ACTIVITY 8.4.1 QUANTITATIVE TITRATION

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Evidence

Part I

mass $KHC_8H_4O_{4(s)}$ used: 0.400 g

average volume of $NaOH_{(aq)}$ used : 19.20 mL

Part II

2.00 mL vinegar

average volume of $NaOH_{(aq)}$ used : 17.50 mL

Analysis

(a) $$C_{NaOH(aq)} = \frac{19.6 \times 10^{-3}\ \text{mol}}{19.20 \times 10^{-3}\ \text{L}}$$

$$= 0.102\ \text{mol/L}$$

The molar concentration of the sodium hydroxide solution was 0.102 mol/L.

(b) $KHC_8H_4O_{4(s)}$ is commonly used in standardization titrations because it is a weak acid and because a known amount (in moles) of it can conveniently and accurately be determined using a balance. Furthermore, potassium hydrogen phthalate can be obtained to a high degree of purity. (The degree of purity is given on the reagent bottle.)

(c) Shaking the solution before titrating dissolves more carbon dioxide into the solution, which then produces carbonic acid. This would slightly increase the acidity of the solution—an unnecessary source of error.

(d) Boiling the water removes dissolved carbon dioxide from the solution.

(e) For the titration of vinegar:

$$HC_2H_3O_{2(aq)} + NaOH_{(aq)} \rightarrow H_2O_{(l)} + NaC_2H_3O_{2(aq)}$$

$$n_{NaOH_{(aq)}} = 17.50\ \text{mL} \times 0.102\ \text{mol/L}$$

$$n_{NaOH_{(aq)}} = 1.785\ \text{mmol} \qquad \text{(extra digit carried)}$$

$$C_{HC_2H_3O_{2(aq)}} = \frac{1.785\ \text{mmol}}{2.00\ \text{mL}}$$

$$C_{HC_2H_3O_{2(aq)}} = 0.893\ \text{mol/L}$$

The molar concentration of acetic acid in vinegar is also 0.893 mol/L.

Evaluation

(g) The Procedure is judged to be adequate because the concentration of the unknown acid was determined with no obvious flaws. One minor improvement would be to use a more concentrated sodium hydroxide solution. This would allow for a larger sample of vinegar to be titrated.

INVESTIGATION 8.5.1 BUFFER ACTION

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Prediction

(a) According to the empirical definition of a buffer, the addition of small amounts of a strong acid or strong base produces only a slight change in the pH of the mixture until the capacity of the buffer is exceeded. The reasoning is that the added hydronium ion or hydroxide ion is quantitatively removed by reacting with a component of the weak acid–weak base buffer mixture, as shown below.

$$H_3O^+_{(aq)} + HPO^{2-}_{4(aq)} \rightarrow H_2O_{(l)} + H_2PO^-_{4(aq)}$$

$$OH^-_{(aq)} + H_2PO^-_{4(aq)} \rightarrow H_2O_{(l)} + HPO^{2-}_{4(aq)}$$

Both of the above reactions are predicted to be quantitative.

Experimental Design

(b) A $H_2PO^-_{4(aq)}/HPO^{2-}_{4(aq)}$ buffer is prepared and then tested by adding small amounts of $HCl_{(aq)}$ and $NaHO_{(aq)}$, one drop at a time. Indicators are used to indicate a sudden change in pH. A $NaCl_{(aq)}$ solution is tested as a control. The independent variable is the volume (number of drops) of acid or base added. The dependent variable is the colour of the indicator (the pH). The controlled variables are temperature, volume, and concentration of buffer used, and volume and concentration of added acid or base.

Materials

(c) lab apron
eye protection
distilled or deionized water
0.10 mol/L $KH_2PO_{4(aq)}$
0.10 mol/L $NaOH_{(aq)}$
0.10 mol/L $NaCl_{(aq)}$
0.10 mol/L $HCl_{(aq)}$
two 50-mL graduated cylinders
150-mL beaker
4 small test tubes or spot plate

Evidence

(d)

	Colour of bromocresol green in buffer solution					
Volume of $HCl_{(aq)}$ (drops)	0	2	4	6	8	10
Colour of buffer solution	blue	blue	blue	blue	blue	blue
Colour of salt solution	blue	green	yellow	yellow	yellow	yellow

	Colour of phenolphthalein in buffer solution					
Volume of $NaOH_{(aq)}$ (drops)	0	2	4	6	8	10
Colour of buffer solution	blue	blue	blue	blue	blue	blue
Colour of salt solution	colourless	pink	red	red	red	red

Analysis

(e) Based on the evidence gathered in this investigation, the slow addition of a strong acid or a strong base to a $H_2PO^{-}_{4(aq)}/HPO^{2-}_{4(aq)}$ buffer does not change the pH significantly.

Evaluation

(f) The Experimental Design is judged to be adequate because there was clearly one independent variable and one dependent variable, and all necessary controls were employed. Using the salt solution as a control was even better than using distilled water. A universal indicator or pH paper might have provided more precise information as to the pH change. The procedure is judged to be adequate because it was simple and efficient. Perhaps the pH of the prepared buffer could be checked with an indicator, pH paper, and pH meter to provide more certainty, and the droppers used to add the strong acid and strong base could be calibrated to give more quantitative evidence. The skills used to carry out the investigation are judged to be adequate because they are very simple. The precision of the measurements is not crucial. Overall I am confident in the evidence gathered by the use of the experimental design, procedure, and skills employed herein.

(g) The Prediction based upon the empirical definition of a buffer is judged to be verified because the evidence supported the predicted effects on the pH. The pH of the buffer solution, as predicted, did not change significantly when a strong acid or base was added. The empirical definition of a buffer is judged to be acceptable because the prediction was verified. The prediction based upon the empirical definition of a buffer was accurate. It appears that we can continue to have confidence in this concept.

CHAPTER 8 SUMMARY

MAKE A SUMMARY

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(Answers may vary, but should include a page for each of the following six sections: The Nature of Acid–Base Equilibria; Weak Acids and Bases; Acid–Base Properties of Salt Solutions; Acid–Base Titration; Buffers; and The Science of Acid Deposition.)

CHAPTER 8 SELF-QUIZ

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1. False. The stronger a Brønsted–Lowry acid is, the weaker its conjugate base.
2. False. Group I metal ions produce neutral solutions.
3. True
4. True
5. False. A solution of the bicarbonate ion is basic.
6. False. The pH of water would be less than 7.
7. True
8. False. Most dyes that act as acid–base indicators are weak acids.
9. True
10. (b)
11. (b)
12. (e)
13. (a)
14. (b)
15. (c)
16. (e)
17. (a)
18. (b)
19. (a)

CHAPTER 8 REVIEW

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Understanding Concepts

1. $$n_{NaOH} = \frac{8.50\text{ g}}{40.00\text{ g/mol}}$$

$$n_{NaOH} = 0.2125\text{ mol} \qquad \text{(extra digits carried)}$$

$$[OH^-] = \frac{0.2125\text{ mol}}{0.500\text{ L}}$$

$$[OH^-] = 0.425\text{ mol/L}$$

$$\text{pOH} = -\log 0.425$$

$$\text{pOH} = -0.372$$

The pOH of sodium hydroxide is –0.372.