Selected Methods of Analysis

This chapter contains detailed directions for performing a variety of chemical analyses. The methods have been chosen to introduce you to analytical techniques that are widely used by chemists. For most of these analyses, the composition of the samples is known to the instructor. Thus, you will be able to judge how well you are mastering these techniques.

Your chances of success in the laboratory will greatly improve if you take time before beginning any analysis to read carefully and understand each step in the method and to develop a plan for how and when you will perform each step. For greatest efficiency, such study and planning should take place before you enter the laboratory.

The discussion in this introduction is aimed at helping you develop efficient work habits in the laboratory; it also provides you with some general information about an analytical chemistry laboratory.

Background Information

Before you start an analysis, you should understand the significance of each step in the procedure to avoid the pitfalls and potential sources of error that exist in all analytical methods. Information about these steps can usually be found in preliminary discussion sections, in earlier chapters that are referred to in the discussion section, and in the “Notes” that follow many of the procedures. If, after reading these materials, you still do not understand the reason for one or more of the steps in the method, consult your instructor before you begin laboratory work.

The Accuracy of Measurements

In looking over an analytical procedure, you should decide which measurements must be made with maximum precision, and thus with maximum care, as opposed to those that can be carried out rapidly with little concern for precision.
Generally, measurements that appear in the equation used to compute the results must be performed with maximum precision. The remaining measurements can and should be made less carefully to conserve time. The words “about” and “approximately” are frequently used to indicate that a measurement does not have to be done carefully. You should not waste time and effort to measure, let us say, a volume to ± 0.02 mL when an uncertainty of ± 0.5 mL or even ± 5 mL will have no discernible effect on the results.

In some procedures, a statement such as “weigh three 0.5-g samples to the nearest 0.1 mg” is encountered. Here, samples of perhaps 0.4 to 0.6 g are acceptable, but their weights must be known to the nearest 0.1 mg. The number of significant figures in the specification of a volume or a weight is also a guide as to the care that should be taken in making a measurement. For example, the statement “add 10.00 mL of a solution to the beaker” indicates that you should measure the volume carefully with a buret or a pipet with the aim of limiting the uncertainty to perhaps ± 0.02 mL. In contrast, if the directions read “add 10 mL,” the measurement can be made with a graduated cylinder.

**Planning Your Time**

You should study carefully the time requirements of the several unit operations involved in an analysis before work is started. Such study will reveal operations that require considerable elapsed, or clock, time but little or no operator time. Examples of such operations include drying a sample in an oven, cooling a sample in a desiccator, or evaporating liquid on a hot plate. The experienced chemist plans to use such periods of waiting to perform other operations or perhaps to begin a new analysis. Some people find it worthwhile to prepare a written time schedule for each laboratory period to avoid periods when no work can be done.

Time planning is also needed to identify places where an analysis can be interrupted for overnight or longer as well as those operations that must be completed without a break.

**Reagents**

Directions for the preparation of reagents accompany many of the procedures. Before preparing such reagents, be sure and check to see if they are already prepared and available on a side shelf for general use.

If a reagent is known to pose a hazard, you should plan in advance of the laboratory period the steps that you should take to minimize injury or damage. Furthermore, you must acquaint yourself with the rules that apply in your
laboratory for the disposal of waste liquids and solids. These rules vary from one location to another and even among laboratories in the same locale.

**Water**

Some laboratories use deionizers to purify water; others employ stills for this purpose. The terms “distilled water” and “deionized water” are used interchangeably in the directions that follow. Either type is satisfactory for the procedures in this chapter.

You should use tap water only for preliminary cleaning of glassware. The cleaned glassware is then rinsed with at least three small portions of distilled or deionized water.

**22A An Introductory Experiment**

The purpose of this experiment is to introduce several of the tools, techniques, and skills necessary for work in the analytical chemistry laboratory. The techniques are considered one at a time, as unit operations. It is important to learn proper techniques and acquire individual skills before attempting additional laboratory experiments.

**22A-1 Using the Analytical Balance**

**Discussion**

In this experiment, you will obtain the mass of five new pennies, first by determining the mass of each penny individually. Then you will determine the mass of all five pennies at once, remove one penny at a time, and calculate the individual masses of the pennies by difference. The pair of masses determined for a particular penny by the two different methods should agree to within a few tenths of a milligram. From the data, you will determine the mean and median values, the standard deviation, and the relative standard deviation of the masses of the pennies.

You will then weigh an unknown aluminum cylinder and report the mass of this unknown.

**Procedure**

1. After you have been instructed in the use of the balance and have become familiar with its use, obtain a set of pennies, an unknown aluminum cylinder, and a pair of tweezers from the instructor.
2. Do not handle the pennies or the cylinder with your fingers; always use tweezers. If you are using a mechanical balance, be sure to have the balance in the “off” or “complete arrest” position whenever you remove anything from or add anything to the balance pan.
22A-2 Making Quantitative Transfers

Discussion
The following experiment is designed to provide experience in the correct use of the volumetric flask.

Procedure

1. Weigh a 50-mL beaker on a triple beam balance.
2. Adjust the balance for an additional 0.4 g, and add solid KMnO₄ to the beaker until the beam is again balanced. Note that chemicals should never be returned to stock bottles, because they may contaminate the bottle.
3. Dissolve the potassium permanganate in about 20 mL of distilled water, stirring gently to avoid loss. This solution is nearly saturated, and some care is required to dissolve the crystals completely.
4. Quantitatively transfer the solution to a 100-mL volumetric flask fitted with a small funnel. To prevent solution from running down the outside of the beaker, pour it down the stirring rod, and then touch the rod to the spout of the beaker to remove the last drop. Add more water to the beaker, stir, and repeat the procedure.
5. Repeat the procedure until no trace of the color of the permanganate remains in the beaker. Note the number of washings required to quantitatively transfer the permanganate from the beaker to the flask.
6. Rinse the last portions of solution from the stirring rod into the volumetric flask with a stream of water from the wash bottle. Rinse the funnel and remove it. Dilute the solution in the flask until the bottom of the meniscus is even with the graduation mark. Stopper, invert, and shake the flask. Return it to the upright position, and allow the air bubble to return all the way to the top of the neck.
7. Repeat until the solution is completely homogeneous; about 10 inversions and shakings are required. Save the solution for Section 22A-3.
**22A-3 Delivering an Aliquot**

**Discussion**
Whenever a buret or pipet is used to deliver a measured volume of solution, the liquid it contains before measurement should have the same composition as the solution to be dispensed. The following operations are designed to illustrate how to rinse and fill a pipet and how to deliver an aliquot of solution.

**Procedure**

1. Fill a pipet with the solution of potassium permanganate, and let it drain.
2. Draw a few milliliters of distilled water from a 50-mL beaker into the pipet, rinse all internal surfaces of the pipet, and discard the rinse solution. Do not fill the pipet completely; this step is wasteful, time consuming, and inefficient. Just draw in a small amount, tilt the pipet horizontally, and turn it to rinse the sides.
3. Determine the minimum number of such rinsings required to completely remove the permanganate color from the pipet. If your technique is efficient, three rinsings will suffice.
4. Again fill the pipet with permanganate solution, and proceed as before. This time, determine the minimum volume of rinse water required to remove the color by collecting the rinsings in a graduated cylinder. Less than 5 mL is enough with efficient technique. In the rinsing operations, was the water in the 50-mL beaker contaminated with permanganate? If a pink color shows that it was, repeat the exercise with more care.
5. As a test of your technique, ask the laboratory instructor to observe and comment on the following operation: Rinse a 10-mL pipet several times with the solution of potassium permanganate you prepared.
6. Pipet 10 mL of the permanganate solution into a 250-mL volumetric flask.
7. Carefully dilute the solution to volume, trying to mix the contents as little as possible.
8. Mix the solution by repeatedly inverting and shaking the flask. Note the effort required to disperse the permanganate color uniformly throughout the solution.
9. Rinse the pipet with the solution in the volumetric flask. Pipet a 10-mL aliquot of the solution into a conical flask.

**22A-4 Calibrating a Pipet**

**Discussion**
The proper technique for calibrating an analytical transfer pipet is readily learned with practice, care, and attention to detail. Note that this manual technique can only be learned through practice. With the possible exception of a mass determination, this experiment has the potential of being the most accurate and precise set of measurements that you will ever make.
Procedure

1. Clean a 10-mL pipet. When a pipet, buret, or other piece of volumetric glassware is cleaned properly, no droplets of reagent remain on the internal surfaces when they are drained, which is very important for accurate and reproducible results. If reagent is retained inside a pipet, you cannot deliver the nominal volume. If you clean a pipet or any other glassware with alcoholic KOH, use the bottle of cleaning solution only inside the sink and rinse it off thoroughly before putting it back on the plastic matting. Do not put the bottle of cleaning solution directly on a benchtop; it may ruin the surface. The solution is very corrosive. If your fingers feel slippery after use, or if some part of your body develops an itch, wash the area thoroughly with water.

2. Obtain a pipetting bulb, a 50-mL Erlenmeyer flask with dry stopper, a 400-mL beaker of distilled water equilibrated to room temperature, and a thermometer.

3. Determine the mass of the flask and stopper, and record it to the nearest 0.1 mg. Do not touch the flask with your fingers after this weighing. Use tongs or a folded strip of waxed paper.

4. Determine and record the temperature of the water.

5. Pipet 10.00 mL of the distilled water into the flask using the technique described on page 45. Stopper the flask, determine the mass of the flask and the water that it contains, and record the mass.

6. In the same way, add a second pipet of water to the flask; remove the stopper just before the addition. Replace the stopper, and once again determine and record the mass of the flask and the water. Following each trial, determine the mass of water added to the flask.

