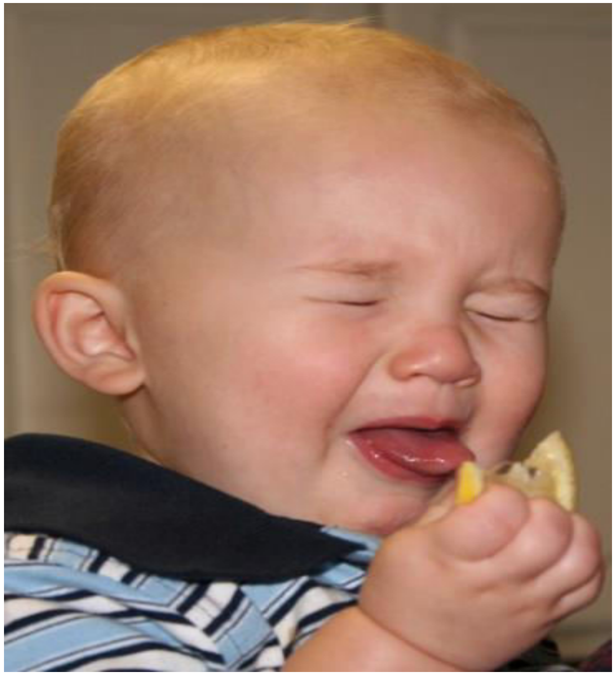
Acids and Bases



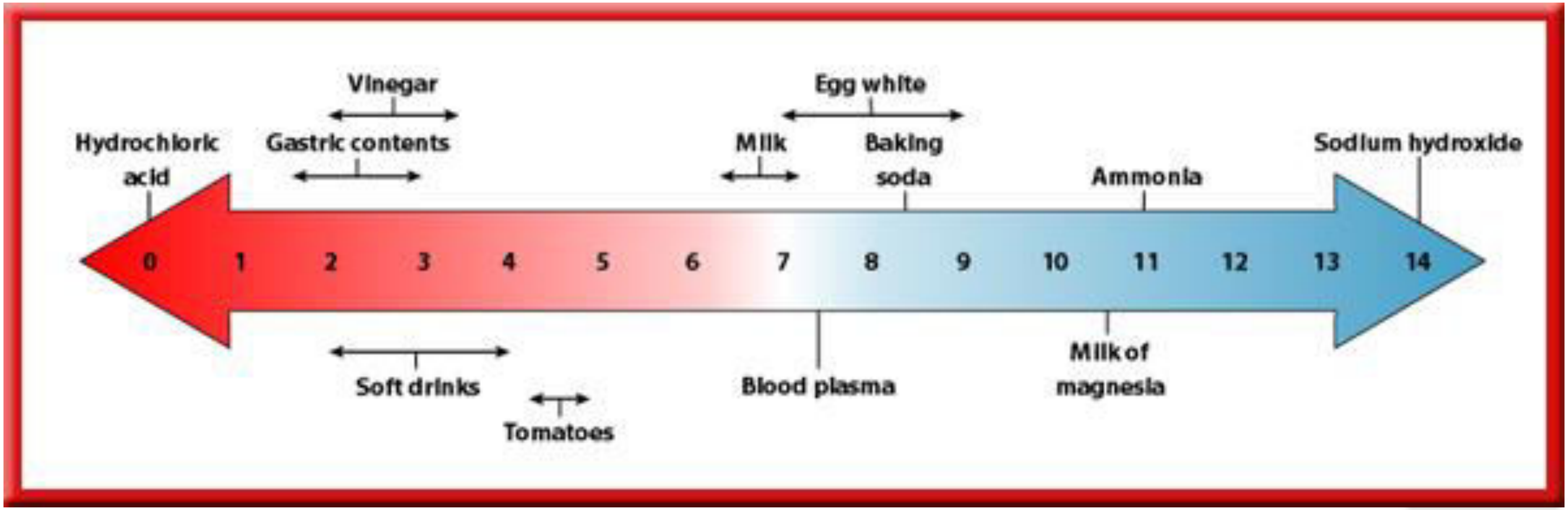
Sour taste

Corrosive

Ant venom – Formic acid

pH Scale

- A change of 1 pH unit represents a **tenfold** change in the acidity of the solution.
- For example, if one solution has a pH of 1 and a second solution has a pH of 2, the first solution is not twice as acidic as the second—it is **ten times** more acidic.



Uses of Acids

- **Acetic Acid** = Vinegar
- **Citric Acid** = lemons, limes, & oranges. It is in many sour candies such as lemonhead & sour patch.
- **Ascorbic acid** = Vitamin C which your body needs to function.
- **Sulfuric acid** is used in the production of fertilizers, steel, paints, and plastics.
- Car batteries

Uses of Bases

- Bases give **soaps**, **ammonia**, and many other cleaning products some of their useful properties.

- Your blood is a basic solution.

- **Caffeine** [$C_8H_{10}N_4O_2$]
  - is a weak base
  - taste bitter and smell like tea,
  - a cardiac stimulant, (boost of energy),
  - mild diuretic, addictive,
  - operates using the same mechanisms that amphetamines, cocaine and heroin use to stimulate the brain.

Arrhenius acids and bases

- **Arrhenius acid** (1880s) : Any substance that, when dissolved in water, increases the concentration of hydronium ion (H_3O^+)

$$HCl(aq) \rightleftharpoons H^+(aq) + Cl^-(aq)$$

- **Arrhenius base (1880s)** : Any substance that, when dissolved in water, increases the concentration of hydroxide ion (OH^-)

$$NaOH(aq) \rightleftharpoons Na^+(aq) + OH^-(aq)$$

Problems with Arrhenius theory

- Only compounds with **OH^-** can be classified as a base. What about ammonia, **NH_3** ?

- Can only be applied to reactions that occur in **water**

- Would **incorrectly** classify some compounds as acids, such as **CH_4**

Brønsted–Lowry acid – base (1923)

an Acid : Proton donor

a Base : Proton acceptor

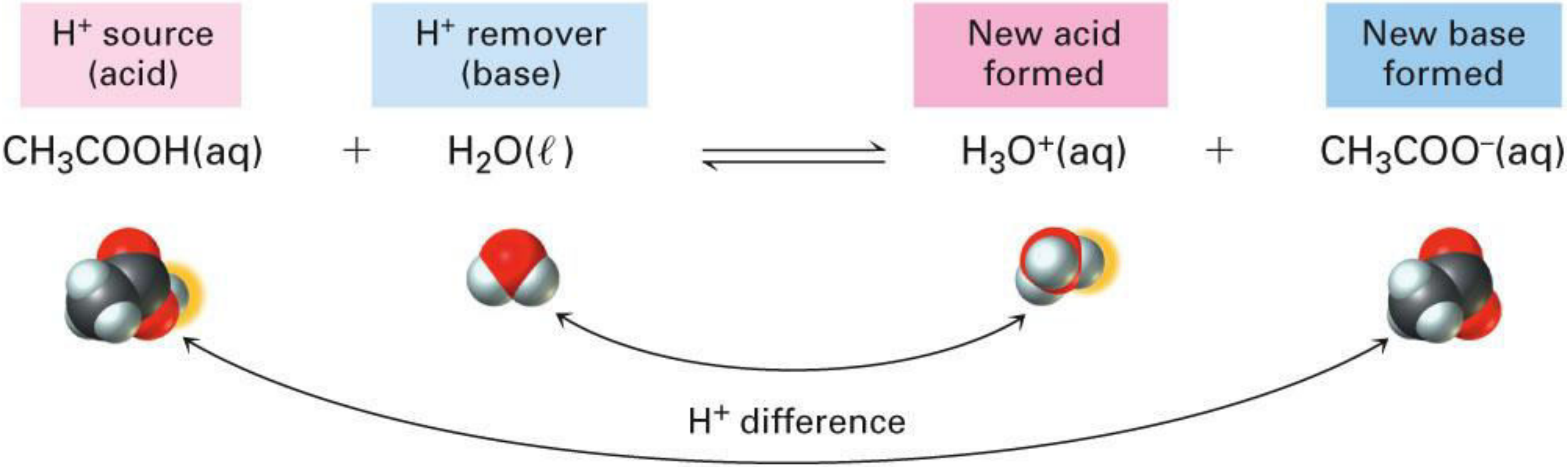
$$HCl(aq) + H_2O \rightleftharpoons H_3O^+(aq) + Cl^-(aq)$$

Which is the acid and which is the base in each of these rxns?

$$HCl + H_2O \rightleftharpoons [Cl^- \cdots H^+ \cdots H_2O] \rightleftharpoons H_3O^+ + Cl^-$$

$$NH_3 + H_2O \rightleftharpoons [NH_3 \cdots H^+ \cdots OH^-] \rightleftharpoons NH_4^+ + OH^-$$

