**Chemistry of the Kitchen – Acids and Bases**

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Section 3

Group 3

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# Ⅰ OBJECTIVES

* Use a homemade indicator to determine whether a common household chemical is acid or base.
* Understand the concept of pH scale.
* Find the difference between strong and weak acids.
* Titrate a vinegar sample and find out its acid concentration.

# Ⅱ INTRODUCTION



**Figure 1** organic compounds turned to distinctive colors

# Many household chemicals belong to the category

# of acids or bases. Acids are of great help in home,

# industry and environment. However, they may also

# do harm to the environment. Acids and bases have

# the ability to turn certain organic compounds to

# distinctive colors (Figure 1). We can use this

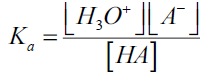
# characteristic to test the acidity of some solutions.

# Ⅲ BACKGROUND

**A. Properties of Acids and Bases**

In history, there are three ways to define an acid or a base: the classical (Arrhenius), the Brønsted-Lowry, and the Lewis. Acids and bases are electrolytes in water. A strong acid/base dissociates completely in water while a weak acid/base dissociate partly. In a chemical equation, we use → to represent the aqueous solution of a strong acid/base andC:\Users\thinkpad\AppData\Roaming\Tencent\Users\1165008570\QQ\WinTemp\RichOle\IMXG)N5{E0V3QBAR`FRLW9R.pngto represent a weak acid/base. In a weak electrolyte solution, very few origin molecules dissociate to ions.

Ka (the dissociation constant) is used to quantify the percentage of acid that is actually dissociated.

At equilibrium, 

Ka tells us how far the right reaction is preceded at equilibrium. Stronger acid has a higher Ka.

**B. The pH Scale**

# To simplify calculations, we use p-scale, a numerical system to turn the concentration into simpler figures. pH=-log[H3O+] The higher the pH, the lower the concentration. The pH of an acid solution is lower than that of a base solution. The pH of an aqueous solution commonly ranges from 0 to 14. In the lab, an acid-base indicator is used to obtain the pH value more precisely.

**C. Calculating the pH**

# Without the pH meter we can also calculate the pH value by setting up a reaction table. We assign values for the concentration of each ion at three stages: initial, change and equilibrium. We solve for the unknown x and we will get the value of [H3O+]. Then we use the equation pH=-log[H3O+] to find out pH. This mathematical process can also be reversed. During calculations, we can first check the known Ka value to omit some comparatively small numbers.

**D. Titration**

# We can use acid-base titration to determine the



# concentration of an acid. First, we place the acidic

# solution of which the concentration is unknown into a flask. The volume of the acid solution is measured. Then a base indicator, usually phenolphthalein is added to the flask. The standard solution of base in the burette is solely added into the flask until the end point (when the indicator changes color permanently) is reached. (Figure 2)In this way, the consumed amount of base can be known. [H3O+] of the acidic solution can thus be calculated by knowing the stoichiometry of the reaction.

at the end point. The indicator is phenolphthalein.

**Figure 2** An acid–base titration

**E. Overview**

# In part A, we will test the pH of different household chemicals by using red cabbage extract and a universal indicator. In part B, we will use the pH meter to find out the difference between strong and weak acids. Then we will determine an unknown concentration of a strong base. In part C, we will use the acid-base titration to calculate the acid concentration of vinegar.

# Ⅳ EXPERIMENTAL PROCEDURES

**Part A Relative Acidity/Basicity of Common Household Products**

|  |  |
| --- | --- |
| Chemicals used | Materials used |
| Universal indicator papers: Strips of  universal pH indicator paper.  Various household chemicals: (example  ammonia, vinegar, Shampoo (colorless), soda  (a colorless variety), milk, lemon juice, liquid  detergent (without dyes), milk of magnesia, tap  H2O, deionized H2O, Bleach) | Several beakers of various sizes  10-mL Graduated cylinder  Glass stirring rod  Test tube w/rubber stopper  Test tube rack  Funnel  Hot plate & Knife |

1. Working alone, select one or two household products to test from the available samples.

Rinse a freshly washed beaker with 5 mL of your selected household chemical(s).

2. Pour 10 mL of the same household product into the beaker(s). Immerse partially a strip of universal pH indicator paper, then let the indicator paper(s) dry and label it with the sample that you tested. Repeat step 1 once again.

3. Label your samples and bring the indicator paper to the front of the lab for observation by the class. Be sure to record results for other household products (color of indicator paper and corresponding pH from the indicator paper box or from the chart posted on laboratory benches).

4. Repeat procedures (2.&3.) to test the pH of a stock solution of 1M NH3.H2O (inside eyedropper vial) but use directly only 1-2 drops.

