

# *ACIDS AND BASES*

## **CHM 101: GENERAL CHEMISTRY 1**

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(Lecture : Ci9- 7)



# *Definition Of Acids And Bases*

Acids and bases shall be defined based on the following theories

- Arrehenius
- Bronsted Lowry
- Lewis theory

B o b b y s a w

# *Definition of acids and bases contd.*

## **Arrhenius definition:**

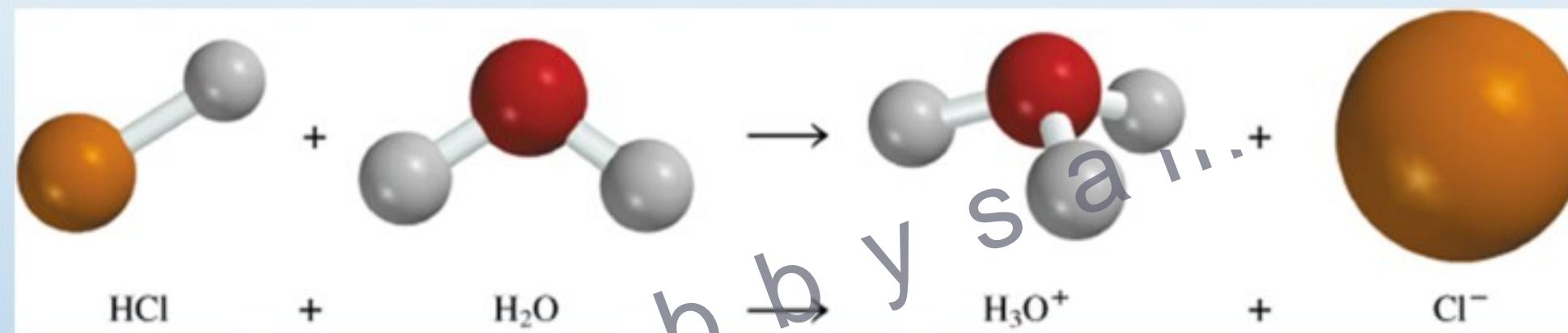
acids: generate  $\text{H}^+$  /  $(\text{H}_3\text{O}^+)$  in water

bases: generate  $\text{OH}^-$  in water

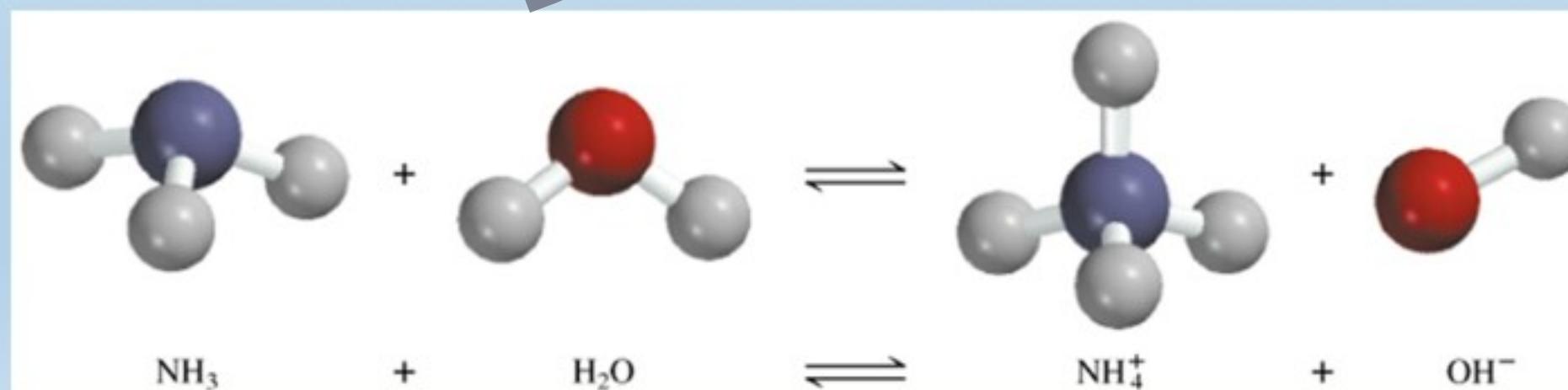
B o b b y s a w

# Arrhenius acids and bases

Arrhenius acid is a substance that produces  $\text{H}^+$  ( $\text{H}_3\text{O}^+$ ) in water



Arrhenius base is a substance that produces OH<sup>-</sup> in water



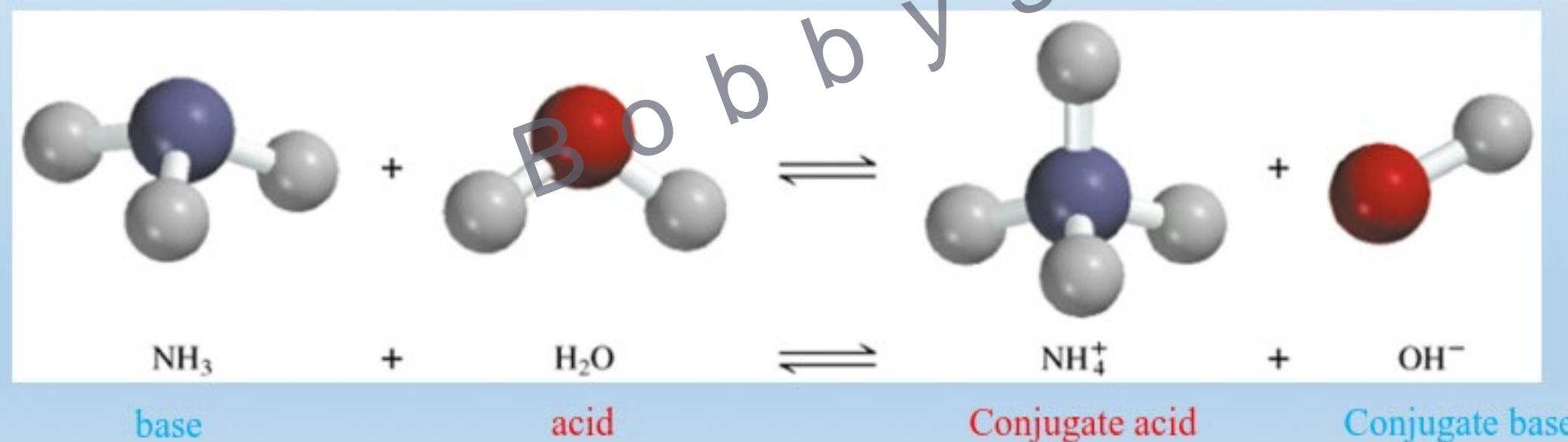
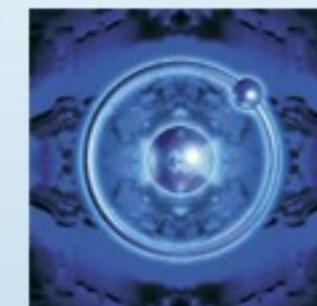
Limitation  
of  
Arrhenius  
definition is  
some bases  
don't have  
hydroxide  
ions