Conjugate Acid-Base Pairs

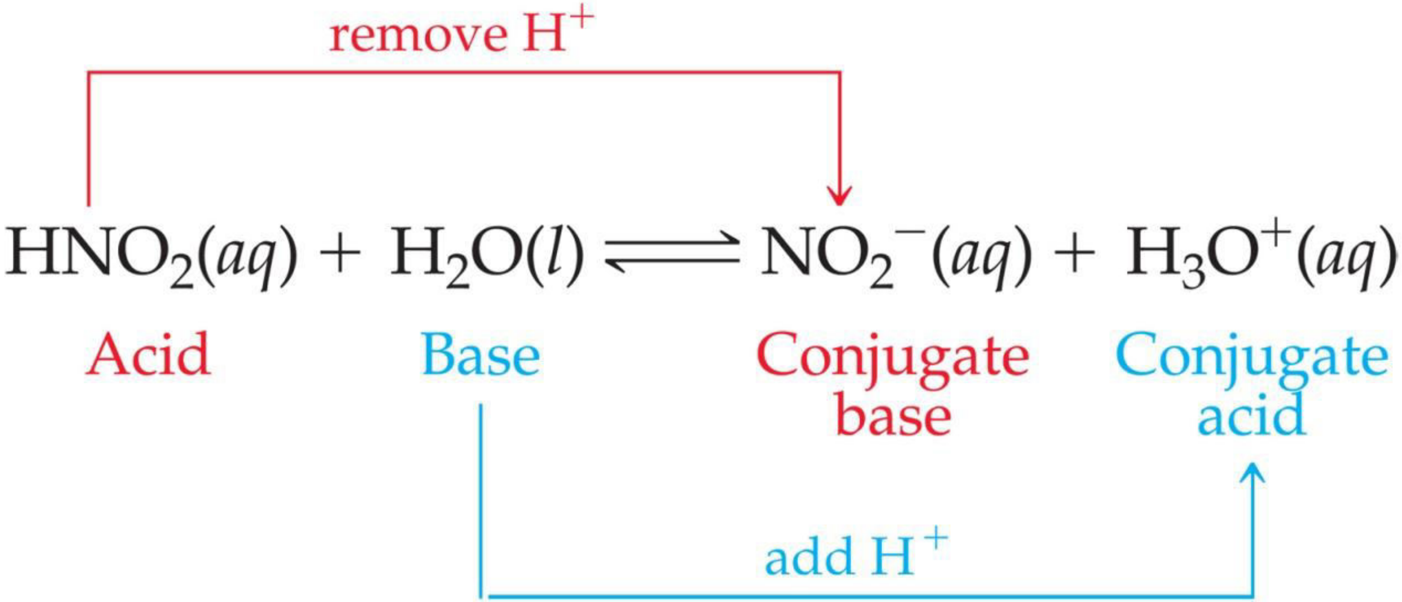


H^+ source (acid)		H^+ remover (base)		New acid formed		New base formed
$CH_3COOH(aq)$	$+$	$H_2O(\ell)$	$\rightleftharpoons$	$H_3O^+(aq)$	$+$	$CH_3COO^-(aq)$

H^+ difference

Conjugate Acids and Bases:

- Reactions between **acids** and **bases** always yield their **conjugate** bases and acids.



remove H^+

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

Acid

Base

Conjugate base

Conjugate acid

add H^+

$$HNO_2 + H_2O \rightleftharpoons [NO_2^- \cdots H^+ \cdots H_2O] \rightleftharpoons NO_2^- + H_30^+$$

A Brønsted–Lowry acid…

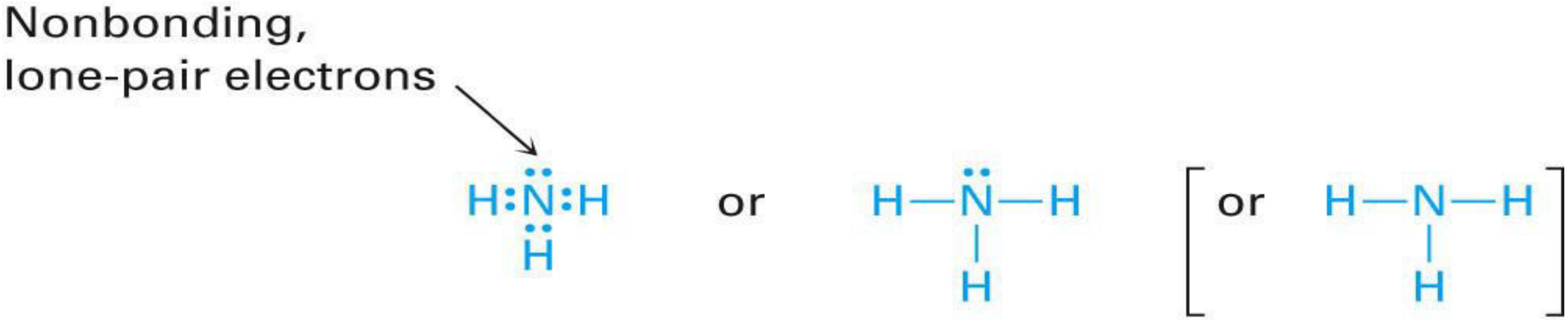
…must have a removable (acidic) proton.

HCl, H_2O, H_2SO_4

A Brønsted–Lowry base…

…must have a pair of nonbonding electrons.

NH_3, H_2O



Ammonia

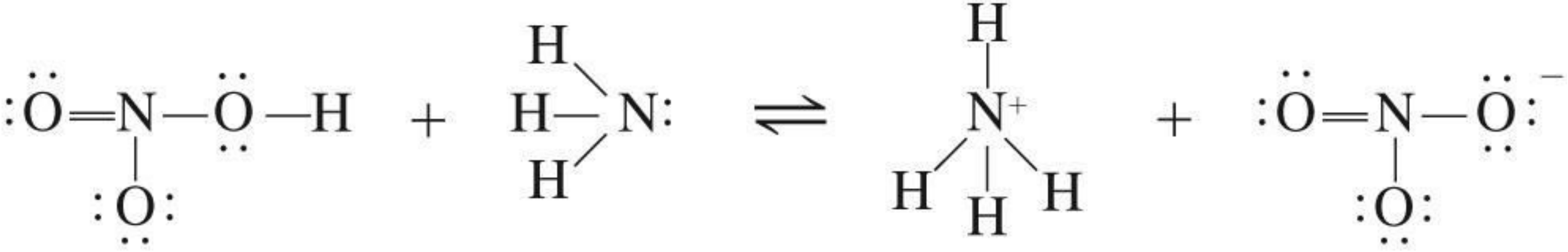
Bases

Brønsted-Lowry Theory of Acids & Bases

Acid: H^+ source

Base: H^+ remover

$$HNO_3(aq) + NH_3(g) \rightleftharpoons NH_4^+(aq) + NO_3^-(aq)$$



Problems with **B-L** theory

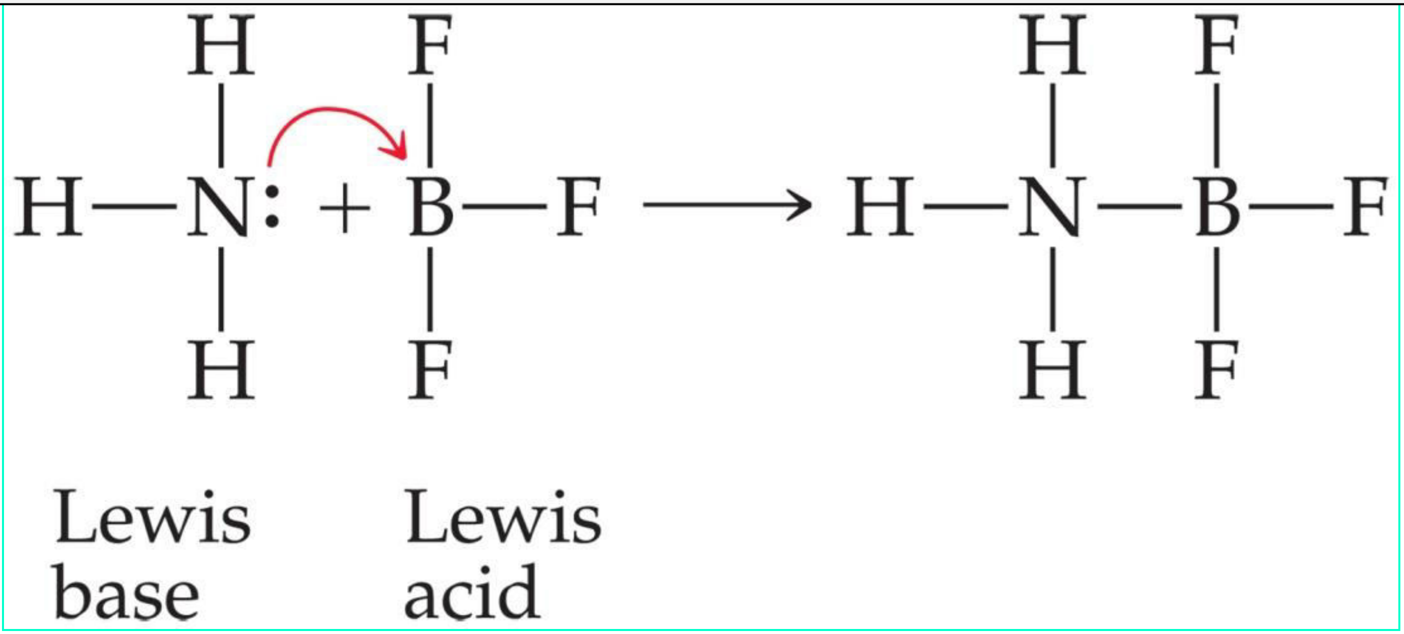
- The theory works very nicely in all **protic** solvent, but fails to explain acid-base behavior in **aprotic** solvents and non-solvent situations.

- A more general concept on acid and base was proposed by **G.N. Lewis** at about the same time Bronsted-Lowry theory was proposed.

