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Acid–Base Titrations

AP* Chemistry Big Idea 6, Investigation 14

An Advanced Inquiry Lab

Introduction

Most products we use every day, including food, beverages, medication and cleaning solutions, have acidic or basic properties. A common question chemists have to answer is how much of a specific substance is present in a sample or a product. The amount or concentration of acid or base in a sample may be determined by acid–base titration. The strength of the acid or base being analyzed plays an important role in the experimental design.

Concepts

- Strong and weak acids
- Strong and weak bases
- Equilibrium constant, K_a
- Titration
- Indicators
- Equivalence point

Background

Titration is a method of *volumetric analysis*—the use of volume measurements to analyze an unknown. In acid–base chemistry, titration is most often used to analyze the amount of acid or base in a sample or solution. Consider a solution containing an unknown amount of hydrochloric acid. In a titration experiment, a known volume of the hydrochloric acid solution would be "titrated" by slowly adding dropwise a *standard* solution of a strong base such as sodium hydroxide. (A standard solution is one whose concentration is accurately known.) The titrant, sodium hydroxide in this case, reacts with and consumes the acid via a neutralization reaction (Equation 1). The exact volume of base needed to react completely with the acid is measured. This is called the equivalence point of the titration—the point at which stoichiometric amounts of the acid and base have combined.

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l) \qquad \textit{Equation 1}$$

Knowing the exact concentration and volume of the added titrant gives the number of moles of sodium hydroxide, which is, in turn, related by the mole ratio to the number of moles of hydrochloric acid initially present in the unknown.

Either acids or bases may be titrated to determine their concentration by choosing an appropriate standard solution as the titrant. Indicators are usually added to acid–base titrations to detect the equivalence point. The endpoint of the titration is the point at which the indicator changes color and signals that the equivalence point has indeed been reached. For example, in the case of the neutralization reaction shown in Equation 1, the pH of the solution would be acidic (< 7) before the equivalence point and basic (> 7) after the equivalence point if excess sodium hydroxide is added. The pH at the equivalence point should be exactly 7, corresponding to the neutral products—sodium chloride and water. An indicator that changes color around pH 7 is therefore a suitable indicator for the titration of a strong acid with a strong base.

The progress of an acid–base titration can also be followed by measuring the pH of the solution being analyzed as a function of the volume of titrant added. A plot of the resulting data is called a pH curve or titration curve. Titration curves allow a precise determination of the equivalence point of the titration without the use of an indicator.

The graph of pH versus volume of NaOH added for the titration of HCl is shown in Figure 1. Note the significant change in pH in the vicinity of the equivalence point.

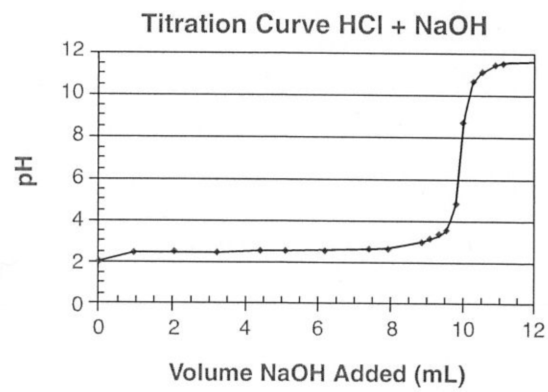

Figure 1.

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When a weak acid is titrated with a strong base, the equivalence point is not at pH 7, but rather is on the basic side. The value of the equilibrium constant for the dissociation of a weak acid can be obtained from its titration curve with a strong base. The shape of the titration curve for a weak acid with a strong base is explored in the *Pre-Lab Questions,* along with the equilibrium constant determination.

Experiment Overview

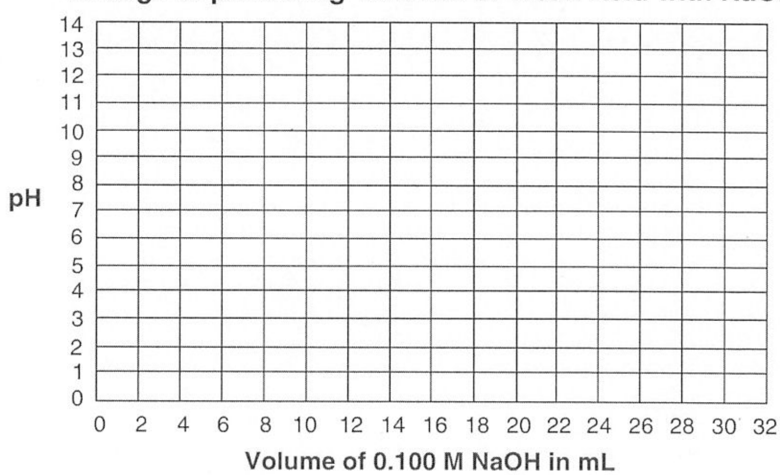
The purpose of this advanced inquiry lab activity is to conduct a series of acid–base titrations to determine the concentration of two unknowns. The lab begins with an introductory activity in which students qualitatively analyze an acid and base using pH paper. "Rough" titrations of the acid and base will be conducted and determination of endpoints will be tested with the indicators provided. At the end of each rough titration, pH paper will be used to determine if the final solution is basic or acidic. The activity provides a model for a guided-inquiry experiment, during which students collect quantitative titration data using a buret and pH meter. Each group uses two acids and two bases. One of the acids will have a known molarity and the other will have an unknown molarity. The same applies to the two bases. Students will graph titration curves from the collected data and determine the concentrations of each unknown. A variety of acids and bases, strong and weak, are provided for the class to perform different combinations of titrations.

