

## Chapter 14: Acids, Bases and Salts Sections 14.1 – 14.5

CHM152 GCC

OpenSTAX: Chemistry

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## Why study acids & bases?

- ❖ Many household substances, including cleaning solutions and food/beverages that we consume, are acids or bases.
- ❖ In the environment, the pH of rain, water and soil can also have significant effects.
- ❖ **Acid – base reactions occurring in our body are essential for life. They are also involved in many industrial processes.**

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## Acid vs. Base Characteristics

### Acids

- Taste sour
- Corrode most metals
- Turn litmus paper red
- Have low pH's
- E.g. vinegar, lemon juice, gastric juice, soft drinks



### Bases

- Taste bitter, chalky
- Feel soapy, slippery
- Turn litmus paper blue
- Have high pH's
- E.g. milk of magnesia, bleach, ammonia, drano



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## Acid-Base Models

CHM 150/151 taught Arrhenius theory

- ♦ **Acids** produce hydrogen ions in water
  - $\text{HA (aq)} \rightarrow \text{H}^+ \text{ (aq)} + \text{A}^- \text{ (aq)}$
- ♦ **Bases** produce hydroxide ions in water
  - $\text{MOH (aq)} \rightarrow \text{M}^+ \text{ (aq)} + \text{OH}^- \text{ (aq)}$

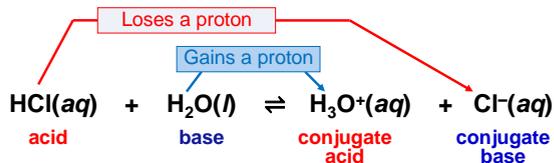
CHM 152 uses Brønsted-Lowry theory

- ♦ **Brønsted Acid:** Substance that can donate  $\text{H}^+$   
⇒ Acid loses  $\text{H}^+$
- ♦ **Brønsted Base:** Substance that can accept  $\text{H}^+$   
⇒ Base gains  $\text{H}^+$

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## Brønsted-Lowry Reactions involve the transfer of 1 $\text{H}^+$

A **conjugate base** is formed when a Brønsted acid loses a proton.



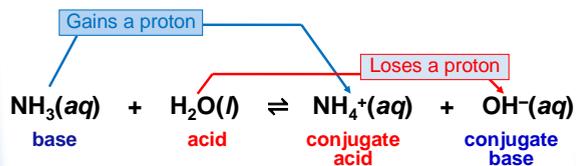
**conjugate pair:** differ by 1  $\text{H}^+$  ion

Conjugate pairs above: 1)  $\text{HCl}$  and  $\text{Cl}^-$  2)  $\text{H}_2\text{O}$  and  $\text{H}_3\text{O}^+$

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## Bronsted Acids and Bases

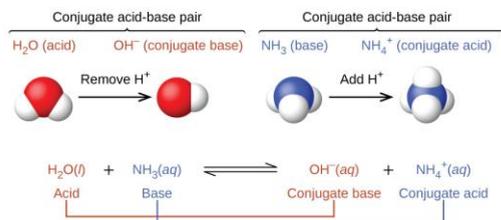
A **conjugate acid** is formed when a Brønsted base accepts a proton.



Note: Water is **amphoteric** – it can act as an acid or a base depending on what its reacting with.

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## Conjugate Pairs



## Identify Conjugate Acid/Base

a) What is the conjugate acid of  $\text{CO}_3^{2-}$ ?

b) What is the conjugate base of  $\text{H}_2\text{PO}_4^-$ ?

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## Bronsted-Lowry Reaction

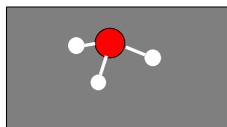
Write an equation for the dissociation of HCN in water. Identify the acid, the base, the conjugate acid, and the conjugate base.

