

- Add more indicator if too pale

5. Add same no of drop of indicator to 1st and 2nd
6. Dilute HCl with water, vol same as 3rd
Dilute NaOH with water, vol same as 3rd
7. Add NaOH 1 cm³ at a time to 3rd, mix
8. Measure pH with pH meter just when color change (compare to 1st)
9. Add NaOH 1 cm³ at a time to 3rd, mix
10. Measure pH until color change is complete (compare with 2nd)
11. Repeat with other indicator

[illegible]

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F.7 Chemistry Laboratory Report

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Experiment 4: Indicator

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Objective

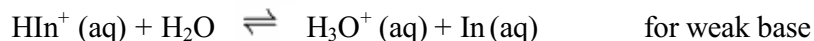
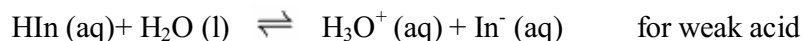
To determine the indicator range of some acid-base indicators

Introduction

In this experiment, the indicator ranges of some acid-base indicators were determined. Indicators are chemicals that would change color as the pH of the solution in which they are dissolved changes within the indicator range.

Indicators are commonly used in acid-base equilibrium in order to determine the concentration of a solution. In the titration, the equivalence point, which is the point at which equal quantities of acid and base have reacted, needed to be determined. As the suitable indicators change color obviously or reach end point, near the equivalence point, the end point could signify the equivalence point. However, indicators would only be applicable within the indicator range. If the pH of the solution at equivalence point was out of this range, the indicator shows no change in color and hence failed to serve its function.

The principle behind the color change of indicator mainly lies on the ratio between the indicator and its conjugate counterparts which had different color. The indicator was a weak acid or a weak base itself which dissociated in water to set up an ionic equilibrium.



Both the indicator and its conjugated counterparts are colored. With the change in equilibrium position, the ratio of the indicator and conjugated counterparts changes so as the ratio of different color. For example, in acidic medium, the equilibrium would shift to the left and the solution. The concentration of conjugated counterpart is higher than the indicator and hence its color masked the original color. The solution would show the color of the dominant species which is HIn for weak acid and HIn^+ for weak base. On the other hand, in alkaline medium, the equilibrium would shift to right and the solution would show the color of the dominant species which is In^- for weak acid and In for weak base.

To be specific, a particular ratio of indicator to conjugated counterparts would give a distinct color.

$$\square K_C = [\text{H}_3\text{O}^+ \text{ (aq)}][\text{In}^- \text{ (aq)}] / [\text{HIn (aq)}][\text{H}_2\text{O (l)}]$$

$$K_{\text{In}} = [\text{H}_3\text{O}^+ \text{ (aq)}][\text{In}^- \text{ (aq)}] / [\text{HIn (aq)}]$$

$$[\text{H}_3\text{O}^+ \text{ (aq)}] = K_{\text{In}} \times [\text{HIn (aq)}] / [\text{In}^- \text{ (aq)}]$$

$$-\log [\text{H}_3\text{O}^+ \text{ (aq)}] = -\log K_{\text{In}} - \log [\text{HIn (aq)}] / [\text{In}^- \text{ (aq)}]$$

$$\text{pH} = \text{p}K_{\text{In}} + \log [\text{In}^- \text{ (aq)}] / [\text{HIn (aq)}]$$

The ratio of $[In^- (aq)] / [HIn (aq)]$ determined the color of the solution.

When $[In^- (aq)] / [HIn (aq)] \leq 1/10$, color of $[HIn (aq)]$ is observed and there is no apparent change of color.

When $[In^- (aq)] / [HIn (aq)] \geq 10$, color of $[In^- (aq)]$ is observed and there is no apparent change of color.

When $[In^- (aq)] = [HIn (aq)]$, an intermediate color between color of $[HIn (aq)]$ and $[In^- (aq)]$ is shown.

Thus, the indicator range of an indicator is the pH when the ratio smaller than or equals to 1/10 to the pH when the ratio is larger than or equals to 10.

