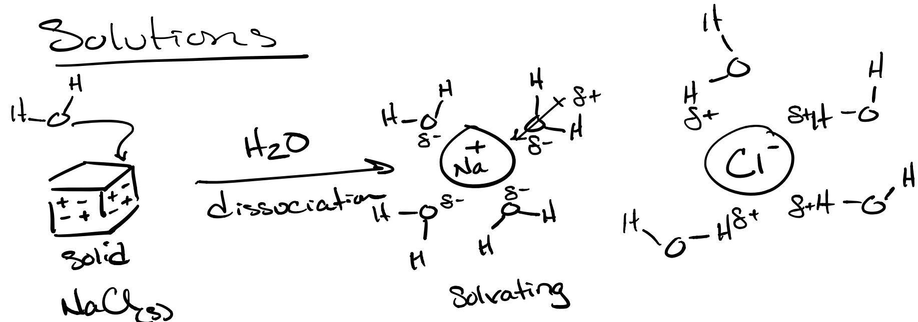
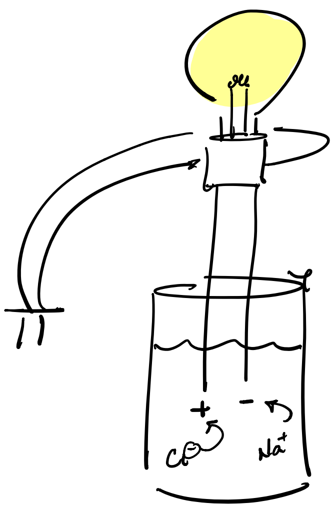
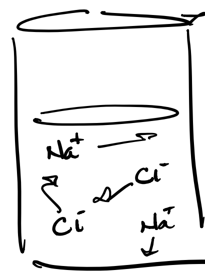
Solutions







Ionic Solution
w/ dissociated ions
Electrolytic Solution

Strong electrolyte = many ions in solution
→ Completely dissociated

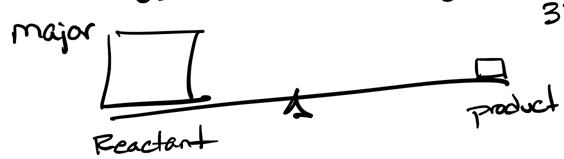
$$NaCl_{(s)} \xrightarrow{H_2O} Na^{+}_{(aq)} + Cl^{-}_{(aq)}$$

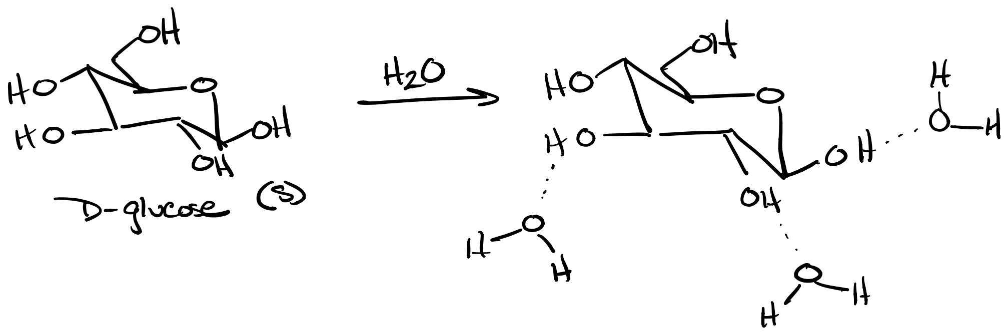
100%

Weak electrolyte = few ions in solution
→ partly dissociated

$$Fe(OH)_{3\,(s)} \overset{H_2O}{\rightleftarrows} Fe^{3+}_{(aq)} + 3OH^{-}_{(aq)}$$

