Using Acid/Base Chemistry: A Quality Control Test

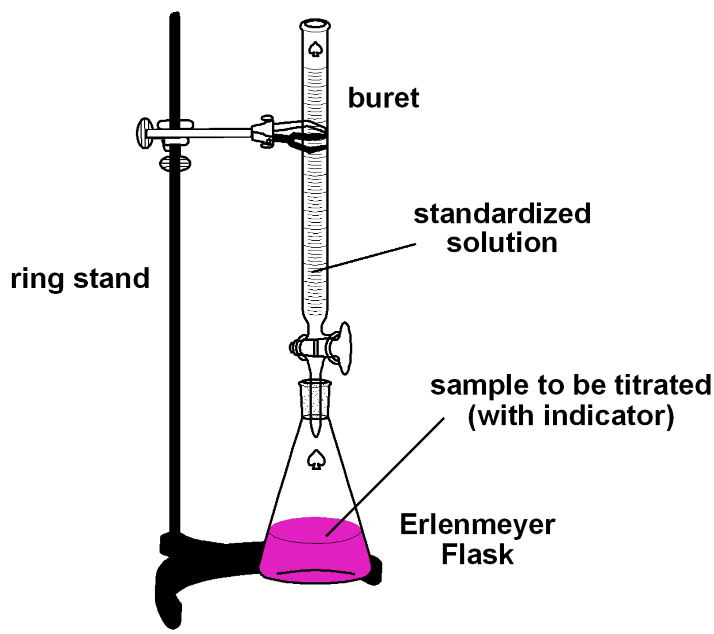
What is a Titration?

A titration is an analytical procedure used to determine the concentration of a sample by reacting it with a standard solution. One type of titration uses a neutralization reaction, in which an acid and a base react to produce a salt and water.

$$HCl\ (aq) + NaOH\ (aq) \longrightarrow NaCl\ (aq) + H_2O\ (l) \qquad \textbf{(1)}$$

In equation 1, the acid is HCl (called hydrochloric acid) and the base is NaOH (called sodium hydroxide). When the acid and base react, they form NaCl (sodium chloride), which is also known as table salt. The titration proceeds until the equivalence point is reached, where the number of moles of acid is equal to the number of moles of base. This point is usually marked by observing a color change in an added indicator.

In a titration, the standard solution goes in a buret, which is a piece of glassware used to measure the volume of solvent to approximately 0.1 mL of accuracy. The solution that you are titrating goes in an Erlenmeyer flask, which should be large enough to accommodate both your sample and the standard solution you are adding.



What is an Indicator and What is it Used For?

An indicator is any substance in solution that changes its color as it reacts with either an acid or a base. Selecting the proper indicator is important because each indicator changes its color over a particular range of pH values. Indicators are either weak acids or weak bases. For example, phenolphthalein is a weak acid (which we will represent as HIn). In aqueous solution, the phenolphthalein dissociates slightly, forming an equilibrium.

$$\underset{\text{clear}}{HIn\ (aq) + H_2O\ (l)} \rightleftharpoons \underset{\text{pink}}{H_3O^+\ (aq) + In^-\ (aq)} \qquad \textbf{(2)}$$

An equilibrium occurs when the amount of reactants and the amount of products are constant. When a system is in equilibrium, it will stay there until something changes the conditions. A famous French chemist, named Le Chatelier, developed a way to predict how changes in equilibrium affect the system. Le Chatelier's principle states that when an equilibrium is disturbed by applying stress, the equilibrium will shift to relieve the stress. **In an acidic solution**, there is an excess of H_3O^+ ions so the equilibrium will shift to the left and favor the formation of HIn, thus we **observe a clear solution**. **In basic solution**, there is an excess of OH^- ions that react with the H_3O^+ ions to form water. This shifts the equilibrium to the right because water is being formed and H_3O^+ ions are being removed, thus **we observe a pink solution**. We can use this color change to determine when the end of the titration has been achieved.

Table 1 lists common indicators and the pH range over which they change colors.

Table 1: Table of Indicators

Indicator	**Color change interval (pH)**	**Acid**	**Base**
thymol blue	1.2 - 2.8	red	yellow
methyl orange	3.1 - 4.4	red	yellow
methyl red	4.4 - 6.2	red	yellow
chlorophenol red	5.4 - 6.8	yellow	Red
bromothymol blue	6.2 -7.6	yellow	Blue
phenol red	6.4 - 8.0	yellow	Red
thymol blue	8.0 - 9.6	yellow	Blue
Phenolphthalein	8.0 - 10.0	colorless	Red
alizarin yellow	10.0 -12.0	yellow	green

Measuring pH: How to Calibrate a pH Meter

pH is a measure of acidity or basicity. An acid has a pH less than 7, a neutral compound (like water) has a pH near 7, and a base has a pH from 7-14. pH can be measured using either litmus (or indicator) paper, which changes color based on the acidity of a solution, or by using a pH meter. A pH meter is a more accurate means of measuring pH because it can be calibrated to measure one tenth of a pH unit, whereas the indicator paper only measures to one pH unit.

