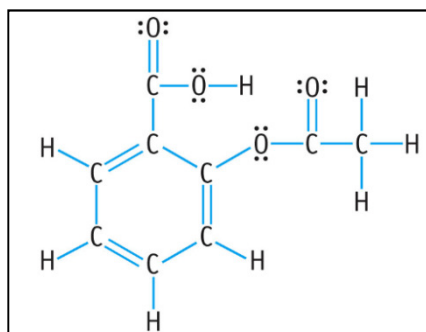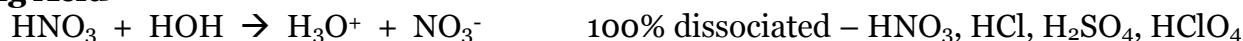These Notes are to <u>SUPPLIMENT</u> the Text, They do NOT Replace reading the Text Book Material. Additional material that is in the Text Book will be on your tests! To get the most information, <u>READ THE CHAPTER</u> prior to the Lecture, bring in <u>these lecture notes and make comments on these notes.</u> <u>These notes alone are NOT enough to pass any test!</u>
<u>NOTE THESE ARE DRAFT LECTURE NOTES!</u>

Asprin – Acetyl Salicylic Acid

Strong Acid

$HNO_3 + HOH \rightarrow H_3O^+ + NO_3^-$ 100% dissociated – HNO_3, HCl, H_2SO_4, $HClO_4$

Weak Acid

CH3COOH $\rightleftarrows$ H_3O^+ + CH3COO⁻ Acetic Acid or vinegar, less than 100% dissociated

Strong Base

NaOH $\rightarrow$ Na⁺ + OH⁻ NaOH, KOH, $Ca(OH)_2$

Weak Base

$NH_3 + H_2O \rightleftarrows NH_4^+ + OH^-$ Ammonia, less than 100% dissociated

Arrhenius Acid: any substance that dissolves in water and increases the concentration of H⁺

HCl $\rightarrow$ H⁺ + Cl⁻ Arrhenius Acid

Arrhenius Base: any substance that dissolves in water and increases the concentration of OH⁻
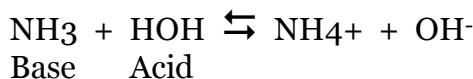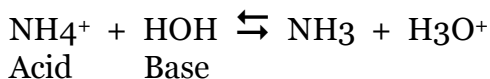
NaOH $\rightarrow$ Na⁺ + OH⁻ Arrhenius Base
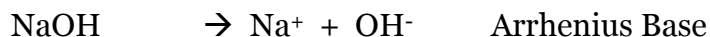
Reaction between and acid and a base involves H⁺ and OH⁻

NaOH + HCl $\rightarrow$ HOH + NaCl
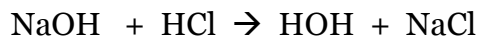
An acid and a base react for form water and a salt

Bronsted-Lowry acid is a Proton (H⁺) Donor

Bronsted-Lowry Base is a Proton (H⁺) Acceptor

HCl + HOH $\rightarrow$ H3O⁺ + Cl⁻ HNO3 + HOH $\rightleftarrows$ NO3⁻ + H3O⁺
Acid Base Hydronium Ion Acid Base

NH4⁺ + HOH $\rightleftarrows$ NH3 + H3O⁺ NH3 + HOH $\rightleftarrows$ NH4+ + OH⁻
Acid Base Base Acid

H_3PO_4 + HOH ⇆ HPO_4^{2-} + H3O+
Acid   Base

$H_2PO_4^-$ + HOH ⇆ HPO_4^{2-} + H3O+
Acid   Base

CO_3^{2-} + HOH ⇆ HCO_3^- + OH-
Which one is the Acid?

$[Fe(H_2O)_6]^{3+}$ + H_2O ⇆ $[Fe(H_2O)_5(OH)]^{2+}$ + H3O+
Which one is the Acid?
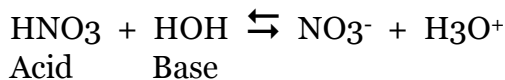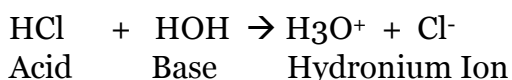
Monoprotic Acid are capable of donating one proton: HCl + HOH ⇆ H3O+ + Cl-

Polyprotic Acids can donate more than one proton: H_2S, H_3PO_4, H_2CO_3, H_2SO_4

H2SO4 + HOH ⇆ H3O+ + HSO_4^- 1st Proton removed

HSO_4^- + HOH ⇆ H3O+ + SO_4^- 2nd Proton removed

CO_3^{2-} + H_2O ⇆ HCO_3^- + OH- HCO_3^- + H_2O ⇆ H_2CO_3 + OH-
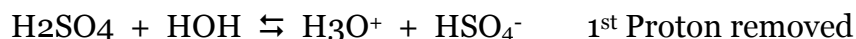
Base   Acid   Base   Acid

Amphiprotic(Amphoteric) Can behave as Bronsted Acid, donate H+, or Bronsted Base, accept H+

$H_2PO_4^-$ + H_2O ⇆ HPO_4^{2-} + H_3O^+ $H_2PO_4^-$ + H_2O ⇆ H_3PO_4 + OH-
Acid   Base
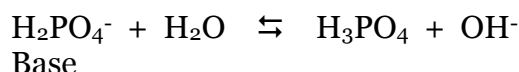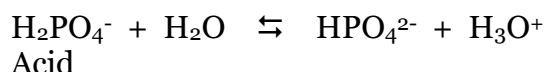
Table 17.1 Polyprotic Acids and Bases

