

More Stuff about Aqueous Equilibria

3/18/19 (1)

A buffer is made by adding 0.300 mol CH_3COOH and 0.300 mol NaCH_3COO to enough water to make 1.00 L of soln. calculate pH before and after 0.020 mol of NaOH is added. pK_a acetic acid = 4.75

$$\text{Before: } \text{pH} = \text{pK}_a + \frac{\log [\text{A}^-]}{[\text{HA}]}$$

$$\text{pH} = \text{pK}_a + \log (1) = 4.75$$

After: Assume that all added OH^- reacts w/ HA to produce A^- .

	$[\text{CH}_3\text{COOH}]$	$[\text{OH}^-]$	$[\text{CH}_3\text{COO}^-]$
I	0.300	0.200	0.300
C	-0.020	-0.020	+0.020
E	0.280	0	0.320

use Henderson - Hasselbach equation

The ratio of molarity can be calculated using a ratio of moles

$$\text{pH} = \text{pK}_a + \log (n\text{A}^- / n\text{HA})$$

$$\text{pH} = 4.75 + \log (0.320 / 0.280) = 4.80$$

pH went up a little bit

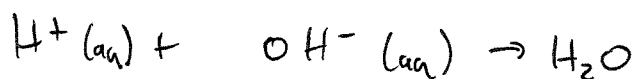
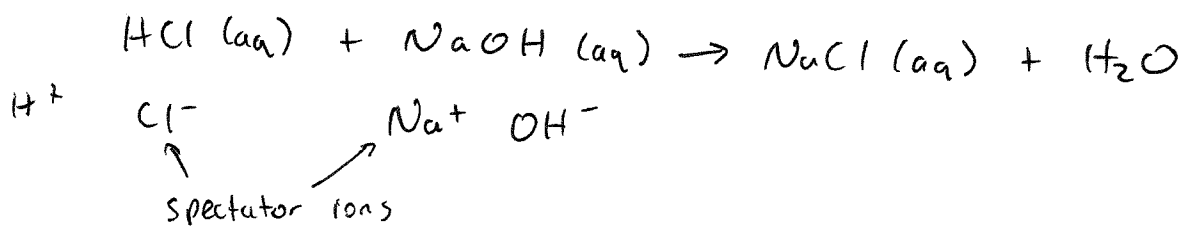
Titration

- The titrant (acid/base of known concentration) is added to analyte (a base/acid unknown concentration)
- An indicator (or pH meter) is used to determine when the soln. has reached the equivalence point, where moles acid = moles base

Titration of SA w/ SB

(2)

50 mL 0.100 M HCl titrated w/ 0.100 mol NaOH. pH during titration?



assume complete reaction for SA w/ SB

Initial (No NaOH added):

$$0.1 \frac{\text{mol}}{\text{L}} \text{HCl} \Rightarrow 0.1 \frac{\text{mol}}{\text{L}} \text{H}^+$$

$$\text{pH} = -\log \text{H}^+ = 1$$

After (NaOH added):

$$\frac{0.1 \text{ mol H}^+}{\text{L}} \times 0.050 \text{ L} = 0.005 \text{ mol H}^+$$

$$\frac{0.1 \text{ mol OH}^-}{\text{L}} \times 0.01 \text{ L} = 0.001 \text{ mol OH}^-$$

$$\text{mol H}^+_{\text{eq}} = \text{mol H}^+_{\text{initial}} - \text{mol OH}^-_{\text{added}}$$

$$[\text{H}^+] = \frac{0.004 \text{ mol H}^+}{0.060 \text{ L}} = 0.067 \text{ M}$$
$$[\text{H}^+] = \frac{0.0025 \text{ mol}}{0.075 \text{ L}} = 0.033 \text{ M}$$

$$\text{pH} = -\log (0.067)$$

$$\text{pH} = -\log (0.033) = 1.48$$

Total (25 mL NaOH)

$$\text{mol H}^+_{\text{eq}} = 0.005 \text{ mol} - (0.1 \frac{\text{mol}}{\text{L}} \times 0.025 \text{ L}) = 0.0025 \text{ mol H}^+$$

50 mL of NaOH added

(3)

the equivalence point!

$$\text{mol H}^+ \text{ initial} = \text{mol NaOH added}$$

For SA-SB $\text{pH} = 7$

100 mL of NaOH added

$$\begin{aligned} \text{mol OH}^-_{\text{eq}} &= \text{mol OH}^-_{\text{added}} - \text{mol H}^+_{\text{initial}} \\ &= \left(0.1 \frac{\text{mol}}{\text{L}} \times 0.1 \text{ L}\right) - 0.005 \text{ mol} = 0.005 \text{ mol OH}^- \end{aligned}$$

$$[\text{OH}^-] = \frac{0.005 \text{ mol}}{0.15 \text{ L}} = 0.033 \text{ M}$$

$$\text{pOH} = 1.5 \quad \text{pH} = 12.5$$

- Initially, the pH goes up slowly

- $[\text{H}^+]$ is calculated by subtracting the moles of OH^- added from the initial moles of H^+ and dividing by the total volume

- Just before the equivalence point, the pH rises rapidly

- At the equivalence point, $\text{pH} = 7$

- After the equivalence point, the pH goes up rapidly but then levels off

- $[\text{OH}^-]$ is calculated by subtracting the initial moles of H^+ from the moles of OH^- added and

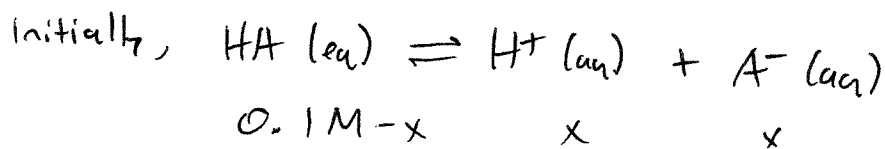
- Titration curve looks similar to the strong acid being titrated by strong base, but it's flipped over

- pH during titration: high pH initially, $\text{pH} = 7$ at equivalence point, low pH at end

Titration WA + SB

(9)

50 mL of 0.1M acetic acid titrated w/ 0.1M NaOH



$$K_a = \frac{x^2}{0.1 - x} = 1.8 \times 10^{-5}$$

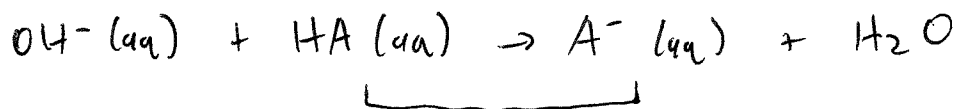
assume $\frac{x^2}{0.1} = 1.8 \times 10^{-5}$

$$x = 1.34 \cdot 10^{-3} [\text{H}^+]$$

$$\text{pH} = 2.87$$

- Initially, only HA is present

After 20 mL NaOH



- Up to equivalence point, both HA and A⁻ are present (buffer region)
- After the equivalence point, only A⁻ is present