A titration is used to determine the concentration of an unknown solution using a solution of known concentration. Acids and bases are common applications of this technique. You would do well to check out titrations in the Brown et. al textbook.

- Strong acid + strong base, endpoint is pH 7.0, phenolphthalein indicator good.
- Strong acid + weak base, endpoint is < 7.0, methyl-red usually suitable
- Weak acid + strong base, endpoint is > 7.0, phenolphthalein indicator suitable

You will perform 2 titrations in this lab twice (I hope!) each. First, you will do a titration of a strong base (NaOH) with a strong acid (HCl). Since the solution will go from very basic to very acidic quickly, the slope between when the solution is basic and acidic is very steep - we can use any number of indicators from about pH 5 to pH 10. Phenolphthalein, that changes from pink (basic) to clear at pH 8.2 works quite nicely. The goal of this first titration is to see how a titration works and to verify mathematical (mol-mol) relationships.

The second titration will involve a weak base (ammonia) with a strong acid (HCl). The endpoint for this titration will occur at a low pH (5.28). In this case, the pH will change a bit more slowly and the range at which we can detect a pH change is not as large. We must choose an indicator that switches at an acidic pH – alizarin Red works well as it changes from red (pH 14 to pH 6) to yellow (below pH 4.6). Your goal is to determine the strength of the ammonia solution using mol-mol relationships.

\[
\text{NaOH} \rightarrow \text{Na}^+ + \text{OH}^-; \text{HCl} \rightarrow + \text{H}^+ + \text{Cl}^-; \text{Net Ionic: } \text{OH}^- + \text{H}^+ \rightarrow \text{H}_2\text{O} \quad \text{(Neutral pH) (note, NaCl remains dissolved)}
\]

Strong base dissociates to \( \text{OH}^- \), reacts with the \( \text{H}^+ \) from the A strong acid to make water and salt.

In water, Ammonia makes ammonium hydroxide, \( \text{NH}_3 + \text{H}_2\text{O} \rightarrow \text{NH}_4^+ + \text{OH}^- \); the H from the acid reacts to form water, and ammonium chloride. Actual Reaction: \( \text{H}^+ + \text{Cl}^- + \text{NH}_3 + \text{OH}^- \rightarrow \text{NH}_4^+ + \text{Cl}^- + \text{H}_2\text{O} \); Ammonium Chloride remains dissolved;

Net Ionic: \( \text{NH}_3 + \text{H}^+ \rightarrow \text{NH}_4^+ + \text{H}_2\text{O} \);

Then the ammonium ion dissociates to produce a hydrogen and the solution is acidic. \( \text{NH}_4^+ \rightarrow \text{NH}_3 + \text{H}^+ \) (Acid)

Thus, when all of the hydrochloric acid and all of the ammonia are “used up”, the ammonium ion that resulted controls the pH.

Solutions include: 0.100 M HCl, approximately 0.100 M NaOH, Ammonia solution of unknown molarity.

1) Obtain a burette, ring stand, double burette clamp, 2 Erlenmeyer flasks (250 mL), beaker (200 o 400 mL).
2) Obtain approximately 100 mL of HCl and 25 mL of NaOH in each of the two flasks.
3) Set up the burette, rinse with distilled water well.
4) Using a funnel, rinse the burette with about 5 mL of acid, then fill with acid. Note the top measure.

Precision is important! Take care when measuring and reading the burette!
5) Measure 10.0 mL of the base (NaOH or NH₃) in a graduated cylinder and add to the beaker. Put a drop of the indicator (phenolphthalein for NaOH, alizarin red for NH₃) in the beaker with the base and then place it under the burette.

6) Add acid from the burette in small amounts to the base solution until such time as the appropriate color change occurs. (Pink to clear for phenolphthalein and red to yellow for alizarin red). When you are getting close, add a drop of the titrant at a time! Record the volume of acid used to the nearest hundredth if possible
   a) Note: if you need more than the burette will hold – fill it before you go below the bottom line…and record how much you used and added each time!
7) Repeat the above procedure a second time for each base.

Results: Remember: Moles = Molarity X Volume (in liters); Molarity = moles/liters

<table>
<thead>
<tr>
<th>NaOH Sample</th>
<th>Initial Burette Volume (ml)</th>
<th>Final Burette Volume (ml)</th>
<th>HCl added (ml)</th>
<th>Moles HCl added</th>
<th>Moles of NaOH neutralized</th>
<th>Actual Molarity of NaOH</th>
<th>Grams of NaOH in sample</th>
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<tr>
<th>NH₃ Sample</th>
<th>Initial Burette Volume (ml)</th>
<th>Final Burette Volume (ml)</th>
<th>HCl added (ml)</th>
<th>Moles HCl acid added</th>
<th>Moles of NH₃ neutralized</th>
<th>Actual Molarity of NH₃</th>
<th>Grams of NH₃ in sample</th>
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_Precision is important! Take care when measuring and reading the burette!_
Titration – Acid of a Base

Name: ____________________  Partner Name(s): ____________________

Answer the following questions on a separate sheet of paper – **show all work for all calculations.**

1. Write the balanced (molecular) equations that include subscripts for state for both reactions. Then write the net ionic equations with subscripts.

2. Calculate the number of moles of HCl added based on the volume added (record in Table 1) for all trials.
   
   - Moles = volume (in Liters) X molarity. \((HCl = 0.100 \text{ M})\)

3. Calculate the moles of NaOH neutralized (record in Table 1) based on volume of NaOH used.
   
   - Moles = volume (in Liters) X molarity.

4. Calculate the molarity of the NaOH solution
   
   - Molarity = moles NaOH/liters of NaOH

5. Calculate the grams of NaOH neutralized
   
   - Grams = moles X molar mass

6. Calculate the moles of ammonia
   
   - Multiply the moles of HCl used by the mole ratio

7. Calculate the Molarity of the ammonia solution
   
   - Moles ammonia/liters of ammonia solution

8. Calculate the grams of NH₃ neutralized
   
   - Grams = moles X molar mass

9. Some common antacids are CaCO₃, NaHCO₃, Mg(OH)₂, Al(OH)₃.
   
   - (A) Write balanced equations for HCl and each of these substances.
   - (B) Determine how many L of a 0.100 M HCl solution would be needed to neutralize 1.00 g of each of the substances.
     
     \(\text{Calculate moles of the base, then multiply by mole ratio to get moles acid, then } (M = \text{mol/L}) \text{ calculate volume of HCl needed.}\)

**Turn in the table, ALL calculations for the table (1-8), and organized work for number 9.**

*Precision is important! Take care when measuring and reading the burette!*