Pre-Lab Questions

1. Predict whether the pH at the equivalence point will be acidic, basic, or neutral for the following classic titrations. Explain based on the properties of the conjugate acid–base pairs: *a)* strong acid with a strong base, *b)* weak acid with a strong base, *c)* strong base with a strong acid, *d)* weak base with a strong acid.
2. Distinguish between a strong acid and a weak acid in terms of their dissociation reactions and equilibrium constants.
3. The following values were collected for the titration of a solid weak acid with 0.100 M NaOH as the titrant. Graph the data in the chart provided and identify the pH at the equivalence point.

Volume NaOH, mL	pH
0.00	2.88
5.00	4.15
10.00	4.58
12.50	4.76
15.00	4.93
20.00	5.36
24.00	6.14
24.90	7.15
25.00	8.73
26.00	11.29
30.00	11.96

Change of pH During Titration of Weak Acid with NaOH



4. The equilibrium constant K_a for dissociation of a weak acid HA can be determined from its titration curve with a strong base. See Equations 2 and 3.

$$HA + H_2O \overset{K_a}{\rightleftarrows} A^- + H_3O^+ \qquad \textit{Equation 2}$$

$$K_a = \frac{[A^-][H_3O^+]}{[HA]} \qquad \textit{Equation 3}$$

When $[A^-] = [HA]$, K_a reduces to $[H_3O^+]$, and $pK_a = -\log K_a = pH$.

This condition is met at the "half-equivalence point" in the titration. When one-half the volume of NaOH needed to reach the equivalence point has been added, the pH of the solution is equal to the pK_a value for the acid. Draw dashed lines on the titration curve in Question 3 to illustrate these relationships, and estimate the pK_a value for the weak acid.

5. The following acid–base indicators are available to follow the titration shown in Question 3. Which indicator would be most appropriate for signaling the endpoint of the titration? Explain, and give the expected color to look for at the endpoint.

Indicator	Acid Form	Base Form	pH Transition Interval
Bromphenol blue	yellow	purple	3.0–4.6
Bromthymol blue	yellow	blue	6.0–7.6
Thymol blue	yellow	blue	8.0–9.2

Materials*

Acetic acid, CH_3CO_2H
Ammonia, NH_3
Calcium hydroxide, $Ca(OH)_2$, 0.200–0.300 g, unknown only
Hydrochloric acid, HCl
Methyl red indicator
Nitric acid, HNO_3
pH paper
Phenolphthalein indicator
Sodium hydroxide, NaOH
Sulfuric acid, H_2SO_4
Thymolphthalein indicator
Water, distilled or deionized

Beakers, 50-mL, 150-mL, and 250-mL
Beral-type pipets, graduated
Buret, 50-mL
Graduated cylinders, 10-mL and 100-mL
Magnetic stirrer and stir bar, or stirring rod
pH sensor or pH meter
Support stand and buret clamp
Test tube rack
Test tubes, medium, 4
Wash bottle

**Amounts and concentrations of chemicals for each group will depend on the two selected acid–base titrations approved by the instructor.*

Safety Precautions

Phenolphthalein and thymolphthalein solutions contain alcohol and are flammable liquids; they are toxic by ingestion. Do not use near flames or other sources of ignition. Dilute sodium hydroxide solution is slightly toxic by ingestion and skin absorption and is irritating to skin and eyes. The ammonia water solution is mildly toxic by ingestion and inhalation, irritating to body tissues, and a lachrymator. Solid calcium hydroxide is toxic by inhalation, irritating to body tissue, and its solution is caustic. Acetic acid may cause respiratory tract irritation. Hydrochloric acid is slightly toxic by inhalation and ingestion, a severe body tissue irritant, and corrosive to eyes. Sulfuric acid is slightly toxic by ingestion and severely irritating to body tissues, especially eyes. Nitric acid is irritating to body tissues. Avoid contact of all chemicals with eyes and skin. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Wash hands thoroughly with soap and water before leaving the lab.

