Strong Acids and Bases________________ Z Ch 7.1-7.4

END OF WEEK SWITCHING TO HARRIS TEXT

FRIDAY QUIZ ON CHEMICAL EQUILIBRIUM – some M ACID/BASE

7.1 The Nature of Acids and Bases
7.2 Acid Strength
7.3 The pH Scale
7.4 Calculating the pH of Strong Acid Solutions
7.6 Bases (just the strong ones)

REVIEW FROM FRIDAY

Le Châtelier's Principle – Change *P* (*T* Constant)

Law of Mass Action

For $aA + bB \Leftrightarrow cC + dD$ the equilibrium constant *K* is

$$K = \frac{P^c_C P^d_D}{P^a_A P^b_B}$$

2) add inert gas (one that does not participate in the chemical equilibrium

Since $P_i = n_i RT/V$, the partial pressures of all gases participating in the equilibrium reaction are unaffected by the presence of the inert gas. Hence so is the equilibrium constant.

REVIEW FROM FRIDAY

Le Châtelier's Principle – Change *P* (*T* Constant)

Law of Mass Action

For aA + bB <=> cC + dD the equilibrium constant *K* is

$$K = \frac{P^c_{\mathrm{C}} P^d_{\mathrm{D}}}{P^a_{\mathrm{A}} P^b_{\mathrm{B}}}$$

3) add gaseous reactant or product

Since $P_i = n_i RT/V$, partial pressures of all other gases participating in the equilibrium reaction are unaffected. But *K* would change and it is a constant! Therefore if added gas were a reactant, *K* would decrease unless equilibrium shifts to products => equilibrium must shift to products. Similarly, were the added gas a product, the equilibrium would shift toward reactants.

Le Châtelier's Principle – Change *P* (*T* Constant)

Law of Mass Action

For aA + bB <=> cC + dD the equilibrium constant *K* is

$$K = \frac{P^c_C P^d_D}{P^a_A P^b_B}$$

4) decrease volume by compression (consider 4-fold decrease)

Since $P_i = x_i P_{TOT}$, each P_i in *K* (actually *Q* until new equilibrium established) increases 4-fold. If total # of moles of reactant gases in balanced chemical equation same as total # of moles of product gases then no effect - powers that P_i are raised to in numerator of *K* same as in denominator. Otherwise, a *P* increase shifts equilibrium to side with fewer moles in balanced equation.

REVIEW FROM FRIDAY

Le Châtelier's Principle – Change *T* (*P* Constant)

Exothermic reactions gives off heat, endothermic reactions require heat

a) the following reaction is exothermic

$$N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g) \qquad + \text{heat}$$

b) the following reaction is endothermic

$$+ \text{heat} \qquad N_2(g) + O_2(g) \rightleftharpoons 2\,NO(g)$$

analogy with the stress imposed by adding or removing a gaseous reactant if one were to consider heat as a "reactant" in the case of an endothermic reaction and as a "product" for an exothermic reaction.

REVIEW FROM FRIDAY

Nature of Acids and Bases

	Lewis	Brønsted-Lowry	Arrhenius
acid:	accepts electron pair	donates H^+	produces H^+
base:	donates electron pair	accepts H^+	produces OH^-

Lewis

coordinate covalent bond

(CHEM 118)

