UNIT 8 – CHAPTER 14 STUDENT NOTES: ACIDS AND BASES

Definitions

Arrhenius acid-base

Acid:

Base:

Bronstad-Lowry

Acid:

Base:

Conjugate acid:

Conjugate base:

$[H_3O^+]$ IS EQUIVALENT TO $[H^+]$

EX: $HA + H_2O \longleftrightarrow H_3O^+ + A^-$

$HC_2H_3O_2 + H_2O \longleftrightarrow H_3O^+ + C_2H_3O_2^-$

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

$H_2S + NH_3 \longleftrightarrow HS^- + NH_4^+$

$H_2O + H_2O \longleftrightarrow H_3O^+ + OH^-$

Acid dissociation constant – is used to determine the strength of an acid.

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

Strong acids – completely dissociate

Weak acids – dissociate incompletely

Page 628 table 14.2 dissociation constants

Strong acid dissociation- $HCl + H_2O \longleftrightarrow H_3O^+ + Cl^-$

100% DISSOCIATION

$$K_a = \frac{[H_3O^+][Cl^-]}{[HCl]}$$

Strong acids are:

HBr, HI, HCl, HNO_3, H_2SO_4, $HClO_4$

Weak acid dissociation $HC_2H_3O_2 + H_2O \longleftrightarrow H_3O^+ + C_2H_3O_2^-$

<100% DISSOCIATION

$$K_a = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

$K_a = 1.8 \times 10^{-5}$ (SMALL NUMBER)

EX 1: The stronger the acid, the weaker the conjugate base. The stronger the base, the weaker the conjugate acid. Arrange the following acids in order of their strength, then arrange the conjugate bases in order of strength.

HOC_6H_5, H_2O, HSO_4^-, NH_4^+, HNO_3

water is a stronger base than the conjugate base of a strong acid

EX 2: Arrange the following species according to their strengths as bases:

H_2O, F^-, Cl^-, NO_2^-, CN^-

Amphoteric – sometimes acts like an acid, others a base depending on what it is with

$$H_2O + H_2O \longleftrightarrow H_3O^+ + OH^-$$

$[H^+] = [OH^-] = 1\times10^{-7}$ (NEUTRAL)
$[H^+] < [OH^-]$ = BASE
$[H^+] > [OH^-]$ = ACID

<u>Formulas</u>

$K_w = [H_3O^+] \cdot [OH^-] = 1 \text{ X } 10^{-14}$ <u>ion product constant</u>

pH = - log $[H_3O^+]$

pOH = - log $[OH^-]$

$[H^+]$ = antilog [-pH]

$[OH^-]$ = antilog [-pOH]

EX 3: Calculate the pH of the following strong acids

$1\times10^{-2} \leftarrow$ a) 0.01 *M* HCl $PH = -LOG[.01] = 2$

$1\times10^{-3} \leftarrow$ b) 0.001 *M* HNO_3 $PH = -LOG[.001] = 3$

$1\times10^{0} \leftarrow$ c) 1 *M* H_2SO_4 $PH = -LOG[1] = 0$

EX 4: Calculate the pH of the following strong bases

$1\times10^{-2} \leftarrow$ a) 0.01 *M* NaOH $* [H^+][OH^-] = 1\times10^{-14}$
$[H^+][.01] = 1\times10^{-14}$

$[H^+] = 1\times10^{-12}$ $PH = -LOG[1\times10^{-12}] = 12$

$1\times10^{-3} \leftarrow$ b) 0.001 *M* KOH $* POH = -LOG[OH^-]$
$POH = -LOG(.001)$

$POH = 3.00 \rightarrow$ $pH + pOH = 14.00$
$pH + 3.00 = 14.00$
$PH = 11$

EX 5: Calculate the $[H_3O^+]$ in the following

a) pH = 2.2 HCl $[H^+] = ANTILOG(-PH) \rightarrow \overset{10^x}{ANTILOG}(-2.2) = 6.2\times10^{-3}\,M$

b) pH = 9.0 NaOH $[H^+] = ANTILOG(-PH) \rightarrow ANTILOG(-9.0) = 1\times10^{-9}\,M$

WHOLE #,
SO WE KNOW THE EXPONENT!