# Bronsted-Lowry acids and bases

Acid: proton ( $H^+$ ) ion donor

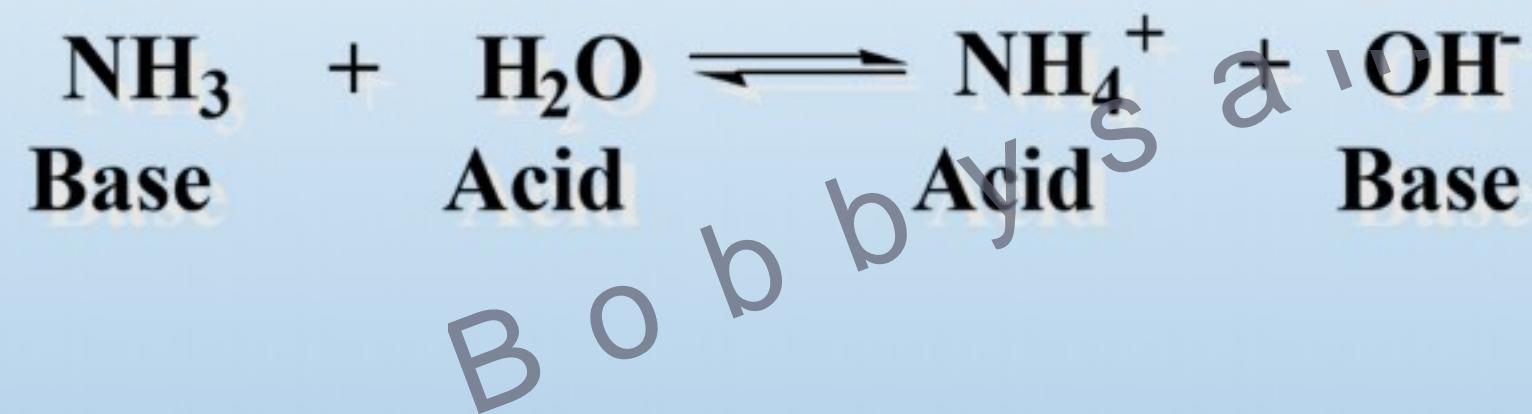
Base: proton ( $H^+$ ) ion acceptor

A “proton” is a hydrogen atom that has lost its electron

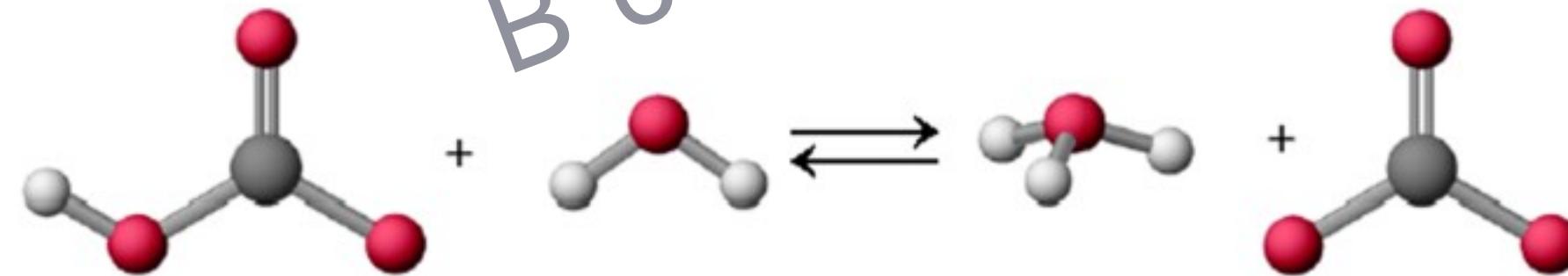
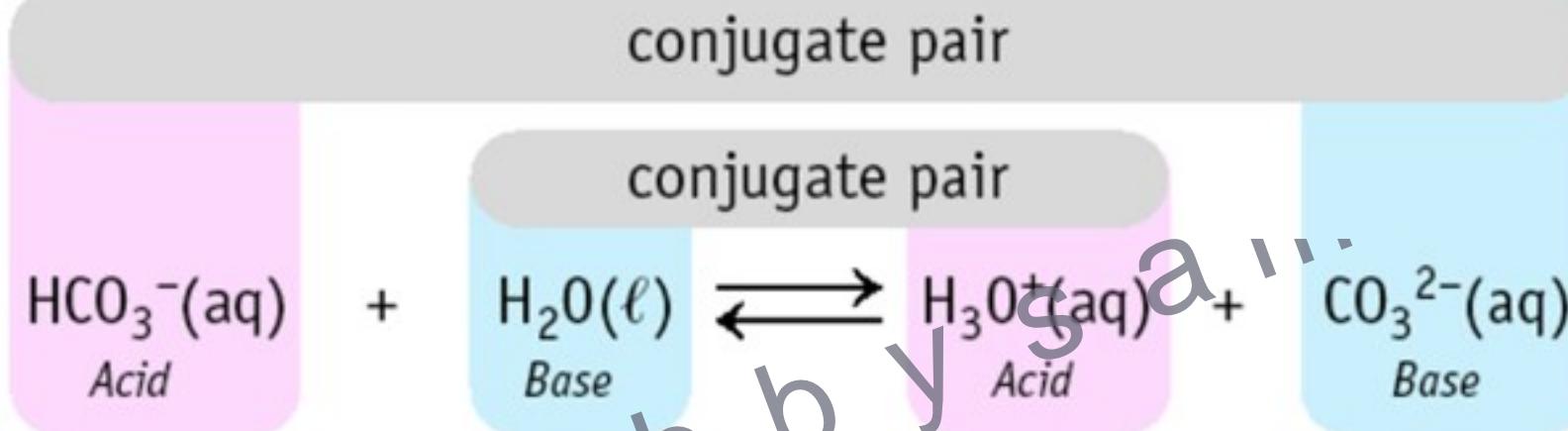


## Bronsted-Lowry acids and bases contd.

The Brønsted definition means NH<sub>3</sub> is a **BASE** in water — and water is itself an **ACID**

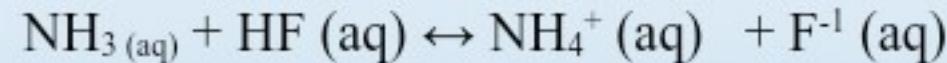


# Conjugate Pairs



# Check out!

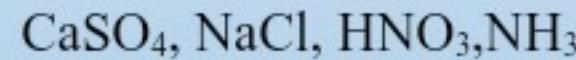
1. Label the acid, base, conjugate acid, and conjugate base in each reaction:



2. Which of the following compounds is a Brønsted-Lowry base



3. Which of the following compounds is a strong acid?



# Lewis definition of acids and bases

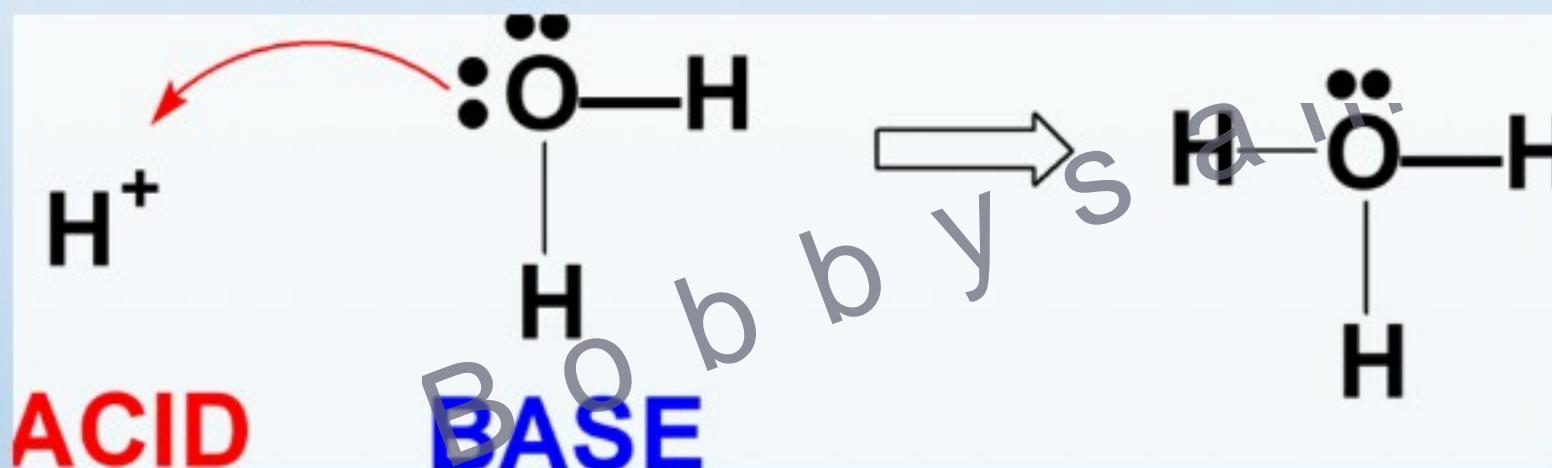
Lewis acid - a substance that accepts an electron pair

Lewis base - a substance that donates an electron pair

Bobby says a ...