Lewis Acids

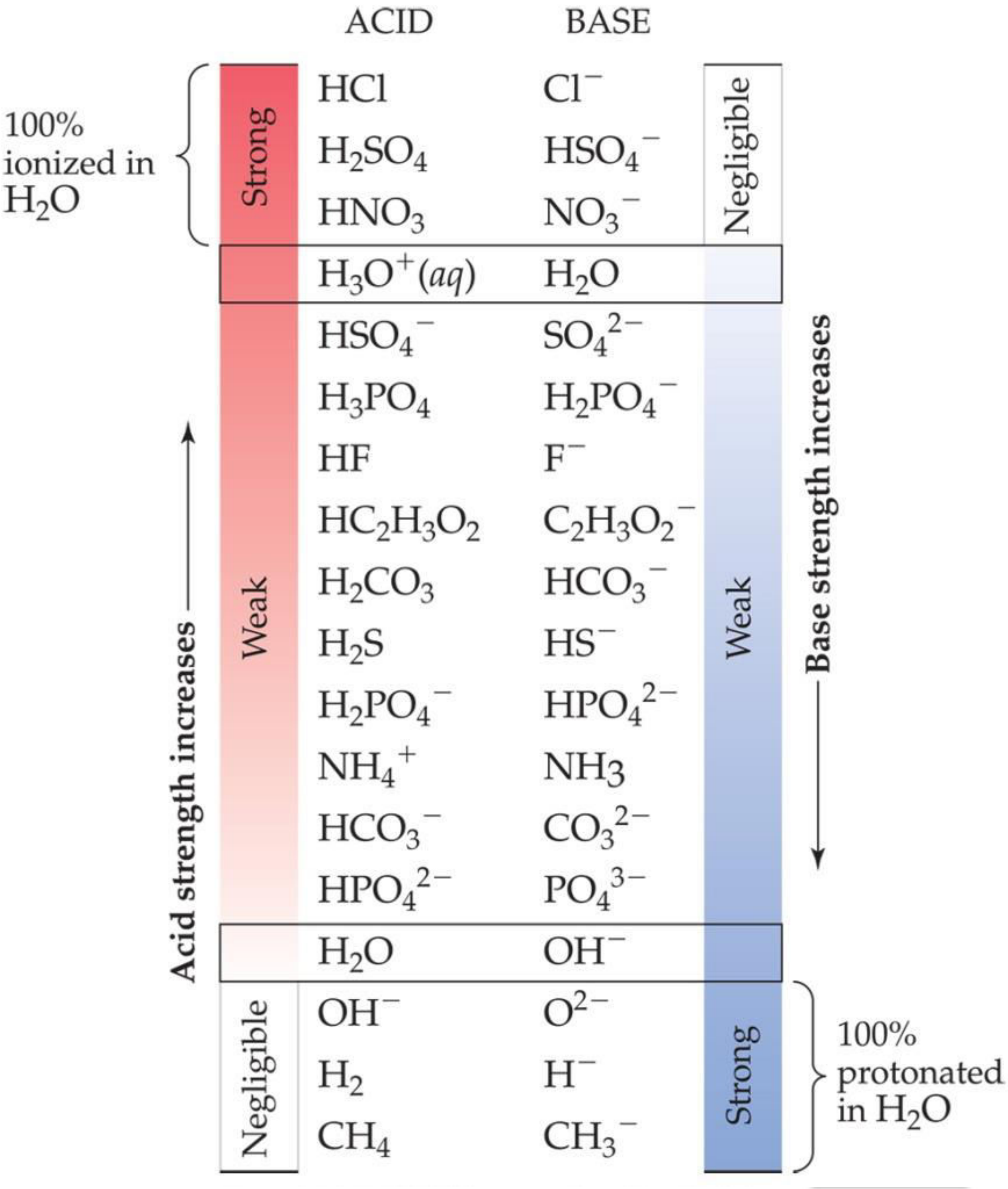
- Lewis acids are defined as **electron-pair acceptors**.
- Atoms with an empty valence orbital can be Lewis acids.

- Lewis bases are defined as **electron-pair donors**.
- Lewis bases can interact with things other than protons.



Acid and Base Strength

- Strong acids are completely dissociated in water.
  - ***Their conjugate bases are quite weak.***
- Weak acids only dissociate partially in water.
  - ***Their conjugate bases are weak bases.***



Acid strength	ACID	BASE	Base strength
Strong (100% ionized in H_2O)	HCl	Cl^-	Negligible
	H_2SO_4	HSO_4^-	
	HNO_3	NO_3^-	
	$H_3O^+(aq)$	H_2O	
Weak	HSO_4^-	SO_4^{2-}	Weak
	H_3PO_4	$H_2PO_4^-$	
	HF	F^-	
	$HC_2H_3O_2$	$C_2H_3O_2^-$	
	H_2CO_3	HCO_3^-	
	H_2S	HS^-	
	$H_2PO_4^-$	HPO_4^{2-}	
	NH_4^+	NH_3	
	HCO_3^-	CO_3^{2-}	
	HPO_4^{2-}	PO_4^{3-}	
	H_2O	OH^-	
Negligible	OH^-	O^{2-}	Strong (100% protonated in H_2O)
	H_2	H^-	
	CH_4	CH_3^-	

Acid strength increases ↑

Base strength increases ↓

Acid and Base Strength

- In any acid-base reaction, the equilibrium favors the reaction **that moves the proton to the stronger base**.

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

- **H_2O** is a much stronger base than **Cl^-**, so the equilibrium lies so far to the right, ***K* is not measured (*K*>>1)**.

Acid and Base Strength

$CH_3COOH(aq) + H_2O \rightleftharpoons H_3O^+(aq) + CH_3COO^-(aq)$

- **Acetate** (CH_3COO^-) is a stronger base than H_2O, so the equilibrium favors the left side (**$K<1$**).

- The stronger base (CH_3COO^-) "**wins**" the proton.

- What happens when **HCl** is dissolved in **acetic acid**? (an acid in an acid)?

$HCl + CH_3COOH(aq) \rightleftharpoons Cl^-(aq) + CH_3COOH_2^+(aq)$

- HCl is a weak acid in acetic acid?

Autoionization of Water

Water is amphoteric

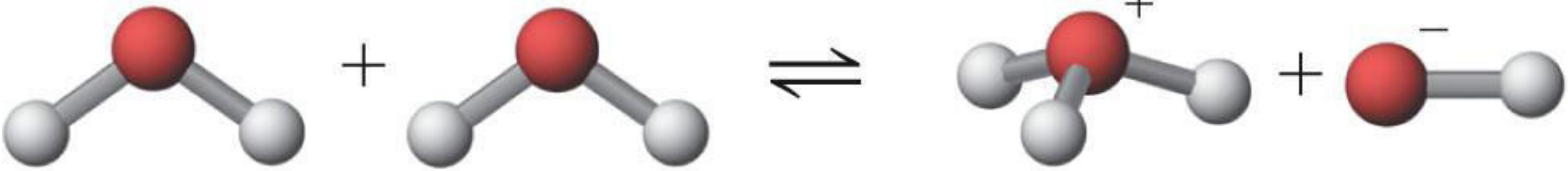
- In pure water, a few molecules act as **bases** and a few act as acids.

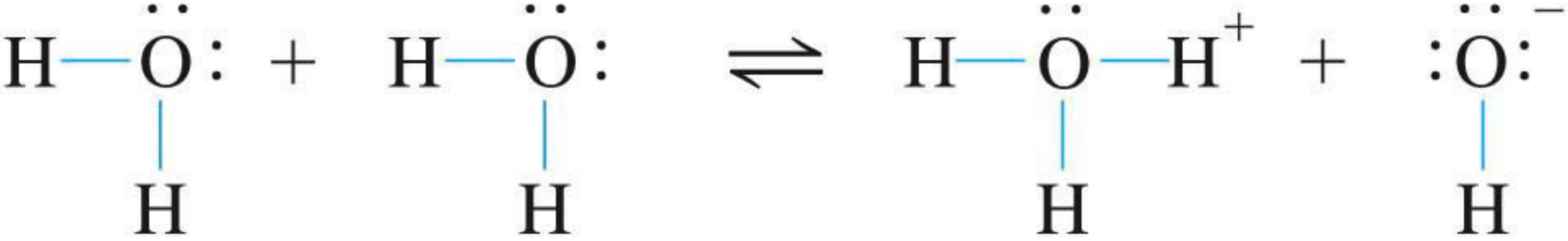
$$H_2O(l) + H_2O(l) \rightleftharpoons OH^-(aq) + H_30^+(aq)$$

This process is called **auto-ionization** .

Water Equilibrium

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





Self-ionization of water

Water molecules autoionize

$$H_2O(l) + H_2O(l) \rightleftharpoons OH^-(aq) + H_30^+(aq)$$

$$K_c = \frac{[H_3O^+]\,[OH^-]}{[H_2O]^2}$$

(**conc** = 1000/18 = **55.6**)

$K_w = [H_2O]^2 \times K_c = 1 \times 10^{-14}$ only at 25°C, it's *T* dependent.

$K_w = 55.6^2 \times K_c = \mathbf{1.0 \times 10^{-14}}$

At 25° C, $K_w = 1.0 \times 10^{-14}$

K_w increases as ***T*** increases, and its value remains the same in the presence of acid or base.

