Name **Period** **Date**

27. Diprotic Titration: Multi-Step Chemical Reactions

Driving Questions

A titration can tell you how much of an acid is present in a solution, but what happens if the acid has more than one hydrogen ion that it can donate? How will the titration curve differ? Why does the titration curve change and can an unknown concentration still be determined?

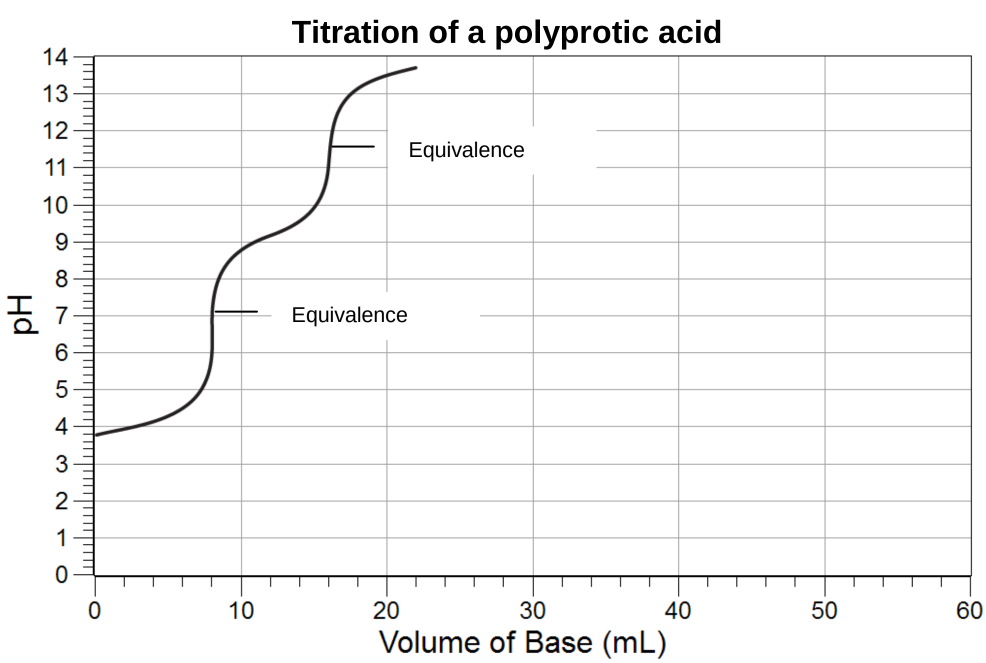
Background

Acids are substances that have a hydrogen ion that can be donated. Bases are the complements to acids in that bases accept hydrogen ions. When an acid gives up a hydrogen ion, the anionic portion left behind is called the conjugate base (because now it acts as a base by being able to take on a new hydrogen ion). When a base accepts a hydrogen ion, the newly created species is called the conjugate acid (because it can, in turn, give up the newly acquired hydrogen ion just like an acid).

Some acids have more than one hydrogen ion they can donate; these are referred to as polyprotic acids. The number of hydrogen ions that can be donated further refines the acid's classification ($1H^+$ = monoprotic, $2H^+$ = diprotic, $3H^+$ = triprotic, and so on.).

A carbonate ion (CO_3^{2-}) is the conjugate base of carbonic acid (H_2CO_3). By making a solution of sodium carbonate (Na_2CO_3), the solution contains carbonate ions that can accept hydrogen ions. These carbonate ions can be titrated with an acid, such as hydrochloric acid. The acceptance of hydrogen ions takes place in two steps: first, the carbonate ions each take on one hydrogen ion until they have all become bicarbonate ions (HCO_3^-); second, the new bicarbonate ions take on one additional hydrogen atom to become carbonic acid. The carbonic acid immediately decomposes into water and carbon dioxide gas (which can be seen as bubbles).

Each step can be seen in a titration curve. As the titration progresses, the pH can be measured with a sensor. The point at which the curve is the steepest indicates an equivalence point (where the number of moles of acid equals the number of moles of base). For polyprotic acids, there will be one equivalence point for each hydrogen ion that can be donated.