EX 6: Calculate the pH of a 0.100 *M* aqueous solution of HOCl (hypochlorous acid) $K_a = 3.5 \text{ X } 10^{-8}$

EX 7: Calculate the pH of a 0.200 M aqueous solution of $C_6H_{18}O_8$, citric acid (HCA) $K_a = 7.45 \times 10^{-4}$

	HCA	$\rightarrow$ H^+	+ CA^-
I	.200	∅	∅
C	−X	X	X
E	.200−X	X	X

$K_a = 7.45\times10^{-4} = \frac{(X)(X)}{(.200-X)} \rightarrow X = .0122$ $pH = -\log(.0122)$ $\boxed{pH = 1.91}$

$\frac{.0122}{.200} \times 100\%$ $=6.1\%$

$K_a = 7.45\times10^{-4} = \frac{X^2}{.200-X}$

$X^2 + .00075X - .00149 = ∅$

$X = \frac{-B \pm \sqrt{B^2 - 4AC}}{2A}$

$X = .0118$ or $X = -1.\cancel{X}8\times10^{-2}$

$PH = -\log(.0118) = 1.93$

EX 8: Calculate the pH of a solution that contains equal amounts of

MAJOR SPECIES:
HCN, HNO_2, H_2O

1.00 M HCN $K_a = 6.2 \times 10^{-10}$

5.00 M HNO_2 $K_a = 4.0 \times 10^{-4}$

* $HOH \rightarrow H^+ + OH^-$ $[H^+] = 1\times10^{-7}$
* $HCN \rightarrow H^+ + CN^-$ $[H^+] = 2.49\times10^{-5}$

$1.00-X$ X   Y

$K_a = 6.2\times10^{-10} = \frac{(X)(X)}{(1.00-X)}$

* $HNO_2 \rightarrow H^+ + NO_2^-$

$5.00-X$ X   X

$K_a = 4.0\times10^{-4} = \frac{(X)(X)}{(5.00-X)}$

$[H^+] = 4.5\times10^{-2}$

$PH = -\log(4.5\times10^{-2})$

$PH = 1.35$

EX 9: Calculate the % dissociation of $HC_2H_3O_2$ at: ($K_a = 1.8 \times 10^{-5}$)

a) 1.00 M $HC_2H_3O_2$

b) 0.100 M $HC_2H_3O_2$

a) $HC_2H_3O_2 \rightarrow H^+ + C_2H_3O_2^-$

$1.00-X$ X   X

$K_a = 1.8\times10^{-5} = \frac{(X)(X)}{(1.00-X)}$ $X = 4.2\times10^{-3}$

$\%\text{ DISSOCIATION} = \frac{4.2\times10^{-3}}{1.00} \times 100\%$ $= .42\%$

b) $HC_2H_3O_2 \rightarrow H^+ + C_2H_3O_2^-$

$.100-X$ X   X

$K_a = 1.8\times10^{-5} = \frac{(X)(X)}{(.100-X)} = 1.3\times10^{-3}$

$\%\text{ DISSOCIATION} = \frac{1.3\times10^{-3}}{.100} \times 100\% = 1.3\%$

EX 10: A 0.500 M solution of $HC_5H_3N_4O_4$ (uric acid) is 1.6% dissociated. Calculate K_a.

1) $HC_5H_3N_4O_4 \rightarrow H^+ + C_5H_3N_4O_4^-$

$.500-X$ X   X

2) $1.6\% = \frac{[C_5H_3N_4O_4^-]}{.500} \times 100$

3) $[C_5H_3N_4O_4^-] = [H^+] = .0080M$

$K_a = \frac{(.0080)(.0080)}{.500 - .0080} = \boxed{1.3\times10^{-4}}$

EX 11: Calculate the K_a for .100 M lactic acid ($HC_3H_5O_3$) that is 3.7% dissociated.