Solutions of strong acids and bases

Strong acids and **strong bases** **completely ionize** in their solutions.

$$HCl\ (aq) + H_2O\ (l) \rightarrow H_3O^+\ (aq) + Cl^-\ (aq)$$
$$KOH\ (aq) \rightarrow K^+\ (aq) + OH^-\ (aq)$$

In a **0.100 M HCl or HNO_3 solution**,

$[H^+] = \mathbf{0.100\ M}$ **(pH = $-\log_{10}[H^+] = 1$)**

and $[OH^-] = \mathbf{1 \times 10^{-14} / 0.100} = 1 \times 10^{-13}$ M at 25°C

In a **0.100 M NaOH or KOH solution**,

$[OH^-] = \mathbf{0.100\ M}$

and $[H^+] = \mathbf{1 \times 10^{-14} / 0.100} = 1 \times 10^{-13}$ M at 25°C

(pH = $-\log_{10}[H^+] = 13$)

In a 0.010 M $Ca(OH)_2$ solution,

$[OH^-] = \mathbf{2} \times 0.010 = \mathbf{0.020\ M}$

$[H^+] = \mathbf{1 \times 10^{-14} / 0.020} = 5 \times 10^{-13}$M at 25°C

(pH = $-\log_{10}[H^+] = 12.3$)

pH

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

- In pure water,

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

- Because in pure water $[H_3O^+] = [OH^-]$,

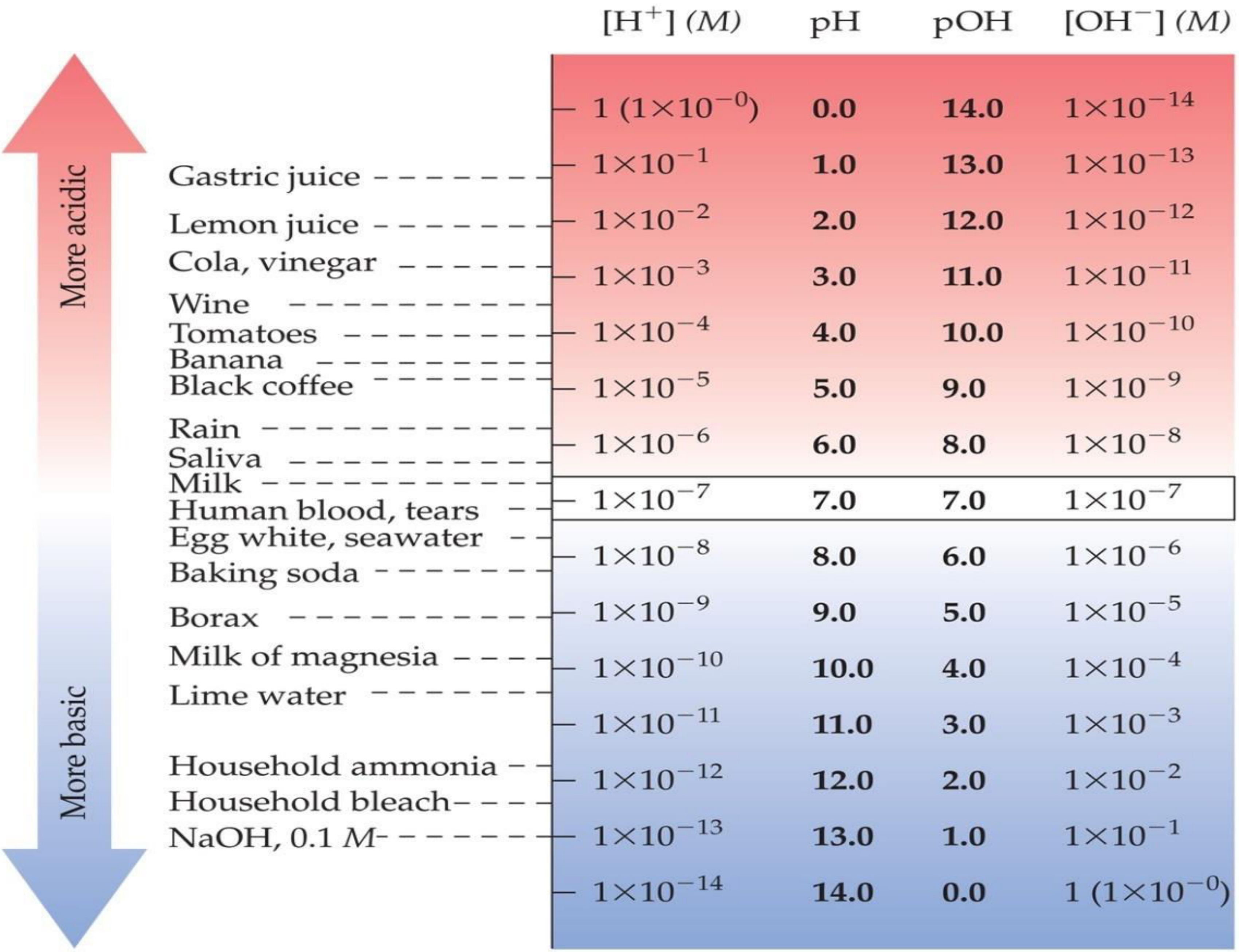
$$[H_3O^+] = (1.0 \times 10^{-14})^{1/2} = \underline{1.0 \times 10^{-7}}$$

- An **acid** has a higher $[H_3O^+]$ than pure water, so its pH is <7
- A **base** has a lower $[H_3O^+]$ than pure water, so its **pH is >7**.

➢ $pOH = -\log [OH^-]$

➢ $pK_w = -\log K_w$

pH values for common substances.



$[H^+]$ (*M*)	pH	pOH	$[OH^-]$ (*M*)
$1\ (1\times10^{-0})$	**0.0**	**14.0**	1×10^{-14}
1×10^{-1}	**1.0**	**13.0**	1×10^{-13}
1×10^{-2}	**2.0**	**12.0**	1×10^{-12}
1×10^{-3}	**3.0**	**11.0**	1×10^{-11}
1×10^{-4}	**4.0**	**10.0**	1×10^{-10}
1×10^{-5}	**5.0**	**9.0**	1×10^{-9}
1×10^{-6}	**6.0**	**8.0**	1×10^{-8}
1×10^{-7}	**7.0**	**7.0**	1×10^{-7}
1×10^{-8}	**8.0**	**6.0**	1×10^{-6}
1×10^{-9}	**9.0**	**5.0**	1×10^{-5}
1×10^{-10}	**10.0**	**4.0**	1×10^{-4}
1×10^{-11}	**11.0**	**3.0**	1×10^{-3}
1×10^{-12}	**12.0**	**2.0**	1×10^{-2}
1×10^{-13}	**13.0**	**1.0**	1×10^{-1}
1×10^{-14}	**14.0**	**0.0**	$1\ (1\times10^{-0})$

Because

$$[H_3O^+]\,[OH^-] = K_w \;\; = \mathbf{1.0 \times 10^{-14}},$$

Taking **–ve log** on both sides of equation:-

$$-\log\,[H_3O^+] + -\log\,[OH^-] = -\log\,K_w$$

or, in other words,

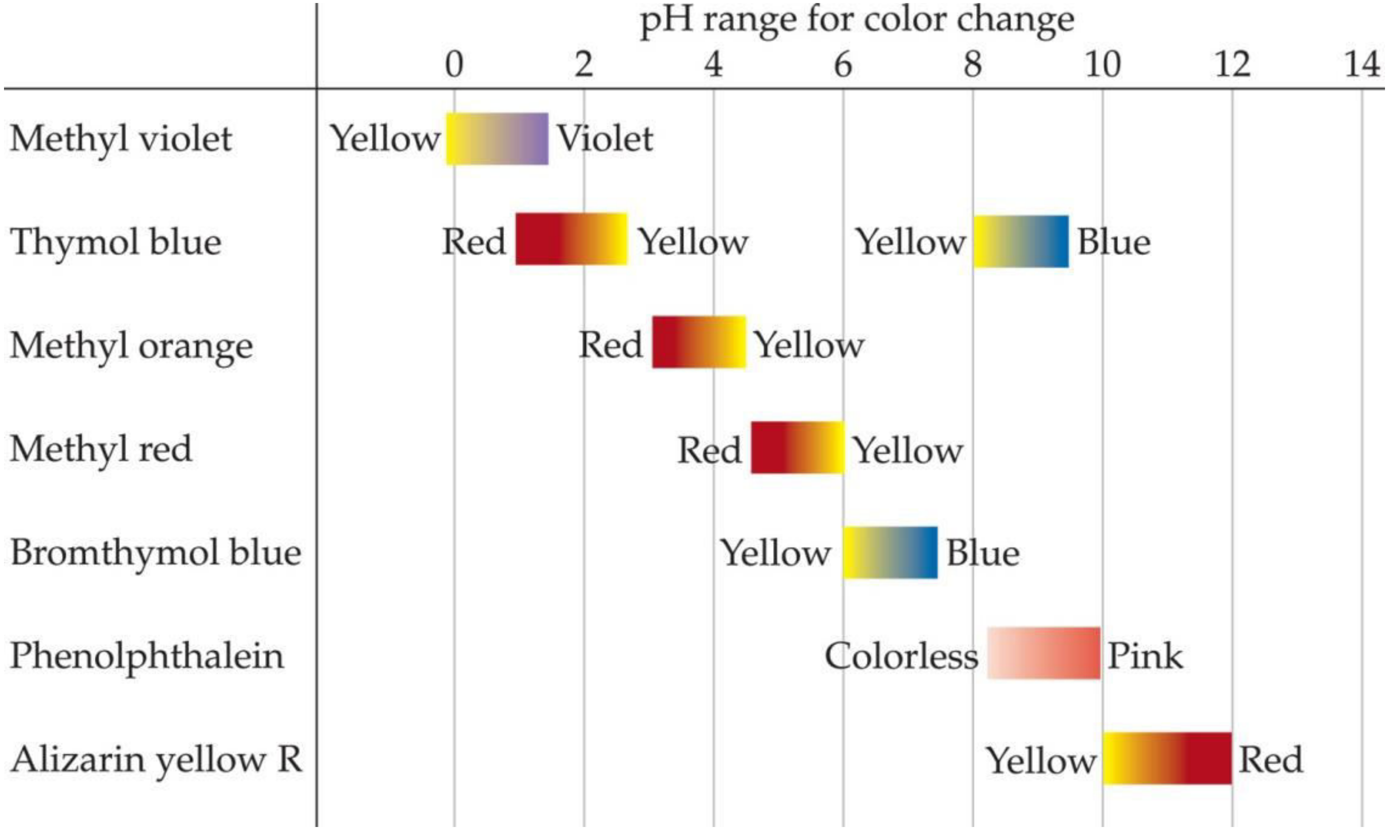
$$\mathbf{pH + pOH = p}K_w \mathbf{= 14.00}$$

If you know one, you know the rest:

$[H^+]$, $[OH^-]$, pH , pOH

Bases

How is pH measured ?