7. Repeat this process until you have determined four consecutive masses of water that agree within a range of 0.02 g. If the determinations of the mass of water delivered by the pipet do not agree within this range, your pipetting technique may be suspect. Consult your instructor for assistance in finding the source of your error, and then repeat the experiment until you are able to deliver four consecutive volumes of water with the precision cited above.

8. Correct the mass for buoyancy and calculate the volume of the pipet in milliliters.

9. Report the mean, the standard deviation, and the relative standard deviation of the volume of your pipet. Calculate and report the 95% confidence interval for the volume of your pipet.

22A-5 Reading Buret Sections

Discussion

The following exercise will give you practice in reading a buret and will confirm the accuracy of your readings.

Procedure

1. Obtain a set of five buret sections from your instructor.

2. Invert each section, and tap the section lightly to remove any solvent that might remain in the sealed tip.
3. Record the number and reading of each buret section on the form provided. Use a buret reading card to observe the readings to the nearest 0.01 mL.

4. Check your readings of the buret sections against the known values provided by your instructor.

22A-6 Reading a Buret

Discussion
The following exercise demonstrates the proper way to use a buret.

Procedure

1. Mount a buret in a buret stand, and fill the buret with distilled water.
2. Wait at least 30 s before taking the initial reading. Use a buret reading card to take readings. A buret reading card can be easily constructed by applying a piece of black electrical tape to a three-by-five-inch card. Never adjust the volume of solution in a buret to exactly 0.00 mL. Attempting to do so will introduce bias into the measurement process and will waste time.
3. Now let about 5 mL run into a 250-mL Erlenmeyer flask. Wait at least 30 s, and take the “final reading.” The amount of solution in the Erlenmeyer flask is equal to the difference between the final reading and the initial reading. Record the final reading in your laboratory notebook, and then ask your instructor to take the final reading. Compare the two readings. They should agree within 0.01 mL. Notice that the final digit in the buret reading is the estimation of the distance between two consecutive 0.1-mL marks on the buret.
4. Refill the buret, and take a new zero reading. Now add 30 drops to the Erlenmeyer flask, and take the final reading. Calculate the mean volume of one drop, then repeat this step except with 40 drops, and again calculate the mean volume of a drop. Record these results and compare them.
5. Finally, practice adding half-drops to the flask. Calculate the mean volume of several half-drops, and compare your results with those that you obtained with full drops. When you perform titrations, you should attempt to determine end points to within half a drop to achieve good precision.

22A-7 Sampling

Discussion
In most analytical methods, only a small fraction of the entire population is analyzed. The results from the determination of an analyte in a laboratory sample are assumed to be similar to the concentration of the analyte in the whole population. Consequently, a laboratory sample taken from the entire batch must be representative of the population.

In this experiment, you will investigate how the sample size influences the uncertainty associated with the sampling step. Generally, the required sample

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size must increase as the sample heterogeneity increases, as the fraction of the analyte decreases, or as the desired uncertainty decreases. The model system used in this experiment consists of a collection of plastic beads that are identical in size, shape, and density but that are different in color. If \( p \) represents the fraction of the particles of the analyte (beads of the first color), then \( 1 - p \) is the fraction of the second type of particles (beads of the second color). If a sample of \( n \) particles is drawn from the population, the number of particles of the analyte in the sample should be \( np \). It can be shown that the standard deviation of the number of particles of analyte \( np \) obtained from a sample of the two-component mixture is \( \sqrt{np(1 - p)} \). The relative standard deviation (RSD) is then

\[
\text{RSD} = \frac{\sqrt{np(1 - p)}}{np} = \sqrt{\frac{1 - p}{np}}
\]

This equation suggests that as the number of particles sampled increases, the relative uncertainty decreases. Using a mixture of beads of two colors, you will determine the uncertainty of sampling as a function of sample size.

### Procedure

1. Stir the container of beads thoroughly, and withdraw a sample of beads using a small beaker. Make sure that the beaker is full to the top but not overflowing.
2. Empty the beads into a counting tray, and count the number of beads of each color.
3. Repeat step 1 using a medium-size beaker and then the larger beaker. Record the total number of beads in your sample and the percentage of beads of a particular color specified by your instructor. Each student in your class will collect and count three similar samples and enter the data on a class chart that will be provided by your instructor. After all data are entered, the chart will be copied and distributed to all students in your class.

### Calculations

1. Using the compiled class data, calculate the mean percentage of beads of the specified color and the relative standard deviation of that percentage for each sample size.
2. Using the equation given above based on sampling theory, calculate the theoretical relative standard deviation using the values of \( p \) and the mean number of particles for each of the three sample sizes.
3. Compare your class data with the theoretical result. Does the relative standard deviation decrease as the sample size increases as predicted by sampling theory?
4. Use the equation for the relative standard deviation to find the number of beads that would have to be sampled to achieve a relative standard deviation of 0.002.
5. Suggest two reasons why the theory above might not be adequate to describe the sampling of many materials for chemical analysis.

### 22B GRAVIMETRIC METHODS OF ANALYSIS

General aspects, calculations, and typical applications of gravimetric analysis are discussed in Chapter 4.
22B-1 The Gravimetric Determination of Chloride in a Soluble Sample

Discussion

The chloride content of a soluble salt can be determined by precipitation as silver chloride:

\[ \text{Ag}^+ + \text{Cl}^- \rightarrow \text{AgCl(s)} \]

The precipitate is collected in a weighed filtering crucible and washed. After the precipitate has been dried to constant mass at 110°C, its mass is then determined.

The solution containing the sample is kept slightly acidic during the precipitation to eliminate possible interference from anions of weak acids (such as \( \text{CO}_3^{2-} \)) that form sparingly soluble silver salts in a neutral environment. A moderate excess of silver ion is needed to diminish the solubility of silver chloride, but a large excess is avoided to minimize coprecipitation of silver nitrate.

Silver chloride forms first as a colloid and is subsequently coagulated with heat. Nitric acid and the small excess of silver nitrate promote coagulation by providing a moderately high electrolyte concentration. Nitric acid in the wash solution maintains the electrolyte concentration and eliminates the possibility of peptization during the washing step; the acid subsequently decomposes to give volatile products when the precipitate is dried. See Section 4A-2 for additional information concerning the properties and treatment of colloidal precipitates.

In common with other silver halides, finely divided silver chloride undergoes photodecomposition:

\[ 2\text{AgCl(s)} \xrightarrow{h\nu} 2\text{Ag(s)} + \text{Cl}_2(g) \]

The elemental silver produced in this reaction is responsible for the violet color that develops in the precipitate. In principle, this reaction leads to low results for chloride ion. In practice, however, its effect is negligible provided you avoid direct and prolonged exposure of the precipitate to sunlight.

If photodecomposition of silver chloride occurs before filtration, the additional reaction

\[ 3\text{Cl}_2(aq) + 3\text{H}_2\text{O} + 5\text{Ag}^+ \rightarrow 5\text{AgCl(s)} + \text{ClO}_3^- + 6\text{H}^+ \]

tends to cause high results.

Some photodecomposition of silver chloride is inevitable as the analysis is ordinarily performed. It is worthwhile to minimize exposure of the solid to intense sources of light as far as possible.

Because silver nitrate is expensive, any unused reagent should be collected in a storage container; similarly, precipitated silver chloride should be retained after the analysis is complete.\(^2\)

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Procedure

Clean three medium-porosity sintered-glass or porcelain filtering crucibles by allowing about 5 mL of concentrated HNO₃ to stand in each crucible for about 5 min. Use a vacuum (Figure 18-16) to draw the acid through the crucible. Rinse each crucible with three portions of tap water, and then discontinue the vacuum. Next, add about 5 mL of 6 M NH₃ and wait about 5 min before drawing it through the filter. Finally, rinse each crucible with six to eight portions of distilled or deionized water. Provide each crucible with an identifying mark. Dry the crucibles to constant mass by heating at 110°C while the other steps in the analysis are being carried out. The first drying should be for at least 1 hr; subsequent heating periods can be somewhat shorter (30 to 40 min). This process of heating and drying should be repeated until the mass becomes constant to within 0.2 to 0.3 mg.

Transfer the unknown to a weighing bottle and dry it at 110°C (Figure 18-9) for 1 to 2 hr; allow the bottle and contents to cool to room temperature in a desiccator. Weigh (to the nearest 0.1 mg) individual samples by difference into 400-mL beakers (note 1). Dissolve each sample in about 100 mL of distilled water to which 2 to 3 mL of 6 M HNO₃ has been added.

Slowly, and with good stirring, add 0.2 M AgNO₃ to each of the cold sample solutions until AgCl is observed to coagulate (notes 2 and 3), and then introduce an additional 3 to 5 mL. Heat almost to boiling, and digest the solids for about 10 min. Add a few drops of AgNO₃ to confirm that precipitation is complete. If more precipitate forms, add about 3 mL of AgNO₃, digest, and again test for completeness of precipitation. Pour any unused AgNO₃ into a waste container (not into the original reagent bottle). Cover each beaker, and store in a dark place for at least 2 hr and preferably until the next laboratory period.

Read the instructions for filtration in Section 18F-2. Decant the supernatant liquids through weighed filtering crucibles. Wash the precipitates several times (while they are still in the beaker) with a solution consisting of 2 to 5 mL of 6 M HNO₃ per liter of distilled water; decant these washings through the filters. Quantitatively transfer the AgCl from the beakers to the individual crucibles with fine streams of wash solution; use rubber policemen to dislodge any particles that adhere to the walls of the beakers. Continue washing until the filtrates are essentially free of Ag⁺ ion (note 4).

