

## P<sub>H</sub> TITRATION CURVE

### Introduction:

One method a chemist can use to investigate acid-base reactions is a titration. A p<sub>H</sub> titration is performed by adding small, accurate amounts of standard base to an acid of unknown concentration. The p<sub>H</sub> is recorded methodically and is plotted vs. the volume of base added to the acid solution. The result of this plot is an "S" shaped curve. The inflection point of this curve (middle of the "S") is indicative of the endpoint. This point occurs when the acid and base in solution are stoichiometrically equivalent. An endpoint can be very useful in determining the concentration of an acid or base.

### Purpose:

The purpose of this experiment is to determine the concentration of an unknown acid solution via the endpoint of an acid-base titration.

### Equipment / Materials:

HCl (unknown conc.)	250-ml. beaker
0.10 M NaOH	10mL. volumetric pipet
magnetic stirrer and stirring bar	pH meter and electrode
50-ml. buret	droppers
pH buffer solutions (pH 4, 7, and 10)	buret clamp
funnel	washbottle
150 ml. beaker	Kimwipes

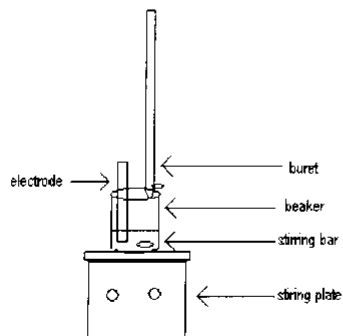
**Safety:**

- Always wear apron and safety glasses in the lab.
- Avoid contact of solutions with skin and clothing. If there is contact, rinse well with water and alert the instructor.
- Sodium hydroxide can cause instantaneous, permanent eye damage.

**Procedure:**

1. Plug in the pH meter and allow it to warm up for about 10 minutes.
2. The temperature knob should be set between 20 and 25°C.
3. Remove the cap from the electrode, and rinse the electrode with DI water. Blot the end with a Kimwipe.
4. Place the electrode in the pH 7 buffer, turn the knob to pH, and adjust the pH to 7.00 with the standardization knob.
5. Place the instrument on standby. Rinse and blot the electrode.
6. Place the electrode in pH 4.00 buffer, turn the knob to pH, and adjust the pH to 4.00 with the slope knob.
7. Place the instrument on standby. Rinse and blot the electrode.
8. Measure 10.00 mL of unknown HCl into a 250-mL beaker and dilute it with ~75 mL of distilled water.
9. Place the beaker on the magnetic stirrer and add a stirring bar.
10. Clean a 50-mL buret and rinse with ~ 5 mL NaOH solution.
11. Fill the buret past 0.00-mL mark with NaOH solution and deliver the excess into a waste beaker. Make sure the buret tip is full of NaOH solution. If the 0.00 mL is passed, use a dropper to get the reading to 0.00 mL.

12. Immerse the electrode in the HCl solution and turn on the stirring bar at a very slow rate. Higher speeds may cause the stirring bar to hit the electrode and damage it. Keep the electrode in the solution but **above** and to the side of the stirring bar. See the diagram.



13. Begin titrating, stopping to record the pH and the volume of NaOH to **two** decimal places at intervals of  $\sim 0.2$  pH units. When near the endpoint, a very small amount of NaOH will change the pH dramatically. When beginning the titration, it may be a good idea to practice adding one drop at a time or even part of a drop - that skill will be helpful when the endpoint is approached.
14. Continue the titration until a pH of 12 is obtained.
15. Plot a titration curve of pH (y axis) versus volume of NaOH (x axis).
16. Determine the endpoint of the titration (inflection point in the graph).
17. Repeat as instructed.
18. When completed rinse the electrode, but do not blot. Replace the cap on the electrode.









**Calculations:**

1. From the graph find the volume of base required to reach the end point.
2. Find the moles of base. (vol. base (L.) x conc base (M/L.) = moles).
3. Write and balance the equation for the reaction of HCl with NaOH.
4. Calculate the moles of acid required to react with the calculated moles of NaOH.
5. Determine the concentration of acid in moles/L.
6. Repeat for all trials and average.



**Questions:**

1. How would the following affect the results?
  - a. The buret is dirty and drops of NaOH cling to the side walls of the buret as it is drained.
  - b. The buret is not rinsed with NaOH prior to filling.
  - c. The buret tip is not filled at the start of the titration.
  - d. 20 mL of unknown acid are used instead of 10 mL.
  - e. The buret is not read at eye level. Explain.
  - f. The base is added too rapidly in the region of rapid pH change.
  - g. 100 mL of distilled water is used instead of 75 mL.
  - h. The original volume of base is not at 0.00 mL, and a student assumes it is.