Watch charges! Acid loses 1  $\text{H}^+$ , base gains 1  $\text{H}^+$

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## Hydronium Ions

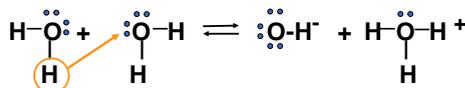
- $\text{HA}(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{A}^-(\text{aq})$
- $\text{H}^+$  is very reactive and will bond with O in  $\text{H}_2\text{O}$  to form  $\text{H}_3\text{O}^+$
- $\text{H}^+$  and  $\text{H}_3\text{O}^+$  are used interchangeably but  $\text{H}_3\text{O}^+$  is a more accurate representation.



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## Autoionization of Water

In water a  $\text{H}^+$  ion can be transferred between  $\text{H}_2\text{O}$  molecules.



water molecules      hydroxide      hydronium  
**Acid**      **Base**      **CB**      **CA**

This equilibrium process is highly reactant favored thus it gives rise to a very small equilibrium constant.

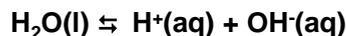
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### Equilibrium Constant for Water Ionization, $K_w$



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

Why is water omitted?



$$\text{or } K_w = [\text{H}^+][\text{OH}^-]$$

$$K_w = 1 \times 10^{-14} \text{ at } 25^\circ\text{C}$$

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### Relationship between acidity, $[\text{H}^+]$ , $[\text{OH}^-]$

$$K_w = [\text{H}^+][\text{OH}^-] = 1 \times 10^{-14}$$

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

Acidic:  $[\text{H}^+] > [\text{OH}^-]$

Basic:  $[\text{H}^+] < [\text{OH}^-]$

Example.  $[\text{H}_3\text{O}^+] = 5.6 \times 10^{-4} \text{ M}$ . What is  $[\text{OH}^-]$ ?  
Is this solution acidic or basic?

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### pH: A measure of Acidity

- pH is the - log of the  $\text{H}^+$  concentration (M)
- Scale was proposed to give us convenient numbers to work with

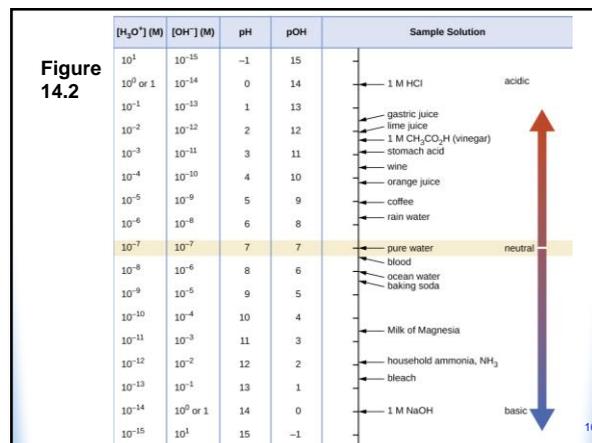
#### Relationship Between pH and Acidity

$\text{pH} < 7$ ,  $[\text{H}^+] > 1.0 \times 10^{-7} \text{ M}$       Acidic

$\text{pH} = 7$ ,  $[\text{H}^+] = 1.0 \times 10^{-7} \text{ M}$       Neutral

$\text{pH} > 7$ ,  $[\text{H}^+] < 1.0 \times 10^{-7} \text{ M}$       Basic

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### pH Scale

Memorize these equations!

$$\text{pH} = -\log [\text{H}_3\text{O}^+] \quad [\text{H}_3\text{O}^+] = 10^{-\text{pH}}$$

$$\text{pOH} = -\log [\text{OH}^-] \quad [\text{OH}^-] = 10^{-\text{pOH}}$$

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1 \times 10^{-14}$$

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

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### pH Sig Figs

2 sig figs

$$[\text{H}_3\text{O}^+] = 5.6 \times 10^{-4} \text{ M} \quad \text{2 decimal places}$$

$$\text{pH} = -\log (5.6 \times 10^{-4}) = 3.25$$

3 related to exponent; tells us acidity

Find  $[\text{H}_3\text{O}^+]$ , pH, & pOH if  $[\text{OH}^-] = 9.8 \times 10^{-9} \text{ M}$     2 sf

$$\text{pOH} = -\log [\text{OH}^-] = -\log 9.8 \times 10^{-9} = 8.01 \text{ (2 dec pl)}$$

$$\text{pH} = 14 - \text{pOH} = 14 - 8.01 = 5.99 \text{ (2 dec pl)}$$

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-5.99} = 1.0 \times 10^{-6} \text{ M} \text{ (2 sf)}$$

