Acids and Bases

strong electrolytes

"strong acids & bases"

mono protic

"one H^+"

di protic

"two H^+"

tri protic

"3 H^+"

HNO_3

later H_2SO_4 * only the 1st ionization "strong"

$H_2SO_4 \rightarrow H^+ + HSO_4^-$ strong

$HSO_4^- \rightarrow H^+ + SO_4^{2-}$ weak

HCl, HBr, HI

$HClO_4$

Strong bases

all group I cations + OH^-

$Ba(OH)_2$ & $Sr(OH)_2$

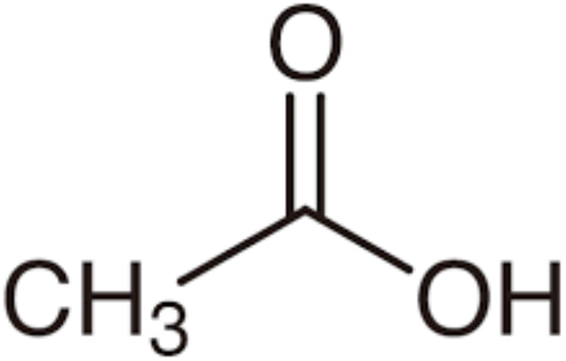
any other base is a weak base

$Mg(OH)_2\ (s)$

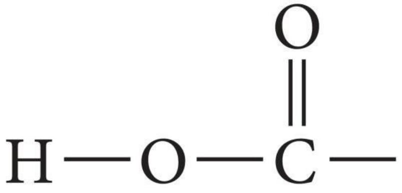
$$Mg(OH)_{2(s)} \rightleftharpoons Mg^{2+}_{(aq)} + 2OH^-_{(aq)}$$

[redacted]

Structure of Acids (Organic Acids)



Acetic acid



$$H-O-\overset{\overset{\displaystyle O}{\|}}{C}-$$

Carboxylic acid group

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

eq. constant for a

acid

weak acid

* only use M for aq. solutions

Indicators

- Indicators are chemicals that change color depending on the solution’s acidity or basicity.

- Demo using butterfly pea flower

Definitions of Acids and Bases

HCl → acid

NaOH → base

- Arrhenius definition (review)
  - Based on H+ and OH-
  - Flawed, does not account for molecular bases such as ammonia (NH_3)
- Brønsted–Lowry definition (review)
  - Based on reactions in which H+ is transferred
  - This theory is used most often in Ch 104
- Lewis definition → later
  - Based on electron transfer
  - Will go over this definition later

Brønsted–Lowry Theory

- The acid is an H⁺ donor.
- The base is an H⁺ acceptor.
- In a Brønsted–Lowry acid–base reaction, the acid molecule donates an H⁺ to the base molecule.

