**Name: Period: Seat#:**

**S-76, 77, 78, 79, 80, 81, 82, 83**

**Directions:** Try these problems. If you can DO them, check the box (🗹).   
If you CANNOT do them, write some notes TO YOURSELF about what you need to study to succeed at these problems.

**S76 – Quick Check #1**

🞎 **Acid-Base Equilibria**

Consider the following equilibrium:

HCN + H2O(*l*) ⮀ H3O+ + CN Ka for HCN = 4.0 x 1010

Identify the two **conjugate acid-base pairs**:

\_\_\_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_\_\_

Draw a box around the more appropriate representation of this equilibrium:

HCN + H2O(*l*) ⮀ H3O+ + CN HCN + H2O(l) ⮀ H3O+ + CN

Which acid is ***weaker***? \_\_\_\_\_\_\_\_\_\_

🞎 **Conjugate Acids and Bases**

What is the…

conjugate ***base*** of H2O? \_\_\_\_\_\_

conjugate ***acid*** of NH3? \_\_\_\_\_\_

conjugate ***base*** of OH? \_\_\_\_\_\_

conjugate ***acid*** of HCO3? \_\_\_\_\_\_

🞎 **pH Calculations**

Fill in the chart below:

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **[H3O+]** | **[OH]** | **pH** | **pOH** | **Acidic or Basic** |
| 2.0 x 103 |  |  |  |  |
|  |  | 6.25 |  |  |
|  | 5.6 x 102 |  |  |  |

**S77 – Quick Check #2**

🞎 **Equilibrium Favors the Weaker Acid/Weaker Base**

Consider this equation: HCN + H2O ⮀ H3O+ + CN Ka = 4.0 x 10-10.

The two bases are: \_\_\_\_\_ and \_\_\_\_\_. The weaker base is \_\_\_\_\_\_\_\_.

🞎 **I.C.E. Box Problem**

Calculate the pH of a 0.100 M HCN solution. Ka for HCN = 4.0 x 10 -10

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
|  | HCN | H2O(*l*) | ⮀ | H3O+ | CN |
| Initial |  |  |  |  |  |
| Change |  |  |  |  |  |
| Equilibrium |  |  |  |  |  |

🞎 **pH Problems**

Calculate the pH of a 0.100 M HBr solution. \_\_\_\_\_\_\_

Calculate the pH of a 0.100 M KOH solution. \_\_\_\_\_\_\_

Calculate the pH of a 0.100 M NH3 solution. \_\_\_\_\_\_\_ Kb for NH3 = 1.8 x 10-5

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
|  | NH3 | H2O(l) | ⮀ | NH4+ | OH |
| Initial |  |  |  |  |  |
| Change |  |  |  |  |  |
| Equilibrium |  |  |  |  |  |

🞎 **Conjugate Bases**

CN- is the conjugate base of the weak acid, HCN. Finish the equation below:

CN + H2O ⮀

**S78 – Quick Check #3**

🞎 **Diprotic Acids**

Sulfurous acid, H2SO3, is a diprotic acid. Write the step-wise dissociation equations for H2SO3.

|  |  |
| --- | --- |
|  | Ka1 = 1.5 x 10-5 |
|  | Ka2 = 1.0 x 10-7 |

🞎 **Lewis Acids and Bases**

Consider the following picture. The Lewis acid is \_\_\_\_\_\_\_\_\_\_\_\_. The Lewis base is \_\_\_\_\_\_\_\_\_\_\_.

Diagram

Description automatically generated

Consider the equation: H+ + OH ⭢ H2O. The Lewis acid is \_\_\_\_\_\_\_. The Lewis base is \_\_\_\_\_\_.

🞎 **Strengths of Acids**

Consider the acids: HClO2, HBrO2, HIO2. Rank them from weakest to strongest.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Weakest |  |  |  | Strongest |

Justification:

Consider the acids: HBrO, HBrO2, HBrO3. Rank them from weakest to strongest.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Weakest |  |  |  | Strongest |

Justification:

Consider the acids: HCl, HBr, HI. Rank them from weakest to strongest.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Weakest |  |  |  | Strongest |

Justification:

🞎 **Diprotic Acid Calculations**

Sulfurous acid, H2SO3, is a diprotic acid. Ka1 = 1.5 x 10-5; Ka2 = 1.0 x 10-7

What is the [SO32-] in a 0.150 M solution of H2SO3? \_\_\_\_\_\_\_\_\_\_\_

Calculate the pH of a 0.150 M solution of H2SO3.

**S79 – Quick Check #4**

🞎 **pH’s are Logarithmic**

Solution A has a pH of 3. Solution B has a pH of 6. Which solution is more acidic? \_\_\_\_\_\_\_\_

How many times more acidic is the more acidic solution? \_\_\_\_\_\_

🞎 **ICE Box with a Twist**

A 0.10 M solution of HF has a pH of 2.10. Calculate the Ka of HF.

**S80 – Quick Check #1b**

🞎 **pH at the Starting Point of a Titration**

What is the pH of a 25.0 mL sample of 0.200 M HCl?

🞎 **How Much Base is Needed to Neutralize an Acid**

How many mL of 0.100 M NaOH solution is needed to titrate a 25.0 mL sample of a 0.200 M HCl.

🞎 **pH at the Endpoint of a Titration**

|  |  |  |
| --- | --- | --- |
| **Acid** | **Base** | **pH at the Endpoint (circle choice)** |
| strong | strong | less than 7 | 7 | more than 7 |
| strong | weak | less than 7 | 7 | more than 7 |
| weak | strong | less than 7 | 7 | more than 7 |

🞎 **pH Beyond the Endpoint of a Titration**

Calculate the pH of a solution made by adding 30.0 mL of 0.100 M NaOH to 10.0 mL of 0.200 M HCl.