Acid Form	Amphiprotic Form	Base Form
H_2S (hydrosulfuric acid or hydrogen sulfide)	HS^- (hydrogen sulfide ion)	S^{2-} (sulfide ion)
H_3PO_4 (phosphoric acid)	$H_2PO_4^-$ (dihydrogen phosphate ion) HPO_4^{2-} (hydrogen phosphate ion)	PO_4^{3-} (phosphate ion)
H_2CO_3 (carbonic acid)	HCO_3^- (hydrogen carbonate ion or bicarbonate ion)	CO_3^{2-} (carbonate ion)
$H_2C_2O_4$ (oxalic acid)	$HC_2O_4^-$ (hydrogen oxalate ion)	$C_2O_4^{2-}$ (oxalate ion)

Conjugate Acid-Base Pairs consist of 2 species that **differ from each other by the presence of one hydrogen ion.** Every reaction between a Bronsted Acid and Base involves a conjugate Acid-Base Pair.
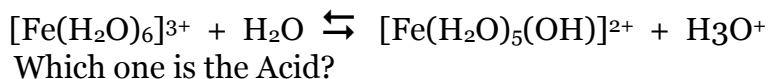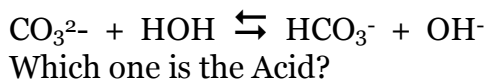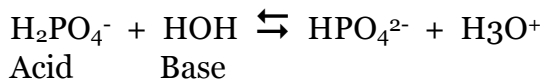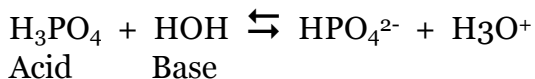

$$HCO_3^-(aq) + H_2O(\ell) \rightleftharpoons OH^-(aq) + H_2CO_3(aq)$$
base   acid   base   acid
conjugate pair 1
conjugate pair 2

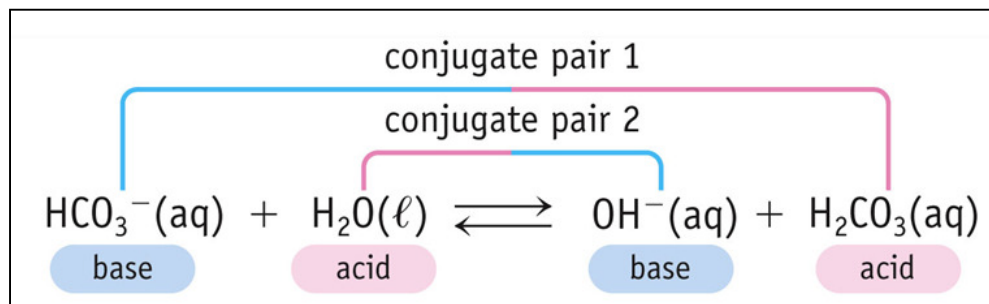**Table 17.2 Acid–Base Reactions and Conjugate Acid–Base Pairs***

Name	Acid 1		Base 2		Base 1		Acid 2
Hydrochloric acid	HCl	+	H_2O	⇌	Cl^-	+	H_3O^+
Nitric acid	HNO_3	+	H_2O	⇌	NO_3^-	+	H_3O^+
Carbonic acid	H_2CO_3	+	H_2O	⇌	HCO_3^-	+	H_3O^+
Acetic acid	CH_3CO_2H	+	H_2O	⇌	$CH_3CO_2^-$	+	H_3O^+
Hydrocyanic acid	HCN	+	H_2O	⇌	CN^-	+	H_3O^+
Hydrogen sulfide	H_2S	+	H_2O	⇌	HS^-	+	H_3O^+
Ammonia	H_2O	+	NH_3	⇌	OH^-	+	NH_4^+
Carbonate ion	H_2O	+	CO_3^{2-}	⇌	OH^-	+	HCO_3^-
Water	H_2O	+	H_2O	⇌	OH^-	+	H_3O^+

*Acid 1 and base 1 are a conjugate pair, as are base 2 and acid 2.

17.3 Water and the pH Scale

Autoionization & K_w $2\,H_2O \rightleftharpoons H_3O^+ + OH^-$ Even ultrapure water conducts electricity

Kw $= [H_3O^+]\,[OH^-] / [H_2O] = [H_3O^+]\,[OH^-] = 1.0 \times 10^{-14}$ at 25°C = Waters Auto Ionization Constant

In pure water: $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$

Since $[H_3O^+] = [OH^-]$, pure water is said to be neutral (not acidic or basic)

$[H_3O^+] > [OH^-]$ solution is acid $[H_3O^+] < [OH^-]$ solution is basic

Le Chatelier's Principle says that if you add acid to water the amount of base will decrease, visa versa

Example 17.1 What are the H_3O and OH^- concentrations in 0.0012 M NaOH at 25°C?

NaOH   →   Na⁺   +   OH⁻
0.0012 M   100% dissociation   0.0012M   0.0012M

$[H_3O^+]\,[OH^-] = 1.0 \times 10^{-14}$ $[H_3O^+] = 1.0 \times 10^{-14} / [OH^-] = 1.0 \times 10^{-14} / 1.2 \times 10^{-3} = 8.3 \times 10^{-12}$ M

The pH Scale: The pH of a solution is the negative of the base¹⁰ log of the $[H_3O^+]$

pH $= -\log ([H_3O^+])$	**pOH** $= -\log ([OH^-])$
Kw $= [H_3O^+]\,[OH^-] = 1.0 \times 10^{-14}$	or **pKw** $= 14 = pH + pOH$
$[OH^-] = Kw / [H_3O^+]$	pH $= 14 - pOH$
$[H_3O^+] = K2 / [OH^-]$	pOH $= 14 - pH$

	pH	$[H_3O^+]$	$[OH^-]$	pOH
Basic	14.00	1.0×10^{-14}	1.0×10^{0}	0.00
	10.00	1.0×10^{-10}	1.0×10^{-4}	4.00
Neutral	7.00	1.0×10^{-7}	1.0×10^{-7}	7.00
Acidic	4.00	1.0×10^{-4}	1.0×10^{-10}	10.00
	0.00	1.0×10^{0}	1.0×10^{-14}	14.00