Brønsted-Lowry

conjugate acid-base pairs

$$A_1 + B_2 \rightarrow A_2 + B_1$$

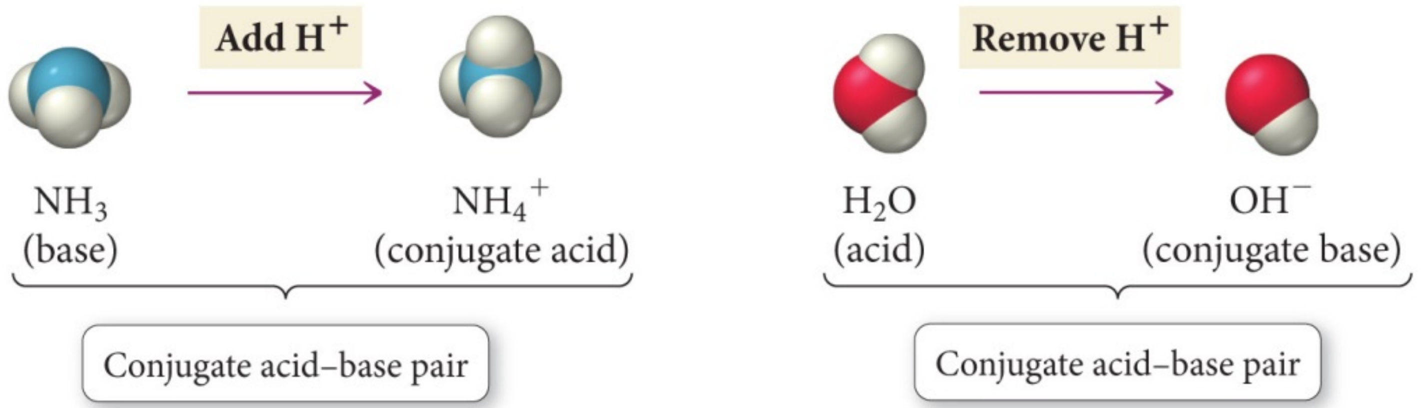
Arrhenius

neutralization

$$H^+ + OH^- \rightarrow H_2O$$

	Lewis	Brønsted-Lowry	Arrhenius
problem:	most general	must have H	need aqueous solution

Brønsted-Lowry: Conjugate Acid-Base Pairs

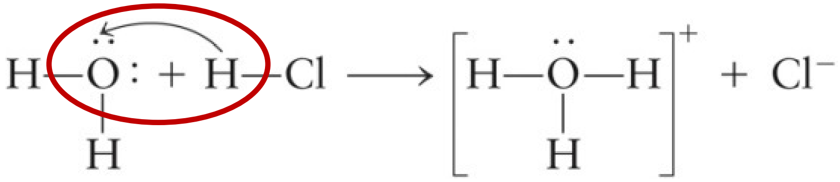


EX 1. For each of the following write the formula of its conjugate.

ACIDS		BASES	
HCl	Cl^-	NH_3	NH_4^+
H_2O	OH^-	H_2S	H_3S^+
NH_3	NH_2^-	$NaOH$	H_2O
CH_3COOH	CH_3COO^-	O^{2-}	OH^-
SH^-	S^{2-}	CN^-	HCN

Brønsted-Lowry: Acid-Base Pair Chemistry

acids – proton donors => form a species [acid – H^+] called **conjugate base**

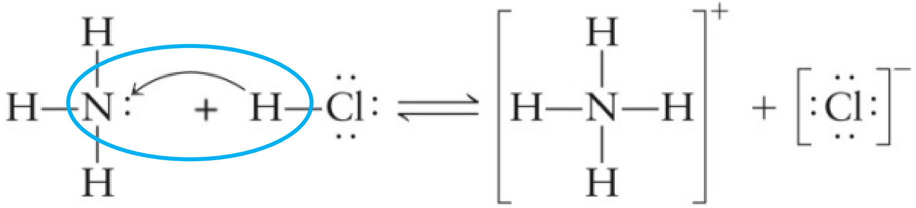


strong acid or base often uses reaction arrow

$$\underset{\text{Acid}}{H_2SO_4(aq)} + \underset{\text{Base}}{H_2O(l)} \longrightarrow \underset{\text{Conjugate base}}{HSO_4^-(aq)} + \underset{\text{Conjugate acid}}{H_3O^+(aq)}$$

reaction when acid dissolved in water

bases – proton acceptors => form a species [base + H^+] called **conjugate acid**