3%





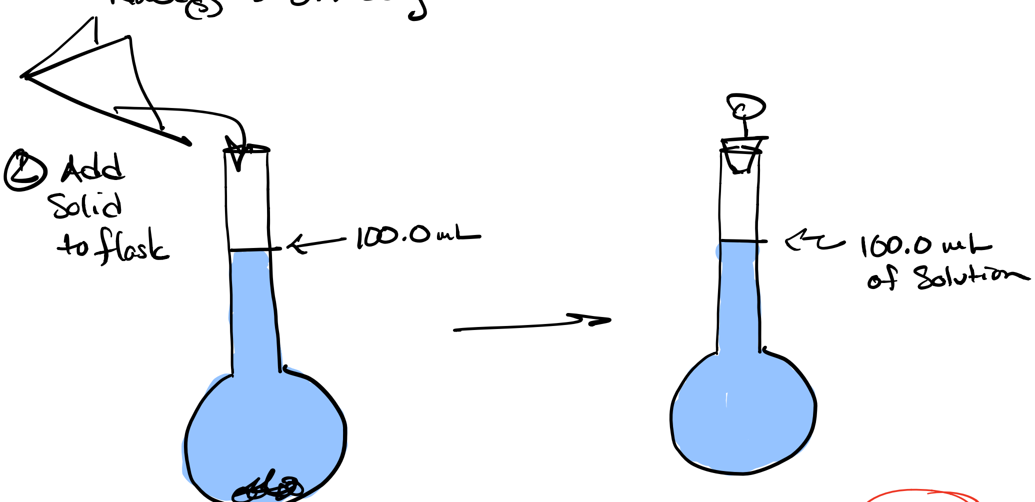
non-electrolyte
Does not conduct electricity
dissolves as a molecule
rather than dissociating

Making a Solution

① Weigh out Solid

$NaCl_{(s)} = 0.236\ g$

② Add Solid to flask



③ Add H_2O to the mark

④ Dissolve

$$\frac{0.236\ \cancel{g\ NaCl}}{100.0\ \cancel{mL\ Sol}} \times \frac{1\ \text{mole NaCl}}{58.44\ \cancel{g\ NaCl}} \times \frac{1000\ \cancel{mL\ Sol}}{1\ L\ sol}$$

$$= 0.0404\ \text{moles NaCl} / L\ sol$$

Molarity $= 0.0404\ mol/L\ NaCl_{(aq)}$

$= 0.0404\ \underline{M}\ NaCl_{(aq)}$

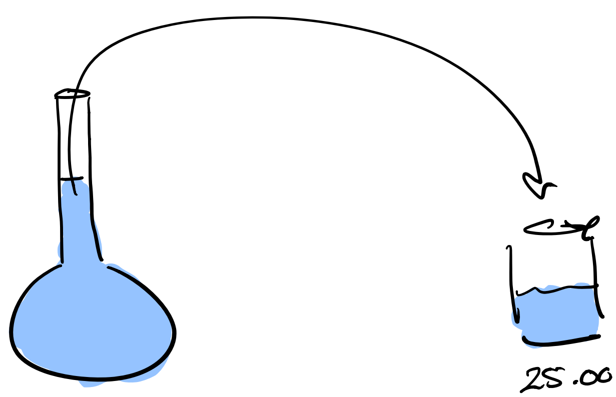
Molarity is a measure of Concentration in moles of solute per liter of Solution

Solute = Material dissolving

Solvent = liquid dissolved in

Solution = Solute + Solvent

How do we use molarity?



0.0404 mols/L NaCl

25.00 mL

How many moles NaCl?

Road Map

mL sol ⟶ L sol ⟶ moles NaCl

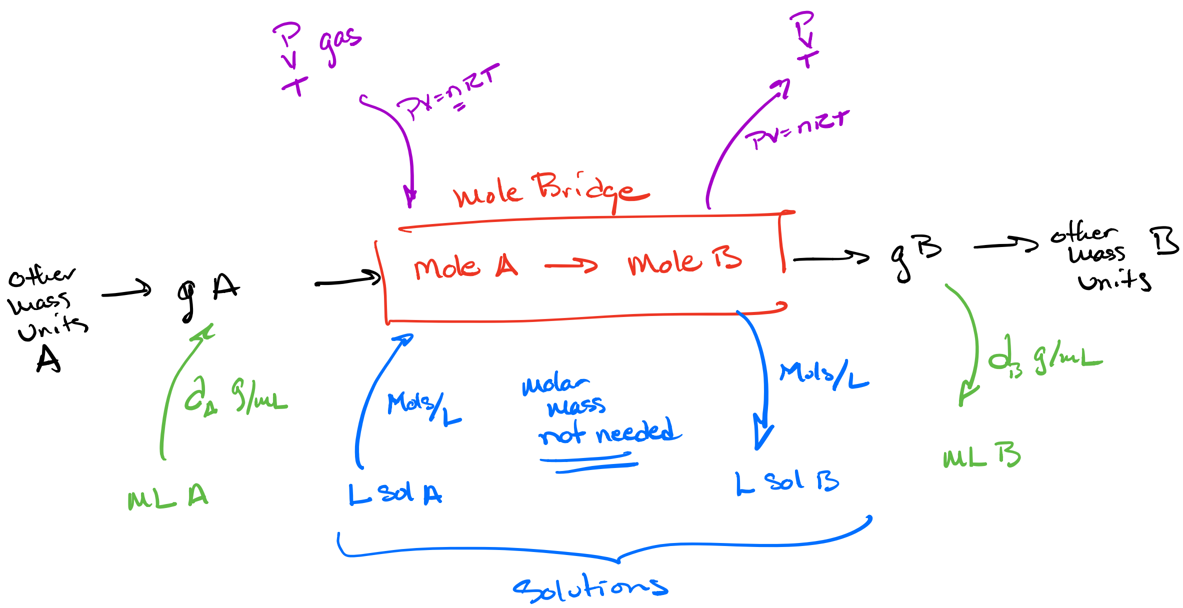
$$25.00\ \cancel{\text{mL sol}} \times \frac{\cancel{\text{L sol}}}{1000\ \cancel{\text{mL sol}}} \times \frac{0.0404\ \text{moles NaCl}}{1\ \cancel{\text{L sol}}} = \boxed{0.00101\ \text{moles NaCl}}$$