Review Check – Students do calculations

1. What is the pH of 0.0012 M NaOH
2. The pH of diet soda is 4.32, what is the Hydronium ion concentration
3. The pH of a strong base is 10.46, what is the Hydronium and Hydroxide ions concentration

17.4 Equilibrium Constants – for weak acids and bases

Weak Acid $HA + H_2O \leftrightarrows H_3O^+ + A^-$ $Ka = [H_3O^+][A^-] / [HA]$

Weak Base $B + H_2O \leftrightarrows BH^+ + OH^-$ $Kb = [BH^+][OH^-] / [B]$
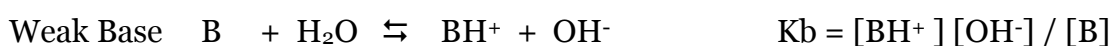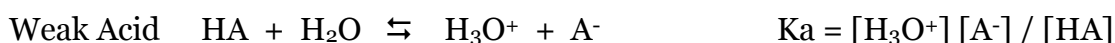
If K is large, then ionization is favored and molecule is a strong acid or strong base

The larger Ka, the smaller Kb   The larger Kb, the smaller Ka

SEE TABLE BELOW

Strongest acids are on left and upper   Strongest bases are lower right

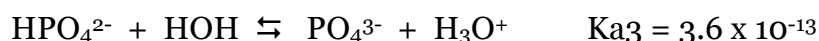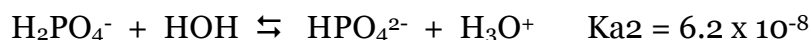Conjugate base is to its right   Conjugate acid to its left

Larger Ka, stronger acid   Larger Kb, stronger base

Table 17.3 Ionization Constants for Some Acids and Their Conjugate Bases at 25 °C

Acid Name	Acid	K_a	Base	K_b	Base Name
Perchloric acid	$HClO_4$	Large	ClO_4^-	Very small	Perchlorate ion
Sulfuric acid	H_2SO_4	Large	HSO_4^-	Very small	Hydrogen sulfate ion
Hydrochloric acid	HCl	Large	Cl^-	Very small	Chloride ion
Nitric acid	HNO_3	Large	NO_3^-	Very small	Nitrate ion
Hydronium ion	H_3O^+	1.0	H_2O	1.0×10^{-14}	Water
Sulfurous acid	H_2SO_3	1.2×10^{-2}	HSO_3^-	8.3×10^{-13}	Hydrogen sulfite ion
Hydrogen sulfate ion	HSO_4^-	1.2×10^{-2}	SO_4^{2-}	8.3×10^{-13}	Sulfate ion
Phosphoric acid	H_3PO_4	7.5×10^{-3}	$H_2PO_4^-$	1.3×10^{-12}	Dihydrogen phosphate ion
Hexaaquairon(III) ion	$[Fe(H_2O)_6]^{3+}$	6.3×10^{-3}	$[Fe(H_2O)_5OH]^{2+}$	1.6×10^{-12}	Pentaaquahydroxoiron(III) ion
Hydrofluoric acid	HF	7.2×10^{-4}	F^-	1.4×10^{-11}	Fluoride ion
Nitrous acid	HNO_2	4.5×10^{-4}	NO_2^-	2.2×10^{-11}	Nitrite ion
Formic acid	HCO_2H	1.8×10^{-4}	HCO_2^-	5.6×10^{-11}	Formate ion
Benzoic acid	$C_6H_5CO_2H$	6.3×10^{-5}	$C_6H_5CO_2^-$	1.6×10^{-10}	Benzoate ion
Acetic acid	CH_3CO_2H	1.8×10^{-5}	$CH_3CO_2^-$	5.6×10^{-10}	Acetate ion
Propanoic acid	$CH_3CH_2CO_2H$	1.3×10^{-5}	$CH_3CH_2CO_2^-$	7.7×10^{-10}	Propanoate ion
Hexaaquaaluminum ion	$[Al(H_2O)_6]^{3+}$	7.9×10^{-6}	$[Al(H_2O)_5OH]^{2+}$	1.3×10^{-9}	Pentaaquahydroxoaluminum ion
Carbonic acid	H_2CO_3	4.2×10^{-7}	HCO_3^-	2.4×10^{-8}	Hydrogen carbonate ion
Hexaaquacopper(II) ion	$[Cu(H_2O)_6]^{2+}$	1.6×10^{-7}	$[Cu(H_2O)_5OH]^+$	6.3×10^{-8}	Pentaaquahydroxocopper(II) ion
Hydrogen sulfide	H_2S	1×10^{-7}	HS^-	1×10^{-7}	Hydrogen sulfide ion
Dihydrogen phosphate ion	$H_2PO_4^-$	6.2×10^{-8}	HPO_4^{2-}	1.6×10^{-7}	Hydrogen phosphate ion
Hydrogen sulfite ion	HSO_3^-	6.2×10^{-8}	SO_3^{2-}	1.6×10^{-7}	Sulfite ion
Hypochlorous acid	$HClO$	3.5×10^{-8}	ClO^-	2.9×10^{-7}	Hypochlorite ion
Hexaaqualead(II) ion	$[Pb(H_2O)_6]^{2+}$	1.5×10^{-8}	$[Pb(H_2O)_5OH]^+$	6.7×10^{-7}	Pentaaquahydroxolead(II) ion
Hexaaquacobalt(II) ion	$[Co(H_2O)_6]^{2+}$	1.3×10^{-9}	$[Co(H_2O)_5OH]^+$	7.7×10^{-6}	Pentaaquahydroxocobalt(II) ion
Boric acid	$B(OH)_3(H_2O)$	7.3×10^{-10}	$B(OH)_4^-$	1.4×10^{-5}	Tetrahydroxoborate ion
Ammonium ion	NH_4^+	5.6×10^{-10}	NH_3	1.8×10^{-5}	Ammonia
Hydrocyanic acid	HCN	4.0×10^{-10}	CN^-	2.5×10^{-5}	Cyanide ion
Hexaaquairon(II) ion	$[Fe(H_2O)_6]^{2+}$	3.2×10^{-10}	$[Fe(H_2O)_5OH]^+$	3.1×10^{-5}	Pentaaquahydroxoiron(II) ion
Hydrogen carbonate ion	HCO_3^-	4.8×10^{-11}	CO_3^{2-}	2.1×10^{-4}	Carbonate ion
Hexaaquanickel(II) ion	$[Ni(H_2O)_6]^{2+}$	2.5×10^{-11}	$[Ni(H_2O)_5OH]^+$	4.0×10^{-4}	Pentaaquahydroxonickel(II) ion
Hydrogen phosphate ion	HPO_4^{2-}	3.6×10^{-13}	PO_4^{3-}	2.8×10^{-2}	Phosphate ion
Water	H_2O	1.0×10^{-14}	OH^-	1.0	Hydroxide ion
Hydrogen sulfide ion*	HS^-	1×10^{-19}	S^{2-}	1×10^5	Sulfide ion
Ethanol	C_2H_5OH	Very small	$C_2H_5O^-$	Large	Ethoxide ion
Ammonia	NH_3	Very small	NH_2^-	Large	Amide ion
Hydrogen	H_2	Very small	H^-	Large	Hydride ion

