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Ninth
Edition

GENERAL CHEMISTRY

Principles and Modern Applications

Chapter 16: Acids and Bases

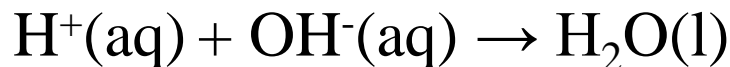
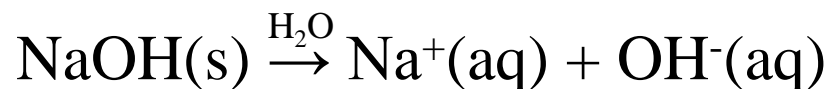
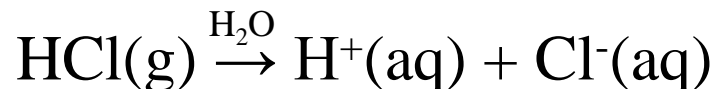
Philip Dutton
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➤ *Focus On Acid Rain*

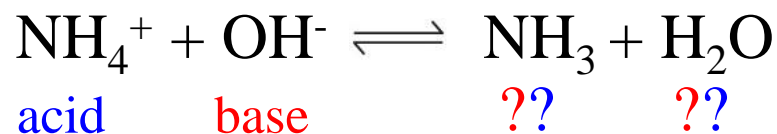
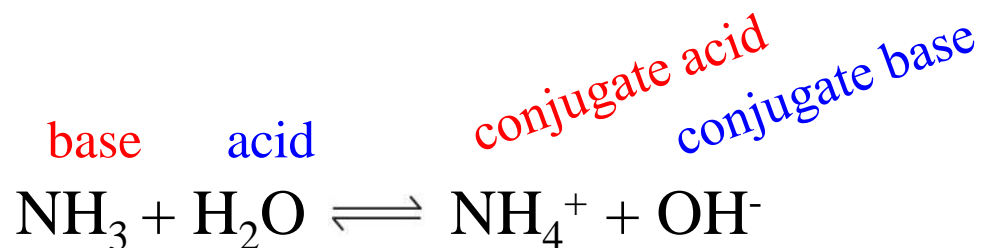
16-1 The Arrhenius Theory: A Brief Review



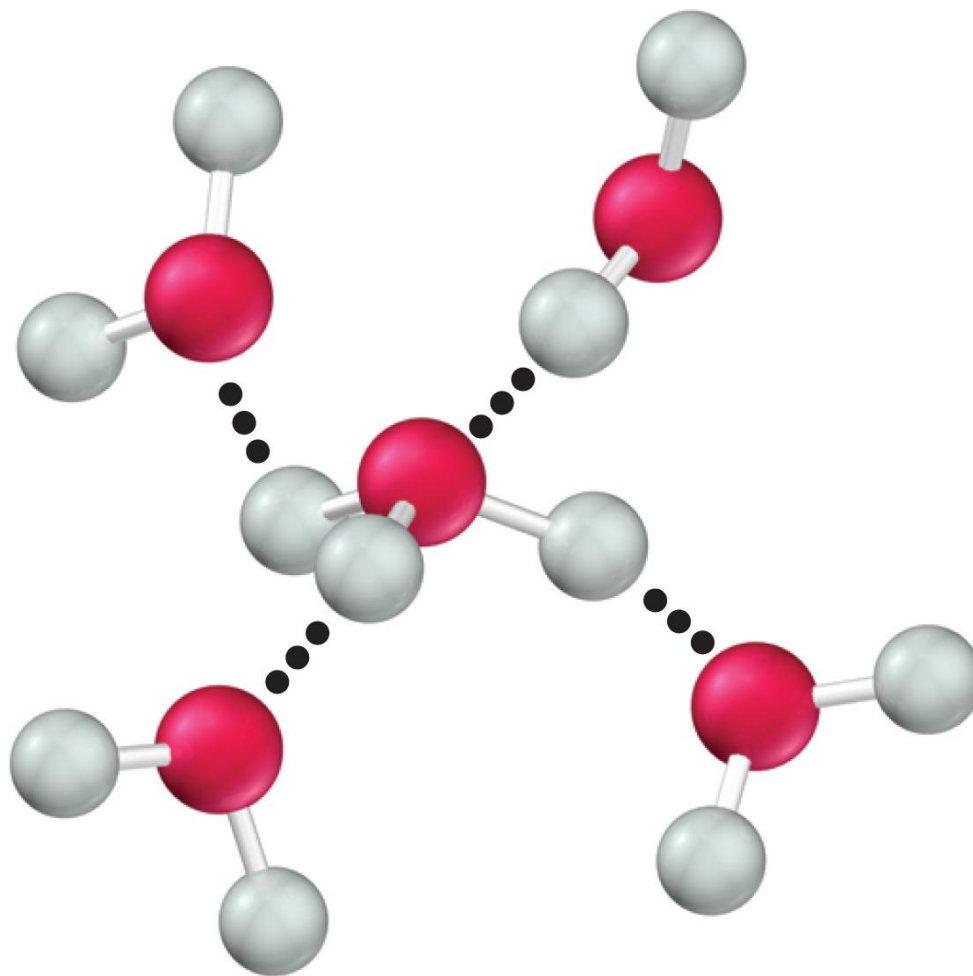
Arrhenius theory did not handle non OH⁻
bases such as ammonia very well.

16-2 Brønsted-Lowry Theory of Acids and Bases

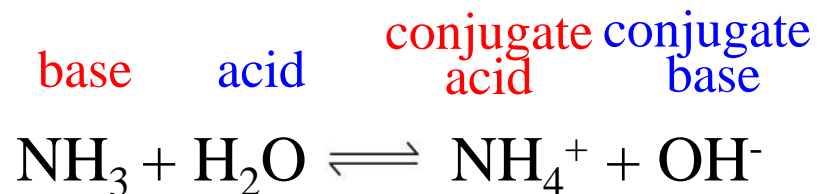
- ◆ An acid is a *proton donor*.
- ◆ A base is a *proton acceptor*.



The Solvated Proton



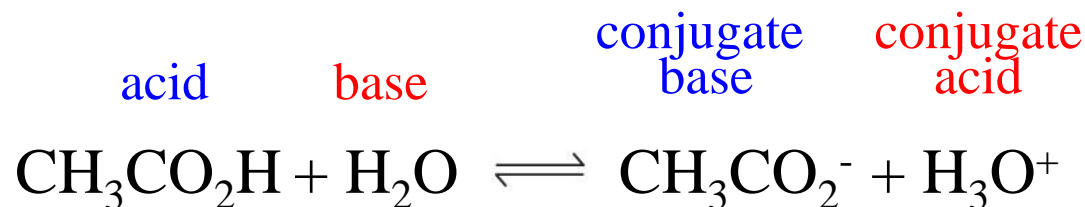
Base Ionization Constant



$$K_c = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3][\text{H}_2\text{O}]}$$

$$K_b = K_c[\text{H}_2\text{O}] = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = 1.8 \times 10^{-5}$$



Acid Ionization Constant



$$K_c = \frac{[\text{CH}_3\text{CO}_2^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{CO}_2\text{H}][\text{H}_2\text{O}]}$$

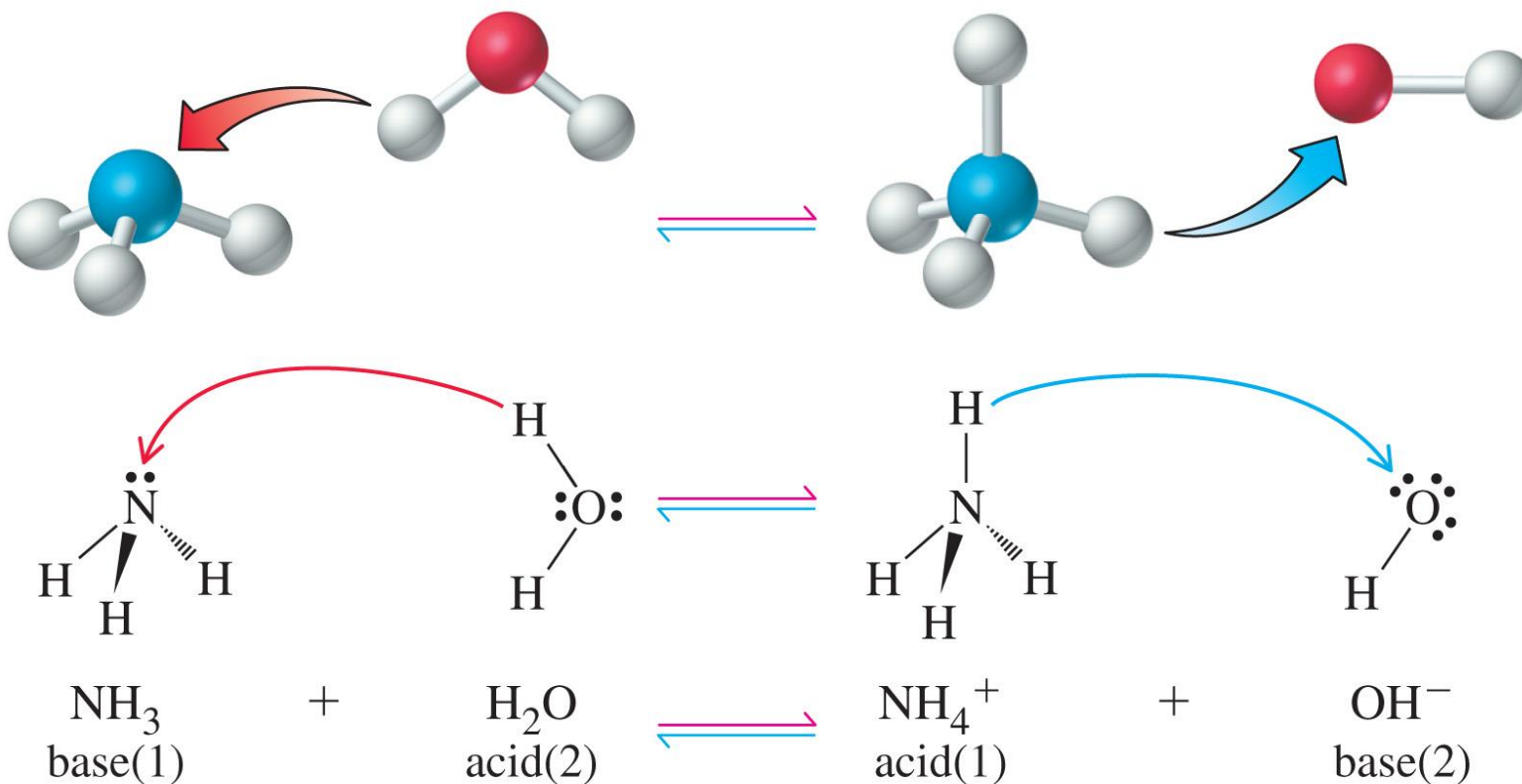
$$K_a = K_c[\text{H}_2\text{O}] = \frac{[\text{CH}_3\text{CO}_2^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{CO}_2\text{H}]} = 1.8 \times 10^{-5}$$

