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# Module 8; Titration

## Introduction

*Titration* is the process wherein a known volume of one solution is completely reacted by adding a carefully measured second solution (the *titrant*). Typically, when the first solution is completely reacted, the mixture will undergo a visible change to indicate the *endpoint* (complete reaction) has been reached. Titrations are used to determine the concentration of an unknown solution.

Although titrations are possible with any type of reaction that produces a visible change, *acid-base titrations* are most common. As the name implies, an acid-base titration is one in which carefully measured volumes of acid and base solutions are used to neutralize each other. A small amount of indicator (a compound that changes color in an acidic or basic environment) is added to signal the endpoint. In this experiment, you will use an HCl solution (acid) to determine the concentration of a NaOH solution (base). The neutralization reaction of HCl and NaOH is:

$$HCl + NaOH \rightarrow NaCl + H_2O$$

You will be given a HCl solution whose concentration is unknown. This will be poured into a burette, a device capable of dispensing precise volumes of liquid (see illustration). A known amount of NaOH solution will be placed in an Erlenmeyer flask using a pipet. A drop or two of phenolphthalein solution will be added. Phenolphthalein is an acid-base indicator that is colorless in acidic solution and pink in basic solution. You will carefully add NaOH to the HCl. When the HCl solution just turns pink, you will have completely neutralized the HCl. Then you will note the volume of the NaOH solution you added.

The concentration of the NaOH will be in molarity (mol NaOH / L of solution). Therefore, the number of moles of NaOH added will be:

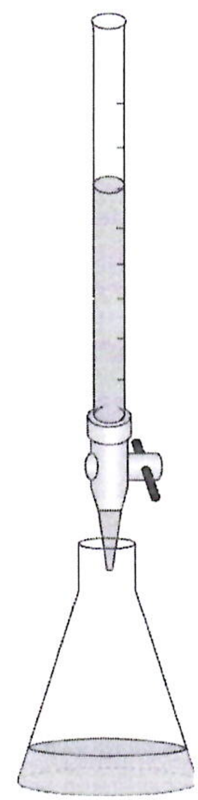
$$\text{mol NaOH} = \text{L sol'n} \times \frac{\text{mol NaOH}}{\text{1 L NaOH}} = \text{mol NaOH}$$

The stoichiometry of the reaction tells you that:

$$\text{mol NaOH} \times \frac{\text{1 mol HCl}}{\text{1 mol NaOH}} = \text{mol HCl}$$

Since you know the volume of HCl solution added, you can then calculate the molarity of the solution:

$$\text{Molarity of HCl sol'n} = \frac{\text{mol HCl}}{\text{L HCl sol'n}}$$



**The titration set-up.** The solution of known concentration (the titrant) is placed in the burette. A known volume of the solution with unknown concentration is placed in the Erlymeyer flask. (The simulator you will be using does it the opposite way; the unknown is placed in the burette. That works too.)

## Pre-Laboratory Questions

Consider the following titration: 0.0250 L of HCl are placed in an Erlenmeyer Flask. 40.05 mL of 0.250M NaOH are added from the buret to reach the end point. What is the concentration of the HCl in the Erlenmeyer Flask?

a. First, figure out how many moles of NaOH were added to the flask. (Start with 40.05 mL. Remember that 0.250 M means 0.250 moles / 1 L.)

| Show your calculation. Use units with numbers. Report your answer with three sig. figs. |
|---|
| $0.250\ \text{mol/L} \cdot \left(40.05\ \text{mL} \cdot \frac{1\ \text{L}}{1000\ \text{mL}}\right) = 1.00 \times 10^{-2}$ mol NaOH |

b. Then figure out how many moles of HCl are in the Erlenmeyer flask. Use the balanced equation mole relationship from the introduction of this lab.