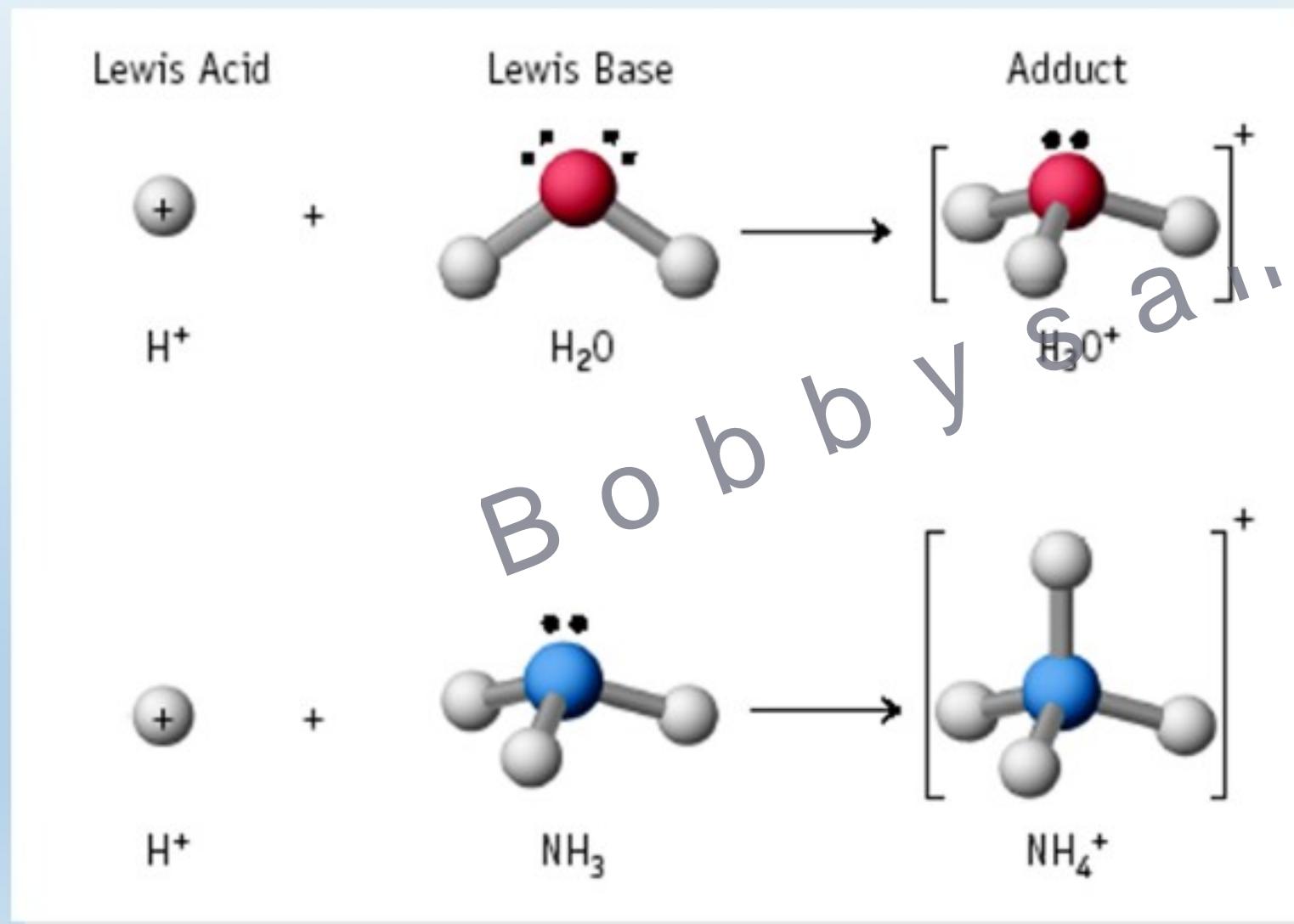
# Lewis Acids & Bases

Formation of hydronium ion is also an excellent example.



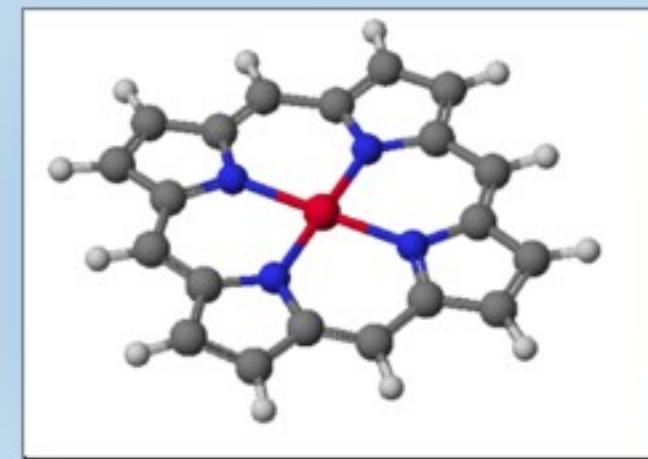
- Electron pair of the new O-H bond originates on the Lewis base.

# Lewis Acid/Base Reaction



# Lewis Acid-Base Interactions in Biology

- The heme group in hemoglobin can interact with O<sub>2</sub> and CO.
- The Fe ion in hemoglobin is a Lewis acid
- O<sub>2</sub> and CO can act as Lewis bases



# *Strengths of Acids and Bases*

Strong acids donate  $\text{H}^+$  ions more easily

The stronger the acid, the weaker the conjugate base associated with that acid

Strong bases accept  $\text{H}^+$  ions more easily

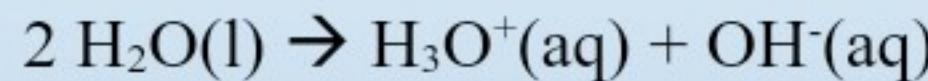
The stronger the base, the weaker the conjugate acid associated with that base

B o b b y s a w

Strong 100% ionized (H <sup>+</sup> completely donated to water)		Extremely weak (negligible H <sup>+</sup> acceptance from water)	
<b>Conjugate acid</b>		<b>Conjugate base</b>	
H <sub>2</sub> SO <sub>4</sub>		HSO <sub>4</sub> <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	
H <sub>2</sub> SO <sub>3</sub> (sulfurous)		HSO <sub>3</sub> <sup>-</sup>	
HSO <sub>4</sub> <sup>-</sup>		SO <sub>4</sub> <sup>2-</sup>	
H <sub>3</sub> PO <sub>4</sub> (phosphoric)		H <sub>2</sub> PO <sub>4</sub> <sup>-</sup>	
[Fe(H <sub>2</sub> O) <sub>5</sub> ] <sup>3+</sup>		[Fe(H <sub>2</sub> O) <sub>5</sub> (OH)] <sup>2+</sup>	
HF (hydrofluoric)		F <sup>-</sup>	
HNO <sub>2</sub> (nitrous)		NO <sub>2</sub> <sup>-</sup>	
CH <sub>3</sub> COOH (acetic)		CH <sub>3</sub> COO <sup>-</sup>	
H <sub>2</sub> CO <sub>3</sub> (carbonic)		HCO <sub>3</sub> <sup>-</sup>	
H <sub>2</sub> S (hydrosulfuric)		HS <sup>-</sup>	
H <sub>2</sub> PO <sub>4</sub> <sup>-</sup>		HPO <sub>4</sub> <sup>2-</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		OH <sup>-</sup>	
OH <sup>-</sup>		O <sup>2-</sup>	
H <sub>2</sub>		H <sup>-</sup>	
CH <sub>4</sub>		CH <sub>3</sub> <sup>-</sup>	

# Autoionization of Water

Water always undergoes some degree of dissociation to form H<sub>3</sub>O<sup>+</sup> ions and OH<sup>-</sup> ions



The equilibrium constant for this process at 25 °C is:

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

In pure water

$$[\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \text{ M}$$

# Autoionization of Water contd.

$K_w$  is temperature dependent—it increases with increasing temperature

Table 1: Temperature Dependence of  $K_w$  for Water

$T$ ( $^{\circ}\text{C}$ )	$K_w$
10	$0.29 \times 10^{-14}$
15	$0.45 \times 10^{-14}$
20	$0.68 \times 10^{-14}$
25	$1.01 \times 10^{-14}$
30	$1.47 \times 10^{-14}$
50	$5.48 \times 10^{-14}$

# The pH Scale

pH is defined as:

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

log is log base 10, not ln (natural log);  $[\text{H}_3\text{O}^+]$  is given in molar units (M)

pH of pure water ( $[\text{H}_3\text{O}^+] = (1.0 \times 10^{-7} \text{ M})$ ):

$$\text{pH} = -\log(1.0 \times 10^{-7}) = 7.0$$

pH of last example ( $[\text{H}_3\text{O}^+] = 1.9 \times 10^{-13} \text{ M}$ ):

$$\text{pH} = -\log(1.9 \times 10^{-13}) = 12.7$$

$$\begin{aligned} [\text{H}_3\text{O}^+] &= K_w / [\text{OH}^-] = 1.0 \times 10^{-14} / 0.053 \\ &= 1.9 \times 10^{-13} \text{ M} \end{aligned}$$

## The pH Scale contd.

Neutral is defined as the pH of pure water:

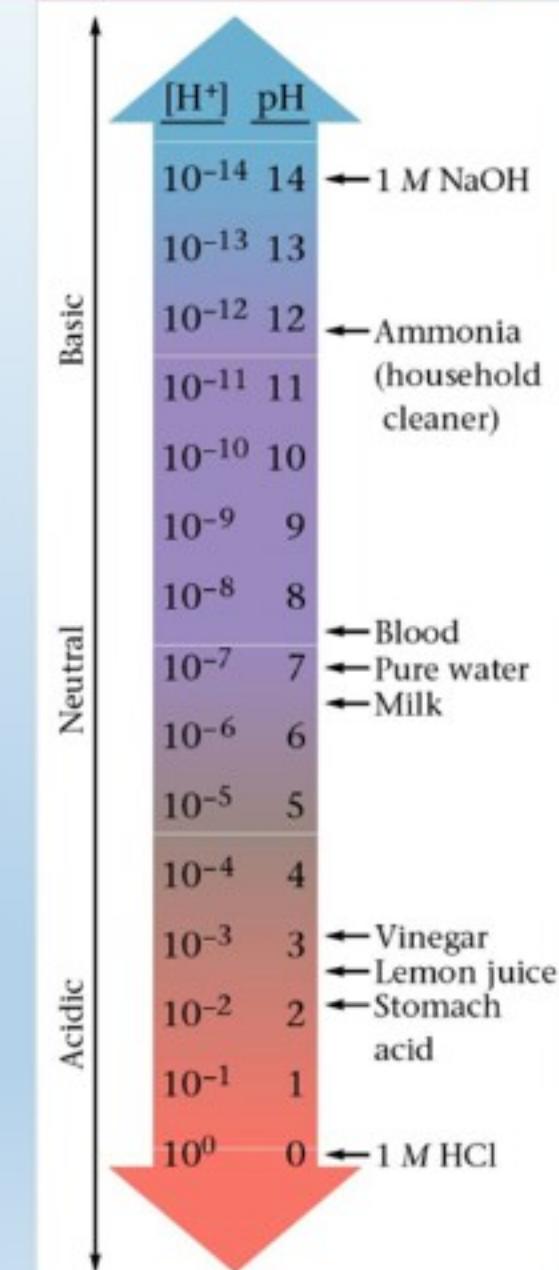
$$\text{pH} = 7$$

Acidic solutions have pH lower than 7:

$$\text{pH} < 7 \rightarrow \text{acidic}$$

Basic solutions have pH larger than 7:

$$\text{pH} > 7 \rightarrow \text{basic}$$



## *The pH Scale contd.*

We can also use pOH to describe a solution

pOH is defined as:

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

The sum of pH and pOH must equal 14

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

assuming room temperature (25 °C )

## *The pH Scale contd.*

Example: Find  $[H_3O^+]$  of a solution that has  $pOH = 9.37$

**Method 1:** Calculate pH, then  $[H_3O^+]$

**Step 1:** Determine pH

$$pH = 14 - pOH = 14.00 - 9.37 = 4.63$$

**Step 2:** Determine  $[H_3O^+]$

$$H_3O^+ = 10^{-pH} = 10^{-4.63} = 2.34 \times 10^{-5} M$$

# *The pH Scale contd.*

Example (con't.): Find  $[H_3O^+]$  of a solution that has  $pOH = 9.37$

**Method 2:** Calculate  $[OH^-]$  then  $[H_3O^+]$  using  $K_w$

**Step 1:** Determine  $[OH^-]$

$$[OH^-] = 10^{-pOH} = 10^{-9.37} = 4.27 \times 10^{-10} M$$

**Step 2:** Determine  $[H_3O^+]$  using  $K_w$

$$\begin{aligned}[H_3O^+] &= K_w/[OH^-] = (1.0 \times 10^{-14})/(4.27 \times 10^{-10}) \\ &= 2.34 \times 10^{-5} M\end{aligned}$$

# pH calculation of solution.

Example: Determine  $[H_3O^+]$  in a 0.053 M NaOH solution

**Step 1:** since NaOH is a strong base, dissociation is complete

$$[OH^-] = 0.053 \text{ M}$$

**Step 2:** Use  $K_w$  to calculate  $[H_3O^+]$

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

$$\begin{aligned}[H_3O^+] &= K_w / [OH^-] = 1.0 \times 10^{-14} / 0.053 \\ &= 1.9 \times 10^{-13} \text{ M}\end{aligned}$$

# Try these!!!

**Problem 1:** The pH of rainwater collected in a certain region of the northeastern Nigeria on a particular day was 4.82. What is the H<sup>+</sup> ion concentration of the rainwater?

**Problem 2:** The OH<sup>-1</sup> ion concentration of a blood sample is  $2.5 \times 10^{-7}$  M. What is the pH of the blood?

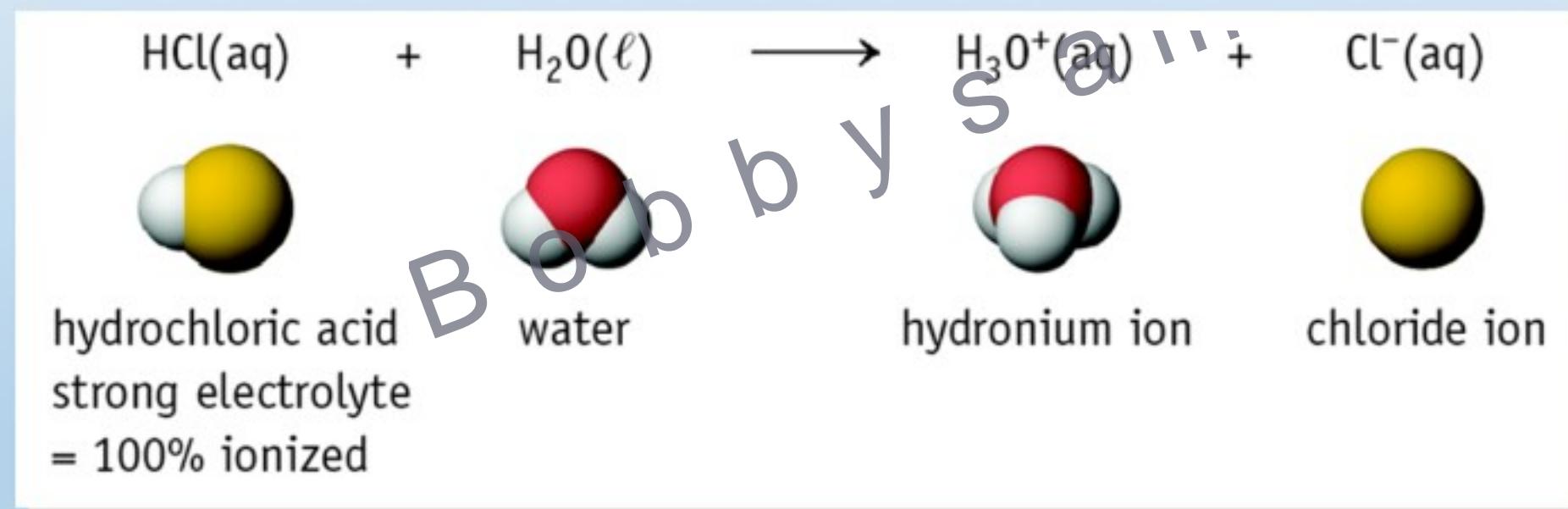
**Problem 3:** A chemist dilutes concentrated hydrochloric acid to make two solutions: (a) 3.0 M and (b) 0.0024 M. Calculate the [H<sub>3</sub>O<sup>+</sup>], pH, [OH<sup>-</sup>], and pOH of the two solutions at 25°C.

**Problem 4:** What is the [H<sub>3</sub>O<sup>+</sup>], [OH<sup>-</sup>], and pOH of a solution with pH = 3.67? Is this an acid, base, or neutral?