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## Complete the following table

pH	[H <sup>+</sup> ] M	[OH <sup>-</sup> ] M	pOH	A, B, N?
	1.0x10 <sup>-9</sup>			
4.815				
			2.30	

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## Strong Acids

A **strong acid** ionizes completely in water to form H<sub>3</sub>O<sup>+</sup> ions (H<sup>+</sup>). SA's are strong electrolytes.



- A one way arrow is used since this reaction is complete, all of the HNO<sub>3</sub> molecules break apart to form H<sub>3</sub>O<sup>+</sup> and NO<sub>3</sub><sup>-</sup> ions.
- These acids have very weak conjugate bases
- For SA, eq lies very far to the right side!

**KNOW 7 Strong acids! Memorize these!**

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

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## Strong Bases

A **strong base** ionizes completely in water to form OH<sup>-</sup> ions. SB's are strong electrolytes!

SB dissociates in water; it doesn't react with it:



- This reaction also goes to completion  
⇒ SB's completely dissociate into ions

**Strong Bases: These Group 1A & 2A hydroxides**

LiOH, NaOH, KOH, RbOH, CsOH, Ca(OH)<sub>2</sub>,

Ba(OH)<sub>2</sub>, Sr(OH)<sub>2</sub> **Memorize the 8 SB's!**

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## Weak Acids

- A **weak acid** ionizes only to a small extent in H<sub>2</sub>O.
- WA's are weak electrolytes!

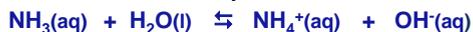


- A two-way arrow is used since this reaction is not complete but instead consists of an equilibrium mixture of HCN molecules, H<sub>3</sub>O<sup>+</sup> and CN<sup>-</sup> ions
- Equilibrium lies to the left since most of the acid molecules have not ionized.
- Common WA's: HF, HNO<sub>2</sub>, HCN, H<sub>3</sub>PO<sub>4</sub>, H<sub>2</sub>CO<sub>3</sub>, carboxylic acids like CH<sub>3</sub>COOH and HCOOH

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## Weak Base

A **weak base** ionizes only to a small extent in H<sub>2</sub>O. WB's are weak electrolytes!

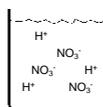


- Equilibrium lies to the left since most of the ammonia molecules have not accepted protons to form NH<sub>4</sub><sup>+</sup> and OH<sup>-</sup> ions.
- Common examples of WB's are ammonia and amines (e.g. CH<sub>3</sub>NH<sub>2</sub>)

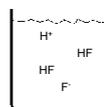
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## SA, WA, SB, WB solutions

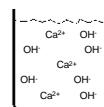
Draw the following in water: HNO<sub>3</sub>, HF, Ca(OH)<sub>2</sub>, and NH<sub>3</sub>. Assume you have 3 formula units or molecules for each substance.



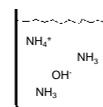
SA



WA



SB



WB

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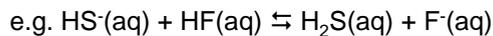
Figure 14.8. Relative Strength of Acids & Bases

