a metal oxide



### F. STRENGTH OF ACIDS AND BASES

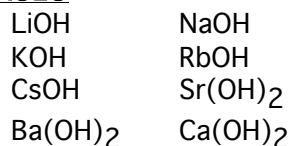
Strong Acids - Those acids which are 100% ionized (reacted with water)

Weak Acids - Those acids which are less than 100% ionized (reacted with water)

Strong Bases - Those bases which are 100% dissociated in water

Weak Bases - Those bases which are less than 100% dissociated in water

#### STRONG BASES



#### STRONG ACIDS

HNO <sub>3</sub>	H <sub>2</sub> SO <sub>4</sub>
HClO <sub>4</sub>	HCl
HBr	HI

## Part 4: Concepts of Acids and Bases

### I. ARRHENIUS CONCEPT

Arrhenius Acid  $\Rightarrow$  A substance that INCREASES the concentration of H<sup>+</sup> in water  
ex.

Arrhenius base  $\Rightarrow$  A substance that INCREASES the concentration of OH<sup>-</sup> in water  
ex.

### II. BRONSTED-LOWRY CONCEPT

#### A. DEFINITIONS

Bronsted Acid  $\Rightarrow$  A substance that can donate a proton (H<sup>+</sup>)

Monoprotic

Polyprotic

Bronsted Base  $\Rightarrow$  A substance that can accept a proton (H<sup>+</sup>)

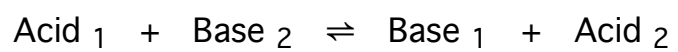
An Acid-Base Reaction:

#### B. CONJUGATED ACID-BASE PAIRS

When a Bronsted-Lowry acid loses a proton, the species formed is a base:

The two species differ by the loss or gain of a proton. These pairs are called: **conjugated acid-base pairs.**

1. Identifying Bronsted Acid and Base Species



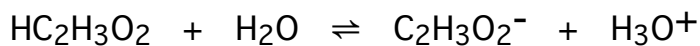
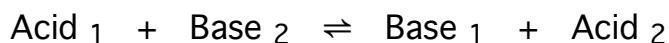
2. Bronsted Amphiprotic substances

Substances which can act as a Bronsted acid or a Bronsted base

### 3. Strengths of Bronsted Acids and Bases

The strength of an **acid** is its tendency to **donate a proton**

The strength of a **base** is its tendency to **accept a proton**



RELATIVE STRENGTHS OF ACIDS AND BASES			
	ACID	BASE	
STRONGEST ACIDS	HClO <sub>4</sub>	ClO <sub>4</sub> <sup>-</sup>	WEAKEST BASES
	H <sub>2</sub> SO <sub>4</sub>	HSO <sub>4</sub> <sup>-</sup>	
	HI	I <sup>-</sup>	
	HBr	Br <sup>-</sup>	
	HCl	Cl <sup>-</sup>	
	HNO <sub>3</sub>	NO <sub>3</sub> <sup>-</sup>	
	H <sub>3</sub> O <sup>+</sup>	H <sub>2</sub> O	
	HSO <sub>4</sub>	SO <sub>4</sub> <sup>-</sup>	
	H <sub>2</sub> SO <sub>3</sub>	HSO <sub>3</sub> <sup>-</sup>	
	H <sub>3</sub> PO <sub>4</sub>	H <sub>2</sub> PO <sub>4</sub> <sup>-</sup>	
	HNO <sub>2</sub>	NO <sub>2</sub> <sup>-</sup>	
	HF	F <sup>-</sup>	
	HC <sub>2</sub> H <sub>3</sub> O <sub>2</sub>	C <sub>2</sub> H <sub>3</sub> O <sub>2</sub> <sup>-</sup>	
	Al(H <sub>2</sub> O) <sub>6</sub> <sup>3+</sup>	Al(H <sub>2</sub> O) <sub>5</sub> <sup>2+</sup>	
	H <sub>2</sub> CO <sub>3</sub>	HCO <sub>3</sub> <sup>-</sup>	
	H <sub>2</sub> S	HS <sup>-</sup>	
	HClO	ClO <sup>-</sup>	
	HBrO	BrO <sup>-</sup>	
	NH <sub>4</sub> <sup>+</sup>	NH <sub>3</sub>	
	HCN	CN <sup>-</sup>	
	HCO <sub>3</sub> <sup>-</sup>	CO <sub>3</sub> <sup>-2</sup>	
	H <sub>2</sub> O <sub>2</sub>	HO <sub>2</sub> <sup>-</sup>	
WEAKEST ACIDS	HS <sup>-</sup>	S <sub>2</sub> <sup>-</sup>	STRONGEST BASES
	H <sub>2</sub> O	OH <sup>-</sup>	

The stronger acid will have the weaker conjugated base

The stronger base will have the weaker conjugated acid

The direction of the acid-base reaction always favors the weaker acid and base:

"The stronger acid plus the stronger base produces the weaker acid and weaker base"

### III. Lewis Acid-Base Concept

#### A. DEFINITION

Lewis Acid  $\Rightarrow$  A substance that is an electron pair acceptor (A covalent bond is made)  
ex.