A pH meter uses an electrode to measure the pH of a solution. The electrode is stored in distilled water in order to keep it at a neutral pH.

To calibrate the pH meter:

1. Remove the electrode from the distilled water and place it in pH 4 buffer, which is pink. Make sure the electrode is completely covered in buffer and swirl the solution around.

2. Set the pH meter to pH 4 and then rinse the electrode with distilled water to remove any excess solution.
3. Place the electrode in pH 10 buffer (which is blue) and swirl it around in the solution.
4. Set the pH meter to pH 10 and rinse the electrode, returning it to the distilled water once you are finished.

The pH meter should now be calibrated to measure any pH accurately.

Standardizing a Sodium Hydroxide (NaOH) Solution

In a titration, it is critical to know the exact concentration of the titrant (the solution in the buret which will be added to the unknown) in order to determine the concentration of the solution being tested. We will standardize the ~0.1 M NaOH solution (the titrant) with potassium hydrogen phthalate (KHP, $KC_8H_4O_4H$) using phenolphthalein as the indicator. KHP is a weak acid and reacts with base in the following way:

$$C_8H_4O_4H^- + OH^- \longrightarrow C_8H_4O_4^{2-} + H_2O\ (l) \qquad \textbf{(3)}$$

To Standardize:

1. Weigh ~0.8 g of dried KHP (MW = 204.23 g/mol) into an Erlenmeyer flask and dissolve in 50-75 mL of distilled water. Record the amount of KHP and water used.
2. Add 4 drops of indicator into the flask and titrate to the first permanent appearance of pink. Near the endpoint, add the NaOH dropwise to determine the total volume most accurately.
3. Calculate the concentration of NaOH in the following way:

   Calculate Concentration of KHP:

   Calculate C

$$\_\_\ \text{g KHP} \left(\frac{\text{1 mol KHP}}{\text{204.23 g}}\right) = \_\_\ \text{moles KHP}$$

$$\_\_\ \text{moles KHP} \underbrace{\left(\frac{\text{1 mol NaOH}}{\text{1 mol KHP}}\right)}_{\text{from balanced eqn}} = \_\_\ \text{moles NaOH}$$

$$\frac{\_\_\ \text{moles NaOH}}{\_\_\ \text{L NaOH used}} = \_\_\ \text{M (mol/L) NaOH}$$

   Remember: There are 1000 mL in a L and 1000 mg in a gram.

4. Report the concentration of NaOH to the class. An average number will be determined to give the most reliable value of NaOH concentration. Do not discard the remaining NaOH – you will use this for the rest of these experiments.

Standardizing an HCl Solution

In a titration, it is critical to know the exact concentration of the titrant (the solution in the buret which will be added to the unknown) in order to determine the concentration of solutions being tested. We will standardize the ~0.1 M HCl solution (the titrant) with sodium carbonate (Na_2CO_3) using phenolphthalein as the indicator. Na_2CO_3 is a base and reacts with the strong acid HCl in the following way:

$$2\ HCl\ (aq) + Na_2CO_3\ (aq) \longrightarrow NaCl\ (aq) + H_2O\ (l) + CO_2\ (g) \qquad \mathbf{(4)}$$

To Standardize:

1. Weigh ~0.2 g Na_2CO_3 into an Erlenmeyer flask and dissolve it in 50 mL of boiled, cooled distilled water. Record the exact amount of Na_2CO_3 used in your notebook. (The water is boiled to expel CO_2 from the solution.)
2. Add 4 drops of phenolphthalein to the solution and record the color.
3. Titrate with the HCl until just before the endpoint (when the solution is very light pink) and then gently boil the solution to expel the CO_2 from solution that has been produced during the reaction (see eq 4).
4. Cool the solution to room temperature and then wash the sides of the flask with a small amount of H_2O to get all of the sample back into solution.
5. Finish the titration (this will take VERY little HCl so go slow!)
6. Record the color of the solution and the volume of HCl used.
7. Calculate the concentration of HCl in the following way:

$$\_\_\ g\ Na_2CO_3 \underbrace{\left(\frac{1\ mol\ Na_2CO_3}{105.99\ g\ Na_2CO_3}\right)}_{MW\ Na_2CO_3} = \_\_\ moles\ Na_2CO_3$$

$$\_\_\ moles\ Na_2CO_3 \underbrace{\left(\frac{2\ mol\ HCl}{1\ mol\ Na_2CO_3}\right)}_{\text{from balanced eqn}} = \_\_\ moles\ HCl$$

$$\frac{\_\_\ moles\ HCl}{\_\_\ L\ HCl\ used} = \_\_\ M\ (mol/L)\ HCl$$

This material was developed through the **Cornell Science Inquiry Partnership** program (http://csip.cornell.edu), with support from the National Science Foundation's Graduate Teaching Fellows in K-12 Education (GK-12) program (DGE # 0231913 and # 9979516) and Cornell University. Any opinions, findings, and conclusions or recommendations expressed in this material are those of the author(s) and do not necessarily reflect the views of the NSF.