Acid			Base	
perchloric acid	HClO <sub>4</sub>	Undergo complete acid ionization in water	ClO <sub>4</sub> <sup>-</sup>	perchlorate ion
sulfuric acid	H <sub>2</sub> SO <sub>4</sub>		HSO <sub>4</sub> <sup>-</sup>	hydrogen sulfate ion
hydrogen iodide	HI		I <sup>-</sup>	iodide ion
hydrogen bromide	HBr		Br <sup>-</sup>	bromide ion
hydrogen chloride	HCl		Cl <sup>-</sup>	chloride ion
nitric acid	HNO <sub>3</sub>		NO <sub>3</sub> <sup>-</sup>	nitrate ion
hydronium ion	H <sub>3</sub> O <sup>+</sup>		H <sub>2</sub> O	water
hydrogen sulfate ion	HSO <sub>4</sub> <sup>-</sup>		SO <sub>4</sub> <sup>2-</sup>	sulfate ion
phosphoric acid	H <sub>3</sub> PO <sub>4</sub>		H <sub>2</sub> PO <sub>4</sub> <sup>-</sup>	dihydrogen phosphate ion
hydrogen fluoride	HF		F <sup>-</sup>	fluoride ion
nitrous acid	HNO <sub>2</sub>	NO <sub>2</sub> <sup>-</sup>	nitrite ion	
acetic acid	CH <sub>3</sub> CO <sub>2</sub> H	CH <sub>3</sub> CO <sub>2</sub> <sup>-</sup>	acetate ion	
carbonic acid	H <sub>2</sub> CO <sub>3</sub>	HCO <sub>3</sub> <sup>-</sup>	hydrogen carbonate ion	
hydrogen sulfide	H <sub>2</sub> S	HS <sup>-</sup>	hydrogen sulfide ion	
ammonium ion	NH <sub>4</sub> <sup>+</sup>	NH <sub>3</sub>	ammonia	
hydrogen cyanide	HCN	CN <sup>-</sup>	cyanide ion	
hydrogen carbonate ion	HCO <sub>3</sub> <sup>-</sup>	CO <sub>3</sub> <sup>2-</sup>	carbonate ion	
water	H <sub>2</sub> O	OH <sup>-</sup>	hydroxide ion	
hydrogen sulfide ion	HS <sup>-</sup>	S <sup>2-</sup>	sulfide ion	
ethanol	C <sub>2</sub> H <sub>5</sub> OH	C <sub>2</sub> H <sub>5</sub> O <sup>-</sup>	ethoxide ion	
ammonia	NH <sub>3</sub>	NH <sub>2</sub> <sup>-</sup>	amide ion	
hydrogen	H <sub>2</sub>	H <sup>-</sup>	hydride ion	
methane	CH <sub>4</sub>	CH <sub>3</sub> <sup>-</sup>	methyl ion	
		Do not undergo acid ionization in water		
		Undergo complete base ionization in water		

## Conjugate Acid-Base Pair Trends

1. A stronger acid loses its proton more readily than a weaker acid and a stronger base gains a proton more readily than a weaker base.
2. The stronger the acid, the weaker its conjugate base. Likewise, the stronger the base, the weaker its conjugate acid.
3. Proton transfer reactions proceed from the stronger acid to the stronger base. Thus, the equilibria lies on the side of the weaker acid-base pair.

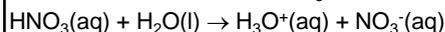
## Strengths of Acids/Bases



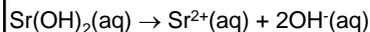
- From Table on slide
- Which acid is stronger?
- Which base is stronger?
- Which side does equilibria lie on?

## pH of Strong Acids/Bases

Calculate pH of 0.103 M HNO<sub>3</sub> solution.

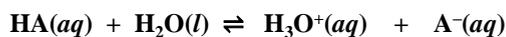


Calculate pH of 0.020 M Sr(OH)<sub>2</sub> soln.



## Weak Acids; Acid-Dissociation Constants (K<sub>a</sub>)

The ionization of a weak acid, HA, in water:



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

K<sub>a</sub> is called the **acid dissociation constant**.

The larger the value of K<sub>a</sub>, the stronger the acid.

Solution (at 25 °C)	K <sub>a</sub>	pH
0.10 M HF	3.5 x 10 <sup>-4</sup>	2.23
0.10 M CH <sub>3</sub> COOH	1.8 x 10 <sup>-5</sup>	2.87

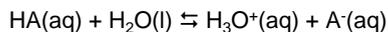
## Weak Acids; Acid-Ionization Constants (K<sub>a</sub>)

