eScience Labs LLC

General Chemistry

Student Manual

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Advanced

Lab 20 Titration Indicators

Lab 20 Titration Indicators

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Introduction

Much like the way the human body indicates change, there are many **chemical indicators** used in chemistry. The indication created by an indicator may vary, but one of the most common indicators is color. Color indicators are halochromic chemicals which reveal the pH of a chemical reaction through color changes. In other words, they provide a visual representation of the amount of hydrogen or hydronium ions which exist in a solution. They are commonly used because they are so easy to view and regulate in an experiment. For example, the chemical litmus (made available as litmus paper) is often used to provide an indication of a solution's pH. Imprecise readings may occur due to the fact that color observations may be subject to individual discrimination and ability. However, color scales can be provided in circumstances which make color differentiation potentially difficult. Additionally, pH meters can be used to determine the numeric pH value of a solution.

Acid-Base Titration Indicators

Color indicators are used in acid-base titrations to reveal the exact amount of acidic and basic solution that is needed to create a neutral solution. This is useful because this allows chemists to determine molar concentration of an unknown solution (provided that the molar concentration of the other solution used in the titration is known). There is a large variety of indicators used in acid-base titrations, all of which change color at a different pH. It is therefore important to select an appropriate indicator for the chemical pair being titrated. Some complex titrations even require multiple indicators for easy visualization.

Figure 1: ***Swimming pool test kits are used as small titration experiments. These kits first use a phenol red indicator to measure the acidity of the pool water. Then, a titrant is slowly added drop by drop until the ideal pH is created. The number of drops added indicates how acidic the pool water was initially.***

End Point and Equivalence Point

The point at which the indicator changes color in titration is known as the **end point** of the titration. This is the moment when color of the indicator changes. However, the

equivalence point is often what a chemist is trying to find (in order to determine molarity). The equivalence point (or, neutralization point) is the point in which the combined acid-base solution reaches a neutral pH. Fortunately, if the correct indicator is chosen, the end point will occur at approximately the same moment as the equivalence point. Skillful technique (slow titration) can provide an accurate depiction of where this point occurs on a titration curve.

Indicator Selection

To better grasp the concept of **indicator selection**, let's take a look at the titration curve in Figure 2. As you can see, the equivalence point (or, midpoint) occurs at a pH of approximately 7. Suppose Indicator A changes color (i.e., hits the end point) around a pH of 12 - 13. This would not be a good indicator for this titration because the end point is so far above from the equivalence point on the pH scale. Indicator C changes color around a pH of 3 - 4. This is also not a good indicator for this titration because the end point is so far below the equivalence point. Indicator B changes color around a pH of 7. This is a good indicator for this titration because the color change (end point) takes place at nearly the same pH as the equivalence point.

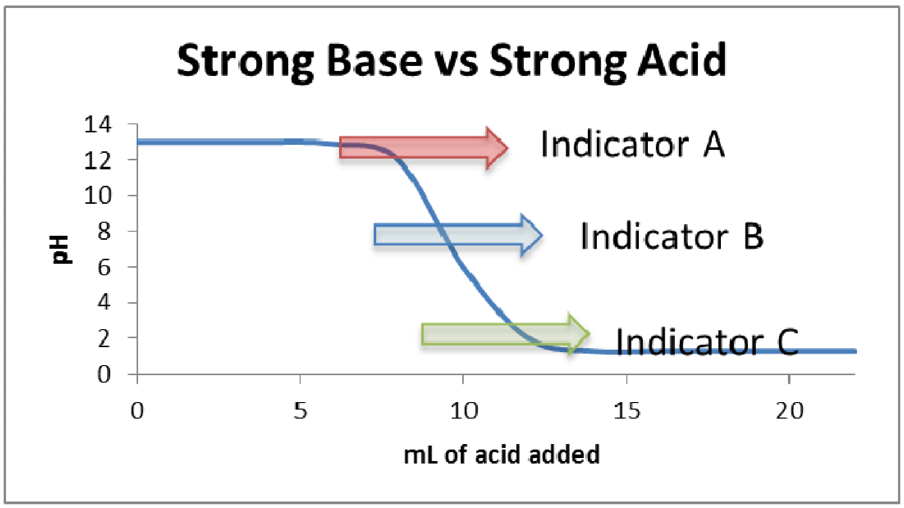

***Figure 2:* Indicators in Acid-Base Titration**

Types of Indicators

Reliable information exists to inform scientists of the different qualities per indicator. This information, along with the types of chemical solutions being titrated and the known molar concentration of one of the solutions, helps scientists select the correct indicator for an experiment. Table 1 lists some of the most common indicators.