What if a reaction requires 0.0200 moles of NaCl. How many mL of 0.0404 M sol will be needed?

moles NaCl $\rightarrow$ L sol $\rightarrow$ mL sol

$$0.0200 \text{ moles } \cancel{\text{NaCl}} \times \frac{1 \cancel{\text{ L sol}}}{0.0404 \cancel{\text{ moles}}} \times \frac{1000 \text{ mL}}{1 \cancel{\text{ L sol}}} = 495 \text{ mL Solution Required}$$

$$\text{moles of Solute} \underset{}{\overset{\text{molarity}}{\rightleftharpoons}} \text{Vol (L) Solution}$$



Acid/Base Reactions $\Rightarrow$ Neutralizations

$$\underset{\text{acid}}{HCl} + \underset{\text{base}}{NaOH} \longrightarrow NaCl + H_2O$$

$Na^+\ Cl^-$ $H^+\ OH^-$

Salt   water

neutral Solution

$\Rightarrow$ no acid $pH = 7$

no base

pH

0 — 7 — 14

Acid — Neutral — Basic

$$H_2SO_4 + 2\,NaOH \longrightarrow Na_2SO_4 + 2\,H_2O$$

H^+ H^+	SO_4^{2-}
2+	2-

Na^+	OH^-

Na^+ Na^+	SO_4^{2-}
2+	2-

H^+	OH^-
+	−

Acid hydrogens always written on left of formula

$\underline{H}Cl$

$\underline{H_2}SO_4$

$\underline{H_3}PO_4$

$\underline{H}C_2H_3O_2$

not acidic

$H_3PO_4 + 3NaOH \longrightarrow Na_3PO_4 + 3H_2O$

$H^+ \quad Na^+OH^- \longrightarrow H_2O$

$H^+ \quad PO_4^{3-} \quad Na^+OH^- \longrightarrow H_2O$ } $3H_2O$

$H^+ \quad Na^+OH^- \longrightarrow H_2O$

HCl	monoprotic	$1\ H^+$	$NaOH$	monobasic	$1\ OH^-$
H_2SO_4	diprotic	$2\ H^+$	$Ca(OH)_2$	dibasic	$2\ OH^-$
H_3PO_4	triprotic	$3\ H^+$	$Fe(OH)_3$	tribasic	$3\ OH^-$

0.026 L of 1.50 mols/L HCl is reacted with 0.025 mols/L NaOH. How many L of Solution are required to react with all of the HCl?

