

## A. DETERMINING THE pH OF ACIDS

## 1. STRONG ACIDS

- Complete dissociation, therefore,  $[H^+] = [HA]$
- $pH = -\text{Log } [H^+]$

## 2. WEAK ACIDS

- Determine equilibrium systems present
- Find  $K_a$  for each system
- Use  $K_a$  expression for system with largest  $K_a$
- Construct a table ( $[ ]_0$ ,  $\Delta[ ]$ ,  $[ ]_{eq}$ )
- Substitute equilibrium  $[ ]$ 's into  $K_a$  expression
- Assume  $[HA]_0 \approx [HA]_{eq}$
- Solve for  $x = [H^+] = \sqrt{K_a C_a}$
- Calculate % Dissociation =  $\left( \frac{[H^+]}{[HA]} \right) \times 100$   
 $< 5\%$ , Solve for pH and other  $[ ]$ 's  
 $> 5\%$ , Solve for  $x$  with the Quadratic Equation

B. DETERMINE  $K_a$ , GIVEN pH AND  $[HA]_0$ 

- Write the equilibrium expression for the reaction
- Construct a table to determine equilibrium concentrations in terms of  $x$
- Determine the  $[H^+] = -pH \text{ Inv Log}$
- Since  $[H^+] = x$ , substitute this value back into the  $K_a$  expression and solve for  $K_a$ ; be sure to subtract the value of  $x$  from  $[HA]$

## C. DETERMINING THE pH OF BASES

## 1. STRONG BASES

- Complete dissociation, therefore,  $[OH^-] = K_b/[B^-]$
- $pH = 14 - \text{Log } [OH^-]$
- An alternative method is to solve for pOH and subtract this value from 14

## 2. WEAK BASES

- $A, 2a$
- Find  $K_b$  for each system; if the base is a conjugate base of a polyprotic acid, look up  $K_a$  for its conjugate acid and determine  $K_b = K_w/K_a$
- Use the  $K_b$  expression for the system with largest  $K_b$
- Construct a table ( $[ ]_0$ ,  $\Delta[ ]$ ,  $[ ]_{eq}$ )
- Substitute equilibrium  $[ ]$ 's into  $K_b$  expression
- Assume  $[B^-]_0 \approx [B^-]_{eq}$
- Solve for  $x = [OH^-] = \sqrt{K_b C_b}$
- Calculate % Dissociation =  $\left( \frac{[OH^-]}{[B^-]} \right) \times 100$   
 $< 5\%$ , Solve for pH and other  $[ ]$ 's (By way of pOH or  $[OH^-] = K_w/[H^+]$ )  
 $> 5\%$ , Solve for  $x$  with the Quadratic Equation

D. DETERMINE  $K_b$ , GIVEN pH AND  $[B^-]_0$ 

- Write the equilibrium expression for the reaction
- Construct a table to determine equilibrium concentrations in terms of  $x$
- Determine the  $[OH^-]$  from pH either by way of pOH or  $[OH^-] = K_w/[H^+]$
- Since  $[OH^-] = x$ , substitute this value back into the  $K_b$  expression and solve for  $K_b$ ; be sure to subtract the value of  $x$  from  $[B^-]$

## E. DETERMINING THE pH OF SALT SOLUTIONS

1. Determine the acidity/basicity/neutrality of the ions
2. If one ion is acidic and the other neutral, or both are acidic, follow A,2a-h
3. If one ion is basic and the other neutral, or both are basic, follow C,2a-h
4. If one ion is acidic and the other is basic,
  - a. Write the equation for the dissolution of the acidic ion in water and the dissolution of the basic ion in water separately
  - b. Determine the respective  $K_a, K_b$  values
  - c. Use the  $K$  expression for the system with the largest  $K$  value
  - d. Follow A,2d-h, or C,2d-h, depending on the dominant system