1) $HC_3H_5O_3 \rightarrow H^+ + C_3H_5O_3^-$

$.100-X$ X   X

2) $3.7\% = \frac{[C_3H_5O_3^-]}{.100} \times 100\%$

3) $[C_3H_5O_3^-] = [H^+] = .0037M$

$K_a = \frac{(.0037)(.0037)}{.100 - .0037}$

$= 1.4\times10^{-4}$

Calculating the pH of bases (K_b are listed on page 647)

EX 12: Calculate the pH of a 0.05 *M* solution of NaOH

pH + pOH = 14

EX 13: Calculate the pH for a 15 *M* solution of NH_3 $K_b = 1.8 \times 10^{-5}$

$NH_3 + HOH \rightarrow NH_4^+ + OH^-$

15 − x x x

$K_b = 1.8 \times 10^{-5} = \frac{(x)(x)}{(15-x)}$ x = .016M

pOH = −log(.016)

pOH = 1.80

pH = 12.20

EX 14: Calculate the pH of a 1.0 *M* solution of methylamine (CH_3NH_3) P.647

$CH_3NH_3 + HOH \rightarrow CH_3NH_4^+ + OH^-$

1.0 − x x x

$K_b = 4.38 \times 10^{-4} = \frac{(x)(x)}{(1.0-x)}$ x = .021

pOH = −log(.021)

pOH = 1.68

pH = 12.32

Acid-base properties of salts

Salt – ionic compound, metal-nonmetal combination

There are 3 types of salt solutions to produce pH

1) Salts that produce neutral solutions

Salts that consist of cations (+) – strong bases; anions (-) – strong acids

EX: $HCl \rightarrow H^+ + Cl^-$

$NaOH \rightarrow Na^+ + OH^-$

These ions have no effect on the [H⁺] in water

2) Salts that produce basic solutions

Salts that contain anions (-) that are conjugate bases of weak acids

EX: $NaC_2H_3O_2 \longleftrightarrow H^+ + C_2H_3O_2^-$

Calculate the pH of a 0.100 *M* solution of $NaC_2H_3O_2$

3) Salts that produce acidic solutions

Salts that contain cations (+) that are conjugate acids of weak bases.

EX: $NH_4Cl \longleftrightarrow NH_4^+ + Cl^-$

Calculate the pH of 0.100 *M* solution of NH_4Cl

EX 15: Calculate the pH of a 0.30 *M* NaF solution

EX 16: Calculate the pH of a 0.500 *M* $NaNO_3$ solution

$K_a \times K_b = 1\times10^{-14}$

EX 17: Calculate the pH of a 0.30 *M* $C_2H_5NH_3Cl$ solution

$C_2H_5NH_3Cl \rightarrow C_2H_5NH_3^+ + Cl^-$

$C_2H_5NH_3^+ \rightarrow C_2H_5NH_2 + H^+$

$.30-x \qquad x \qquad x$

$(K_a)(5.6\times10^{-4}) = 1\times10^{-14}$

$K_b = 5.6\times10^{-4}$

$K_a = 1.8\times10^{-11} = \frac{(x)(x)}{.30-x} = 2.3\times10^{-6}$

$pH = -\log(2.3\times10^{-6})$

$pH = 5.6$

If both ions can contribute to acidity or basicity, how do you determine if the solution is acidic or basic?

$K_a > K_b$	pH < 7, acid
$K_a < K_b$	pH > 7, basic
$K_a = K_b$	pH = 7, neutral

Predict whether an aqueous solution of each of the following salts are acid, base, or neutral

a) $NH_4C_2H_3O_2$

$NH_4C_2H_3O_2 \rightarrow NH_4^+ + C_2H_3O_2^-$

$NH_4^+ : K_a = 5.6\times10^{-10}$

$C_2H_3O_2^- : K_b = 5.6\times10^{-10}$

$K_a = K_b$, SO NEUTRAL

b) NH_4CN

$NH_4CN \rightarrow NH_4^+ + CN^-$

$NH_4^+ : K_a = 5.6\times10^{-10}$

$CN^- : K_b = (6.2\times10^{-10})(K_b) = 1\times10^{-14} \qquad K_b = 1.6\times10^{-5}$

$K_a < K_b$, SO BASE

<u>Calculating the pH of polyprotic acids</u>

EX 18: Calculate the pH of a 5.0 *M* H_3PO_4 solution and the equilibrium [ ] of the species [H_3PO_4], [$H_2PO_4^-$], [HPO_4^{2-}], and [PO_4^{3-}]