

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
 - $\text{MOH (aq)} \rightarrow \text{M}^+ \text{ (aq)} + \text{OH}^- \text{ (aq)}$

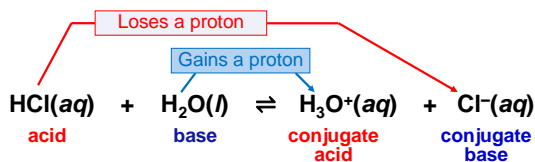
CHM 152 uses Brønsted-Lowry theory

- ♦ **Brønsted Acid:** Substance that can donate H^+
⇒ Acid loses H^+
- ♦ **Brønsted Base:** Substance that can accept H^+
⇒ Base gains H^+

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Brønsted-Lowry Reactions involve the transfer of 1 H^+

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



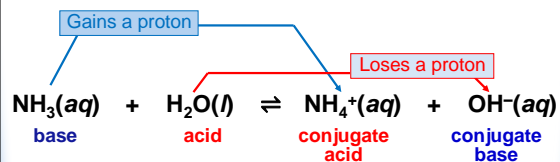
conjugate pair: differ by 1 H^+ ion

Conjugate pairs above: 1) HCl and Cl^- 2) H_2O and H_3O^+

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Brønsted Acids and Bases

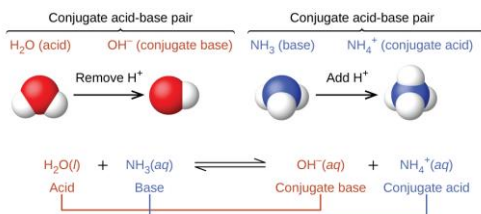
A **conjugate acid** is formed when a Brønsted base accepts a proton.



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 CO_3^{2-} ?

b) What is the conjugate base of H_2PO_4^- ?

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.

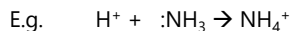
Watch charges! Acid loses 1 H^+ , base gains 1 H^+

Lewis Acids and Bases

- Broadest acid base definition since Lewis acids don't have to contain H^+ .
- Looks at electron transfer instead of H^+ transfer
- Organic CHM primarily uses this AB theory.

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

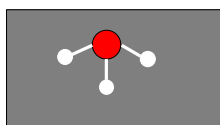
A **Lewis acid** is an electron pair acceptor.



⇒ Ammonia donates the electron pair to H^+ to make the bond; H^+ accepts the electron pair

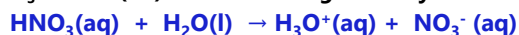
Hydronium Ions

- $\text{HA}(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{A}^-(\text{aq})$
- H^+ is very reactive and will bond with O in H_2O to form H_3O^+
- H^+ and H_3O^+ are used interchangeably but H_3O^+ is a more accurate representation.



Strong Acids

A **strong acid** ionizes completely in water to form H_3O^+ ions (H^+). SA's are strong electrolytes.



- ♦ A one way arrow is used since this reaction is complete, all of the HNO_3 molecules break apart to form H_3O^+ and NO_3^- 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, HNO₃, H₂SO₄, HClO₄, HClO₃

Strong Bases

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

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



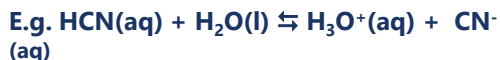
- This reaction also goes to completion
 \Rightarrow SB's completely dissociate into ions

Strong Bases: These Group 1A & 2A hydroxides
 LiOH , NaOH , KOH , RbOH , CsOH , Ca(OH)_2 ,
 Ba(OH)_2 , Sr(OH)_2 **Memorize the 8 SB's!**

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Weak Acids

- A **weak acid** ionizes only to a small extent in H_2O .
- 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_3O^+ and CN^- ions
- Equilibrium lies to the left since most of the acid molecules have not ionized.
- Common WA's: HF , HNO_2 , HCN , H_3PO_4 , H_2CO_3 , carboxylic acids like CH_3COOH and HCOOH

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Weak Base

A **weak base** ionizes only to a small extent in H_2O . WB's are weak electrolytes!

