Acids and Bases

Chapter 15

1

Acids

Have a sour taste. Vinegar owes its taste to acetic acid. Citrus fruits contain citric acid.

React with certain metals to produce hydrogen gas.

React with carbonates and bicarbonates to produce carbon dioxide gas

Bases

Have a bitter taste.

Feel slippery. Many soaps contain bases.

A Brønsted acid is a proton donor
A Brønsted base is a proton acceptor

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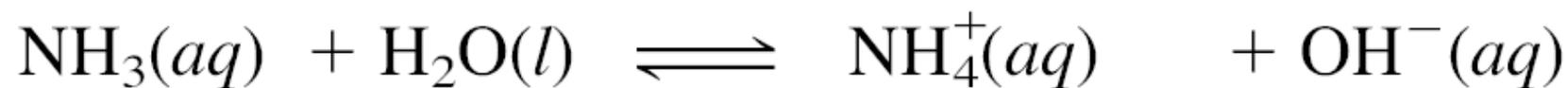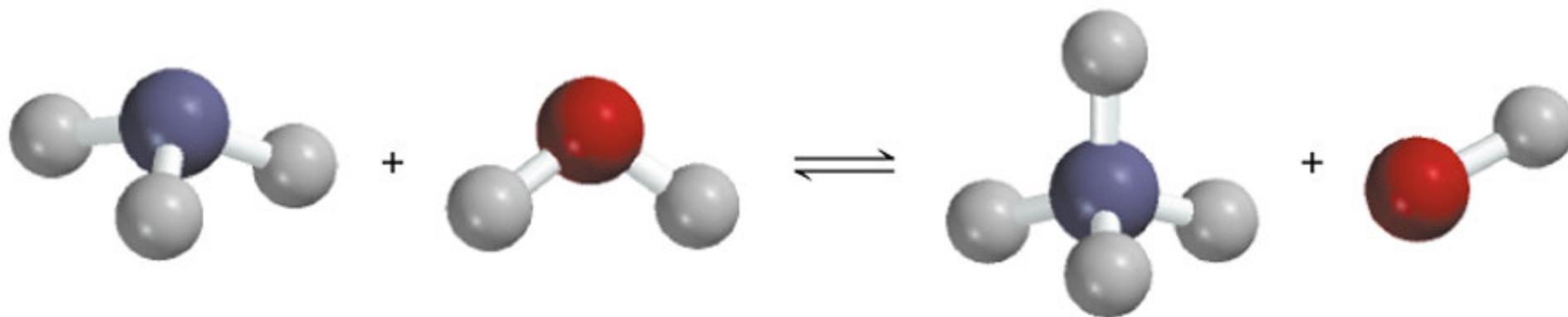
base	acid		acid	base
base	acid		conjugate acid	conjugate base

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Acid-Base Properties of Water

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

autoionization of water

base

conjugate
acid

$$H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-$$

acid

conjugate
base

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The Ion Product of Water

$$H_2O \; (l) \rightleftharpoons H^+ \; (aq) + OH^- \; (aq) \qquad K_c = \frac{[H^+][OH^-]}{[H_2O]} \qquad [H_2O] = \text{constant}$$
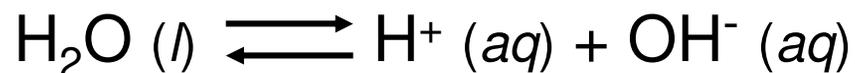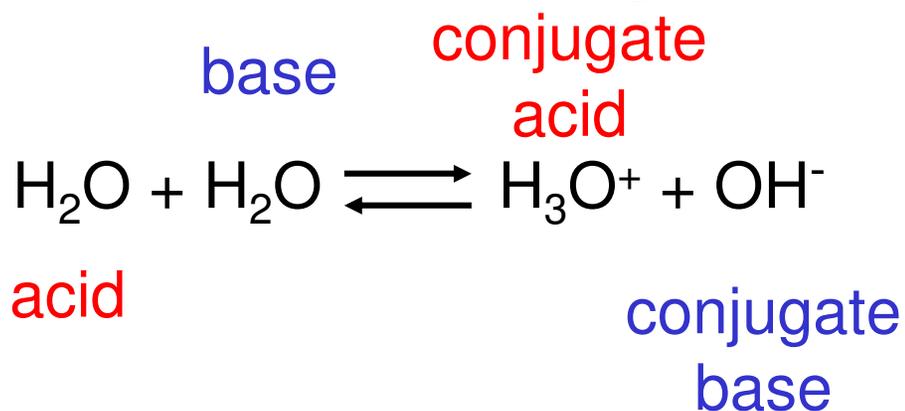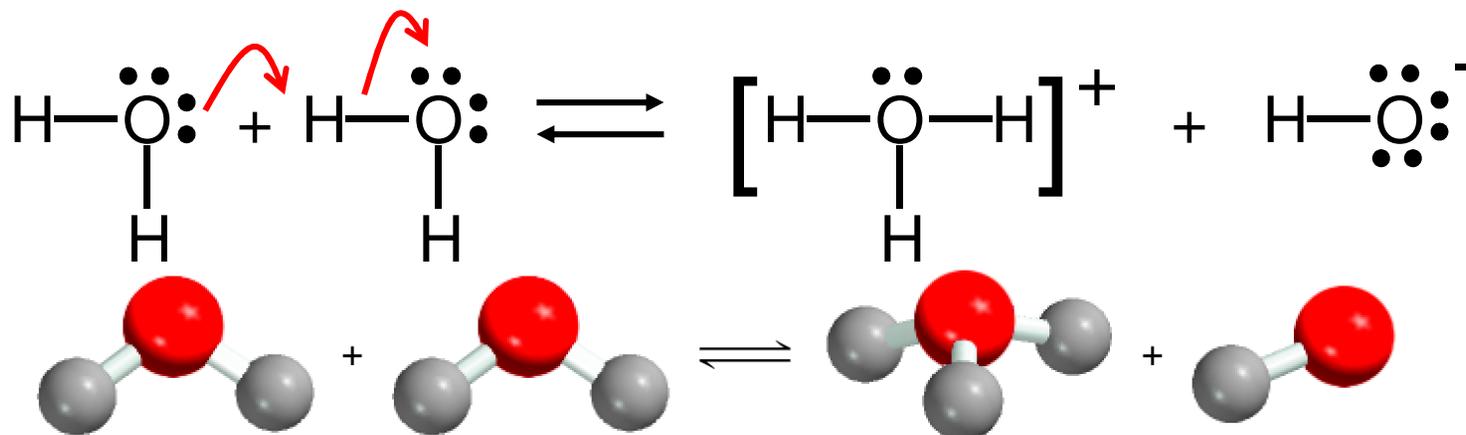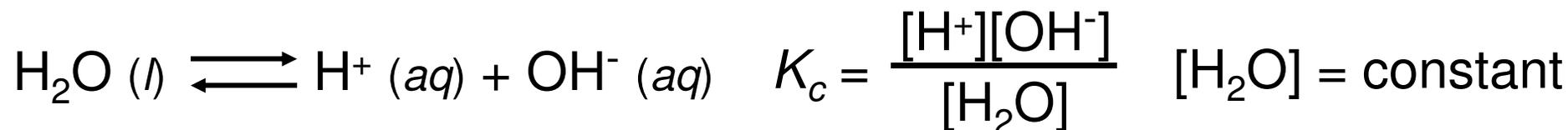
$$K_c[H_2O] = K_w = [H^+][OH^-]$$

The ***ion-product constant*** (K_w) is the product of the molar concentrations of H^+ and OH^- ions **at a particular temperature.**

	Solution Is
At 25^0C $K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$	$[H^+] = [OH^-]$ neutral $[H^+] > [OH^-]$ acidic $[H^+] < [OH^-]$ basic

What is the concentration of OH^- ions in a HCl solution whose hydrogen ion concentration is 1.3 M?