Therefore, when the ratio is just larger than 1/10, the color of the solution starts to change. Within the range between $[In^- (aq)] / [HIn (aq)] \leq 1/10$ to ≥ 10 , the color of the solution keeps changing. If $[In^- (aq)]$ is smaller than $[HIn (aq)]$, the color would be closer to that of $[HIn (aq)]$. If $[In^- (aq)]$ is larger than $[HIn (aq)]$, the color would be closer to that of $[In^- (aq)]$. And when the ratio is just larger than 10, the solution shows the color of $[In^- (aq)]$ and has no further color change.

Procedure

Preparation of solutions

To start with, three burettes were filled with deionized water, 0.1M sodium hydroxide solution and 0.1M hydrochloric acid. Three beakers were labeled as beaker 1, 2 and 3. Beaker 1 was then filled with 25cm³ of 0.1M hydrochloric acid and 10 cm³ of deionized water. Beaker 2 was filled with 25cm³ of 0.1M sodium hydroxide solution and 10 cm³ of deionized water. And beaker 3 was filled with 25cm³ of buffer solution. The solutions should be mixed thoroughly.

About 2 drops of indicator should be added to beaker 3 and if the color is too pale, more drops could be added. However, it should be noted that, the minimum amount of indicator should be added to give the distinct color. After that, the same number of drop of indicator was added to beaker 1 and beaker 2.

Determination of indicator range

1cm³ of 0.1 M sodium hydroxide solution was added each at a time to beaker 3 and the contents were mixed thoroughly. With comparison to beaker 1, the pH of the mixture was measured with a pH meter at the first sign of color change. This would be the lower point of the range.

Then, 1cm³ of 0.1 M sodium hydroxide solution was continuously added each at a time to beaker 3. With comparison to beaker 2, the pH of the mixture was measured with a pH meter at the point when there is no further change in color. The experiment was then repeated for other indicators to determine their corresponding indicator range.

In this process, the solution in beaker 1 and beaker 2 could be diluted from time to time in order to give fair comparison with color of beaker 3 with the same volume.

Results

Indicator	pH (color change start)	pH (color change complete)	Color 1	Color 2
Litmus	6.8	9.9	Pink	Blue
Phenolphthalein	7.1	10.4	Colorless	Pale pink
Methyl red	3.3	5.3	Red	Pale yellow
Methyl orange	4.3	5.6	Red	Pale yellow
Bromophenol blue	2.6	3.8	Pale yellow	Pale purple

Discussion

Measured in buffer solution

The indicator range of indicator was determined by adding sodium hydroxide solution to buffer solution in beaker 3. Beaker 3 should be filled with buffer solution but not deionized water for easy measurement of pH at particular point.

To begin with, a buffer solution is a solution which tends to resist pH change when it is diluted or a small amount of acid or bases is added to it. With the ions inside the buffer solution, the acid and base added would be converted into weak acid or weak bases which would only slightly ionized into water.

This was because the buffer solution is made of weak acid or weak base with its salt. The salt in buffer solution completely ionized in water to give the salt of the acid or base and the other oppositely charged ion. The weak acid or base slightly ionized in water and the ion from the dissociation of salt further shifted the equilibrium position of dissociation to the left. The pH changes slightly when sodium hydroxide was added to the buffer solution as the hydroxide ions react with the weak acid or base in buffer to form water and ion of weak acid or base. There was no big change in concentration of hydroxide or hydrogen ions. Thus, every time when the sodium hydroxide solution was added, most of the hydroxide ion inside would be consumed and resulted in a slight change in pH.

It is advantageous to have a slight change in pH. With a drastic change in pH, and hence a large amount of hydroxide ion added in each addition, it promoted the reaction with the indicator and cause a large change in the concentration of indicator and conjugated counterpart in the same period of time. Thus, the color of the indicator changed very quickly. After spotting the color change, the pH may already changes a lot and the measurement of pH by pH meter may not be exactly at the time when the ratio equals to 1/10 or 10.