$1HCl + 1NaOH \longrightarrow NaCl + H_2O$

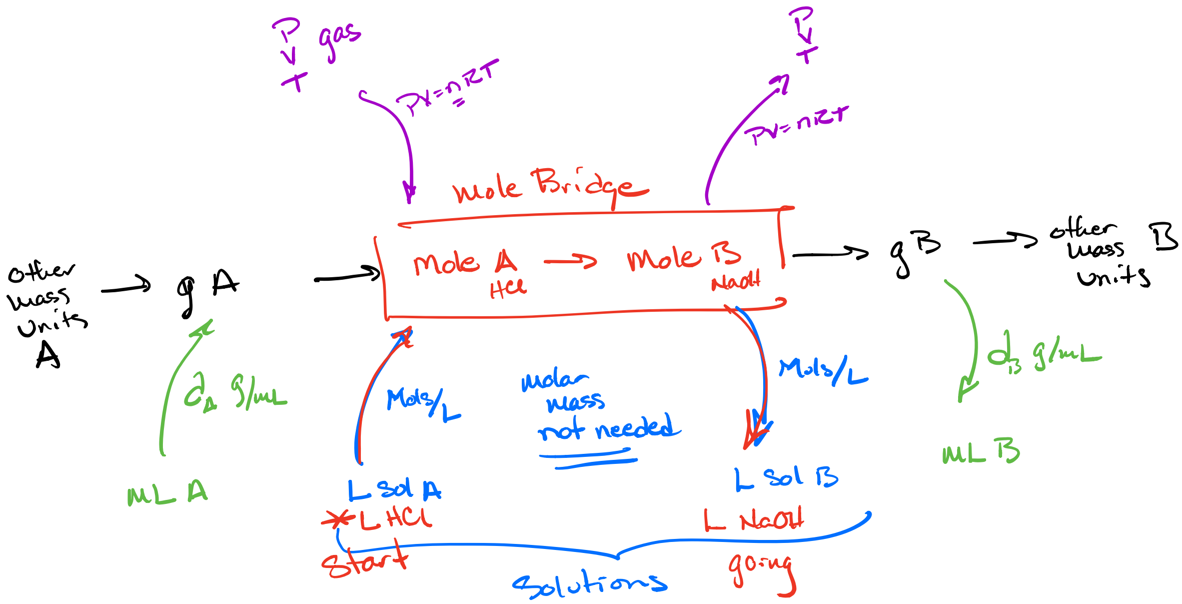
0.026 L sol   ? L

1.50 mols/L   0.025 mols/L

mols HCl $\longrightarrow$ mols NaOH

Road map

$L\ HCl \xrightarrow{\text{mols/L}} \text{moles } HCl \xrightarrow{\text{Coefficients from Eq}} \text{moles } NaOH \xrightarrow{\text{mols/L}} L\ NaOH$



0.026 L of 1.50 mols/L HCl is reacted with 0.025 mols/L NaOH. How many L of Solution are required to react with all of the HCl?

$$1\,HCl + 1\,NaOH \longrightarrow NaCl + H_2O$$

0.026 L sol   ? L

1.50 mols/L   0.025 mols/L

mols HCl → mols NaOH

Road map

$$L\ HCl \xrightarrow{\text{mols/L}} \text{moles HCl} \xrightarrow{\text{Coefficients from Eq}} \text{moles NaOH} \xrightarrow{\text{mols/L}} L\ NaOH$$

$$0.026\,L\,HCl \times \frac{1.50\text{ mols HCl}}{1\text{ L sol}} \times \frac{1\text{ mole NaOH}}{1\text{ mole HCl}} \times \frac{1\text{ L NaOH}}{0.025\text{ mole NaOH}} = 1.56\text{ L NaOH}$$

2   3   2

$$\boxed{1.6\text{ L NaOH}}$$

Titration

use a solution of known molarity to measure the molarity of an unknown solution such as an acid base reaction.

$$NaOH_{(aq)} + HC_2H_3O_{2(aq)} \longrightarrow NaC_2H_3O_{2(aq)} + H_2O$$

⟹ Known Conc.