**Part B Concentration of unknown molarity of NaOH solution using KHP titration**

CNaOH VNaOH= mKHP/MWKHP

KHC8H4O4 (aq) + NaOH (aq) → KNaC8H4O4 (aq) + H2O (Ɩ)

|  |  |
| --- | --- |
| Chemicals used | Materials used |
| Unknown molarity NaOH solution inside  bottle labelled as 0.1M NaOH.  KHP  pH indicator (phenolphthalein) | Provided glassware:  beakers, Erlenmeyer flasks, & graduation  cylinders  Weighing papers & desiccators |

1. At a designated area, wear protective gloves and wash the provided glassware of your group using detergent, brush and tap water and then rinse with de-ionized water, 3 times each. Be careful not to break or crack the glassware and not to injure yourself. Any broken or cracked glassware must be reported to the instructor and the glassware must be disposed carefully in the provided glass disposal box and not into the trash can or sink.

2. Using the Mettler balance, carefully weigh out 0.4000-0.6000g KHP into 3 labeled freshly washed Erlenmeyer flasks, record each weight. Add no more than 30-40 mL of de-ionized water to entirely solve KHP and 2-3 drops of phenolphthalein (to one of the flasks).

3. Run about 25.0 mL of deionized water through a 50-mL burette and then rinse the burette with about 5.0 mL of the labelled 0.1 M NaOH solution. Mount the burette in the burette stand.

4. Use a funnel to add roughly 30.0 mL of the labelled 0.10 M NaOH solution into the burette.

Drain any air bubbles from the tip of the burette into a waste beaker. Record the initial volume V0 of the burette (-/+0.02mL).

5. Slowly add the NaOH solution from the burette drop-wise to the KHP solution in the flask with the indicator, swirling the flask after each addition. Continue until the endpoint is reached. Record the final volume V1.

6. Repeat steps 4-5 for a second and third trial using the other flasks with KHP solution.

**Part C Acid-Base Titration of Vinegar Solution**

Calculation: 10 CNaOHVNaOH = CHAcVHAc

|  |  |
| --- | --- |
| Chemicals used | Materials used |
| Same solution bottle of NaOH of Part B.  Vinegar with unknown concentration of acetic  acid (CH3COOH)  Phenolphthalein | Burette and burette stand  250-mL Erlenmeyer flask  100-mL Beaker (2)  25-mL Pipet w/bulb & graduated cylinder |

1. Run about 25.0 mL of deionized water through a 50-mL burette and then rinse the burette with about 5.0 mL of the labelled 0.1 M NaOH solution. Mount the burette in the burette stand. (No need to rinse if you are using the same burette and stock solution bottle of NaOH from Part B)

2. Use a funnel to add the labelled 0.1 M NaOH solution (from Part B) into the burette and fill it up to near below the top mark. Drain any air bubbles from the tip of the burette into a waste beaker. Record the initial volume V1 (±0.02 mL).

3. Pipet 25.00mL of the **vinegar** into a 250.00mL volumetric flask, add de-ionized water to the calibration line and shake upside down several times.

4. Pipet 25.00 mL of the solution in step 3, into 3 freshly washed & rinsed Erlenmeyer flasks and add 2 – 3 drops of phenolphthalein to the flask.

5. Slowly add the NaOH solution, drop wise from the burette to the vinegar, swirling the flask after each addition. Continue until the endpoint is reached. Record the final volume V2 (± 0.02 mL).

6. Repeat steps 2 – 4 for a second trial and third trial.

**Part D Closing your Lab Session**

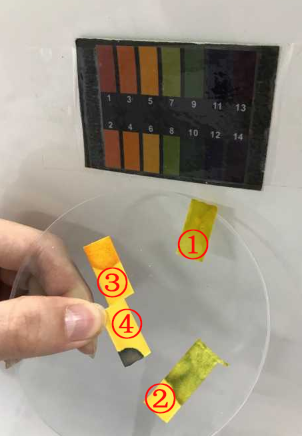
1. Wear protective gloves and wash the used glassware using detergent, brush and tap water and then rinse with de-ionized water. Be careful not to break or crack the glassware. Any broken or cracked glassware must only be disposed carefully into the provided glass disposal box

2. Clean up your designated working area thoroughly so the next lab session can use safely.

3. A designated team from the groups will be instructed just before dismissal of the laboratory to inspect and insure the general cleaning and safety of the laboratory.