Table 1: Common Titration Indicators			
	pH Range	Base Color (High on the pH Scale)	Acid Color (Low on the pH Scale)
Methyl Yellow	**2.9 - 4.0**	**Yellow**	**Red**
Methyl Orange	**3.2 - 4.4**	**Yellow**	**Red**
Congo Red	**3.0 - 4.6**	**Red**	**Blue-Violet**
Bromothymol Blue	**6.0 - 7.6**	**Blue**	**Yellow**
Phenol Red	**6.8 - 8.4**	**Red**	**Yellow**
Phenolphthalein	**8.3 - 10.0**	**Fuschia**	**Colorless**

Indicator Example

Suppose your teacher gives you an assignment in which you need to determine the molar concentration of a hydrochloride solution (acidic). You choose to do this through an acid-base titration using a 0.1M sodium hydroxide primary standard solution. To begin, you place the solution with the unknown molarity in a flask or beaker below the buret or syringe. Then, you must add the indicator. Based on your knowledge of hydrochloride and sodium hydroxide, you select phenolphthalein as your indicator and add a few drops of the indicator to the hydrochloride solution. Because hydrochloride is acidic, the phenolphthalein is in a low pH environment and it will remain in its colorless state.

Make sure the stopcock is closed, pour the sodium hydroxide into a buret, and begin to slowly titrate it into the hydrochloride by opening the stopcock. You continue in this fashion, looking for a pink-fuschia color change. As soon as you see a faint trace of pink appear in the hydrochloride solution. Your reaction has reached the endpoint and the acid has been neutralized. You can then check on the amount of sodium hydroxide required to neutralize the solution, and use stoichiometry to determine the molarity using the equation:

$$N_a \times M_a \times V_a = N_b \times M_b \times V_b$$

Where:

- $\mathbf{N_a}$ = Moles of acid
- $\mathbf{N_b}$ = Moles of base
- $\mathbf{M_a}$ = Molarity of the acid
- $\mathbf{M_b}$ = Molarity of the base
- $\mathbf{V_a}$ = Volume of the acid
- $\mathbf{V_b}$ = Volume of the base

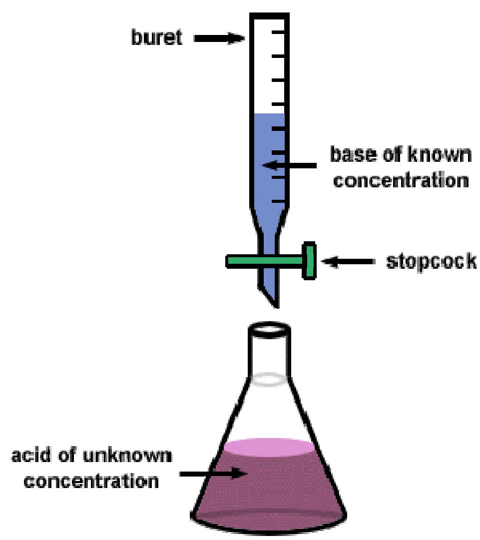

***Figure 3:* Sample acid-base titration set-up**

Lab 20 Titration Indicators

Pre-Lab Questions

1. What is the difference between an end point and an equivalence point?

2. Table 1 lists several indicators that are commonly used along with their equivalence points and the color change that accompanies the reaction. Review the information, and answer the questions below.

    a. Suppose a titration was performed in which a base of pH 6 was being titrated. The equivalence point of the titration was at pH near 8. What indicators should be added to the base solution before the titration is carried out?

    b. What is the color of the basic solution, before any acid is added?

    c. What is the color of the solution after the equivalence point is reached at a pH of 8?

Materials

3 mL Turmeric indicator

(4) 100 mL Beaker

Funnel

Permanent Marker

100 mL Sudsy ammonia

20 Drops of Bromothymol Blue

20 Drops of Methyl Orange

30 mL Syringe

Syringe stopcock

100 mL 0.5 M Citric Acid, $C_6H_8O_7$

pH meter

Ring Stand

Ring clamp

100 mL Graduated cylinder

Stir rod

1 Pipette

*30 mL Distilled water

*You must provide

Experiment 1: Getting Acquainted with Indicators

Procedure

Part 1

1. Set up three 100 mL beakers. Use the permanent marker to label the beakers 1 - 3.
2. Use the graduated cylinder to measure and pour approximately 20 mL of sudsy ammonia into each of the beakers. When measuring the sudsy ammonia, it is best to slightly tilt the graduated cylinder for the majority of the time while pouring the sudsy ammonia to avoid bubble formation. Be sure to return the graduated cylinder to a vertical position when you near the end of the pour to obtain an accurate measurement.
3. Transfer 5 drops of bromothymol blue into Beaker 1. Use a stir rod to mix the indicator into the solution. This should create a blue solution.
4. Pipette 0.5 mL of the turmeric indicator into Beaker 2. Rinse the stir rod and use it to mix the indicator into solution. This should create a red or amber colored solution.
5. Transfer 5 drops of methyl orange into Beaker 3. Rinse the stir rod and use it to mix the indicator into solution. This should create an orange solution.
6. Record the color of each beaker in Table 2.
7. Fasten the stopcock onto the threaded tip of the syringe. Make sure the stopcock is closed by rotating the handle on the stopcock to a position that is perpendicular to the syringe.
8. Connect the ring clamp onto the ring stand, and secure the syringe into the clamp.
9. Fill the syringe with 20 mL of 0.5 M citric acid.
10. Place Beaker 1 under the syringe.
11. Remove the cap from the pH meter, and turn the pH meter on by moving the switch at the top to the right. Rinse the pH meter probe with approximately 10 mL distilled water.
12. Place the pH meter into Beaker 1 and record the initial pH of the solution in Table 2. The pH meter reading may take up to 30 seconds to stabilize.
13. Slowly open the stopcock by rotating the handle to a position that is parallel to the syringe. Slowly add the citric acid from the syringe , while continually swirling the solution in Beaker 1. Pay special attention to the **color** and **pH** of the solution.