- **Litmus paper**
  - “Red” paper turns **blue** above ~pH = 8
  - “Blue” paper turns **red** below ~pH = 5
- **An indicator**
  - Compound that changes color in solution

How Do We Measure pH?

pH meters: Measure the **voltage** in the solution

Strong Acids

- The seven strong acids are HCl, HBr, HI, HNO_3, H_2SO_4, $HClO_3$, and $HClO_4$.

- These are strong electrolytes and exist totally as **ions** in aqueous solution.

- For the ***monoprotic*** strong acids, $[\mathbf{H_3O^+}] = [\mathbf{Acid}]$.
- For the ***diprotic*** strong acids, $[\mathbf{H_3O^+}] = \mathbf{2x}[\mathbf{Acid}]$.
- For the ***triprotic*** strong acids, $[\mathbf{H_3O^+}] = \mathbf{3x}\ [\mathbf{Acid}]$.

Strong Bases

- Strong bases are the soluble hydroxides, which are the alkali metal (**NaOH**, **KOH**)and heavier alkaline earth metal hydroxides e.g. **$Ca(OH)_2$**.

- These substances dissociate completely in aqueous solution.

$[OH^-]$ = [hydroxide added] x equivalent of OH^- released

- Are strong electrolytes and exist totally as **ions** in aqueous solution.

Dissociation Constants

- For a generalized acid dissociation in water,

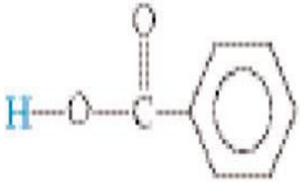
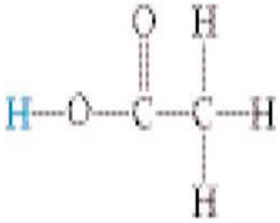
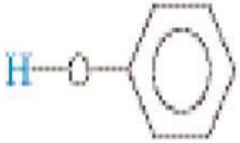
$$HA(aq) + H_2O(l) \rightleftharpoons A^-(aq) + H_3O^+(aq)$$

the equilibrium expression is

$$K_c = K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

- This equilibrium constant is called the acid-dissociation constant, ***K_a***.

The greater the value of K_a, the stronger the acid.

Acid	Structural Formula	Conjugate Base	Equilibrium Reaction	K_a
Hydrofluoric (HF)	H—F	F^-	$HF(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + F^-(aq)$	6.8×10^{-4}
Nitrous (HNO_2)	H—O—N=O	NO_2^-	$HNO_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + NO_2^-(aq)$	4.5×10^{-4}
Benzoic ($HC_7H_5O_2$)		$C_7H_5O_2^-$	$HC_7H_5O_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + C_7H_5O_2^-(aq)$	6.3×10^{-5}
Acetic ($HC_2H_3O_2$)		$C_2H_3O_2^-$	$HC_2H_3O_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + C_2H_3O_2^-(aq)$	1.8×10^{-5}
Hypochlorous (HClO)	H—O—Cl	ClO^-	$HClO(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + ClO^-(aq)$	3.0×10^{-8}
Hydrocyanic (HCN)	H—C≡N	CN^-	$HCN(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CN^-(aq)$	4.9×10^{-10}
Phenol (HC_6H_5O)		$C_6H_5O^-$	$HC_6H_5O(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + C_6H_5O^-(aq)$	1.3×10^{-10}

*The proton that ionizes is shown in blue.

Calculating K_a from the pH

- The pH of a **0.10 M** solution of formic acid, HCOOH, at 25° C is **2.38**. Calculate K_a for formic acid at this temperature.

$$HCOOH + H_2O \rightleftharpoons HCOO^- + H_3O^+$$

- To calculate ***K_a***, we need all **equilibrium concentrations**.

- We can find [**H_3O^+**], which is the same as [**$HCOO^-$**], from the **pH**.

$$K_a = \frac{[H_3O^+]\,[HCOO^-]}{[HCOOH]}$$

Calculating K_a from the pH *cont.*

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

$$2.38 = -\log[H_3O^+]$$

$$-2.38 = \log[H_3O^+]$$

$$\boxed{[H_3O^+] = \text{antilog } {-2.38} = 10^{-2.38}}$$

$[H_3O^+] = [HCOO^-] =$ **4.2×10^{-3} M**

Calculating K_a from pH

$$HCOOH + H_2O \rightleftharpoons HCOO^- + H_3O^+$$

In table form:

	[HCOOH], M	[H_3O^+], M	[$HCOO^-$], M
Initially	0.10	0	0
Change	-4.2×10^{-3}	$+4.2 \times 10^{-3}$	$+4.2 \times 10^{-3}$
At Equilibrium	**$0.10 - 4.2 \times 10^{-3}$**	4.2×10^{-3}	4.2×10^{-3}
At Equilibrium	**= 0.0958 ≈ 0.10**	**4.2×10^{-3}**	**4.2×10^{-3}**

$$K_a = \frac{[H_3O^+][HCOO^-]}{[HCOOH]}$$

Calculating K_a from pH

$$HCOOH + H_2O \rightleftharpoons HCOO^- + H_3O^+$$

$$K_a = \frac{[H_3O^+]\,[HCOO^-]}{[HCOOH]}$$

$$Ka \;=\; \frac{4.2 \times 10^{-3} \text{ x } 4.2 \times 10^{-3}}{0.10 - 4.2 \times 10^{-3}}$$

$$K_a \approx \frac{[4.2 \times 10^{-3}]\,[4.2 \times 10^{-3}]}{[0.10]}$$

$$K_a \approx 1.8 \times 10^{-4}$$

Calculating Percent Ionization

$$HCOOH + H_2O \rightleftharpoons HCOO^- + H_3O^+$$

$$Percent\ ionization = \frac{amount\ ionized}{total\ in\ solution} \times 100$$

In the example:

$[HCOO^-]_{eq} = [H_3O^+]_{eq} = \underline{4.2 \times 10^{-3}\ M}$

$[HCOOH]_{initial} = \underline{0.10\ M}$

$$\text{Percent Ionization} = \frac{4.2 \times 10^{-3}}{0.10} \times 100$$

$= 4.2\%$

Calculating pH from K_a

$\mathrm{p}K_a = -\log K_a$ $K_a = 10^{-\mathrm{p}K_a}$

Calculate the pH of a <u>0.30 *M*</u> solution of acetic acid, $C_2H_3O_2H$, at 25° C.

$$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$$

K_a for acetic acid at 25° C is **1.8×10^{-5}**.

Q: Is acetic acid more or less ionized than formic acid (K_a=1.8×10^{-4})?

Calculating pH from K_a

$$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$$

The equilibrium constant expression is:

$$K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]}$$

Calculating pH from K_a

$$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$$

Use the ICE table:

	$[C_2H_3O_2]$, *M*	$[H_3O^+]$, *M*	$[C_2H_3O_2^-]$, *M*
Initial	0.30	0	0
Change	$-x$	$+x$	$+x$
Equilibrium	$0.30 - x$	x	x

$$\boxed{K_a = \frac{x^2}{0.30 - x}}$$

Simplify: how big is x relative to 0.30? *i.e.,*

Equilibrium	$0.30 - x \approx 0.30$	$\boldsymbol{x}$	$\boldsymbol{x}$

Calculating pH from K_a

$$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$$

$$\boxed{K_a = \frac{x^2}{0.30 - x} \simeq \frac{x^2}{0.30}}$$

NB: $x = [H_3O^+]$

$K_a \times 0.3 = \boldsymbol{x^2}$

$(1.8 \times 10^{-5})\ (0.30) = x^2$

$5.4 \times 10^{-6} = x^2$

$2.3 \times 10^{-3} = x$

$\text{pH} = -\log [H_3O^+]$

$\text{pH} = -\log (2.3 \times 10^{-3})$

pH = **2.64**

Determine K_a and percent ionization

Nicotinic acid, HNic, is a monoprotic acid. A solution containing 0.012 M HNic, has a pH of 3.39. What is its K_a? What is the percent of ionization?

Solution:

	HNic	=	H^+	+	Nic^-
Initial	0.012	=	0		0
Eqm	0.012 - x	=	x		x

$$x = [H^+] = 10^{-pH} = 10^{-3.39} = 4.1 \times 10^{-4}$$

$$[HNic] = 0.012 - 0.00041 = 0.012$$

$$K_a = \frac{(4.1 \times 10^{-4})^2}{0.012} = 1.4 \times 10^{-5}$$

Degree of ionization = 0.00041 / 0.012 = 0.034 = **3.4%**

Application of K_a

The K_a of nicotinic acid, HNic, is 1.4×10^{-5}. A solution containing 0.22 M HNic. What is its pH? What is the degree of ionization?