Acids	K <sub>a</sub> values at 25 °C
HF	3.5 x 10 <sup>-4</sup>
HNO <sub>2</sub>	4.6 x 10 <sup>-4</sup>
HNCO	2.0 x 10 <sup>-4</sup>
HCO <sub>2</sub> H	1.8 x 10 <sup>-4</sup>
CH <sub>3</sub> CO <sub>2</sub> H	6.3 x 10 <sup>-5</sup>
HClO	3.5 x 10 <sup>-8</sup>
HBrO	2.8 x 10 <sup>-9</sup>
HCN	4.9 x 10 <sup>-10</sup>
C <sub>6</sub> H <sub>5</sub> COOH	6.3 x 10 <sup>-5</sup>

**NOTE:** The reported constants vary – make sure to use the value provided in the question.

Additional K<sub>a</sub> values found in appendix in the OpenSTAX book.

## Equilibria of Weak Acids



- $\text{pK}_a = -\log K_a$ : stronger acid = lower  $\text{pK}_a$

$$\text{percent ionization} = \frac{[\text{H}_3\text{O}^+]_{\text{eq}}}{[\text{HA}]_0} \times 100\%$$

- For a stronger acid, indicate if the following will be higher or lower:
  - $[\text{H}_3\text{O}^+]$  ↑
  - pH ↓
  - % ionization ↑
  - $K_a$  ↑
  - $\text{pK}_a$  ↓

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## Weak Acid Problems

- Set up ICE tables &  $K_a$
- With small K values, assume x will be small to simplify math.

WA problem types:

- Given  $[\text{HA}]_i$  and  $K_a$ , find all equilibrium concentrations and pH
- Given  $[\text{HA}]_i$  and pH, find all equilibrium concentrations and  $K_a$
- Given  $[\text{HA}]_i$  and % dissociation, find eq concentrations, pH and  $K_a$

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## Find pH of Weak Acid

Find the equilibrium concentrations of all substances and the pH of a 0.50 M solution of HClO (hypochlorous acid).  $K_a = 3.5 \times 10^{-8}$

- Set up ICE Table

	HClO	+ H <sub>2</sub> O	⇌	H <sub>3</sub> O <sup>+</sup>	+ ClO <sup>-</sup>
I					
C					
E					

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## Find pH of Weak Acid

- Write  $K_a$  expression for this reaction
- Plug eq terms into  $K_a$  and solve for x:
- $x =$ 
  - Check approximation:  $(x / [\text{HA}]_i) \times 100 < 5\%$

$$x = [\text{H}_3\text{O}^+] = [\text{OCl}^-] =$$

$$[\text{HOCl}] = 0.50 - x =$$

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

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## Find $K_a$ for weak acid

Find the  $K_a$  of a 1.25 M solution of nitrous acid,  $\text{HNO}_2$ . The pH of this solution is measured to be 1.62.

Write hydrolysis reaction &  $K_a$  expression!

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} =$$

(keep 2 sf since pH has 2 dec pl)

$$[\text{NO}_2^-] =$$

$$[\text{HNO}_2] =$$

\*Plug in x ( $\text{H}_3\text{O}^+$ ) since we know value from pH!

$$K_a =$$

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## Percent Ionization

% dissociation is the same calculation as checking assumption of a small x.

$$\text{\% dissociation} = \left( \frac{[\text{H}_3\text{O}^+]_{\text{eq}}}{[\text{HA}]_i} \right) \times 100$$

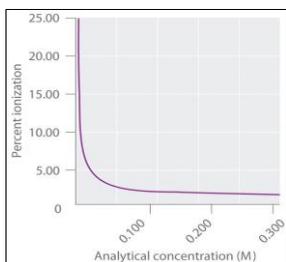
- Calculate the percent dissociation from the previous problem.

$$\text{\%dissociation} =$$

Will the percent dissociated increase or decrease as an acid is diluted?

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## Effect of Dilution on % Ionization



As we add water, we can dissociate more.  
More dilute = more dissociation.

CC-BY-NA-3.0: <http://2012books.lardbucket.org/books/principles-of-general-chemistry-v1.0/s20-aqueous-acid-base-equilibriums.html>

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## Find $K_a$ for weak acid

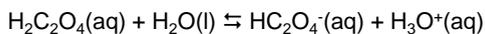
A 0.200 M solution of a weak acid, HA, is 9.4 percent dissociated. Using this information, calculate  $[\text{H}_3\text{O}^+]$ ,  $[\text{A}^-]$ ,  $[\text{HA}]_{\text{eq}}$ , and  $K_a$ .