Increasing Acid Strength (left vertical axis)

Increasing Base Strength (right vertical axis)

*The values of K_a for HS^- and K_b for S^{2-} are estimates.

Ka for Polyprotic Acids

$$H_3PO_4 + HOH \rightleftharpoons H_2PO_4^- + H_3O^+ \qquad Ka_1 = 7.5 \times 10^{-3}$$

$$H_2PO_4^- + HOH \rightleftharpoons HPO_4^{2-} + H_3O^+ \qquad Ka_2 = 6.2 \times 10^{-8}$$

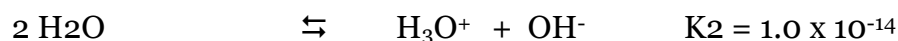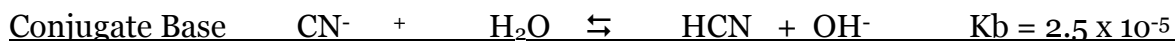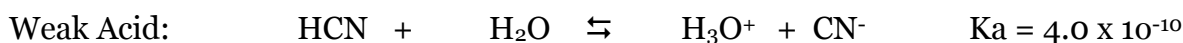$$HPO_4^{2-} + HOH \rightleftharpoons PO_4^{3-} + H_3O^+ \qquad Ka_3 = 3.6 \times 10^{-13}$$

For each ionization, Ka gets smaller so it's harder to take off the next Hydronium Ion

pKa = - log Ka   pKa can be used to report the strength of an acid

Acetic Acid has Ka = 1.8 x 10^{-5} pKa = - log (1.8 x 10^{-5}) = 4.74

Ionization constants for an Acid and it's Conjugate Base: **Ka * Kb = Kw**
This is used to solve for Ka if Kb is given

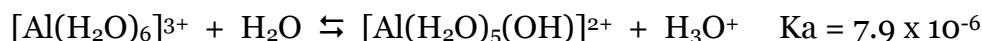| Weak Acid: | HCN + | H_2O ⇄ | H_3O^+ + CN^- | Ka = 4.0 x 10^{-10} |
| Conjugate Base | CN^- + | H_2O ⇄ | HCN + OH^- | Kb = 2.5 x 10^{-5} |
| | $2 H_2O$ | ⇄ | H_3O^+ + OH^- | K2 = 1.0 x 10^{-14} |

17.5 Acid-Base Properties of Salts

Anions can be a Bronsted base because they can accept a proton

CO_3^{2-} (anion) + H_2O ⇄ HCO_3^- + OH^- Kb = 2.1 x 10^{-4}

Metal cations (M^{2+} or M^{3+}, are hydrated in water) can be Bronsted acids, they donate protons

$[Al(H_2O)_6]^{3+}$ + H_2O ⇄ $[Al(H_2O)_5(OH)]^{2+}$ + H_3O^+ Ka = 7.9 x 10^{-6}

Anions that are a conjugate base of a strong acid are weak bases and do not effect pH, Cl^-, NO_3^-

HCl (strong acid) → H^+ + Cl^- Cl^- + HOH → NR
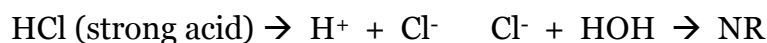
Anions that are a conjugate base of a weak acid will raise the pH of a solution

H_2CO_3 → CO_3^{2-} + $2 H^+$ CO_3^{2-} + HOH ⇄ HCO_3^- + OH^- Hydroxide is a base

Alkali metal and alkaline earth cations have no measureable effect on solution pH

Na^+ + HOH → NR

Acidic Cations (metal cations with 2+ or 3+ charge) and ammonium ions (and organic derivitives)

area weak acids $[Al(H_2O)_6]^{3+}$ + H_2O ⇄ $[Al(H_2O)_5(OH)]^{2+}$ + H_3O^+ Ka = 7.9 x 10^{-6}
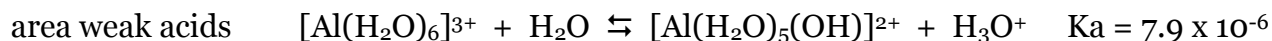
Polyprotic acids – acid-base behavior depends on the amount of deprotonation:

H_2CO_3 is acidic, CO_3^{2-} is basic HCO_3^- is amphiprotic

H_2CO_3 + H_2O ⇄ HCO_3^- + H_3O^+ HCO_3^- + H_2O ⇄ CO_3^{-2} + H_3O^+
Acid →   ← Base   Acid →   ← Base

Table 17.4 Acid and Base Properties of Some Ions in Aqueous Solution

	Neutral		Basic			Acidic	
Anions	Cl^-	NO_3^-	$CH_3CO_2^-$	CN^-	SO_4^{2-}	HSO_4^-	
	Br^-	ClO_4^-	HCO_2^-	PO_4^{3-}	HPO_4^{2-}	$H_2PO_4^-$	
	I^-		CO_3^{2-}	HCO_3^-	SO_3^{2-}	HSO_3^-	
			S^{2-}	HS^-	OCl^-		
			F^-	NO_2^-			
Cations	Li^+		$[Al(H_2O)_5(OH)]^{2+}$ (for example)			$[Al(H_2O)_6]^{3+}$ and hydrated	
	Na^+	Ca^{2+}				transition metal cations	
	K^+	Ba^{2+}				(such as $[Fe(H_2O)_6]^{3+}$)	
						NH_4^+	

Interactive Example 17.2   Are these acid, basic or neutral water solutions

1. $NaNO_3$ Neutral Na^+ and NO_3^- are very weak conjugate base of a strong acid

2. K_3PO_4 Basic PO_4^{3-} is the conjugate base of a weak acid, K^+ and Na^+ are N/A
                                      $PO_4^{3-} + H_2O \leftrightharpoons HPO_4^{2-} + OH^-$ Goes to the Right

3. $FeCl_2$ Weak Acid Fe^{2+} in water gives $[Fe(H_2O)_6]^{2+}$ which is a Bronsted Acid
                                      Cl^- is a very weak conjugate base of the strong acid HCl

4. $NaHCO_3$ Amphiprotic See above **Polyprotic acids**

Increasing Acid Strength →

Acid	HCO_3^-	$HClO$	HF
K_a	4.8×10^{-11}	3.5×10^{-8}	7.2×10^{-4}

← Increasing Base Strength

Base	CO_3^{2-}	ClO^-	F^-
K_b	2.1×10^{-4}	2.9×10^{-7}	1.4×10^{-11}

17.6 Predicting the direction of Acid-Base Reactions

Acid + Base $\leftrightharpoons$ Conjugate base of the acid + Conjugate acid of the base

**Reaction/Equilibrium Direction is towards the WEAKER Acid and Base**

Predicting the Direction of Acid–Base Reactions

- According to the Brønsted–Lowry theory, all acid–base reactions can be written as equilibria involving the acid and base and their conjugates.

Acid + Base ⮂ Conjugate base of the acid + Conjugate acid of the base

conjugate pair 1
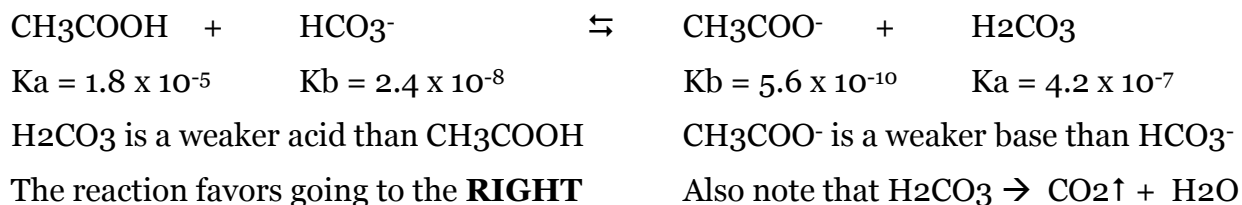
conjugate pair 2

$HCl(aq)$ + $H_2O(\ell)$ $\underset{K > 1}{\rightleftharpoons}$ $H_3O^+(aq)$ + $Cl^-(aq)$

stronger acid than H_3O^+ | stronger base than Cl^- | weaker acid than HCl | weaker base than H_2O

- All proton transfer reactions proceed from the stronger acid and base to the weaker acid and base.

$CH_3COOH + HOH ⇌ CH_3COO^- + H_3O^+$ $K = 1.8 \times 10^{-5}$ **What's its pH?**
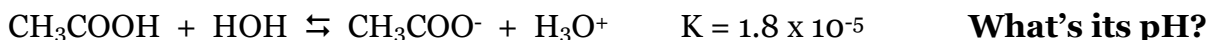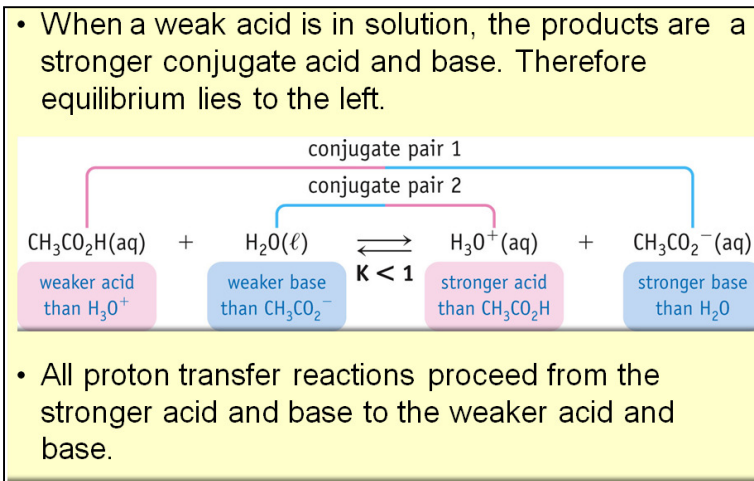
In 0.01 M acetic acid, ionization is to about 0.001 M or 99% of the acid is NOT ionized!