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

$$[H^+] = 1.3 \ M$$

$$[OH^-] = \frac{K_w}{[H^+]} = \frac{1 \times 10^{-14}}{1.3} = 7.7 \times 10^{-15} \ M$$

pH – A Measure of Acidity

$$pH = -\log [H^+]$$

Solution Is		**At 25^{0}C**	
neutral	$[H^+] = [OH^-]$	$[H^+] = 1 \times 10^{-7}$	$pH = 7$
acidic	$[H^+] > [OH^-]$	$[H^+] > 1 \times 10^{-7}$	$pH < 7$
basic	$[H^+] < [OH^-]$	$[H^+] < 1 \times 10^{-7}$	$pH > 7$

pH ↑ $[H^+]$ ↓

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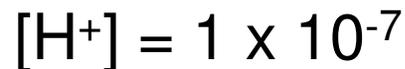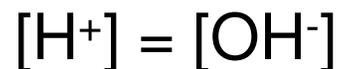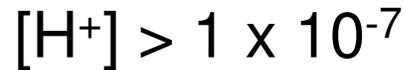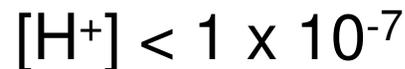
TABLE 15.1

The pHs of Some Common Fluids

Sample	pH Value
Gastric juice in the stomach	1.0–2.0
Lemon juice	2.4
Vinegar	3.0
Grapefruit juice	3.2
Orange juice	3.5
Urine	4.8–7.5
Water exposed to air*	5.5
Saliva	6.4–6.9
Milk	6.5
Pure water	7.0
Blood	7.35–7.45
Tears	7.4
Milk of magnesia	10.6
Household ammonia	11.5

*Water exposed to air for a long period of time absorbs atmospheric CO_2 to form carbonic acid, H_2CO_3.

Other important relationships

$$pOH = -\log [OH^-]$$

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

$$-\log [H^+] - \log [OH^-] = 14.00$$

$$pH + pOH = 14.00$$

pH Meter

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The pH of rainwater collected in a certain region of the northeastern United States on a particular day was 4.82. What is the H+ ion concentration of the rainwater?

$$pH = -\log [H^+]$$

$$[H^+] = 10^{-pH} = 10^{-4.82} = 1.5 \times 10^{-5} \; M$$

The OH- ion concentration of a blood sample is $2.5 \times 10^{-7} \; M$. What is the pH of the blood?

$$pH + pOH = 14.00$$

$$pOH = -\log [OH^-] = -\log (2.5 \times 10^{-7}) = 6.60$$

$$pH = 14.00 - pOH = 14.00 - 6.60 = 7.40$$

Strong Electrolyte – 100% dissociation

$$NaCl \ (s) \ \xrightarrow{\text{H}_2\text{O}} \ Na^+ \ (aq) + Cl^- \ (aq)$$

Weak Electrolyte – not completely dissociated

$$CH_3COOH \rightleftharpoons CH_3COO^- \ (aq) + H^+ \ (aq)$$
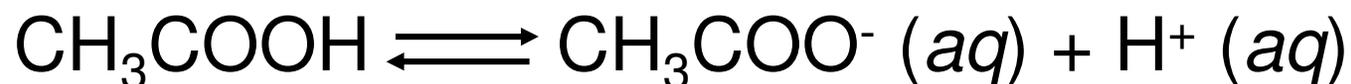
Strong Acids are strong electrolytes

$$HCl \ (aq) + H_2O \ (l) \longrightarrow H_3O^+ \ (aq) + Cl^- \ (aq)$$

$$HNO_3 \ (aq) + H_2O \ (l) \longrightarrow H_3O^+ \ (aq) + NO_3^- \ (aq)$$

$$HClO_4 \ (aq) + H_2O \ (l) \longrightarrow H_3O^+ \ (aq) + ClO_4^- \ (aq)$$

$$H_2SO_4 \ (aq) + H_2O \ (l) \longrightarrow H_3O^+ \ (aq) + HSO_4^- \ (aq)$$
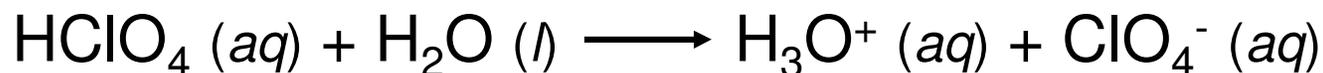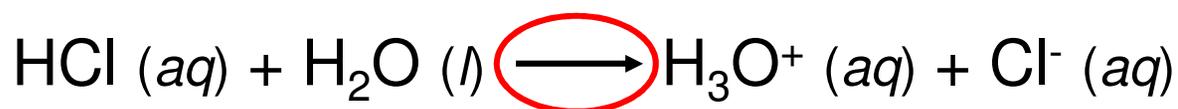
Weak Acids are weak electrolytes

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

$HNO_2\ (aq) + H_2O\ (l) \rightleftharpoons H_3O^+\ (aq) + NO_2^-\ (aq)$

$HSO_4^-\ (aq) + H_2O\ (l) \rightleftharpoons H_3O^+\ (aq) + SO_4^{2-}\ (aq)$

$H_2O\ (l) + H_2O\ (l) \rightleftharpoons H_3O^+\ (aq) + OH^-\ (aq)$
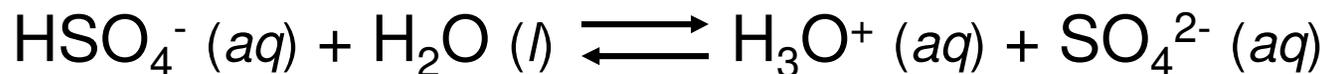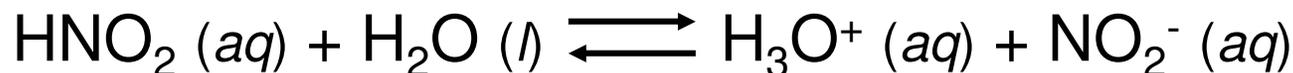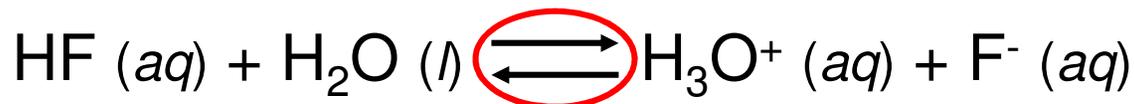
Strong Bases are strong electrolytes

$NaOH\ (s) \xrightarrow{H_2O} Na^+\ (aq) + OH^-\ (aq)$

$KOH\ (s) \xrightarrow{H_2O} K^+\ (aq) + OH^-\ (aq)$

$Ba(OH)_2\ (s) \xrightarrow{H_2O} Ba^{2+}\ (aq) + 2OH^-\ (aq)$

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**Weak Bases** are weak electrolytes

$$F^- \ (aq) + H_2O \ (l) \rightleftharpoons OH^- \ (aq) + HF \ (aq)$$

$$NO_2^- \ (aq) + H_2O \ (l) \rightleftharpoons OH^- \ (aq) + HNO_2 \ (aq)$$
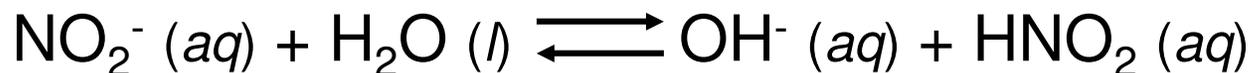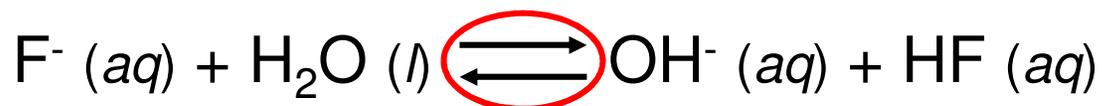
Conjugate acid-base pairs:

- The conjugate base of a strong acid has no measurable strength.

- H_3O^+ is the strongest acid that can exist in aqueous solution.

- The OH^- ion is the strongest base that can exist in aqeous solution.

