

Name \_\_\_\_\_ KEY \_\_\_\_\_

Section \_\_\_\_\_

1. (3 pts) Which of the following is TRUE about a weak base/strong acid titration?

- a) **The strong acid completely ionizes and the weak base establishes some equilibrium.**
- b) The weak base completely ionizes and the strong acid establishes some equilibrium.
- c) The pH of the equivalence point is greater than 7 (basic.)
- d) There will be less  $H^+$  present at the equivalence points.

2. (3 pts) Match each titration with its titration curve.

- a) Strong Acid/Strong Base Titration   2      pH = 7.0 at Eq Pt.
- b) Weak Acid/Strong Base Titration   1      pH = 9.0 (basic) at Eq. Pt.
- c) Strong Acid/Weak Base Titration   3      pH = 5.0 (acidic) at Eq. Pt.

3. (2 pts) Based on the information of **GRAPH 1** above, choose the proper indicator.Indicator Name:   Phenolphthalein  

| Indicator         | Color     |               | pH Range* |
|-------------------|-----------|---------------|-----------|
|                   | In Acid   | In Base       |           |
| Thymol blue       | Red       | Yellow        | 1.2–2.8   |
| Bromophenol blue  | Yellow    | Bluish purple | 3.0–4.6   |
| Methyl orange     | Orange    | Yellow        | 3.1–4.4   |
| Methyl red        | Red       | Yellow        | 4.2–6.3   |
| Chlorophenol blue | Yellow    | Red           | 4.8–6.4   |
| Bromothymol blue  | Yellow    | Blue          | 6.0–7.6   |
| Cresol red        | Yellow    | Red           | 7.2–8.8   |
| Phenolphthalein   | Colorless | Reddish pink  | 8.3–10.0  |

\*The pH range is defined as the range over which the indicator changes from the acid color to the base color.

4. (5 pts) Calculate the **pH** of the following solution after adding the LiOH.

Note: the solution will have a total volume of 60.0 mL!!!

**Strong Acid/Strong Base Titration**

Consider titrating 40.0 mL of 0.16 M HBr with 20.0 mL 0.16 M LiOH

$$\text{Moles H}^+ = (0.16 \text{ M})(0.040 \text{ L}) = 0.0064 \text{ moles H}^+$$

$$\text{Moles OH}^- = (0.16 \text{ M})(0.020 \text{ L}) = 0.0032 \text{ moles OH}^-$$

Net Ionic Equation for ANY Strong Acid / Strong Base Titration:



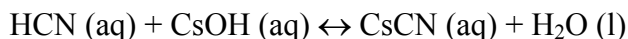
Based on this balanced equation, you have a 1:1 ratio of acid to base. Therefore, when all the base is used up (0.0032 moles), you will still have 0.0032 moles of acid left.

$$[\text{H}^+] = 0.0032 \text{ moles} / 0.060 \text{ L} = 0.053 \text{ M}$$

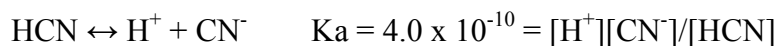
$$\text{pH} = -\log[\text{H}^+] = -\log(0.053) = \mathbf{1.27}$$

**Weak Acid/Strong Base Titration**

Consider titrating 20.0 mL of 0.20 M of HCN ( $K_a = 4.0 \times 10^{-10}$ ) with 0.137 M CsOH



5a. Calculate the pH of the HCN solution BEFORE any base is added.



$$4.0 \times 10^{-10} = x^2 / (0.2 - x)$$

$$x = 8.94 \times 10^{-6}$$

$$[\text{H}^+] = 8.94 \times 10^{-6} \text{ M} \quad \text{pH} = \mathbf{5.05}$$

5b. How many milliliters of CsOH are needed to reach the equivalence point?

At the Eq. Pt, we will have reacted all the moles of Acid with our Base. We calculate the TOTAL MOLES of acid present

$$\text{Moles HCN} = (0.20 \text{ M})(0.02 \text{ L}) = 4.0 \times 10^{-3} \text{ moles HCN}$$

$$4.0 \times 10^{-3} \text{ moles HCN} = \text{moles OH}^- \text{ at Eq. Pt.}$$

$$\text{Volume CsOH} = (4.0 \times 10^{-3} \text{ moles OH}^-) / (0.137 \text{ M}) = 2.9 \times 10^{-2} \text{ L} = \mathbf{29.20 \text{ mL}}$$

5c. Which of the following is a reasonable pH for the equivalence point of this titration?

- a) 7.0   b) 4.0   c) 2.0   d) 1.0   e) **10.0**