

## **Properties of Acids**

□ Sour taste Change the color of indicators litmus - blue to red bromcresol - green to yellow phenolphthalein - rose to colorless React with metals to liberate hydrogen gas  $\square 2HCl + Mg \rightarrow MgCl_2 + H_2$ 

## **Properties of Acids**

□ React with basic metal oxides & hydroxides to form a salt & H<sub>2</sub>O
 □ 2HBr + CuO(s) → CuBr<sub>2</sub> + 2H<sub>2</sub>O
 □ React with salts of weaker or volatile acids to give a new salt and a new acid
 □ 2HClO<sub>4</sub> + FeS → H<sub>2</sub>S + Fe(ClO<sub>4</sub>)<sub>2</sub> + 2H<sub>2</sub>O

### **Properties of Bases**

Bitter taste
Change the color of indicators

litmus - red to blue
bromcresol - yellow to green
phenolphthalein - colorless to rose

Neutralize acids

Acid-Base Concepts: The Brønsted-Lowry Theory Arrhenius Acid: A substance that dissociates in water to produce hydrogen ions, H<sup>+</sup>.

 $HA(aq) \rightleftharpoons H^+(aq) + A^-(aq)$ 

Arrhenius Base: A substance that dissociates in water to produce hydroxide ions, OH<sup>-</sup>.

 $MOH(aq) \rightleftharpoons M^+(aq) + OH^-(aq)$ 

Acid-Base Concepts: The Brønsted-Lowry Theory Brønsted-Lowry Acid: A substance that can transfer hydrogen ions, H<sup>+</sup>. In other words, a proton donor.

**Brønsted-Lowry Base**: A substance that can accept hydrogen ions, H<sup>+</sup>. In other words, a proton acceptor.



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**Conjugate Acid-Base Pairs**: Chemical species whose formulas differ only by one hydrogen ion, H<sup>+</sup>.

## Acid-Base Concepts: The Brønsted-Lowry Theory Acid-Dissociation Equilibrium



## Acid-Base Concepts: The Brønsted-Lowry Theory Base-Dissociation Equilibrium



Acid Strength and Base Strength $HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$ Acid Base Acid Base

With equal concentrations of reactants and products, what will be the direction of reaction?

Stronger acid + Stronger base → Weaker acid + Weaker base

## Acid Strength

Strong acid: near 100% ionizationWeak acid: 10% or less ionization

# Acid Strength and Base Strength Weak Acid: An acid that is only partially dissociated in

water and is thus a weak electrolyte.



	ə , .	
	Acid, HA	Base, A <sup>-</sup>
Stronger acid	$ \begin{array}{c} \text{HClO}_4 \\ \text{HCl} \\ \text{H}_2\text{SO}_4 \\ \text{HNO}_3 \end{array} \end{array} \begin{array}{c} \textbf{Strong acids:} \\ 100\% \text{ dissociated} \\ \text{in aqueous} \\ \text{solution.} \end{array} $	$\begin{array}{c} ClO_4^{-} \\ Cl^{-} \\ HSO_4^{-} \\ NO_3^{-} \end{array} \end{array} \begin{array}{c} Very \ weak \ bases: \\ Negligible \ tendency \\ to \ be \ protonated \ in \\ aqueous \ solution. \end{array} Weaker \\ base$
	$H_{3}O^{+}$ $HSO_{4}^{-}$ $H_{3}PO_{4}$ $HNO_{2}$ $HF$ $CH_{3}CO_{2}H$ $H_{2}CO_{3}$ $H_{2}S$ $NH_{4}^{+}$ $HCN$ $HCO_{3}^{-}$ $H_{2}O$ $Weak acids:$ $Exist in solution$ $as a mixture of$ $HA, A^{-}, and H_{3}O^{+}.$	H <sub>2</sub> O SO <sub>4</sub> <sup>2-</sup> H <sub>2</sub> PO <sub>4</sub> - NO <sub>2</sub> - $F^-$ CH <sub>3</sub> CO <sub>2</sub> - HCO <sub>3</sub> - HS <sup>-</sup> NH <sub>3</sub> CN <sup>-</sup> CO <sub>3</sub> <sup>2-</sup> OH <sup>-</sup> Weak bases: Moderate tendency to be protonated in aqueous solution.
Weaker acid	$ \begin{array}{c} \text{NH}_{3} \\ \text{OH}^{-} \\ \text{H}_{2} \end{array} \end{array} \begin{array}{c} \text{Very weak acids:} \\ \text{Negligible tendency} \\ \text{to dissociate.} \end{array} $	$ \begin{array}{c} \mathrm{NH_2}^- \\ \mathrm{O^{2-}} \\ \mathrm{H}^- \end{array} \end{array} \begin{array}{c} \textbf{Strong bases:} \\ 100\% \text{ protonated in} \\ \text{aqueous solution.} \end{array} \begin{array}{c} \textbf{Stronger} \\ \textbf{base} \end{array} $