Example 17.3 Write balanced equation for reaction between acetic acid and sodium bicarbonate and state which way the reaction will go?

$CH3COOH$ + $HCO3^-$ ⇌ $CH3COO^-$ + $H2CO3$

$Ka = 1.8 \times 10^{-5}$ $Kb = 2.4 \times 10^{-8}$ $Kb = 5.6 \times 10^{-10}$ $Ka = 4.2 \times 10^{-7}$

$H2CO3$ is a weaker acid than $CH3COOH$ $CH3COO^-$ is a weaker base than $HCO3^-$

The reaction favors going to the **RIGHT**   Also note that $H2CO3 \rightarrow CO2\uparrow + H2O$

Student Question How will this reaction go? $H_3PO_4 + CH_3COO^- ⇌ H_2PO_3^- + CH_3COOH$
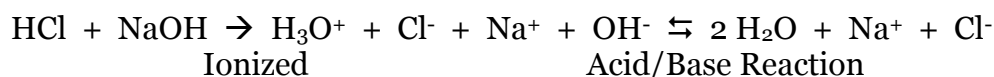
17.7 Acid-Base Reactions

Table 17.5 Characteristics of Acid–Base Reactions

Type	Example	Net Ionic Equation	Species Present after Equal Amounts Are Mixed; pH
Strong acid + strong base	HCl + NaOH	$H_3O^+(aq) + OH^-(aq) ⇌ 2 H_2O(\ell)$	Cl^-, Na^+, pH = 7
Strong acid + weak base	HCl + NH_3	$H_3O^+(aq) + NH_3(aq) ⇌ NH_4^+(aq) + H_2O(\ell)$	Cl^-, NH_4^+, pH < 7
Weak acid + strong base	HCO_2H + NaOH	$HCO_2H(aq) + OH^-(aq) ⇌ HCO_2^-(aq) + H_2O(\ell)$	HCO_2^-, Na^+, pH > 7
Weak acid + weak base	HCO_2H + NH_3	$HCO_2H(aq) + NH_3(aq) ⇌ HCO_2^-(aq) + NH_4^+(aq)$	HCO_2^-, NH_4^+, pH dependent on K_a and K_b of conjugate acid and base

Strong Acid and Strong Base   Strong = 100% ionized in solution

$HCl + NaOH \rightarrow H_3O^+ + Cl^- + Na^+ + OH^- ⇌ 2 H_2O + Na^+ + Cl^-$
              Ionized   Acid/Base Reaction

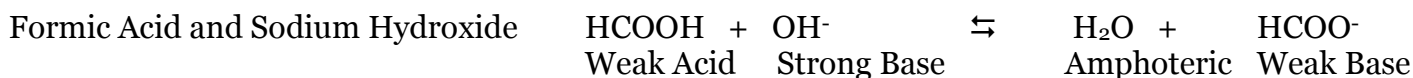The Net Ionic Equation $H_3O^+ + OH^- ⇌ 2 H_2O$ $K = 1 / [H_3O^+][OH^-] = 1 / Kw = 1.0 \times 10^{14}$

H_3O^+ (Ka = 1.0) and OH^- (Kb = 1.0) – data from acid strength table
Reactions of a strong acid and a strong base are called a **Neutralization**
*Mixing equal molar amounts of a strong acid and strong base results in a **neutral pH around 7.0***

Weak Acid and Strong Base   Weak Acid is not fully ionized

Formic Acid and Sodium Hydroxide $HCOOH$ + OH^- $\leftrightarrows$ H_2O + $HCOO^-$
                                      Weak Acid   Strong Base   Amphoteric   Weak Base

NaOH (Kb = 1.0) is a stronger base than $HCOO^-$ (Kb = 5.6 x 10^{-11}), Reaction goes to the **RIGHT**
*Mixing equal molar amounts of a strong base and a weak acid gives a salt whose anion is the conjugate base of a weak acid. The solution is **Basic***

**Strong Acid and Weak Base**   Weak Base is not fully ionized

Hydrochloric acid and Ammonia H_3O^+ + NH_3 $\leftrightarrows$ H_2O + NH_4^+

Hydronium Ion (H_3O^+) (Ka = 1.0) is stronger acid than NH_4^+ (Ka = 5.6 x 10^{-10})

NH_3 (Kb = 1.8 x 10^{-5}) is a strong base than H_2O (Kb = 1.0 x 10^{-14}), Reaction goes to the **RIGHT**
*Mixing equal molar amounts of a strong acid and a weak base gives a salt whose cation is the conjugate acid of the weak base. The solution is **Acidic***

Reaction of a Weak Acid and a Weak Base   Neither fully ionized

Acetic acid and ammonia CH_3COOH + NH_3 $\leftrightarrows$ NH_4^+ + CH_3COO^-
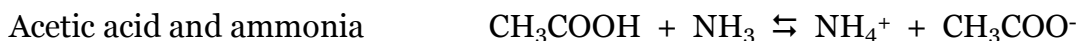
CH_3COOH (Ka = 1.8 x 10^{-5}) is a stronger acid than NH_4^+ (Ka = 5.6 x 10^{-10})

NH3 (Kb = 1.8 x 10^{-5}) is a stronger base than CH_3COO^- (Ka = 5.6 x 10^{-10})

Equilibrium is on the side of the weaker acid / base = to the **RIGHT**
Acetic acid and ammonia yields Ammonium Acetate.
Is the solution Acidic or basic? NH_4^+ (Ka = 5.6 x 10^{-10}) and CH_3COO^- (Ka = 5.6 x 10^{-10})
The solution will be neutral
Mixing equal molar amounts o a weak acid and weak base yields a salt, the **pH depends** on the relative Ka / Kb values.