Materials and Equipment

For each student or group:

- Data collection system
- Drop counter
- pH sensor
- Micro stir bar
- Magnetic stirrer
- Beaker (2), 50-mL
- Beaker, 250-mL
- Graduated cylinder, 50-mL
- Graduated cylinder, 100-mL
- Transfer pipet
- Buret, 50-mL
- Buret clamp
- Ring stand
- Right-angle clamp
- Funnel
- Waste container
- Wash bottle filled with distilled (deionized) water
- Buffer solution, pH 4, 25 mL
- Buffer solution, pH 10, 25 mL
- Distilled (deionized) water, 200 mL
- Sodium carbonate (Na_2CO_3) solution, 40 mL
- 1.0 M Hydrochloric acid (HCl), 110 mL

Safety

Add these important safety precautions to your normal laboratory procedures:

- Hydrochloric acid is a strong acid. Avoid contact with the skin and eyes.
- Be sure that all acids and bases are neutralized before disposal down the drain.

Sequencing Challenge

The steps below are part of the Procedure for this lab activity. They are not in the right order. Determine the proper order and write numbers in the circles that put the steps in the correct sequence.

○	○	○	○	○
Rinse the buret with 1.0 M hydrochloric acid and practice adjusting the stopcock. Collect the HCl in a waste container.	Set up the titration equipment and calibrate the pH sensor.	Place a 250-mL beaker containing diluted analyte (sodium carbonate) on the magnetic stirrer.	Repeat the titration to collect a second set of data.	Titrate the sodium carbonate solution with HCl until the pH stabilizes at a pH less than 2.

Procedure

After you complete a step (or answer a question), place a check mark in the box (☐) next to that step.

Note: When you see the symbol "�" with a superscripted number following a step, refer to the numbered Tech Tips listed in the Tech Tips appendix that corresponds to your PASCO data collection system. There you will find detailed technical instructions for performing that step. Your teacher will provide you with a copy of the instructions for these operations.

Set Up

1. ☐ Start a new experiment on the data collection system. �(1.2)

2. ☐ Connect a pH sensor to the data collection system. �(2.1)

3. ☐ Place 25 mL of pH 4 buffer solution in a 50-mL beaker and 25 mL of pH 10 buffer solution in a second 50-mL beaker. Use these solutions to calibrate the pH sensor. �(3.6)

4. ☐ Will a pH sensor that is not calibrated give precise results? Why must the pH sensor be calibrated?

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5. ☐ Connect a drop counter to the data collection system. �(2.2)

6. ☐ Display pH versus Drop Count (drops) on a graph. �(7.1.1)

7. ☐ Assemble the titration apparatus, using the steps below and the illustration as a guide.

    a. Assemble the ring stand.

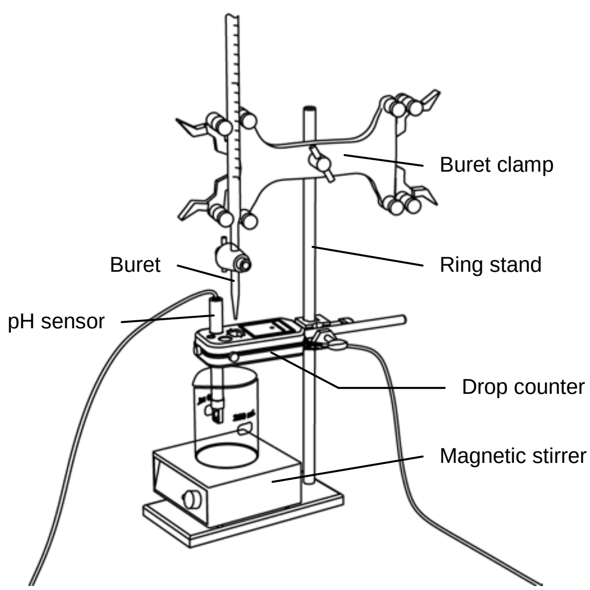
    b. Position the magnetic stirrer on (or next to) the base of the ring stand.

    c. Place a waste container on the magnetic stirrer.

    d. Use the buret clamp to attach the buret to the ring stand.

    e. Position the drop counter over the waste container and attach it to the ring stand using the right-angle clamp.

    f. Place the pH sensor through one of the slots in the drop counter.