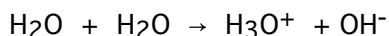
Lewis Base  $\Rightarrow$  A substance that is an electron pair donator (A covalent bond is made)  
ex.

An Acid-Base Reaction:

### Part 5: Autoionization of Water

#### I. THE CONCEPT

Experiments have shown that a very small percentage of water will undergo the following ionization to produce ions:



Experimentally, it was determined that the product between the molar concentrations of the hydronium ion and hydroxide ion is a constant:

$$[\text{H}_3\text{O}^+] \times [\text{OH}^-] = 10^{-14} \quad \leftarrow \text{ion product for water (constant)}$$

$$\text{or } [\text{H}^+] \times [\text{OH}^-] = 10^{-14} \quad \text{as an abbreviation}$$

Therefore, in pure water,  $[\text{H}_3\text{O}^+] = [\text{OH}^-] = 10^{-7}$

Problems:

1. What is the hydrogen (hydronium) ion concentration when  $[\text{OH}^-] = 2 \times 10^{-3}\text{M}$ ?
2. What is the hydroxide ion concentration in a 0.002 M HCl solution?
3. What is the hydrogen ion concentration in a  $1.00 \times 10^{-2}$  M NaOH solution?

## II. THE pH SCALE

Hydrogen ion concentrations,  $[H^+]$  and hydroxide ion concentrations,  $[OH^-]$  are usually very small numbers...

Example:  $[H^+] = 2 \times 10^{-1}M$  and  $[H^+] = 1 \times 10^{-11}M$ .

The pH scale was developed to handle these very small numbers over a wide range.

( $2 \times 10^{-1}$  is 20 **trillion** times larger than  $1 \times 10^{-11}$ !)

$$pH = -\log [H_3O^+]$$

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

Problems:

1. What is the pH when  $[H^+] = 10^{-3}M$ ?
2. What is the what is the pH of a 0.002 M HCl solution?
3. What is the pH of a  $1.00 \times 10^{-2}$  M NaOH solution?
4. What is the  $[H^+]$  if the pH = 2.56
5. What is the  $[H^+]$  of a NaOH solution with a pOH of 9.983?

Acidic, Basic, and Neutral solutions:

$[H_3O^+]$	$[OH^-]$	pH	acidic/basic or/neutral
$10^0$	$10^{-14}$	0	
$10^{-1}$	$10^{-13}$	1	
$10^{-2}$	$10^{-12}$	2	
$10^{-3}$	$10^{-11}$	3	
$10^{-4}$	$10^{-10}$	4	
$10^{-5}$	$10^{-9}$	5	
$10^{-6}$	$10^{-8}$	6	
$10^{-7}$	$10^{-7}$	7	
$10^{-8}$	$10^{-6}$	8	
$10^{-9}$	$10^{-5}$	9	
$10^{-10}$	$10^{-4}$	10	
$10^{-11}$	$10^{-3}$	11	
$10^{-12}$	$10^{-2}$	12	
$10^{-13}$	$10^{-1}$	13	
$10^{-14}$	$10^0$	14	

Problem: Is  $2 \times 10^{-10}$  M  $[H^+]$  a basic, acidic or neutral solution?



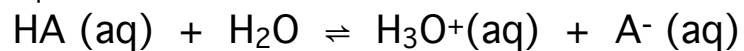
### **III. INDICATORS**

Indicators are often used to determine the approximate pH of solutions. Indicators are weak acids or weak bases whose conjugated species is a different color than it is:

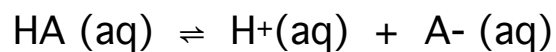
## Part 6: Acid - Base Equilibria

### I. WEAK ACID IONIZATION

#### A. Equilibria of a weak acid.



Or



ex.