Solution:

$$\mathrm{HNic} = \mathrm{H^+} + \mathrm{Nic^-}$$

$$0.22 - x = x + x$$

$$K_a = \frac{x^2}{0.22 - \mathbf{x}} \ (\approx \mathbf{0.22}) = 1.4 \times 10^{-5}$$

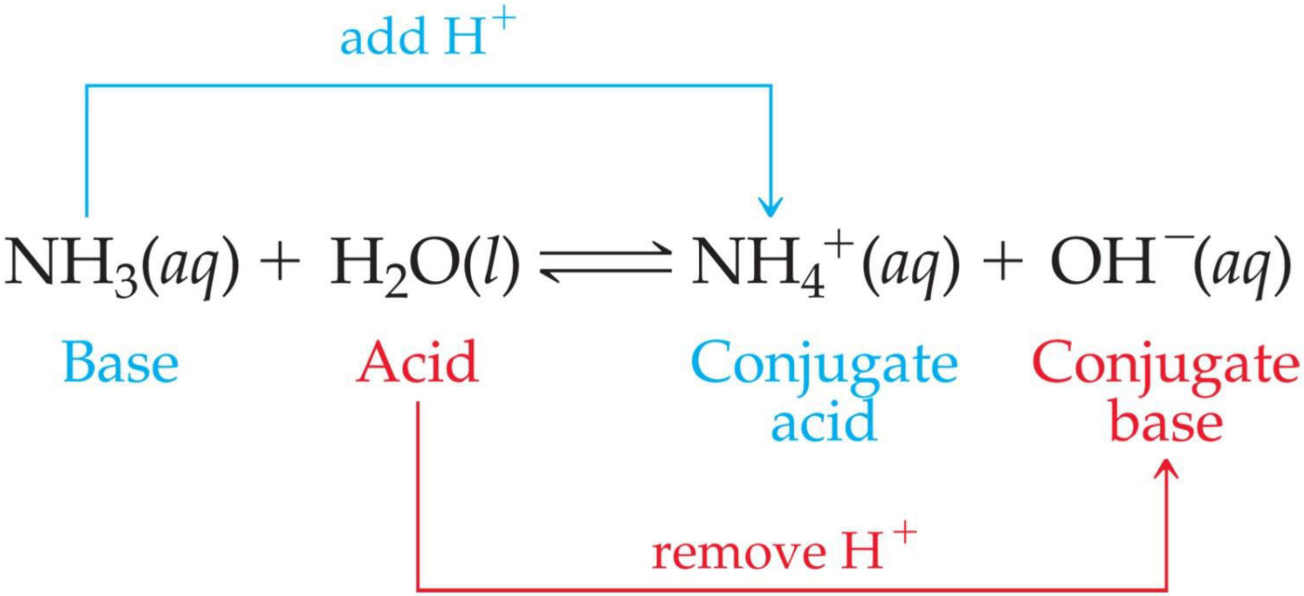
$$x = \sqrt{(0.22 * 1.4 \times 10^{-5})} = \underline{\mathbf{0.0018}}$$

$$\mathrm{pH} = -\log(0.0018) = \underline{\mathbf{2.76}}$$

Weak Bases

$$\ddot{B} + H_2O \rightleftharpoons HB^+ + OH^-$$

Bases react with water to produce hydroxide ion.



add H^+

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

Base

Acid

Conjugate acid

Conjugate base

remove H^+

Weak Bases

$$\ddot{B} + H_2O \rightleftharpoons HB^+ + OH^-$$

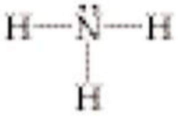
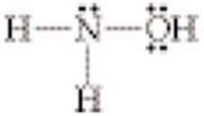
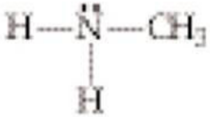
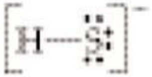
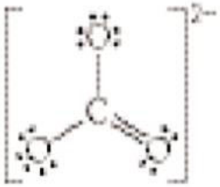
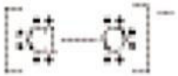
The equilibrium constant expression for this reaction is

$$K_c = K_b = \frac{[HB^+][OH^-]}{[B]}$$

where K_b is the base-dissociation constant.

Weak Bases

K_b can be used to find [OH$^-$] and, through it, pH.

Base	Lewis Structure	Conjugate Acid	Equilibrium Reaction	K_b
Ammonia (NH_3)		NH_4^+	$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$	1.8×10^{-5}
Pyridine (C_5H_5N)		$C_5H_5NH^+$	$C_5H_5N + H_2O \rightleftharpoons C_5H_5NH^+ + OH^-$	1.7×10^{-9}
Hydroxylamine (H_2NOH)		H_3NOH^+	$H_2NOH + H_2O \rightleftharpoons H_3NOH^+ + OH^-$	1.1×10^{-8}
Methylamine (NH_2CH_3)		$NH_3CH_3^+$	$NH_2CH_3 + H_2O \rightleftharpoons NH_3CH_3^+ + OH^-$	4.4×10^{-4}
Hydrosulfide ion (HS^-)		H_2S	$HS^- + H_2O \rightleftharpoons H_2S + OH^-$	1.8×10^{-7}
Carbonate ion (CO_3^{2-})		HCO_3^-	$CO_3^{2-} + H_2O \rightleftharpoons HCO_3^- + OH^-$	1.8×10^{-4}
Hypochlorite ion (ClO^-)		$HClO$	$ClO^- + H_2O \rightleftharpoons HClO + OH^-$	3.3×10^{-7}

pH of Basic Solutions

What is the pH of a **0.15 *M*** solution of NH_3?

$$\ddot{N}H_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = 1.8 \times 10^{-5}$$

	$[NH_3]$, *M*	$[NH_4^+]$, *M*	$[OH^-]$, *M*
Initial	0.15	0	0
Equilibrium	0.15 - ***x*** ≈ 0.15	***x***	***x***

Simplify: How big is **x** relative to **0.15?**

$$1.8 \times 10^{-5} = \frac{(x)^2}{(0.15)}$$

$$x^2 = (1.8 \times 10^{-5})\,(0.15)$$

$$x^2 = 2.7 \times 10^{-6}$$

$$x = 1.6 \times 10^{-3}$$

$$\boldsymbol{x} = [\mathrm{OH^-}] = \underline{\mathbf{1.6 \times 10^{-3}\ M}}$$

Therefore,

$[\mathrm{OH^-}] = 1.6 \times 10^{-3}\ M$

$\mathrm{pOH} = -\log\,(1.6 \times 10^{-3})$

$\mathrm{pOH} = 2.80$

$\mathrm{pH} = \mathbf{pKw - pOH^-}$

$= 14.00 - 2.80$

pH = <u>11.20</u>

K_a and K_b are linked:

$$\ddot{N}H_3 + H_2O \rightleftharpoons NH_4^+ + OH^- \quad K_b$$

$$NH_4^+ + H_2O \rightleftharpoons \ddot{N}H_3 + H_3O^+ \quad K_a$$

Combined reaction = ?

K_a and K_b are linked:

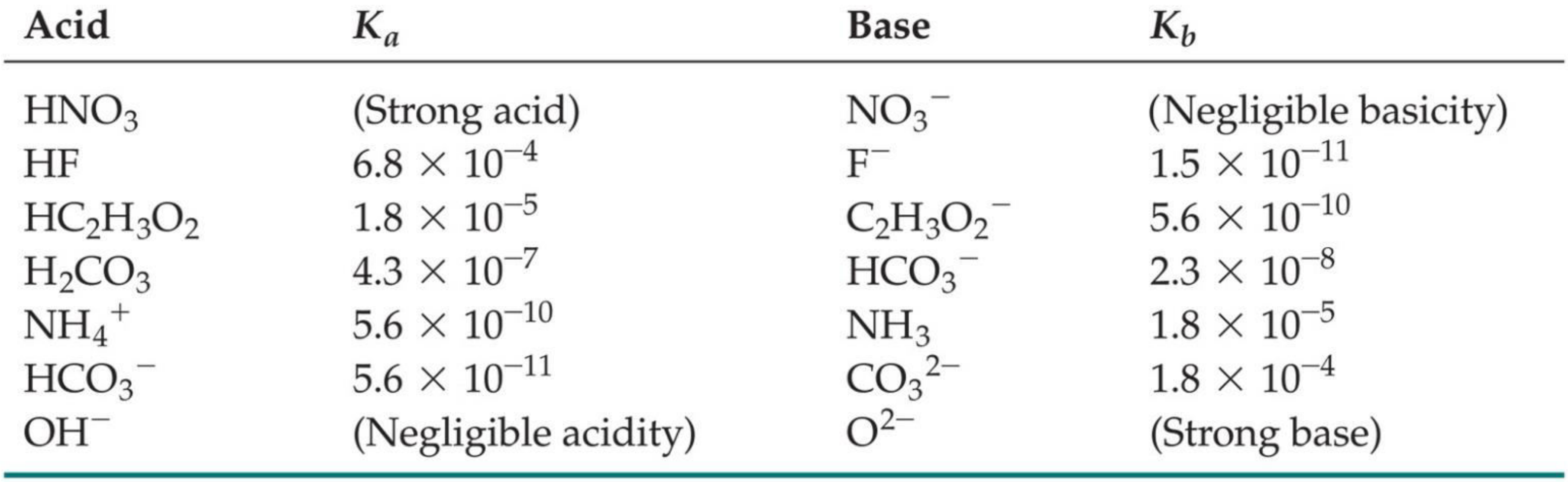
$$\ddot{N}H_3 + H_2O \rightleftharpoons NH_4^+ + OH^- \quad K_b$$

