

UNIT 8 – CHAPTER 14 STUDENT NOTES: ACIDS AND BASES

Definitions

Arrhenius acid-base

Acid:

Base:

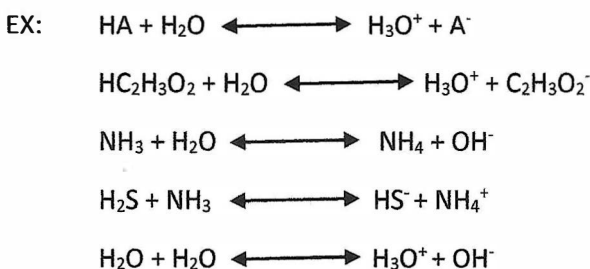
Bronstad-Lowry

Acid:

Base:

Conjugate acid:

Conjugate base:



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

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

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

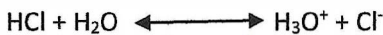
Strong acids – completely dissociate

Weak acids – dissociate incompletely

Page 628 table 14.2 dissociation constants

Strong acid dissociation-

100% DISSOCIATION



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

Strong acids are:

HBr, HI, HCl, HNO₃, H₂SO₄, HClO₄

Weak acid dissociation

<100% DISSOCIATION



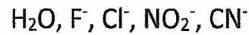
$$K_a = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} \quad K_a = 1.8 \times 10^{-5} \text{ (small number)}$$

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.



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:



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



$[H^+] = [OH^-] = 1 \times 10^{-7}$ (NEUTRAL)
 $[H^+] < [OH^-] =$ BASE
 $[H^+] > [OH^-] =$ ACID

Formulas

$K_w = [H_3O^+] \cdot [OH^-] = 1 \times 10^{-14}$ ion product constant

$pH = -\log [H_3O^+]$

$pOH = -\log [OH^-]$

$[H^+] = \text{antilog} [-pH]$

$[OH^-] = \text{antilog} [-pOH]$

EX 3: Calculate the pH of the following strong acids

- 1×10^{-2} ← a) 0.01 M HCl $pH = -\log [0.01] = 2$
- 1×10^{-3} ← b) 0.001 M HNO₃ $pH = -\log [0.001] = 3$
- 1×10^0 ← c) 1 M H₂SO₄ $pH = -\log [1] = 0$

EX 4: Calculate the pH of the following strong bases

- 1×10^{-2} ← a) 0.01 M NaOH $* [H^+][OH^-] = 1 \times 10^{-14}$
 $[H^+][0.01] = 1 \times 10^{-14}$
 $* pOH = -\log [OH^-]$
 $pOH = -\log (0.01)$
 - 1×10^{-3} ← b) 0.001 M KOH
- $[H^+] = 1 \times 10^{-12}$ $pH = -\log [1 \times 10^{-12}] = 12$
- $pOH = 3.00 \rightarrow pH + pOH = 14.00$
 $pH + 3.00 = 14.00$
 $pH = 11$

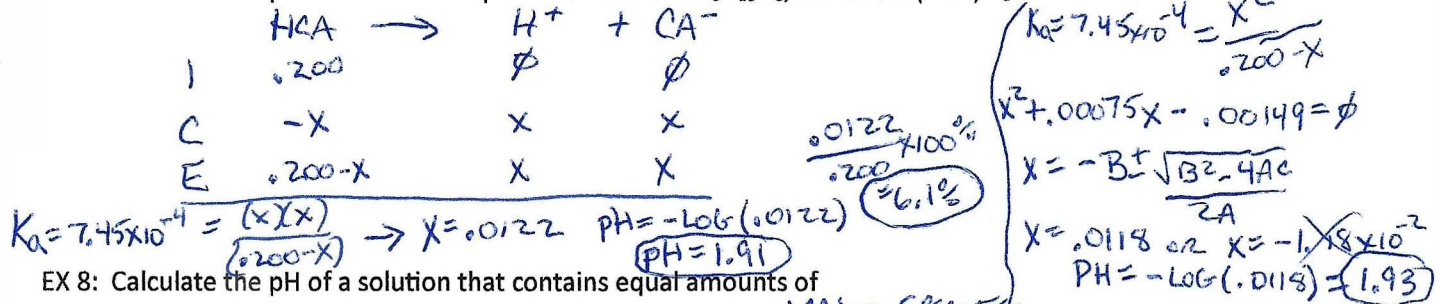
EX 5: Calculate the [H₃O⁺] in the following

