Experiment E5

Preparation and Properties of a buffer

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Objective

To prepare a buffer solution and observe the properties of a buffer

Introduction

Buffer is a solution that can resist dramatic pH change when a small amount of acid or alkali is added to it. It is made up of equal amount of a weak acid/alkali and its conjugate base/acid. The working principle of buffer can be demonstrated as below:

Ethanoic acid is a weak acid that undergoes slight dissociation,

$$CH_3COOH\ (aq) + H_2O(l) \rightleftharpoons CH_3CHCOO^-(aq) + H_3O^+(aq)$$

where the equilibrium position mainly lies on the left side.

At the same time, sodium ethanoate undergoes complete dissociation,

$$CH_3COONa\ (aq) \rightleftharpoons CH_3COO^-\ (aq) + Na^+\ (aq)$$

in which its equilibrium position lies mainly on the right side.

When a small amount of acid is added to the solution, the addition of $H_3O^+(aq)$ will shift the equilibrium of $CH_3COOH\ (aq) + H_2O(l) \rightleftharpoons CH_3CHCOO^-(aq) + H_3O^+(aq)$ to the left. The increase of $[CH_3COOH]$ does not cause a significant change in pH, thus the pH of the solution is kept almost constant.

On the other hand, if a small amount of base is added to the solution, $CH_3COOH(aq)$ will react with the base, giving out salt i.e. CH_3COONa. This shifts the equilibrium of $CH_3COONa\ (aq) \rightleftharpoons CH_3COO^-\ (aq) + Na^+\ (aq)$ to the right. The increase of $[CH_3COO^-]$ does not contribute to any change in pH. Therefore, the pH of the solution remains almost constant.

Procedures

Preparation of Solution A(Buffer)

1. The pH meter was calibrated with a buffer of pH 10.
2. 25 cm^3 of 0.1M NaOH was measured with a measuring cylinder and poured into a 100 cm^3 beaker.
3. 50 cm^3 of 0.1M ethanoic acid was measured using a measuring cylinder and poured into the same beaker.
4. The reaction mixture was stirred with a clean glass rod and marked as solution A.
5. The pH of solution A was measured with the calibrated pH meter.

Preparation of Solution B(Unbuffered solution)

1. 25 cm^3 of 0.1M NaOH was measured with a measuring cylinder and poured into a 100 cm^3 beaker.
2. A small amount of 0.1M HC*l* was added to the beaker.
3. The reaction mixture was stirred with a clean glass rod.
4. The pH value was measured with the calibrated pH meter.
5. 0.1M HC*l* was added to the beaker until the pH of the solution was the same as solution A.
6. The solution was marked Solution B.

Testing of the buffer capacity of Solution A and B

1. 20 cm^3 of Solution A was measured using a clean measuring cylinder and poured into a 100cm^3 beaker.
2. 1 cm^3 of 0.1M HC*l* was measured with a 10 cm^3 measuring cylinder and poured to the beaker.
3. The reaction mixture was stirred with a clean glass rod. The pH of the solution was measured with the calibrated pH meter.
4. 4 cm^3 of HC*l* was added to the beaker. The reaction mixture was stirred with a clean glass rod.
5. The pH value was measured with the pH meter.
6. The above 5 steps were repeated using 0.1M NaOH instead.
7. The same procedures were conducted with solution B.

Testing of effect of dilution on pH

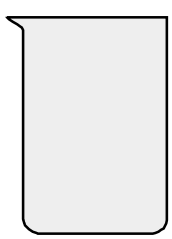
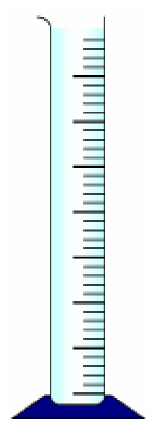
1. 5 cm^3 of Solution A was measured with a clean measuring cylinder and poured into a 100 cm^3 beaker.
2. 45 cm^3 of distilled water was measured using clean measuring cylinder and poured into the beaker.
3. The diluted reaction mixture was stirred with a clean glass rod.
4. The pH of the reaction mixture was measured.
5. The above procedures were repeated using solution B.