### A. Determine acidity/basicity/neutrality of cation/anion pair

1. Split the salt into cation and anion
2. Ask: Is the cation a metal of Group 1 or 2?
  - a. Yes, the cation is a spectator ion
  - b. No, the cation is acidic (there are no basic cations)
3. Ask: Is the anion strong acid related? (Is it  $\text{Cl}^-, \text{Br}^-, \text{I}^-, \text{ClO}_4^-, \text{NO}_3^-, \text{SO}_4^{2-}$ ?)
  - a. Yes, the anion is a spectator ion
  - b. No, is the anion amphoteric (can donate/accept  $\text{H}^+$ 's)
    - (1) No, the anion is basic
    - (2) Yes, do the following:
      - (a) Write out the equation for the dissociation of the species as a weak acid; look up the  $K_a$  for the weak acid
      - (b) Write out the equation for the dissociation of the species as a weak base; look up the  $K_b$  for the weak base
      - (c) If  $K_a > K_b$ , the species is acidic; if not then it is basic

### B. Determining the acidity of the salt once the acidity/basicity/neutrality of the individual ions is known

1. Salt solutions are neutral if both ions are spectator ions.
2. Salt solutions are acidic if one ion is a spectator ion and the other is acidic.
3. Salt solutions are basic if one ion is a spectator ion and other is basic.
4. The acidity/basicity of a salt solution made up of one acidic ion and one basic ion cannot be determined without further information. The information needed is the  $K_a$  of the acidic ion and the  $K_b$  of the basic ion; if  $K_a > K_b$ , then the salt is acidic – otherwise it is basic.

TITRATION CURVES

STRATEGIES FOR IDENTIFYING THE TYPE OF SYSTEM, WHERE YOU ARE IN THE SYSTEM, AND HOW TO DETERMINE THE pH AT ANY POINT IN THE SYSTEM:

<u>SYSTEM/ TITRANT</u>	<u>STRONG ACID/ STRONG BASE</u>	<u>WEAK ACID/ STRONG BASE</u>	<u>WEAK BASE/ STRONG ACID</u>
Acid/Base	$HA \rightarrow H^+ + A^-$	$HA \rightleftharpoons H^+ + A^-$	$A^- + H_2O \rightleftharpoons HA + OH^-$
Initially, Before Titration	$[H^+] = [HA]$ $pH = -\text{Log}[H^+]$	$K_a = \frac{[H^+][A^-]}{[HA]}$ $[H^+] = \sqrt{K_a C_a}$ $pH = -\text{Log}[H^+]$	$K_b = \frac{[OH^-][HA]}{[A^-]}$ $[OH^-] = \sqrt{K_b C_b}$ $pOH = -\text{Log}[OH^-]$ $pH = 14 - pOH$
Before Equivalence Point	$pH = -\text{Log}[H^+]$	Buffer Situation $pH = pK_a + \frac{[A^-]}{[HA]}$	Buffer Situation $pH = pK_a + \frac{[A^-]}{[HA]}$
1/2 Way to Equivalence Point	$pH = -\text{Log}[H^+]$	$pH = pK_a$	$pH = pK_a$
@ Equivalence Point	$pH = 7$	$pH > 7$ $K_b = \frac{[OH^-][HA]}{[A^-]}$ $[OH^-] = \sqrt{K_b C_b}$ $pOH = -\text{Log}[OH^-]$ $pH = 14 - pOH$	$pH < 7$ $K_a = \frac{[H^+][A^-]}{[HA]}$ $[H^+] = \sqrt{K_a C_a}$ $pH = -\text{Log}[H^+]$
After Equivalence Point	$pH > 7$ Base $[OH^-]$ + Salt $pOH = -\text{Log}[OH^-]$ $pH = 14 - pOH$	$pH > 7$ Base $[OH^-]$ + Salt $pOH = -\text{Log}[OH^-]$ $pH = 14 - pOH$	$pH < 7$ Acid $[H^+]$ + Salt $pH = -\text{Log}[H^+]$

# 45 TITRATION CURVES