- a) pH = 2.2 HCl $[H^+] = \text{ANTILOG} (-pH) \rightarrow \text{ANTILOG} (-2.2) = 6.2 \times 10^{-3} M$
- b) pH = 9.0 NaOH $[H^+] = \text{ANTILOG} (-pH) \rightarrow \text{ANTILOG} (-9.0) = 1 \times 10^{-9} M$

WHOLE #,
 SO WE KNOW THE EXPONENT!

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

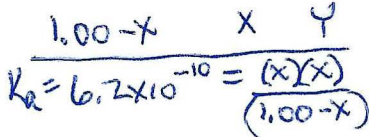
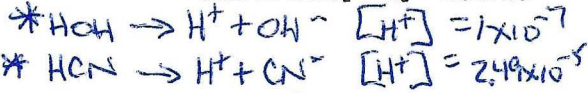
EX 7: Calculate the pH of a 0.200 M aqueous solution of $C_6H_8O_8$, citric acid (HCA) $K_a = 7.45 \times 10^{-4}$



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

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

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



$[H^+] = 4.5 \times 10^{-2}$

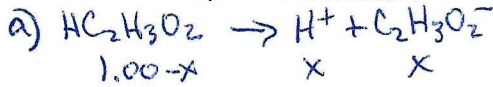
$pH = -\log(4.5 \times 10^{-2})$
 $pH = 1.35$

EX 9: Calculate the % dissociation of $HC_2H_3O_2$ at:

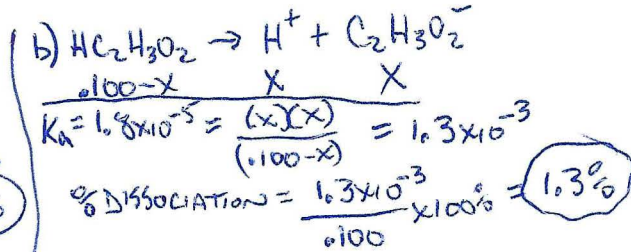
($K_a = 1.8 \times 10^{-5}$)

a) 1.00 M $HC_2H_3O_2$

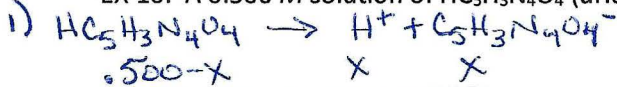
b) 0.100 M $HC_2H_3O_2$



$K_a = 1.8 \times 10^{-5} = \frac{(X)(X)}{(1.00-X)}$
 $X = 4.2 \times 10^{-3}$
 $\% \text{ DISSOCIATION} = \frac{4.2 \times 10^{-3}}{1.00} \times 100\% = 0.42\%$