**Problem 5:** Problem #4 with pH = 8.05?

# Strong and Weak Acids/Bases

The strength of an acid (or base) is determined by the amount of **IONIZATION**.



$\text{HNO}_3$ ,  $\text{HCl}$ ,  $\text{H}_2\text{SO}_4$  and  $\text{HClO}_4$  are among the only known strong acids.

# **Strong and Weak Acids/Bases** contd.

Generally acids and bases can be either STRONG or WEAK .



$\text{HNO}_3$  is about 100% dissociated in water.

**Weak acids** are much less than 100% ionized in water.

One of the best known is acetic acid =  $\text{CH}_3\text{CO}_2\text{H}$

# *Strong and Weak Acids/Bases contd.*

**Strong Base:** 100% dissociated in water.



Other common strong bases include KOH and  $\text{Ca}(\text{OH})_2$ .



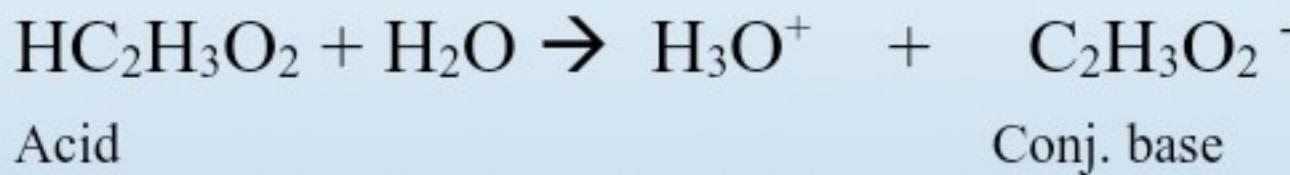
**Weak base:** less than 100% ionized in water

One of the best known weak bases is ammonia



# Equilibria Involving Weak Acids and Bases

Consider acetic acid,  $\text{HC}_2\text{H}_3\text{O}_2$  ( $\text{HOAc}$ )



$$K_a = \frac{[H_3O^+][OAc^-]}{[HOAc]} = 1.8 \times 10^{-5}$$

(K is designated  $K_a$  for dissociation constant of ACID)

K gives the ratio of ions (split up) to molecules (don't split up)

Table 2: Ionization Constant for some Acids and Their Conjugate Bases

## ACIDS

Increase strength

Acid Name	Acid	$K_a$	Base	$K_b$	Base Name
Perchloric acid	$\text{HClO}_4$	large	$\text{ClO}_4^-$	very small	perchlorate ion
Sulfuric acid	$\text{H}_2\text{SO}_4$	large	$\text{HSO}_4^-$	very small	hydrogen sulfate ion
Hydrochloric acid	$\text{HCl}$	large	$\text{Cl}^-$	very small	chloride ion
Nitric acid	$\text{HNO}_3$	large	$\text{NO}_3^-$	very small	nitrate ion
Hydronium ion	$\text{H}_3\text{O}^+$	1.0	$\text{H}_2\text{O}$	$1.0 \times 10^{-14}$	water
Sulfurous acid	$\text{H}_2\text{SO}_3$	$1.2 \times 10^{-2}$	$\text{HSO}_3^-$	$8.3 \times 10^{-13}$	hydrogen sulfite ion
Hydrogen sulfate ion	$\text{HSO}_4^-$	$1.2 \times 10^{-2}$	$\text{SO}_4^{2-}$	$8.3 \times 10^{-13}$	sulfate ion
Phosphoric acid	$\text{H}_3\text{PO}_4$	$7.5 \times 10^{-3}$	$\text{H}_2\text{PO}_4^-$	$1.3 \times 10^{-12}$	dihydrogen phosphate ion
Hexaaquaiiron(III) ion	$\text{Fe}(\text{H}_2\text{O})_6^{3+}$	$6.3 \times 10^{-3}$	$\text{Fe}(\text{H}_2\text{O})_5\text{OH}^{2+}$	$1.6 \times 10^{-12}$	pentaaqua hydroxoholiron(III) ion
Hydrofluoric acid	$\text{HF}$	$7.2 \times 10^{-4}$	$\text{F}^-$	$1.4 \times 10^{-11}$	fluoride ion
Nitrous acid	$\text{HNO}_2$	$4.5 \times 10^{-4}$	$\text{NO}_2^-$	$2.2 \times 10^{-11}$	nitrite ion
Formic acid	$\text{HCO}_2\text{H}$	$1.8 \times 10^{-4}$	$\text{HCO}_2^-$	$5.6 \times 10^{-11}$	formate ion
Benzoic acid	$\text{C}_6\text{H}_5\text{CO}_2\text{H}$	$6.3 \times 10^{-5}$	$\text{C}_6\text{H}_5\text{CO}_2^-$	$1.6 \times 10^{-10}$	benzoate ion
Acetic acid	$\text{CH}_3\text{CO}_2\text{H}$	$1.8 \times 10^{-5}$	$\text{CH}_3\text{CO}_2^-$	$5.6 \times 10^{-10}$	acetate ion
Propanoic acid	$\text{CH}_3\text{CH}_2\text{CO}_2\text{H}$	$1.3 \times 10^{-5}$	$\text{CH}_3\text{CH}_2\text{CO}_2^-$	$7.7 \times 10^{-10}$	propanoate ion
Hexaaquaaluminum ion	$\text{Al}(\text{H}_2\text{O})_6^{3+}$	$7.9 \times 10^{-6}$	$\text{Al}(\text{H}_2\text{O})_5\text{OH}^{2+}$	$1.3 \times 10^{-9}$	pentaaqua hydroxoholuminum ion
Carbonic acid	$\text{H}_2\text{CO}_3$	$4.2 \times 10^{-7}$	$\text{HCO}_3^-$	$2.4 \times 10^{-8}$	hydrogen carbonate ion
Hexaaquacopper(II) ion	$\text{Cu}(\text{H}_2\text{O})_6^{2+}$	$1.6 \times 10^{-7}$	$\text{Cu}(\text{H}_2\text{O})_5\text{OH}^+$	$6.25 \times 10^{-8}$	pentaaqua hydroxocopper(II) ion
Hydrogen sulfide	$\text{H}_2\text{S}$	$1 \times 10^{-7}$	$\text{HS}^-$	$1 \times 10^{-7}$	hydrogen sulfide ion
Dihydrogen phosphate ion	$\text{H}_2\text{PO}_4^-$	$6.2 \times 10^{-8}$	$\text{HPO}_4^{2-}$	$1.6 \times 10^{-7}$	hydrogen phosphate ion
Hydrogen sulfite ion	$\text{HSO}_3^-$	$6.2 \times 10^{-8}$	$\text{SO}_3^{2-}$	$1.6 \times 10^{-7}$	sulfite ion
Hypochlorous acid	$\text{HClO}$	$3.5 \times 10^{-8}$	$\text{ClO}^-$	$2.9 \times 10^{-7}$	hypochlorite ion
Hexaaqualead(II) ion	$\text{Pb}(\text{H}_2\text{O})_6^{2+}$	$1.5 \times 10^{-8}$	$\text{Pb}(\text{H}_2\text{O})_5\text{OH}^+$	$6.7 \times 10^{-7}$	pentaaqua hydroxolead(II) ion
Hexaaquacobalt(II) ion	$\text{Co}(\text{H}_2\text{O})_6^{2+}$	$1.3 \times 10^{-9}$	$\text{Co}(\text{H}_2\text{O})_5\text{OH}^+$	$7.7 \times 10^{-6}$	pentaaqua hydroxocobalt(II) ion
Boric acid	$\text{B}(\text{OH})_3(\text{H}_2\text{O})$	$7.3 \times 10^{-10}$	$\text{B}(\text{OH})_4^-$	$1.4 \times 10^{-5}$	tetrahydroxoborate ion
Ammonium ion	$\text{NH}_4^+$	$5.6 \times 10^{-10}$	$\text{NH}_3$	$1.8 \times 10^{-5}$	ammonia
Hydrocyanic acid	$\text{HCN}$	$4.0 \times 10^{-10}$	$\text{CN}^-$	$2.5 \times 10^{-5}$	cyanide ion
Hexaaquaiiron(II) ion	$\text{Fe}(\text{H}_2\text{O})_6^{2+}$	$3.2 \times 10^{-10}$	$\text{Fe}(\text{H}_2\text{O})_5\text{OH}^+$	$3.1 \times 10^{-5}$	pentaaqua hydroxoholiron(II) ion
Hydrogen carbonate ion	$\text{HCO}_3^-$	$4.8 \times 10^{-11}$	$\text{CO}_3^{2-}$	$2.1 \times 10^{-4}$	carbonate ion
Hexaaquanickel(II) ion	$\text{Ni}(\text{H}_2\text{O})_6^{2+}$	$2.5 \times 10^{-11}$	$\text{Ni}(\text{H}_2\text{O})_5\text{OH}^+$	$4.0 \times 10^{-4}$	pentaaqua hydroxonickel(II) ion
Hydrogen phosphate ion	$\text{HPO}_4^{2-}$	$3.6 \times 10^{-11}$	$\text{PO}_4^{3-}$	$2.8 \times 10^{-2}$	phosphate ion
Water	$\text{H}_2\text{O}$	$1.0 \times 10^{-14}$	$\text{OH}^-$	1.0	hydroxide ion
Hydrogen sulfide ion*	$\text{HS}^-$	$1 \times 10^{-15}$	$\text{S}^{2-}$	$1 \times 10^5$	sulfide ion
Ethanol	$\text{C}_2\text{H}_5\text{OH}$	very small	$\text{C}_2\text{H}_5\text{O}^-$	large	ethoxide ion
Ammonia	$\text{NH}_3$	very small	$\text{NH}_2^-$	large	amide ion
Hydrogen	$\text{H}_2$	very small	$\text{H}^-$	large	hydride ion