8. ☐ Rinse the buret with several milliliters of the standardized HCl solution. Follow the steps below to complete this step.

   a. Ensure that the stopcock is closed and use a transfer pipet to rinse the inside of the buret with several milliliters of the HCl solution.

   b. Open the stopcock on the buret and drain the HCl rinse into the waste container.

   c. Repeat this process two more times.

9. ☐ Make sure the stopcock is in the "off" position and then use a funnel to fill the buret with about 50 mL of the HCl solution (titrant).

10. ☐ Drain a small amount of the titrant through the drop counter into the waste beaker.

11. ☐ Why is it necessary to drain a small amount of titrant through the drop counter before you begin a titration?

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12. ☐ Practice adjusting the stopcock on the buret so that the titrant goes through the drop counter in distinguishable drops that fall at about 2 to 3 drops per second.

Note: It is important that you have good control of adjusting the stopcock. If you accidently open the stopcock too far and the HCl flows out (as opposed to drops out), then you will have to start over.

13. ☐ How can the green light on the drop counter help you determine if the HCl is falling in distinguishable drops?

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14. ☐ Close the stopcock and then remove the waste container.

Data Collection

Table 1: Titration data

Measurement	Trial 1	Trial 2
Concentration of the HCl solution (M)		
Volume of Na_2CO_3 used (mL)		
Initial volume of HCl in the buret (mL)		
Final volume of HCl in the buret (mL)		

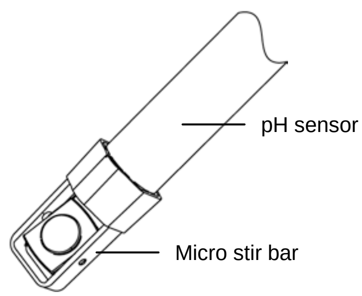
15. ☐ Record the exact concentration of the standardized HCl solution in Table 1 above.

16. ☐ Using the 100-mL graduated cylinder, measure 100.0 mL of distilled water and add it to a 250-mL beaker.

17. ☐ Using the 50-mL graduated cylinder, measure approximately 20.0 mL of the sodium carbonate (Na_2CO_3) solution and add to the 100 mL of distilled water in the beaker.

18. ☐ Record the exact volume of Na_2CO_3 used in Table 1 above.

19. ☐ Add the micro stir bar to the end of the pH sensor.



20. ☐ Position the 250-mL beaker containing the sodium carbonate on the magnetic stirrer with the pH sensor submerged in the solution.

21. ☐ Ensure that the bulb of the pH sensor is fully submerged in the solution and then turn on the magnetic stirrer and begin stirring at a slow-to-medium speed.

22. ☐ Determine the initial volume of the titrant (HCl) in the buret to a precision of 0.01 mL, and record this in Table 1 above.

23. ☐ Start recording data. �(6.2)

24. ☐ Carefully open the stopcock on the buret so that 2 to 3 drops per second are released. If needed, rescale the axes so that you can see the changes taking place. �(7.1.2)

25. ☐ Continue to record data until the pH of the solution stabilizes at a pH less than 2. Do not let the volume of the titrant fall below the last mark on the buret.

26. ☐ Why do you need to ensure that the volume of the titrant does not fall below the zero mark on the buret?

27. ☐ Which reaction step is taking place at the beginning of the titration between the initial pH (~12) and a pH value of 8? Describe the reaction that is occurring.

28. ☐ What begins to happen when the pH of the solution is lower than 8? How can these observations be explained?