TABLE 15.2 Relative Strengths of Conjugate Acid-Base Pairs

	Acid	Conjugate Base
Strong acids	$HClO_4$ (perchloric acid)	ClO_4^- (perchlorate ion)
	HI (hydroiodic acid)	I^- (iodide ion)
	HBr (hydrobromic acid)	Br^- (bromide ion)
	HCl (hydrochloric acid)	Cl^- (chloride ion)
	H_2SO_4 (sulfuric acid)	HSO_4^- (hydrogen sulfate ion)
	HNO_3 (nitric acid)	NO_3^- (nitrate ion)
	H_3O^+ (hydronium ion)	H_2O (water)
Weak acids	HSO_4^- (hydrogen sulfate ion)	SO_4^{2-} (sulfate ion)
	HF (hydrofluoric acid)	F^- (fluoride ion)
	HNO_2 (nitrous acid)	NO_2^- (nitrite ion)
	HCOOH (formic acid)	$HCOO^-$ (formate ion)
	CH_3COOH (acetic acid)	CH_3COO^- (acetate ion)
	NH_4^+ (ammonium ion)	NH_3 (ammonia)
	HCN (hydrocyanic acid)	CN^- (cyanide ion)
	H_2O (water)	OH^- (hydroxide ion)
	NH_3 (ammonia)	NH_2^- (amide ion)

Acid strength increases →

Base strength increases →

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Strong Acid (HCl)

Weak Acid (HF)

Cl⁻

HF

F⁻

H_2O

H_3O^+
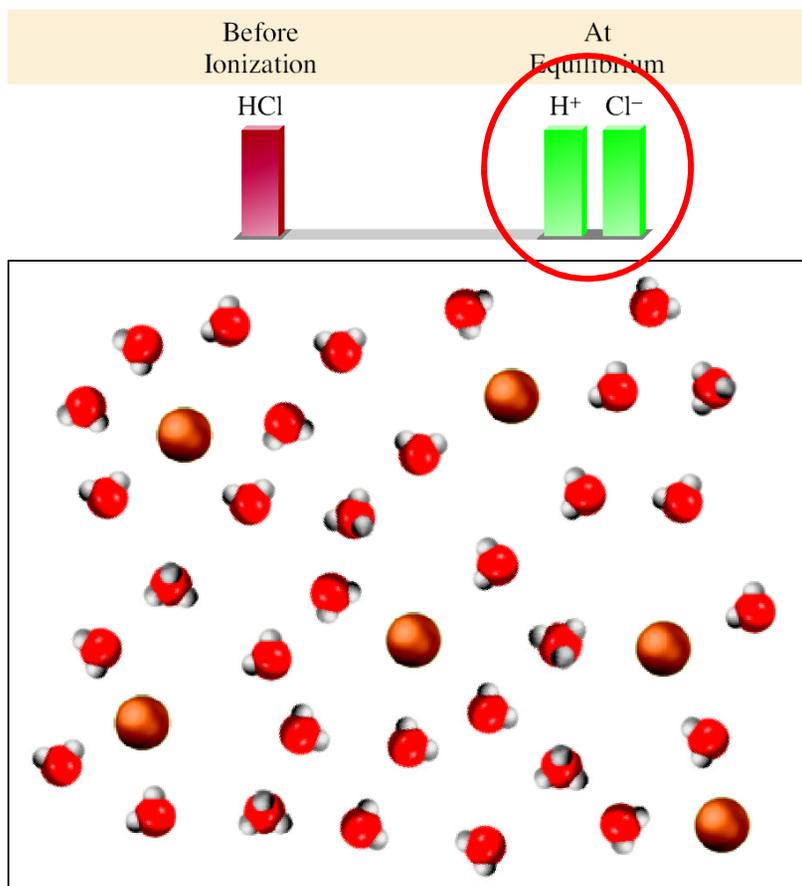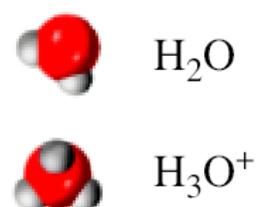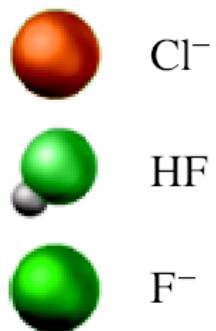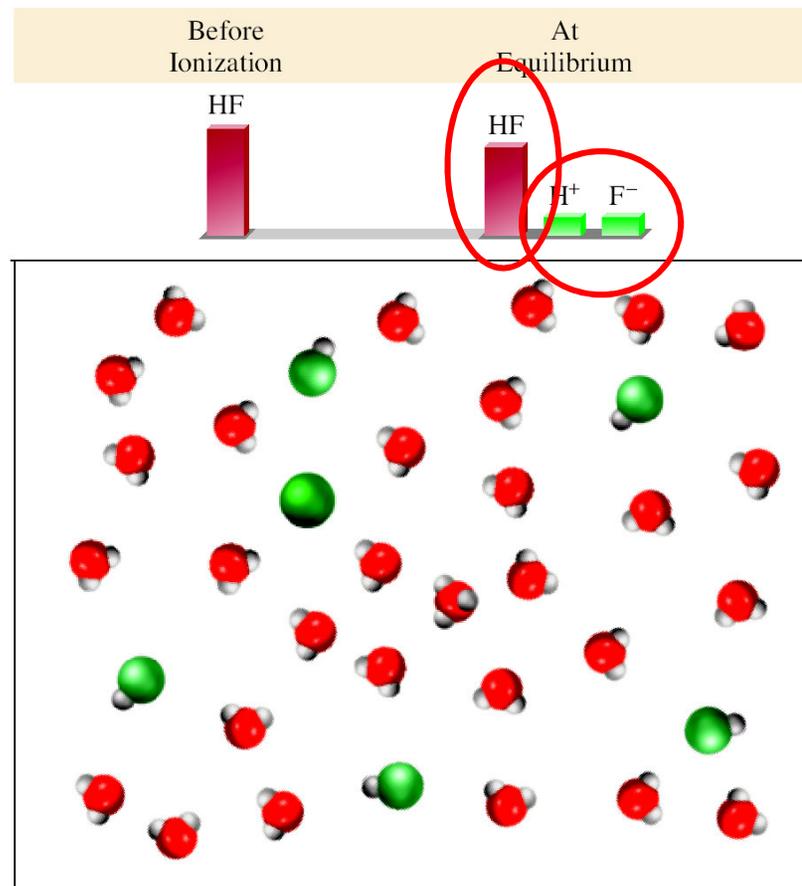
What is the pH of a 2×10^{-3} M HNO_3 solution?

HNO_3 is a strong acid – 100% dissociation.

Start   0.002 M 0.0 M 0.0 M

HNO_3 (aq) + H_2O (l) $\longrightarrow$ H_3O^+ (aq) + NO_3^- (aq)

End   0.0 M 0.002 M 0.002 M
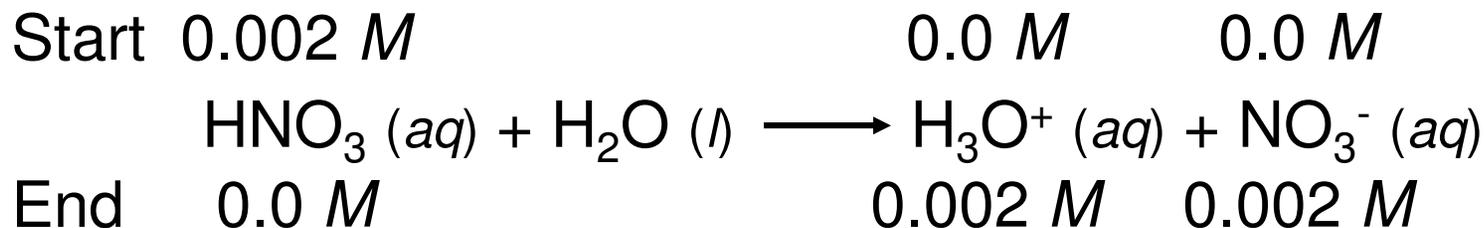
$$pH = -\log[H^+] = -\log[H_3O^+] = -\log(0.002) = 2.7$$

What is the pH of a 1.8×10^{-2} M $Ba(OH)_2$ solution?