TABLE 16.1 Relative Strengths of Some Common Brønsted–Lowry Acids and Bases

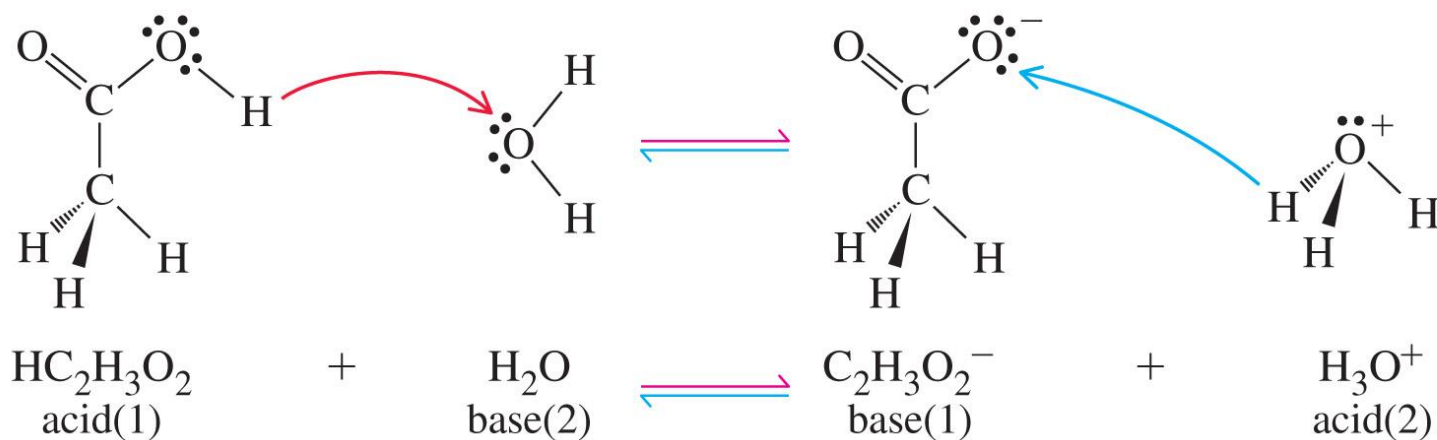
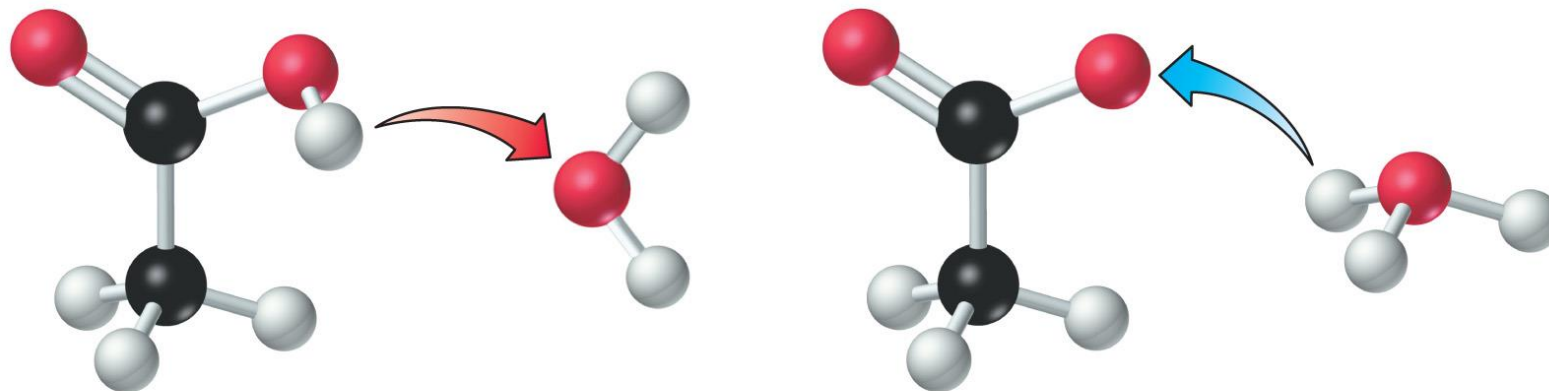
Acid		Conjugate Base		
 Increasing acid strength	Perchloric acid	HClO_4	Perchlorate ion	ClO_4^-
	Hydroiodic acid	HI	Iodide ion	I^-
	Hydrobromic acid	HBr	Bromide ion	Br^-
	Hydrochloric acid	HCl	Chloride ion	Cl^-
	Sulfuric acid	H_2SO_4	Hydrogen sulfate ion	HSO_4^-
	Nitric acid	HNO_3	Nitrate ion	NO_3^-
	Hydronium ion ^a	H_3O^+	Water ^a	H_2O
	Hydrogen sulfate ion	HSO_4^-	Sulfate ion	SO_4^{2-}
	Nitrous acid	HNO_2	Nitrite ion	NO_2^-
	Acetic acid	$\text{HC}_2\text{H}_3\text{O}_2$	Acetate ion	$\text{C}_2\text{H}_3\text{O}_2^-$
	Carbonic acid	H_2CO_3	Hydrogen carbonate ion	HCO_3^-
	Ammonium ion	NH_4^+	Ammonia	NH_3
	Hydrogen carbonate ion	HCO_3^-	Carbonate ion	CO_3^{2-}
	Water	H_2O	Hydroxide ion	OH^-
	Methanol	CH_3OH	Methoxide ion	CH_3O^-
	Ammonia	NH_3	Amide ion	NH_2^-
				 Increasing base strength

^aThe hydronium ion–water combination refers to the ease with which a proton is passed from one water molecule to another; that is, $\text{H}_3\text{O}^+ + \text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{O} + \text{H}_3\text{O}^+$

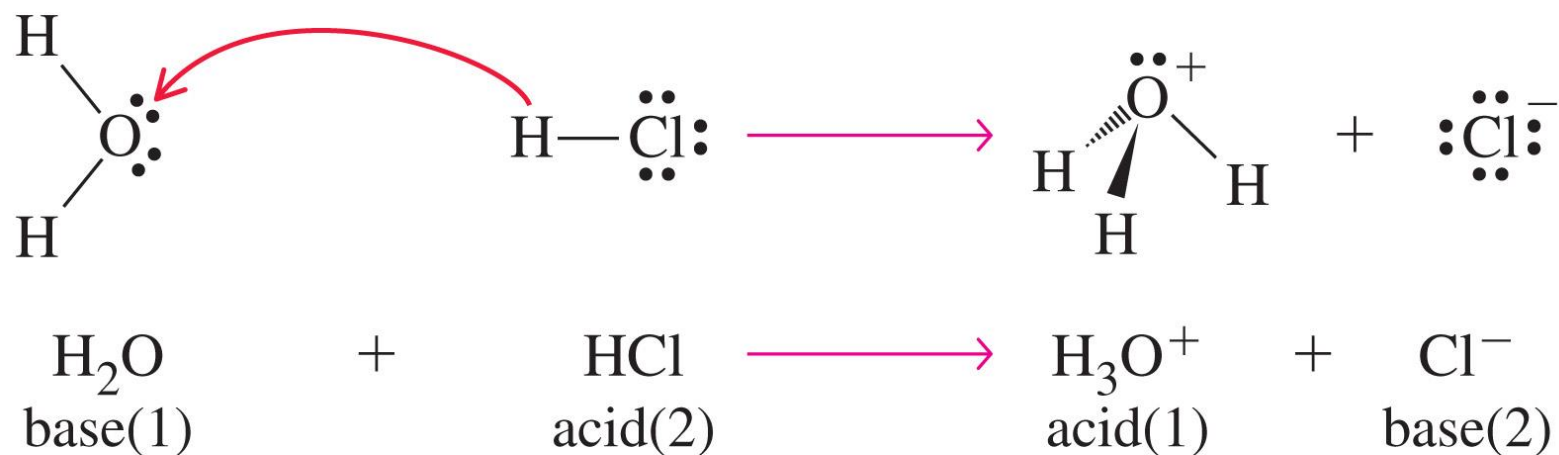
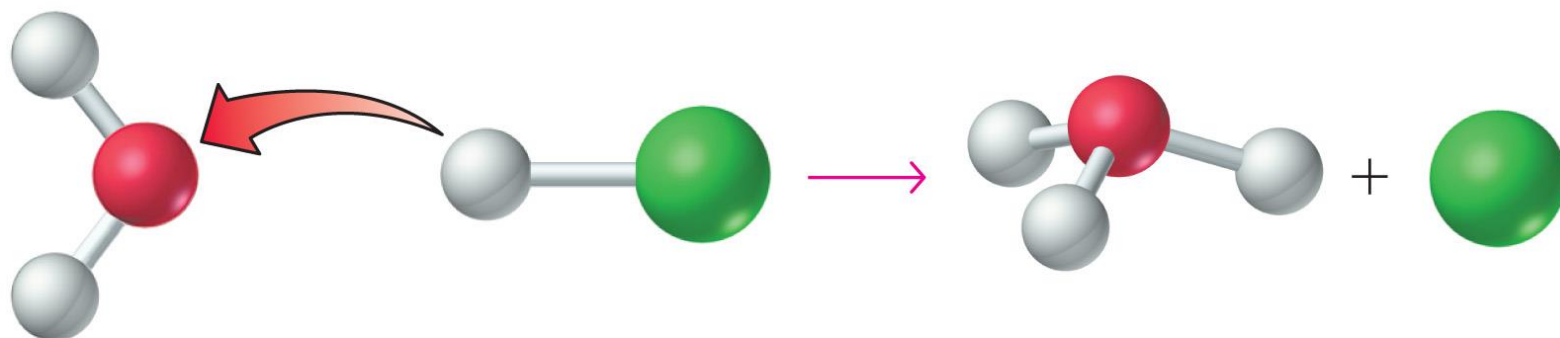
A Weak Base