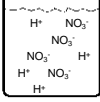

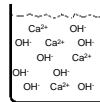
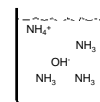


- Equilibrium lies to the left since most of the ammonia molecules have not accepted protons to form NH_4^+ and OH^- ions.
- Common examples of WB's are ammonia and amines (e.g. CH_3NH_2)

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SA, WA, SB, WB solutions

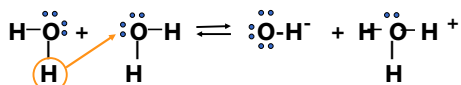
Draw the following in water: HNO_3 , HF , Ca(OH)_2 , and NH_3 . Assume you have 4 formula units or molecules for each substance.

			
SA	WA	SB	WB
All ions: H^+ , A^- ions	mostly HA molecules; a few H^+ , A^- ions	All ions: OH^- , M^+ ions	mostly B molecules; a few OH^- , HB^- ions

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Autoionization of Water

In water a H^+ ion can be transferred between H_2O 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.

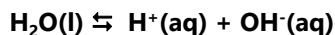
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Equilibrium Constant for Water Ionization, K_w



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

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, [H⁺], [OH⁻]

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

Neutral: [H⁺] = [OH⁻] = 1.0x10⁻⁷ M

Acidic: [H⁺] > [OH⁻]

Basic: [H⁺] < [OH⁻]

Example. [H₃O⁺] = 5.6 x 10⁻⁴ M. What is [OH⁻]?
Is this solution acidic or basic?

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pH: A measure of Acidity

- pH is the - log of the H⁺ concentration (M)
- Scale was proposed to give us convenient numbers to work with

Relationship Between pH and Acidity

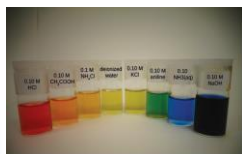
pH < 7, [H⁺] > 1.0x10⁻⁷ M Acidic

pH = 7, [H⁺] = 1.0x10⁻⁷ M
Neutral

pH > 7, [H⁺] < 1.0x10⁻⁷ M Basic

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Ways to measure pH



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Figure 14.2

[H ₃ O ⁺] (M)	[OH ⁻] (M)	pH	pOH	Sample Solution
10 ¹	10 ⁻¹⁵	-1	15	
10 ⁰ or 1	10 ⁻¹⁴	0	14	1 M HCl
10 ⁻¹	10 ⁻¹³	1	13	gastric juice
10 ⁻²	10 ⁻¹²	2	12	lime juice
10 ⁻³	10 ⁻¹¹	3	11	1 M CH ₃ CO ₂ H (vinegar)
10 ⁻⁴	10 ⁻¹⁰	4	10	stomach acid
10 ⁻⁵	10 ⁻⁹	5	9	wine
10 ⁻⁶	10 ⁻⁸	6	8	orange juice
10 ⁻⁷	10 ⁻⁷	7	7	coffee
10 ⁻⁸	10 ⁻⁶	8	6	rain water
10 ⁻⁹	10 ⁻⁵	9	5	pure water
10 ⁻¹⁰	10 ⁻⁴	10	4	blood
10 ⁻¹¹	10 ⁻³	11	3	ocean water
10 ⁻¹²	10 ⁻²	12	2	baking soda
10 ⁻¹³	10 ⁻¹	13	1	Milk of Magnesia
10 ⁻¹⁴	10 ⁰ or 1	14	0	household ammonia, NH ₃
10 ⁻¹⁵	10 ¹	15	-1	bleach
				1 M NaOH

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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 [H₃O⁺], pH, & pOH if [OH⁻] = 9.8x10⁻⁹ M 2 sf

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

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

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

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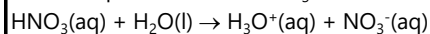
Complete the following table

pH	[H ⁺] M	[OH ⁻] M	pOH	A, B, N?
	1.0x10 ⁻⁹			
4.815				
			2.30	

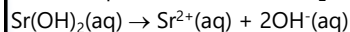
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pH of Strong Acids/Bases

Calculate pH of 0.103 M HNO₃ solution.