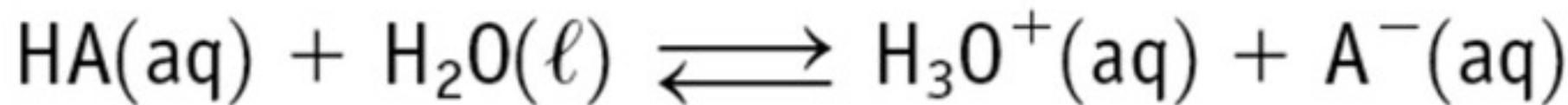
\*The values of  $K_a$  for  $\text{HS}^-$  and  $K_b$  for  $\text{S}^{2-}$  are estimates.

## Conjugate Bases

Increase strength

# Equilibrium Constants for Weak Acids

contd.

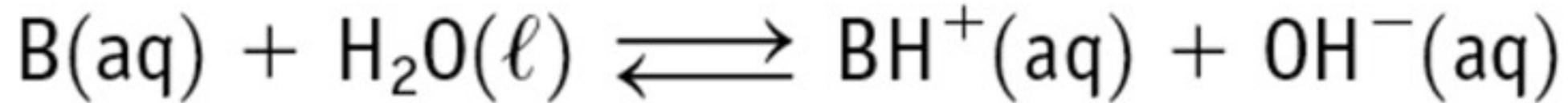


$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

Weak acid has  $K_a < 1$

Leads to small  $[\text{H}_3\text{O}^+]$  and a pH of 2 - 7

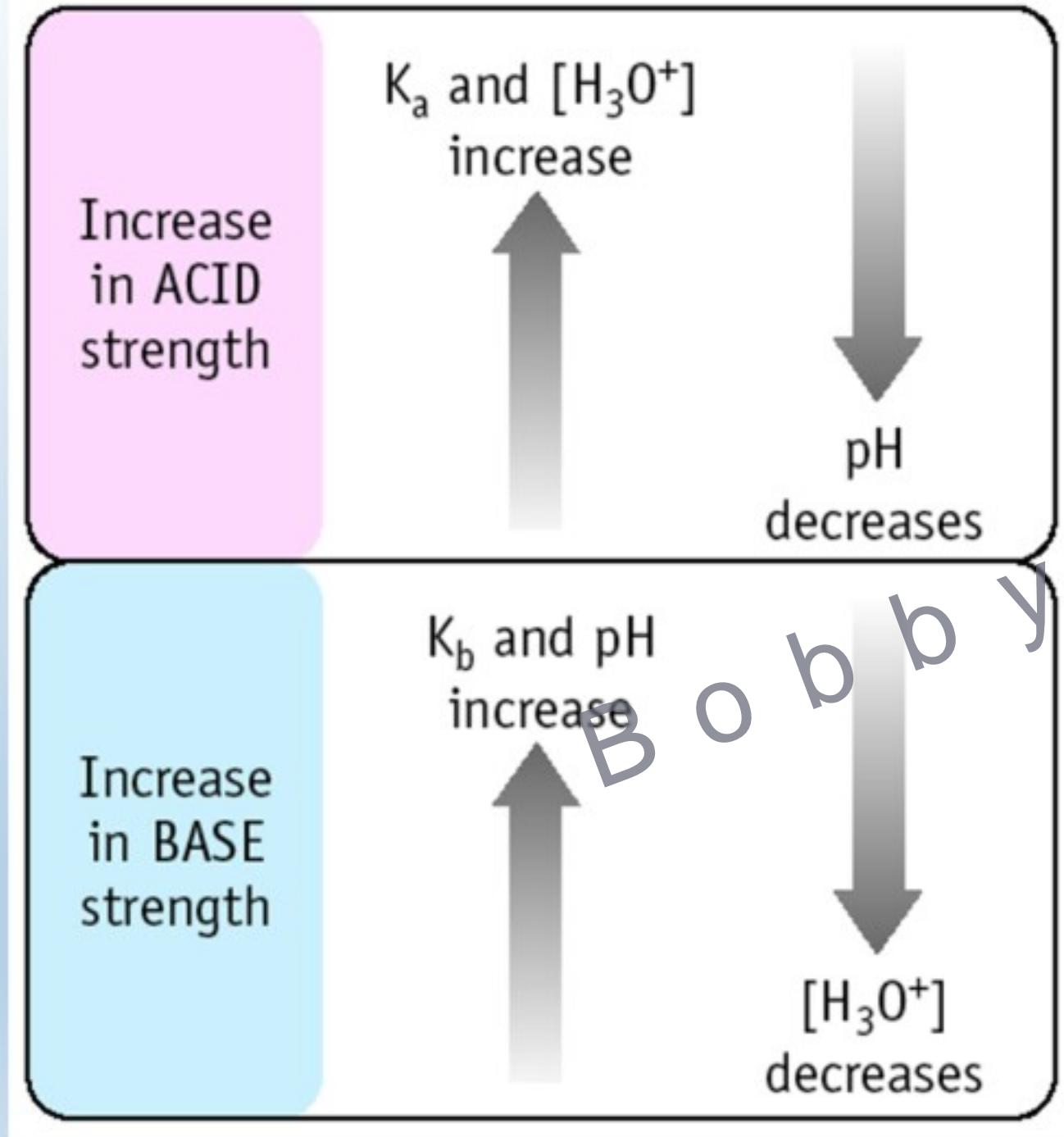
## Equilibrium Constants for Weak Bases contd.



$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

Weak base has  $K_b < 1$

Leads to small  $[OH^-]$  and a pH of 7 - 12



Relation of  
 $K_a$ ,  $K_b$ ,  
 $[H_3O^+]$  and  
pH

# Equilibria Involving A Weak Acid calculation

- Calculate the equilibrium concs. of HOAc, H<sub>3</sub>O<sup>+</sup>, OAc<sup>-</sup>, and the pH of 1.00 M HOAc.

**Step 1:** Define equilibrium concs. in Ionization constant equilibrium (ICE) table.

		$[HOAc] \rightarrow$	$b$	$b$	$y$	$s$	$a$
initial	(M)		1.00	0	0		
Change (M)			$-x$	$+x$	$+x$		
Equilib (M)			$1.00-x$	$x$	$x$		

# Equilibria Involving A Weak Acid calculation contd.

Calculate the equilibrium concs. of HOAc, H<sub>3</sub>O<sup>+</sup>, OAc<sup>-</sup>, and the pH of 1.00 M HOAc.

Step 2: Write K<sub>a</sub> expression

$$K_a = 1.8 \times 10^{-5} = \frac{[H_3O^+][OAc^-]}{[HOAc]} = \frac{x^2}{1.00 - x}$$

This is a quadratic. Solve using quadratic formula.

$$\frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

or you can make an approximation if x is very small! (Rule of thumb: 10<sup>-5</sup> or smaller is ok)

## Equilibria Involving A Weak Acid calculation contd.

- Calculate the equilibrium concs. of HOAc, H<sub>3</sub>O<sup>+</sup>, OAc<sup>-</sup>, and the pH of 1.00 M HOAc.

