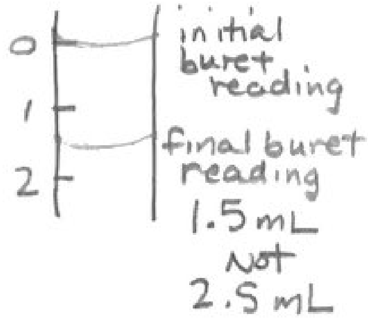
notes from Thurs, Mar. 25

# Titration

One purpose of a titration is ① to find the concentration of a solution
② to find the $K_a$ or $K_b$ of an acid or base (graph)

Acid or base of a Known concentration is gradually added to a given amount of base or acid of an unknown concentration until the equivalence point is reached.

Titrant – • in the buret
• titrate with titrant




Buret – • 50.00 mL
• Know procedures for filling a buret
• tells you the amount delivered

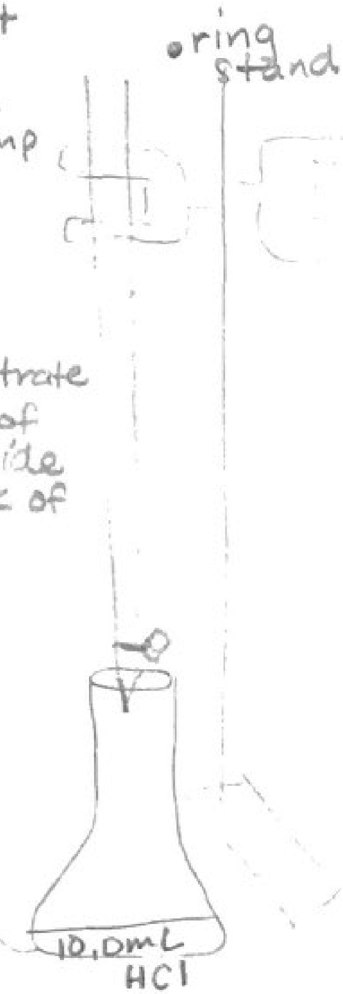
Erlenmeyer flask –  shape → swirl
NEVER used to measure.

Indicator – • chemical (usually a weak acid) that changes color as pH changes.

Equivalence point – (stoichiometric point) [H+] = [OH-]

Endpoint – color change

→ Goal: at same point




colorless → acid
pink → base

10.0 mL of HCl of an unknown concentration are placed in an Erlenmeyer flask along with a few drops of phenolphthalein indicator. The acid is titrated with .10 M NaOH until the endpoint is observed. The ~~original~~ final buret reading was 47.51 mL and the ~~final~~ initial buret reading was 23.45 mL. What is the concentration of the acid?

Always titrate with tip of buret inside the neck of E. flask.

Balanced Equation: $HCl + NaOH \rightarrow NaCl + H_2O$

Method 1: 47.51 mL − 23.45 mL = 24.06 mL

$$\frac{24.06 \text{ mL NaOH}}{1} \times \frac{.10 \text{ mol NaOH}}{1000 \text{ mL}} \times \frac{1 \text{ mol HCl}}{1 \text{ mol NaOH}} = .0024 \text{ mol HCl}$$

$$\frac{.0024 \text{ mol HCl}}{.0100 \text{ L}} = .24 \text{ M HCl}$$

Method 2: $M_aV_a = M_bV_b$

$(M_a)(10.0 \text{ mL}) = (.10 \text{ M})(24.06 \text{ mL})$

$M_a = .24 \text{ M HCl}$

1. NaOH with a concentration of 0.033 M is used to titrate 25.00 mL of $H_2SO_4$ to a phenolphthalein endpoint. The initial buret reading is 38.6 mL and the final buret reading is 16.5 mL. What is the concentration of the acid?

Balanced equation: $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$

Method 1: 38.6 mL − 16.5 mL = 22.1 mL

$$\frac{22.1\,mL\,NaOH}{1} \times \frac{.033\,mol\,NaOH}{1000\,mL\,NaOH} \times \frac{1\,mol\,H_2SO_4}{2\,mol\,NaOH} = \frac{.00036\,mol\,H_2SO_4}{.02500\,L} = .014\,M\,H_2SO_4$$

diprotic acid

Method 2: $(2)M_aV_a = M_bV_b$

$(2)(M_a)(25.00\,mL) = (.033\,M)(22.1\,mL)$

$M_a = .014\,M$

$[H^+] = [OH^-]$

1 mole acid → 2 moles $H^+$ (only need ½ as much acid)

2-4 Home-work!

(2) How many mL of 0.15 M HCl are required to completely neutralize 25.0 mL of 0.20 M $Ca(OH)_2$?

Balanced equation: $2HCl + Ca(OH)_2 \rightarrow CaCl_2 + 2H_2O$

Method 1: $\frac{25.0\,mL\,B}{1} \times \frac{.20\,mol\,B}{1000\,mL\,B} \times \frac{2\,mol\,A}{1\,mol\,B} \times \frac{1000\,mL\,A}{.15\,mol\,A} =$ **66.7 mL HCl**

Method 2: $M_aV_a = (2)M_bV_b$

$(.15)(V_a) = (2)(.20)(25.0)$

$V_a =$ **66.7 mL**

(3) What is the concentration of a solution of nitric acid if 23.0 mL are completely neutralized by 32.00 mL of 0.100 M potassium hydroxide solution?

Balanced equation: $HNO_3 + KOH \rightarrow KNO_3 + H_2O$

Method 1: $\frac{32.00\,mL\,B}{1} \times \frac{.100\,mol\,B}{1000\,mL\,B} \times \frac{1\,mol\,A}{1\,mol\,B} = \frac{.0032\,mol\,HNO_3}{.0230\,L\,HNO_3} =$ **.139 M $HNO_3$**

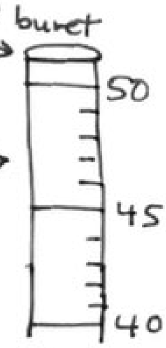
Method 2: $M_aV_a = M_bV_b$

$(M_a)(23.0) = (.100)(32.00)$

$M_a =$ **.139 M $HNO_3$**

(4) 10.0 mL of HCl are titrated to the equivalence point with 0.100 M NaOH. If the concentration of the acid is .075 M and the intial buret reading is 50.00 mL, what will be the final buret reading?

assumes buret has 50.00 at top and zero at bottom.




Balanced equation: $HCl + NaOH \rightarrow NaCl + H_2O$

Method 1: $\frac{10.0\,mL\,A}{1} \times \frac{.075\,mol\,A}{1000\,mL\,A} \times \frac{1\,mol\,B}{1\,mol\,A} \times \frac{1000\,mL\,B}{.100\,mol\,B} = 7.5\,mL$ of NaOH used

Method 2: $M_aV_a = M_bV_b$

$(.075)(10.0) = (.100)(V_b)$

$V_b = 7.5$ mL NaOH used in the titration

50.00 − 7.5 = **42.5 mL**