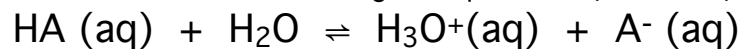
#### B. Acid Ionization constant, $K_a$

#### C. $\text{PK}_a = -\log K_a$

#### D. Degree of Ionization, $\alpha$

The degree of ionization,  $\alpha$ , of a weak acid or base in water is the fraction of the total concentration that has formed ions:

A 0.10 M HA solution undergoes equilibrium (ionization)



## Ka TABLE

## ACID-IONIZATION CONSTANTS AT 25°C\*

Substance	Formula	Ka
Acetic acid	HC <sub>2</sub> H <sub>3</sub> O <sub>2</sub>	1.7 X 10 <sup>-5</sup>
Benzoic acid	HC <sub>7</sub> H <sub>5</sub> O <sub>5</sub>	6.3 X 10 <sup>-5</sup>
Boric acid	H <sub>3</sub> BO <sub>3</sub>	Ka <sub>1</sub> 5.9 X 10 <sup>-10</sup>
Carbonic acid	H <sub>2</sub> CO <sub>3</sub>	Ka <sub>1</sub> = 4.3 X 10 <sup>-7</sup>
	HCO <sub>3</sub> <sup>-</sup>	Ka <sub>2</sub> = 4.8 X 10 <sup>-11</sup>
Cyanic acid	HOCN	3.5 X 10 <sup>-4</sup>
Formic acid	HCHO <sub>2</sub>	1.7 X 10 <sup>-4</sup>
Hydrocyanic acid	HCN	4.9 X 10 <sup>-10</sup>
Hydrofluoric acid	HF	6.8 X 10 <sup>-4</sup>
Hydrogen sulfate ion	HSO <sub>4</sub> <sup>-</sup>	1.1 X 10 <sup>-2</sup>
Hydrogen sulfide	H <sub>2</sub> S	Ka <sub>1</sub> = 8.9 X 10 <sup>-8</sup>
	HS <sup>-</sup>	Ka <sub>2</sub> = 1.2 X 10 <sup>-13</sup> †
Hypochlorous acid	HClO	3.5 X 10 <sup>-8</sup>
Nitrous acid	HNO <sub>2</sub>	4.5 X 10 <sup>-4</sup>
Oxalic acid	H <sub>2</sub> C <sub>2</sub> O <sub>4</sub>	Ka <sub>1</sub> = 5.6 X 10 <sup>-2</sup>
	HC <sub>2</sub> O <sub>4</sub> <sup>-</sup>	Ka <sub>2</sub> = 5.1 X 10 <sup>-5</sup>
Phosphoric acid	H <sub>3</sub> PO <sub>4</sub>	Ka <sub>1</sub> = 6.9 X 10 <sup>-3</sup>
	H <sub>2</sub> PO <sub>4</sub> <sup>-</sup>	Ka <sub>2</sub> = 6.2 X 10 <sup>-8</sup>
	HPO <sub>4</sub> <sup>2-</sup>	Ka <sub>3</sub> = 4.8 X 10 <sup>-13</sup>
Phosphorous acid	H <sub>3</sub> PO <sub>3</sub>	Ka <sub>1</sub> = 1.6 X 10 <sup>-2</sup>
	H <sub>2</sub> PO <sub>3</sub> <sup>-</sup>	Ka <sub>2</sub> = 7 X 10 <sup>-7</sup>
Propionic acid	HC <sub>3</sub> H <sub>5</sub> O <sub>2</sub>	1.3 X 10 <sup>-5</sup>
Pyruvic acid	HC <sub>3</sub> H <sub>3</sub> O <sub>3</sub>	1.4 X 10 <sup>-4</sup>
Sulfurous acid	H <sub>2</sub> SO <sub>3</sub>	Ka <sub>1</sub> = 1.3 X 10 <sup>-2</sup>
	HSO <sub>3</sub> <sup>-</sup>	Ka <sub>2</sub> = 6.3 X 10 <sup>-8</sup>

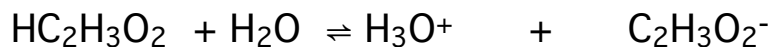
\*The ionization constants for polyprotic acids are for successive ionizations. Thus, for H<sub>3</sub>PO<sub>4</sub>, the equilibrium is H<sub>3</sub>PO<sub>4</sub> ⇌ H<sup>+</sup> + H<sub>2</sub>PO<sub>4</sub><sup>-</sup>. For H<sub>2</sub>PO<sub>4</sub><sup>-</sup>, the equilibrium is H<sub>2</sub>PO<sub>4</sub><sup>-</sup> ⇌ H<sup>+</sup> + HPO<sub>4</sub><sup>2-</sup>.

†This value is in doubt. Some evidence suggests that it is about  $10^{-19}$ .