$$H{-}A + :B \Leftrightarrow :A^- + \boxed{H{-}B^+}$$

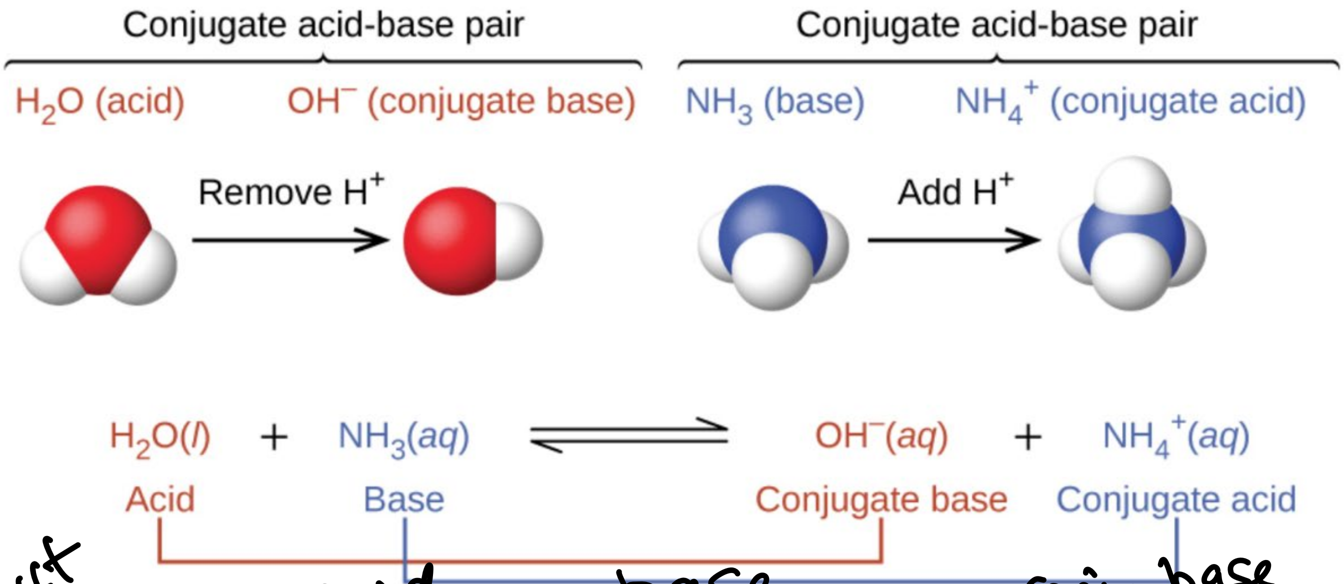
conjugate base

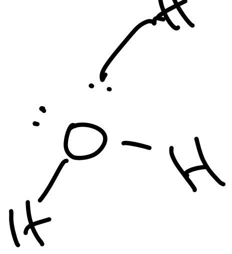
$$\underset{\text{acid}}{H{-}Cl} + \underset{\text{base}}{H{-}\underset{\underset{H}{|}}{\ddot{N}}{-}H} \rightarrow Cl^- + \underset{\text{conjugate acid}}{NH_4^+}$$

Conjugate Acid–Base Pairs

- In a Brønsted-–Lowry acid–base reaction,
  - the original base becomes an acid in the reverse reaction.
  - the original acid becomes a base in the reverse process.

- Each reactant and the product it becomes is called conjugate pair





$$\overset{\text{acid}}{\underline{H}Cl_{(aq)}} + \overset{\text{base}}{H_2O_{(l)}} \longrightarrow \overset{\text{conj. base}}{Cl^-_{(aq)}} + H_3O^+_{(aq)}$$

Bronsted Lowry

acid base conj. acid conj. base

- HF (aq) + H_2O ➔ H_3O^+ (aq) + F^- (aq)

acid conj. base conj. acid

- NH_3(aq) + H_2O → OH^- (aq) + NH_4^+

base

- What is the acid and base, conjugate base and conjugate acid?

Amphoteric Substances

- **Amphoteric substances** can act as either an acid or a base because they have both a transferable H and an atom with lone pair electrons.

HCO_3^-

acid

conj. acid

base

$$HCO_{3\,(aq)}^- + H_2O_{(l)} \rightleftharpoons H_2CO_{3\,(aq)} + OH^-_{(aq)}$$