 TABLE 14.1
 Relative Strengths of Conjugate Acid–Base Pairs

# Hydrated Protons and Hydronium Ions

### $HA(aq) \rightleftharpoons H^+(aq) + A^-(aq)$

Due to high reactivity of the hydrogen ion, it is actually hydrated by one or more water molecules.

#### For our purposes, H<sup>+</sup> is equivalent to H<sub>3</sub>O<sup>+</sup>.



**Dissociation of Water:** 

 $2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$ 

### **Dissociation of Water**

**Ion-Product Constant for Water**:  $K_w = [H_3O^+][OH^-]$ 

at 25°C:  $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7} M$ 

So...

 $K_{\rm w} = (1.0 \text{ x } 10^{-7})(1.0 \text{ x } 10^{-7}) = 1.0 \text{ x } 10^{-14}$ 

Dissociation of Water  $K_{\rm w} = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$ 



or

$$[OH^{-}] = \frac{K_{w}}{[H_{3}O^{+}]} = \frac{1.0 \times 10^{-14}}{[H_{3}O^{+}]}$$

## **Dissociation of Water**





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# The pH Scale

Basic solution:pH > 7Neutral solution:pH = 7Acidic solution:pH < 7

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

and

 $[H_3O^+] = 10^{-pH}$ 

The hydronium ion concentration for lemon juice is approximately 0.0025. What is the pH when  $[H_3O^+] = 0.0025$  M?

2 significant figures pH = -log(0.0025) = 2.6

Calculate the pH of an aqueous ammonia solution that has an OH<sup>-</sup> concentration of 0.0019 M.

$$[H_3O^+] = \frac{1.0 \times 10^{-14}}{[OH^-]} = \frac{1.0 \times 10^{-14}}{0.0019} = 5.3 \times 10^{-12} M$$

 $pH = -log(5.3 \times 10^{-12}) = 11.28$ 

Acid rain is a matter of serious concern because most species of fish die in waters having a pH lower than 4.5–5.0. Calculate  $[H_3O^+]$  in a lake that has a pH of 4.5.

 $[H_3O^+] = 10^{-4.5} = 3.2 \text{ x } 10^{-5} \text{ M}$ 

# Measuring pH

Acid-Base Indicator: A substance that changes color in a specific pH range. Indicators exhibit pH-dependent color changes because they are weak acids and have different colors in their acid (HIn) and conjugate base (In<sup>1-</sup>) forms.

 $HIn(aq) + H_2O(l) \iff H_3O^{1+}(aq) + In^{1-}(aq)$ Color A Color B

# Measuring pH



## Measuring pH

![](_page_24_Figure_1.jpeg)

The pH in Solutions of Strong Acids and Strong Bases What is the pH of a 0.025 M solution of HNO<sub>3</sub>?

 $HNO_3(aq) + H_2O(l) \xrightarrow{100\%} H_3O^+(aq) + NO_3^-(aq)$ 

Since HNO<sub>3</sub> is a strong acid,  $[H_3O^+] = [HNO_3]$ .

 $pH = -log([H_3O^+]) = -log(0.025) = 1.60$ 

The pH in Solutions of Strong Acids and Strong Bases What is the pH of a 0.025 M solution of NaOH?