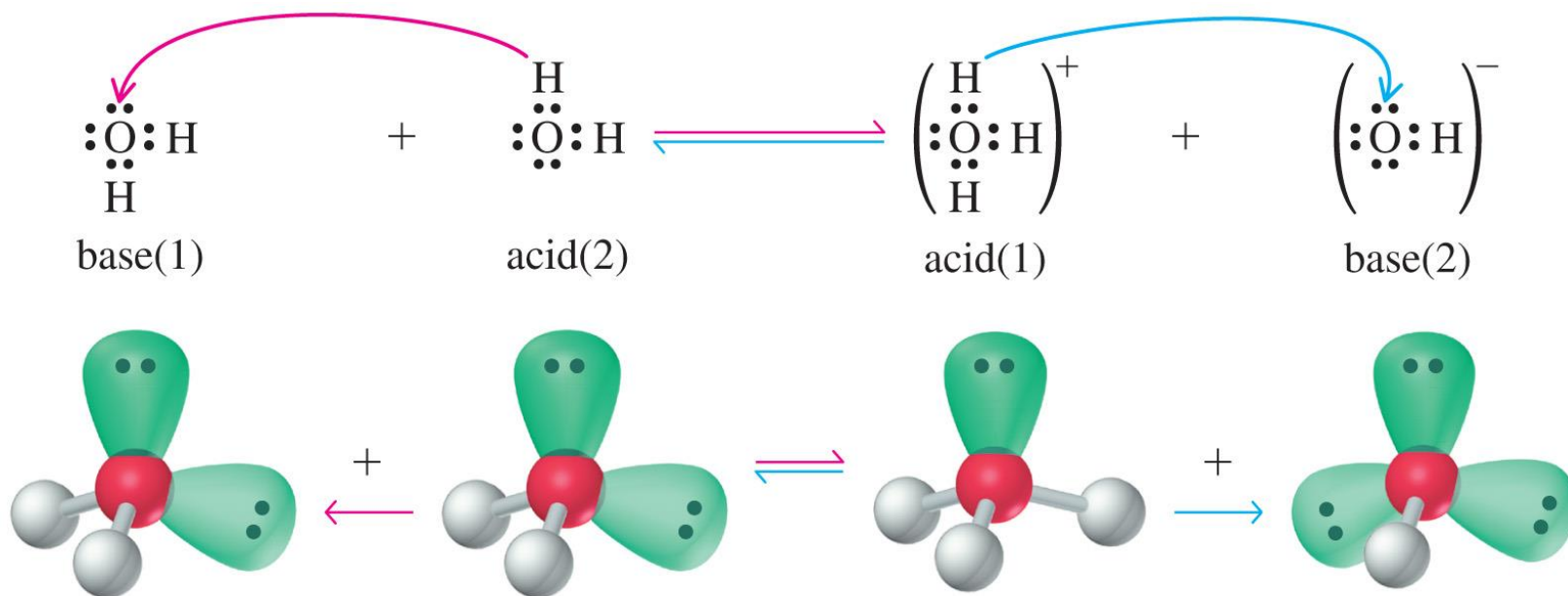
A Weak Acid



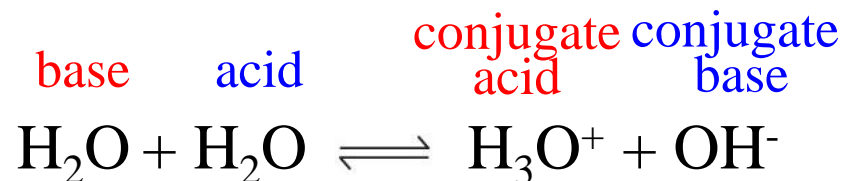
A Strong Acid



16-3 The Self-Ionization of Water and the pH Scale



Ion Product of Water



$$K_c = \frac{[\text{H}_3\text{O}^+][\text{OH}^-]}{[\text{H}_2\text{O}][\text{H}_2\text{O}]}$$

$$K_w = K_c[\text{H}_2\text{O}][\text{H}_2\text{O}] = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

pH and pOH

- ◆ The potential of the hydrogen ion was defined in 1909 as *the negative of the logarithm of $[H^+]$* .

$$pH = -\log[H_3O^+] \qquad pOH = -\log[OH^-]$$

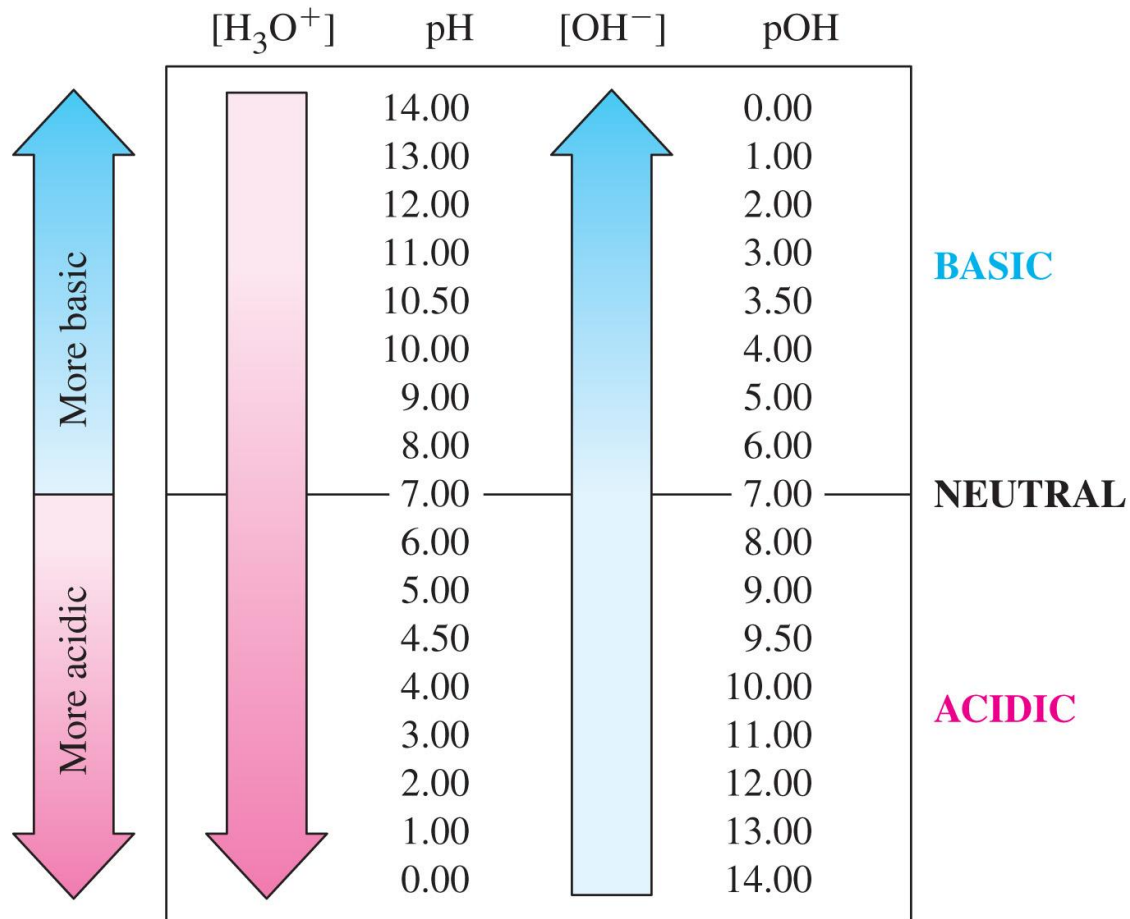
$$K_W = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

$$-\log K_W = -\log[H_3O^+] - \log[OH^-] = -\log(1.0 \times 10^{-14})$$

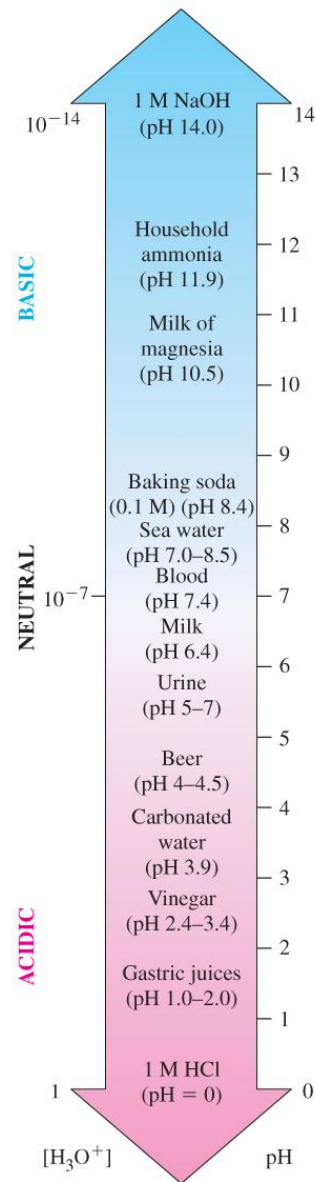
$$pK_W = pH + pOH = -(-14)$$

$$pK_W = pH + pOH = 14$$

pH and pOH Scales



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16-4 Strong Acids and Bases

TABLE 16.2
The Common Strong
Acids and Strong Bases

Acids	Bases
HCl	LiOH
HBr	NaOH
HI	KOH
HClO ₄	RbOH
HNO ₃	CsOH
H ₂ SO ₄ ^a	Mg(OH) ₂
	Ca(OH) ₂
	Sr(OH) ₂
	Ba(OH) ₂

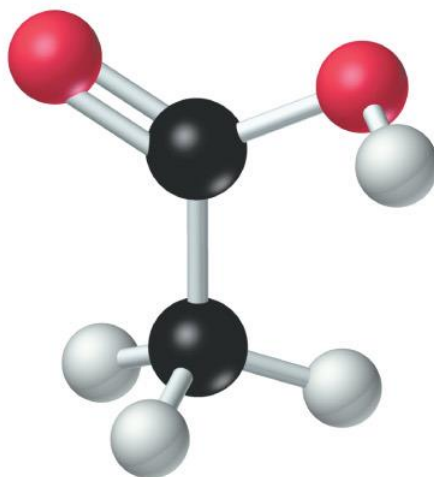
^aH₂SO₄ ionizes in two distinct steps. It is a strong acid only in its first ionization