$$NH_4^+ + H_2O \rightleftharpoons \ddot{N}H_3 + H_3O^+ \quad K_a$$

Combined reaction

$$2H_2O \rightleftharpoons H_3O^+ + OH^- \quad K_a \times K_b = K_w$$

K_a and K_b

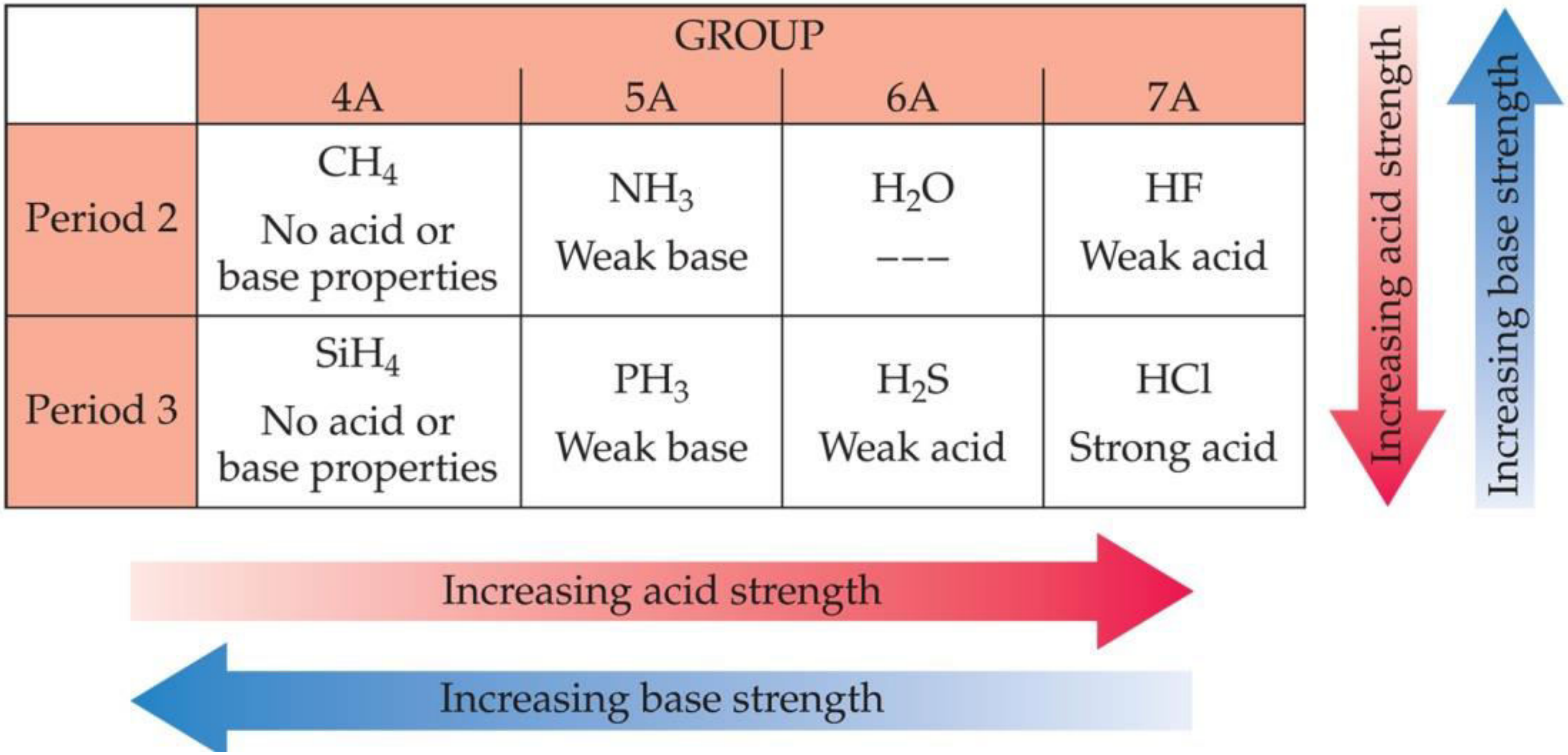


Acid	K_a	Base	K_b
HNO_3	(Strong acid)	NO_3^-	(Negligible basicity)
HF	6.8×10^{-4}	F^-	1.5×10^{-11}
$HC_2H_3O_2$	1.8×10^{-5}	$C_2H_3O_2^-$	5.6×10^{-10}
H_2CO_3	4.3×10^{-7}	HCO_3^-	2.3×10^{-8}
NH_4^+	5.6×10^{-10}	NH_3	1.8×10^{-5}
HCO_3^-	5.6×10^{-11}	CO_3^{2-}	1.8×10^{-4}
OH^-	(Negligible acidity)	O^{2-}	(Strong base)

K_a and K_b are related : $K_a \times K_b = K_w$

Therefore, if you know one of them, you can calculate the other.

Factors Affecting Acid Strength

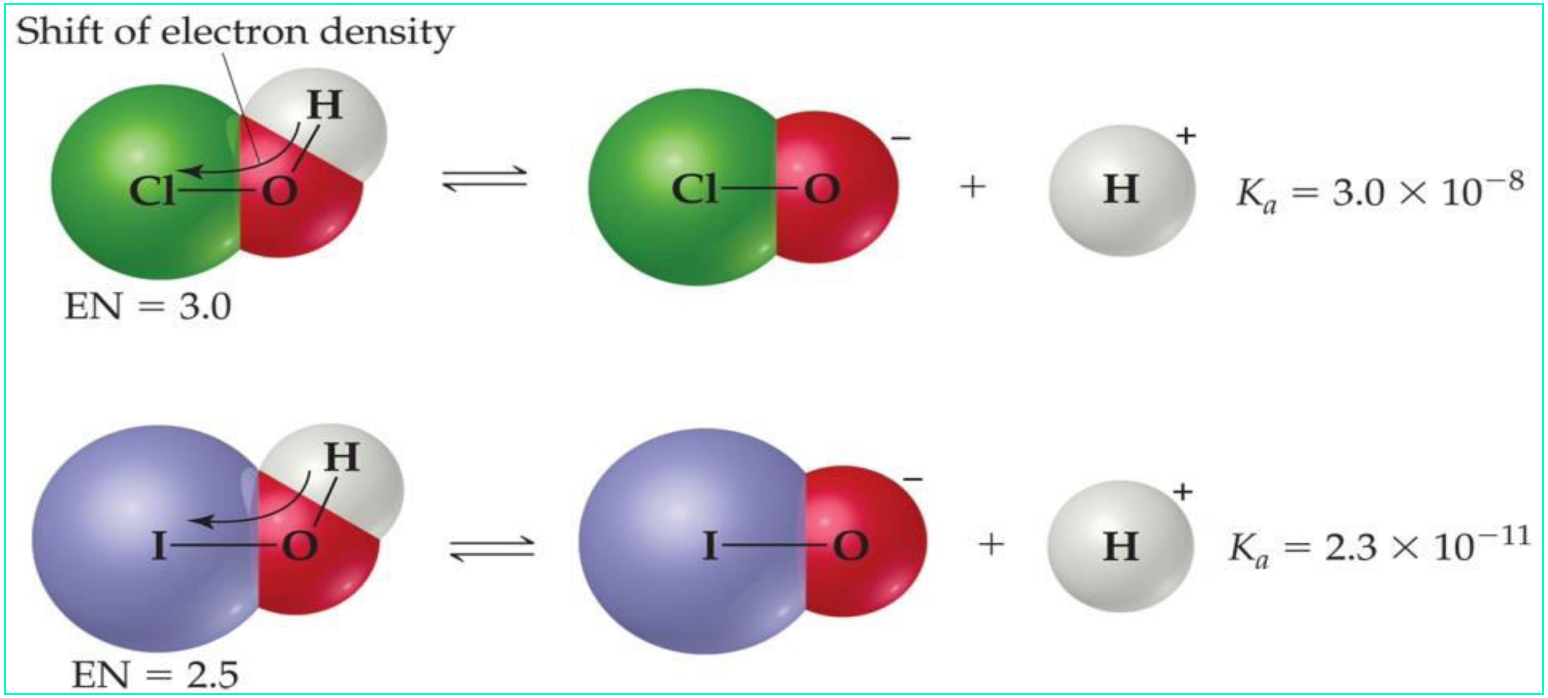


	GROUP			
	4A	5A	6A	7A
Period 2	CH_4 No acid or base properties	NH_3 Weak base	H_2O ---	HF Weak acid
Period 3	SiH_4 No acid or base properties	PH_3 Weak base	H_2S Weak acid	HCl Strong acid

- The more **polar** the **H-X bond** and/or the **weaker the H-X bond**, the more acidic the compound.
- Acidity increases from left to right across a Period and from top to bottom down a group.