Calculate pH of 0.020 M Sr(OH)₂ soln.



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Figure 14.8. Relative Strength of Acids & Bases

Acid		Base	
perchloric acid	HClO ₄	ClO ₄ ⁻	perchlorate ion
sulfuric acid	H ₂ SO ₄	HSO ₄ ⁻	hydrogen sulfate ion
hydrogen iodide	HI	I ⁻	iodide ion
hydrogen bromide	HBr	Br ⁻	bromide ion
hydrogen chloride	HCl	Cl ⁻	chloride ion
nitric acid	HNO ₃	NO ₃ ⁻	nitrate ion
hydronium ion	H ₃ O ⁺	H ₂ O	water
hydrogen sulfate ion	HSO ₄ ⁻	SO ₄ ²⁻	sulfate ion
phosphoric acid	H ₃ PO ₄	H ₂ PO ₄ ⁻	dihydrogen phosphate ion
hydrogen fluoride	HF	F ⁻	fluoride ion
nitrous acid	HNO ₂	NO ₂ ⁻	nitrite ion
acetic acid	CH ₃ CO ₂ H	CH ₃ CO ₂ ⁻	acetate ion
carbonic acid	H ₂ CO ₃	HCO ₃ ⁻	hydrogen carbonate ion
hydrogen sulfide	H ₂ S	HS ⁻	hydrogen sulfide ion
ammonium ion	NH ₄ ⁺	NH ₃	ammonia
hydrogen cyanide	HCN	CN ⁻	cyanide ion
hydrogen carbonate ion	HCO ₃ ⁻	CO ₃ ²⁻	carbonate ion
water	H ₂ O	OH ⁻	hydroxide ion
hydrogen sulfide ion	HS ⁻	S ²⁻	sulfide ion
ethanol	C ₂ H ₅ OH	C ₂ H ₅ O ⁻	ethoxide ion
ammonia	NH ₃	NH ₂ ⁻	amide ion
hydrogen	H ₂	H ⁻	hydride ion
methane	CH ₄	CH ₃ ⁻	methide ion

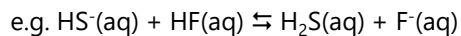
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Conjugate Acid-Base Pair Trends

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.

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Strengths of Acids/Bases



From Table 14.8

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

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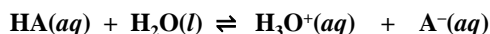
Which has higher pH?

- 0.10 M HCl or 0.10 M HCOOH
- 0.10 M HCl or 0.0010 M HCl
- 0.10 M NaOH or 0.10 M Ca(OH)₂

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Weak Acids; Acid-Dissociation Constants (K_a)

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



omit H_2O since it's a liquid

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

K_a is called the **acid dissociation constant**.

The larger the value of K_a , the stronger the acid.

Solution (at 25 °C)	K_a	pH
0.10 M HF	3.5×10^{-4}	2.23
0.10 M CH_3COOH	1.8×10^{-5}	2.87

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Weak Acids; Acid-Ionization Constants (K_a)

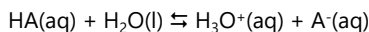
Acids	K_a values at 25 °C
HF	3.5×10^{-4}
HNO_2	4.6×10^{-4}
HNCO	2.0×10^{-4}
HCO_2H	1.8×10^{-4}
$\text{CH}_3\text{CO}_2\text{H}$	1.8×10^{-5}
HClO	3.5×10^{-8}
HBrO	2.8×10^{-9}
HCN	4.9×10^{-10}
$\text{C}_6\text{H}_5\text{COOH}$	6.3×10^{-5}

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

Additional K_a values found in appendix in the OpenSTAX book.