Chemicals used

0.1M sodium hydroxide
0.1M ethanoic acid
0.1M hydrochloric acid

Apparatus

Beaker
50 mL Measuring cylinder
10 mL Measuring cylinder
pH meter
Glass rod

Beaker	**pH meter**	**Measuring cylinder**
		

Result table

	pH value	
	Solution A	Solution B
Original solution	4.5	4.5
After addition of 1 cm^3 of 0.1M HCl	4.5	2.3
After addition of a total 5 cm^3 of 0.1M HCl	3.9	1.8
After addition of 1 cm^3 of 0.1M NaOH	4.8	11.4
After addition of a total of 5 cm^3 of 0.1M NaOH	5.5	12.2
After addition of 45 cm^3 of distilled water	4.6	5.9

Data analysis:

The pH values measured are mostly coherent to the calculated results within a range of 0.1 to 0.2 only.
For solution A, the change in pH was insignificant when an acid, a base or water was added to it.
For solution B, the change in pH was dramatic even for a small amount of acid or base added. Dilution of the solution also changed its pH value.

Data Analysis

Calculation

	pH value	
	Solution A	Solution B
Original solution	$pH=-\log K_a$ $=-\log 1.74\times10^{-5}$ $=\underline{4.76}$	4.76
After addition of 1 cm^3 of 0.1M HCl	$pH=-\log 1.74\times10^{-5}+\log\frac{(0.1\div3\times0.02-0.001\times0.1)\div0.021}{(0.1\div3\times0.02+0.001\times0.1)\div0.021}$ $=\underline{4.63}$	$pH=-\log(1\times10^{-3}+0.1\times0.001\div0.021)$ $=\underline{2.32}$
After addition of a total of 5 cm^3 of 0.1M HCl	$pH=-\log 1.74\times10^{-5}+\log\frac{(0.1\div3\times0.02-0.005\times0.1)\div0.025}{(0.1\div3\times0.02+0.005\times0.1)\div0.025}$ $=\underline{3.91}$	$pH=-\log(1\times10^{-3}+0.1\times0.005\div0.025)$ $=\underline{1.70}$
After addition of 1 cm^3 of 0.1M NaOH	$pH=-\log 1.74\times10^{-5}+\log\frac{(0.1\div3\times0.02+0.001\times0.1)\div0.021}{(0.1\div3\times0.02-0.001\times0.1)\div0.021}$ $=\underline{4.89}$	$pH=-\log(1\times10^{-3}-0.1\times0.001\div0.021)$ $=\underline{11.68}$
After addition of a total of 5 cm^3 of 0.1M NaOH	$pH=-\log 1.74\times10^{-5}+\log\frac{(0.1\div3\times0.02+0.005\times0.1)\div0.025}{(0.1\div3\times0.02-0.005\times0.1)\div0.025}$ $=\underline{5.60}$	$pH=14+\log(0.1\times0.005\div0.025)$ $=\underline{12.3}$
After addition of 45 cm^3 of distilled water	$pH=-\log 1.74\times10^{-5}+\log\frac{(0.1\div3\times0.005)\div0.05}{(0.1\div3\times0.005)\div0.05}$ $=\underline{4.76}$	$pH=-\log(1.74\times10^{-5}\times0.005\div0.05)$ $=\underline{5.76}$

Conclusion

In the experiment, solution A was the buffer. Solution A contains ethanoic acid and its conjugate base, sodium ethanoate, in equal amount, from the neutralization between sodium hydroxide and ethanoic acid. Solution B is only a solution containing excess hydrochloric acid.

When a small amount (1 cm^3) of hydrochloric acid was added to solution A, the pH remained almost the same. This change was insignificant that even the pH meter could not show the change. On the contrary, when a small amount (1 cm^3) of sodium hydroxide was added to it, the pH increases from 4.5 to 4.8 only. The change was small. Dilution of solution A almost causes no change in pH. This shows solution A, a buffer, can resist the pH change upon addition of acid, base or water.

The calculated and measured values of pH vary within a range of 0.1 to 0.2. One of the possible errors is that the experiment was not performed at room temperature, i.e. 25℃/298K. The temperature in the laboratory was about 20.9℃. The value of dissociation constant, $K_a = 1.74 \times 10^{-5}$ provided is not applicable in the experiment. The actual dissociation of hydroxonium ions was less than the standard one.

Measuring cylinders were used in the experiment. There might be inaccuracy in the reading on the cylinder. Thus, the solutions prepared might not be of known concentration or the volume of acid or base added was not exactly 1/5 cm^3. Pipette might be used to transfer the solution instead as it has a higher accuracy.

pH meter was used to measure the pH values of solutions. However, the accuracy and precision of pH meter varies. If the precision of the pH meter is not high enough, a slight change in pH value may not be shown. An inaccurate pH meter also gives invalid pH values. A pH meter with higher precision may be used and calibration of pH meter helps to improve the accuracy.