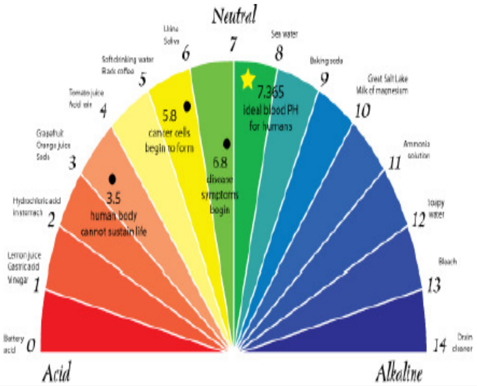
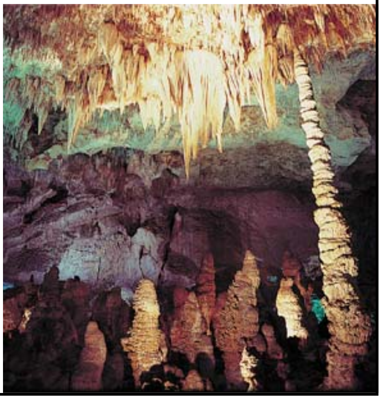
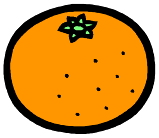
Acids and Bases

Chapter 15





Acids



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

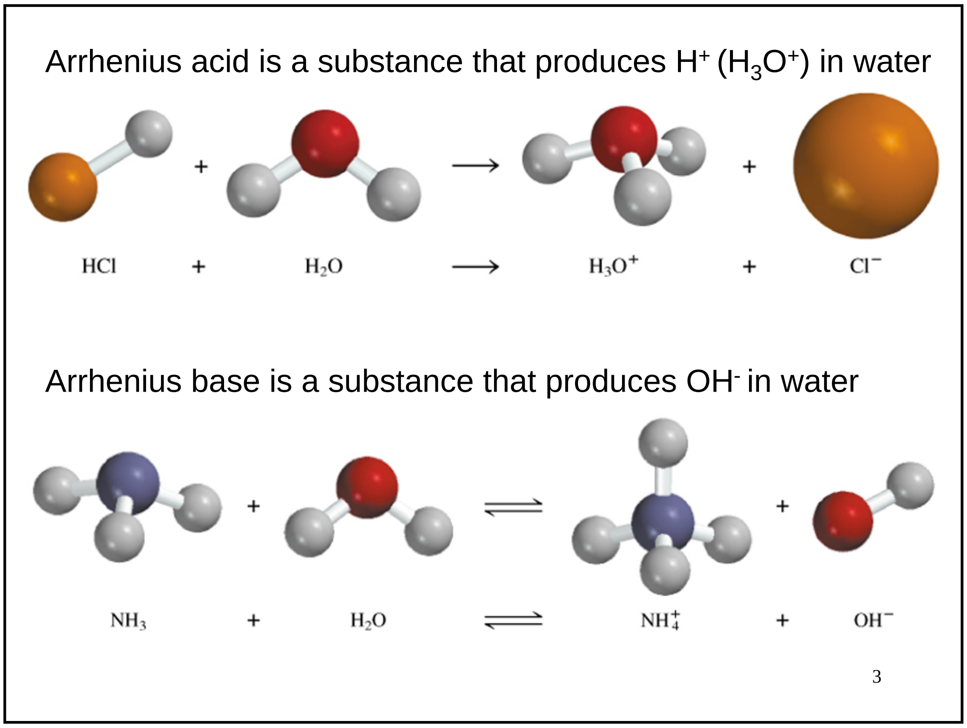
React with certain metals to produce hydrogen gas.

React with carbonates and bicarbonates to produce carbon dioxide gas

Bases

Have a bitter taste.

Feel slippery. Many soaps contain bases.