nonaqueous reaction in liquid ammonia

weak acid or base uses equilibrium arrow

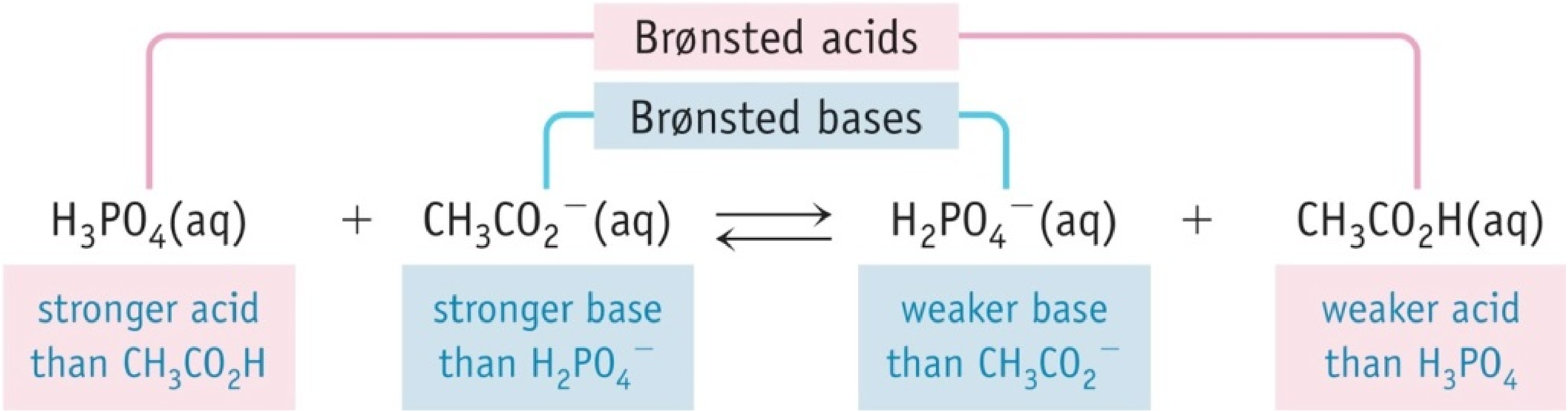
$$\underset{\text{Base}}{HCO_3^-(aq)} + \underset{\text{Acid}}{H_2O(l)} \rightleftharpoons \underset{\text{Conjugate acid}}{H_2CO_3(aq)} + \underset{\text{Conjugate base}}{OH^-(aq)}$$

reaction when base dissolved in water

REVIEW FROM FRIDAY

Brønsted-Lowry: Acid-Base Pair Chemistry

two acids competing to give up H^+ – the stronger acid "wins"



two bases competing for the acidic proton – the stronger base "wins"

EX 3. From the Brønsted-Lowry point of view, which is the stronger acid in the following reaction:

$$H_2SO_4(aq) + HNO_3(aq) <=> HSO_4^-(aq) + H_2NO_3^+(aq)$$

Amphoteric Nature of Water

Water acting as a base

$$\underbrace{HCl(aq)}_{\text{Acid (proton donor)}} + \underbrace{H_2O(l)}_{\text{Base (proton acceptor)}} \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

Water acting as an acid

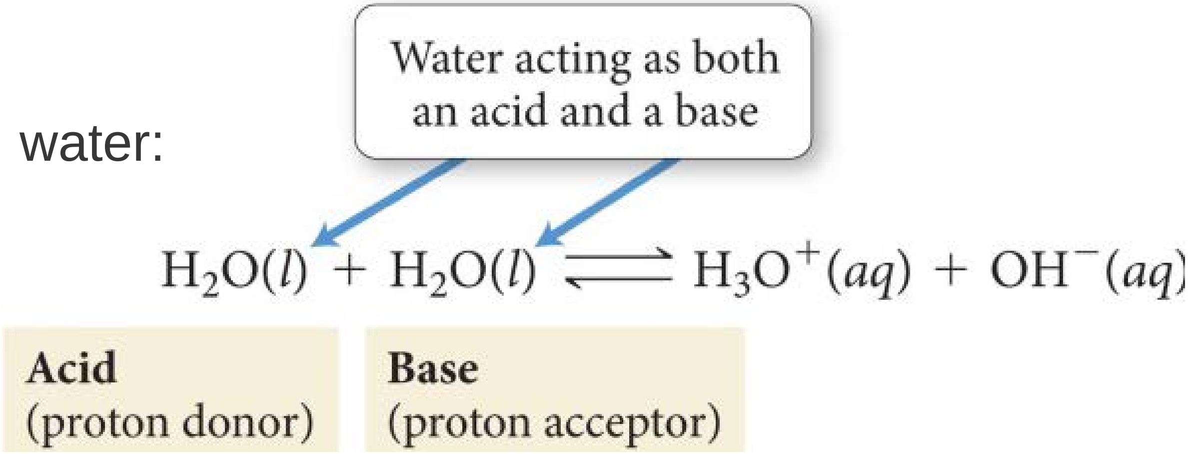
$$\underbrace{NH_3(aq)}_{\text{Base (proton acceptor)}} + \underbrace{H_2O(l)}_{\text{Acid (proton donor)}} \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