$Ba(OH)_2$ is a strong base – 100% dissociation.

Start   0.018 M 0.0 M 0.0 M

$Ba(OH)_2$ (s) $\longrightarrow$ Ba^{2+} (aq) + $2OH^-$ (aq)

End   0.0 M 0.018 M 0.036 M
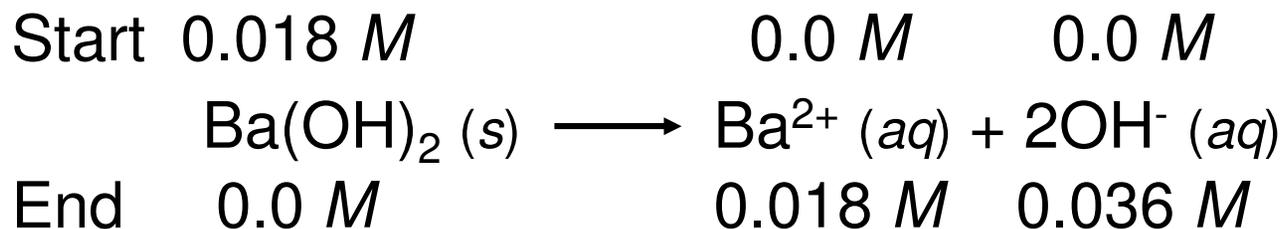
$$pH = 14.00 - pOH = 14.00 + \log(0.036) = 12.6$$

Weak Acids (HA) and Acid Ionization Constants

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

$$HA\ (aq) \rightleftharpoons H^+\ (aq) + A^-\ (aq)$$
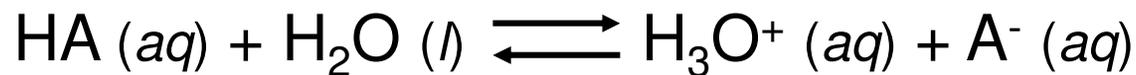
$$K_a = \frac{[H^+][A^-]}{[HA]}$$

K_a is the **acid ionization constant**

K_a ↑   weak acid strength ↑

Name of Acid	Formula	Structure	K_a	Conjugate Base	$K_b^\dagger$
Hydrofluoric acid	HF	H—F	7.1×10^{-4}	F^-	1.4×10^{-11}
Nitrous acid	HNO_2	O=N—O—H	4.5×10^{-4}	NO_2^-	2.2×10^{-11}
Acetylsalicylic acid (aspirin)	$C_9H_8O_4$		3.0×10^{-4}	$C_9H_7O_4^-$	3.3×10^{-11}
Formic acid	HCOOH		1.7×10^{-4}	$HCOO^-$	5.9×10^{-11}
Ascorbic acid*	$C_6H_8O_6$		8.0×10^{-5}	$C_6H_7O_6^-$	1.3×10^{-10}
Benzoic acid	C_6H_5COOH		6.5×10^{-5}	$C_6H_5COO^-$	1.5×10^{-10}
Acetic acid	CH_3COOH		1.8×10^{-5}	CH_3COO^-	5.6×10^{-10}
Hydrocyanic acid	HCN	H—C≡N	4.9×10^{-10}	CN^-	2.0×10^{-5}
Phenol	C_6H_5OH		1.3×10^{-10}	$C_6H_5O^-$	7.7×10^{-5}
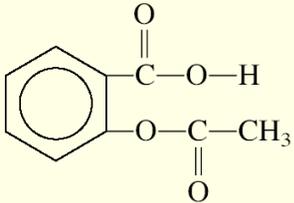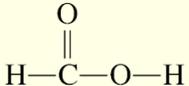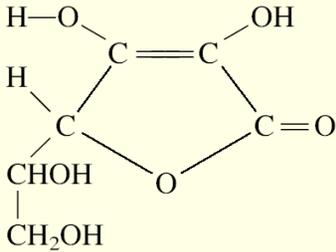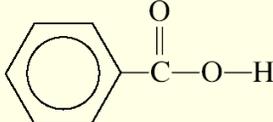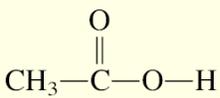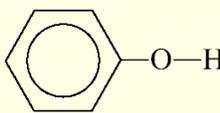
*For ascorbic acid it is the upper left hydroxyl group that is associated with this ionization constant.
†The base ionization constant K_b is discussed in Section 15.6.

What is the pH of a 0.5 M HF solution (at 25^{0}C)?

$$HF_{(aq)} \rightleftharpoons H^+_{(aq)} + F^-_{(aq)} \qquad K_a = \frac{[H^+][F^-]}{[HF]} = 7.1 \times 10^{-4}$$

$$HF_{(aq)} \rightleftharpoons H^+_{(aq)} + F^-_{(aq)}$$
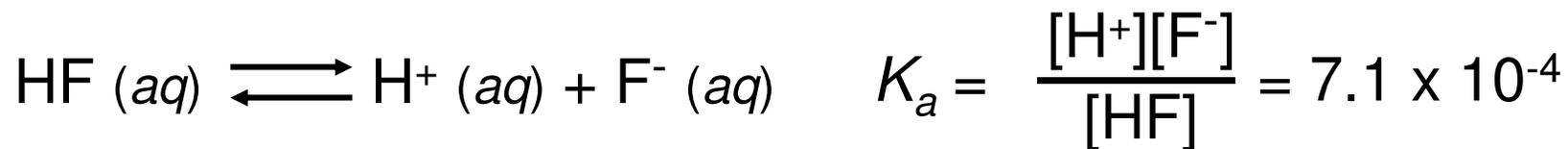
Initial (M)	0.50	0.00	0.00
Change (M)	-x	+x	+x
Equilibrium (M)	0.50 - x	x	x

$$K_a = \frac{x^2}{0.50 - x} = 7.1 \times 10^{-4} \qquad K_a << 1 \quad 0.50 - x \approx 0.50$$

$$K_a \approx \frac{x^2}{0.50} = 7.1 \times 10^{-4} \qquad x^2 = 3.55 \times 10^{-4} \qquad x = 0.019 \ M$$

$[H^+] = [F^-] = 0.019 \ M \qquad\qquad pH = -log[H^+] = 1.72$

$[HF] = 0.50 - x = 0.48 \ M$

When can I use the approximation?

$$K_a \ll 1 \quad 0.50 - x \approx 0.50$$

When x is less than 5% of the value from which it is subtracted.

$x = 0.019$ $\quad \dfrac{0.019\ M}{0.50\ M}$ x 100% = 3.8% $\qquad$ Less than 5%
Approximation ok.

What is the pH of a 0.05 M HF solution (at 25^0C)?

$$K_a \approx \dfrac{x^2}{0.05} = 7.1 \text{ x } 10^{-4} \quad x = 0.006\ M$$

$\dfrac{0.006\ M}{0.05\ M}$ x 100% = 12% $\qquad$ More than 5%
Approximation **not** ok.

Must solve for x exactly using quadratic equation or method of successive approximations.

Solving **weak acid** ionization problems:

1. Identify the major species that can affect the pH.

   - In most cases, you can ignore the autoionization of water.

   - Ignore $[OH^-]$ because it is determined by $[H^+]$.

2. Use ICE to express the equilibrium concentrations in terms of single unknown x.

3. Write K_a in terms of equilibrium concentrations. Solve for x by the approximation method. If approximation is not valid, solve for x exactly.

4. Calculate concentrations of all species and/or pH of the solution.

What is the pH of a 0.122 M monoprotic acid whose K_a is 5.7×10^{-4}?