 $NaOH(aq) \longrightarrow Na^+(aq) + OH^{1-}(aq)$ 

Since NaOH is a strong base,  $[OH^{1-}] = [NaOH].$   $[H_3O^+] = \frac{1.0 \times 10^{-14}}{[OH^-]} = \frac{1.0 \times 10^{-14}}{0.025} = 4.0 \times 10^{-13} \text{ M}$  $pH = -log([H_3O^+]) = -log(4.0 \times 10^{-13}) = 12.40$ 

# Equilibria in Solutions of Weak Acids

### $HA(aq) + H_2O(l) \implies H_3O^+(aq) + A^-(aq)$

![](_page_27_Picture_3.jpeg)

# Equilibria in Solutions of Weak Acids

The pH of 0.250 M HF is 2.036. What are the values of  $K_a$  and p $K_a$  for hydrofluoric acid?

 $HF(aq) + H_2O(l) \Longrightarrow H_3O^+(aq) + F^-(aq)$ 

conc <sub>i</sub>	0.250	行行的	≈0	0
change	-x	22月19日	+x	+x
conc <sub>eq</sub>	0.250 - x		x	x

 $x = [H_3O^+] = 10^{-2.036} = 0.00920 \text{ M}$ 

# Equilibria in Solutions of Weak Acids $K_a = \frac{[H_3O^+][F^-]}{[HF]}$

 $[F^{-}] = [H_3O^{+}] = 0.00920 \text{ M}$ 

[HF] = 0.250 - x = 0.250 - 0.00920 = 0.241 M

$$K_{\rm a} = \frac{[{\rm H}_{3}{\rm O}^{+}][{\rm F}^{-}]}{[{\rm H}{\rm F}]} = \frac{(0.00920)(0.00920)}{0.241} = 3.51 \times 10^{-4}$$
$$pK_{\rm a} = -\log(K_{\rm a}) = -\log(3.51 \times 10^{-4}) = 3.455$$

Calculating Equilibrium Concentrations for Weak Acids Calculate the pH of a 0.10 M HCN solution. At 25 °C,  $K_a = 1.4 \times 10^{-9}$ .

 $HCN(aq) + H_2O(l) \Longrightarrow H_3O^+(aq) + CN^-(aq)$ 

conc <sub>i</sub>	0.10		≈0	0
change	-X		+X	+X
conc <sub>eq</sub>	0.10 - x	1944	x	x

$$K_{\rm a} = \frac{[\rm H_3O^+][\rm CN^-]}{[\rm HCN]}$$

Calculating Equilibrium Concentrations for Weak Acids  $1.4 \ge 10^{-9} = \frac{(x)(x)}{(0.10 - x)} \approx \frac{x^2}{0.10}$ 

 $x = [H_3O^+] = 1.2 \times 10^{-5} M$ 

 $pH = -log([H_3O^+]) = -log(1.2 \times 10^{-5}) = [4.9]$ 

![](_page_32_Figure_0.jpeg)

# Percent Dissociation in Solutions of Weak Acids

Percent dissociation =

[HA] dissociated [HA] initial x 100%

![](_page_33_Figure_3.jpeg)

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## **Classification of Acids**

MonoproticDiproticTriprotic

Polyprotic Acids  $H_2CO_3(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$  $K_{a1} = \frac{[H_3O^+] [HCO_3^-]}{[H_2CO_3]} = 4.3 \times 10^{-7}$ 

 $HCO_{3}(aq) + H_{2}O(l) \rightleftharpoons H_{3}O^{+}(aq) + CO_{3}(aq)$  $K_{a2} = \frac{[H_{3}O^{+}][CO_{3}(2)]}{[HCO_{3}(2)]} = 5.6 \times 10^{-11}$ 