conj. base

acid

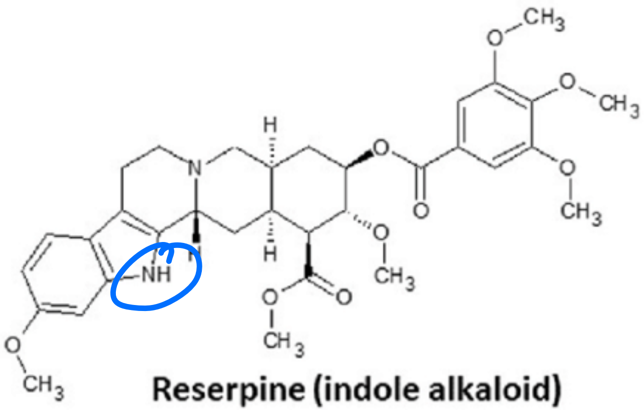
conj. base

$$HCO_3^- + H_2O \rightleftharpoons CO_{3\,(aq)}^{2-} + H_3O^+_{(aq)}$$

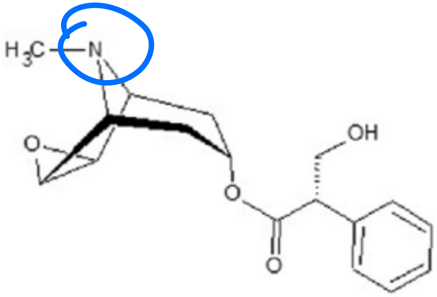
base

conj. acid

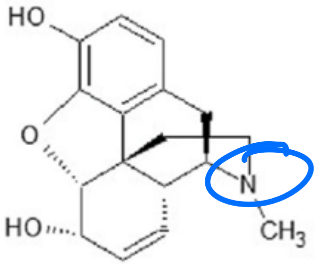
Application to medicine (alkaloids)



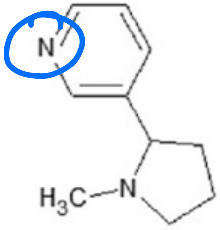
Reserpine (indole alkaloid)



Scopolamine (tropane alkaloid



Morphine (benzylisoquinoline alkaloid)



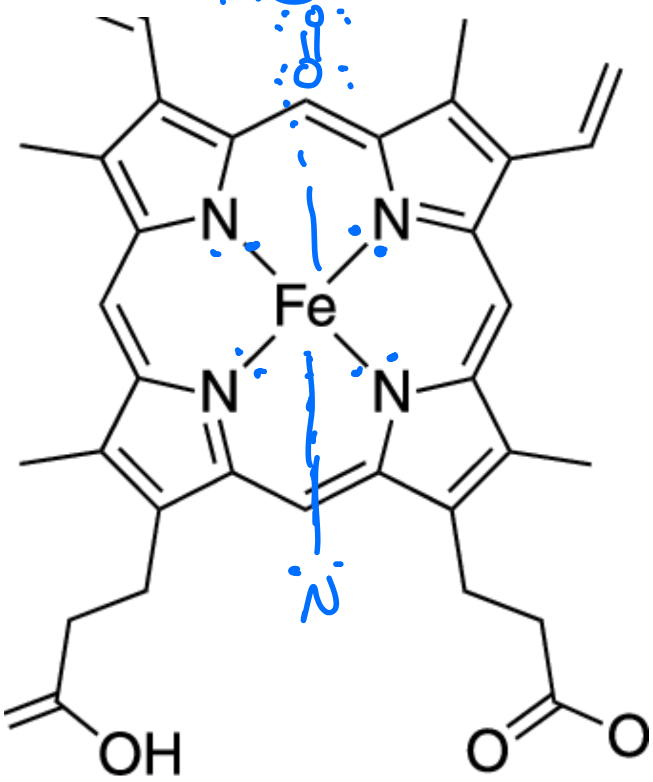
Nicotine (pyridine alkaloid)

Lewis Acid–Base Theory

- Lewis acid–base theory focuses on transferring an electron pair.
- Does NOT require H atoms
- The electron donor is called the Lewis base.
  - Electron rich; therefore nucleophile organic
- The electron acceptor is called the Lewis acid.
  - Electron deficient; therefore electrophile organic