Introductory Activity

Selecting an Indicator and Estimating the Endpoint

1. Choose an acid and a base. Obtain about 5 mL (50–100 mg of solid $Ca(OH)_2$ and add to 2–5 mL deionized water) of each in separate test tubes and test each solution with pH paper. Determine the initial pH of each solution. Pour the solutions together and measure the pH of the final solution using pH paper or a pH meter.
2. Identify the acid and base as either a strong or weak acid, and a strong or weak base, respectively, according to its name, structure and initial pH.
3. Based on the pH of the combined solution, predict whether the acid and the base have approximately equal concentrations, or if they are different. Explain your reasoning.

4. Review the list of indicators in the *Materials* section. Look up the acidic and basic colors of each indicator, and its pH transition range. Enter the information in a table, and choose a suitable indicator for titration of the acid and base in this activity.
5. Pour 5.0 mL of the acid into a clean test tube and add 1–2 drops of the appropriate indicator. Note the color.
6. Obtain 10.0 mL of the base in a clean graduated cylinder.
7. Using a graduated pipet, add the base in 1-mL increments to the acid in the test tube.
8. Observe the indicator color, and record the amount of base that has been added when the indicator color changes.
9. Estimate the *relative* concentrations of the acid and the base. *Hint:* Consider the structures of the acid and base, that is, if they are mono- or diprotic.

Guided-Inquiry Design and Procedure

Acid–Base Titration Curves and the Concentrations of Unknowns

1. With the instructor's approval, select two acids and two bases to analyze by titration. Record the name and formula of each compound, and identify each as a strong or weak acid and base, respectively.
2. Obtain the known concentration of one each of the acids and bases. These will be used as the *titrants* in this investigation.
3. Select an appropriate indicator for titration of the unknown acid and base, respectively.
4. Set up a buret with a clamp and support stand, a pH meter or sensor, a beaker, or flask and a magnetic stirrer, if available (see Figure 2). Explain why it is desirable to clean and rinse the buret with the titrant before beginning the titration.
5. Is it necessary to know the exact volume of the "unknown" acid or base to be titrated? Explain.
6. It is helpful to occasionally rinse the sides of the beaker or flask with distilled water during the titration procedure. Explain why or why not it is necessary to measure the volume of rinse water used during the procedure.

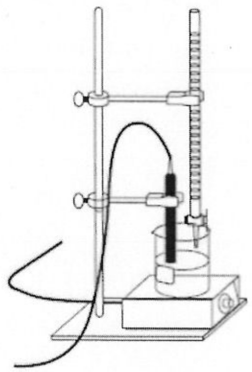


Figure 2.

7. Write a detailed, step-by-step procedure for titrating each unknown and obtaining the data for the titration curve. Include the chemicals needed, the indicators, and the safety precautions required. Describe the glassware, equipment, and techniques needed to ensure accuracy and precision in the quantitative analysis. Choose an amount (volume) of unknown for a convenient titration. Review the hazards of all chemicals and write appropriate safety precautions that must be followed.
8. Perform the titrations and graph the data collected. Label the important regions of each titration curve.
9. Obtain permission to repeat each titration a second time. *Note:* It is not necessary to use the pH meter for the second titration.
10. Average the titration data for each unknown and determine its concentration. Show all calculations.

Opportunities for Inquiry

Acid–Base Titration Curves

Randomly assign various combinations of strong and weak acids and bases in the guided-inquiry activity. Incorporate the titration curve results into a collaborative class project comparing and contrasting the behavior of strong versus weak acids. Students should share their data in larger groups to analyze four classic titration curves: (1) strong acid/strong base; (2) weak acid/strong base; (3) strong base/strong acid; and (4) weak base/strong acid. Students should compare the initial pH of the solution, the pH at the equivalence point, and the pH at the midpoint of the titration curve (for weak acids and weak bases, respectively). They may also reflect on the general shape of the titration curve and its relevance to the properties of buffers. This is an excellent culminating-type activity to assess student understanding of the principles of acid–base chemistry.

AP Chemistry Review Questions

Vinegar is a dilute aqueous solution of acetic acid produced by the fermentation of apple juice (cider vinegar), grapes (wine vinegar), or barley malt (malt vinegar). Federal regulations require that vinegar contain at least 4% acetic acid by mass. If the amount of acetic acid is less than 4%, the acidity level may not be high enough to prevent the growth of bacteria in pickled or canned foods. The amount of acetic acid in vinegar can be determined by microscale titration with a standard solution of sodium hydroxide.

1. Write the balanced chemical equation for the reaction of acetic acid with sodium hydroxide.

In the microscale titration, the exact number of drops of sodium hydroxide of known molarity (0.50 M) needed to react completely with a measured number of drops of vinegar was counted. Assume the amount of vinegar was 15 drops for each trial.

Brand of Vinegar	White Cider Vinegar
Titration Trial	**Number of Drops of NaOH Added**
1	24 drops
2	27 drops
3	24 drops
4	26 drops
5	26 drops

2. Calculate the average molarity of acetic acid in the vinegar using the data provided.

3. Calculate the average percent acetic acid in the vinegar using the following equation.

$$\%\text{ acetic acid} = \frac{\text{g (acetic acid)}}{\text{mL (vinegar)}} \times 100\%$$