Dry the precipitate at 110°C for at least 1 hr. Store the crucibles in a desiccator while they cool. Determine the mass of the crucibles and their contents. Repeat the cycle of heating, cooling, and weighing until consecutive weighings agree to within 0.2 mg. Calculate the percentage of Cl⁻ in the sample.

Upon completion of the analysis, remove the precipitates by gently tapping the crucibles over a piece of glazed paper. Transfer the collected AgCl to a container for silver wastes. Remove the last traces of AgCl by filling the crucibles with 6 M NH₃ and allowing them to stand.

Notes

1. Consult with the instructor concerning an appropriate sample size.
2. Determine the approximate amount of AgNO₃ needed by calculating the volume that would be required if the unknown were pure NaCl.
3. Use a separate stirring rod for each sample, and leave it in its beaker throughout the determination.
4. To test the washings for Ag⁺, collect a small volume in a test tube and add a few drops of HCl. Washing is judged complete when little or no turbidity develops.
The nickel in a steel sample can be precipitated from a slightly alkaline medium with an alcoholic solution of dimethylglyoxime (Section 4D-3). Interference from iron(III) is eliminated by masking with tartaric acid. The product is freed of moisture by drying at 110°C.

The bulky character of nickel dimethylglyoxime limits the mass of nickel that can be accommodated conveniently and thus the sample mass. Care must also be taken to control the excess of alcoholic dimethylglyoxime used. If too much is added, the alcohol concentration becomes sufficient to dissolve appreciable amounts of the nickel dimethylglyoxime, which leads to low results. If the alcohol concentration becomes too low, however, some of the reagent may precipitate and cause a positive error.

Preparation of Solutions

1. **Dimethylglyoxime, 1% (w/v).** Dissolve 10 g of dimethylglyoxime in 1 L of ethanol. (This solution is sufficient for about 50 precipitations.)

2. **Tartaric acid, 15% (w/v).** Dissolve 225 g of tartaric acid in sufficient water to give 1500 mL of solution. Filter before use if the solution is not clear. (This solution is sufficient for about 50 precipitations.)

Procedure

Clean and mark three medium-porosity sintered-glass crucibles (note 1); bring them to constant mass by drying at 110°C for at least 1 hr.

Weigh (to the nearest 0.1 mg) samples containing between 30 and 35 mg of nickel into individual 400-mL beakers (note 2). In the hood, dissolve each sample in about 50 mL of 6 M HCl with gentle warming. Carefully add approximately 15 mL of 6 M HNO₃, and boil gently to expel any oxides of nitrogen that may have been produced. Dilute to about 200 mL, and heat to boiling. Introduce about 30 mL of 15% tartaric acid and sufficient concentrated NH₃ to produce a faint odor of NH₃ in the vapors over the solutions (note 3); then add another 1 to 2 mL of NH₃. If the solutions are not clear at this stage, proceed as directed in note 4. Make the solutions acidic with HCl (no odor of NH₃), heat to 60 to 80°C, and add about 20 mL of the 1% dimethylglyoxime solution. With good stirring, add 6 M NH₃ until a slight excess exists (faint odor of NH₃) plus an additional 1 to 2 mL. Digest the precipitates for 30 to 60 min, cool for at least 1 hr, and filter.

Wash the solids with water until the washings are free of Cl⁻ (note 5). Bring the crucibles and their contents to constant mass at 110°C. Report the percentage of nickel in the sample. The dried precipitate has the composition Ni(C₄H₇O₂N₂)₂ (288.92 g/mol).

Notes

1. Medium-porosity porcelain filtering crucibles or Gooch crucibles with glass pads can be substituted for sintered-glass crucibles in this determination.

2. Use a separate stirring rod for each sample, and leave it in the beaker throughout.
3. The presence or absence of excess NH₃ is readily established by odor; use a waving motion with your hand to waft the vapors toward your nose.

4. If Fe₂O₃·xH₂O forms upon addition of NH₃, acidify the solution with HCl, introduce additional tartaric acid, and neutralize again. Alternatively, remove the solid by filtration. Thorough washing with a hot NH₃/NH₄Cl solution is required; the washings are combined with the solution containing the bulk of the sample.

5. Test the washings for Cl⁻ by collecting a small portion in a test tube, acidifying with HNO₃, and adding a drop or two of 0.1 M AgNO₃. Washing is judged complete when little or no turbidity develops.

### 22C NEUTRALIZATION TITRATIONS

Neutralization titrations are performed with standard solutions of strong acids or bases. Although a single solution (of either acid or base) is sufficient for the titration of a given type of analyte, it is convenient to have standard solutions of both acid and base available in the event that a back-titration is needed to locate end points more exactly. The concentration of one solution is established by titration against a primary standard; the concentration of the other is then determined from the acid/base ratio (that is, the volume of acid needed to neutralize 1.000 mL of the base).

**22C-1 The Effect of Atmospheric Carbon Dioxide on Neutralization Titrations**

Water in equilibrium with the atmosphere is about 1 × 10⁻⁵ M in carbonic acid as a consequence of the equilibrium

\[
\text{CO}_2(g) + \text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{CO}_3(aq)
\]

At this concentration level, the amount of 0.1-M base consumed by the carbonic acid in a typical titration is negligible. With more dilute reagents (<0.05 M), however, the water used as a solvent for the analyte and in the preparation of reagents must be freed of carbonic acid by boiling for a brief period.

Water that has been purified by distillation rather than by deionization is often supersaturated with carbon dioxide and may thus contain sufficient acid to affect the results of an analysis.³ The instructions that follow are based on the assumption that the amount of carbon dioxide in the water supply can be neglected without causing serious error. For further discussion on the effects of carbon dioxide in neutralization titrations, see Section 9A-3.

³Water that is to be used for neutralization titrations can be tested by adding 5 drops of phenolphthalein to a 500-mL portion. Less than 0.2 to 0.3 mL of 0.1 M OH⁻ should suffice to produce the first faint pink color of the indicator. If a larger volume is needed, the water should be boiled and cooled before it is used to prepare standard solutions or to dissolve samples.
**22C-2 Preparation of Indicator Solutions for Neutralization Titrations**

**Discussion**
The theory of acid/base indicators is discussed in Section 7A-2. An indicator exists for virtually any pH range between 1 and 13.\(^4\) Directions follow for the preparation of indicator solutions suitable for most neutralization titrations.

**Procedure**
Stock solutions ordinarily contain between 0.5 and 1.0 g of indicator per liter. (One liter of indicator is sufficient for hundreds of titrations.)

1. *Bromocresol green*. Dissolve the sodium salt directly in distilled water.
2. *Phenolphthalein, thymolphthalein*. Dissolve the solid indicator in a solution consisting of 800 mL of ethanol and 200 mL of distilled or deionized water.

---

**22C-3 Preparation of Dilute Hydrochloric Acid Solutions**

**Discussion**
The preparation and standardization of acids are considered in Sections 9A-1 and 9A-2.

**Procedure**
For a 0.1-M solution, add about 8 mL of concentrated HCl to about 1 L of distilled water (note). Mix thoroughly, and store in a glass-stoppered bottle.

**Note**
It is advisable to eliminate CO\(_2\) from the water by a preliminary boiling if very dilute solutions (<0.05 M) are being prepared.

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**22C-4 Preparation of Carbonate-Free Sodium Hydroxide**

**Discussion**
See Sections 9A-3 and 9A-4 for information concerning the preparation and standardization of bases.

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Procedure

Transfer 1 L of distilled water to a polyethylene storage bottle (see the note in Section 27C-3). Decant 4 to 5 mL of 50% NaOH into a small container (note 2), add it to the water, and mix thoroughly. Use extreme care in handling 50% NaOH, which is highly corrosive. If the reagent comes into contact with skin, immediately flush the area with copious amounts of water.

Protect the solution from unnecessary contact with the atmosphere.

Notes

1. A solution of base that will be used up within two weeks can be stored in a tightly capped polyethylene bottle. After each removal of base, squeeze the bottle while tightening the cap to minimize the air space above the reagent. The bottle will become embrittled after extensive use as a container for bases.

2. Be certain that any solid Na₂CO₃ in the 50% NaOH has settled to the bottom of the container and that the decanted liquid is absolutely clear. If necessary, filter the base through a glass mat in a Gooch crucible; collect the clear filtrate in a test tube inserted in the filter flask.

22C-5  The Determination of the Acid/Base Ratio

Discussion

If both acid and base solutions have been prepared, it is useful to determine their volumetric combining ratio. Knowledge of this ratio and the concentration of one solution permits calculation of the molarity of the other.

Procedure

Instructions for placing a buret into service are given in Sections 18G-4 and 18G-6; consult these instructions if necessary. Place a test tube or small beaker over the top of the buret that holds the NaOH solution to minimize contact between the solution and the atmosphere.

Record the initial volumes of acid and base in the burets to the nearest 0.01 mL. Deliver 35 to 40 mL of the acid into a 250-mL conical flask. Touch the tip of the buret to the inside wall of the flask, and rinse down with a little distilled water. Add two drops of phenolphthalein (note 1) and then sufficient base to render the solution a definite pink. Introduce acid dropwise to discharge the color, and again rinse down the walls of the flask. Carefully add base until the solution again acquires a faint pink hue that persists for at least 30 s (notes 2, 3). Record the final buret volumes (again, to nearest 0.01 mL). Repeat the titration. Calculate the acid/base volume ratio. The ratios for duplicate titrations should agree to within 1 to 2 ppt. Perform additional titrations, if necessary, to achieve this order of precision.

