Acid and Base Equilibrium

Chapter 8

Arrhenius

- Acid: A substance that make H^+ (H_3O^+) when dissolved in water.
- Base: A substance that makes OH^- when dissolved in water.
- An acid/base reaction occurs when and H^+ from an acid reacts with an OH^- from a base.

Acids

- Strong acids: Dissociate completely when dissolved in water.
  - HCl, HNO_3
- Weak acids only dissociate a little bit.
  - CH_3CO_2H

Base

- Strong base: dissociates completely when dissolved in water.
  - NaOH, KOH
- Weak base: Makes only a little bit of OH^-
  - -NH_3

1.00 M Acetic acid, a weak acid

$CH_3COOH_{(aq)} + H_2O_{(l)} \leftrightarrows H_3O^+_{(aq)} + CH_3COO^-_{(aq)}$

$$K = \frac{[H_3O^+][CH_3CO_2^-]}{[CH_3CO_2H]}$$

- The K_a for this reaction is 1.8 x 10^{-5}. Since little of the CH_3CO_2H dissociates, we can call it 1 M.
- For every H_3O^+ there will be one CH_3COO^-. Let these concentrations = x
- x=$[H^+]$ = $[CH_3COO^-]$ = .00042

Brønsted-Lowry Definition

- Acid: A proton donor.
- Base: A proton acceptor.
- An acid base reaction is one where there is a proton transfer.
- A broader definition than the Arrhenius definition.
- conjugate base: The acid without an H^+.
- conjugate acid: The base with an H^+.

Table 8.2

TABLE 8.2 Some Acids and Their Conjugate Bases, in Decreasing Order of Acid Strength

	Acid		Conjugate Base		
Strong Acids	HI	Hydroiodic acid	I^-	Iodide ion	Weak Bases
	H_2SO_4	Sulfuric acid	HSO_4^-	Hydrogen sulfate ion	
	HCl	Hydrochloric acid	Cl^-	Chloride ion	
	HNO_3	Nitric acid	NO_3^-	Nitrate ion	
	H_3O^+	Hydronium ion	H_2O	Water	
	HSO_4^-	Hydrogen sulfate ion	SO_4^{2-}	Sulfate ion	
	H_3PO_4	Phosphoric acid	$H_2PO_4^-$	Dihydrogen phosphate ion	
	CH_3COOH	Acetic acid	CH_3COO^-	Acetate ion	
	H_2CO_3	Carbonic acid	HCO_3^-	Bicarbonate ion	
	H_2S	Hydrogen sulfide	HS^-	Hydrogen sulfide ion	
	$H_2PO_4^-$	Dihydrogen phosphate ion	HPO_4^{2-}	Hydrogen phosphate ion	
	NH_4^+	Ammonium ion	NH_3	Ammonia	
	C_6H_5OH	Phenol	$C_6H_5O^-$	Phenoxide ion	
	HCO_3^-	Bicarbonate ion	CO_3^{2-}	Carbonate ion	
	HPO_4^{2-}	Hydrogen phosphate ion	PO_4^{3-}	Phosphate ion	
	H_2O	Water	OH^-	Hydroxide ion	
Weak Acids	C_2H_5OH	Ethanol	$C_2H_5O^-$	Ethoxide ion	
	NH_3	Ammonia	NH_2^-	Amide ion	Strong Bases

A standard acid/base reaction

Acid + Base ⇆ c. base + c. acid

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

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

More on Acid Base reactions

- Water is amphoteric or amphiprotic, it can act as both an acid and a base.
- All acid base reactions are equilibrium reactions
- The equilibrium lies to the side of the weaker acid.
- When the equilibrium lies to the right, a lot of reaction occurs and there is often heat released or there is a color change….
- When the equilibrium lies to the left very little reaction occurs. (no heat…)

Water

- Water auto-ionizes

$$\underset{\text{acid}}{H_2O} + \underset{\text{base}}{H_2O} \leftrightarrows \underset{[10^{-7}M]}{H_3O^+} + \underset{[10^{-7}M]}{OH^-}$$

$$K = \frac{[H_3O^+][OH^-]}{\cancel{[H_2O]^2}} = [H_3O^+][OH^-] = 1\times10^{-14}$$

Acids and pH

	0.1 M HCl	pure water	0.1 M NaOH
$[H^+]$ or $[H_3O^+]$			
$[OH^-]$			
pH			
pOH			
acid or basic			

The big six

1. $[H^+]\,[OH^-] = 10^{-14}$
2. $pH = -\log [H^+]$
3. $pOH = -\log [OH^-]$
4. $[H^+] = 10^{-pH}$
5. $[OH^-] = 10^{-pOH}$
6. $pH + pOH = 14$

The [H+] of 0.1 M NH_3

- The pH is 11.13. What is the [H+] ?
- pH =-log [H+]
- [H+]=10-pH
- [H+]=10-11.13

Filling out a table

	$[H^+]$	$[OH^-]$	pH	pOH
0.08 M HCl				
0.08 M Acetic Acid				

Buffer solution

A buffer solution keeps the pH approximately the same even upon the addition of a strong acid or strong base.

- Need a weak acid.
- Its conjugate base.
- Present in a large enough quantity to resist the pH changes.

$CH_3CO_2H/CH_3CO_2^-$

- If you add an acid, H+, The base of the buffer reacts.
  - $H^+ + CH_3CO_2^- \rightarrow CH_3CO_2H$
  - $H_3O^+ + CH_3CO_2^- \rightarrow H_2O + CH_3CO_2H$
- Of you add a base, OH-, the acid of the buffer system reacts.
  - $CH_3CO_2H + OH^- \rightarrow H_2O + CH_3CO_2^-$