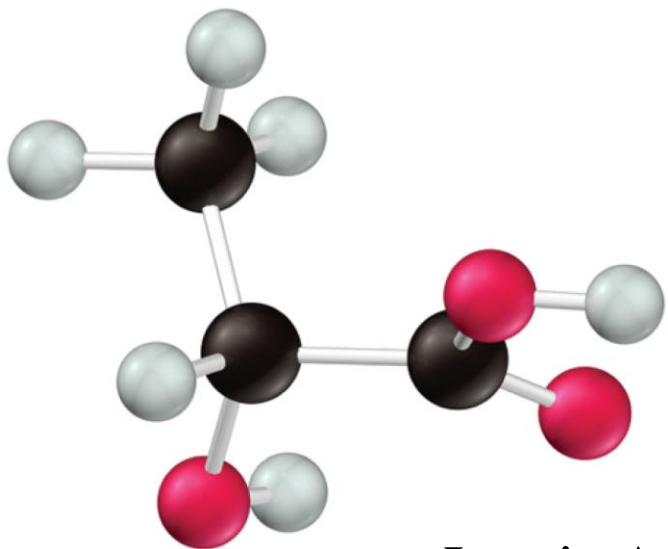
Thymol Blue Indicator



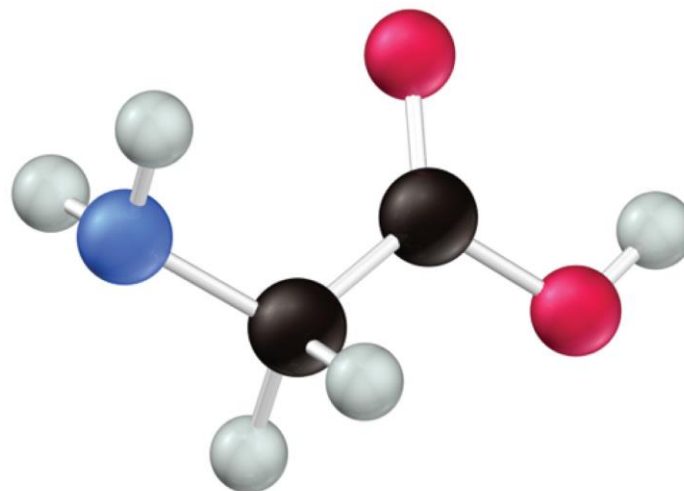
16-5 Weak Acids and Bases



Acetic Acid

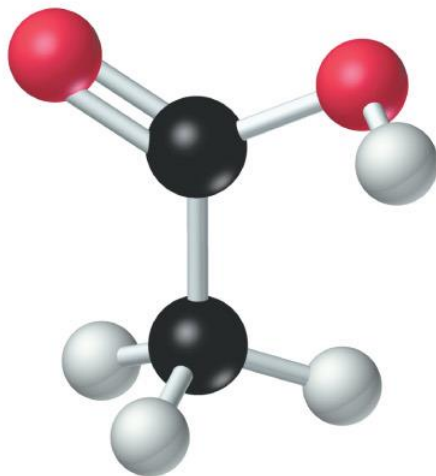


Lactic Acid



Glycine

Weak Acids

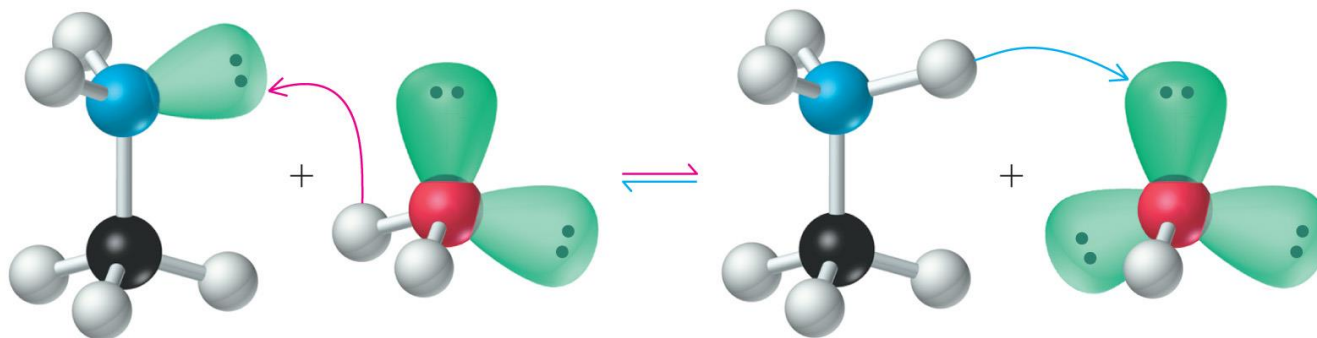
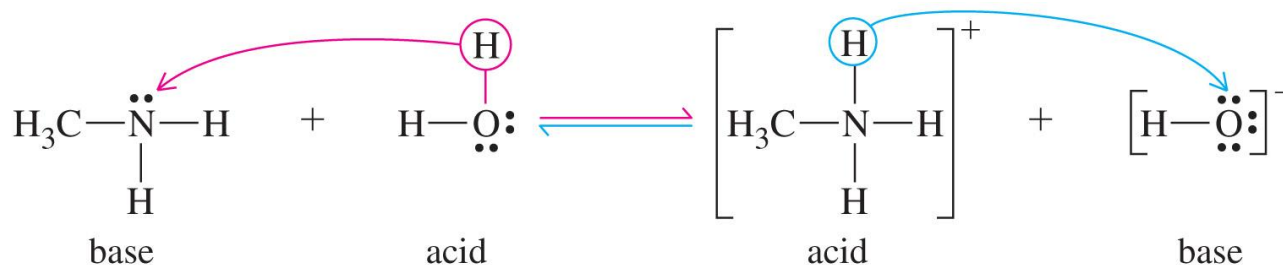


Acetic Acid

$$K_a = \frac{[\text{CH}_3\text{CO}_2^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{CO}_2\text{H}]} = 1.8 \times 10^{-5}$$

$$\text{p}K_a = -\log(1.8 \times 10^{-5}) = 4.74$$

Weak Bases



$$K_b = \frac{[\text{CH}_3\text{NH}_3^+][\text{HO}^-]}{[\text{CH}_3\text{NH}_2]} = 4.3 \times 10^{-4}$$

$$\text{p}K_b = -\log(4.2 \times 10^{-4}) = 3.37$$

Table 16.3 Ionization Constants of Weak Acids and Bases

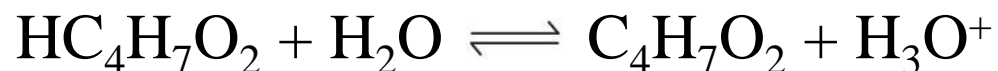
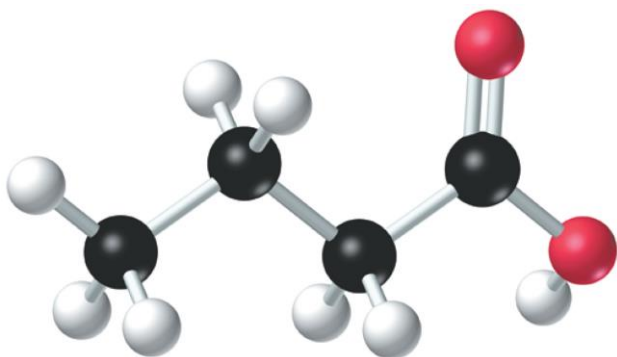
Ionization Equilibrium			Ionization Constant K	pK
Acid			$K_a =$	$pK_a =$
Iodic acid	$\text{HIO}_3 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{IO}_3^-$		1.6×10^{-1}	0.80
Chlorous acid	$\text{HClO}_2 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{ClO}_2^-$		1.1×10^{-2}	1.96
Chloroacetic acid	$\text{HC}_2\text{H}_2\text{ClO}_2 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{C}_2\text{H}_2\text{ClO}_2^-$		1.4×10^{-3}	2.85
Nitrous acid	$\text{HNO}_2 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{NO}_2^-$		7.2×10^{-4}	3.14
Hydrofluoric acid	$\text{HF} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{F}^-$		6.6×10^{-4}	3.18
Formic acid	$\text{HCHO}_2 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{CHO}_2^-$		1.8×10^{-4}	3.74
Benzoic acid	$\text{HC}_7\text{H}_5\text{O}_2 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{C}_7\text{H}_5\text{O}_2^-$		6.3×10^{-5}	4.20
Hydrazoic acid	$\text{HN}_3 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{N}_3^-$		1.9×10^{-5}	4.72
Acetic acid	$\text{HC}_2\text{H}_3\text{O}_2 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{C}_2\text{H}_3\text{O}_2^-$		1.8×10^{-5}	4.74
Hypochlorous acid	$\text{HOCl} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OCl}^-$		2.9×10^{-8}	7.54
Hydrocyanic acid	$\text{HCN} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{CN}^-$		6.2×10^{-10}	9.21
Phenol	$\text{HOC}_6\text{H}_5 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{C}_6\text{H}_5\text{O}^-$		1.0×10^{-10}	10.00
Hydrogen peroxide	$\text{H}_2\text{O}_2 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{HO}_2^-$		1.8×10^{-12}	11.74
Base			$K_b =$	$pK_b =$
Diethylamine	$(\text{C}_2\text{H}_5)_2\text{NH} + \text{H}_2\text{O} \rightleftharpoons (\text{C}_2\text{H}_5)_2\text{NH}_2^+ + \text{OH}^-$		6.9×10^{-4}	3.16
Ethylamine	$\text{C}_2\text{H}_5\text{NH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{C}_2\text{H}_5\text{NH}_3^+ + \text{OH}^-$		4.3×10^{-4}	3.37
Ammonia	$\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$		1.8×10^{-5}	4.74
Hydroxylamine	$\text{HONH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{HONH}_3^+ + \text{OH}^-$		9.1×10^{-9}	8.04
Pyridine	$\text{C}_5\text{H}_5\text{N} + \text{H}_2\text{O} \rightleftharpoons \text{C}_5\text{H}_5\text{NH}^+ + \text{OH}^-$		1.5×10^{-9}	8.82
Aniline	$\text{C}_6\text{H}_5\text{NH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{C}_6\text{H}_5\text{NH}_3^+ + \text{OH}^-$		7.4×10^{-10}	9.13

Acid strength

Base strength

EXAMPLE 16-5

Determining a Value of K_A from the pH of a Solution of a Weak Acid. Butyric acid, $\text{HC}_4\text{H}_7\text{O}_2$ (or $\text{CH}_3\text{CH}_2\text{CH}_2\text{CO}_2\text{H}$) is used to make compounds employed in artificial flavorings and syrups. A 0.250 M aqueous solution of $\text{HC}_4\text{H}_7\text{O}_2$ is found to have a pH of 2.72. Determine K_A for butyric acid.



$$K_a = ?$$

EXAMPLE 16-5

Solution:

For $\text{HC}_4\text{H}_7\text{O}_2$ K_A is likely to be much larger than K_W . Therefore assume self-ionization of water is unimportant.