Minimum amount of indicator should be added

A minimum amount of indicator should be added. This is because indicators are weak acid and weak base themselves that would dissociate slightly in the solution to give hydrogen ion or hydroxide ion. If too much indicators were added, the amount of hydrogen ion or hydroxide ion dissociated from it would be significant. For a weak acid serving as indicator, more hydrogen ion would be detected by pH meter than the solution originally have, and hence the measured pH would be lower. As for a base serving as indicator, the measured pH would be higher than if small amount of indicator was added.

Comparison with beaker 1 and beaker 2

The indicator range was not determined only by observing the color change of the buffer solution react with the sodium hydroxide solution. Instead, the color change was compared with the color of diluted hydrochloric acid and diluted sodium hydroxide solution.

The lower point of the pH range was determined at the first sign of color change when compared to the acid. As the pH of hydrochloric acid, a strong acid, should be well below the pH range of the indicator. Thus, it could be assumed that the color of buffer solution before color change equaled to that of the acid. To ensure the color of buffer solution and acid was originally the same, a small amount of hydrochloric acid could be added to the buffer solution first. Once passed this point, the color kept changing and the pH was within the range of indicator. The upper point of the range was determined by comparing to the color of a strong base. For buffer having pH further from the range, the color remained unchanged up to pH14. Thus, the color above the upper limit would be the same as base at much higher pH. When, the colors of two beakers were the same, the upper point of the indicator range was found.

It should also be noted that the volume of beaker 1, 2 and 3 should be approximately the same. It was because even with the same pH and hence same color of indicator, the color would be paler if the solution was diluted or deeper for a more concentrated solution. This made it difficult to compare the color of the solutions.

Errors and improvements

The experimental ranges usually have a pH of ± 1 different from the theoretical indicator range. This showed that certain errors exist which may affect the accuracy of the experiment.

Indicator	Indicator range	Theoretical indicator range
Litmus	6.8-9.9	4.5-8.3
Phenolphthalein	7.1-10.4	8.2-10.2
Methyl red	3.3-5.3	4.4-6.2
Methyl orange	4.3-5.6	3.1-4.4
Bromophenol blue	2.6-3.8	3.0-4.6

Firstly, the pH meter may not be precise enough. By immersing the pH meter into the neutral distilled water, it was found that the pH was not equal to 7. So, the determination of pH later when color changed may not be accurate also. In order to have a higher accuracy, accurate calibration of pH meter was needed. The pH meter was first dipped into standard buffer solutions of known pH before use and then the meter was adjusted to match this standard pH. Secondly, rinsing of pH meter would be required every time after measuring the pH. This prevent the residue solution from sticking on the surface of probe that would affect the pH. The pH meter should be immersed in water without using. But as to reduce contamination, this beaker of water should be exchanged every time after the measurement of pH.

Secondly, only if the point of color starts to change and stop the change were determined accurately, can we know the corresponding pH at that moment. Otherwise, the pH measured by pH meter would just be the pH within the range. Thus, the determination of color change is important. Yet, determination by human eyes is not accurate and quite objective. To improve this situation, the color of beaker 3 could be compared with beaker 1 and 2 as stated above. Also, white tile could be placed under the beakers to make the color change more conspicuous.

Thirdly, equal number of drops of indicators was added to the three beakers to have fair comparison of the colors. If more drops of indicator are added, the color of the solution would be deeper and did not resemble the initial color of beaker 3. This may lead to an inaccurate determination of the point of color change and a measurement of pH at wrong time.

Fourthly, as the sodium hydroxide solution was added 1 cm^3 each time, the accuracy could only be corrected to 1 cm^3 . If color change occurred when an additional amount of 0.5 cm^3 added to the solution, an extra amount of 0.5 cm^3 of sodium hydroxide solution may already have been added. Thus, a higher pH would be measured. Therefore, sodium hydroxide solution could be added drop by drop when the color was about to change or stop to change. The pH meter should also be dipped immediately when color started to change or stopped to change so as to measure the pH exactly at these points.

Conclusion

In this experiment, the indicator range of the following indicators were measured.

Indicator	Indicator range
Litmus	6.8-9.9
Phenolphthalein	7.1-10.4
Methyl red	3.3-5.3
Methyl orange	4.3-5.6
Bromophenol blue	2.6-3.8