14. As you notice the color of the solution begin to change, close the stopcock and take another pH measurement. When you re-open the stopcock, make sure the citric acid flow is slow; drop-wise flow is ideal.

15. Continue this process until the color has completely changed and the pH has leveled out. Record the final pH in Table 2.

16. Record the final color of the solution in Table 2. Also record the pH range at which the color change of the solution occurred.

17. Repeat Steps 10 - 17 for Beaker 2 and Beaker 3.

18. Rinse out Beakers 1, 2, and 3 and the graduated cylinder. Keep the syringe with the citric acid in the titration apparatus.

Part 2: Titration of Sudsy Ammonia with 0.5 M Citric Acid

1. Use a clean graduated cylinder to measure and pour 20 mL of sudsy ammonia into a clean 100 mL beaker. When measuring the sudsy ammonia, it is best to slightly tilt the graduated cylinder for the majority of the time while pouring the sudsy ammonia to avoid bubble formation. Be sure to return the graduated cylinder to a vertical position when you near the end of the pour to obtain an accurate measurement.

2. Add enough 0.5 M citric acid to the syringe so that the meniscus meets the 30.0 mL line of the syringe.

3. Remove the cap from the pH meter, and turn the pH meter on by moving the switch at the top to the right. Rinse the pH meter probe with approximately 10 mL distilled water.

4. Place your pH meter into the beaker with sudsy ammonia. Swirl it around and record the initial pH in Table 3. This establishes the initial reading of the solution.

5. Begin to add 0.5 mL (approximately 8 drop) aliquots of the 0.5 M citric acid from the syringe. Swirl the beaker to mix the solution and record the pH of the solution for each aliquot added in Table 3.

6. Slow down the flow of citric acid to a dropwise technique when you notice the pH begin to drop significantly. This should take approximately 2 - 4 mL.

7. Record the final pH of the solution when the reaction is complete (i.e., when the pH has dropped significantly.

8. Using the recorded data, make a plot of the titration curve. Include your graph in Post-Lab Question 1.

9. Answer Post-Lab Questions 2 and 3. You will need this information to proceed with Part 3 of the procedure.

10. Clean the beaker and the graduated cylinder.

Part 3: Titration of Sudsy Ammonia with 0.50 M Citric Acid Using the Appropriate Indicator

1. Place 20 mL of sudsy ammonia into a clean 100 mL beaker. When measuring the sudsy ammonia, it is best to slightly tilt the graduated cylinder for the majority of the time while pouring the sudsy ammonia to avoid bubble formation. Be sure to return the graduated cylinder to a vertical position when you near the end of the pour to obtain an accurate measurement.

2. Top off the 30 mL syringe with citric acid so that the meniscus meets the 30.0 mL mark of the syringe.

3. Place the correct amount of the appropriate indicator you chose (the indicator you chose in Post– Lab Question 3) into the sudsy ammonia. Swirl the solution and record the initial pH and color of the solution in Table 3.

4. Slowly add the citric acid from syringe in 0.5 mL (approximately 8 drop) aliquots, continually swirling the solution in the beaker. Pay special attention to the **color** and **pH** of the solution. Record the pH and color of the solution in Table 3.

5. As you notice the color of the solution begin to change, slightly close the stopcock to slow the citric acid flow. A drop-wise flow is ideal.

Table 2: Part 1 - Indicator, pH Range, and Color Change

Flask	Indicator	Amount Needed for Reaction to Complete	pH Range	pH at Color Change	Color (Initial : Final)
1					
2					
3					

Table 3: Titration Data

Part 2: mL Citric Acid Added	Part 2: pH	Part 3: mL Citric Acid Added	Part 3: pH	Part 3: Color
0.0		0.0		
Total mL Added:		Total mL Added:		

6. Record the final pH and color of the solution when the reaction is complete (i.e., when the color change is complete).

Post-Lab Questions

1. Create a graph of the titration curve of 20 mL sudsy ammonia with 0.5 M citric acid. Be sure to title your graph and label your axes. You may also create the graph using a graphing software program.

2. What is the pH of the solution at the equivalence point shown by the titration curve?

3. Which indicator shows a color change at about the same pH as the equivalence point?

4. Fill in the following data:

   a. Initial color of the solution:

   b. Amount of citric acid added to reach the end point:

   c. Color of solution after end point was reached:

5. Identify two reasons why you did not choose the other two indicators to be used in the titration.

6. How did the mL of titrant needed to reach the endpoint using the indicator you chose compare with the mL of titrant needed to reach the equivalence point?

7. Is the indicator you chose a good indicator for this titration? Explain your answer.

More Than An Experiment, An Experience

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