Answer:  $[\text{H}_3\text{O}^+] = 0.019 \text{ M}$ ,  $[\text{A}^-] = 0.019 \text{ M}$ ,  
 $[\text{HA}] = 0.181 \text{ M}$ , and  $K_a = 2.0 \times 10^{-3}$

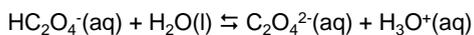
See Worksheet key

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## Polyprotic Acids



$$K_{a1} = \frac{[\text{HC}_2\text{O}_4^-][\text{H}_3\text{O}^+]}{[\text{H}_2\text{C}_2\text{O}_4]} = 5.9 \times 10^{-2}$$



$$K_{a2} = \frac{[\text{C}_2\text{O}_4^{2-}][\text{H}_3\text{O}^+]}{[\text{HC}_2\text{O}_4^-]} = 6.4 \times 10^{-5}$$

In general,  $K_{a1} > K_{a2} > K_{a3} \dots$  Why?

- pH of a polyprotic acid solution primarily arises from  $[\text{H}^+]$  formed in the 1<sup>st</sup> step.

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## Weak Bases; Base-Ionization Constants ( $K_b$ )

The ionization of a *weak base* in water:



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

$K_b$  is called the *base dissociation constant*.

The larger the value of  $K_b$ , the stronger the base.

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## Weak Bases

Similar to WA problems but base hydrolysis reaction forms  $\text{OH}^-$  (this is x!) and use  $K_b$

- $\text{NH}_3(\text{aq}) + \text{H}_2\text{O(l)} \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$

- $K_b$  Values for Weak Bases in appendix
- Why are amines (derivatives of  $\text{NH}_3$ ) basic?

$\Rightarrow$  Lone e- pair on N attracts H from water.  
N pulls  $\text{H}^+$  off of water.

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## Weak Bases; $K_b$ Constants

Base	Formula	$K_b$	$\text{p}K_b$
phosphate ion	$\text{PO}_4^{3-}$	$2.1 \times 10^{-2}$	1.68
dimethylamine	$(\text{CH}_3)_2\text{NH}$	$5.4 \times 10^{-4}$	3.27
methylamine	$\text{CH}_3\text{NH}_2$	$4.6 \times 10^{-4}$	3.34
trimethylamine	$(\text{CH}_3)_3\text{N}$	$6.3 \times 10^{-5}$	4.20
ammonia	$\text{NH}_3$	$1.8 \times 10^{-5}$	4.75
pyridine	$\text{C}_5\text{H}_5\text{N}$	$1.7 \times 10^{-9}$	8.77
aniline	$\text{C}_6\text{H}_5\text{NH}_2$	$7.4 \times 10^{-10}$	9.13

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## Weak Bases

Calculate the pH of a 0.50 M dimethylamine ((CH<sub>3</sub>)<sub>2</sub>NH) solution.  $K_b = 5.4 \times 10^{-4}$

Write reaction, set up ICE, and  $K_b$ .

x =

Check approximation!

Plug in x!

$$[(\text{CH}_3)_2\text{NH}_2^+] = [\text{OH}^-] =$$

$$[(\text{CH}_3)_2\text{NH}] =$$

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## Weak Bases

2 ways to find pH

1) Convert [OH<sup>-</sup>] to [H<sub>3</sub>O<sup>+</sup>]:

$$x = [\text{OH}^-] =$$

$$[\text{H}_3\text{O}^+] = K_w / [\text{OH}^-] =$$

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

2) Convert [OH<sup>-</sup>] to pOH

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

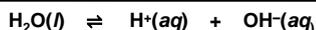
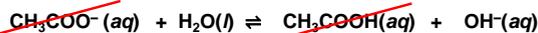
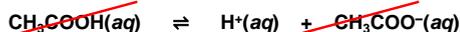
$$\text{pH} = 14 - \text{pOH} =$$

**this is easiest way**

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## Relationship between $K_a$ and $K_b$

The ionization constant of a weak acid ( $K_a$ ) and the ionization constant of a weak base ( $K_b$ ) are related:



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

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

$$K_a \times K_b = K_w$$

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## Calculate $K_b$

What is the  $K_b$  for acetate if  $K_a$  is  $1.8 \times 10^{-5}$  for acetic acid?