17.8 Calculations and Equilibrium Constants.

Determine K from initial concentration and pH.
Interactive Example 17.4 0.10 M Lactic Acid CH3-CHOH-COOH has a pH of 2.43. What is Ka?

$CH3-CHOH-COOH$ + $H2O$ $\leftrightarrows$ $CH3-CHOH-COO^-$ + H_3O^+
Lactic Acid   Lactate Ion
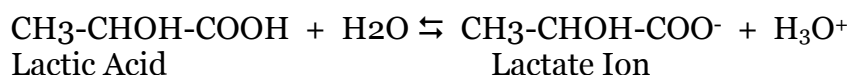
From pH = 2.43, $[H_3O^+]$ = $10^{-2.43}$ = 3.7 x 10^{-3} M

	$CH3-CHOH-COOH$ + $H2O$ $\leftrightarrows$	$CH3-CHOH-COO^-$ +	H_3O^+
Initial	0.10	0.0	0.0
Change	-x	+x	+x
Equilibrium	0.10 – x	x	x

Ka   = $[H_3O^+]$ $[CH3-CHOH-COO^-]$ / $[CH3-CHOH-COOH]$

     =   x * x / (0.10 –x)

At equilibrium:
    $[CH3-CHOH-COO^-]$ = $[H_3O^+]$ = 3.7 x 10^{-3} M

    x = the amount of Lactic Acid ionized, so the [Lactic Acid] = 0.10 M - 3.7 x 10^{-3} M

    Ka = 3.7 x 10^{-3} * 3.7 x 10^{-3} / (0.10 - 3.7 x 10^{-3}) = 1.4 x 10^{-4}

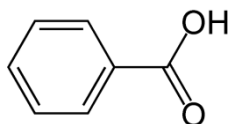**Note** that $(0.10 - 3.7 \times 10^{-3}) = $ approx $= 0.10$. Therefore (see discussion in book) for a weak acid:

$[HA]_{equilibrium} = [HA]_0 - [H_3O^+] \approx [HA]_0$ when $[HA]_0$ is $>= 100 * K_a$

pH of a weak acid or base

Interactive Example 17.5 **?** is pH of 0.020 M benzoic acid (C_6H_5COOH), $K_a = 6.3 \times 10^{-5}$

	C_6H_5COOH +	HOH $\leftrightarrows$	$C_6H_5COO^-$ +	H_3O^+
Initial	0.020		0	0
Change	-x		+x	+x
Equilibrium	(0.020 – x)		x	x

$K_a = [C_6H_5COO^-][H_3O^+] / [C_6H_5COOH] = x * x / (0.20-x)$

Since K_a is so small, $(0.020 - x) \approx 0.020$

$K_a = 6.3 \times 10^{-5} = x^2 / 0.020$

$6.3 \times 10^{-5} * 0.020 = x^2$ Now take Square Root of both sides

$x = 1.1 \times 10^{-3} = 0.0011$ M

$C_6H_5COOH = (0.020 - x) = (0.020 - 0.0011) = 0.019$ M

$[H_3O^+] = x = 1.1 \times 10^{-3}$ M

pH $= - \log([H_3O^+]) = - \log(1.1 \times 10^{-3}) = 2.96$

Interactive Example 17.7 **?** is the pH of 0.015 M Sodium Acetate?

Sodium Acetate is basic because acetate is the conjugate base of a weak acid, acetic acid, which reacts with water to form OH^-. $\{CH3COO^- + H_2O \leftrightarrows CH_3COOH + OH^-\}$ ***also show short cut!***

	$CH3COO^-$ + H_2O $\leftrightarrows$	CH_3COOH +	OH^-
Initial	0.015	0	0
Change	-x	-x	-x
Equilibrium	(0.015 – x)	x	x

$K_b = [CH_3COOH][OH^-] / [CH3COO^-] = 5.6 \times 10^{-10} = x * x / (0.015 - x)$

Acetate is a weak base as K_a is small, so x is very small, so $(0.015 - x) \approx 0.015$

$5.6 \times 10^{-10} = x * x / 0.015$ $x^2 = 0.015 * 5.6 \times 10^{-10}$ **x = 2.9 \times 10^{-6} M**

$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$ $[H_3O^+] = 1.0 \times 10^{-14}$ M $/ 2.9 \times 10^{-6}$ M $= 3.5 \times 10^{-9}$ M
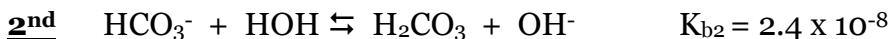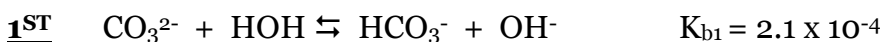
pH $= - \log([H_3O^+]) = - \log(3.5 \times 10^{-9}) = 8.46$

17.9 Polyprotic Acid and Base

The loss of successive protons is 10^4 to 10^6 smaller – harder to pull off the 2nd or 3rd proton!