🞎 **Same problem broken into steps for you**

Calculate the pH of a solution made by adding 30.0 mL of 0.100 M NaOH to 10.0 mL of 0.200 M HCl.

***Flow Map:***

|  |  |  |  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| Δ mL  to L | 🢡 | Moles OH | 🢡 | Moles  H+ | 🢡 | XS moles of OH  or H+ | 🢡 | Add Volumes | 🢡 | [H+] or [OH] | 🢡 | pH or pOH |

V⋅M = mol V⋅M = mol

Step 1: Change your volumes to Liters.

Vol NaOH = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ Vol HCl = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Step 2: Calculate moles of OH. *(Note: volume x molarity = moles)*

Step 3: Calculate moles of H+.

Step 4: Subtract moles of OH and moles H+ to determine excess moles.

*(Note: You are forming H2O until H+ or OH runs out. Be sure to label your answer as H+ or OH.)*

Step 5: Calculate the total volume (in Liters).

Step 6: Determine the concentration of H+ or OH (whichever is in excess).  *(Note: Molarity = moles/Liters)*

Step 7: Determine the pH or pOH. *(Note: This may involve one or two steps. State the equation used.)*

**S81 – Quick Check #2b**

🞎 **pH of a weak acid solution**

Calculate the pH of a 0.200 M solution of nitrous acid, HNO2. Ka of HNO2 = 4.0 x 104.

🞎 **Salt solutions**

A solution of NaNO2 will be \_\_\_\_\_\_\_\_\_\_\_ (acidic, basic, neutral).

Write the ***net*** equation for the equilibrium involved when NaNO2 dissolves in water.

Write the equilibrium expression for the above equation. Should this be labeled Kc, Ka, Kb, Keq?

Calculate the pH of a 0.100 M solution of NaNO2.

🞎 **Acid-Base Neutralization**

Write the balanced net equation for:

A solution of sulfurous acid is added to a suspension of magnesium hydroxide

**S82 – Quick Check #3b**

🞎 **Buffer basics**

What could you mix with 100. mL of 2.00 M HNO2 (nitrous acid) to make a buffer?

What is the pH of the best buffer made from nitrous acid, HNO2? Ka of HNO2 = 4.0 x 104.

🞎 **Adjusting the pH of a buffer**

The weak acid, HCN, could be used to make a buffer. The Ka for HCN is 6.2 x 1010.

What is the pH of the best buffer made from HCN?

What ratio of [HCN] to [CN] is needed to have a buffer with pH = 9.00?

🞎 **Multiple Choice Question:**

Which of the following mixtures would result in a buffer solution?

I. 10 mL 0.20 M HCl and 10 mL 0.40 M NH3

II. 10 mL 0.20 M HF and 10 mL 0.20 M NaF

III. 10 mL 0.40 M HC2H3O2 and 10 mL 0.20 M NaOH

A) II only

B) II and III only

C) I and III only

D) I, II and III

E) None of these will result in a buffer solution.

**S83 – Quick Check #4b**

1. Which statement is a logical consequence of the fact that a 0.10 molar solution of potassium acetate, KC2H3O2, is less basic than a 0.10 molar solution of potassium cyanide, KCN?
   1. Hydrocyanic acid (HCN) is a weaker acid than acetic acid.
   2. Hydrocyanic acid is less soluble in water than acetic acid.
   3. Cyanides are less soluble than acetates.
   4. Acetic acid is a weaker acid than hydrocyanic acid.
2. Which solution would show the least change in pH upon addition of 3.0 mL of 1.0 M KOH? (Assume equal volumes of each solution are used. Ka for HC2H3O2 = 1.8 x10-5)
   1. A solution that is 0.50 M acetic acid and 0.50 M sodium acetate.
   2. A solution that is 0.10 M acetic acid and 0.10 M sodium acetate.
   3. A solution that is 1.0 M acetic acid.
   4. A solution that is 0.50 M sodium acetate.
3. A strong monoprotic acid is being titrated with a 0.500 M NaOH solution. Which statement is true for this titration?

(A) The pH at the equivalence point cannot be determined without knowing the identity of the acid.

(B) The pH at the equivalence point cannot be determined unless the concentration of the acid is known.

(C) The pH at the equivalence point depends on neither the identity of the acid nor the concentration of the acid.

1. Which of the following would not make a good buffering system?

|  |  |
| --- | --- |
| (A) SO42– and H2SO4 | (B) HCO3– and H2CO3 |
| (C) NH3 and NH4+ | (D) CH3COO– and CH3COOH |

1. The amount (in grams) of sodium acetate (MW = 82.0) to be added to 500.0 mL of 0.200 molar acetic acid (Ka = 1.80 x 10-5) in order to make a buffer with pH = 5.000 is

(A) 69 (B) 0.180 (C) 14.9 (D) 29.5 (E) None of these

1. Determine the pH of a solution in which 1.00 mol H2CO3 (Ka = 4.2 x 10-7) and 1.00 mole NaHCO3 are dissolved in enough water to form 1.00 L of solution.
2. How many grams of Mg(OH)2 are required to neutralize 50.0 ml of a 3.00 M HCl solution?
3. A sample of 20.0 mL of a 0.100-molar HCN solution is titrated with a 0.150-molar NaOH solution. (Ka HCN = 6.2 x 10-10)
   1. What volume of NaOH is used in the titration in order to reach the equivalence point?
   2. What is the molar concentration of CN- at the equivalence point?
   3. What is the pH of the solution at the equivalence point?

**ANSWERS**:

A / A / C / A / C / 6.4 / 4.37 / a. 0.0133L / b. 0.060M / c. 11.0