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

	HA	H⁺	A⁻
Initial (M)	0.122	0.00	0.00
Change (M)	-x	+x	+x
Equilibrium (M)	0.122 - x	x	x

$$K_a = \frac{x^2}{0.122 - x} = 5.7 \times 10^{-4} \qquad K_a \ll 1 \quad 0.122 - x \approx 0.122$$

$$K_a \approx \frac{x^2}{0.122} = 5.7 \times 10^{-4} \qquad x^2 = 6.95 \times 10^{-5} \qquad x = 0.0083\ M$$

$$\frac{0.0083\ M}{0.122\ M} \times 100\% = 6.8\%$$

More than 5%

Approximation **not** ok.

$$K_a = \frac{x^2}{0.122 - x} = 5.7 \times 10^{-4} \qquad x^2 + 0.00057x - 6.95 \times 10^{-5} = 0$$

$$ax^2 + bx + c = 0 \qquad\qquad x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$x = 0.0081 \qquad\qquad x = -0.0081$$

$$HA_{(aq)} \rightleftharpoons H^+_{(aq)} + A^-_{(aq)}$$

	HA (aq)	H$^+$ (aq)	A$^-$ (aq)
Initial (M)	0.122	0.00	0.00
Change (M)	-x	+x	+x
Equilibrium (M)	0.122 - x	x	x

$$[H^+] = x = 0.0081 \; M \qquad\qquad pH = -\log[H^+] = 2.09$$

$$\textbf{\textit{percent ionization}} = \frac{\text{Ionized acid concentration at equilibrium}}{\text{Initial concentration of acid}} \times 100\%$$

For a monoprotic acid HA

$$\text{Percent ionization} = \frac{[H^+]}{[HA]_0} \times 100\% \qquad [HA]_0 = \text{initial concentration}$$
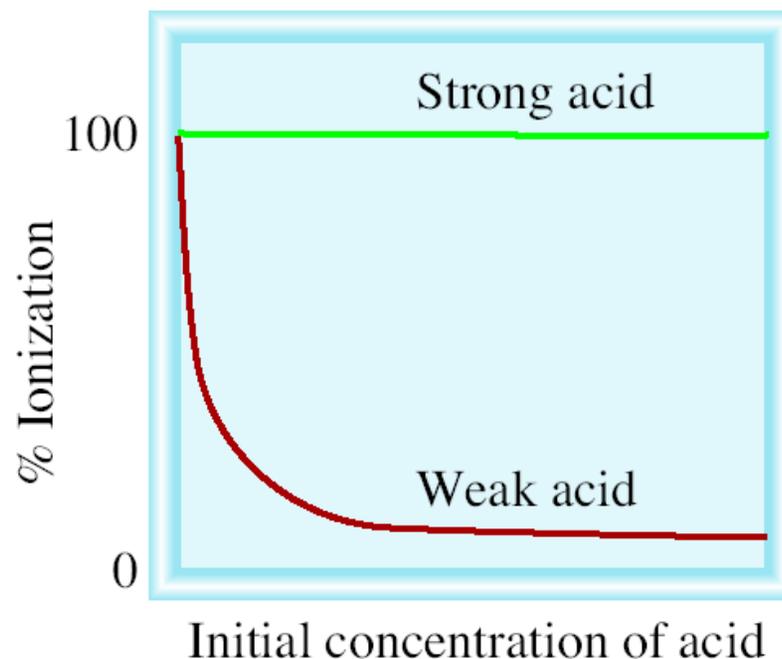
Weak Bases and Base Ionization Constants

$$NH_3 \ (aq) + H_2O \ (l) \rightleftharpoons NH_4^+ \ (aq) + OH^- \ (aq)$$
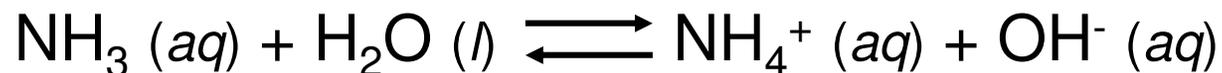
$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

K_b is the **base ionization constant**

$$K_b \uparrow \qquad \text{weak base} \uparrow$$
$$\text{strength}$$

Solve weak base problems like weak acids **except** solve for [OH-] instead of [H$^+$].

TABLE 15.4   **Ionization Constants of Some Weak Bases and Their Conjugate Acids at 25°C**

Name of Base	Formula	Structure	K_b*	Conjugate Acid	K_a
Ethylamine	$C_2H_5NH_2$	$CH_3-CH_2-\overset{..}{\underset{H}{N}}-H$	5.6×10^{-4}	$C_2H_5\overset{+}{N}H_3$	1.8×10^{-11}
Methylamine	CH_3NH_2	$CH_3-\overset{..}{\underset{H}{N}}-H$	4.4×10^{-4}	$CH_3\overset{+}{N}H_3$	2.3×10^{-11}
Ammonia	NH_3	$H-\overset{..}{\underset{H}{N}}-H$	1.8×10^{-5}	NH_4^+	5.6×10^{-10}
Pyridine	C_5H_5N		1.7×10^{-9}	$C_5H_5\overset{+}{N}H$	5.9×10^{-6}
Aniline	$C_6H_5NH_2$		3.8×10^{-10}	$C_6H_5\overset{+}{N}H_3$	2.6×10^{-5}
Caffeine	$C_8H_{10}N_4O_2$		5.3×10^{-14}	$C_8H_{11}\overset{+}{N}_4O_2$	0.19
Urea	$(NH_2)_2CO$	$H-\overset{..}{\underset{H}{N}}-\overset{O}{\overset{\|}{C}}-\overset{..}{\underset{H}{N}}-H$	1.5×10^{-14}	$H_2NCO\overset{+}{N}H_3$	0.67
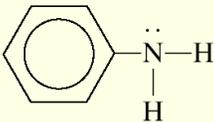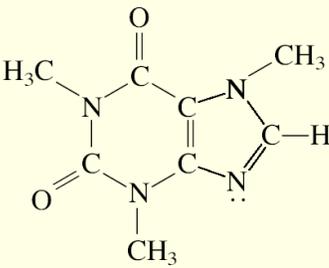
*The nitrogen atom with the lone pair accounts for each compound's basicity. In the case of urea, K_b can be associated with either nitrogen atom.

Ionization Constants of Conjugate Acid-Base Pairs

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

$$A^-\ (aq) + H_2O\ (l) \rightleftharpoons OH^-\ (aq) + HA\ (aq) \qquad K_b$$

$$H_2O\ (l) \rightleftharpoons H^+\ (aq) + OH^-\ (aq) \qquad K_w$$
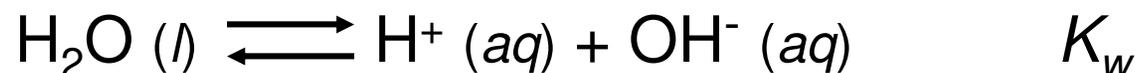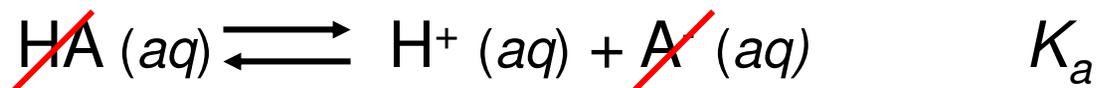
$$\boxed{K_a K_b = K_w}$$