Examples of Lewis Acid–Base Reactions

heme



Heme group

$Ag^+_{(aq)}$ + $_{(aq)}$ 2 :$NH_{3(aq)}$ → $Ag(NH_3)_2^+$

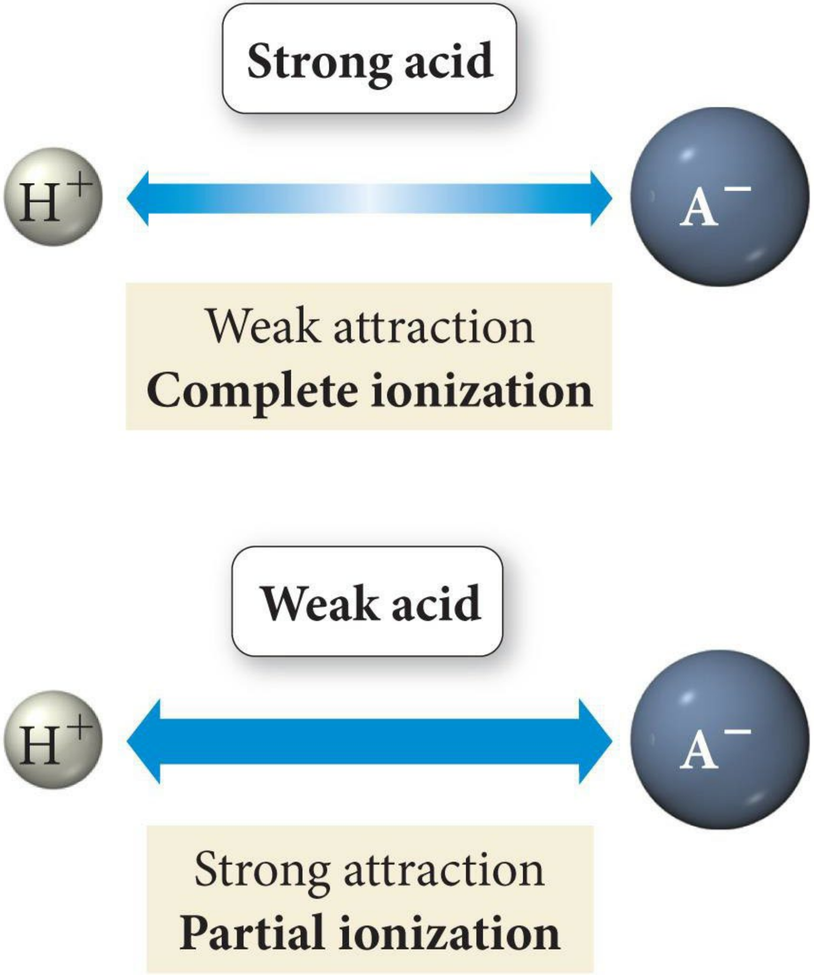
Lewis Acid

Lewis Base

Complex ion!

Fe^{2+} Lewis acid

Ionic Attraction and Acid Strength



Strong acid

Weak attraction
Complete ionization

Weak acid

Strong attraction
Partial ionization

Strong

$HClO_4$

HCl

HI

HBr

HNO_3

H_2SO_4 1st ionization

Relationship between Bond Strength and Acidity

Acid	Bond Energy kJ/mol	Type of Acid
HF	565	weak
HCl	431	strong
HBr	364	strong

Strong versus weak acids

strong

$$HCl + H_2O \xrightarrow[\text{ionized}]{100\%} H_3O^+ + Cl$$

acetic acid

<<100%

$$HC_2H_3O_2 + H_2O \rightleftharpoons H_3O^+ + C_2H_3O_2^-$$

"hydrolysis"

reaction with water

Autoionization of Water

- Water is amphoteric; it can act either as an acid or a base.
  - Therefore, there must be a few ions present.

$$2H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$$

- All aqueous solutions contain **both** H_3O^+ and OH^-.
  - The concentration of H_3O^+ and OH^- are equal in water.
  - $[H_3O^+] = [OH^-] = 10^{-7}$M at 25 °C, Kw = $[H_3O^+][OH\text{-}]$

$\boxed{Kw = 1.00 \times 10^{-14}}$ (equilibrium constant for autoionization of water)

$$K_w = \frac{[H_3O^+][OH^-]}{1}$$

	$2H_2O_{(l)}$	$\rightleftharpoons$	$H_3O^+_{(aq)}$	+	$OH^-_{(aq)}$
I			0		0
C			+x		+x
E			$\boxed{x}$		$\boxed{x}$

$K_w = x \cdot x$ $\qquad K_w = x^2$

@ 25.0°C $\quad 1.00 \times 10^{-14}$

$1.00 \times 10^{-14} = x^2$

$x = \sqrt{1.00 \times 10^{-14}}$

$x = [H_3O^+] = [OH^-] = \boxed{1.0 \times 10^{-7} M}$

Measuring Acidity: pH

- The acidity or basicity of a solution is often expressed as **pH**.