# **Polyprotic Acids**

# TABLE 14.3Stepwise Dissociation Constants for Polyprotic<br/>Acids at 25 °C

Formula	K <sub>a1</sub>	K <sub>a2</sub>	K <sub>a3</sub>
H <sub>2</sub> CO <sub>3</sub>	$4.3  imes 10^{-7}$	$5.6  imes 10^{-11}$	
$H_2S$	$1.0 \times 10^{-7}$	$\sim 10^{-19}$	
$H_2C_2O_4$	$5.9 \times 10^{-2}$	$6.4  imes 10^{-5}$	
H <sub>3</sub> PO <sub>4</sub>	$7.5 \times 10^{-3}$	$6.2  imes 10^{-8}$	$4.8 \times 10^{-13}$
$H_2SO_4$	Very large	$1.2 \times 10^{-2}$	
$H_2SO_3$	$1.5 \times 10^{-2}$	$6.3  imes 10^{-8}$	
	Formula $H_2CO_3$ $H_2S$ $H_2C_2O_4$ $H_3PO_4$ $H_2SO_4$ $H_2SO_4$ $H_2SO_3$	Formula $K_{a1}$ $H_2CO_3$ $4.3 \times 10^{-7}$ $H_2S$ $1.0 \times 10^{-7}$ $H_2C_2O_4$ $5.9 \times 10^{-2}$ $H_3PO_4$ $7.5 \times 10^{-3}$ $H_2SO_4$ Very large $H_2SO_3$ $1.5 \times 10^{-2}$	Formula $K_{a1}$ $K_{a2}$ $H_2CO_3$ $4.3 \times 10^{-7}$ $5.6 \times 10^{-11}$ $H_2S$ $1.0 \times 10^{-7}$ $\sim 10^{-19}$ $H_2C_2O_4$ $5.9 \times 10^{-2}$ $6.4 \times 10^{-5}$ $H_3PO_4$ $7.5 \times 10^{-3}$ $6.2 \times 10^{-8}$ $H_2SO_4$ Very large $1.2 \times 10^{-2}$ $H_2SO_3$ $1.5 \times 10^{-2}$ $6.3 \times 10^{-8}$

\*Because of its very small size,  $K_{a2}$  for  $H_2S$  is difficult to measure and its value is uncertain.

# Polyprotic Acids Calculate the pH of a 0.020 M H<sub>2</sub>CO<sub>3</sub> solution. At 25 °C, $K_{a1} = 4.3 \times 10^{-7}$ .

 $H_2CO_3(aq) + H_2O(l) \Longrightarrow H_3O^+(aq) + HCO_3^-(aq)$ 

conc <sub>i</sub>	0.020	COM I	≈0	0
change	-x	11 Alex F	+X	+x
conc <sub>eq</sub>	0.020 - x		x	x

$$K_{a1} = \frac{[H_3O^+][HCO_3^-]}{[H_2CO_3]}$$

# Polyprotic Acids $4.3 \ge 10^{-7} = \frac{(x)(x)}{(0.020 - x)} \approx \frac{x^2}{0.020}$

 $x = [H_3O^+] = 9.3 \times 10^{-5} M$ 

 $pH = -log([H_3O^+]) = -log(9.3 \times 10^{-5}) = 4.0$ 

 $B(aq) + H_2O(l) \Longrightarrow BH^+(aq) + OH^-(aq)$ 

Base Acid Acid Base

**Base-Dissociation Constant:** 

[BH<sup>+</sup>][OH<sup>-</sup>] [B]

 $K_{\rm b} =$ 

 $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$ 

 $K_{\rm b} = \frac{[\rm NH_4^+][\rm OH^-]}{[\rm NH_3]}$ 

TABLE 14.4 $K_b$  Values for Some Weak Bases and  $K_a$  Values for TheirConjugate Acids at 25 °C

Base	Formula, B	K <sub>b</sub>	Conjugate Acid, BH <sup>+</sup>	Ka
Ammonia	NH <sub>3</sub>	$1.8 \times 10^{-5}$	$NH_4$ +	$5.6 \times 10^{-10}$
Aniline	$C_6H_5NH_2$	$4.3 \times 10^{-10}$	$C_6H_5NH_3^+$	$2.3 \times 10^{-5}$
Dimethylamine	(CH <sub>3</sub> ) <sub>2</sub> NH	$5.4  imes 10^{-4}$	$(CH_3)_2NH_2^+$	$1.9 \times 10^{-11}$
Hydrazine	$N_2H_4$	$8.9 \times 10^{-7}$	$N_2H_5$ <sup>+</sup>	$1.1 \times 10^{-8}$
Hydroxylamine	NH <sub>2</sub> OH	$9.1 \times 10^{-9}$	NH <sub>3</sub> OH <sup>+</sup>	$1.1 \times 10^{-6}$
Methylamine	CH <sub>3</sub> NH <sub>2</sub>	$3.7 \times 10^{-4}$	$CH_3NH_3^+$	$2.7 \times 10^{-11}$

Calculate the pH of a 0.40 M NH<sub>3</sub> solution. At 25 °C,  $K_b = 1.8 \times 10^{-5}$ .