| Show your calculation. Use units with numbers. Report your answer with three sig. figs. |
|---|
| 1 mol HCl : 1 mol NaOH<br>$0.280\ \text{mol/L} \cdot 0.0250\ \text{L} = 0.00625$ mol HCl |

c. Finally, calculate the concentration of HCl in molarity. Use the moles you calculated for HCl (b.) and use 0.0250 L since that is the volume of HCl. Remember that M = mol/L or: molarity is equal to moles ÷ L.

| Show your calculation. Use units with numbers. Report your answer with three sig. figs. |
|---|
| molarity HCl $= \frac{0.00625\ \text{mol}}{0.0280\ \text{L}}$<br>$= 0.25\ \text{mol/L}$ |

## Procedures and Data

Use the following link to watch the titration procedure carefully. It is a three minute video with music only (no speaking). I suggest you use a split screen so that you can follow directions while recording data for the experiment.
https://www.youtube.com/watch?v=8UiuE7Xx5l8

Fill in this table:

1) You will be using two pieces of glassware in this experiment. Name them (refer to introduction if needed).

| | |
|---|---|
| The "top" glassware | Burette |
| The "bottom" glassware | Conical Flask |

2) Answer these questions regarding the titration experiment

| | |
|---|---|
| What color* is the base, NaOH, in the buret? | Colorless/Clear |
| What color is the acid, HCl, in the Erlenmeyer flask (or conical flask)? | Colorless/Clear |
| What is the name of the indicator used? | Phenolphthalein |
| What does a "pink" color mean when the indicator is used? | The solution is above pH 10 (very basic) |
| Why does the pink color disappear when indicator is added? | The pH of the current substance is under 10 pH. |
| What does it mean at the end point when the color no longer disappears? | The pH of the final solution is above 10 pH (very basic) |

*Remember: "clear" is not a color.

**Data and Calculations**

Now you can calculate the concentration of your unknown Acid: You will need three pieces of data from your experiment in the video. Fill them in the table here:

| | |
|---|---|
| 1) Molarity of Base in **M** (take a look at the small beaker from which the presenter filled the burette). | 0.10 M |
| 2) Volume of Acid in mL (take a look at the Erlenmeyer flask/Conical flask). Recall that 1cm$^3$ = 1 mL | 25 mL |
| 3) Volume of base added in mL (You learned about uncertainty in measurement in Module 1). Apply that knowledge to this question. The presenter in the video missed reading the uncertainty (the last estimated digit between the minor lines on the burette). He indicated the uncertainty in the burette to one decimal place (as seen at the end of the video), which is incorrect. Using good chemistry measurement skills, provide the correct value with all significant figures and the unit for the base volume after titration is complete. | 25.39 mL |

Now do your calculations. Show your work here:

| | | |
|---|---|---|
| Liters of base | Hint: You know how many mL of base was used (from #3 above). Recall that 1000 mL = 1 L. Do a conversion starting with mL base. | $25.39\,mL \cdot \frac{1\,L}{1000\,mL} = 2.539 \times 10^{-2}\,L$ |
| moles of base | Hint: Start with the L of base you just found. Use the molarity (from #1 above) as a conversion factor. (Recall M = moles/L) | $0.10\,mol/L \cdot 2.539 \times 10^{-2}\,L = 2.539 \times 10^{-3}\,mol$ |
| moles of acid | Hint: This is a conversion. If you look at the balanced equation in the introduction, you will see that 1 mole NaOH =1 mole HCl | 1 mol NaOH = 1 mol HCl<br>$2.539 \times 10^{-3}\,mol\,NaOH \times \frac{1\,mol\,HCl}{1\,mol\,NaOH} = 2.539 \times 10^{-3}\,mol\,HCl$ |
| L of acid | Hint: You recorded the volume of acid above in #2. Convert that to L. | $25\,mL \cdot \frac{1\,L}{1000\,mL} = 2.5 \times 10^{-2}\,L$ |
| Molarity of unknown acid. | Hint: Molarity is defined as moles/L. You've calculated moles of acid and L of acid. Now calculate molarity of acid. Show work with units. | $M = \frac{mol}{L} = \frac{2.539 \times 10^{-3}\,mol\,HCl}{2.5 \times 10^{-2}\,L\,HCl}$<br>$= 0.1\,mol/L\,HCl$ |