p “−log“

- $pH = -\log[H_3O^+]$
- pH < 7 is acidic; pH > 7 is basic. pH = 7 is neutral.
- $\boxed{[H_3O^+] = 10^{-pH}}$

TABLE 15.6 The pH of Some Common Substances

Substance	pH
Gastric juice (human stomach)	1.0–3.0
Limes	1.8–2.0
Lemons	2.2–2.4
Soft drinks	2.0–4.0
Plums	2.8–3.0
Wines	2.8–3.8
Apples	2.9–3.3
Peaches	3.4–3.6
Cherries	3.2–4.0
Beers	4.0–5.0
Rainwater (unpolluted)	5.6
Human blood	7.3–7.4
Egg whites	7.6–8.0
Milk of magnesia	10.5
Household ammonia	10.5–11.5
4% NaOH solution	14

Example Problems Strong Acids and Bases

given on eq. sheet $K_w = 1.00 \times 10^{-14}$

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

1. Determine the $[H_3O^+]$, $[OH^-]$ and pH of a 0.100 M HCl solution.

$$HCl + H_2O \xrightarrow{100\%} H_3O^+ + Cl^-$$

$[H_3O^+] = 0.100\ M$

$K_w = [H_3O^+][OH^-]$ @25.0°C

$$\frac{1.00\times10^{-14}}{0.100} = \frac{(\cancel{0.100})([OH^-])}{\cancel{0.100}}$$

$[OH^-] = 1.00\times10^{-13} M$

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

$pH = -\log(0.10)$

$\boxed{pH = 1}$

1 → 14

on exam

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

$pOH = -\log[OH^-]$ $pH + pOH = 14$

? pOH: 14 - 1 = ⑬

strong base

2) What is the pH of a 0.200 M solution of $Ba(OH)_2$

$[OH^-] = 2 \times 0.200 M =$

$0.400 M$

$K_w = [H_3O^+][OH^-]$

$1.00 \times 10^{-14} = (0.400)([H_3O^+]$

$[H_3O^+] = 2.5 \times 10^{-14}$

$pH = -\log(2.5 \times 10^{-14})$

$pH = 13.\underline{6}$

Weak acids (Example Problem)

use ICE

weak acids << 100% ionized

① Calculate the pH of 1.0 M HF solution.

The Ka of HF is 6.3×10^{-4}

math check

$\frac{1.0}{6.3\times10^{-4}} = 1,587$ ignore " $-x$ "

$$HF_{(aq)} + H_2O_{(l)} \rightleftharpoons F^-_{(aq)} + H_3O^+_{(aq)}$$

I	1.0 M		0	0
C	~~−x~~		+x	+x
E	1.0 M		x	x

$$K_a = \frac{[F^-][H_3O^+]}{[HF]}$$

$$6.3\times10^{-4} = \frac{(x)(x)}{1.0}$$

$$\sqrt{(6.3\times10^{-4})(1.0)} = \sqrt{x^2}$$

$$x = \sqrt{6.3\times10^{-4}}$$

$$x = 0.025 = [H_3O^+]$$

$$pH = -\log(0.025)$$

$$\boxed{pH = 1.6}$$

$[HCl] = 0.025 M$

version finding pH of a weak acid solution version 2

Finding the Ka of a weak acid given the pH of the solution.

If the pH of a 0.100M solution of a generic weak acid "HA" is 5.0, what is the Ka?

$$HA_{(aq)} + H_2O_{(l)} \rightleftharpoons A^-_{(aq)} + H_3O^+_{(aq)}$$

I	0.100M		0	0
C	$-x$		$+x$	$+x$
E	$0.100 - 1.0\times10^{-5} = 0.09999$		1.0×10^{-5}	1.0×10^{-5} M

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

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

$10^{-5} = 1.0\times10^{-5}M$

$$Ka = \frac{(1.0\times10^{-5})^2}{0.0999}$$

$$\boxed{Ka = 1.00\times10^{-9}}$$

Weak Bases (example problems)

What is the pH of a 0.200 M NH_3 solution?