# Ⅴ CALCULATION/ANALYSIS/DATA PROCESSING

**Part A Relative Acidity/Basicity of Common Household Products**

 **Figure 3** color of indicator paper compared with the chart

As is shown in **Figure 3**, we tested four chemicals:

1. Water ②example ammonia ③vitamin C solution ④NaOH solution

By comparing with the pH chart, we found that the pH of ①water is 7, ②example ammonia is 8,③vitamin C solution is 6,④NaOH solution is 13.

**Part B Concentration of unknown molarity of NaOH solution using KHP titration**

|  |  |  |  |
| --- | --- | --- | --- |
| Trial | 1 | 2 | 3 |
| mKHP (g) | 0.4866 | 0.4091 | 0.4185 |
| Total Volume of NaOH(mL) | 27.90 | 21.45 | 22.00 |
| CNaOH (mol/L) | 0.0855 | 0.0935 | 0.0907 |
| Average CNaOH (mol/L) |  | 0.0907 |  |

**Table 1 experiment data**

Using the equation CNaOH (mol/L) = mKHP \*1000/(MWKHP\*VNaOH), MWKHP=204g/mol. We calculated the concentration of NaOH in each trial.

For trial 1, CNaOH (mol/L) = 0.0855mol/L

For trial 2, CNaOH (mol/L) = 0.0935mol/L

For trial 3, CNaOH (mol/L) = 0.0932mol/L

The average value of CNaOH in the three trials is 0.0907mol/L

Considering the NaOH solution is labeled 0.1mol/L,

the relative error is =-9.3%

**Part C Acid-Base Titration of Vinegar**

|  |  |  |  |
| --- | --- | --- | --- |
|  | Initial Vol. of NaOH(mL) | Final Vol. of NaOH(mL) | Total Vol. of NaOH(mL) |
| Trial 1 | 0.50 | 23.35 | 22.85 |
| Trial 2 | 0.50 | 23.20 | 22.70 |
| Trial 3 | 0.80 | 23.80 | 23.00 |
| Average consumption of NaOH(mL) |  | 22.85 |  |

**Table 2 experiment data**

From the equation 10×CNaOHVNaOH = CHAcVHAc,

We can get CHAc=, and we plug the values CNaOH=0.1M and VNaOH=22.85mL into this equation.

CHAc=

Then we calculate the % by mass of acetic acid in vinegar.

Using the equation %=

The value of the bottle of vinegar is 5g/100mL,

which is 100%5g/(100mL1g/mL)=5%

Therefore, the relative error is

**Ⅵ DISCUSSION**

In Part B, the analysis of our experiment data indicates that the measured molarity of NaOH solution is smaller than labeled. During discussion, we’ve come up with several factors causing this error:

1. When the reaction was about to reach the end point, we didn’t control the speed properly so that NaOH was more than needed causing the indicator turning to a deeper color.
2. When transferring KHP into the Erlenmeyer flasks in step2, there might be a loss in the mass of KHP.
3. The volume of consumed NaOH might be misread. The initial volume might be read smaller than reality and the final volume might be read larger than reality.

In Part C, the analysis of our experiment data indicates that the measured % by mass of acetic acid in vinegar is larger than labeled. Here are some possible reasons:

1. The concentration of NaOH is in fact smaller than labeled.
2. The volume of consumed NaOH might be misread. The initial volume might be read smaller than reality and the final volume might be read larger than reality.

**Ⅶ CONCLUSIONS ＆ RECOMMENDATIONS**

In Experiment 1, I’ve learned the property of acids and bases to change the color of indicator. Also, I’ve learned the differences between strong and weak acids. I use the knowledge to titrate and calculate the concentration of two solutions.

In Part A, we’ve examined the pH of several household chemicals using the universal pH indicator paper. The pH of water, example ammonia, vitamin C solution and NaOH solution is 7, 8, 6, 13 relatively.

In Part B, we used the solution of KHP of known mass to titrate NaOH solution of unknown concentration. According to our experiment data and calculation, the concentration is 0.0907mol/L.

In Part C, we’ve calculated that the acid concentration of vinegar in mass is 5.2%.

I strongly recommend that experimenter should have be familiar with the experiment procedure and the operation of each laboratory apparatus. This will be of good help for saving time and avoiding accidents.

When conducting experiment B&C, experimenters should remember to add phenolphthalein into the solution, or the whole experiment will bear no fruit. Also, experimenters should carefully control the rate of drop. A single extra drop of NaOH solution may cause the indicator to turn to deeper color.

**Ⅷ REFERENCES**

-1. Peter Atkins, *Chemical Principles The Quest for Insight Seventh Edition*, Macmillan education, 2016.

-2. VC211 Laboratory Manual, UM-SJTU JI &SJTU Chemistry Department, 2018-2019.