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

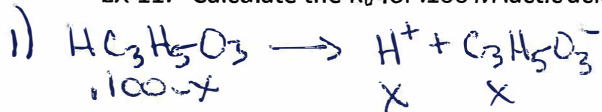


2) $1.6\% = \frac{[C_5H_3N_4O_4^-]}{0.500} \times 100$

3) $[C_5H_3N_4O_4^-] = [H^+] = 0.0080 M$

$K_a = \frac{(0.0080)(0.0080)}{0.500 - 0.0080} = 1.3 \times 10^{-4}$

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



2) $3.7\% = \frac{[C_3H_5O_3^-]}{0.100} \times 100\%$

3) $[C_3H_5O_3^-] = [H^+] = 0.0037 M$

$K_a = \frac{(0.0037)(0.0037)}{0.100 - 0.0037}$

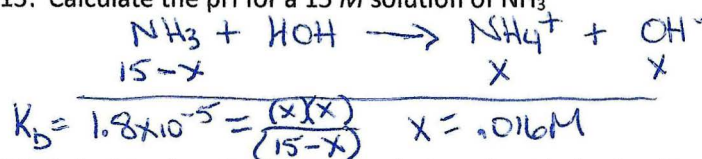
$= 1.4 \times 10^{-4}$

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

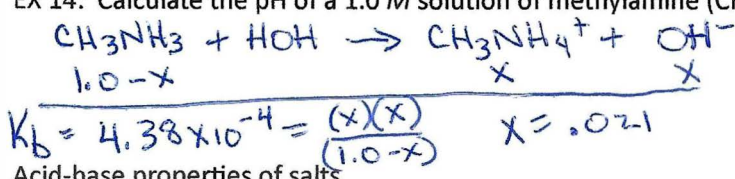
$pH + pOH = 14$

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



$pOH = -\log(0.016)$
 $pOH = 1.80$
 $pH = 12.20$

EX 14: Calculate the pH of a 1.0 M solution of methylamine (CH_3NH_2) $K_b = 4.38 \times 10^{-4}$



$pOH = -\log(0.021)$
 $pOH = 1.68$
 $pH = 12.32$

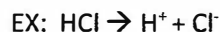
Acid-base properties of salts

Salt – ionic compound, metal-nonmetal combination

There are 3 types of salt solutions to produce pH

1) Salts that produce neutral solutions

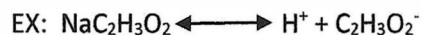
Salts that consist of cations (+) – strong bases; anions (-) – strong acids



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

2) Salts that produce basic solutions

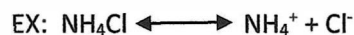
Salts that contain anions (-) that are conjugate bases of weak acids



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.

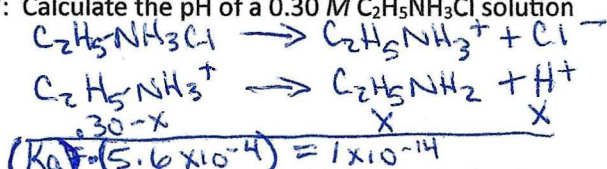


Calculate the pH of 0.100 M solution of NH_4Cl

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

EX 16: Calculate the pH of a 0.500 M NaNO₃ solution

EX 17: Calculate the pH of a 0.30 M C₂H₅NH₃Cl solution

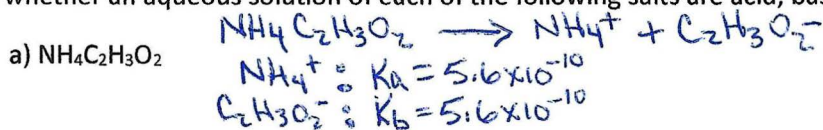


$$\begin{aligned} K_b &= 5.6 \times 10^{-4} \\ K_a &= 1.8 \times 10^{-11} = \frac{(x)(x)}{0.30 - x} \\ K_a \times K_b &= 1 \times 10^{-14} \\ &= 2.3 \times 10^{-6} \\ \text{pH} &= -\log(2.3 \times 10^{-6}) \\ \text{pH} &= 5.6 \end{aligned}$$

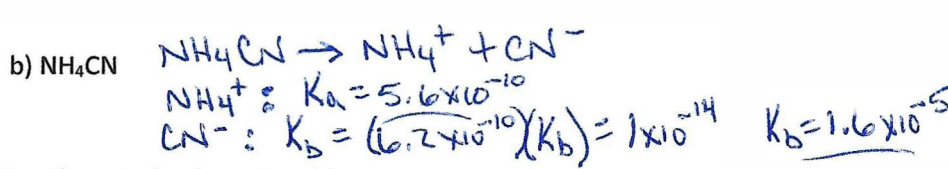
If both ions can contribute to acidity or basicity, how do you determine if the solution is acidic or basic?

- $K_a > K_b$ 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



$K_a = K_b$,
SO NEUTRAL



$K_a < K_b$,
SO BASE

Calculating the pH of polyprotic acids

EX 18: Calculate the pH of a 5.0 M H₃PO₄ solution and the equilibrium [] of the species [H₃PO₄], [H₂PO₄], [HPO₄²⁻], and [PO₄³⁻]