Sodium hydroxide

Vol of unknown Concentration acetic acid

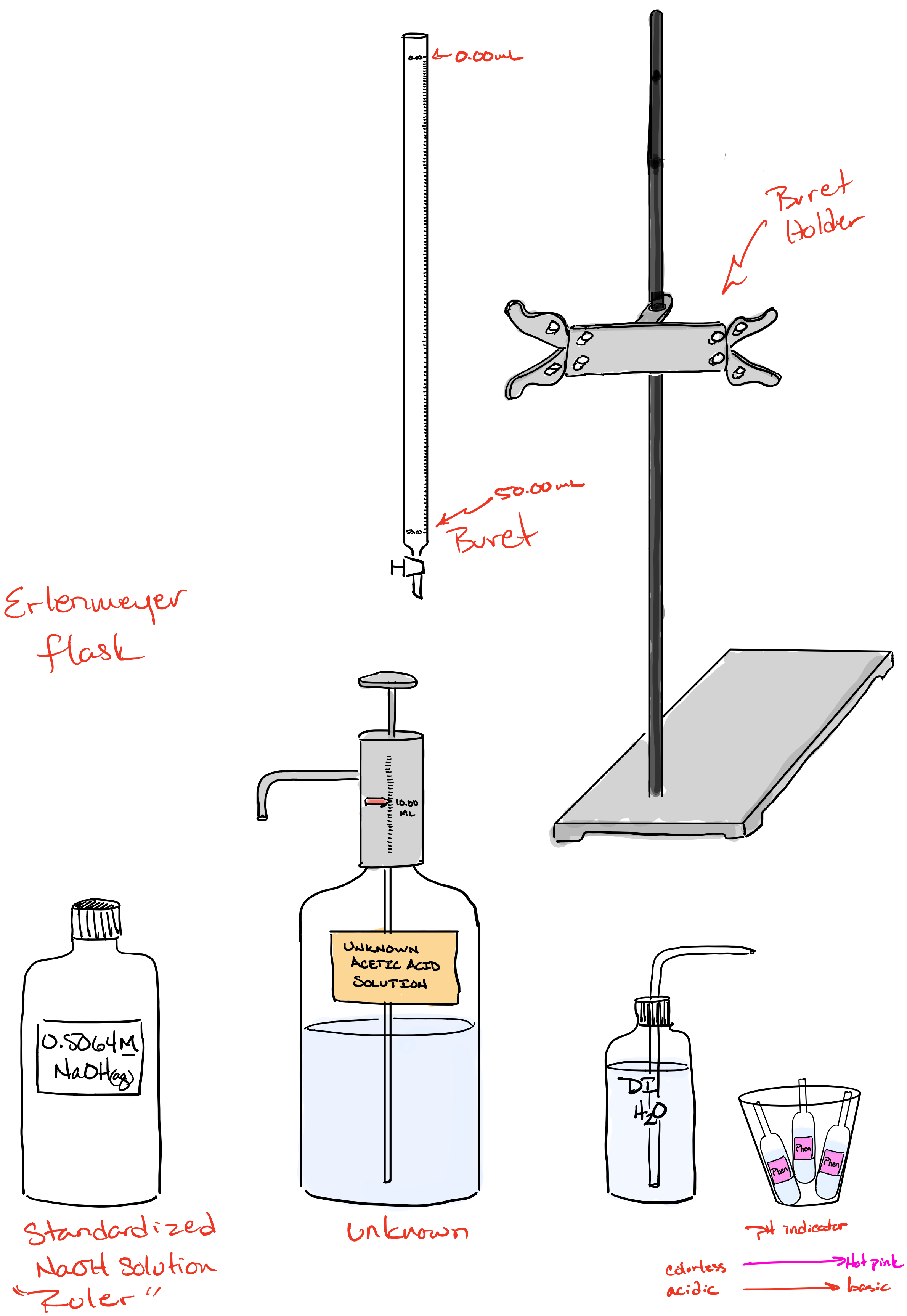
Add NaOH measuring Volume Added

until pH = 7

Neutralization

$$\boxed{\frac{\text{moles } NaOH = \text{moles } HC_2H_3O_2}{\text{Vol Starting}}} = \text{mols}/L = \text{molarity}$$

or Concentration



0.00mL
Buret Holder
50.00mL
Buret
Erlenmeyer flask
10.00 ML
UNKNOWN ACETIC ACID SOLUTION
0.5064M NaOH(aq)
DI H2O
Phen
Standardized NaOH Solution "Ruler"
Unknown
pH indicator
colorless → Hot pink
acidic → basic

① Add 10.00 mL of unknown $HC_2H_3O_2$ into erlenmeyer

② Add DI H_2O to flask ~ 10-20 mL

③ fill buret w/ standardized NaOH solution

④ Add 4 drops phen (indicator)