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

$K_b = 1.8 \times 10^{-5}$ math check $\frac{0.2}{1.8\times10^{-5}} = 11{,}000$

$$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH_{4(aq)}^{+} + OH^{-}_{(aq)}$$

I	0.200 M	0	0
C	$-x$	$+x$	$+x$
E	0.200	$+x$	$+x$

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

$$1.8\times10^{-5} = \frac{x^2}{0.200}$$

$$x = \sqrt{1.8 \times 10^{-5} \cdot 0.200}$$

$$\boxed{1.9 \times 10^{-3}}$$

$$x = [OH^-] = 1.9 \times 10^{-3}$$

$$pOH = 2.7$$

$$pH + pOH = 14$$

$$pH = 14.0 - 2.7$$

$$= \boxed{11.3}$$

For acids $pH < 7$

$pH = 7$ neutral

$pH > 7$ bases

Hydrolysis of Salts

determine the pH of a salt solution using a pH meter.

? in a lecture problem

$$\begin{tikzcd} & NaCl \arrow[dl] \arrow[dr] & \\ \underline{Na^+} & & \underline{Cl^-} \end{tikzcd}$$

strong base strong acid

$NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O$

$\cancel{Na^+_{(aq)}} + OH^-_{(aq)} + H^+_{(aq)} + \cancel{Cl^-_{(aq)}} \rightarrow \cancel{Na^+_{(aq)}} + \cancel{Cl^-_{(aq)}} + \boxed{H_2O_{(l)}}$

neutral
b/c autoioni. of water
pH is 7.0

Is a solution of the salt $NaC_2H_3O_2$ acidic, basic or neutral

$$NaC_2H_3O_2$$

$\swarrow \underline{Na^+}$ $\qquad \searrow \boxed{C_2H_3O_2^-}$

weak

strong

$NaOH_{(aq)} + HC_2H_3O_{2(aq)} \rightarrow$

strong

$NaC_2H_3O_{2(aq)} + H_2O_{(l)}$

$\cancel{Na^+_{(aq)}} + OH^-_{(aq)} + HC_2H_3O_{2(aq)} \rightarrow$

$\cancel{Na^+_{(aq)}} + C_2H_3O_2^-{}_{(aq)} + H_2O_{(l)}$

netionic $OH^-_{(aq)} + HC_2H_3O_{2(aq)} \rightarrow$

$C_2H_3O_2^-{}_{(aq)} + H_2O_{(l)}$

$\uparrow [OH^-]$ basic

$C_2H_3O_2^- + H_2O \rightleftharpoons OH^- + HC_2H_3O_2$

Hydrolysis of Salts Example Problem

Calculate the pH of a 0.100 M sodium acetate Solution

$NaC_2H_3O_2$

$\cancel{Na^+}$ $C_2H_3O_2^-$

$C_2H_3O_{2(aq)}^- + H_2O_{(l)} \rightleftharpoons OH^-_{(aq)} + HC_2H_3O_2$

K_b for acetate ion is 5.6×10^{-10}

$$C_2H_3O_2^- + H_2O_{(l)} \rightleftharpoons OH^-_{(aq)} + HC_2H_3O_{2\,(aq)}$$

I	0.100 M	0	0
C	~~$-x$~~	$+x$	x
E	0.100	x	x

$$K_b = \frac{[OH^-][HC_2H_3O_2]}{[C_2H_3O_2^-]}$$

$$5.6\times10^{-10} = \frac{(x)(x)}{0.100}$$

$$x^2 = \sqrt{0.100\cdot(5.6\times10^{-10})}$$

$$x = 7.5\times10^{-6}$$

$$x = [OH^-] = 7.5\times10^{-6}\,M$$

$$pOH = -\log(7.5\times10^{-6})$$

$$pOH = 5.1$$

$$pH = 14.0 - 5.1$$

Diprotic Acids (more than one ionization)

$\boxed{pH = 8.9}$

Example: What is the pH of a 0.0100 M solution of sulfuric acid?