

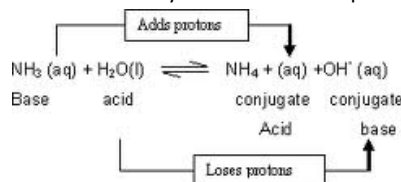
The Concepts of Acids and Bases

Concept	Definition	Equation Example
Arrhenius	Acids: substances that when dissolved in water, increase concentration of H ⁺ ions Bases: substances that when dissolved in water, increase concentration of OH ⁻ ions	$\text{HCl}_{(g)} \rightarrow \text{H}^+_{(aq)} + \text{Cl}^-_{(aq)}$
Bronsted-Lowry	Acids: substances that can transfer a proton. It must have a hydrogen atom. Base: substances that can accept a proton. It must have a nonbonding pair of electrons.	Acid: $\text{HCl}_{(g)} + \text{H}_2\text{O}_{(l)} \rightarrow \text{H}_3\text{O}^+_{(aq)} + \text{Cl}^-_{(aq)}$ Base: $\text{NH}_3_{(aq)} + \text{H}_2\text{O}_{(l)} \leftrightarrow \text{NH}_4^+_{(aq)} + \text{OH}^-_{(aq)}$
Lewis Acids and Bases	Acids: electron pair acceptors Bases: electron pair donors	$\begin{array}{c} \text{H} & & \text{F} \\ \vdots & & \vdots \\ \text{H} : \text{N} : & & \text{B} : \text{F} \\ \vdots & & \vdots \\ \text{H} & & \text{F} \end{array} \longrightarrow \begin{array}{c} \text{H} & \text{F} \\ \vdots & \vdots \\ \text{H} : \text{N} : \text{B} : \text{F} \\ \vdots & \vdots \\ \text{H} & \text{F} \end{array}$ <p>Lewis Base Lewis Acid</p>

Amphoteric: Substances capable of acting as either a base or an acid
Example: H₂O (as shown in Bronsted-Lowry example)

Acid-Base Pairs

- Conjugate Acid-Base Pair:** And acid and base that only differ in the absence or presence of a proton
 - Conjugate Base:** Formed by removal of a proton from the acid.
 - Conjugate Acid:** Formed by the addition of a proton to a base.



Relative Strengths of Acids and Bases

- The stronger the acid, the weaker its conjugate base
- The stronger the base, the weaker its conjugate acid

The pH Scale

$\text{pH} = -\log[\text{H}^+]$	pH=3	$3 = -\log[\text{H}^+]$	$10^{-3} = \text{H}^+$
$\text{pOH} = -\log[\text{OH}^-]$	$\text{OH}^- = 1 \times 10^{-4}$	$\text{pOH} = -\log(1 \times 10^{-4})$	pOH = 4

* **Strong Acids:** HCl, HBr, HI, HNO₃, HClO₃, HClO₄, H₂SO₄

Types of Acids

- Monoprotic:** Only one hydrogen can be liberated (Ex. HCl, HBr, HI)
- Polyprotic:** More than one acidic hydrogen can be liberated (Ex. H₂SO₄, H₃PO₄)
- Oxyacids (HaXbOc):** Proton attached to oxygen of an ion (Ex. HNO₃, H₃PO₄)
- Organic Acids:** Contain carboxyl (-COOH) with H attached to O (Ex. CH₃COOH)

Weak Acids	Partially ionize in aqueous solutions	Acid Dissociation Constant (K_a): Equilibrium constant for the ionization of an acid	The larger the value of K _a , the stronger the acid
Weak Bases	a chemical base that does not ionize fully in an aqueous solution	Base Dissociation Constant (K_b): Equilibrium where a base reacts with water to form a conjugate acid and OH ⁻	The larger the value of K _b , the stronger the base

	ACID	BASE		
↑ Acid strength increases	Strong	HCl, H ₂ SO ₄ , HNO ₃	Cl ⁻ , HSO ₄ ⁻ , NO ₃ ⁻	
	Weak	H ₃ O ⁺ (aq)	H ₂ O	
		HSO ₄ ⁻	SO ₄ ²⁻	
		H ₃ PO ₄	H ₂ PO ₄ ⁻	
		HF	F ⁻	
		HC ₂ H ₃ O ₂	C ₂ H ₃ O ₂ ⁻	
		H ₂ CO ₃	HCO ₃ ⁻	
		H ₂ S	HS ⁻	
		H ₂ PO ₄ ⁻	HPO ₄ ²⁻	
		NH ₄ ⁺	NH ₃	
HCO ₃ ⁻		CO ₃ ²⁻		
HPO ₄ ²⁻	PO ₄ ³⁻			
↓ Base strength increases	H ₂ O	OH ⁻		
	OH ⁻	O ²⁻		
	H ₂	H ⁻		
	CH ₄	CH ₃ ⁻		