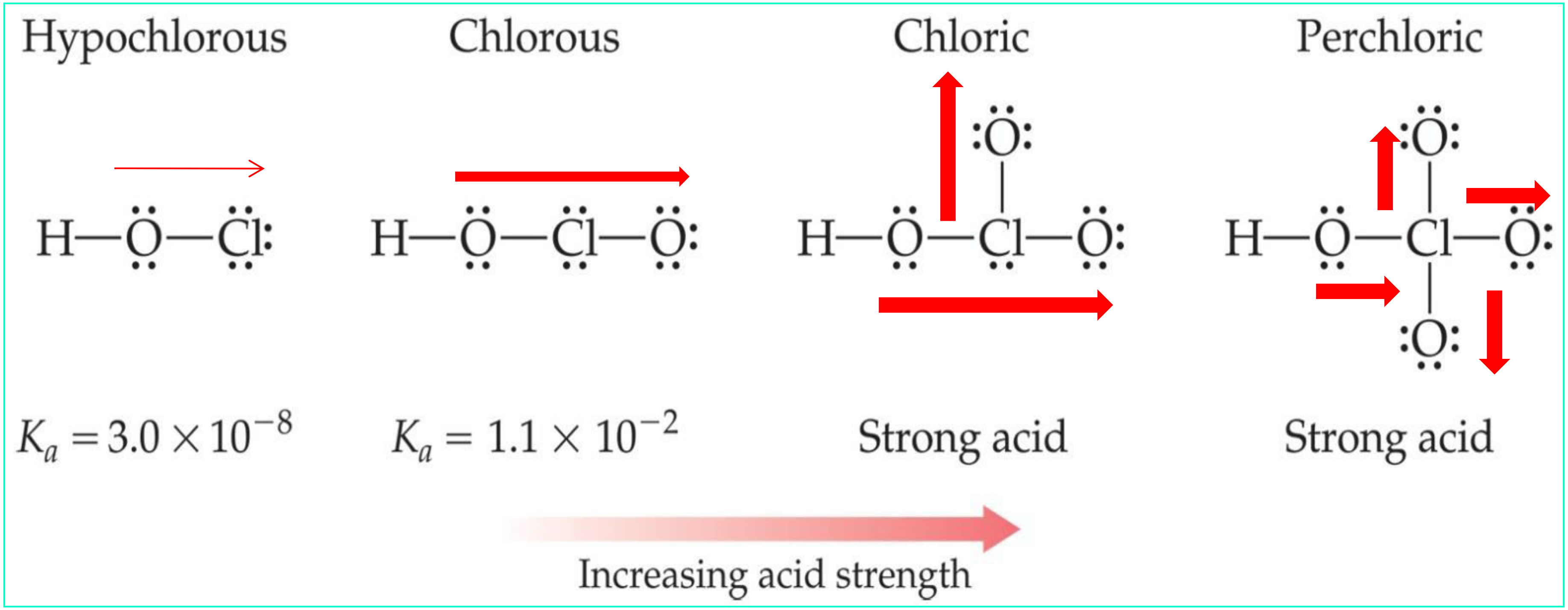
Factors Affecting Acid Strength

In **oxyacids**, in which an **OH** is bonded to another atom, **Y**, the **more electronegative** **Y** is, the more acidic the acid.



Acid	EN of Y	K_a
HClO	3.0	3.0×10^{-8}
HBrO	2.8	2.5×10^{-9}
HIO	2.5	2.3×10^{-11}

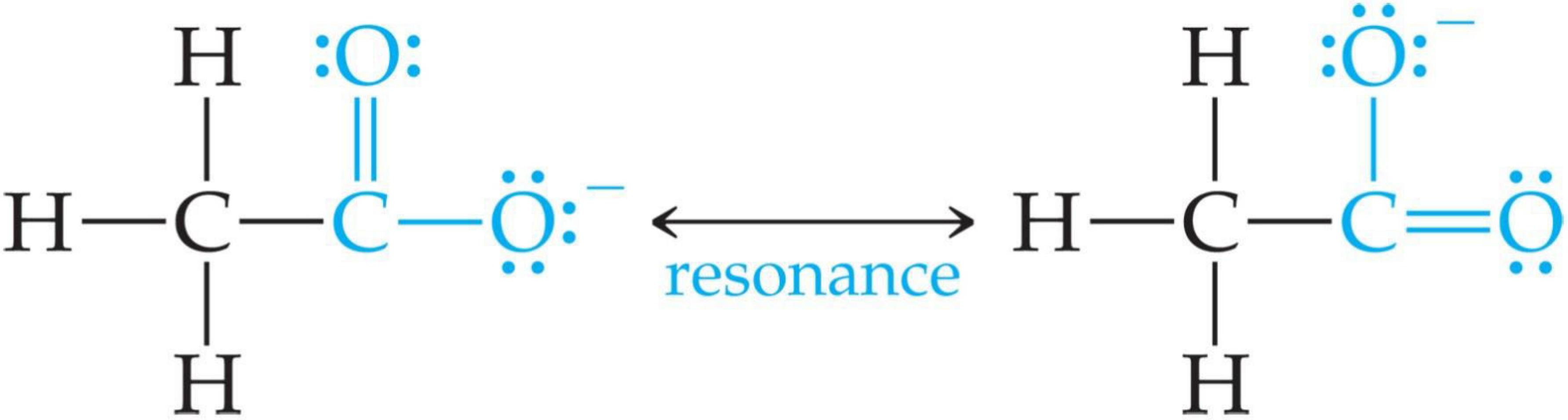
Factors Affecting Acid Strength



For a series of oxyacids, acidity increases with the **number** of **oxygens**.

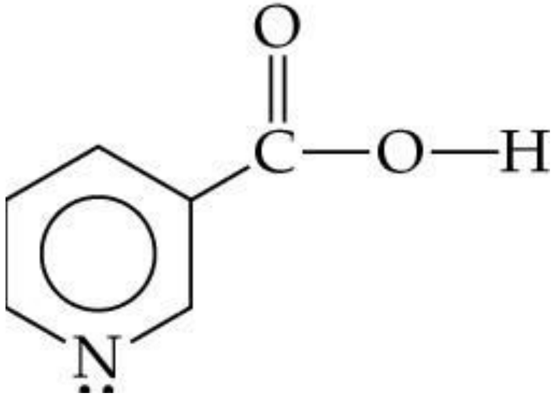
Factors Affecting Acid Strength

Resonance in the conjugate bases of carboxylic acids **stabilizes** the base and makes the conjugate acid more acidic.



PRACTICE EXERCISES

Niacin, one of the B vitamins, has the following molecular structure.

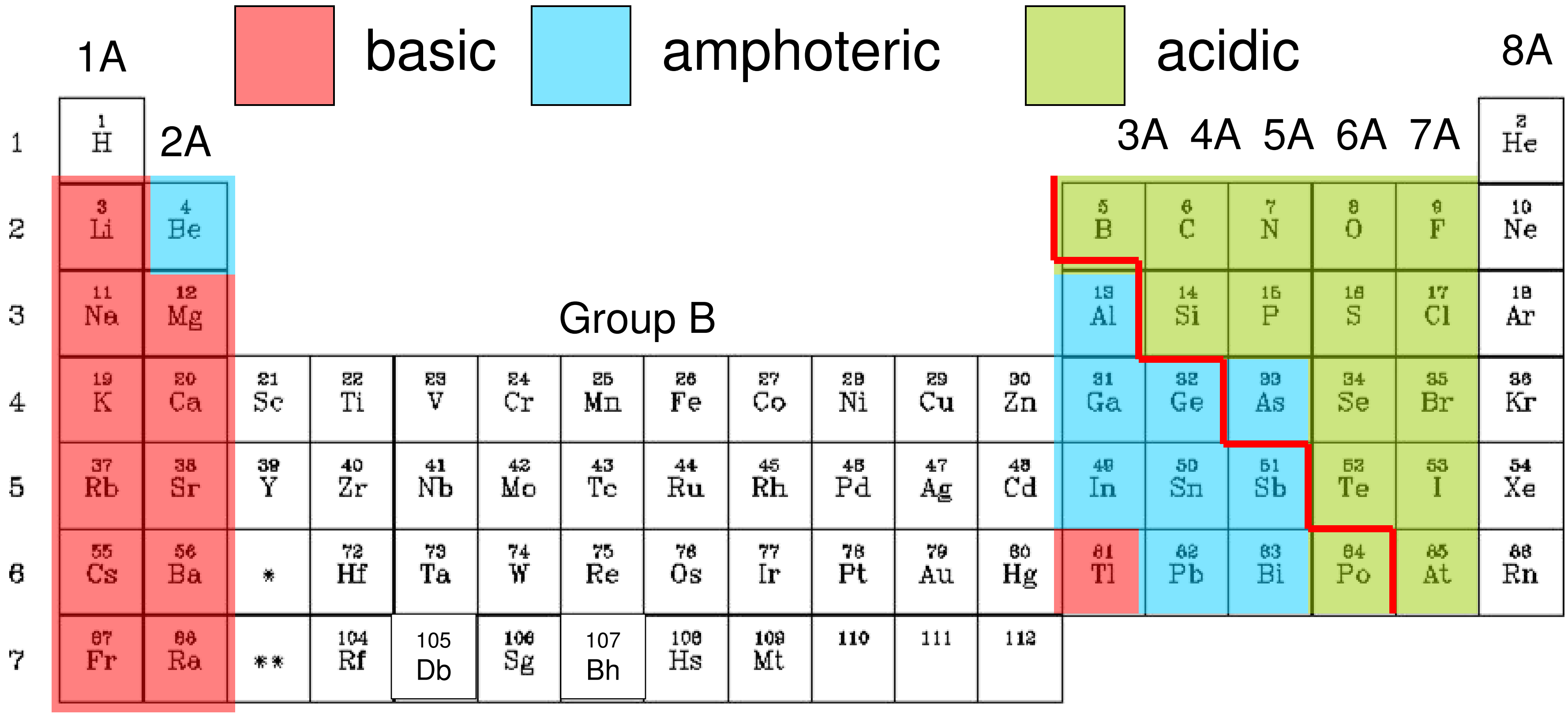


1. A 0.020 *M* solution of niacin has a pH of 3.26.
   (a) What percentage of the acid is ionized in this solution?
   (b) What is the acid-dissociation constant, K_a, for niacin?

2. Calculate the percentage of HF molecules ionized in **(a)** a 0.10 *M* HF solution, **(b)** a 0.010 *M* HF solution. **Ka for HF is 6.8 x10^{-4}.**

3. A solution of acetic acid is 2% ionized at 25° C. **K_a=1.8x10^{-5}**. What was the original concentration of the acid?

Behavior of oxides in water– Group A



🟥 basic 🟦 amphoteric 🟩 acidic

	1A	2A	Group B										3A	4A	5A	6A	7A	8A
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	*	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	**	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110	111	112						

basic: $Na_2O + H_2O \rightarrow 2NaOH$

($O^{2-} + H_2O \rightarrow 2OH^{-}$)

acidic: $CO_2 + H_2O \rightarrow H_2CO_3$