 $NH_3(aq) + H_2O(l) \Longrightarrow NH_4^+(aq) + OH^-(aq)$ 

conc <sub>i</sub>	0.40	0	≈0
change	-X	+X	+x
conc <sub>eq</sub>	0.40 - x	×	x

 $K_{\rm b} = \frac{[\rm NH_4^+][\rm OH^-]}{[\rm NH_3]}$ 

$$1.8 \ge 10^{-5} = \frac{(x)(x)}{(0.40 - x)} \approx \frac{x^2}{0.40}$$

 $x = [OH^{-}] = 0.0027 M$ 

 $[H_3O^+] = \frac{1.0 \times 10^{-14}}{0.0027} = 3.7 \times 10^{-12} \text{ M}$ 

 $pH = -log([H_3O^+]) = -log(3.7 \times 10^{-12}) = [11.4]$ 

Relation Between  $K_a$  and  $K_b$   $NH_4^+(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + NH_3(aq)$  $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$ 

 $2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$ 

Ka

K<sub>b</sub>

 $K_{\rm w}$ 

![](_page_44_Figure_2.jpeg)

 $= (5.6 \times 10^{-10})(1.8 \times 10^{-5}) = 1.0 \times 10^{-14}$ 

# Relation Between $K_a$ and $K_b$ $K_a \ge K_b = K_w$

conjugate acid-base pair

![](_page_45_Picture_2.jpeg)

![](_page_45_Picture_3.jpeg)

 $pK_{a} + pK_{b} = pK_{w} = 14.00$ 

## Acid-Base Properties of Salts

![](_page_46_Figure_1.jpeg)

### Acid-Base Properties of Salts Salts That Yield Neutral Solutions

The following ions do not react appreciably with water to produce either  $H_3O^+$  or  $OH^-$  ions:

#### Cations from strong bases:

- Alkali metal cations of group 1a (Li<sup>+</sup>, Na<sup>+</sup>, K<sup>+</sup>)
- Alkaline earth metal cations of group 2a (Mg<sup>2+</sup>, Ca<sup>2+</sup>, Sr<sup>2+</sup>, Ba<sup>2+</sup>), except for Be<sup>2+</sup>
- Anions from strong monoprotic acids:
  - $Cl^-$ ,  $Br^-$ ,  $I^-$ ,  $NO_3^-$ , and  $ClO_4^-$

### Acid-Base Properties of Salts Salts That Yield Acidic Solutions

Salts such as  $NH_4Cl$  that are derived from a weak base ( $NH_3$ ) and a strong acid (HCl) yield acidic solutions.

 $NH_4^+(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + NH_3(aq)$ 

Ammonium ion  $(NH_4^+)$  is the conjugate acid of the weak base ammonia  $(NH_3)$  while chloride ion  $(Cl^-)$  is neither acidic nor basic.

## Acid-Base Properties of Salts

#### **Salts That Yield Acidic Solutions**

# Hydrated cations of small, highly charged metal ions, such as Al<sup>3+</sup>.

![](_page_49_Figure_3.jpeg)

### Acid-Base Properties of Salts Salts That Yield Acidic Solutions

Hydrated cations of small, highly charged metal ions, such as Al<sup>3+</sup>.

![](_page_50_Picture_2.jpeg)

### Acid-Base Properties of Salts Salts That Yield Acidic Solutions

The acidity of hydrated main-group cations increases from left to right in the periodic table as the metal ion charge increases and the metal ion size decreases:

> $Li^+ < Be^{2+}$ Na<sup>+</sup> < Mg^{2+} < Al^{3+}

### Acid-Base Properties of Salts Salts That Yield Basic Solutions

Salts such as NaCN that are derived from a strong base (NaOH) and a weak acid (HCN) yield basic solutions.

 $CN^{-}(aq) + H_2O(l) \Longrightarrow HCN(aq) + OH^{-}(aq)$ 

Cyanide ion  $(CN^{-})$  is the conjugate base of the weak acid hydrocyanic acid (HCN) while sodium ion  $(Na^{+})$  is neither acidic nor basic.