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Equilibria of Weak Acids



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

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

- As acid strength \uparrow , indicate if each is higher or lower:
 - $[\text{H}_3\text{O}^+]$
 - pH
 - % ionization
 - K_a
 - $\text{p}K_a$

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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 ₂ O	\rightleftharpoons	H ₃ O ⁺	+ ClO ⁻
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:

$$\begin{aligned}
 & \bullet X = \\
 & \quad \bullet \text{Check approximation: } (x / [\text{HA}]_i) \times 100 < 5\% \\
 & x = [\text{H}_3\text{O}^+] = [\text{ClO}^-] = \\
 & \quad [\text{HClO}] = 0.50 - x = \\
 & \quad \text{pH} = -\log [\text{H}_3\text{O}^+] =
 \end{aligned}$$

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Find K_a for weak acid

Find the K_a of a 1.25 M solution of nitrous acid, 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.

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

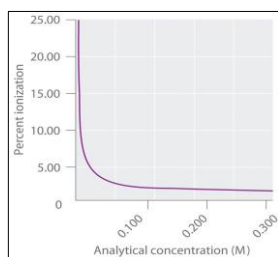
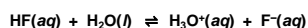
- Calculate the percent dissociation from the previous problem.

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

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

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Effect of Dilution on % Ionization

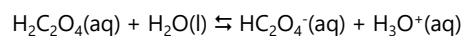


As we add water, we can dissociate more. More dilute = more dissociation.

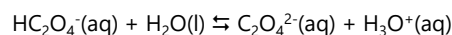
CC-BY-NA-3.0: <http://2012books.lardbucket.org/books/principles-of-general-chemistry-v1.0/s20-aqueous-acid-base-equilibriums.html>

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Polyprotic Acids



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



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

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

- pH of a polyprotic acid solution primarily arises from $[\text{H}_3\text{O}^+]$ formed in the 1st step.

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Weak Bases; Base-Ionization Constants (K_b)

The ionization of a *weak base* in water:



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

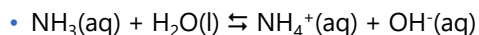
K_b is called the *base dissociation constant*.

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 OH^- (this is x!) and use K_b



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

- K_b Values for Weak Bases

- Why are amines (derivatives of NH_3) basic?

⇒ Lone e- pair on N attracts H from water.

N pulls H^+ off of water.

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Weak Bases; K_b Constants

Base	Formula	K_b	pK_b
phosphate ion	PO_4^{3-}	2.1×10^{-2}	1.68
dimethylamine	$(\text{CH}_3)_2\text{NH}$	5.4×10^{-4}	3.27
methylamine	CH_3NH_2	4.6×10^{-4}	3.34
trimethylamine	$(\text{CH}_3)_3\text{N}$	6.3×10^{-5}	4.20
ammonia	NH_3	1.8×10^{-5}	4.75
pyridine	$\text{C}_5\text{H}_5\text{N}$	1.7×10^{-9}	8.77
aniline	$\text{C}_6\text{H}_5\text{NH}_2$	7.4×10^{-10}	9.13

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Calculate pH of a Weak Base

Calculate the pH of a 0.50 M dimethylamine $((\text{CH}_3)_2\text{NH})$ solution. $K_b = 5.4 \times 10^{-4}$

Write reaction, set up K_b , fill in ICE table, solve for x!

$x = ?$

Check approximation! Is $x < 5\%$?

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

How do we find pH?

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Calculate pH of a Weak Base

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

2) **this is easiest way**
 $\text{pOH} = -\log [\text{OH}^-] =$
 $\text{pH} = 14 - \text{pOH} =$

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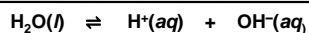
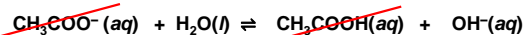
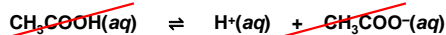
Calculate K_b for a Weak Base

A 0.065 M solution of methylamine, CH_3NH_2 , has a pH of 11.70. What is K_b for CH_3NH_2 ?