29. ☐ Close the stopcock on the buret when the pH of the solution stabilizes at a pH less than 2.

30. ☐ Stop data recording. �(6.2)

31. ☐ Name the data run "trial 1". �(8.2)

32. ☐ Determine the final volume of the titrant in the buret and record the volume to 0.01 mL in Table 1 above.

33. ☐ Turn off the magnetic stirrer.

34. ☐ Remove the beaker and dispose of its contents according to the teacher's instructions.

35. ☐ Place the waste container under the pH sensor and use the wash bottle to thoroughly clean the micro stir bar and the pH sensor.

36. ☐ Dispose of this waste according to the teacher's instructions.

37. ☐ Thoroughly clean the 250-mL beaker so that you can reuse it for trial 2.

38. ☐ Collect a second set of data. Refer to the steps above as needed. Record the data in the Trial 2 column of Table 1 above.

39. ☐ Name the second run of data "trial 2".

40. ☐ Why is it necessary to take multiple sets of data?

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41. ☐ Save your data file and clean up according to the teacher's instructions. �(11.1)

Data Analysis

1. ☐ Determine the total volume of HCl titrant used in each trial. Record the total volume used in Table 2 below.

Total volume used = Final volume HCl – Initial volume HCl

Table 2: Total volumes of HCl titrant used in each trial

	Trial 1	**Trial 2**
Final volume of HCl (mL)		
Initial volume of HCl (mL)		
Total Volume HCl used (mL)		

2. ☐ Use the pH versus Drop Count (drops) graph to determine the total number of drops used in each trial. Follow the steps below to complete this on your data collection system.

   a. Display the run of data you want to analyze. �(7.1.7)

   b. Find the final drop count by finding the coordinates of the last data point collected. �(9.1)

   c. Record the final drop count for each trial in Table 3 below.

Table 3: Final drop count

	Trial 1	**Trial 2**
Final drop count		

3. ☐ Create a calculation to convert drop count to volume (mL) for each trial. Follow the steps below to do this on your data collection system.

   a. Write the mathematical equation that can be used to convert drop count to volume. The general equation is given below, but you need to replace "total volume of titrant used" and "final drop count" with the numerical values determined above. Write the mathematical equation for each trial in Table 4 below.

$$\text{calcvolume} = [\text{Drop Count}] * (\text{total volume of titrant used} / \text{final drop count})$$

Note: In the equation above "calcvolume" stands for calculated volume. You will have a different calculation for each trial so the two calculated volumes need to have different names.

   b. Enter the equations you determined above into the data collection system. �(10.3)

Table 4: Calculations for converting HCl drop count to volume

Mathematical equation for trial 1:
Mathematical equation for trial 2:

4. ☐ Determine the volume of HCl at both equivalence points for each trial. Follow the steps below to do this on your data collection system.

   a. Change the units on the x-axis to the calculated volume for that run of data you want to analyze. �(7.1.9)

   b. Display the run of data you want to analyze. �(7.1.7)

   c. Find the coordinates of each of the equivalence points. �(9.3)

   d. Record the volume of HCl at each of the equivalence points in Table 5 below.

Table 5: Volume of titrant at equivalence points

	Trial 1	**Trial 2**
Volume at Equivalence Point 1 (mL)		
Volume at Equivalence Point 2 (mL)		

5. ☐ Calculate the molar concentration of the sodium carbonate solution using the first equivalence point. Use the following steps as a guide and record your work for each step in Table 6 below.

   a. Determine the number of moles of HCl added at the first equivalence point using the volume at equivalence point #1 and the molarity of the HCl solution.

   b. Convert from moles of HCl to moles of sodium carbonate using the balanced chemical equation for the first reaction step.

   c. Use the moles of sodium carbonate calculated and the volume of sodium carbonate used as the analyte to determine the molarity of the base.

   d. Repeat for Trial 2.