The pH scale in water depends upon this

Autoionization

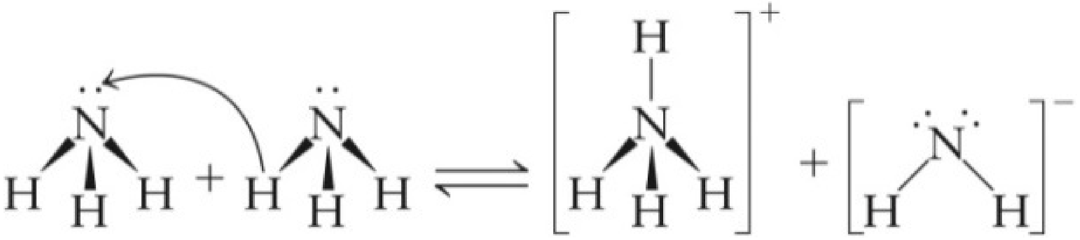
A self-ionization process which depends upon the amphoteric nature of the solvent.. It is exactly this process which defines what acidity and basicity in a particular solvent is (via the autoionization reaction of the solvent).

of water:



of ammonia:

$NH_3(l) + NH_3(l)$ <=> $NH_4^+(aq) + NH_2^-(aq)$



base acid conjugate acid conjugate base

Strong Acids and Bases to Know

seven strong acids to know	
hydrochloric acid	HCl
hydrobromic acid	HBr
hydroiodic acid	HI
perchloric acid	$HClO_4$
chloric acid	$HClO_3$
sulfuric acid	H_2SO_4
nitric acid	HNO_3

soluble strong bases to know	
lithium hydroxide	$LiOH$
sodium hydroxide	$NaOH$
potassium hydroxide	KOH
rubidium hydroxide	$RbOH$
cesium hydroxide	$CsOH$
barium hydroxide	$Ba(OH)_2$

strong bases - all Group I and Group II hydroxides except Be

Acid Strength

HA – generic way of writing a monoprotic acid (one acidic hydrogen)

acid strength – determined by extent of reaction of acid with water to form $H_3O^+(aq)$, or the extent of its ionization or dissociation, as shown by the magnitude of its equilibrium constant, K_a – then for any hydrogen-containing compound, HA

$$HA(aq) \quad \boxed{+ \quad H_2O(l) \quad \rightleftharpoons \quad H_3O^+(aq) \quad +} \quad A^-(aq)$$

EQUATION FOR ACIDITY FOR ANY HA

EQUILIBRIUM CONSTANT

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

Various Ways to Describe Acid Strength

Property	Strong Acid	Weak Acid
K_a value	K_a is large	K_a is small
Position of the dissociation equilibrium	Far to the right	Far to the left
Equilibrium concentration of H^+ compared with original concentration of HA	$[H^+] \approx [HA]_0$	$[H^+] \ll [HA]_0$
Strength of conjugate base compared with that of water	A^- much weaker base than H_2O	A^- much stronger base than H_2O

Base Strength

B: – generic way of writing a monobasic base (one basic site)

base strength – (aside from Group I and II hydroxides) determined by extent of reaction of base with water to form $OH^-(aq)$, or extent of its ionization, as shown by the magnitude of its equilibrium constant, K_b – then for any base B:

$$B{:}(aq) \quad \boxed{+ \quad H_2O(l) \quad \rightleftharpoons \quad OH^-(aq) \quad +} \quad BH^+(aq)$$

EQUATION FOR BASICITY FOR ANY B:

