### ACID-BASE TITRATION

## LAB PH 2.COMP

From *Chemistry with Computers*, Vernier Software & Technology, 2000.

### INTRODUCTION

A titration is a process used to determine the volume of a solution needed to react with a given amount of another substance. In this experiment, you will titrate a hydrochloric acid solution, HCl, with a basic sodium hydroxide solution, NaOH. The concentration of the NaOH solution is given and you will determine the unknown concentration of the HCl. Hydrogen ions from the HCl react with hydroxide ions from the NaOH in a one-to-one ratio to produce water in the overall reaction:

H+(aq) + Cl–(aq) + Na+(aq) +OH–(aq)  H2O(l) + Na+(aq) + Cl–(aq)

When an HCl solution is titrated with an NaOH solution, the pH of the acidic solution is initially low. As base is added, the change in pH is quite gradual until close to the equivalence point, when equimolar amounts of acid and base have been mixed. Near the equivalence point, the pH increases very rapidly, as shown in Figure 1. The change in pH then becomes more gradual again, before leveling off with the addition of excess base.

In this experiment, you will use a computer to monitor pH as you titrate. The region of most rapid pH change will then be used to determine the equivalence point. The volume of NaOH titrant used at the equivalence point will be used to determine the molarity of the HCl.

Figure 1 volume vs ph graph

Figure 1

### Purpose

The purpose of this experiment is to determine the molarity of a HCl solution by performing a titration, using a pH sensor to monitor the progress of the reaction. The procedure will illustrate the concept of an equivalence point.

**EQUIPMENT/Materials**

|  |  |
| --- | --- |
| Laptop computer with Logger *Pro* | wash bottle |
| LabPro with AC adapter | 50-mL burette |
| LabPro → computer cable | ring stand |
| Vernier pH Sensor | 2 utility clamps |
| HCl solution, unknown concentration | 10-mL pipet |
| 0.100 M NaOH solution | pipet bulb or pump |
| magnetic stirrer (if available) | 250-mL beaker |
| stirring bar | distilled water |

### Safety

* Always wear an apron and goggles in the lab.
* Use caution handling the hydrochloric acid and the sodium hydroxide. Avoid contact with skin and clothing.

**Procedure**

1. Use a pipet bulb (or pipet pump) to pipet 10 mL of the HCl solution into a 250-mL beaker. Add 50 mL of distilled water. **CAUTION:** *Handle the hydrochloric acid with care. It can cause painful burns if it comes in contact with the skin.*
2. Place the beaker on a magnetic stirrer and add a stirring bar. If no magnetic stirrer is available, you need to stir with a stirring rod during the titration.

figure 2 ph and magnetic stirrer apparatus setup illustration

Figure 2

1. Use a utility clamp to suspend a pH Sensor on a ring stand as shown in Figure 2. Position the pH Sensor in the HCl solution and adjust its position so that it is not struck by the stirring bar.
2. Obtain a 50-mL burette and rinse the burette with a few mL of the ~0.1 M NaOH solution. Use a utility clamp to attach the burette to the ring stand as shown in Figure 2. Fill the burette a little above the 0.00-mL level of the burette with ~0.1 M NaOH solution. Drain a small amount of NaOH solution so it fills the burette tip *and* leaves the NaOH at the 0.00-mL level of the burette. Record the precise concentration of the NaOH solution in your data table. Dispose of the waste solution from this step as directed by your teacher. CAUTION**:** *Sodium hydroxide solution is caustic. Avoid spilling it on your skin or clothing.*
3. Prepare the computer for data collection by opening the file in the Experiment 24 folder of *Chemistry with Computers.* The vertical axis has pH scaled from 0 to 14 pH units. The horizontal axis has volume scaled from 0 to 25 mL. Check to see that the Meter window shows a pH value between 2 and 3.
4. Before adding NaOH titrant, click  and monitor pH for 5-10 seconds. Once the displayed pH reading has stabilized, click . In the edit box, type “0” (for 0 mL added). Press the ENTER key to store the first data pair for this experiment.
5. You are now ready to begin the titration. This process goes faster if one person manipulates and reads the burette while another person operates the computer and enters volumes.
6. Add the next increment of NaOH titrant (enough to raise the pH about 0.15 units). When the pH stabilizes, again click . In the edit box, type the current burette reading, to the nearest 0.01 mL. Press ENTER. You have now saved the second data pair for the experiment.
7. Continue adding NaOH solution in increments that raise the pH by about 0.15 units and enter the burette reading after each increment.
8. When a pH value of approximately 3.5 is reached, change to a one-drop increment. Enter a new burette reading after each increment. Note: It is important that all increment volumes in this part of the titration be equal; that is, one-drop increments.
9. After a pH value of approximately 10 is reached, again add larger increments that raise the pH by about 0.15 pH units, and enter the burette level after each increment.
10. Continue adding NaOH solution until the pH value remains constant.
11. When you have finished collecting data, click . Dispose of the beaker contents as directed by your teacher.
12. Print a copy of the Table window. Enter your name(s) and the number of copies of the table.
13. Print a copy of the Graph window. Enter your name(s) and the number of copies of the graph.
14. If time permits, repeat the procedure.

### PROCESSING THE DATA

1. Use your graph and data table to determine the volume of NaOH titrant used in each trial. Examine the data to find the largest increase in pH values upon the addition of 1 drop of NaOH solution. Find and record the NaOH volume just *before* this jump.
2. Find and record the NaOH volume *after* the drop producing the largest pH increase was added.
3. Determine the volume of NaOH added at the equivalence point. To do this, add the two NaOH values determined above and divide by two.
4. Calculate the number of moles of NaOH used.
5. See the equation for the neutralization reaction given in the introduction. Determine the number of moles of HCl used.
6. Recall that you pipeted out 10.0 mL of the unknown HCl solution for each titration. Calculate the HCl concentration.
7. If you did two titrations, determine the average [HCl] in mol/L.

**DATA SHEET** Name \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Name \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Period \_\_\_\_\_\_\_ Class \_\_\_\_\_\_\_\_\_\_\_

Date \_\_\_\_\_\_\_\_\_\_\_

### ACID-BASE TITRATION

## DATA TABLE

|  |  |  |
| --- | --- | --- |
|  | Trial #1 | Trial #2 |

|  |  |  |
| --- | --- | --- |
| Concentration of NaOH | M | M |
| NaOH volume added *before* the largest pH increase | mL | mL |
| NaOH volume added *after* the largest pH increase | mL | mL |

|  |  |  |
| --- | --- | --- |
| Volume of NaOH added at equivalence point | mL | mL |
| Moles NaOH | mol | mol |
| Moles HCl | mol | mol |
| Concentration of HCl | mol/L | mol/L |
| Average [HCl] |  |  |