Initial conc.	0.250 M	0	0
Changes	$-x$ M	$+x$ M	$+x$ M
Equilibrium Concentration	$(0.250-x)$ M	x M	x M

EXAMPLE 16-5



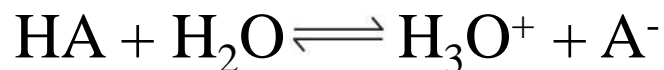
$$\text{Log}[\text{H}_3\text{O}^+] = -\text{pH} = -2.72$$

$$[\text{H}_3\text{O}^+] = 10^{-2.72} = 1.9 \times 10^{-3} = x$$

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{C}_4\text{H}_7\text{O}_2^-]}{[\text{HC}_4\text{H}_7\text{O}_2]} = \frac{1.9 \times 10^{-3} \cdot 1.9 \times 10^{-3}}{(0.250 - 1.9 \times 10^{-3})}$$

$$K_a = 1.5 \times 10^{-5} \quad \text{Check assumption: } K_a \gg K_w.$$

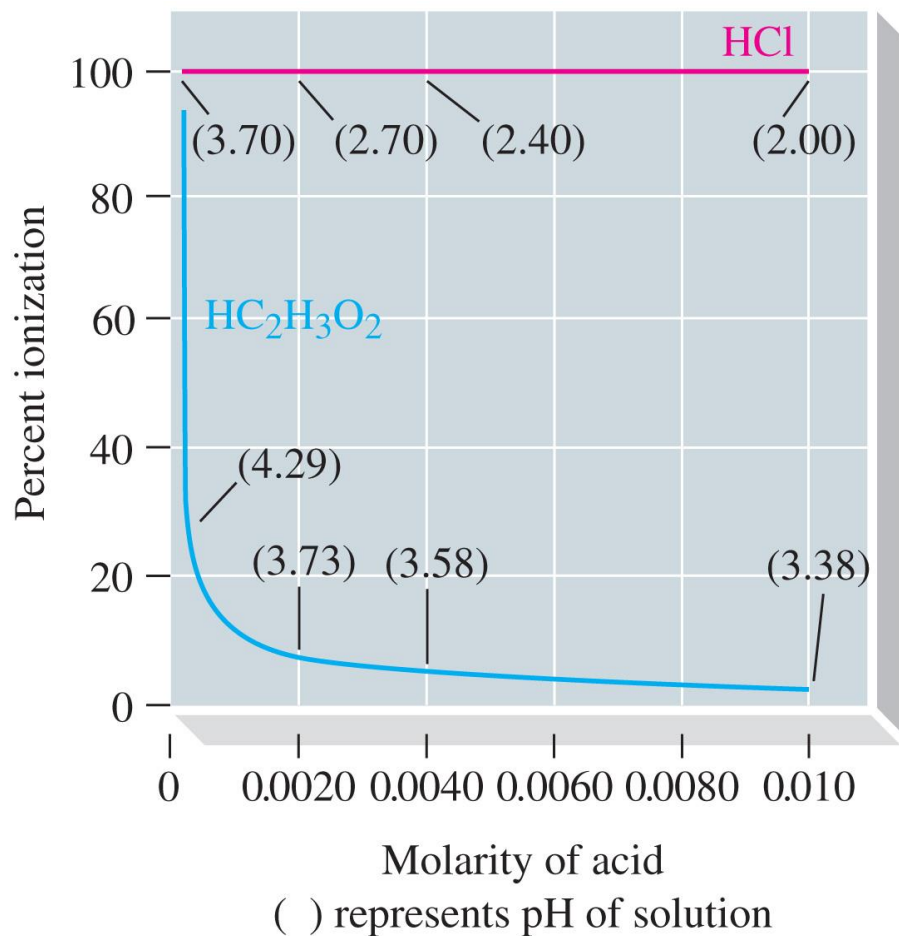
Percent Ionization



$$\text{Degree of ionization} = \frac{[\text{H}_3\text{O}^+] \text{ from HA}}{[\text{HA}] \text{ originally}}$$

$$\text{Percent ionization} = \frac{[\text{H}_3\text{O}^+] \text{ from HA}}{[\text{HA}] \text{ originally}} \times 100\%$$

Percent Ionization



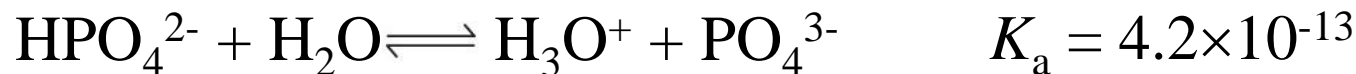
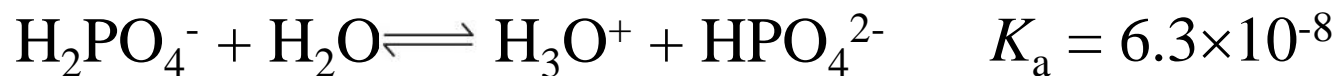
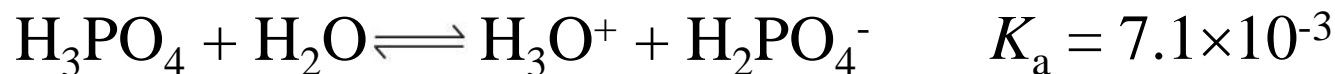
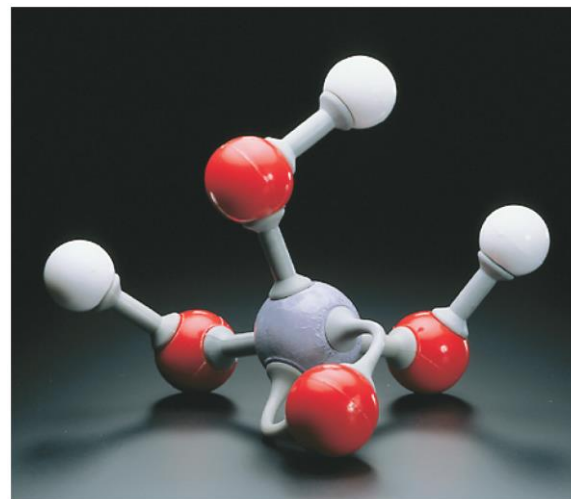
$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

$$K_a = \frac{n_{\text{H}_3\text{O}^+} n_{\text{A}^-}}{n_{\text{HA}}} \frac{1}{V}$$

16-6 Polyprotic Acids

Phosphoric acid:

A triprotic acid.



Phosphoric Acid

- ◆ $K_{a1} \gg K_{a2}$
 - All H_3O^+ is formed in the first ionization step.
- ◆ H_2PO_4^- essentially does not ionize further.
 - Assume $[\text{H}_2\text{PO}_4^-] = [\text{H}_3\text{O}^+]$.
- ◆ $[\text{HPO}_4^{2-}] \approx K_{a2}$ regardless of solution molarity.

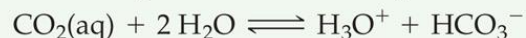
TABLE 16.4 Ionization Constants of Some Polyprotic Acids

Acid	Ionization Equilibria	Ionization Constants, K	pK
Hydrosulfuric ^a	$\text{H}_2\text{S} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{HS}^-$	$K_{a_1} = 1.0 \times 10^{-7}$	$pK_{a_1} = 7.00$
	$\text{HS}^- + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{S}^{2-}$	$K_{a_2} = 1 \times 10^{-19}$	$pK_{a_2} = 19.0$
Carbonic ^b	$\text{H}_2\text{CO}_3 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{HCO}_3^-$	$K_{a_1} = 4.4 \times 10^{-7}$	$pK_{a_1} = 6.36$
	$\text{HCO}_3^- + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{CO}_3^{2-}$	$K_{a_2} = 4.7 \times 10^{-11}$	$pK_{a_2} = 10.33$
Phosphoric	$\text{H}_3\text{PO}_4 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{H}_2\text{PO}_4^-$	$K_{a_1} = 7.1 \times 10^{-3}$	$pK_{a_1} = 2.15$
	$\text{H}_2\text{PO}_4^- + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{HPO}_4^{2-}$	$K_{a_2} = 6.3 \times 10^{-8}$	$pK_{a_2} = 7.20$
	$\text{HPO}_4^{2-} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{PO}_4^{3-}$	$K_{a_3} = 4.2 \times 10^{-13}$	$pK_{a_3} = 12.38$
Sulfurous ^c	$\text{H}_2\text{SO}_3 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{HSO}_3^-$	$K_{a_1} = 1.3 \times 10^{-2}$	$pK_{a_1} = 1.89$
	$\text{HSO}_3^- + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{SO}_3^{2-}$	$K_{a_2} = 6.2 \times 10^{-8}$	$pK_{a_2} = 7.21$
Sulfuric ^d	$\text{H}_2\text{SO}_4 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{HSO}_4^-$	$K_{a_1} = \text{very large}$	$pK_{a_1} < 0$
	$\text{HSO}_4^- + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{SO}_4^{2-}$	$K_{a_2} = 1.1 \times 10^{-2}$	$pK_{a_2} = 1.96$

Acid strength

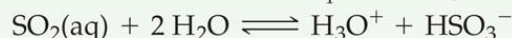
^aThe value for K_{a_2} of H_2S most commonly found in older literature is about 1×10^{-14} , but current evidence suggests that the value is considerably smaller.

^b H_2CO_3 cannot be isolated. It is in equilibrium with H_2O and dissolved CO_2 . The value given for K_{a_1} is actually for the reaction



Generally, aqueous solutions of CO_2 are treated *as if* the $\text{CO}_2(\text{aq})$ were first converted to H_2CO_3 , followed by ionization of the H_2CO_3 .

^c H_2SO_3 is a hypothetical, nonisolatable species. The value listed for K_{a_1} is actually for the reaction



^d H_2SO_4 is completely ionized in the first step.