Step 3: Solve K<sub>a</sub> expression

$$K_a = 1.8 \times 10^{-5} = \frac{[H_3O^+][OAc^-]}{[HOAc]} = \frac{x^2}{1.00 - x}$$

First assume x is very small because K<sub>a</sub> is so small.

$$K_a = 1.8 \times 10^{-5} = \frac{x^2}{1.00}$$

# Equilibria Involving A Weak Acid calculation contd.

- Calculate the equilibrium concs. of HOAc, H<sub>3</sub>O<sup>+</sup>, OAc<sup>-</sup>, and the pH of 1.00 M HOAc.

Step 3: Solve K<sub>a</sub> approximate expression

$$K_a = 1.8 \times 10^{-5} = \frac{x^2}{1.00}$$

$$x = [H_3O^+] = [OAc^-] = 4.2 \times 10^{-3} M$$

$$pH = -\log [H_3O^+] = -\log (4.2 \times 10^{-3}) = 2.37$$

$$[HOAc] = 1.00 - 4.2 \times 10^{-3} = 0.9958 M$$

# Try the questions!!!

- **Problem 1:** Calculate the pH of a 0.0010 M solution of formic acid, HCO<sub>2</sub>H.



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

Approximate solution

$$[\text{H}_3\text{O}^+] = 4.2 \times 10^{-4} \text{ M}, \text{ pH} = 3.37$$

Exact Solution

$$[\text{H}_3\text{O}^+] = [\text{HCO}_2^-] = 3.4 \times 10^{-4} \text{ M}$$

$$[\text{HCO}_2\text{H}] = 0.0010 - 3.4 \times 10^{-4} = 0.0007 \text{ M}$$

$$\text{pH} = 3.47$$

- **Problem 2:** Calculate the pH of a 0.2 M acetic acid solution ( For Approximation and Exact)

# Equilibria Involving A Weak Base calculation

- Calculate the pH of 0.010 M  $\text{NH}_3$  solution .



$$K_b = 1.8 \times 10^{-5}$$

**Step 1.** Define equilibrium concs. in ICE table

	[ $\text{NH}_3$ ]	[ $\text{NH}_4^+$ ]	[ $\text{OH}^-$ ]
initial (M)	0.010	0	0
Change (M)	-x	x	x
Equilib (M)	$0.010 - x$	x	x

# Equilibria Involving A Weak Base

## calculation contd.

Calculate the pH of 0.010 M NH<sub>3</sub> solution



$$K_b = 1.8 \times 10^{-5}$$

**Step 2.** Solve the equilibrium expression

$$K_b = 1.8 \times 10^{-5} = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = \frac{x^2}{0.010 - x}$$

Assume x is small, so

$$x = [\text{OH}^-] = [\text{NH}_4^+] = 4.2 \times 10^{-4} \text{ M}$$

$$\text{and } [\text{NH}_3] = 0.010 - 4.2 \times 10^{-4} \approx 0.010 \text{ M}$$

The approximation is valid!!!

## Equilibria Involving A Weak Base calculation contd.

Calculate the pH of 0.010 M NH<sub>3</sub> solution



$$K_b = 1.8 \times 10^{-5}$$

**Step 3.** Calculate pH

$$[\text{OH}^-] = 4.2 \times 10^{-4} \text{ M}$$

$$\text{so } \text{pOH} = -\log [\text{OH}^-] = 3.37$$

Because pH + pOH = 14,

$$\text{pH} = 10.63$$

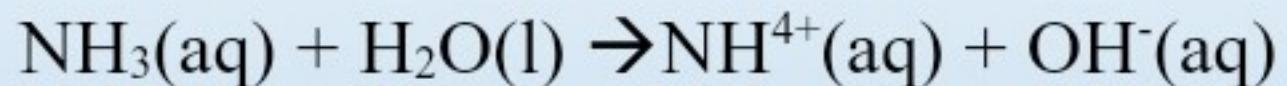
# Types of Acid/Base Reactions

Table 3: Characteristics of Acid-Base Reactions

Type	Example	Net Ionic Equation	Species Present After Equal Molar Amounts Are Mixed; pH
Strong acid + strong base	HCl + NaOH	$\text{H}_3\text{O}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightleftharpoons 2 \text{H}_2\text{O}(\ell)$	$\text{Cl}^-$ , $\text{Na}^+$ , pH = 7
Strong acid + weak base	HCl + NH <sub>3</sub>	$\text{H}_3\text{O}^+(\text{aq}) + \text{NH}_3(\text{aq}) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{H}_2\text{O}(\ell)$	$\text{Cl}^-$ , $\text{NH}_4^+$ , pH < 7
Weak acid + strong base	$\text{HCO}_2\text{H} + \text{NaOH}$	$\text{HCO}_2\text{H}(\text{aq}) + \text{OH}^-(\text{aq}) \rightleftharpoons \text{HCO}_2^-(\text{aq}) + \text{H}_2\text{O}(\ell)$	$\text{HCO}_2^-$ , $\text{Na}^+$ , pH > 7
Weak acid + weak base	$\text{HCO}_2\text{H} + \text{NH}_3$	$\text{HCO}_2\text{H}(\text{aq}) + \text{NH}_3(\text{aq}) \rightleftharpoons \text{HCO}_2^-(\text{aq}) + \text{NH}_4^+(\text{aq})$	$\text{HCO}_2^-$ , $\text{NH}_4^+$ , pH dependent on $K_a$ and $K_b$ of conjugate acid and base.

# TRY THE QUESTIONS!!!

**Problem 1:** Determine [B] in a  $1.82 \times 10^{-3}$  M solution of NH<sub>3</sub>



$$K_b = 1.8 \times 10^{-5}$$

$$\text{Answer} = 1.65 \times 10^{-3} \text{ M}$$

**Problem 2:** What is the pH of a 0.40 M ammonia solution

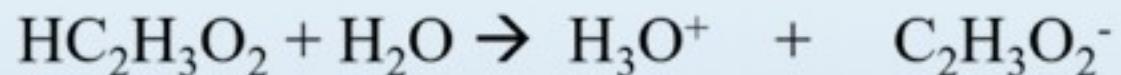
$$\text{Answer} = 11.43$$

**Problem 3 :** Calculate the pH of a 0.26 M methylamine (CH<sub>3</sub>NH<sub>2</sub>) solution

$$K_b = 4.4 \times 10^{-4}$$

$$\text{Answer} = 12.03$$

# The Ionization Constants of Acids and their Conjugate Bases Relationship



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

The conjugate base  $\text{C}_2\text{H}_3\text{O}_2^-$  supplied by sodium acetate ( $\text{CH}_3\text{COONa}$ ) solution reacts with water according to the equation



The base ionization can be written as:

$$K_b = \frac{[\text{CH}_3\text{COOH}][\text{OH}^-]}{[\text{CH}_3\text{COO}^-]}$$

The product of these two ionization constants is given by

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

# Polyprotic Acids

- Some acids contain more than one hydrogen atom that may be donated to form  $\text{H}^+$  ion. These are called *Polyprotic acids*

Examples include:

$\text{H}_2\text{SO}_4$  sulfurous acid (2  $\text{H}^+$  ions)

$\text{H}_2\text{SO}_4$  sulfuric acid (2  $\text{H}^+$  ions)

$\text{C}_2\text{H}_2\text{O}_4$  Oxalic acid (2  $\text{H}^+$  ions)

$\text{H}_3\text{PO}_4$  phosphoric acid (3  $\text{H}^+$  ions)

$\text{H}_2\text{CO}_3$  carbonic acid (2  $\text{H}^+$  ions)

# *Strengths of Acids*

Acid strength is determined by a combination of factors:

Bond polarity—the H-A bond must be polar in order for the H atom to be transferred to water as H<sup>+</sup>

The H atom in CH<sub>4</sub> is non-acidic because the C-H bond is not polar

The H-Cl bond in HCl is polar, and HCl is a strong acid

## Strengths of Acids contd.

Acid strength is determined by a combination of factors:

Bond strength—the stronger the bond, the weaker the acid—it is harder to pull away the H atom to form H<sup>+</sup>