Therefore, the pH is depends on the loss of the 1st proton

Example 17.9 Calculate the pH of 0.10 M Sodium Carbonate

1ST CO_3^{2-} + HOH $\leftrightarrows$ HCO_3^- + OH^- K_{b1} = 2.1 x 10^{-4}

2nd HCO_3^- + HOH $\leftrightarrows$ H_2CO_3 + OH^- K_{b2} = 2.4 x 10^{-8}

1ST	CO_3^{2-} + HOH	$\leftrightarrows$	HCO_3^- +	OH^-
Initial	0.10		0	0
Change	-x		+x	+x
Equilibrium	0.10 – x		x	x

K_b = 2.1 x 10^{-4} = [OH^-] [HCO_3^-] / [CO_3^{2-}] = x * x / (0.10 – x)

Since K_b is small, x is small and (0.10 – x)) $\approx$ 0.10

2.1 x 10^{-4} = x^2 / 0.10 x^2 = 2.1 x 10^{-4} * 0.10 = 2.1 x 10^{-5} **x = 4.6 x 10^{-3}**

pOH = - log (4.6 x 10^{-3}) = 2.34   **pH** = 14.00 – pOH = 14.00 – 2.34 = **11.66**

2nd	HCO_3^- + HOH	$\leftrightarrows$	H_2CO_3 +	OH^-
Initial	x = 4.6 x 10^{-3}		0	4.6 x 10^{-3}
Change	-y		+y	+y
Equilibrium	4.6 x 10^{-3}–y		y	y + 4.6 x 10^{-3}

K_{b2} = 2.4 x 10^{-8} = [H_2CO_3] [OH^-] / [HCO_3^-] = y (4.6 x 10^{-3}–y) / (y + 4.6 x 10^{-3})

Since K_{b2} is very small, smaller than K_b, the amount of product produced in 2nd step is very small

Therefore [HCO_3^-] = (4.6 x 10^{-3}–y) and [OH^-] = (y + 4.6 x 10^{-3}) is close to 4.6 x 10^{-3}

K_{b2} = 2.4 x 10^{-8} = y * 4.6 x 10^{-3} / 4.6 x 10^{-3}

y = (2.4 x 10^{-8}) * (4.6 x 10^{-3}) / 4.6 x 10^{-3} = 2.4 x 10^{-8} M

[OH^-] from the first step = 4.6 x 10^{-3} M, from the second step 2.4 x 10^{-8} M = No Effect!!

Acid Strength of Hydrogen Halides: HI > HBr > HCl >> HF

OxoAcids – acids that contain more than one Oxygen Atom. Acid strength increases with an increase in the number of Oxygen's.   HNO3 is a stronger acid than HNO2.

HOCl > HOClO > HOClO2 > HOClO4   H2SO4 > H2SO3

CH3-COOH (Acetic Acid) < ClCH2-COOH < Cl2CH-COOH < Cl3C-COOH
The extra chlorine stabilizes the negative charge of the acid ion (R-COO⁻)

Carboxylic Acids (R-COOH) are Bronsted Acids   R-COOH + H2O $\leftrightarrows$ R-COO⁻ + H3O⁺
        Formic Acid H-COOH   Acetic Acid CH3-COOH

Hydrated Metal Cations are Bronsted Acids:
        [$Cu(H_2O)_6$]$^{2+}$ + H_2O $\leftrightarrows$ [$Cu(H_2O)_5(OH)^-$]$^{2+}$ + H_3O^+
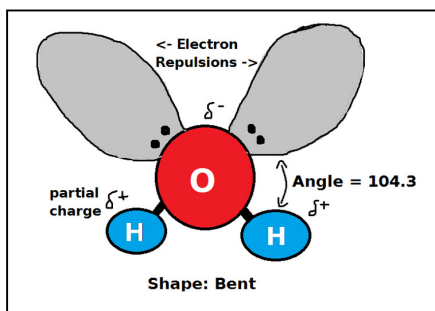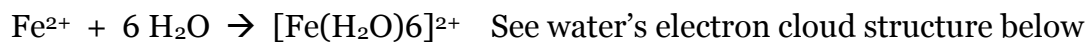
Lewis Acid and Base   Used for non-aqueous solutions!
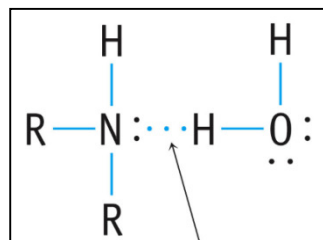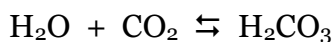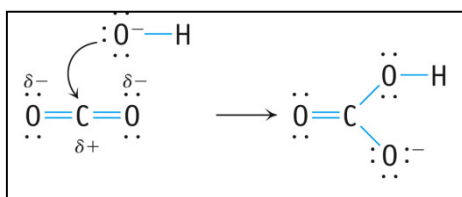Lewis Acid can accept a pair of electrons.   Lewis Base donates a pair of electrons

        A (acid) + B: (base) $\leftrightarrows$ A:B or A->B   Also called a coordinate covalent bond

        Ammonia H_3N: is a good example of a Lewis Base

Metals and water from Complex Ions also called Coordinate Complexes

$Fe^{2+} + 6\ H_2O \rightarrow [Fe(H_2O)6]^{2+}$ See water's electron cloud structure below

<- Electron
Repulsions ->

δ^-

O

partial charge δ^+

Angle = 104.3
δ^+

H   H

Shape: Bent

Picture from Wikimedia commons

Water pictured above. Assume the Hydrogen's are in the field of this piece of paper. The 2 pair of electrons form orbits going in and out of the plane of the paper.

$$:O^- — H$$

$$\delta^- \qquad \delta^-$$
$$O=C=O \qquad \longrightarrow \qquad O=C$$
$$\delta^+$$
$$O — H$$
$$:O:^-$$

$H_2O + CO_2 \leftrightarrows H_2CO_3$

H   H
| |
R—N ⋯ H—O:
|
R

Amine (Nitrogen containing) will hydrogen bond with water

Amphoteric compounds can behave as an acid or a base:

$Al(OH)_3 \downarrow + OH^- \rightarrow [Al(OH)_4]^-$ $Al(OH)_3 \downarrow + 3\ H_3O+ \rightarrow Al^{3+} + 6\ H_2O$
Acid   base   Base   Acid