EXAMPLE 16-9

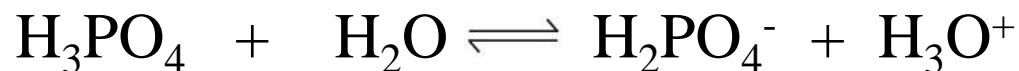
Calculating Ion Concentrations in a Polyprotic Acid

Solution. For a 3.0 M H_3PO_4 solution, calculate:

(a) $[\text{H}_3\text{O}^+]$; (b) $[\text{H}_2\text{PO}_4^-]$; (c) $[\text{HPO}_4^{2-}]$ (d) $[\text{PO}_4^{3-}]$

	H_3PO_4	$+$	H_2O	\rightleftharpoons	H_2PO_4^-	$+$	H_3O^+
Initial conc.	3.0 M				0		0
Changes	$-x$ M				$+x$ M		$+x$ M
Equilibrium Concentration	$(3.0-x)$ M				x M		x M

EXAMPLE 16-9



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{H}_2\text{PO}_4^-]}{[\text{H}_3\text{PO}_4]} = \frac{x \cdot x}{(3.0 - x)} = 7.1 \times 10^{-3}$$

Assume that $x \ll 3.0$

$$x^2 = (3.0)(7.1 \times 10^{-3}) \qquad x = 0.14 \text{ M}$$

$$[\text{H}_2\text{PO}_4^-] = [\text{H}_3\text{O}^+] = 0.14 \text{ M}$$

EXAMPLE 16-9



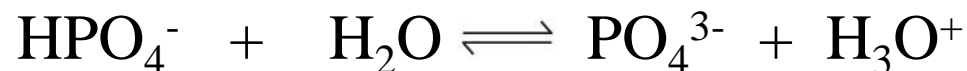
Initial conc.	0.14 M	0	0.14 M
Changes	-y M	+y M	+y M
Equilibrium Concentration	(0.14 - y) M	y M	(0.14 + y) M

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]} = \frac{y \cdot (0.14 + y)}{(0.14 - y)} = 6.3 \times 10^{-8}$$

$$y \ll 0.14 \text{ M}$$

$$y = [\text{HPO}_4^{2-}] = 6.3 \times 10^{-8}$$

EXAMPLE 16-9



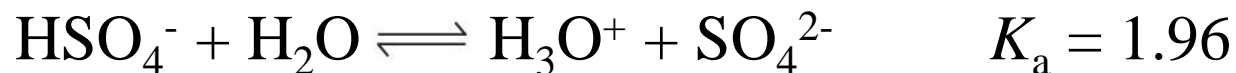
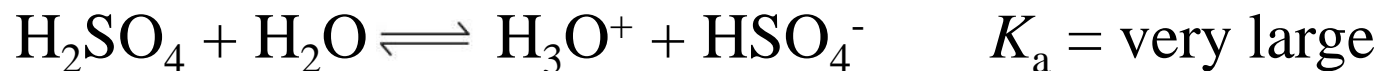
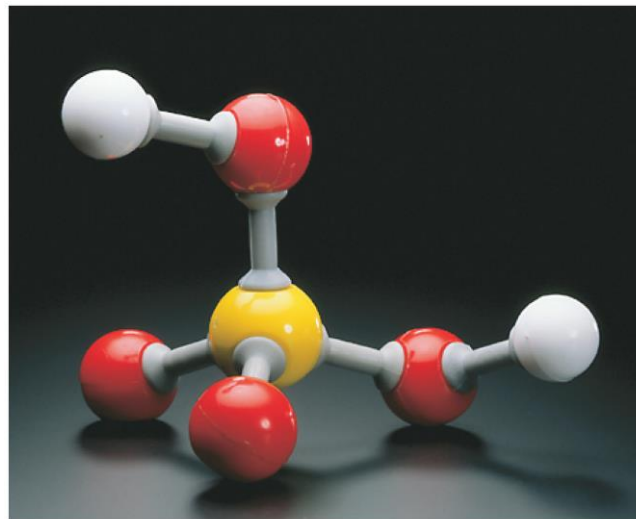
$$K_a = \frac{[\text{H}_3\text{O}^+][\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]} = \frac{(0.14)[\text{PO}_4^{3-}]}{6.3 \times 10^{-8}} = 4.2 \times 10^{-13} \text{ M}$$

$$[\text{PO}_4^{3-}] = 1.9 \times 10^{-19} \text{ M}$$

Sulfuric Acid

Sulfuric acid:

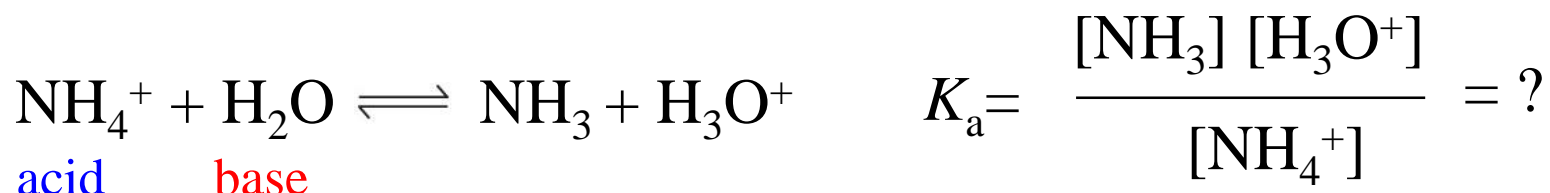
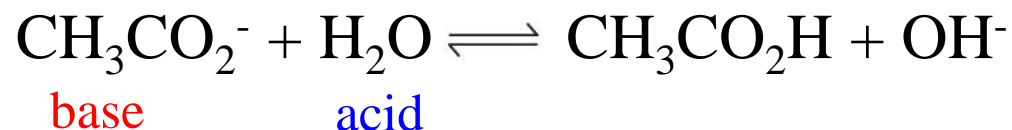
A diprotic acid.



General Approach to Solution Equilibrium Calculations

- ◆ Identify species present in any significant amounts in solution (excluding H_2O).
- ◆ Write equations that include these species.
 - Number of equations = number of unknowns.
 - Equilibrium constant expressions.
 - Material balance equations.
 - Electroneutrality condition.
- ◆ Solve the system of equations for the unknowns.

16-7 Ions as Acids and Bases

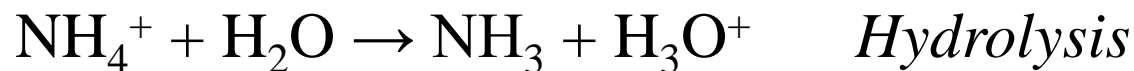


$$K_a = \frac{[\text{NH}_3] [\text{H}_3\text{O}^+] [\text{OH}^-]}{[\text{NH}_4^+] [\text{OH}^-]} = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$$

$$K_a K_b = K_w$$

Hydrolysis

- ◆ Water (*hydro*) causing cleavage (*lysis*) of a bond.



16-8 Molecular Structure and Acid-Base Behavior

- ◆ Why is HCl a strong acid, but HF is a weak one?
- ◆ Why is $\text{CH}_3\text{CO}_2\text{H}$ a stronger acid than $\text{CH}_3\text{CH}_2\text{OH}$?
- ◆ There is a relationship between molecular structure and acid strength.
- ◆ Bond dissociation energies are measured in the gas phase and *not* in solution.

Strengths of Binary Acids

	HI		HBr		HCl		HF	
<i>Bond length</i>	160.9	>	141.4	>	127.4	>	91.7	pm
<i>Bond energy</i>	297	<	368	<	431	<	569	kJ/mol
<i>Acid strength</i>	10^9	>	10^8	>	1.3×10^6	>>	6.6×10^{-4}	



ion pair *free ions*
H-bonding

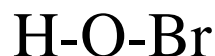
Strengths of Oxoacids

- ◆ Factors promoting electron withdrawal from the OH bond to the oxygen atom:
 - High electronegativity (EN) of the central atom.
 - A large number of terminal O atoms in the molecule.



$$EN_{\text{Cl}} = 3.0$$

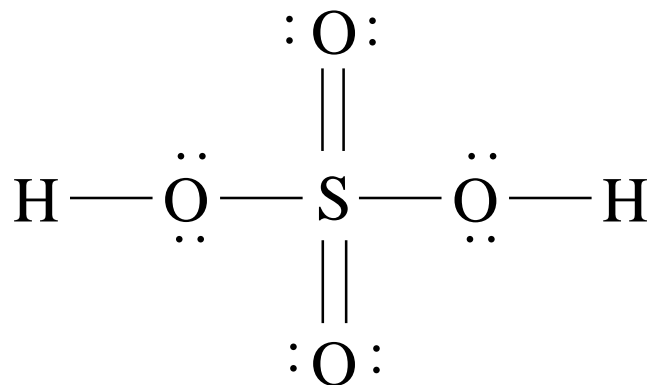
$$K_{\text{a}} = 2.9 \times 10^{-8}$$



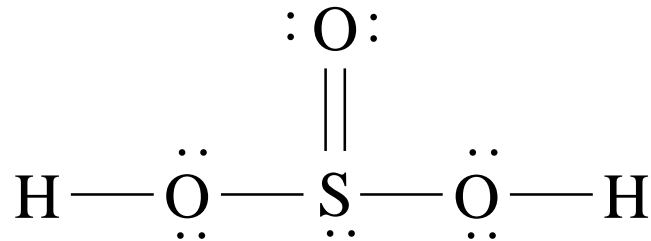
$$EN_{\text{Br}} = 2.8$$

$$K_{\text{a}} = 2.1 \times 10^{-9}$$

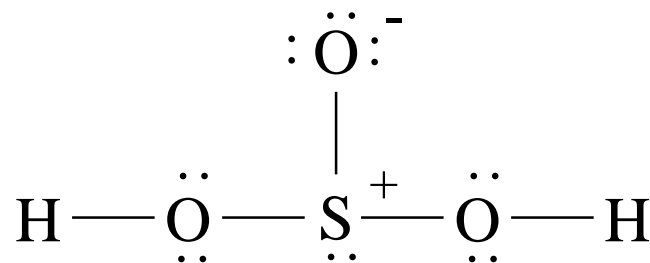
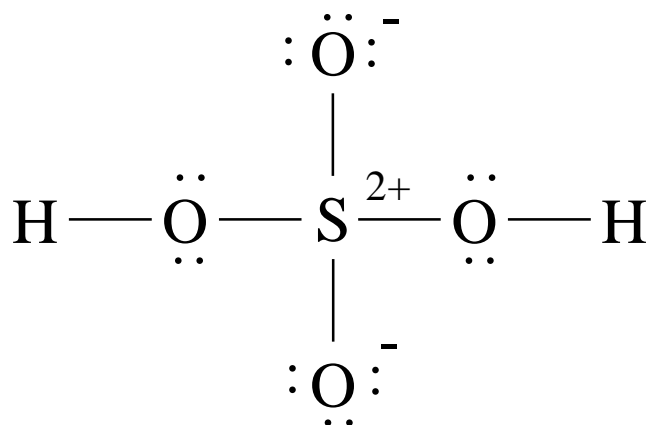
Strengths of Oxoacids