Acid	Bond Energy	K <sub>a</sub>	Acid strength
HF	617 kJ/mol	7.2 x 10 <sup>-4</sup>	weak
HCl	427 kJ/mol	~106	strong
HBr	362 kJ/mol	~108	strong
HI	295 kJ/mol	~109	strong

# *Strengths of Acids* contd.

Oxoacids are those with a H-O-Z linkage

Electronegativity of Z—the higher the electronegativity of Z, the stronger the acid

Acid	Z	Ka
HOCl	3.0	$3.5 \times 10^{-8}$
HOBr	2.8	$2.5 \times 10^{-9}$
HOI	2.5	$2.3 \times 10^{-11}$

Bobby saw

# Strengths of Acids contd.

Oxoacids are those with a H-O-Z linkage (the same central atom but different numbers of attached groups)

The more O atoms attached to Z, the stronger the acid

Acid	O atoms	K <sub>a</sub>
HOCl	1	$3.5 \times 10^{-8}$
HOClO	2	$1.1 \times 10^{-2}$
HOClO <sub>2</sub>	3	$\sim 10^3$
HOClO <sub>3</sub>	4	$\sim 10^8$

NOTE: Concentration is not the same as strength

Concentration deals with the amount of hydronium ions in the solution, compared to the amount of water in the solution.  
→ More acid or base and less water = more concentrated  
→ More ions and less molecules = stronger

# Practice Exercises!!!

**Problem 1:** Predict the relative strengths of the oxoacids in each of the following groups

- a) HClO, HBrO, HIO
- b) HNO<sub>3</sub> and HNO<sub>2</sub>

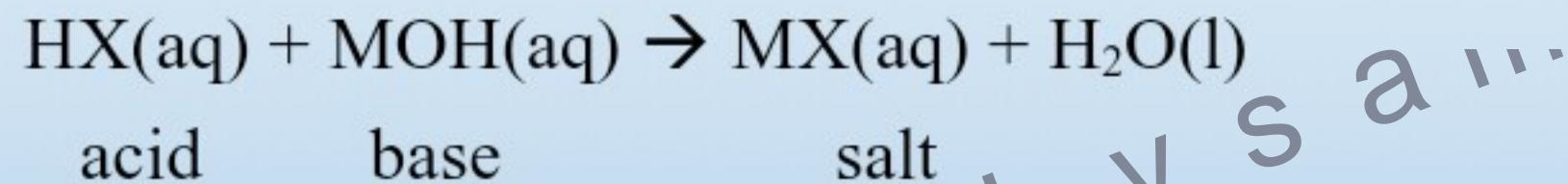
**Problem 2:** Which of the following acids is weaker H<sub>3</sub>PO<sub>3</sub> or H<sub>3</sub>PO<sub>4</sub>

**Problem 3:** Compare the strengths of the following pairs of acids

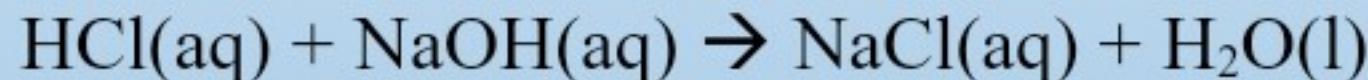
H<sub>2</sub>SO<sub>4</sub> and H<sub>2</sub>SeO<sub>4</sub>; H<sub>3</sub>PO<sub>4</sub> and H<sub>3</sub>AsO<sub>4</sub>

# Acids and Bases Properties of Salts

When an acid and a base undergo an exchange reaction, the result is a salt and water:



If a strong base is neutralized with a strong acid, the resulting solution contains only the salt



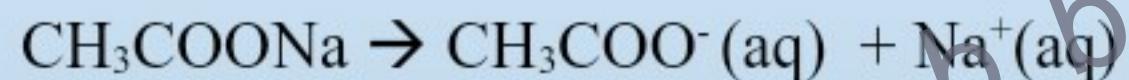
# Acids and Bases Properties of Salts contd.

## Weak acid and salt of a strong base:

If stoichiometric amounts are combined, the solution will be slightly basic

Determine pH of a solution with 0.15 M solution of sodium acetate acetic ( $\text{CH}_3\text{COONa}$ ).

**Step 1:** The sodium acetate dissociates into cation and anion; the initial conc. of the ions are equal to 0.15 M



0.15 M      0.15 M

**Step 2:** Let  $x$  be the equilibrium conc. of  $\text{CH}_3\text{COO}^-$  and  $\text{OH}^-$  ions in mol/L



Initial (M)	0.15 M	0.00	0.00
Change (M)	-x	+x	+x
Equilibrium (M)	0.15 - x	x	x

# Acids and Bases Properties of Salts contd.

Step 3: The base ionization constant

$$K_b = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]}$$
$$= \frac{x^2}{0.15-x}$$
$$x^2 / 0.15 - x = 5.6 \times 10^{-10}$$

Assume  $0.15-x = 0.15$  because  $K_b$  is very small

$$x^2 / 0.15 - x = 5.6 \times 10^{-10}$$

$$X = 9.2 \times 10^{-6}$$

## Acids and Bases Properties of Salts contd.

**Step 4:** At equilibrium

$$[\text{OH}] = 9.2 \times 10^{-6} \text{ M}$$

$$\text{pOH} = -\log (9.2 \times 10^{-6} \text{ M})$$

$$= 5.04$$

$$\text{pH} = 14.00 - 5.04$$

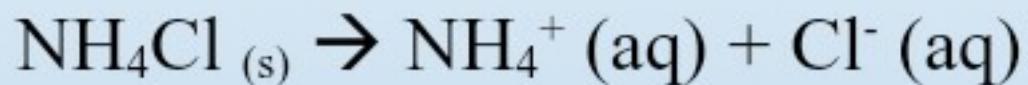
$$= 8.96$$

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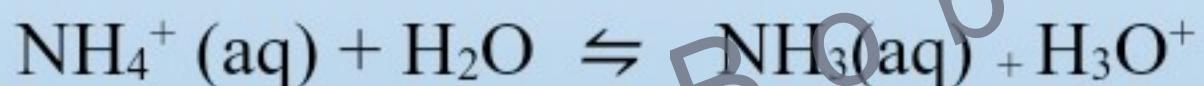
# Acids and Bases Properties of Salts contd.

Strong acid and salt of a weak base:

When a salt derived from a strong acid such as HCl and a weak base such as NH<sub>3</sub> dissolves in water, the solution becomes acidic.



The ~~b~~ is the weak conjugate acid of the weak base NH<sub>3</sub> and ionizes as follows:



Or simply

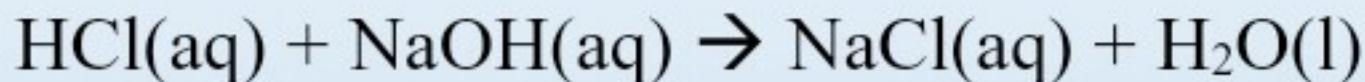


Note the pH of the solution decreases. The equilibrium for the process is given by

$$K_a = \frac{[\text{NH}_3]_{\text{H}^+} [\text{NH}_4^+]_{\text{H}^+}}{[\text{NH}_3]_{\text{H}^+} [\text{NH}_4^+]_{\text{H}^+}} = K_w K_b \quad K_w K_b = 1.0 \times 10^{-14}$$

# *Neutralization and Titration*

- Neutralization is a reaction between an acid and base to produce salt and water



- Titrations

Titrations are performed with acids and bases to determine their concentrations.

At the equivalence point, the number of moles of the acid will equal the number of moles of the base.

This indicates that the reaction has been neutralized.

- Neutralization: moles of acid = moles of base

# Assignments

1. Read up hydrolysis of metal to produce acidic solution.
2. What are indicators in acid /base titrations.
3. Read up on buffer solutions?
4. Calculate the pH of a 0.020 M  $\text{Al}(\text{NO}_3)_3$  solution.
5. What is the pH of a 0.050 M  $\text{AlCl}_3$  solution.
6. A 30 mL of 1.00 M NaOH is needed to titrate 60 mL of an HCl solution. Calculate the molarity of HCl needed to achieve this neutralisation.

Bobby saw

## Recommended Textbook

Chemistry by Raymond Chang, Published by Mc Graw Hill Higher Education  
( 7<sup>th</sup> Edition )