EQUILIBRIUM CONSTANT

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

Various Ways to Describe Base Strength

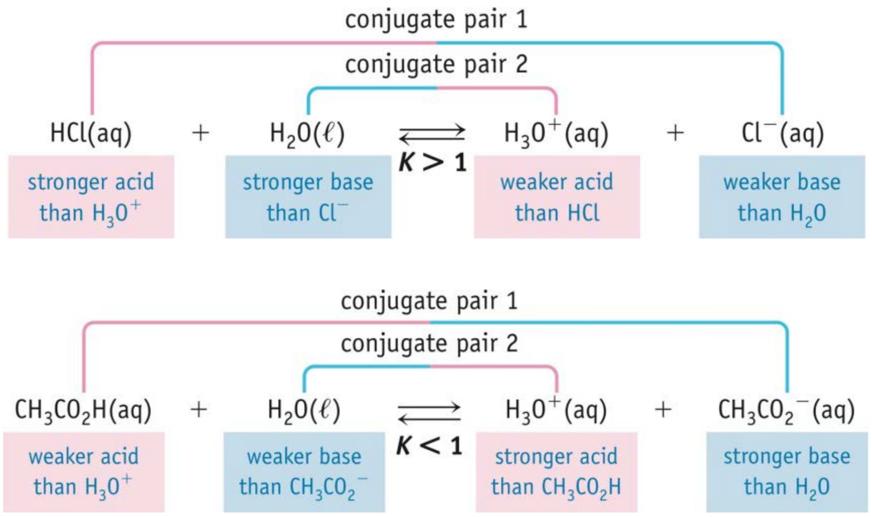
Property	Strong Base	Weak Base
K_b value	either 1) dissociates to give	K_b is small
Position of the dissociation equilibrium	OH^- ions to the solution	Far to the left
Equilibrium concentration of OH^- compared with original concentration of B:	or 2) reacts with water in either case:	$[OH^-] \ll [B:]_0$
Strength of conjugate acid compared with that of water	$[OH^-] = [B:]_0$	BH^+ much stronger acid than water