Notes

1. The volume ratio can also be determined with an indicator that has an acid transition range, such as bromocresol green. If the NaOH is contaminated with carbonate, the ratio obtained with this indicator will differ significantly from the value obtained with phenolphthalein. In general, the acid/base ratio should be evaluated with the indicator that is to be used in subsequent titrations.
22C Neutralization Titrations

2. Fractional drops can be formed on the buret tip, touched to the wall of the flask, and then rinsed down with a small amount of water.

3. The phenolphthalein end point fades as CO$_2$ is absorbed from the atmosphere.

22C-6 Standardization of Hydrochloric Acid against Sodium Carbonate

Discussion

See Section 9A-2.

Procedure

Dry a quantity of primary-standard Na$_2$CO$_3$ for about 2 hr at 110°C (Figure 18-9), and cool in a desiccator. Weigh individual 0.20- to 0.25-g samples (to the nearest 0.1 mg) into 250-mL conical flasks, and dissolve each in about 50 mL of distilled water. Introduce 3 drops of bromocresol green, and titrate with HCl until the solution just begins to change from blue to green. Boil the solution for 2 to 3 min, cool to room temperature (note 1), and complete the titration (note 2).

Determine an indicator correction by titrating approximately 100 mL of 0.05 M NaCl and 3 drops of indicator. Boil briefly, cool, and complete the titration. Subtract any volume needed for the blank from the titration volumes. Calculate the concentration of the HCl solution.

Notes

1. The indicator should change from green to blue as CO$_2$ is removed during heating. If no color change occurs, an excess of acid was added originally. This excess can be back-titrated with base provided that the acid/base combining ratio is known; otherwise, the sample must be discarded.

2. It is permissible to back-titrate with base to establish the end point with greater certainty.

22C-7 Standardization of Sodium Hydroxide against Potassium Hydrogen Phthalate

Discussion

See Section 9A-4.

Procedure

Dry a quantity of primary-standard potassium hydrogen phthalate (KHP) for about 2 hr at 110°C (Figure 18-9), and cool in a desiccator. Weigh individual 0.7- to 0.8-g samples (to the nearest 0.1 mg) into 250-mL conical flasks, and dissolve each in 50 to 75 mL of distilled water. Add 2 drops of phenolphthalein; titrate with base until the pink color of the indicator persists for 30 s (note). Calculate the concentration of the NaOH solution.
22C-8  Determining Potassium Hydrogen Phthalate in an Impure Sample

Discussion
The unknown is a mixture of KHP and a neutral salt. This analysis is conveniently performed concurrently with the standardization of the base.

Procedure
Consult with the instructor concerning an appropriate sample size. Then follow the directions in Section 22C-7.

22C-9  Determining the Acid Content of Vinegars and Wines

Discussion
The total acid content of a vinegar or a wine is readily determined by titration with a standard base. It is customary to report the acid content of vinegar in terms of acetic acid, the principal acidic constituent, even though other acids are present. Similarly, the acid content of a wine is expressed as percent tartaric acid, even though there are other acids in the sample. Most vinegars contain about 5% acid (w/v) expressed as acetic acid; wines ordinarily contain somewhat under 1% acid (w/v) expressed as tartaric acid.

Procedure
1. If the unknown is a vinegar (note 1), pipet 25.00 mL into a 250-mL volumetric flask, and dilute to the mark with distilled water. Mix thoroughly, and pipet 50.00-mL aliquots into 250-mL conical flasks. Add about 50 mL of water and 2 drops of phenolphthalein (note 2) to each, and titrate with standard 0.1 M NaOH to the first permanent (≈30 s) pink color. Report the acidity of the vinegar as percent (w/v) CH₃COOH (60.053 g/mol).
2. If the unknown is a wine, pipet 50.00-mL aliquots into 250-mL conical flasks, add about 50 mL of distilled water and 2 drops of phenolphthalein to each (note 2), and titrate to the first permanent (≈30 s) pink color. Express the acidity of the sample as percent (weight/volume) tartaric acid C₂H₄O₂ (COOH)₂ (150.09 g/mol) (note 3).
Notes
1. The acidity of bottled vinegar tends to decrease on exposure to air. It is recom-
mended that unknowns be stored in individual vials with snug covers.
2. The amount of indicator used should be increased as necessary to make the color change visible in colored samples.
3. Tartaric acid has two acidic hydrogens, both of which are titrated at a phenolph-
thalein end point.

22C-10 The Determination of Sodium Carbonate in an Impure Sample

Discussion
The titration of sodium carbonate is discussed in Section 9A-2 in connection with its use as a primary standard; the same considerations apply for the determination of carbonate in an unknown that has no interfering contaminants.

Procedure
Dry the unknown at 110°C for 2 hr, and then cool in a desiccator. Consult with the instructor on an appropriate sample size. Then follow the instructions in Section 22C-6. Report the percentage of Na₂CO₃ in the sample.

22C-11 The Determination of Amine Nitrogen by the Kjeldahl Method

Discussion
These directions are suitable for the Kjeldahl determination of protein in materials such as blood meal, wheat flour, pasta products, dry cereals, and pet foods. A simple modification permits the analysis of unknowns that contain more highly oxidized forms of nitrogen.⁵

In the Kjeldahl method (see Section 9B-1), the organic sample is digested in hot concentrated sulfuric acid, which converts amine nitrogen in the sample to ammonium sulfate. After cooling, the sulfuric acid is neutralized by the addition of an excess of concentrated sodium hydroxide. The ammonia liberated by this treatment is then distilled into a measured excess of a standard solution of acid; the excess is determined by back-titration with standard base.

Figure 22-1 illustrates typical equipment for a Kjeldahl distillation. The longnecked container, which is used for both digestion and distillation, is called a Kjeldahl flask. In the apparatus in Figure 22-1a, the base is added slowly by

partially opening the stopcock from the NaOH storage vessel, the liberated ammonia is then carried to the receiving flask by steam distillation.

In an alternative method (Figure 22-1b), a dense, concentrated sodium hydroxide solution is carefully poured down the side of the Kjeldahl flask to form a second, lower layer. The flask is then quickly connected to a spray trap and an ordinary condenser before loss of ammonia can occur. Only then are the two layers mixed by gentle swirling of the flask.

Quantitative collection of ammonia requires the tip of the condenser to extend into the liquid in the receiving flask throughout the distillation step. The tip must be removed before heating is discontinued, however. Otherwise, the liquid will be drawn back into the apparatus.

Two methods are commonly used for collecting and determining the ammonia liberated from the sample. In one, the ammonia is distilled into a measured volume of standard acid. After the distillation is complete, the excess acid is back-titrated with standard base. An indicator with an acidic transition range is required because of the acidity of the ammonium ions present at equivalence. A convenient alternative, which requires only one standard solution, involves the collection of the ammonia in an unmeasured excess of boric acid, which retains the ammonia by the reaction

\[
\text{H}_3\text{BO}_3 + \text{NH}_3 \rightarrow \text{NH}_4^+ + \text{H}_2\text{BO}_3^- 
\]

The dihydrogen borate ion produced is a reasonably strong base that can be titrated with a standard solution of hydrochloric acid:

\[
\text{H}_2\text{BO}_3^- + \text{H}_3\text{O}^+ \rightarrow \text{H}_3\text{BO}_3 + \text{H}_2\text{O} 
\]

At the equivalence point, the solution contains boric acid and ammonium ions; an indicator with an acidic transition interval (such as bromocresol green) is again required.
22C Neutralization Titrations

Procedure

Preparing Samples
Consult with the instructor on sample size. *If the unknown is powdered* (such as blood meal), weigh samples onto individual 9-cm filter papers (note 1). Fold the paper around the sample, and drop each into a Kjeldahl flask (the paper keeps the samples from clinging to the neck of the flask). *If the unknown is not powdered* (such as breakfast cereals or pasta), the samples can be weighed by difference directly into the Kjeldahl flasks.

Add 25 mL of concentrated H$_2$SO$_4$, 10 g of powdered K$_2$SO$_4$, and the catalyst (note 2) to each flask.

Digestion
Clamp the flasks in a slanted position in a hood or vented digestion rack. Heat carefully to boiling. Discontinue heating briefly if foaming becomes excessive; never allow the foam to reach the neck of the flask. Once foaming ceases and the acid is boiling vigorously, the samples can be left unattended; prepare the distillation apparatus during this time. Continue digestion until the solution becomes colorless or faint yellow; 2 to 3 hr may be needed for some materials. If necessary, cautiously replace the acid lost by evaporation.

When digestion is complete, discontinue heating and allow the flasks to cool to room temperature; swirl the flasks if the contents show signs of solidifying. Cautiously add 250 mL of water to each flask and again allow the solution to cool to room temperature.

Distillation of Ammonia
Arrange a distillation apparatus similar to that shown in Figure 22-1. Pipet 50.00 mL of standard 0.1 M HCl into the receiver flask (note 3). Clamp the flask so that the tip of the adapter extends below the surface of the standard acid. Circulate water through the condenser jacket.

Hold the Kjeldahl flask at an angle, and gently introduce about 60 mL of 50% (w/v) NaOH solution, taking care to minimize mixing with the solution in the flask. *The concentrated caustic solution is highly corrosive and should be handled with great care* (note 4). Add several pieces of granulated zinc (note 5) and a small piece of litmus paper. *Immediately* connect the Kjeldahl flask to the spray trap. Cautiously mix the contents by gentle swirling. The litmus paper should be blue after mixing is complete, indicating that the solution is basic.