For conjugate pairs:  $K_a \cdot K_b = 10^{-14}$

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## Properties of Salts

• Acid + Base → Salt + Water



• Salts are ionic products of acid-base reaction.

• **Cation in salt:** comes from the base

• **Anion in salt:** comes from the acid

• Salts can be neutral, acidic, or basic depending on the strength of the acid or base from which they're made.

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## General Ion Categories

1) **Acidic ions:** Most cations (except 1A & 2A)

E.g.  $\text{NH}_4^+$ ,  $\text{Al}^{3+}$ ,  $\text{Cu}^{2+}$

2) **Neutral ions:** cations from strong bases and anions from strong acids

**Cations:** Group 1A & 2A (except Be)

**Anions:**  $\text{NO}_3^-$ ,  $\text{Cl}^-$ ,  $\text{Br}^-$ ,  $\text{I}^-$ ,  $\text{ClO}_4^-$ ,  $\text{ClO}_3^-$  (except  $\text{HSO}_4^-$ )

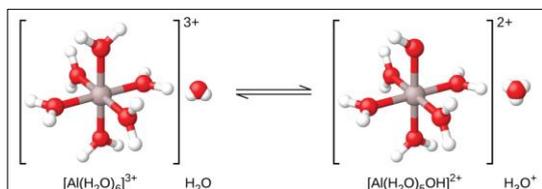
3) **Basic ions:** Most anions derived from weak acids.

E.g.  $\text{F}^-$ ,  $\text{CN}^-$ ,  $\text{CH}_3\text{COO}^-$ ,  $\text{NO}_2^-$

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## Some Metal Ions form Acidic Salts

Small, highly charged metal ions can react with water to produce acidic solution. Easier to break O-H bonds and form hydronium's.



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## Types of Salts

### Neutral Salts – salt formed in SA-SB rxn

- ♦ cation from SB and anion from SA are neutral ions
- ♦ Neutral ions don't hydrolyze!
- ♦ E.g. NaCl



### Acidic Salts – typically contain acidic cation and neutral anion (from SA)

Cation is

- 1) conjugate acid of WB (e.g.  $\text{NH}_4^+$ ) or
- 2) small, highly charged metal ion (e.g.  $\text{Al}^{3+}$ )

E.g.  $\text{NH}_4\text{Cl}$



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## Types of Salts

### Basic Salts – typically contain basic anion (from weak acid) and neutral cation (from SB)

- Anion is conjugate base of WA (e.g.  $\text{F}^-$ )
- E.g. NaF



### Salts containing Acidic cations and basic anions

- Compare  $K_a$  and  $K_b$  values to tell if its acidic or basic. (If  $K_a$  is larger, salt is acidic and visa versa.)
- ♦ E.g.  $\text{NH}_4\text{NO}_2$

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## Classify Salts

- Classify the following salts as acidic, basic, or neutral. For acidic or basic salts, write the reaction of hydrolysis.

♦ KBr

♦ NaF

♦ LiCN

♦  $\text{NH}_4\text{Cl}$

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## Example Salt Problems

- Is KF an acidic, basic or neutral salt?
  - ♦ Write the hydrolysis reaction and calculate the pH of a 0.10 M KF solution.  $K_a(\text{HF}) = 3.5 \times 10^{-4}$

Is  $\text{NH}_4\text{Cl}$  acidic, basic, or neutral?