## PH TITRATION CURVE

### TEACHER NOTES

**Lab Time:** 45 minutes

#### Preparations:

**Time:** 40 minutes

Prepare one liter of 0.10 M HCl by diluting 8.3 mL of concentrated HCl to one liter with deionized water. Each lab group will use at least 10 mL.

Prepare one liter of 0.10 M NaOH by dissolving 4.0 g of the solid to one liter with deionized water. Allow 100 mL per lab group per trial.

Turn on the pH meters about 10 minutes before class.

#### Answers to Questions:

1. How would the following affect your results?

- a. The buret is dirty and drops of NaOH cling to the side walls of the buret as it is drained.

*The actual volume of NaOH will be lowered and concentration of acid will be high.*

- b. The buret is not rinsed with NaOH prior to filling.

*If there was water in the buret the NaOH would be diluted, lowering the moles of base and the concentration of acid.*

- c. The buret tip is not filled at the start of the titration. The volume of NaOH will appear to be larger than it really is.

*This will make the concentration of the acid appear greater than it is.*

- d. 20 mL of unknown acid are used instead of 10 mL.

*If the student records the actual volume of acid used there would be no effect on the final concentration.*

- e. The base is added too rapidly in the region of rapid pH change.

*The end point will not be accurate, causing random errors.*

- f. The original volume of base is not at 0.00 mL, and a student assumes it is.

*It depends on if the liquid level was above or below the 0.00 mL line. If the level was below the 0.00 mL line, then less base was used causing a high acid concentration reading. If the level was above the 0.00 mL line, then more base was used causing a low acid reading.*

**Considerations:**

This lab serves as an introduction to titration with a pH meter and the construction of a titration curve. Be sure to remind students that the tip of the electrode is delicate and to keep it away from the stirring bar at all times. Students will need to take time with the titration since they have no idea where the end point will occur. If the pH meter does not standardize, be sure the buffer solution is fresh and the temperature is set correctly.

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## THE THEORY OF PH

The pH scale was suggested by Sorensen almost a century ago. We can think of it as a sort of shorthand for specifying the concentration of protons ( $H^+$  ions) in (usually aqueous) solution. Sorensen defined  $pH = -\log[H^+]$ . (Remember that [ ] means concentration or molarity.) In water, pH values from 0 (very acidic) to 14 (very basic) are possible. Thus pH 7, being right in the middle of the range, is neutral, that is, neither acidic nor basic. There are two tricky things about this scale: it's nonlinear and it seems "to go the wrong way" (high  $H^+$  corresponds to low pH). Both of these odd features stem from the fact that pH involves logarithms. First of all, a negative logarithm indicates an inversion or reciprocal of a number; that is why the scale seems "to go the wrong way." Also because of logarithms, the numbers seem much smaller:  $\log(10^7) = 7$ ,  $\log(10^{-2}) = -2$ ,  $\log(4.87 \times 10^{-3}) = 3.68$ . So a change the pH from 7 to 5, increases the proton concentration by 100 times!

Why does pH warrant its own rather weird scale? Mostly its due to the central role that water plays in chemistry, biochemistry, and biology. In as much as water is the solvent in which most chemical reactions occur, it's very important to consider ionization of water. In addition to un-ionized  $H_2O$ ,  $OH^-$  and  $H^+$  can be present. The concentration of  $OH^-$  ions can be expressed by pOH which is defined as  $pOH = -\log [OH^-]$ . There is a simple relationship between these ions:  $pH + pOH = 14$ . The pH specifies the ionic state of the most important solvent. Another reason for the widespread use of the pH scale is the ease of constructing electrochemical devices ("pH meters") to measure it.

The pH is a scale of numbers that simplifies manipulation, calculation, and discussions involving the wide range of hydrogen ions in a solution. Determination of pH is a determination of emf (voltage) difference in a pH cell. pH is dependent upon temperature; a temperature compensator varies the instrument definition of a pH unit from 54.20 mV to 66.10 mV at 60°C.

A pH meter consists of an electrode to sense pH, a reference electrode to complete the circuit while not affecting the pH measurement, an electronics module to link the electrode to the readout, and a display unit which could be a meter or a computer screen. Some pH meters also have a third electrode to sense the temperature in order to compensate for thermal effects. Often the electrodes are packaged into a "combination electrode" which is more convenient to use. The most frequent cause of pH meter malfunction is poorly-maintained electrodes.

In modern chemistry the "p" ("negative log of") notation has been extended to other species. Thus a chemist may use  $pS^-$ ,  $pNO_2^-$ ,  $pNa^+$ , and even  $p(\text{urea})$ . So-called ion-selective electrodes have been developed to measure all of those and many more.