Bring the solution to a boil, and distill at a steady rate until one-half to one-third of the original volume remains. Control the rate of heating to prevent the liquid in the receiver flask from being drawn back into the Kjeldahl flask. After distillation is judged complete, lower the receiver flask to bring the adapter well clear of the liquid. Discontinue heating, disconnect the apparatus, and rinse the inside of the condenser with small portions of distilled water, collecting the washings in the receiver flask. Add 2 drops of bromocresol green to the receiver flask, and titrate the residual HCl with standard 0.1 M NaOH to the color change of the indicator.

Report the percentage of nitrogen and the percentage of protein (note 6) in the unknown.

Notes
1. If filter paper is used to hold the sample, carry a similar piece through the analysis as a blank. Acid-washed filter paper is frequently contaminated with measurable amounts of ammonium ion and should be avoided if possible.
2. Any of the following catalyze the digestion: a crystal of CuSO$_4$, 0.1 g of selenium, or 0.2 g of CuSeO$_3$. The catalyst can be omitted, if desired.

3. A modification of this procedure uses about 50 mL of 4% boric acid solution in lieu of the standard HCl in the receiver flask. After distillation is complete, the ammonium borate produced is titrated with standard 0.1 M HCl, with 2 to 3 drops of bromocresol green as indicator.

4. If any sodium hydroxide solution comes into contact with your skin, wash the affected area immediately with copious amounts of water.

5. Granulated zinc (10 to 20 mesh) is added to minimize bumping during the distillation; it reacts slowly with the base to give small bubbles of hydrogen that prevent superheating of the liquid.

6. The percentage of protein in the unknown is calculated by multiplying the % N by an appropriate factor: 5.70 for cereals, 6.25 for meats, and 6.38 for dairy products.

### 22D PRECIPITATION TITRATIONS

As noted in Section 10B-2, most precipitation titrations make use of a standard silver nitrate solution as titrant. Directions follow for the volumetric titration of chloride ion using an adsorption indicator.

### 22D-1 Preparing a Standard Silver Nitrate Solution

**Procedure**

Use a top-loading balance to transfer the approximate mass of AgNO$_3$ to a weighing bottle (note 1). Dry at 110°C for about 1 hr but not much longer (note 2), and then cool to room temperature in a desiccator. Weigh the bottle and contents (to the nearest 0.1 mg). Transfer the bulk of the AgNO$_3$ to a volumetric flask using a powder funnel. Cap the weighing bottle and reweigh it and any solid that remains. Rinse the powder funnel thoroughly. Dissolve the AgNO$_3$, dilute to the mark with water, and mix well (note 3). Calculate the molar concentration of this solution.

**Notes**

1. Consult with the instructor concerning the volume and concentration of AgNO$_3$ to be prepared. The mass of AgNO$_3$ to be taken is as follows:

<table>
<thead>
<tr>
<th>Silver Ion Concentration, M</th>
<th>Approximate Mass (g) of AgNO$_3$ Needed to Prepare</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>1000 mL</td>
</tr>
<tr>
<td>0.10</td>
<td>16.9</td>
</tr>
<tr>
<td>0.05</td>
<td>8.5</td>
</tr>
<tr>
<td>0.02</td>
<td>3.4</td>
</tr>
</tbody>
</table>

2. Prolonged heating causes partial decomposition of AgNO$_3$. Some discoloration may occur, even after only 1 hr at 110°C; the effect of this decomposition on the purity of the reagent is ordinarily imperceptible.

3. Silver nitrate solutions should be stored in a dark place when not in use.
22D-2 The Determination of Chloride by Titration with an Adsorption Indicator

Discussion
In this titration, the anionic adsorption indicator dichlorofluorescein is used to locate the end point. With the first excess of titrant, the indicator becomes incorporated in the counter-ion layer surrounding the silver chloride and imparts color to the solid (page 287). To obtain a satisfactory color change, it is desirable to maintain the particles of silver chloride in the colloidal state. Dextrin is added to the solution to stabilize the colloid and prevent its coagulation.

Preparation of Solutions

Dichlorofluorescein indicator (sufficient for several hundred titrations). Dissolve 0.2 g of dichlorofluorescein in a solution prepared by mixing 75 mL of ethanol and 25 mL of water.

Procedure

Dry the unknown at 110°C for about 1 hr; allow it to return to room temperature in a desiccator. Weigh individual samples (to the nearest 0.1 mg) into individual conical flasks, and dissolve them in appropriate volumes of distilled water (note 1). To each, add about 0.1 g of dextrin and 5 drops of indicator. Titrate (note 2) with AgNO₃ to the first permanent pink color of silver dichlorofluoresceinate. Report the percentage of Cl⁻ in the unknown.

Notes
1. Use 0.25-g samples for 0.1 M AgNO₃ and about half that amount for 0.05 M reagent. Dissolve the former in about 200 mL of distilled water and the latter in about 100 mL. If 0.02 M AgNO₃ is to be used, weigh a 0.4-g sample into a 500-mL volumetric flask, and take 50-mL aliquots for titration.
2. Colloidal AgCl is sensitive to photodecomposition, particularly in the presence of the indicator; attempts to perform the titration in direct sunlight will fail. If photodecomposition appears to be a problem, establish the approximate end point with a rough preliminary titration, and use this information to estimate the volumes of AgNO₃ needed for the other samples. For each subsequent sample, add the indicator and dextrin only after most of the AgNO₃ has been added, and then complete the titration without delay.

22E COMPLEX-FORMATION TITRATIONS WITH EDTA

See Section 10D for a discussion of the analytical uses of EDTA as a chelating reagent. Directions follow for a direct titration of magnesium and a determination of the hardness of a natural water.
22E-1  Preparation of Solutions

**Procedure**

A pH-10 buffer and an indicator solution are needed for these titrations.

1. *Buffer solution, pH 10* (sufficient for 80 to 100 titrations). Dilute 57 mL of concentrated NH₃ and 7 g of NH₄Cl in sufficient distilled water to give 100 mL of solution.

2. *Eriochrome Black T (Erio T) indicator* (sufficient for about 100 titrations). Dissolve 100 mg of the solid in a solution containing 15 mL of ethanolamine and 5 mL of absolute ethanol. This solution should be freshly prepared every two weeks; refrigeration slows its deterioration.

22E-2  Preparing a Standard 0.01 M EDTA Solution

**Discussion**

See Section 10D-1 for a description of the properties of reagent-grade Na₂H₂Y·2H₂O and its use in the direct preparation of standard EDTA solutions.

**Procedure**

Dry about 4 g of the purified dihydrate Na₂H₂Y·2H₂O (note 1) at 80°C to remove superficial moisture. Cool to room temperature in a desiccator. Weigh (to the nearest milligram) about 3.8 g into a 1-L volumetric flask (note 2). Use a powder funnel to ensure quantitative transfer; rinse the funnel well with water before removing it from the flask. Add 600 to 800 mL of water (note 3) and swirl periodically. Dissolution may take 15 min or longer. When all the solid has dissolved, dilute to the mark with water and mix well (note 4). In calculating the molarity of the solution, correct the weight of the salt for the 0.3% moisture it ordinarily retains after drying at 80°C.

**Notes**


2. The solution can be prepared from the anhydrous disodium salt, if desired. The weight taken should be about 3.6 g.

3. Water used in the preparation of standard EDTA solutions must be totally free of polyvalent cations. If any doubt exists concerning its quality, pass the water through a cation-exchange resin before use.

4. As an alternative, an EDTA solution that is approximately 0.01 M can be prepared and standardized by direct titration against a Mg²⁺ solution of known concentration (using the directions in Section 22E-3).

22E-3  The Determination of Magnesium by Direct Titration

**Discussion**

See Section 10D-7.
22F  Titrations with Potassium Permanganate

Procedure

Submit a clean 500-mL volumetric flask to receive the unknown, dilute to the mark with water, and mix thoroughly. Transfer 50.00-mL aliquots to 250-mL conical flasks, add 1 to 2 mL of pH 10 buffer and 3 to 4 drops of Erio T indicator to each. Titrate with 0.01 M EDTA until the color changes from red to pure blue (notes 1, 2).

Express the results as parts per million of Mg$^{2+}$ in the sample.

Notes

1. The color change tends to be slow in the vicinity of the end point. Care must be taken to avoid overtitration.
2. Other alkaline earths, if present, are titrated along with the Mg$^{2+}$; removal of Ca$^{2+}$ and Ba$^{2+}$ can be accomplished with (NH$_4$)$_2$CO$_3$. Most polyvalent cations are also titrated. Precipitation as hydroxides or the use of a masking reagent may be needed to eliminate this source of interference.

22E-4  The Determination of Hardness in Water

Discussion

See Section 10D-9.

Procedure

Acidify 100.0-mL aliquots of the sample with a few drops of HCl, and boil gently for a few minutes to eliminate CO$_2$. Cool, add 3 to 4 drops of methyl red, and neutralize with 0.1 M NaOH. Introduce 2 mL of pH 10 buffer, 3 to 4 drops of Erio T, and titrate with standard 0.01 M Na$_2$H$_2$Y to a color change from red to pure blue (note).

Report the results in terms of milligrams of CaCO$_3$ per liter of water.

Note

The color change is sluggish if Mg$^{2+}$ is absent. In this event, add 1 to 2 mL of 0.1 M MgY$^{2-}$ before starting the titration. This reagent is prepared by adding 2.645 g of MgSO$_4$·7H$_2$O to 3.722 g of Na$_2$H$_2$Y·2H$_2$O in 50 mL of distilled water. This solution is rendered faintly alkaline to phenolphthalein and diluted to 100 mL. A small portion, mixed with pH 10 buffer and a few drops of Erio T indicator, should have a dull violet color. A single drop of 0.01 M EDTA solution should cause a color change to blue, whereas an equal volume of 0.01 Mg$^{2+}$ should cause a change to red. If necessary, adjust the composition with EDTA or with Mg$^{2+}$ until these criteria are met.