F. ACID EQUILIBRIUM CALCULATIONS:

1. Calculation of equilibrium concentrations from  $K_a$

Calculate the pH of a 1.0 M Acetic acid solution, using approximations.  $K_a = 1.8 \times 10^{-5}$



*Initial*

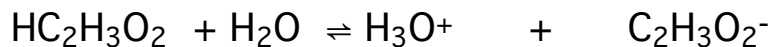
*Change*

*at Equilibrium*

\_\_\_\_\_

2. Calculation of species concentrations from  $K_a$ , using the quadratic formula

Calculate the pH of a 0.000010 M Acetic acid solution



*Initial*

*Change*

*at Equilibrium*

\_\_\_\_\_

3. Calculation of Ka from the pH of a weak acid solution

Calculate the Ka of HNO<sub>2</sub> if a 0.10 M HNO<sub>2</sub> solution has a pH of 2.187

*Initial*

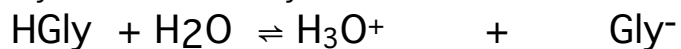
*Change*

*at Equilibrium*

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4. Calculation of Ka from the percent ionization.

Calculate the Ka of Glycine if a 0.10 M Glycine solution is  $4.1 \times 10^{-3}$  ionized



*Initial*

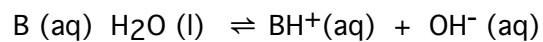
*Change*

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**AT EQUILIBRIUM**

## II. WEAK BASE IONIZATION

### A. Equilibria of a weak base.



ex.

### B. Types of weak bases

1. Insoluble
2. Soluble  
Molecular-Nitrogen based Bases

### C. Base Ionization constant, $K_b$

### D. Ammonia Equilibrium

### F. Degree of Ionization, $\alpha$

The degree of ionization,  $\alpha$ , of a weak acid or base in water is the fraction of the total concentration that has formed ions: A 1.0 M B solution undergoes equilibrium (ionization)



Kb table

DISSOCIATION CONSTANTS FOR BASES

Base	Formula	Dissociation Constants at 25°C
Ammonia	NH <sub>3</sub>	1.76 X 10 <sup>-5</sup>
Aniline	C <sub>6</sub> H <sub>5</sub> NH <sub>2</sub>	3.94 X 10 <sup>-10</sup>
1-Butylamine	CH <sub>3</sub> (CH <sub>2</sub> ) <sub>2</sub> CH <sub>2</sub> NH <sub>2</sub>	4.0 X 10 <sup>-4</sup>
Dimethylamine	(CH <sub>3</sub> ) <sub>2</sub> NH	5.9 X 10 <sup>-4</sup>
Ethanolamine	HOC <sub>2</sub> H <sub>4</sub> NH <sub>2</sub>	3.18 X 10 <sup>-5</sup>
Ethylamine	CH <sub>3</sub> CH <sub>2</sub> NH <sub>2</sub>	4.28 X 10 <sup>-4</sup>
Ethylenediamine	NH <sub>2</sub> C <sub>2</sub> H <sub>4</sub> NH <sub>2</sub>	K = 8.5 X 10 <sup>-5</sup> K = 7.1 X 10 <sup>-8</sup>
Hydrazine	H <sub>2</sub> NNH <sub>2</sub>	1.3 X 10 <sup>-6</sup>
Hydroxylamine	HONH <sub>2</sub>	1.07 X 10 <sup>-8</sup>
Methylamine	CH <sub>3</sub> NH <sub>2</sub>	4.8 X 10 <sup>-4</sup>
Piperidine	C <sub>5</sub> H <sub>11</sub> N	1.3 X 10 <sup>-3</sup>
Pyridine	C <sub>5</sub> H <sub>5</sub> N	1.7 X 10 <sup>-9</sup>
Trimethyl amine	(CH <sub>3</sub> ) <sub>3</sub> N	6.25 X 10 <sup>-5</sup>

G. Base Equilibrium Calculations:

1. Calculation of species concentrations from Kb

Calculate the pH of a 0.100 M NH<sub>3</sub> solution Kb = 1.8 x 10<sup>-5</sup>

*Initial*

*Change*

*at Equilibrium*

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2. Calculation of  $K_b$  from species concentrations

Calculate the  $K_b$  for methylamine,  $\text{CH}_3\text{NH}_2$ , if the pOH of a 0.0100 M  $\text{CH}_3\text{NH}_2$  solution is 7.66

*Initial*

*Change*

*at Equilibrium*

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3. Calculation of species concentrations from  $K_b$

Calculate the pH of a 0.100 M  $\text{C}_5\text{H}_5\text{N}$   $K_b = 1.3 \times 10^{-3}$

*Initial*

*Change*

*at Equilibrium*

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**III. THE EFFECT OF DILUTION ON THE % IONIZATION**

The percent ionization of a weak acid or base increase as the acid/base concentration decreases (dilution)

1. Calculate the percent ionization for 0.10 M HF

2. Calculate the percent ionization for 0.00010 M HF

Equilibrium problems