Table 6: Molar concentration of the sodium carbonate solution at the first equivalence point

	Trial 1 (show your work)	**Trial 2**
Moles of HCl at the first equivalence point (mol)		
Balanced chemical equation for the first reaction		
Moles of sodium carbonate in solution (mol)		
Concentration of Na_2CO_3 solution (M)		

6. ☐ Calculate the molar concentration of the sodium carbonate solution using the second equivalence point. Use the following steps as a guide and record your work for each step in Table 7 below.

   a. Determine the number of moles of HCl added at the second equivalence point using the volume at equivalence point 2 and the molarity of the HCl solution.

   b. Convert from moles of HCl to moles of sodium carbonate using the overall balanced chemical equation.

   c. Use the moles of sodium carbonate calculated and the volume of sodium carbonate used as the analyte to determine the molarity of the base.

   d. Repeat for Trial 2.

Table 7: Molar concentration of the sodium carbonate solution at the second equivalence point

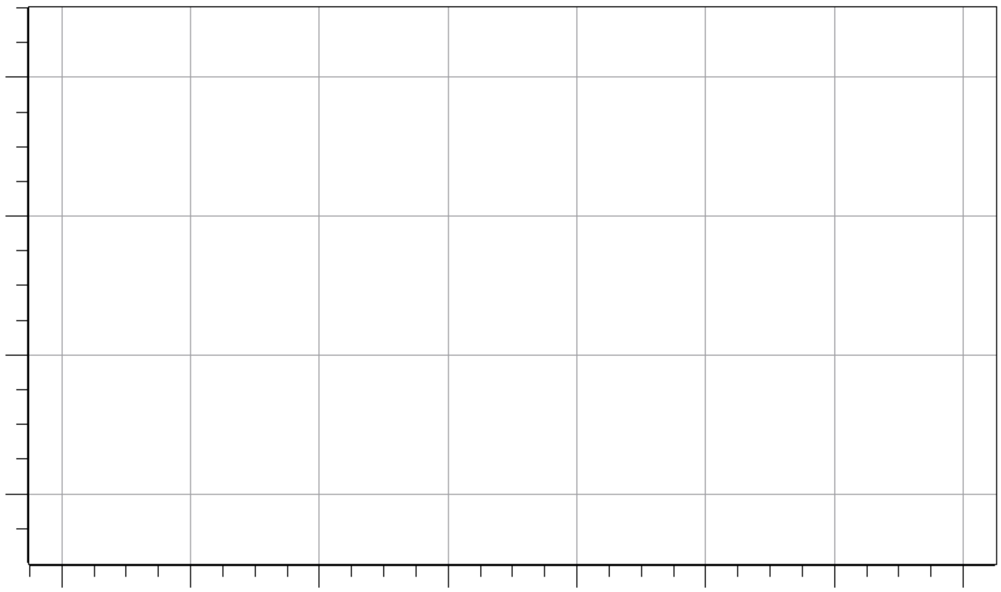
	Trial 1 (show your work)	**Trial 2**
Moles of HCl at the second equivalence point (mol)		
Balanced chemical equation for the overall reaction		
Moles of sodium carbonate in solution (mol)		
Concentration of Na_2CO_3 solution (M)		

7. ☐ Calculate the average molarity of the sodium carbonate solution using all four calculated values.

8. ☐ On your data collection system, create a graph of pH versus Volume of HCl (mL) with both runs of data displayed on the same set of axes. �(7.1.3)

Note: Use either of the calculated volumes on the y-axis. This graph is for comparison purposes only.

9. ☐ Sketch or print a copy of the graph of pH versus Volume of HCl (mL) with both runs of data on one set of axes. Label each run of data as well as the overall graph, the x-axis, the y-axis, and include numbers on the axes. �(11.2)



Analysis Questions

1. **When did the bubbles produced by the formation of carbon dioxide gas start to become visible?**

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2. **Write the two chemical equations as separate steps that add together to give the overall reaction: $2HCl(aq) + Na_2CO_3(aq) \rightarrow CO_2(g) + 2NaCl(aq) + H_2O(l)$**

Step 1: ____

Step 2: ____

3. **Before the first equivalence point, the beaker contained a mixture of both carbonate ions and bicarbonate ions. Which of these two ions accept hydrogen ions easier? How do you know?**

4. **After the first equivalence point, the carbon dioxide production was rapid. Were there any carbonate ions remaining in solution?**

5. **What do you notice about the volume of acid needed to reach each equivalence point?**

6. **Why do you think the bubbling stopped after the second equivalence point?**

Synthesis Questions

Use available resources to help you answer the following questions.

