All Titration Indicators are Not Created Equal—A Lecture Demonstration

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Abstract: This demonstration provides a simple method for emphasizing the importance of indicator selection to determine the end point of a titration involving the common weak base household ammonia. By using phenolphthalein and methyl red as the indicators in the titration of two samples of aqueous household ammonia with hydrochloric acid, students can discover that there is a difference between the two apparent termination points. There is clearly a visible difference in the sharpness of the color changes that take place, and the presence of ammonia at the apparent phenolphthalein endpoint is also detected by olfactory means. By showing the different termination ranges on a titration curve, the students can readily see that phenolphthalein does a poor job of showing the proper termination point,; whereas methyl red shows the proper titration end point.

Introduction

Ask a typical undergraduate chemistry student what should be used to determine the completion of a titration, and almost invariably you will receive the response "phenolphthalein." Though this is a very useful indicator, it is not, as most students think, the panacea of the titration world. In this article we describe a demonstration that illustrates the important concepts involved in the choice of an indicator and the errors that can be introduced by using an inappropriate one. An advantage of this demonstration is its simplicity and ease of set-up compared to a full-scale titration [1]. This demonstration is most appropriate for introductory courses in general or analytical chemistry, and it has proven to be a useful teaching tool even in courses for non-majors.

Procedure

The titration may be carried out directly in a wide-bore (to facilitate mixing of the solutions) graduated cylinder. For larger classes, it is convenient to use a graduated cylinder to measure the ammonia and place it in a 400-mL beaker on an overhead projector stage to enable students to observe the color changes in the projected image.

Measure 25 mL of the household ammonia with a graduated cylinder and transfer it to a 400-mL beaker. Add several drops of phenolphthalein until there is a visible red tint to the ammonia mixture. Use a plastic wash bottle to add 1.0 M HCl until the pink tint nearly disappears and have the class assess whether the color change was sharp or gradual. It is convenient to use a plastic squeeze bottle when adding the HCl; however, a burette can also be employed if greater precision is desired. Pour the solution back into the graduated cylinder and note the amount of HCl that was required to effect the color change, then return the solution (total volume approximately 70 mL; exact values will of course depend on the strength of the household ammonia) to the original beaker. Now, measure out a second 25-mL sample of ammonia and place it in another 400-mL beaker with sufficient methyl red for the yellow tint to be readily visible. Again, add the HCl solution until the color of the solution changes from yellow to red. Prompt the class to comment on the "sharpness" of the color change. Pour this solution into the graduated cylinder and note the amount of HCl that was required for the termination of the titration. There should be a difference of approximately 3 mL between the two titrations.

Another method of demonstrating that phenolphthalein underestimates the amount of HCl required to take the reaction to completion is by a simple olfactory test [2]. Invite an unfortunate student to testify to the class after having taken a cautious sniff of the ammonium solution titrated with phenolphthalein as an indicator that there is definitely still ammonia present in the solution. (An alternative presentation suggested by a reviewer would involve use of a blindfolded student as an independent reporter of the endpoint.) With methyl red, after the proper termination point is reached, no scent of ammonia can be detected because the reaction is complete and hydrolysis is not significant at this pH.

A final demonstration to show the students that the reaction end point has not been reached is to combine the contents of the two 400-mL beakers. If they were both reacted to completion, a simple dilution of the methyl red color with the equal volume of clear liquid would be expected. Due to the unreacted ammonia in the phenolphthalein solution, however, the combined solutions immediately turn yellow and require the amount by which the first (phenolphthalein) titration was underestimated in order to return to the methyl red color.

A variation of this demonstration that can be used to demonstrate the difference in termination points is to add both indicators simultaneously to the same ammonia solution. Explain to the students (possibly by use of a titration curve plot, see Figure 1) that with the addition of HCl, the apparent termination point of phenolphthalein will be reached and the red color will turn to a yellow color. From this point until the methyl red changes the solution from yellow back to a red color, all of the HCl added represents the error in using phenolphthalein improperly as the indicator.

Results and Discussion

This divergent behavior of the two indicators can be readily explained by use of the calculated titration curve (Figure 1).

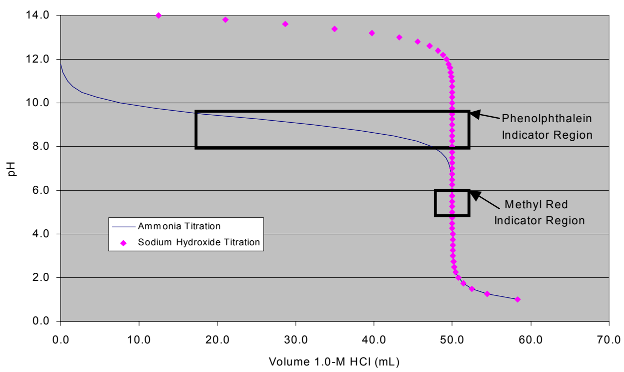

Figure 1. Calculated titration curves for 25.0 mL of 2.0 M NaOH and 25.0 mL of 2.0 M NH_3 with 1.0 M HCl.

Although textbooks [3, 4] frequently discuss the importance of indicator selection and the consequence of an improper choice, this demonstration is an effective method of illustrating this concept using a household chemical with which students are familiar [5]. For a more advanced class, the students can be assigned, before or after the presentation, to generate the appropriate titration curve for this reaction. Instructions for generating a titration curve using a spreadsheet application can be found in Harris [6] as well as other sources [7, 8]. This demonstration provides a useful introduction to more sophisticated experiments, such as potentiometric measurement of acid–base titration errors [9], using spectrophotometry to find the pK_a of an indicator [10], and characterization of various indictors pH transition range and their color changes [11].

Figure 1 shows that when a titration between a strong acid and a strong base is performed (e.g., NaOH and HCl), it makes little difference which indicator is used for the titration. However, when the titration is that of a strong acid with a weak base (or vice versa), the termination region can narrow significantly. Methyl red (color change at pH 4.8–6.0 [12]), rather than phenolphthalein (pH 8.0–9.6), is clearly the appropriate indicator for the titration of a weak base with a strong acid.

Also by referring to Figure 1, it is easy to understand why the exact endpoint is so difficult to determine with phenolphthalein indicator. It is a challenge to see when the phenolphthalein color change is complete, because the change from red to clear is gradual. The slow change is due to the fact that within the phenolphthalein transition range (pH of 8–9.6) the pH of the solution changes slowly, over a large volume of titrant (because of the buffering that occurs in the ammonia/ammonium ion solution). In contrast, a steep drop in the curve is noted at the methyl red endpoint.

References and Notes

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