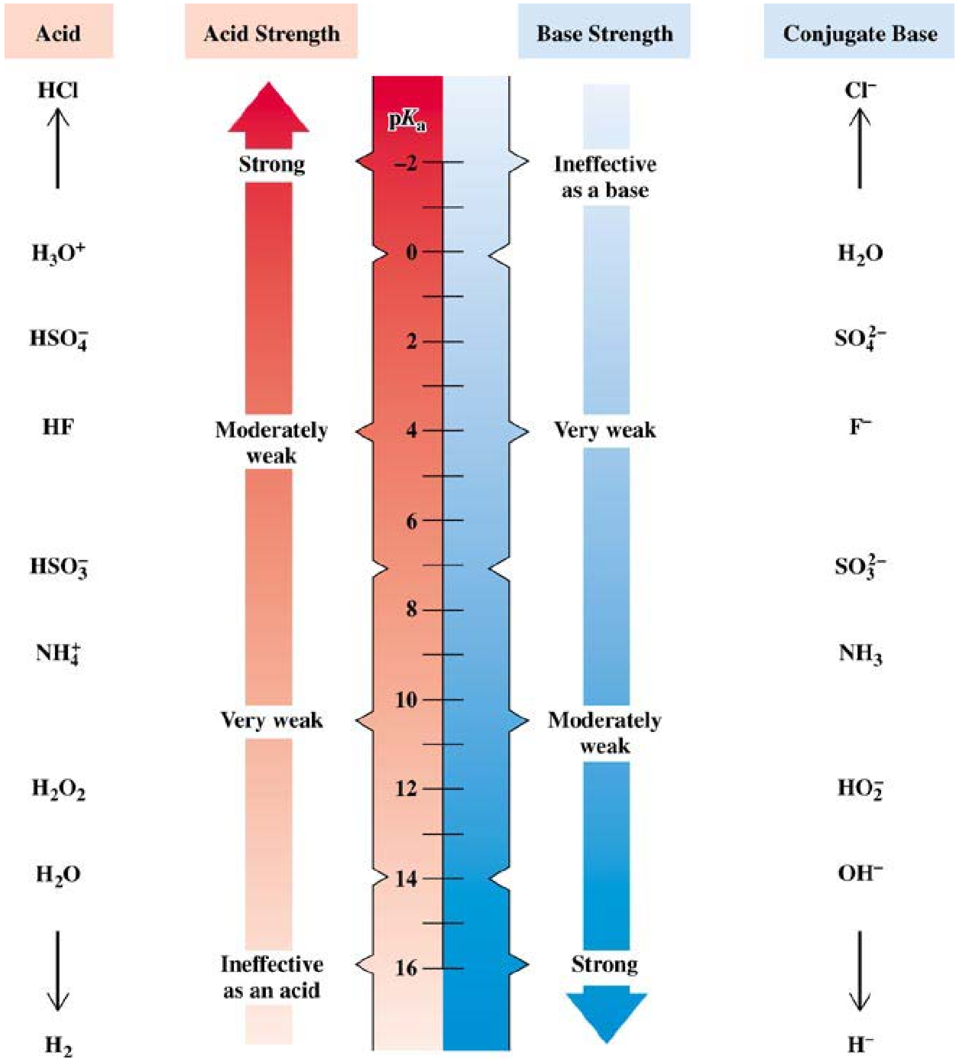
Conjugate Acid/Base Pairs

The conjugate base of a weak acid is a weak base. The weaker the acid, the stronger the base. However, if one member of a conjugate pair is weak, so is its conjugate.

The relation between K_a for an acid and K_b for its conjugate base in aqueous solution is $K_w = K_a \times K_b$.

When a strong acid (or base) is added to a weak base (or acid), they react nearly completely.





Water and the pH Scale

Temperature Dependence of K_w	
Temperature (°C)	K_w
0	0.114×10^{-14}
10	0.292×10^{-14}
20	0.681×10^{-14}
25	1.01×10^{-14}
30	1.47×10^{-14}
40	2.92×10^{-14}
50	5.47×10^{-14}
60	9.61×10^{-14}

water autoionization: $2\ H_2O(l) \Leftarrow\Rightarrow H_3O^+(aq) + OH^-(aq)$

$$K_w = [H_3O^+][OH^-] = 1.01 \times 10^{-14} \text{ (at 25°C)}$$

K_w is an equilibrium constant which depends upon temperature.

pH is temperature dependent

condition	concentrations	pH (only at 25°C)
acidic	$[H_3O^+] > [OH^-]$	pH < 7
neutral	$[H_3O^+] = [OH^-]$	pH = 7
basic	$[H_3O^+] < [OH^-]$	pH > 7

EX 6. $K_w = 2.4 \times 10^{-14}$ at body temperature (98.6°C = 37.0°C).

a) What is the hydrogen ion concentration?

$$K_w = [H_3O^+][OH^-] = x^2 \Rightarrow x = [H_3O^+] = \sqrt{K_w} = \sqrt{(2.4 \times 10^{-14})} = \mathbf{1.5 \times 10^{-7}\ M}$$

b) What is the pH?

$$pH = -\log_{10}[H_3O^+] = -\log_{10}(1.549 \times 10^{-7}) = 6.80989 \Rightarrow \mathbf{6.81}$$

significant figures for logarithms:

3 significant digits   3 decimal places

$$\log 1.00 \times 10^{-3} = 3.000$$

Water and the pH Scale

other "p" functions:

$pH = -\log_{10}[H_3O^+]$

$pOH = -\log_{10}[OH^-]$

$pK_w = -\log_{10}K_w$

$pK_a = -\log_{10}K_a$

$pK_b = -\log_{10}K_b$

$pK_{sp} = -\log_{10}K_{sp}$

EX 7. Answer each of the following

a) $pH = 9.3$, $[H_3O^+] = ?$

$pH = -\log_{10}[H_3O^+] \Rightarrow [H_3O^+] = 10^{-pH} = 10^{-9.3} = 5.01 \times 10^{-10} \Rightarrow \mathbf{5 \times 10^{-10}}$

b) 0.40 moles of $Ba(OH)_2$ is dissolved in a liter of water, pOH?

$Ba(OH)(s) \rightarrow Ba^{2+}(aq) + 2\ OH^-(aq) \Rightarrow 0.80\ M\ [OH^-]$