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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 1.8×10^{-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
- E.g. $\text{HCl(aq)} + \text{NaOH(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)}$
- 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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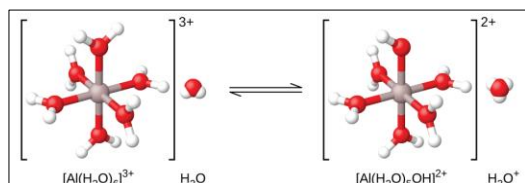
General Ion Categories

- 1) **Acidic ions:** Most cations (except 1A & 2A)
E.g. NH_4^+ , Al^{3+} , Cu^{2+}
- 2) **Neutral ions:** cations from strong bases and anions from strong acids
Cations: Group 1A & 2A (except Be)
Anions: NO_3^- , Cl^- , Br^- , I^- , ClO_4^- , ClO_3^- (except HSO_4^-)
- 3) **Basic ions:** Most anions derived from weak acids.
E.g. F^- , CN^- , CH_3COO^- , 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 react with water!
- ♦ E.g. NaCl $\text{Na}^+ + \text{H}_2\text{O} \rightarrow \text{NR}$; $\text{Cl}^- + \text{H}_2\text{O} \rightarrow \text{NR}$

Acidic Salts – formed in SA-WB reaction

- ♦ typically contain acidic cation

Cation is

- 1) conjugate acid of WB (e.g. NH_4^+) or
- 2) small, highly charged metal ion (e.g. Al^{3+})

E.g. NH_4Cl $\text{Cl}^- + \text{H}_2\text{O} \rightarrow \text{NR}$ (chloride is neutral ion)



An acidic ion loses H^+ ions as it forms hydronium ions

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Types of Salts

Basic Salts – formed in WA-SB titration

- typically contain basic anion
 - Anion is conjugate base of WA (e.g. F^-)
 - E.g. NaF Na^+ is neutral ($\text{Na}^+ + \text{H}_2\text{O} \rightarrow \text{NR}$)
 $\text{F}^-(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{HF}(\text{aq}) + \text{OH}^-(\text{aq})$
- A basic ion gains H^+ ions as it forms OH^- ions**

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. NH_4NO_2

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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 NH_4Cl acidic, basic, or neutral?
 - ♦ Write the hydrolysis reaction and calculate the pH of a 0.10 M NH_4Cl solution.
 $K_b(\text{NH}_3) = 1.8 \times 10^{-5}$

See Worksheet key

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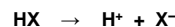
Classify Salts

- Classify the following salts as acidic, basic, or neutral. For acidic or basic salts, write the reaction of hydrolysis.
 - KBr
 - Ba(NO₂)₂
 - LiCN
 - NH₄Cl

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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	6.3x10 ⁻⁴		1.3x10 ⁶		7.9x10 ⁸		2.0x10 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

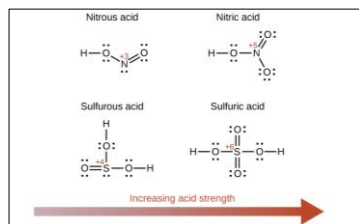
14		15		16		17	
6	CH ₄ neither acid nor base	7	NH ₃ Weak base K _b = 1.8 × 10 ⁻⁵	8	H ₂ O Neutral	9	HF Weak acid K _a = 6.8 × 10 ⁻⁴
14	SiH ₄ neither acid nor base	15	PH ₃ Very weak base K _b = 4 × 10 ⁻¹⁴	16	H ₂ S Weak acid K _a = 9.5 × 10 ⁻⁸	17	HCl Strong acid

↑ (increases from bottom) electronegativity (increases) ↓
← (increases from right) increasing acid strength
→ (increases from left) 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 × 10 ⁻⁸	7.40
HOBr	2.8	2.8 × 10 ⁻⁹	8.55
HOI	2.5	3.2 × 10 ⁻¹¹	10.5

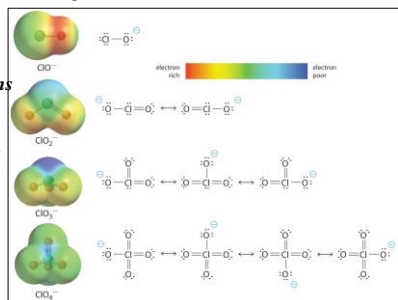
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2) Oxoacids with same central atom, but different numbers of O atoms.

•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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