⑤ measure initial volume on buret

$V_1 = 0.00$ mL

⑥ Add NaOH to flask w/ swirling until end point reached

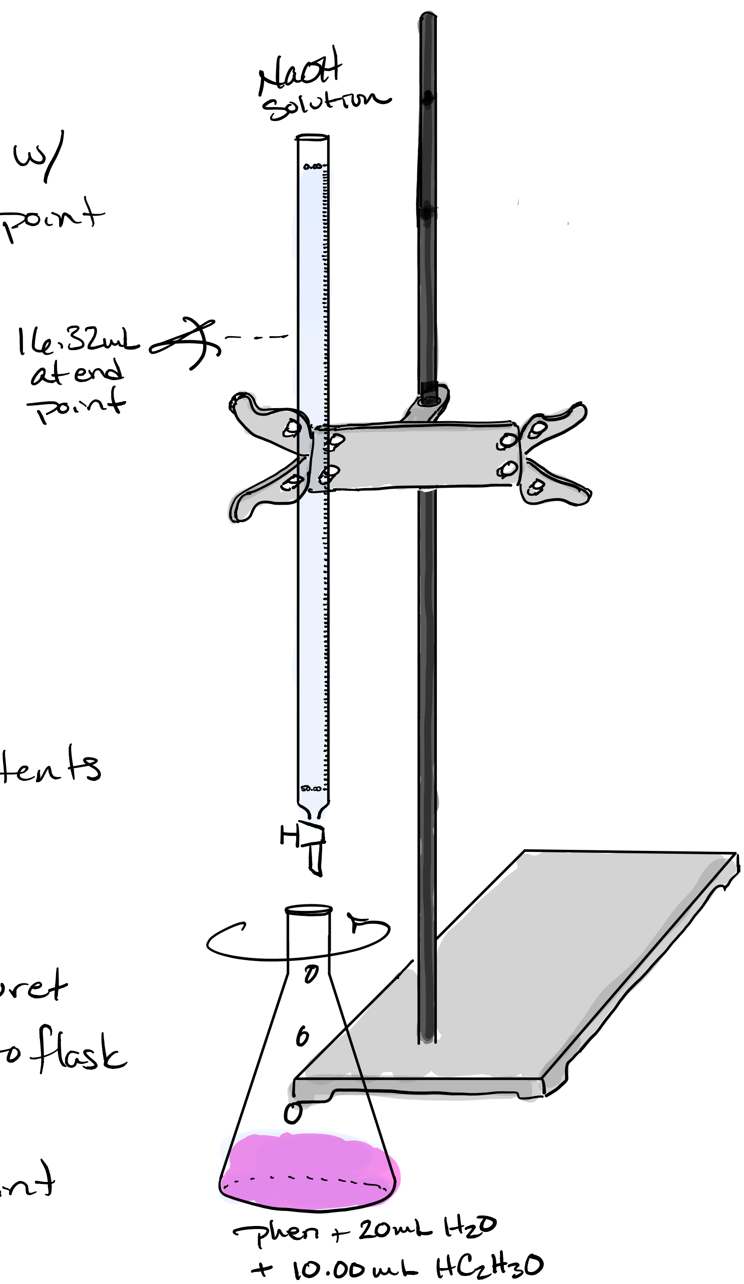
⑦ Read buret for final volume

16.32 mL NaOH

⑧ Reset

Erlenmeyer flask contents goes in (aq) waste

- Rinse w/ H_2O
- Record new V_1 on buret
- Add 10.00 mL acid to flask
- Add 3 drops phen
- Titrate to end point



$$\underline{1}\,NaOH_{(aq)} + \underline{1}\,HC_2H_3O_{2(aq)} \rightarrow NaC_2H_3O_{2(aq)} + H_2O_{(l)}$$

Conc — 0.5064 mols/L — ? mols/L

Vol — $V_f - V_I$ = Vol NaOH — 10.00 mL

16.32mL − 0.00mL

= 16.32mL used

$$\frac{\text{mL NaOH} \xrightarrow{1000\text{mL}=1\text{L}} \text{L NaOH} \xrightarrow{0.5064\text{ mols/L}} \text{moles NaOH} \xrightarrow{1\text{ mol NaOH}=1\text{ mole }HC_2H_3O_2} \text{moles } HC_2H_3O_2}{\text{mL } HC_2H_3O_2 \xrightarrow{1000\text{ mL}=1\text{ L}} \text{L } HC_2H_3O_2} = \text{mols/L}$$

$$\frac{\overset{4}{16.32\text{ mL }\cancel{\text{NaOH sol}}}}{\underset{4}{10.00\text{mL }\cancel{HC_2H_3O\text{ sol}}}} \times \overset{\text{def}}{\frac{1\text{ L }\cancel{\text{NaOH}}}{1000\text{ mL }\cancel{\text{NaOH}}}} \times \overset{4}{\frac{0.5064\text{ moles }\cancel{\text{NaOH}}}{1\text{ L }\cancel{\text{NaOH}}}} \times \overset{\text{counted}}{\frac{1\text{ mole }HC_2H_3O_2}{1\text{ mole }\cancel{\text{NaOH}}}} \times \overset{\text{def}}{\frac{1000\text{ mL }\cancel{HC_2H_3O_2}}{1\text{ L }HC_2H_3O_2}} =$$

