

## Acids and Bases

### Brønsted-Lowry Definition

Acid is a proton ( $\text{H}^+$ ) donor

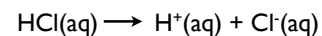
Base is a proton ( $\text{H}^+$ ) acceptor

## Strong Acids and Bases

"Strong" means one thing

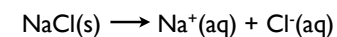
The substance dissociates 100% in water

### Strong Acid



$$K_a = \frac{[\text{H}^+][\text{Cl}^-]}{[\text{HCl}]} \approx \infty$$

### Strong Electrolyte



$$K_{sp} = [\text{Na}^+][\text{Cl}^-] \approx \infty$$

## Strong Acids

HCl	Hydrochloric
HBr	Hydrobromic
HI	Hydroiodic
$\text{HClO}_4$	Perchloric
$\text{HClO}_3$	Chloric
$\text{H}_2\text{SO}_4$	Sulfuric
$\text{HNO}_3$	Nitric

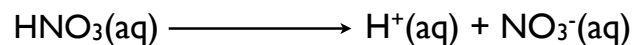
All Dissociate 100%

## Strong Bases

Lithium Hydroxide	LiOH
Sodium Hydroxide	NaOH
Potassium Hydroxide	KOH
Rubidium Hydroxide	RbOH
Cesium Hydroxide	CsOH
Calcium Hydroxide	$\text{Ca(OH)}_2$
Barium Hydroxide	$\text{Ba(OH)}_2$
Strontium Hydroxide	$\text{Sr(OH)}_2$

All Dissociate 100%

What is the pH of a 0.1 M solution of Nitric Acid



$[\text{H}^+] = C_a$   $C_a$  is the concentration of the acid

0.1 M acid makes a solution with  $[\text{H}^+] = 0.1\text{M}$

$$\text{pH} = -\log(0.1) = 1$$

What is the pH of a 0.5M solution of HBr?

A. 0.5

B. 1

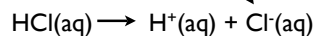
C. 0.3

D. 0

E. 12

$$[\text{H}^+] = 0.5 \quad 0 < \text{pH} < 1$$

We can ignore the “conjugate base” of a strong acid



$$K_a = \frac{[\text{H}^+][\text{Cl}^-]}{[\text{HCl}]} \approx \infty$$

equilibrium constant is so large,  
even if we add  $\text{Cl}^-$  the shift back to  
HCl will be negligible

“spectator ions”

If you know  $[\text{H}^+]$  you know  $[\text{OH}^-]$

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

$$\log(K_w) = \log([\text{H}^+][\text{OH}^-])$$

$$\log(K_w) = \log[\text{H}^+] + \log[\text{OH}^-]$$

$$\log(10^{-14}) = \log[\text{H}^+] + \log[\text{OH}^-]$$

$$-14 = -\text{pH} - \text{pOH}$$

$$14 = \text{pH} + \text{pOH}$$

### Weak Acid



	HA	H <sup>+</sup>	A <sup>-</sup>
I	C	0	0
C	-x	+x	+x
E	C-x	+x	+x

$$K_a = \frac{[\text{H}^{\text{+}}][\text{A}^{-}]}{[\text{HA}]} = \frac{(x)(x)}{C-x}$$

assuming  $x \ll C$   
 $x \sim \sqrt{K_a C}$

really 10<sup>-7</sup>

This is a great simple result

$$[\text{H}^{\text{+}}] \approx \sqrt{K_a C_a}$$

$C_a$  is the concentration of the acid  
 $K_a$  is the equilibrium constant for the acid

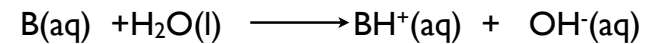
This assumes the concentration is large  
 and that  $K_a$  is small

What is the pH of a 1M solution of weak acid with a  $K_a = 10^{-6}$ ?

- A. 1
- B. 3
- C. 7
- D. 8
- E. 9

$$[\text{H}^{\text{+}}] = \sqrt{1 \times 10^{-6}} = 10^{-3}$$

### Weak Base

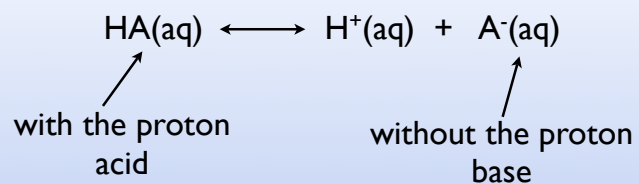


$$K_b = \frac{[\text{BH}^{\text{+}}][\text{OH}^{-}]}{[\text{B}]}$$

identical result as before (same assumptions)

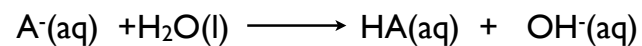
$$[\text{OH}^{-}] = \sqrt{K_b C_b}$$

### Weak Acids



HA weak acid  
A<sup>-</sup> weak base

### Weak Base

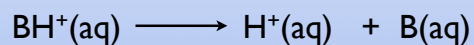
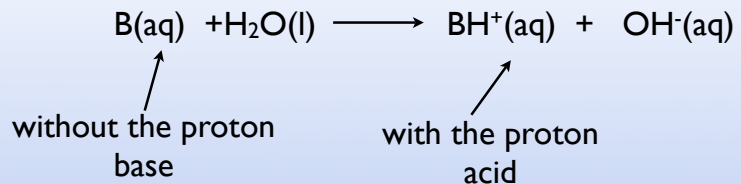


$$K_b = \frac{[\text{HA}][\text{OH}^-]}{[\text{A}^-]}$$

identical result as before (same assumptions)

$$[\text{OH}^-] = \sqrt{K_b C}$$

### Same with the base



### Weak acids

HA and BH<sup>+</sup>

Name is acid HA "acetic acid"

BH<sup>+</sup> has a positive charge and an "extra" proton NH<sub>4</sub><sup>+</sup>

### Weak bases

B and A<sup>-</sup>

A<sup>-</sup> is negative usually name ends in "ate" CH<sub>3</sub>COO<sup>-</sup> acetate

B is hardest to identify it is not one of the other three often it is an "amine"

What is the pH of a 1M solution of sodium benzoate?

- A. 4.9
- B. 5.1
- C. 6.2
- D. 7
- E. 9.09

benzoate is a weak base ( $A^-$ )  
only pH of 9.09 is basic

$$\Delta_R G^\circ(T) = -RT \ln K$$

$$\Delta T = iK_b m_{\text{solute}} \quad \ln \frac{K_2}{K_1} = \frac{\Delta_R H^\circ}{R} \left[ \frac{1}{T_2} - \frac{1}{T_1} \right]$$

$$\Delta T = -iK_f m_{\text{solute}}$$

$$\Delta P = X_{\text{solute}} P^\circ$$

$$\Pi = iMRT$$