22F  TITRATIONS WITH POTASSIUM PERMANGANATE

The properties and uses of potassium permanganate are described in Section 12H-1. Directions follow for the determination of iron in an ore and calcium in a limestone.
22F-1 Preparation of 0.02 M Potassium Permanganate

Discussion
See page 377 for a discussion of the precautions needed in the preparation and storage of permanganate solutions.

Procedure

Dissolve about 3.2 g of KMnO₄ in 1 L of distilled water. Keep the solution at a gentle boil for about 1 hr. Cover and let stand overnight. Remove MnO₂ by filtration (note 1) through a fine-porosity filtering crucible (note 2) or through a Gooch crucible fitted with glass mats. Transfer the solution to a clean glass-stoppered bottle; store in the dark when not in use.

Notes
1. The heating and filtering can be omitted if the permanganate solution is standardized and used on the same day.
2. Remove the MnO₂ that collects on the fritted plate with 1 M H₂SO₄ containing a few milliliters of 3% H₂O₂, followed by a rinse with copious quantities of water.

22F-2 Standardization of Potassium Permanganate Solutions

Discussion
See Section 12H-1 for a discussion of primary standards for permanganate solutions. Directions follow for standardization with sodium oxalate.

Procedure

Dry about 1.5 g of primary-standard Na₂C₂O₄ at 110°C for at least 1 hr. Cool in a desiccator; weigh (to the nearest 0.1 mg) individual 0.2- to 0.3-g samples into 400-mL beakers. Dissolve each in about 250 mL of 1 M H₂SO₄. Heat each solution to 80 to 90°C, and titrate with KMnO₄ while stirring with a thermometer. The pink color imparted by one addition should be permitted to disappear before any further titrant is introduced (notes 1, 2). Reheat if the temperature drops below 60°C. Take the first persistent (≈30 s) pink color as the end point (notes 3, 4). Determine a blank by titrating an equal volume of the 1 M H₂SO₄.

Correct the titration data for the blank, and calculate the concentration of the permanganate solution (note 5).

Notes
1. Promptly wash any KMnO₄ that spatters on the walls of the beaker into the bulk of the liquid with a stream of water.
2. Finely divided MnO₂ will form along with Mn²⁺ if the KMnO₄ is added too rapidly and will cause the solution to acquire a faint brown discoloration. Precipitate formation is not a serious problem so long as sufficient oxalate remains to reduce the MnO₂ to Mn²⁺; the titration is simply discontinued until the brown color disappears. The solution must be free of MnO₂ at the end point.
3. The surface of the permanganate solution rather than the bottom of the meniscus can be used to measure titrant volumes. Alternatively, backlighting with a flashlight or a match will permit reading of the meniscus in the conventional manner.

4. A permanganate solution should not be allowed to stand in a buret any longer than necessary because partial decomposition to MnO₂ may occur. Freshly formed MnO₂ can be removed from a glass surface with 1 M H₂SO₄ containing a small amount of 3% H₂O₂.

5. As noted on page 379, this procedure yields molarities that are a few tenths of a percent low. For more accurate results, introduce from a buret sufficient permanganate to react with 90 to 95% of the oxalate (about 40 mL of 0.02 M KMnO₄ for a 0.3-g sample). Let the solution stand until the permanganate color disappears. Then warm to about 60°C and complete the titration, taking the first permanent pink (≈30 s) as the end point. Determine a blank by titrating an equal volume of the 1 M H₂SO₄.

22F-3 Determining Calcium in a Limestone

Discussion

In common with a number of other cations, calcium is conveniently determined by precipitation with oxalate ion. The solid calcium oxalate is filtered, washed free of excess precipitating reagent, and dissolved in dilute acid. The oxalic acid liberated in this step is then titrated with standard permanganate or some other oxidizing reagent. This method is applicable to samples that contain magnesium and the alkali metals. Most other cations must be absent since they either precipitate or coprecipitate as oxalates and cause positive errors in the analysis.

Factors Affecting the Composition of Calcium Oxalate Precipitates. It is essential that the mole ratio between calcium and oxalate be exactly unity in the precipitate and thus in solution at the time of titration. A number of precautions are needed to ensure this condition. For example, the calcium oxalate formed in a neutral or ammoniacal solution is likely to be contaminated with calcium hydroxide or with a basic calcium oxalate, either of which will cause low results. The formation of these compounds is prevented by adding the oxalate to an acidic solution of the sample and slowly forming the desired precipitate by the dropwise addition of ammonia. The coarsely crystalline calcium oxalate that is produced under these conditions is readily filtered. Losses resulting from the solubility of calcium oxalate are negligible above pH 4, provided washing is limited to freeing the precipitate of excess oxalate.

Coprecipitation of sodium oxalate becomes a source of positive error in the determination of calcium whenever the concentration of sodium in the sample exceeds that of calcium. The error from this source can be eliminated by reprecipitation.

Magnesium, if present in high concentration, may also be a source of contamination. An excess of oxalate ion helps prevent this interference through the formation of soluble oxalate complexes of magnesium. Prompt filtration of the calcium oxalate can also help prevent interference because of the pronounced tendency of magnesium oxalate to form supersaturated solutions from which precipitate formation occurs only after an hour or more. These measures do not suf-
office for samples that contain more magnesium than calcium. Here, reprecipitation of the calcium oxalate becomes necessary.

**The Composition of Limestones.** Limestones are composed principally of calcium carbonate; dolomitic limestones contain large amounts of magnesium carbonate as well. Calcium and magnesium silicates, along with the carbonates and silicates of iron, aluminum, manganese, titanium, sodium, and other metals, are also present in smaller amounts.

Hydrochloric acid is an effective solvent for most limestones. Only silica, which does not interfere with the analysis, remains undissolved. Some limestones are more readily decomposed after they have been ignited; a few yield only to a carbonate fusion.

The method that follows is remarkably effective for determining calcium in most limestones. Iron and aluminum, in amounts equivalent to that of calcium, do not interfere. Small amounts of manganese and titanium can also be tolerated.

### Procedure

**Sample Preparation**

Dry the unknown for 1 to 2 hr at 110°C, and cool in a desiccator. If the material is readily decomposed in acid, weigh 0.25- to 0.30-g samples (to the nearest 0.1 mg) into 250-mL beakers. Add 10 mL of water to each sample, and cover with a watch glass. Add 10 mL of concentrated HCl dropwise, taking care to avoid losses due to spattering as the acid is introduced.

**Precipitation of Calcium Oxalate**

Add 5 drops of saturated bromine water to oxidize any iron in the samples, and boil gently (HOOD) for 5 min to remove the excess Br₂. Dilute each sample solution to about 50 mL, heat to boiling, and add 100 mL of hot 6% (w/v) \((\text{NH}_4)_2\text{C}_2\text{O}_4\) solution. Add 3 to 4 drops of methyl red, and precipitate \(\text{CaC}_2\text{O}_4\) by slowly adding 6 M \(\text{NH}_3\). As the indicator starts to change color, add the \(\text{NH}_3\) at a rate of one drop every 3 to 4 s. Continue until the solutions turn to the intermediate yellow-orange color of the indicator (pH 4.5 to 5.5). Allow the solutions to stand for no more than 30 min (note) and filter; medium-porosity filtering crucibles or Gooch crucibles with glass mats are satisfactory. Wash the precipitates with several 10-mL portions of cold water. Rinse the outside of the crucibles to remove residual \((\text{NH}_4)_2\text{C}_2\text{O}_4\), and return them to the beakers in which the \(\text{CaC}_2\text{O}_4\) was formed.

**Titration**

Add 100 mL of water and 50 mL of 3 M \(\text{H}_2\text{SO}_4\) to each of the beakers containing the precipitated calcium oxalate and the crucible. Heat to 80 to 90°C, and titrate with 0.02 M permanganate. The temperature should be above 60°C throughout the titration; re-heat if necessary.

Report the percentage of CaO in the unknown.

**Note**

The period of standing can be longer if the unknown contains no Mg²⁺.
22F-4 The Determination of Iron in an Ore

Discussion

The common ores of iron are hematite \((\text{Fe}_2\text{O}_3)\), magnetite \((\text{Fe}_3\text{O}_4)\), and limonite \((2\text{Fe}_2\text{O}_3 \cdot 3\text{H}_2\text{O})\). Steps in the analysis of these ores are (1) dissolution of the sample, (2) reduction of iron to the divalent state, and (3) titration of iron(II) with a standard oxidant.

The Decomposition of Iron Ores. Iron ores often decompose completely in hot concentrated hydrochloric acid. The rate of attack by this reagent is increased by the presence of a small amount of tin(II) chloride. The tendency of iron(II) and iron(III) to form chloro complexes accounts for the effectiveness of hydrochloric acid over nitric or sulfuric acid as a solvent for iron ores.

Many iron ores contain silicates that may not be entirely decomposed by treatment with hydrochloric acid. Incomplete decomposition is indicated by a dark residue that remains after prolonged treatment with the acid. A white residue of hydrated silica, which does not interfere in any way, is indicative of complete decomposition.

The Prereduction of Iron. Because part of or all the iron is in the trivalent state after decomposition of the sample, prereduction to iron(II) must precede titration with the oxidant. Any of the methods described in Section 12F-1 can be used. Perhaps the most satisfactory prereductant for iron is tin(II) chloride:

\[
2\text{Fe}^{3+} + \text{Sn}^{2+} \rightarrow 2\text{Fe}^{2+} + \text{Sn}^{4+}
\]

The only other common species reduced by this reagent are the high oxidation states of arsenic, copper, mercury, molybdenum, tungsten, and vanadium.