100% protonated in H₂O

***Sig figs of pH:** the number of sig figs in the lead number is the number of decimal places for the pH value.

Ion Product of Water =

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

Relationship Between K_a and K_b

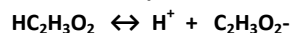
$$K_a \times K_b = K_w$$

$$\text{p}K_a + \text{p}K_b = \text{p}K_w = 14.00$$

% Dissociation =

$$\frac{\text{amount dissociated}}{\text{initial concentration}} \times 100$$

EXAMPLE 1: Calculate pH of 2 M acetic acid (HC₂H₃O₂) with a K_a of 1.8 x 10⁻⁵.



2.0 M	0	0
-X	+X	+X
2.0-X	X	X

$$K_a = \frac{[\text{X}][\text{X}]}{[2.0-\text{X}]} \quad \text{5\% rule makes it } K_a = \frac{X^2}{2.0} \quad K_a = 1.8 \times 10^{-5}$$

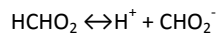
$$1.8 \times 10^{-5} = \frac{X^2}{2.0} \quad \sqrt{2 \times 1.8 \times 10^{-5}} = X \quad \text{H}^+ = X = .006 \text{ M}$$

$$\text{pH} = -\log(.006) = \mathbf{2.22}$$

Find percent dissociation

$$\% \text{ dissociation} = \left(\frac{.006 \text{ M}}{2.0 \text{ M}} \right) = \mathbf{.3\%}$$

EXAMPLE 2: Calculate the K_a of a 0.10 M solution HCHO_2 if the pH at standard temperature is 2.38.



$$K_a = \frac{[\text{H}^+][\text{CHO}_2^-]}{[\text{HCHO}_2]}$$

$$\text{pH} = -\log[\text{H}^+] = 2.38$$

$$10^{-2.3} = \text{H}^+$$

$$\text{H}^+ = 4.2 \times 10^{-3} \text{ M}$$



0.10 M	0	0
-4.2×10^{-3}	$+4.2 \times 10^{-3}$	$+4.2 \times 10^{-3}$
$0.10 - 4.2 \times 10^{-3}$	4.2×10^{-3}	4.2×10^{-3}

$$0.10 - 4.2 \times 10^{-3} \approx 0.10$$

$$K_a = (4.2 \times 10^{-3})^2 / 0.10 = 1.8 \times 10^{-4}$$

Find K_b

$$K_a \times K_b = K_w$$

$$(1.0 \times 10^{-14}) / (1.8 \times 10^{-4}) = K_b$$

$$K_b = 5.56 \times 10^{-11}$$

Polyprotic Acids

- The second K_a (K_{a2}) is much smaller than K_{a1} because it is easier to remove the first proton.

EXAMPLE 3: Solve for the concentration of CO_3^{2-} given that the solubility of CO_2 in pure water at standard conditions is .0037 M,

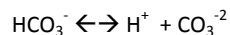
$K_{a1} = 4.3 \times 10^{-7}$, and $K_{a2} = 5.6 \times 10^{-11}$ ($\text{CO}_2 + \text{H}_2\text{O} \leftrightarrow \text{H}_2\text{CO}_3$)



0.0037 M	0	0
-X	+X	+X
.0037-X	X	X

$$K_a = 4.3 \times 10^{-7} = (X)(X) / .0037 - X$$

$$X = 4.0 \times 10^{-5}$$



4.0×10^{-5}	4.0×10^{-5}	0
-Y	+Y	+Y
$4.0 \times 10^{-5} - Y$	$4.0 \times 10^{-5} + Y$	Y

$$K_{a2} = 5.6 \times 10^{-11} = (4.0 \times 10^{-5})(Y) / (4.0 \times 10^{-5})$$

$$\text{CO}_3^{2-} = Y = 5.6 \times 10^{-11}$$

Strengths of Different Acids

Binary Acids	Acids made up of 2 elements	Acid strength increases as you move down group
H-O-Y Oxyacids	Acids in which OH groups are bound to a central atom and some atom Y	Acid strength increases as electronegativity of Y increases
H-Y-On Oxyacids	Acids in which H is bounded to some atom Y and multiple O	Acid strength increases as the number of oxygen atoms attached to Y increases
Carboxylic Acids	Acids that contain a carboxyl group	Acid strength increases as the number of electronegative atoms in the acid increase