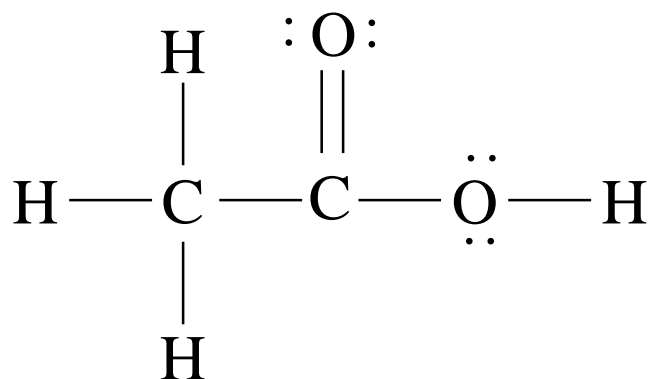
$$K_a \approx 10^3$$



$$K_a = 1.3 \times 10^{-2}$$

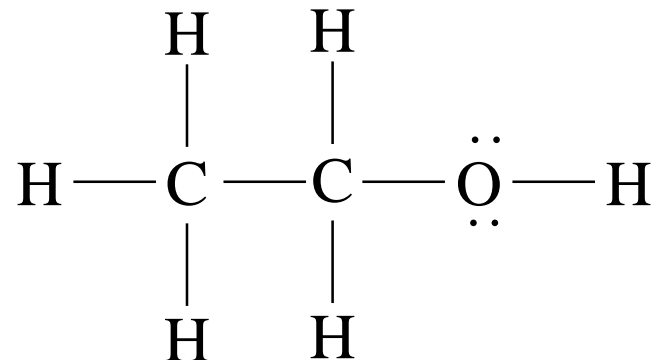


Strengths of Organic Acids



acetic acid

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

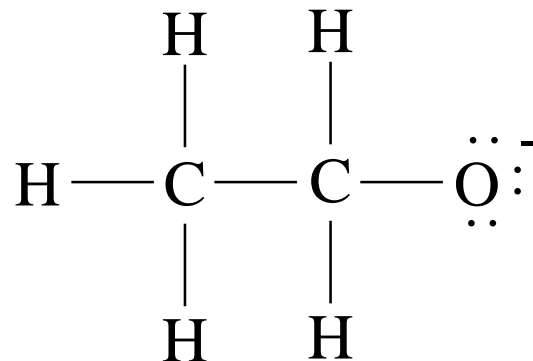


ethanol

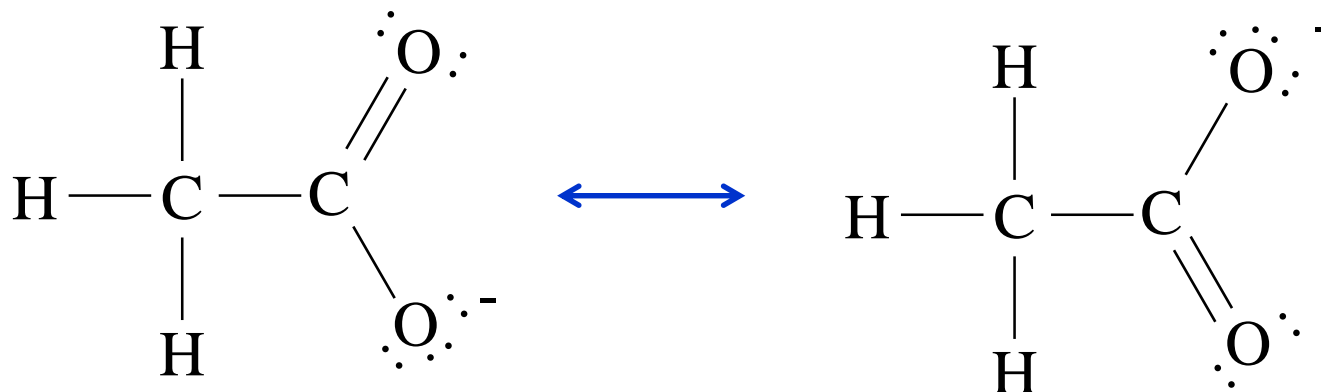
$$K_a = 1.3 \times 10^{-16}$$

Focus on the Anions Formed

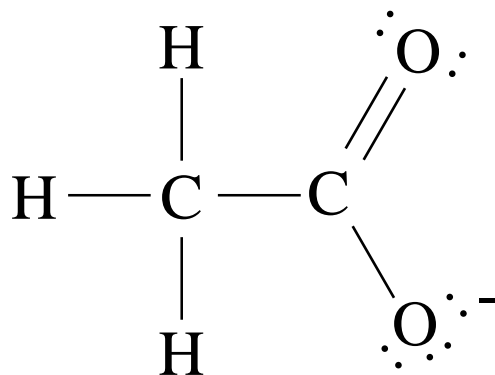
Ethoxide



Acetate

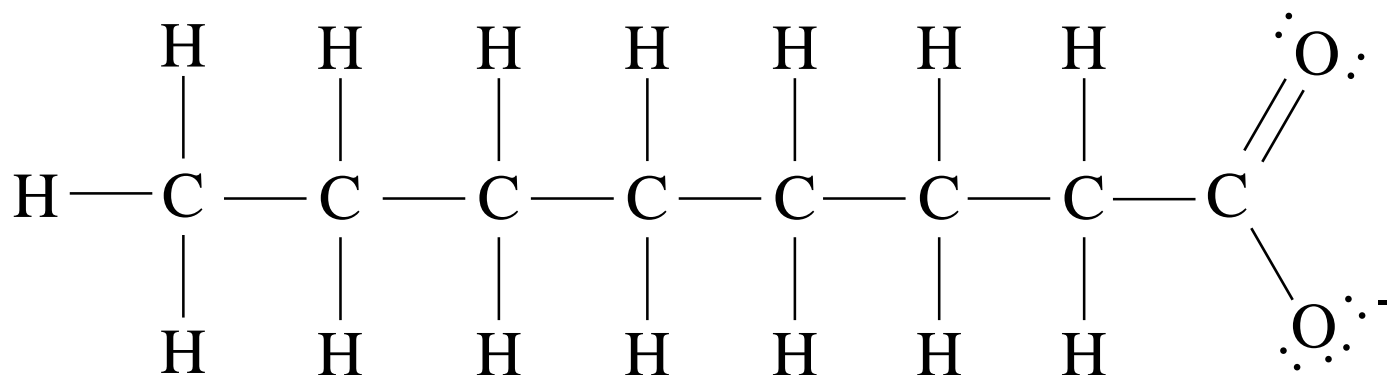


Structural Effects

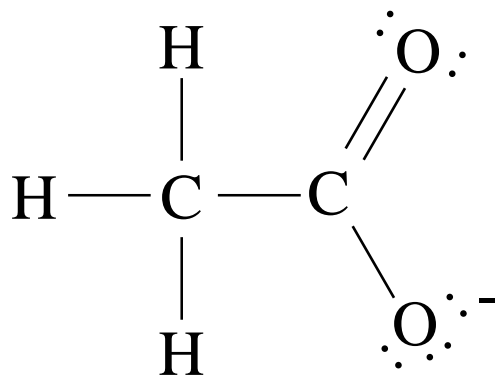


$$K_a = 1.8 \times 10^{-5}$$

$$K_a = 1.3 \times 10^{-5}$$

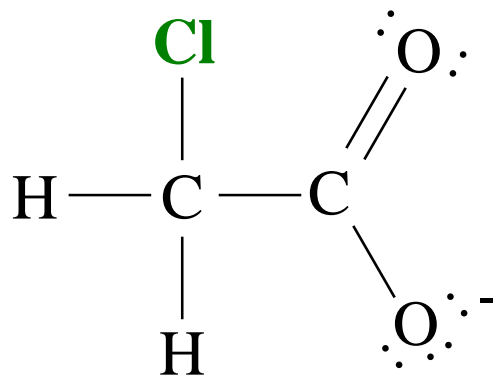


Structural Effects

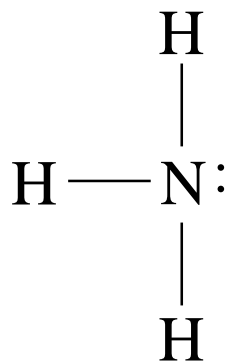


$$K_a = 1.8 \times 10^{-5}$$

$$K_a = 1.4 \times 10^{-3}$$

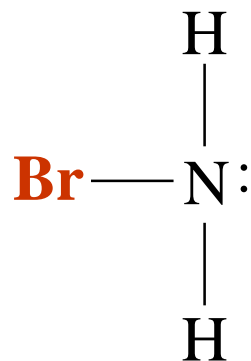


Strengths of Amines as Bases



ammonia

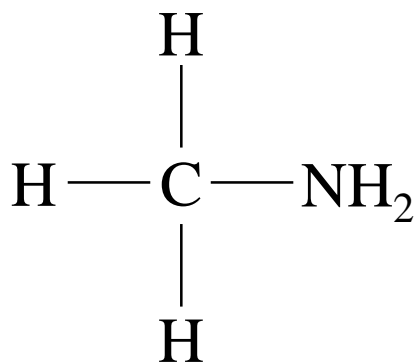
$$\text{p}K_{\text{b}} = 4.74$$



bromamine

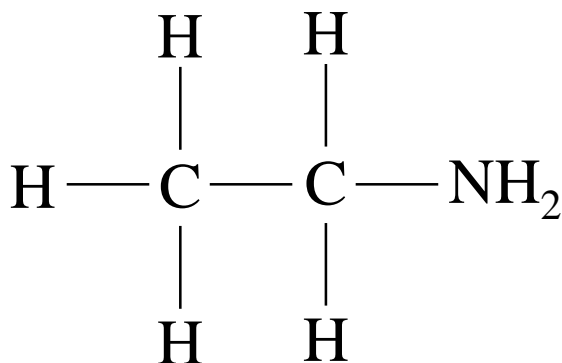
$$\text{p}K_{\text{a}} = 7.61$$

Strengths of Amines as Bases



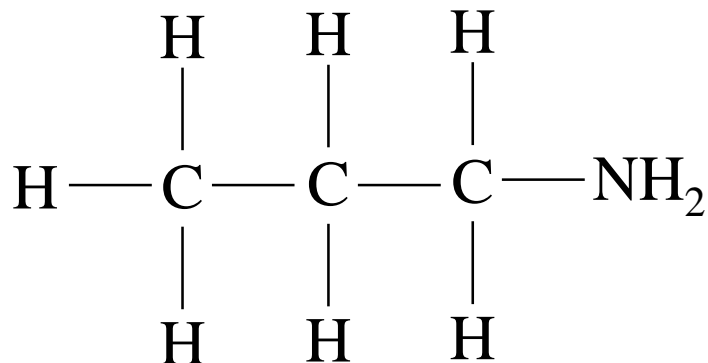
methylamine

$$\text{p}K_{\text{b}} = 4.74$$



ethylamine