0.8264448 moles/L $HC_2H_3O_2$

$$= \boxed{0.8264\text{ mols/L } HC_2H_3O_2}$$

Trial # 1

Chem 3L
05/05/20

Titration Lab Exercise

Acquired Data

Trial 1

A 10.00 mL sample of acetic acid ($HC_2H_3O_2$) is placed in a 100-mL erlenmeyer flask and ~20 mL of deionized water is added to increase the volume. Three drops of phenolphthalein indicator is added to the erlenmeyer flask. A buret is filled with an aqueous 0.5064 mols/L solution of NaOH. The initial reading on the buret is 0.00 mL. Sodium hydroxide is carefully added to the erlenmeyer flask while swirling until the colorless solution turns a faint hint of pink that persists for 1 min. The volume on the buret reads 16.32 mL. Use the data from the experiment to calculate the molarity of the acetic acid solution for trial 1.

Molarity NaOH = 0.5064 mols/L

Volume of acid titrated = 10.00 mL

Initial reading on buret = 0.00 mL

final reading on buret = 16.32 mL

we just did this in class

Trial 2

The contents of the erlenmeyer flask from the first trial are disposed of in the aqueous waste and the flask rinsed with deionized water. A 10.00 mL sample of acetic acid ($HC_2H_3O_2$) is again placed into the erlenmeyer flask and ~15 mL of deionized water is added to increase the volume. Three drops of phenolphthalein indicator is added to the erlenmeyer flask. The initial reading on the buret is 16.32 mL. Sodium hydroxide is carefully added to the erlenmeyer flask while swirling until the colorless solution turns a faint hint of pink that persists for 1 min. The volume on the buret reads 32.30 mL. Use the data from the experiment to calculate the molarity of the acetic acid solution for trial 2.

Molarity NaOH = 0.5064 mols/L

Volume of acid titrated = 10.00 mL

Initial reading on buret = 16.32 mL

final reading on buret = 32.30 mL

vol NaOH

32.30 - 16.32 = 15.98 mL

Trial 3

The contents of the erlenmeyer flask from the second trial are disposed of in the aqueous waste and the flask rinsed with deionized water. A 10.00 mL sample of acetic acid ($HC_2H_3O_2$) is again placed into the erlenmeyer flask and ~25 mL of deionized water is added to increase the volume. Three drops of phenolphthalein indicator is added to the erlenmeyer flask. The initial reading on the buret is 32.30 mL. Sodium hydroxide is carefully added to the erlenmeyer flask while swirling until the colorless solution turns a faint hint of pink that persists for 1 min. The volume on the buret reads 48.94 mL. Use the data from the experiment to calculate the molarity of the acetic acid solution for trail 3.

Molarity NaOH = 0.5064 mols/L

Volume of acid titrated = 10.00 mL

Initial reading on buret = 32.30 mL

final reading on buret = 48.94 mL

Average molarity of three trials

Use the calculated concentrations of the acetic acid solution from trials 1 - 3 to calculate the average acetic acid concentration in mols/liter.

Percent by mass acetic acid in vinegar

The unknown acetic acid solution you have been titrating is actually just consumer strength vinegar solution from the grocery store. Use the average concentration of acetic acid you measured in mols/liter to calculate the percent by mass of acetic acid in the vinegar solution. The calculation is provided below:

$$\text{Molarity } HC_2H_3O_2 = \text{x Mols/L} \quad \text{or} \quad \frac{\text{x Mols } HC_2H_3O_2}{\text{1 L } HC_2H_3O_2}$$

$$\frac{\text{x Mols } HC_2H_3O_2}{\text{1 L solution}} \times \frac{\text{60.05 g } HC_2H_3O_2}{\text{1 Mol } HC_2H_3O_2} \times \frac{\text{1 L solution}}{\text{1000 mL solution}} \times \frac{\text{1.00 g solution}}{\text{1 mL solution}} \times 100 =$$