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# Module 8; pH Lab

## Data and Procedures

Use the following link to access an online simulation. I suggest you use a split screen so that you can follow directions while running the experiment.
https://phet.colorado.edu/sims/html/acid-base-solutions/latest/acid-base-solutions_en.html

## Part 1: pH values

1. Click "Introduction" to begin.

2. The lab has 2 tools that allow you to test for pH values: A probe, and pH paper. Use each one by dipping it into the solution to be tested. Try all the given types of solutions and fill in the Data Chart with the pH value 0-14.

3. The circuit with a battery and bulb as shown: is the tool used to test for conduction of a solution. By dipping the wire leads into the solution, the bulb with either **remain unlit**, be **dimly lit**, be **somewhat bright** or **very bright**. Test each solution and record your observation for the bulbs brightness in the chart below.

| **Part 1: Data** | pH Value from Probe | Color & pH Value from pH Paper | Observations from Circuit Tool<br>Describe the brightness |
|---|---|---|---|
| Water ($H_2O$) | 7.00 | Light Green 7 pH | Non-conductive No brightness |
| Strong Acid (*HA*) | 2.00 | Bright orange 2 pH | Incredibly Bright |
| Weak Acid (*A*) | 4.50 | yellow-ish green (4-5) | somewhat conductive less bright |
| Strong Base (*MOH*) | 12.00 | Navy Blue (12 pH) | Incredibly Bright/ Very conductive |
| Weak Base (*B*) | 9.50 | Blueish Green (9-10) | somewhat conductive less bright |

## Part 2: Concentration of $H_3O^+$

1. Choose "My Solution" on the bottom toolbar.
2. Follow the directions below and record your data in the table.

| | |
|---|---|
| Put the pH probe in the beaker. Choose acid and toggle from **"stronger to weaker"** to find the pH value range for acids. Repeat the procedure for bases. | |
| Record the range for acids: | 2.00 to 6.00 |
| Record the range for bases: | 12.00 to 8.00 |
| Once again, chose acid and observe pH ranges as you toggle across the scale: but this time, toggle across from "**weaker to stronger**" acids. | |
| What happens to the amount of $H_3O^+$ ions: | There are more $H_3O^+$ ions |
| What happens to the pH: | The pH decreases |
| The pH of a solution is the measure of $H_3O^+$ ions. In the acid image, those ions are red. Watch what happens as you **go from lower to higher concentration of an acid**? (Toggle across the initial concentration scale and look at the red $H_3O^+$ ions.) | |
| What happens to the amount of $H_3O^+$ ions: | There are more $H_3O^+$ ions |
| What happens to the pH: | The pH decreases |
| **Conclusion**<br>Write in your own words what will happen if you go from lower concentration to higher concentration of an acid. | As one goes from a lower concentration of the acid to a higher concentration, the acid gets stronger. |
| **Conclusion**<br>Write a sentence relating three terms: pH, $H_3O^+$ ions, acidity: | The pH of an acid/base measures the acidity and basicity of a substance by measuring the number of $H_3O^+$ ions inversely. |

## Part 3: pH calculations:

Look up in textbook, Chapter # 9, you learned how to use these formulas below.

1. pH = - **log $H_3O^+$**
2. $H_3O+$ = inv.log -pH or $[H_3O^+]$ = **$10^{-pH}$**

Using this formula, complete the following table:

| $H_3O^+$ | pH |
|---|---|
| 9.63 X $10^{-6}$ M → | 5.02 |
| $9.77 \times 10^{-7}$ M ← | 6.01 |

To check your work, stay in the "my solutions" page. Make sure the pH probe is in the solution. Adjust the concentration or strength of the acid until you hit the desired pH. Select "Graph from the views. See what concentration is given for $H_3O^+$ ions. If your values in the table are incorrect, try recalculating.