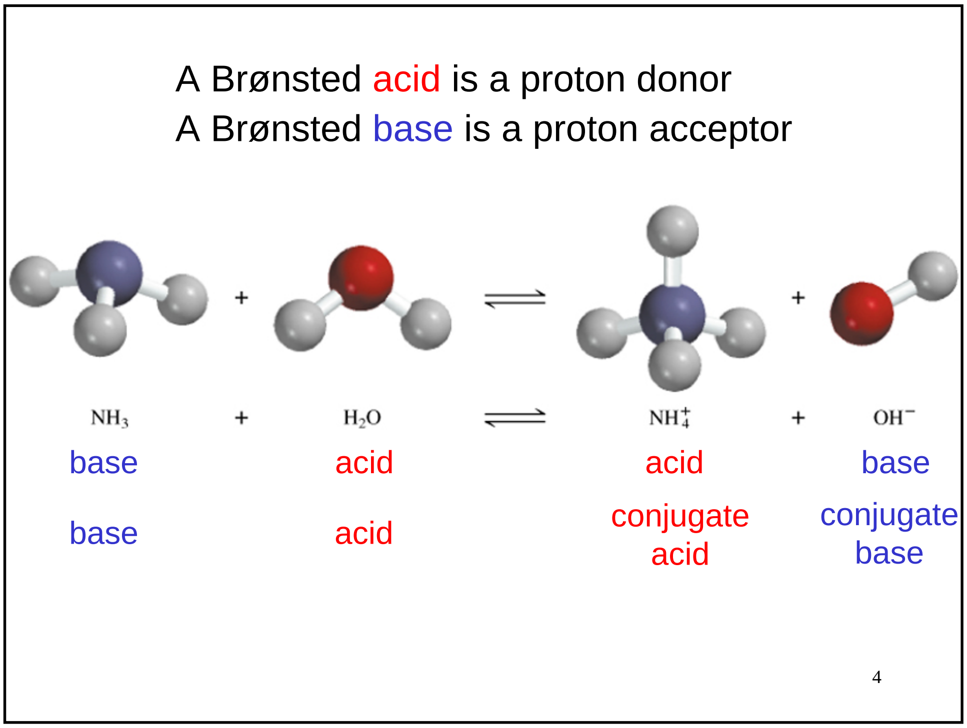


Arrhenius acid is a substance that produces H^+ (H_3O^+) in water

$HCl + H_2O \longrightarrow H_3O^+ + Cl^-$

Arrhenius base is a substance that produces OH^- in water

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$



A Brønsted acid is a proton donor

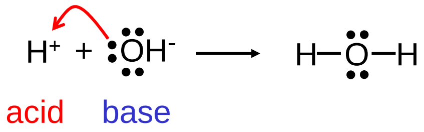
A Brønsted base is a proton acceptor

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

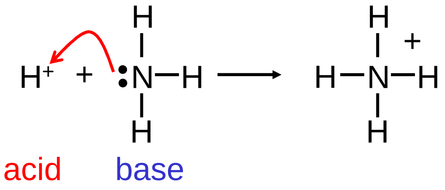
NH_3	H_2O	NH_4^+	OH^-
base	acid	acid	base
base	acid	conjugate acid	conjugate base

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



$H^+ + OH^- \longrightarrow H{-}O{-}H$

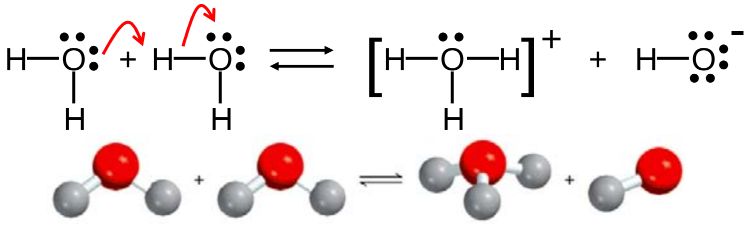


$H^+ + NH_3 \longrightarrow NH_4^+$

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

The Ion Product of Water

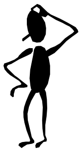
$$H_2O\ (l) \rightleftharpoons H^+\ (aq) + OH^-\ (aq) \qquad K_c = \frac{[H^+][OH^-]}{[H_2O]} \qquad [H_2O] = \text{constant}$$

$$K_c[H_2O] = K_w = [H^+][OH^-]$$

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

At 25^0C
$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$

	Solution Is
$[H^+] = [OH^-]$	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

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

Solution Is		**At 25^0C**	
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

$\text{pH} \uparrow \qquad [H^+] \downarrow$

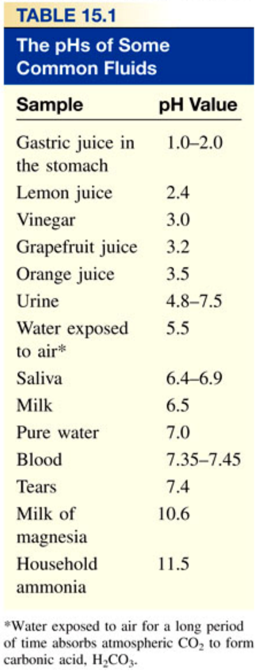


TABLE 15.1

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

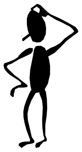
*Water exposed to air for a long period of time absorbs atmospheric CO_2 to form carbonic acid, H_2CO_3.