Weak Acid and Its Conjugate Base

$$K_a = \frac{K_w}{K_b} \qquad\qquad K_b = \frac{K_w}{K_a}$$

Diprotic and Triprotic Acids

- May yield more than one hydrogen ion per molecule.
- Ionize in a stepwise manner; that is, they lose one proton at a time.
- An ionization constant expression can be written for each ionization stage.
- Consequently, two or more equilibrium constant expressions must often be used to calculate the concentrations of species in the acid solution.
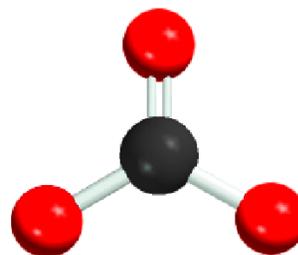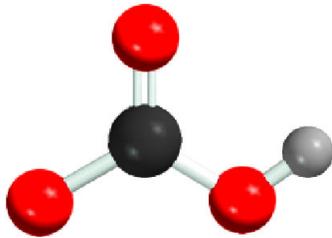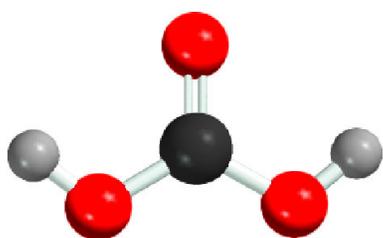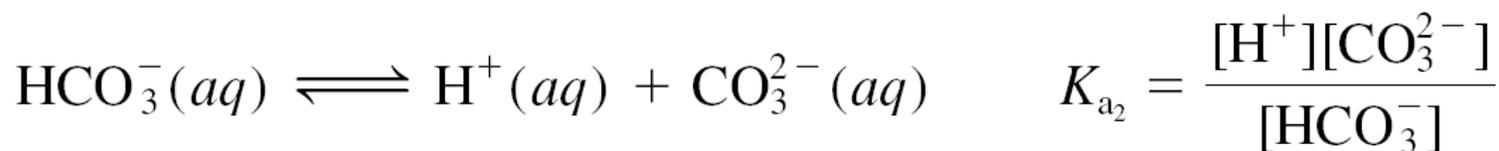
$$H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq) \qquad K_{a_1} = \frac{[H^+][HCO_3^-]}{[H_2CO_3]}$$

$$HCO_3^-(aq) \rightleftharpoons H^+(aq) + CO_3^{2-}(aq) \qquad K_{a_2} = \frac{[H^+][CO_3^{2-}]}{[HCO_3^-]}$$

| Name of Acid	Formula	Structure	K_a	Conjugate Base	K_b
Sulfuric acid	H_2SO_4	$H{-}O{-}\overset{\overset{O}{\|}}{\underset{\underset{O}{\|}}{S}}{-}O{-}H$	very large	HSO_4^-	very small
Hydrogen sulfate ion	HSO_4^-	$H{-}O{-}\overset{\overset{O}{\|}}{\underset{\underset{O}{\|}}{S}}{-}O^-$	1.3×10^{-2}	SO_4^{2-}	7.7×10^{-13}
Oxalic acid	$H_2C_2O_4$	$H{-}O{-}\overset{\overset{O}{\|}}{C}{-}\overset{\overset{O}{\|}}{C}{-}O{-}H$	6.5×10^{-2}	$HC_2O_4^-$	1.5×10^{-13}
Hydrogen oxalate ion	$HC_2O_4^-$	$H{-}O{-}\overset{\overset{O}{\|}}{C}{-}\overset{\overset{O}{\|}}{C}{-}O^-$	6.1×10^{-5}	$C_2O_4^{2-}$	1.6×10^{-10}
Sulfurous acid*	H_2SO_3	$H{-}O{-}\overset{\overset{O}{\|}}{S}{-}O{-}H$	1.3×10^{-2}	HSO_3^-	7.7×10^{-13}
Hydrogen sulfite ion	HSO_3^-	$H{-}O{-}\overset{\overset{O}{\|}}{S}{-}O^-$	6.3×10^{-8}	SO_3^{2-}	1.6×10^{-7}
Carbonic acid	H_2CO_3	$H{-}O{-}\overset{\overset{O}{\|}}{C}{-}O{-}H$	4.2×10^{-7}	HCO_3^-	2.4×10^{-8}
Hydrogen carbonate ion	HCO_3^-	$H{-}O{-}\overset{\overset{O}{\|}}{C}{-}O^-$	4.8×10^{-11}	CO_3^{2-}	2.1×10^{-4}
Hydrosulfuric acid	H_2S	$H{-}S{-}H$	9.5×10^{-8}	HS^-	1.1×10^{-7}
Hydrogen sulfide ion†	HS^-	$H{-}S^-$	1×10^{-19}	S^{2-}	1×10^{5}
Phosphoric acid	H_3PO_4	$H{-}O{-}\overset{\overset{O}{\|}}{\underset{\underset{\underset{H}{\|}}{\underset{O}{\|}}}{P}}{-}O{-}H$	7.5×10^{-3}	$H_2PO_4^-$	1.3×10^{-12}
Dihydrogen phosphate ion	$H_2PO_4^-$	$H{-}O{-}\overset{\overset{O}{\|}}{\underset{\underset{\underset{H}{\|}}{\underset{O}{\|}}}{P}}{-}O^-$	6.2×10^{-8}	HPO_4^{2-}	1.6×10^{-7}
Hydrogen phosphate ion	HPO_4^{2-}	$H{-}O{-}\overset{\overset{O}{\|}}{\underset{\underset{\underset{O^-}{\|}}{}}{P}}{-}O^-$	4.8×10^{-13}	PO_4^{3-}	2.1×10^{-2}

*H_2SO_3 has never been isolated and exists in only minute concentration in aqueous solution of SO_2. The K_a value here refers to the process $SO_2(g) + H_2O(l) \rightleftharpoons H^+(aq) + HSO_3^-(aq)$.
†The ionization constant of HS^- is very low and difficult to measure. The value listed here is only an estimate.

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Molecular Structure and Acid Strength

$$H\text{—}X \longrightarrow H^+ + X^-$$

The
stronger
the bond

The
weaker
the acid

HF << HCl < HBr < HI

acidity
increases

TABLE 15.6	Bond Enthalpies for Hydrogen Halides and Acid Strengths for Hydrohalic Acids	
Bond	**Bond Enthalpy (kJ/mol)**	**Acid Strength**
H—F	568.2	weak
H—Cl	431.9	strong
H—Br	366.1	strong
H—I	298.3	strong

Molecular Structure and Oxoacid Strength

$$\begin{array}{c} \end{array} Z - \overset{\delta^-}{O} - \overset{\delta^+}{H} \longrightarrow \begin{array}{c} \end{array} Z - O^- + H^+$$
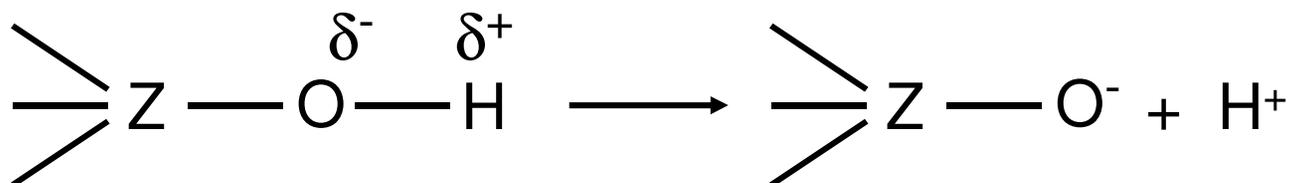
The O-H bond will be more polar and easier to break if:

- Z is very electronegative or

- Z is in a high oxidation state
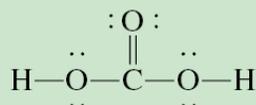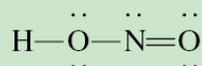
Carbonic acid   Nitrous acid   Nitric acid
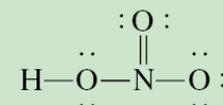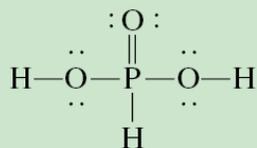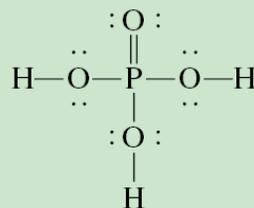
Phosphorous acid   Phosphoric acid   Sulfuric acid   31

Molecular Structure and Oxoacid Strength

1. Oxoacids having different central atoms (Z) that **are from the same group** and that have the **same oxidation number**.