1. **Often, the product of one chemical reaction can become the reactant in another. Give an example of how this statement is true using the reactions seen in this experiment.**

2. **Calculate the number of moles of carbon dioxide produced in one of the trials. Identify the limiting reactant. Show your work, including the balanced chemical equation used.**

3. **Calculate the number of liters of carbon dioxide produced. Assume standard temperature and pressure.**

Multiple Choice Questions

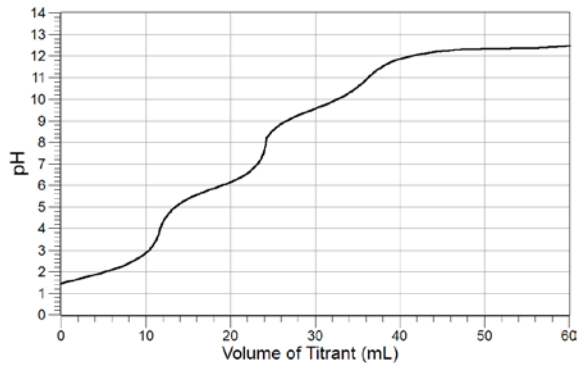
Select the best answer or completion to each of the questions or incomplete statements below.

1. **In an acid-base titration, the equivalence point of the curve shows:**
    A. The point where there are the same number of moles of acid and base
    B. The point where the buret has run out of acid
    C. The point where pH is a maximum
    D. The point where pH is a minimum

2. **Carbonic acid is known to be a *diprotic* acid. Why?**
    A. Because it reacts with a base
    B. Because it has two hydrogen ions that can be donated
    C. Because it makes carbon dioxide
    D. Because it can act as an acid or a base

3. **How many equivalence points are there on a titration of a diprotic acid?**
    A. 0
    B. 1
    C. 2
    D. 3

4. **The titration curve below represents what type of acid?**



A. Monoprotic acid

B. Diprotic acid

C. Triprotic acid

D. Tetraprotic acid

5. **What gas is released when an acid reacts with sodium carbonate?**

A. Hydrogen gas

B. Oxygen gas

C. Methane gas

D. Carbon dioxide gas

Key Term Challenge

Fill in the blanks from the list of words in the Key Term Challenge Word Bank.

1. The Brønsted-Lowry definition of acids and bases states that ________________________ donate hydrogen ions and ________________________ accept hydrogen ions. When an acid ________________________ a hydrogen ion, the remaining anionic portion is called the ________________________. The carbonate ion used in this experiment has a negative two charge. It can accept ________________________ hydrogen ions to become carbonic acid and is an example of a ________________________ acid. When the carbonate ions react with acid in the first step, they initially become ________________________ ions (write both the name and formula). These ions can then react in the second step with additional acid to become ________________________ (write both the name and formula). Carbonic acid decomposes into ________________________ and bubbles of ________________________. The bubbles begin to form after the ________________________ equivalence point. The bubbles will continue until the original ________________________ (write both the name and formula) is completely consumed.

Key Term Challenge Word Bank

Paragraph 1

accepts

acids

bases

bicarbonate (HCO_3^-)

carbon dioxide

carbonate (CO_3^{2-})

carbonic acid (H_2CO_3)

conjugate acid

conjugate base

diprotic

donates

first

hydrochloric acid (HCl)

hydrogen

monoprotic

one

oxygen

second

sodium carbonate (Na_2CO_3)

three

triprotic

two

water