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

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

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

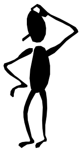
$$\text{pH} + \text{pOH} = 14.00$$



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

$$NaCl\ (s) \xrightarrow{H_2O} Na^+\ (aq) + Cl^-\ (aq)$$

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

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

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

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

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.

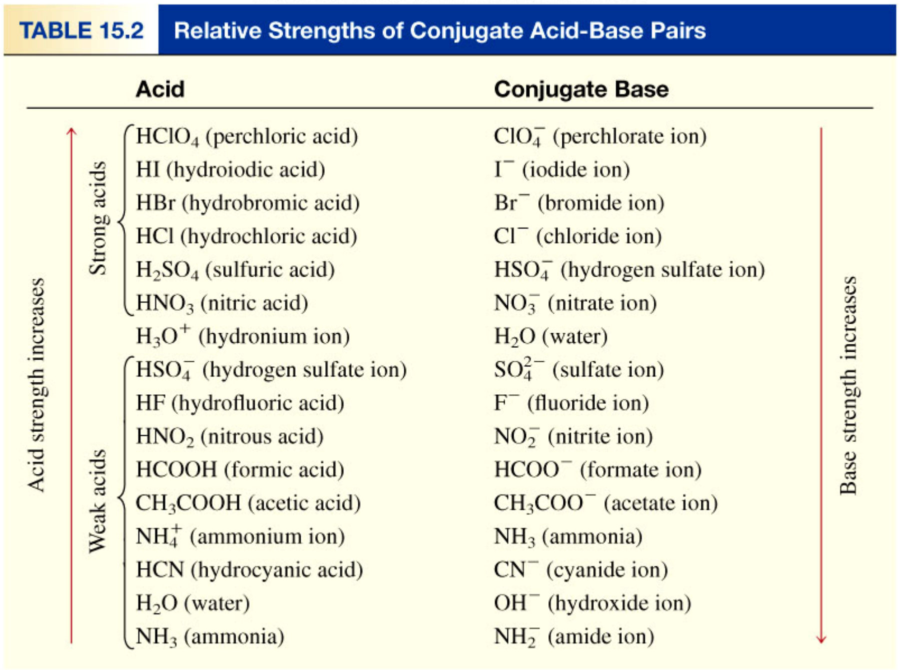
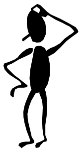


TABLE 15.2 Relative Strengths of Conjugate Acid-Base Pairs

	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



What is the pH of a 2 x 10^{-3} *M* HNO_3 solution?

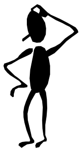
HNO_3 is a strong acid – 100% dissociation.

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 x 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

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

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

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

K_a is the ***acid ionization constant***

K_a ↑ weak acid strength ↑

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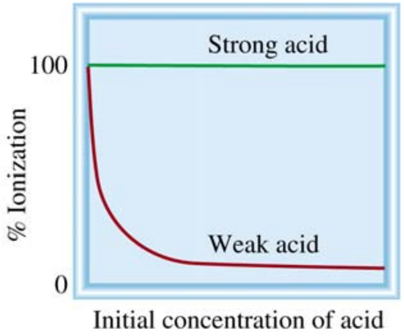
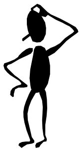
percent ionization = $\frac{\text{Ionized acid concentration at equilibrium}}{\text{Initial concentration of acid}}$ x 100%

For a monoprotic acid HA

Percent ionization = $\frac{[H^+]}{[HA]_0}$ x 100% $[HA]_0$ = initial concentration





What is the pH of a 0.5 *M* HF solution (at 25⁰C)?

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

$$HF\ (aq) \rightleftharpoons H^+\ (aq) + F^-\ (aq)$$

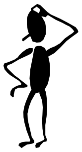
Initial (M)	0.50	0.00	0.00
Change (M)	$-x$	$+x$	$+x$
Equilibrium (M)	$0.50 - x$	x	x

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

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

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

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



When can I use the approximation?

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

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

$$x = 0.019 \qquad \frac{0.019\ M}{0.50\ M} \times 100\% = 3.8\%$$

Less than 5%

Approximation ok.

What is the pH of a 0.05 *M* HF solution (at 25⁰C)?

$$K_a \approx \frac{x^2}{0.05} = 7.1 \times 10^{-4} \qquad x = 0.006\ M$$

$$\frac{0.006\ M}{0.05\ M} \times 100\% = 12\%$$

More than 5%

Approximation **not** ok.

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

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 ↑ weak base strength ↑

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