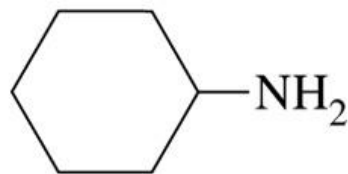
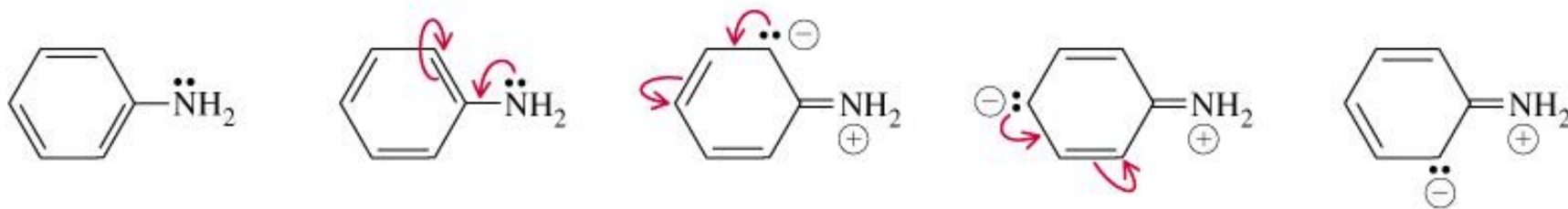
$$\text{p}K_{\text{a}} = 3.38$$



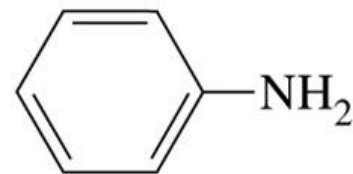
propylamine

$$\text{p}K_{\text{b}} = 3.37$$

Resonance Effects



Cyclohexylamine, $\text{p}K_{\text{b}} = 3.36$

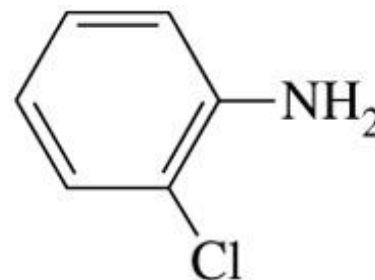


Aniline, $\text{p}K_{\text{b}} = 9.13$

Inductive Effects



para-Chloroaniline, $pK_b = 10.01$



ortho-Chloroaniline, $pK_b = 11.36$

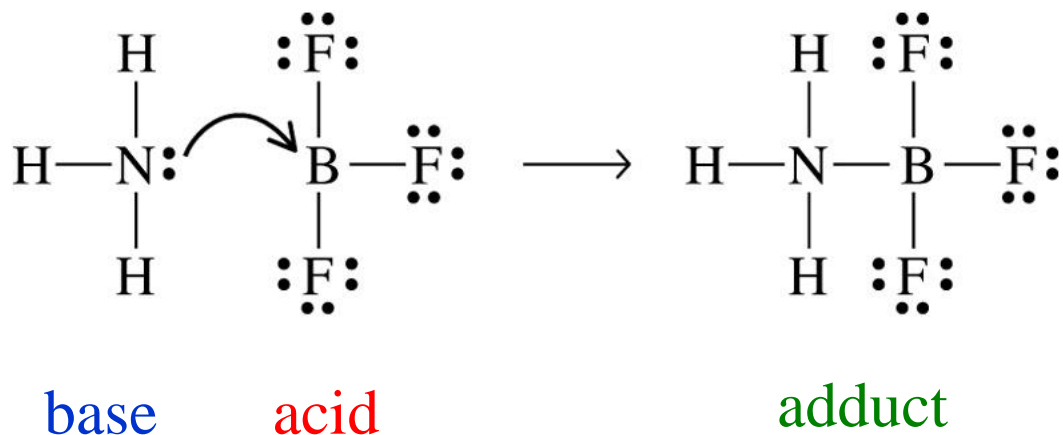
16-9 Lewis Acids and Bases

◆ Lewis Acid

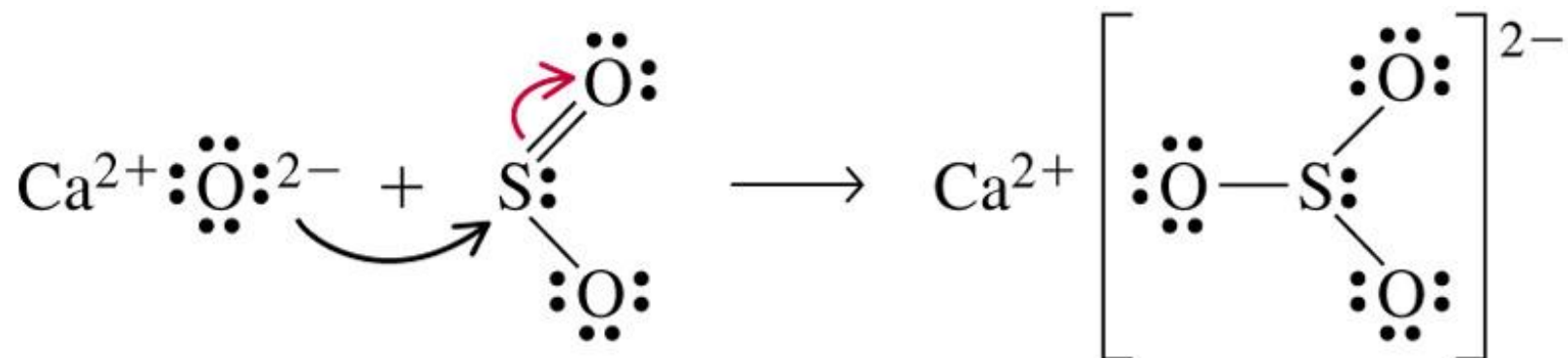
- A species (atom, ion or molecule) that is an electron pair *acceptor*.

◆ Lewis Base

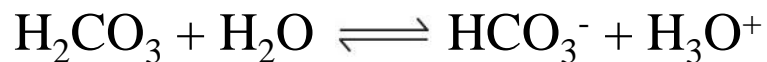
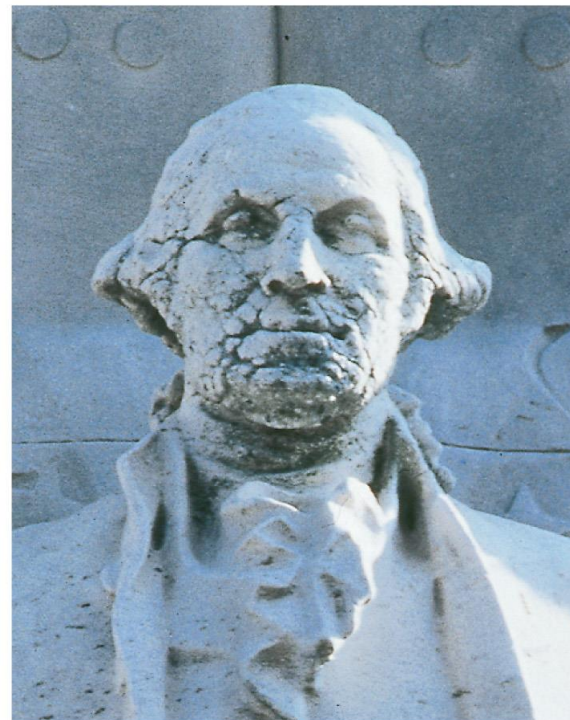
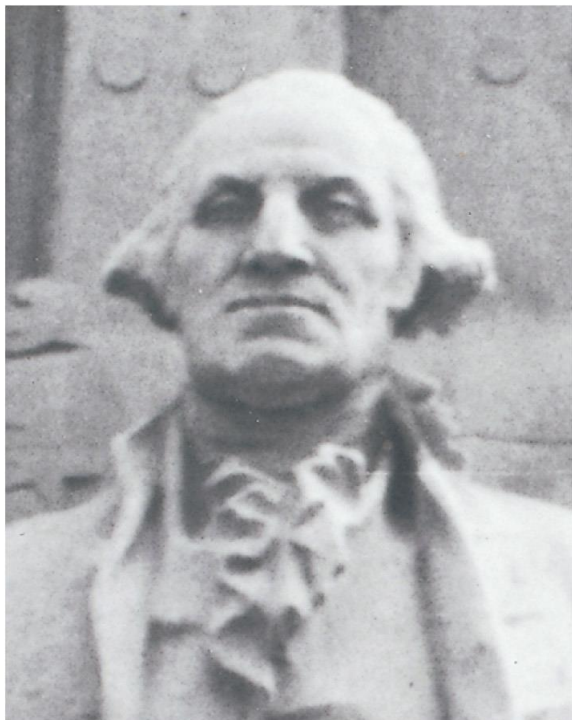
- A species that is an electron pair *donor*.



Showing Electron Movement



Focus On Acid Rain



End of Chapter Questions

- ◆ When you see the correct approach to a problem after going down various wrong paths, you learn about the decision making process that leads to good solutions.
- ◆ Do not go straight to a solution manual after only one attempt at a problem. You will not learn how to make decisions.