### Acid-Base Properties of Salts Salts That Contain Acidic Cations and Basic Anions

The pH of an ammonium carbonate solution,  $(NH_4)_2CO_3$ , depends on the relative acid strength of the cation and the relative base strength of the anion.

Is it acidic or basic?

Acid-Base Properties of Salts Salts That Contain Acidic Cations and Basic Anions (NH<sub>4</sub>)<sub>2</sub>CO<sub>3</sub>:

 $NH_4^+(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + NH_3(aq) \qquad K_a$  $CO_3^{2-}(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + OH^-(aq) \qquad K_b$ 

Three possibilities:

- $K_a > K_b$ : The solution will contain an excess of H<sub>3</sub>O<sup>+</sup> ions (pH < 7).
- $K_a < K_b$ : The solution will contain an excess of OHions (pH > 7).
- $K_a \approx K_b$ : The solution will contain approximately equal concentrations of H<sub>3</sub>O<sup>+</sup> and OH<sup>-</sup> ions (pH  $\approx$  7).

Acid-Base Properties of Salts Salts That Contain Acidic Cations and Basic Anions (NH<sub>4</sub>)<sub>2</sub>CO<sub>3</sub>:

 $NH_4^+(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + NH_3(aq) \qquad K_a$  $CO_3^{2-}(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + OH^-(aq) \qquad K_b$ 

$$K_{a} \text{ for NH}_{4}^{+} = \frac{K_{w}}{K_{b} \text{ for NH}_{3}} = \frac{1.0 \text{ x } 10^{-14}}{1.8 \text{ x } 10^{-5}} = 5.6 \text{ x } 10^{-10}$$
$$K_{b} \text{ for CO}_{3}^{2-} = \frac{K_{w}}{K_{a} \text{ for HCO}_{3}^{1-}} = \frac{1.0 \text{ x } 10^{-14}}{5.6 \text{ x } 10^{-11}} = 1.8 \text{ x } 10^{-4}$$

Basic,  $K_{\rm a} < K_{\rm b}$ 

## Acid-Base Properties of Salts

#### TABLE 14.5 Acid–Base Properties of Salts

Type of Salt	Examples	Ions That React with Water	pH of Solution
Cation from strong base; anion from strong acid	NaCl, KNO <sub>3</sub> , BaI <sub>2</sub>	None	~7
Cation from weak base; anion from strong acid	NH <sub>4</sub> Cl, NH <sub>4</sub> NO <sub>3</sub> , [(CH <sub>3</sub> ) <sub>3</sub> NH]Cl	Cation	<7
Small, highly charged cation; anion from strong acid	AlCl <sub>3</sub> , Cr(NO <sub>3</sub> ) <sub>3</sub> , Fe(ClO <sub>4</sub> ) <sub>3</sub>	Hydrated cation	<7
Cation from strong base; anion from weak acid	NaCN, KF, Na <sub>2</sub> CO <sub>3</sub>	Anion	>7
Cation from weak base; anion from weak acid	NH <sub>4</sub> CN, NH <sub>4</sub> F, (NH <sub>4</sub> ) <sub>2</sub> CO <sub>3</sub>	Cation and anion	<7 if $K_a > K_b$ >7 if $K_a < K_b$ ~7 if $K_a \approx K_b$

![](_page_57_Figure_1.jpeg)

**Bond Polarity** 

![](_page_58_Figure_2.jpeg)

Oxoacids

![](_page_59_Figure_2.jpeg)

Oxyacids

![](_page_60_Figure_2.jpeg)

![](_page_60_Figure_3.jpeg)

**Oxidation number of Cl** 

## Lewis Acids and Bases

Lewis Acid: An electron-pair acceptor.

Lewis Base: An electron-pair donor.

![](_page_61_Figure_3.jpeg)

Lewis acid Lewis base

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## Lewis Acids and Bases

![](_page_62_Picture_1.jpeg)

![](_page_63_Figure_0.jpeg)

![](_page_63_Figure_1.jpeg)

![](_page_63_Figure_2.jpeg)

![](_page_63_Picture_3.jpeg)

![](_page_64_Picture_0.jpeg)