The excess reducing agent is eliminated by the addition of mercury(II) chloride:

\[
\text{Sn}^{2+} + 2\text{HgCl}_2 \rightarrow \text{Hg}_2\text{Cl}_2(s) + \text{Sn}^{4+} + 2\text{Cl}^{-}
\]

The slightly soluble mercury(I) chloride \((\text{Hg}_2\text{Cl}_2)\) does not reduce permanganate, nor does the excess mercury(II) chloride \((\text{HgCl}_2)\) reoxidize iron(II). Care must be taken, however, to prevent the occurrence of the alternative reaction

\[
\text{Sn}^{2+} + \text{HgCl}_2 \rightarrow \text{Hg}(l) + \text{Sn}^{4+} + 2\text{Cl}^{-}
\]

Elemental mercury reacts with permanganate and causes the results of the analysis to be high. The formation of mercury, which is favored by an appreciable excess of tin(II), is prevented by careful control of this excess and by the rapid addition of excess mercury(II) chloride. A proper reduction is indicated by the appearance of a small amount of a silky white precipitate after the addition of mercury(II). Formation of a gray precipitate at this juncture indicates the presence of metallic mercury; the total absence of a precipitate indicates that an insufficient amount of tin(II) chloride was used. In either event, the sample must be discarded.

The Titration of Iron(II). The reaction of iron(II) with permanganate is smooth and rapid. The presence of iron(II) in the reaction mixture, however,
induces the oxidation of chloride ion by permanganate, a reaction that does not ordinarily proceed rapidly enough to cause serious error. High results are obtained if this parasitic reaction is not controlled. Its effects can be eliminated through removal of the hydrochloric acid by evaporation with sulfuric acid or by introduction of Zimmermann–Reinhardt reagent, which contains manganese(II) in a fairly concentrated mixture of sulfuric and phosphoric acids.

The oxidation of chloride ion during a titration is believed to involve a direct reaction between this species and the manganese(III) ions that form as an intermediate in the reduction of permanganate ion by iron(II). The presence of manganese(II) in the Zimmermann–Reinhardt reagent is believed to inhibit the formation of chlorine by decreasing the potential of the manganese(III)/manganese(II) couple. Phosphate ion is believed to exert a similar effect by forming stable manganese(III) complexes. Moreover, phosphate ions react with iron(III) to form nearly colorless complexes so that the yellow color of the iron(II)/chloro complexes does not interfere with the end point.6

### Preparation of Reagents

The following solutions suffice for about 100 titrations.

1. Tin(II) chloride, 0.25 M. Dissolve 60 g of iron-free SnCl₂·2H₂O in 100 mL of concentrated HCl; warm if necessary. After the solid has dissolved, dilute to 1 L with distilled water and store in a well-stoppered bottle. Add a few pieces of mossy tin to help preserve the solution.

2. Mercury(II) chloride, 5% (w/v). Dissolve 50 g of HgCl₂ in 1 L of distilled water.

3. Zimmermann–Reinhardt reagent. Dissolve 300 g of MnSO₄·4H₂O in 1 L of water. Cautiously add 400 mL of concentrated H₂SO₄, add 400 mL of 85% H₃PO₄, and dilute to 3 L.

### Procedure

**Sample Preparation**

Dry the ore at 110°C for at least 3 hr, and then allow it to cool to room temperature in a desiccator. Consult with the instructor for a sample size that will require from 25 to 40 mL of standard 0.02 M KMnO₄. Weigh samples into 500-mL conical flasks. To each, add 10 mL of concentrated HCl and about 3 mL of 0.25 M SnCl₂ (note 1). Cover each flask with a small watch glass or Tuttle flask cover. Heat the flasks in a hood at just below boiling until the samples are decomposed and the undissolved solid, if any, is pure white (note 2). Use another 1 to 2 mL of SnCl₂ to eliminate any yellow color that may develop as the solutions are heated. Heat a blank consisting of 10 mL of HCl and 3 mL of SnCl₂ for the same amount of time.

After the ore has been decomposed, remove the excess Sn(II) by the dropwise addition of 0.02 M KMnO₄ until the solutions become faintly yellow. Dilute to about 15 mL. Add sufficient KMnO₄ solution to impart a faint pink color to the blank, and then decolorize with one drop of the SnCl₂ solution.

---

6The mechanism by which Zimmermann–Reinhardt reagent acts has been the subject of much study. For a discussion of this work, see H. A. Laitinen, Chemical Analysis (New York: McGraw-Hill, 1960), pp. 369–372.
Take samples and blank individually through subsequent steps to minimize air oxidation of iron(II).

Reduction of Iron
Heat the sample solution nearly to boiling, and make dropwise additions of 0.25 M SnCl₂ until the yellow color just disappears; then add two more drops (note 3). Cool to room temperature, and rapidly add 10 mL of 5% HgCl₂ solution. A small amount of silky white Hg₂Cl₂ should precipitate (note 4). The blank should be treated with the HgCl₂ solution.

Titration
Following addition of the HgCl₂, wait 2 to 3 min. Then add 25 mL of Zimmermann–Reinhardt reagent and 300 mL of water. Titrate immediately with standard 0.02 M KMnO₄ to the first faint pink that persists for 15 to 20 s. Do not add the KMnO₄ rapidly at any time. Correct the titrant volume for the blank. Report the percentage of Fe₂O₃ in the sample.

Notes
1. The SnCl₂ hastens decomposition of the ore by reducing iron(III) oxides to iron(II). Insufficient SnCl₂ is indicated by the appearance of yellow iron(III)/chloride complexes.
2. If dark particles persist after the sample has been heated with acid for several hours, filter the solution through ashless paper, wash the residue with 5 to 10 mL of 6 M HCl, and retain the filtrate and washings. Ignite the paper and its contents in a small platinum crucible. Mix 0.5 to 0.7 g of Na₂CO₃ with the residue, and heat until a clear melt is obtained. Cool, add 5 mL of water, and then cautiously add a few milliliters of 6 M HCl. Warm the crucible until the melt has dissolved, and combine the contents with the original filtrate. Evaporate the solution to 15 mL and continue the analysis.
3. The solution may not become entirely colorless but instead may acquire a faint yellow-green hue. Further additions of SnCl₂ will not alter this color. If too much SnCl₂ is added, it can be removed by adding 0.2 M KMnO₄ and repeating the reduction.
4. The absence of precipitate indicates that insufficient SnCl₂ was used and that the reduction of iron(III) was incomplete. A gray residue indicates the presence of elemental mercury, which reacts with KMnO₄. The sample must be discarded in either event.
5. These directions can be used to standardize a permanganate solution against primary standard iron. Weigh (to the nearest 0.1 mg) 0.2-g lengths of electrolytic iron wire into 250-mL conical flasks, and dissolve in about 10 mL of concentrated HCl. Dilute each sample to about 75 mL. Then take each individually through the reduction and titration steps.

22G TITRATIONS WITH IODINE
The oxidizing properties of iodine, the composition and stability of triiodide solutions, and the applications of this reagent in volumetric analysis are discussed in Section 12H-3. Starch is ordinarily employed as an indicator for iodimetric titrations.
22G-1 Preparation of Reagents

Procedure

(a) Iodine approximately 0.05 M. Weigh about 40 g of KI into a 100-mL beaker. Add 12.7 g of I₂ and 10 mL of water. Stir for several minutes (note 1). Introduce an additional 20 mL of water, and stir again for several minutes. Carefully decant the bulk of the liquid into a storage bottle containing 1 L of distilled water. It is essential that any undissolved iodine remain in the beaker (note 2).

(b) Starch indicator (sufficient for about 100 titrations). Rub 1 g of soluble starch and 15 mL of water into a paste. Dilute to about 500 mL with boiling water, and heat until the mixture is clear. Cool; store in a tightly stoppered bottle. For most titrations, 3 to 5 mL of the indicator is used.

The indicator is readily attacked by airborne organisms and should be freshly prepared every few days.

Notes

1. Iodine dissolves slowly in the KI solution. Thorough stirring is needed to hasten the process.
2. Any solid I₂ inadvertently transferred to the storage bottle will cause the concentration of the solution to increase gradually. Filtration through a sintered-glass crucible eliminates this potential source of difficulty.

22G-2 Standardization of Iodine Solutions

Discussion

Arsenic(III) oxide, long a favored primary standard for iodine solutions, is now seldom used because of the elaborate federal regulations governing the use of even small amounts of arsenic-containing compounds. Barium thiosulfate monohydrate and anhydrous sodium thiosulfate have been proposed as alternative standards.7 Perhaps the most convenient method for determining the concentration of an iodine solution is the titration of aliquots with a sodium thiosulfate solution that has been standardized against pure potassium iodate. Instructions for this method follow:

Preparation of Reagents

1. Sodium thiosulfate, 0.1 M. Follow the directions in Sections 22H-1 and 22H-2 for the preparation and standardization of this solution.
2. Starch indicator. See Section 22G-1(b).

### Procedure

Transfer 25.00-mL aliquots of the iodine solution to 250-mL conical flasks, and dilute to about 50 mL. Take each aliquot individually through subsequent steps. Introduce approximately 1 mL of 3 M H$_2$SO$_4$, and titrate immediately with standard sodium thiosulfate until the solution becomes a faint straw yellow. Add about 5 mL of starch indicator, and complete the titration, taking as the end point the change in color from blue to colorless (note).

**Note**

The blue color of the starch/iodine complex may reappear after the titration has been completed, owing to the air oxidation of iodide ion.