Acid strength increases with increasing electronegativity of Z

Cl is more electronegative than Br

$HClO_3 > HBrO_3$

acidity increases
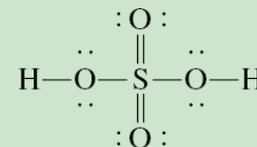
Molecular Structure and Acid Strength

2. Oxoacids having the same central atom (Z) but different numbers of attached groups.

Acid strength increases as the oxidation number of Z increases.
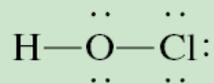
Hypochlorous acid (+1)
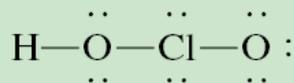
Chlorous acid (+3)
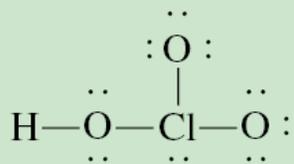
Chloric acid (+5)

Perchloric acid (+7)

$HClO_4 > HClO_3 > HClO_2 > HClO$

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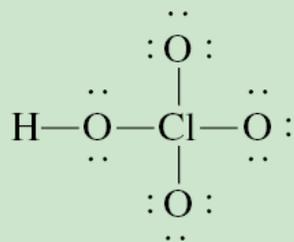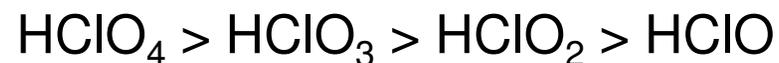
Acid-Base Properties of Salts

Neutral Solutions:

Salts containing an alkali metal or alkaline earth metal ion (except Be^{2+}) **and** the conjugate base of a **strong** acid (*e.g.* Cl^-, Br^-, and NO_3^-).

$$NaCl \ (s) \xrightarrow{\text{H}_2\text{O}} Na^+ \ (aq) + Cl^- \ (aq)$$

Basic Solutions:

Salts derived from a strong base **and** a **weak** acid.

$$NaCH_3COOH \ (s) \xrightarrow{\text{H}_2\text{O}} Na^+ \ (aq) + CH_3COO^- \ (aq)$$

$$CH_3COO^- \ (aq) + H_2O \ (l) \rightleftharpoons CH_3COOH \ (aq) + OH^- \ (aq)$$
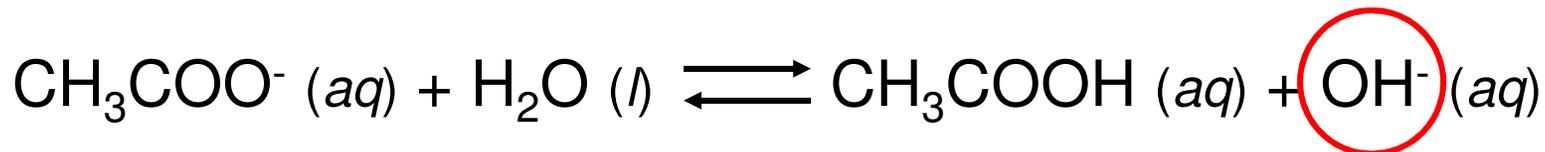
Acid-Base Properties of Salts

<u>Acid Solutions:</u>

Salts derived from a strong acid and a weak base.

$$NH_4Cl \ (s) \xrightarrow{\ H_2O\ } NH_4^+ \ (aq) + Cl^- \ (aq)$$

$$NH_4^+ \ (aq) \rightleftharpoons NH_3 \ (aq) + H^+ \ (aq)$$
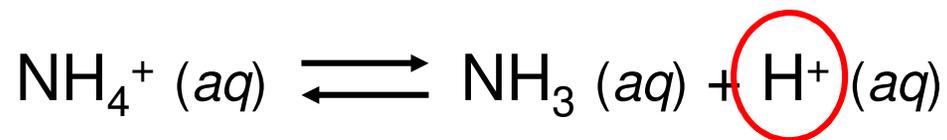
Salts with small, highly charged metal cations (*e.g.* Al^{3+}, Cr^{3+}, and Be^{2+}) and the conjugate base of a strong acid.

$$Al(H_2O)_6^{3+}(aq) \rightleftharpoons Al(OH)(H_2O)_5^{2+}(aq) + H^+ \ (aq)$$

Acid Hydrolysis of Al^{3+}

$$Al(H_2O)_6^{3+} \quad + \quad H_2O \quad \longrightarrow \quad Al(OH)(H_2O)_5^{2+} \quad + \quad H_3O^+$$
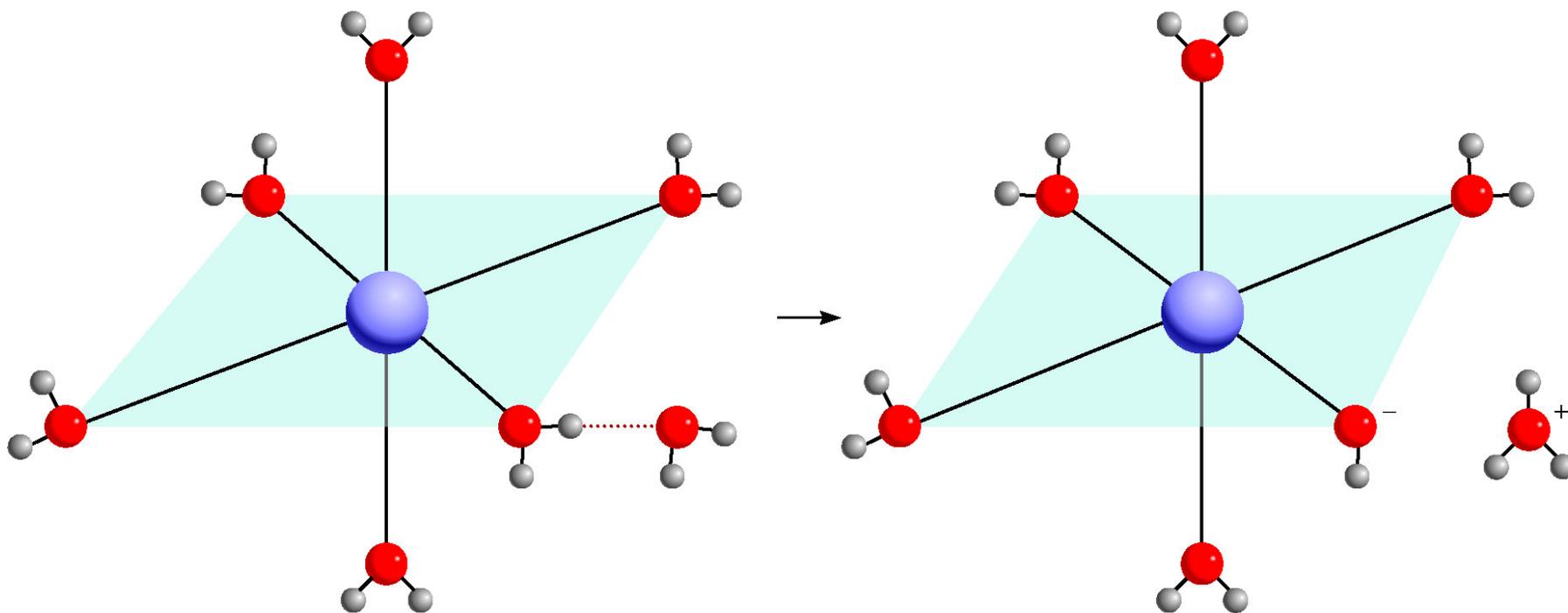
Acid-Base Properties of Salts

Solutions in which both the cation and the anion hydrolyze:

- K_b for the anion $>$ K_a for the cation, solution will be basic

- K_b for the anion $<$ K_a for the cation, solution will be acidic

- K_b for the anion $\approx$ K_a for the cation, solution will be neutral

TABLE 15.7 Acid-Base Properties of Salts

Type of Salt	Examples	Ions That Undergo Hydrolysis	pH of Solution
Cation from strong base; anion from strong acid	$NaCl$, KI, KNO_3, $RbBr$, $BaCl_2$	None	≈ 7
Cation from strong base; anion from weak acid	CH_3COONa, KNO_2	Anion	> 7
Cation from weak base; anion from strong acid	NH_4Cl, NH_4NO_3	Cation	< 7
Cation from weak base; anion from weak acid	NH_4NO_2, CH_3COONH_4, NH_4CN	Anion and cation	< 7 if $K_b < K_a$ ≈ 7 if $K_b \approx K_a$ > 7 if $K_b > K_a$
Small, highly charged cation; anion from strong acid	$AlCl_3$, $Fe(NO_3)_3$	Hydrated cation	< 7

Oxides of the Representative Elements
In Their Highest Oxidation States

Legend:
- Basic oxide
- Acidic oxide
- Amphoteric oxide

1 1A	2 2A											13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
Li_2O	BeO											B_2O_3	CO_2	N_2O_5		OF_2	
Na_2O	MgO	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	Al_2O_3	SiO_2	P_4O_{10}	SO_3	Cl_2O_7	
K_2O	CaO											Ga_2O_3	GeO_2	As_2O_5	SeO_3	Br_2O_7	
Rb_2O	SrO											In_2O_3	SnO_2	Sb_2O_5	TeO_3	I_2O_7	
Cs_2O	BaO											Tl_2O_3	PbO_2	Bi_2O_5	PoO_3	At_2O_7	

$$Na_2O\ (s) + H_2O\ (l) \longrightarrow 2NaOH\ (aq)$$

$$CO_2\ (g) + H_2O\ (l) \rightleftharpoons H_2CO_3\ (aq)$$

$$N_2O_5\ (g) + H_2O\ (l) \rightleftharpoons 2HNO_3\ (aq)$$

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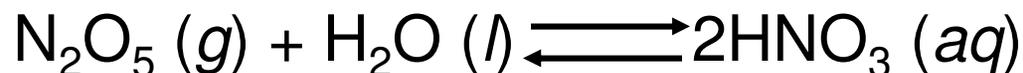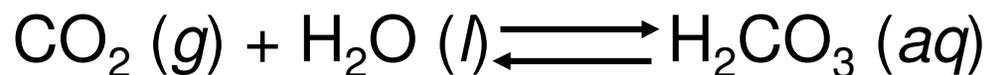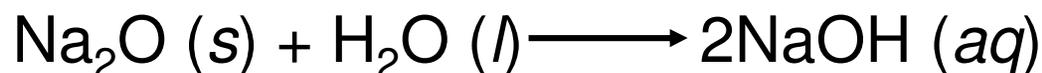
Definition of An Acid

*Arrhenius acid* is a substance that produces H^+ (H_3O^+) in water

A *Brønsted acid* is a proton donor

A *Lewis acid* is a substance that can accept a pair of electrons

A *Lewis base* is a substance that can donate a pair of electrons

$$H^+ + \overset{..}{\underset{..}{O}}H^- \longrightarrow H-\overset{..}{\underset{..}{O}}-H$$

acid   base

$$H^+ + \ :N \begin{matrix} H \\ | \\ -H \\ | \\ H \end{matrix} \longrightarrow H-\underset{\underset{H}{|}}{\overset{\overset{H}{|}}{N}}-H \ ^+$$

acid   base

Lewis Acids and Bases

$$F{-}\overset{\displaystyle F}{\underset{\displaystyle F}{B}} \quad + \quad :\overset{\displaystyle H}{\underset{\displaystyle H}{N}}{-}H \quad \longrightarrow \quad F{-}\overset{\displaystyle F}{\underset{\displaystyle F}{B}}{-}\overset{\displaystyle H}{\underset{\displaystyle H}{N}}{-}H$$
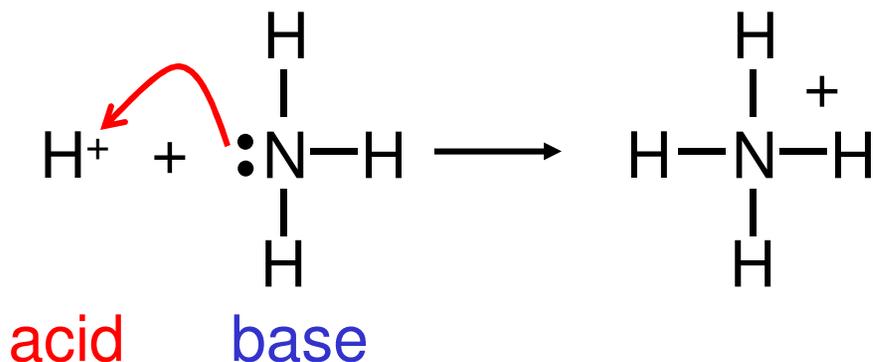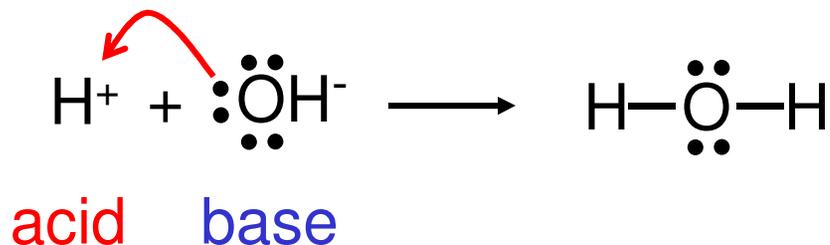
acid   base
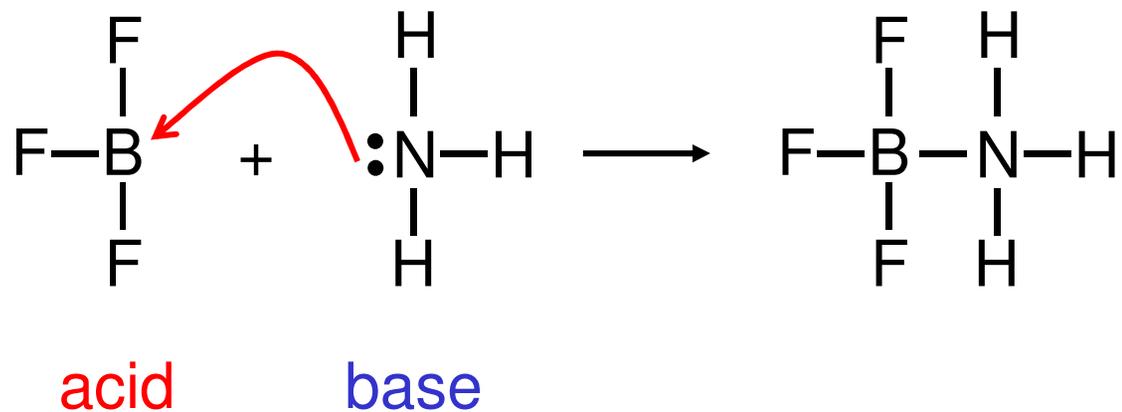
No protons donated or accepted!

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Chemistry In Action: Antacids and the Stomach pH Balance

Some Common Commercial Antacid Preparations

Commercial Name	Active Ingredients
Alka-2	Calcium carbonate
Alka-Seltzer	Aspirin, sodium bicarbonate, citric acid
Bufferin	Aspirin, magnesium carbonate, aluminum glycinate
Buffered aspirin	Aspirin, magnesium carbonate, aluminum hydroxide-glycine
Milk of magnesia	Magnesium hydroxide
Rolaids	Dihydroxy aluminum sodium carbonate
Tums	Calcium carbonate

$NaHCO_3$ (aq) + HCl (aq) $\longrightarrow$

$NaCl$ (aq) + H_2O (l) + CO_2 (g)

$Mg(OH)_2$ (s) + 2HCl (aq) $\longrightarrow$

$MgCl_2$ (aq) + $2H_2O$ (l)

Food

Mucous membran

Blood plasma

HCl(aq)

To intestines

Blood plasma

Cl^- H^+ (active transport)

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