- ♦ Write the hydrolysis reaction and calculate the pH of a 0.10 M  $\text{NH}_4\text{Cl}$  solution.
- $K_b(\text{NH}_3) = 1.8 \times 10^{-5}$

See Worksheet key

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## Factors that Affect Acid Strength

**Acid strength** is a measure of how completely it ionizes.



### 1) Acid strength ↑ as strength of H-X bond ↓

- easier to break weak H-X bonds

### 2) Acid strength ↑ as polarity of H-X bond ↑

- Polarity ↑ when X is more electronegative (easier to ionize)



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## Strength Hydrohalic acids: HF << HCl < HBr < HI

- Why is HF a weak acid?
- Determined by the strength of H-X bond.
- H-F has strongest bonds (F is small so its closer to H)
- Acid Strength ↑ as Bond Strength ↓

Relative Acid Strength	HF	<	HCl	<	HBr	<	HI
H-X Bond Energy (kJ/mol)	570		432		366		298
$K_a$	$6.3 \times 10^{-4}$		$1.3 \times 10^6$		$7.9 \times 10^8$		$2.0 \times 10^9$
$pK_a$	3.20		-6.1		-8.9		-9.3

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## Acid Strength & Bond Polarity

As we move to the right across a row, the acid strength is related to the **polarity of the H-X bond**.

C has the lowest electronegativity, so C-H is least polar and least acidic. F is most electronegative, so F-H is most polar, and most acidic (easiest to ionize).

Figure 14.13

14	15	16	17
6 CH <sub>4</sub> Neither acid nor base	7 NH <sub>3</sub> Weak base $K_b = 1.8 \times 10^{-5}$	8 H <sub>2</sub> O Neutral	9 HF Weak acid $K_a = 6.8 \times 10^{-4}$
14 SiH <sub>4</sub> Neither acid nor base	15 PH <sub>3</sub> Very weak base $K_b = 4.4 \times 10^{-28}$	16 H <sub>2</sub> S Weak acid $K_a = 9.5 \times 10^{-8}$	17 HCl Strong acid

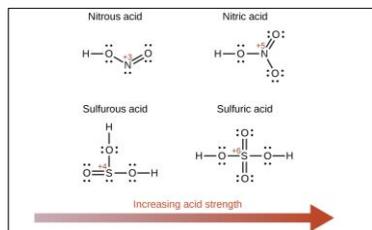
↑ ultimate proton donation  
↑ ultimate proton donation

← Increasing acid strength  
→ Increasing base strength

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## Oxoacids – Common Examples

**Oxoacids:** An oxoacid contains hydrogen, oxygen, and a central nonmetal atom (acids made from polyatomic ions).



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## Oxoacid Strength in Same Group

**1) Oxoacids having different central atoms that are from the same group.**

Acid strength ↑ as electronegativity ↑

Cl is more electronegative so O—H bond is more polar.

Thus HOCl > HOBr > HOI

HOX	Electronegativity of X	$K_a$	$pK_a$
HOCl	3.0	$4.0 \times 10^{-8}$	7.40
HOBr	2.8	$2.8 \times 10^{-9}$	8.55
HOI	2.5	$3.2 \times 10^{-11}$	10.5

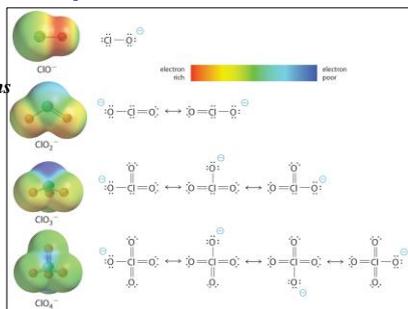
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**2) Oxoacids with same central atom, but different numbers of O atoms.**

• Acidity ↑ as the # of O's ↑

As the  $e^-$  density on the oxygen atoms decreases, so does their affinity for a proton, making the anion less basic.



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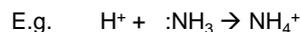
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## Lewis Acids and Bases

- Broadest acid base definition since Lewis acids don't have to contain H<sup>+</sup>.
- Looks at electron transfer instead of H<sup>+</sup> transfer

A **Lewis base** is a an electron pair donor.

A **Lewis acid** is an electron pair acceptor.



⇒ Ammonia donates the electron pair to H<sup>+</sup> to make the bond; H<sup>+</sup> accepts the electron pair

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