### 22G-3 The Determination of Antimony in Stibnite

#### Discussion

The analysis of stibnite, a common antimony ore, is a typical application of iodimetry and is based on the oxidation of Sb(III) to Sb(V):

\[
\text{SbO}_3^{2-} + \text{I}_2 + \text{H}_2\text{O} \rightleftharpoons \text{SbO}_4^{2-} + 2\text{I}^- + 2\text{H}^+
\]

The position of this equilibrium is strongly dependent on the hydrogen ion concentration. To force the reaction to the right, it is common practice to carry out the titration in the presence of an excess of sodium hydrogen carbonate, which consumes the hydrogen ions as they are produced.

Stibnite is an antimony sulfide ore containing silica and other contaminants. Provided that the material is free of iron and arsenic, the analysis of stibnite for its antimony content is straightforward. Samples are decomposed in hot concentrated hydrochloric acid to eliminate sulfide as gaseous hydrogen sulfide. Care is needed to prevent loss of volatile antimony(III) chloride during this step. The addition of potassium chloride helps by favoring formation of nonvolatile chloro complexes such as SbCl$_4^{-}$ and SbCl$_3^2-$.

Sparingly soluble basic antimony salts, such as SbOCl, often form when the excess hydrochloric acid is neutralized; these react incompletely with iodine and cause low results. The difficulty is overcome by adding tartaric acid, which forms a soluble complex (SbOC$_3$H$_4$O$_5$) from which antimony is rapidly oxidized by the reagent.

#### Procedure

Dry the unknown at 110°C for 1 hr, and allow it to cool in a desiccator. Weigh individual samples (note 1) into 500-mL conical flasks. Introduce about 0.3 g of KCl and 10 mL of concentrated HCl to each flask. Heat the mixtures (HOOD) just below boiling until only white or slightly gray residues of SiO$_2$ remain.

Add 3 g of tartaric acid to each sample, and heat for an additional 10 to 15 min. Then, with good swirling, add water (note 2) from a pipet or buret until the volume is about 100 mL. If reddish Sb$_2$S$_3$ forms, discontinue dilution and heat further to eliminate H$_2$S; add more HCl if necessary.
Add 3 drops of phenolphthalein, and neutralize with 6 M NaOH to the first faint pink of the indicator. Discharge the color by the dropwise addition of 6 M HCl, and then add 1 mL in excess. Introduce 4 to 5 g of NaHCO₃, taking care to avoid losses of solution by spattering during the addition. Add 5 mL of starch indicator, rinse down the inside of the flask, and titrate with standard 0.05 M I₂ to the first blue color that persists for 30 s.

Report the percentage of Sb₂S₃ in the unknown.

**Notes**

1. Samples should contain between 1.5 and 2 mmol of antimony; consult with the instructor for an appropriate sample size. Weighings to the nearest milligram are adequate for samples larger than 1 g.
2. The slow addition of water, with efficient stirring, is essential to prevent the formation of SbOCl.

### 22H TITRATIONS WITH SODIUM THIOSULFATE

Numerous methods are based upon the reducing properties of iodide ion:

\[
2I^- \rightarrow I_2 + 2e^-
\]

Iodine, the reaction product, is ordinarily titrated with a standard sodium thiosulfate solution, with starch serving as the indicator:

\[
I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}
\]

A discussion of thiosulfate methods is found in Section 12H-2.

#### 22H-1 Preparation of 0.1 M Sodium Thiosulfate

**Procedure**

Boil about 1 L of distilled water for 10 to 15 min. Allow the water to cool to room temperature; then add about 25 g of Na₂S₂O₃·5H₂O and 0.1 g of Na₂CO₃. Stir until the solid has dissolved. Transfer the solution to a clean glass or plastic bottle, and store in a dark place.

#### 22H-2 Standardizing Sodium Thiosulfate against Potassium Iodate

**Discussion**

Solutions of sodium thiosulfate are conveniently standardized by titration of the iodine produced when an unmeasured excess of potassium iodide is added to a known volume of an acidified standard potassium iodate solution. The reaction is

\[
IO_3^- + 5I^- + 6H^+ \rightarrow 3I_2 + 3H_2O
\]
Note that each mole of iodate results in the production of three moles of iodine. The procedure that follows is based on this reaction.

### Preparation of Solutions

1. **Potassium iodate, 0.0100 M.** Dry about 1.2 g of primary-standard KIO₃ at 110°C for at least 1 hr, and cool in a desiccator. Weigh (to the nearest 0.1 mg) about 1.1 g into a 500-mL volumetric flask; use a powder funnel to ensure quantitative transfer of the solid. Rinse the funnel well, dissolve the KIO₃ in about 200 mL of distilled water, dilute to the mark, and mix thoroughly.
2. **Starch indicator.** See Section 22G-1(b).

### Procedure

Pipet 50.00-mL aliquots of standard iodate solution into 250-mL conical flasks. Treat each sample individually from this point to minimize error resulting from the air oxidation of iodide ion. Introduce 2 g of iodate-free KI, and swirl the flask to hasten solution. Add 2 mL of 6 M HCl, and immediately titrate with thiosulfate until the solution becomes pale yellow. Introduce 5 mL of starch indicator, and titrate with constant stirring to the disappearance of the blue color. Calculate the molarity of the thiosulfate solution.

### 22H-3 Standardization of Sodium Thiosulfate against Copper

#### Discussion

Thiosulfate solutions can also be standardized against pure copper wire or foil. This procedure is advantageous when the solution is to be used for the determination of copper because any systematic error in the method tends to be canceled.

Copper(II) is reduced quantitatively to copper(I) by iodide ion:

\[
2\text{Cu}^{2+} + 4\text{I}^- \rightarrow 2\text{CuI(s)} + \text{I}_2
\]

The importance of CuI formation in forcing this reaction to completion can be seen from the following standard electrode potentials:

- \(\text{Cu}^{2+} + \text{e}^- \rightleftharpoons \text{Cu}^+\) \(E^0 = 0.15\) V
- \(\text{I}_2 + 2\text{e}^- \rightleftharpoons 2\text{I}^-\) \(E^0 = 0.54\) V
- \(\text{Cu}^{2+} + \text{I}^- + \text{e}^- \rightleftharpoons \text{CuI(s)}\) \(E^0 = 0.86\) V

The first two potentials suggest that iodide should have no tendency to reduce copper(II); the formation of CuI, however, favors the reduction. The solution must contain at least 4% excess iodide to force the reaction to completion. Moreover, the pH must be less than 4 to prevent the formation of basic copper species that react slowly and incompletely with iodide ion. The acidity of the solution cannot be greater than about 0.3 M, however, because of the tendency of iodide ion to undergo air oxidation, a process catalyzed by copper salts. Nitrogen oxides also catalyze the air oxidation of iodide ion. A common source of these
oxides is the nitric acid ordinarily used to dissolve metallic copper and the other copper-containing solids. Urea is used to scavenge nitrogen oxides from solutions:

\[(\text{NH}_2\text{CO})_2\text{HNO}_2 + 2\text{HNO}_2 \rightarrow 2\text{N}_2(g) + \text{CO}_2(g) + 3\text{H}_2\text{O}\]

The titration of iodine by thiosulfate tends to yield slightly low results owing to the adsorption of small but measurable quantities of iodine upon solid CuI. The adsorbed iodine is released only slowly, even when thiosulfate is in excess; transient and premature end points result. This difficulty is largely overcome by the addition of thiocyanate ion. The sparingly soluble copper(I) thiocyanate replaces part of the copper iodide at the surface of the solid:

\[\text{CuI(s)} + \text{SCN}^- \rightarrow \text{CuSCN(s)} + \text{I}^{-}\]

Accompanying this reaction is the release of the adsorbed iodine, which thus becomes available for titration. The addition of thiocyanate must be delayed until most of the iodine has been titrated to prevent interference from a slow reaction between the two species, possibly

\[2\text{SCN}^- + I_2 \rightarrow 2I^- + (\text{SCN})_2\]

### Preparation of Solutions

1. **Urea, 5% (w/v).** Dissolve about 5 g of urea in sufficient water to give 100 mL of solution. Approximately 10 mL will be needed for each titration.
2. **Starch indicator.** See Section 22G-1(b).

### Procedure

Use scissors to cut copper wire or foil into 0.20- to 0.25-g portions. Wipe the metal free of dust and grease with a filter paper; do not dry it. The pieces of copper should be handled with paper strips, cotton gloves, or tweezers to prevent contamination by contact with the skin.

Use a weighed watch glass or weighing bottle to obtain the mass of individual copper samples (to the nearest 0.1 mg) by difference. Transfer each sample to a 250-mL conical flask. Add 5 mL of 6 M HNO₃, cover with a small watch glass, and warm gently (HOOD) until the metal has dissolved. Dilute with about 25 mL of distilled water, add 10 mL of 5% (w/v) urea, and boil briefly to eliminate nitrogen oxides. Rinse the watch glass, collecting the rinsings in the flask. Cool.

Add concentrated NH₃ dropwise and with thorough mixing to produce the intensely blue Cu(NH₃)₂⁺: the solution should smell faintly of ammonia (note). Make dropwise additions of 3 M H₂SO₄ until the color of the complex just disappears, and then add 2.0 mL of 85% H₃PO₄. Cool to room temperature.

_Treat each sample individually from this point on to minimize the air oxidation of iodide ion._ Add 4.0 g of KI to the sample, and titrate immediately with Na₂S₂O₃ until the solution becomes pale yellow. Add 5 mL of starch indicator, and continue the titration until the blue color becomes faint. Add 2 g of KSCN; swirl vigorously for 30 s. Complete the titration, using the disappearance of the blue starch/I₂ color as the end point.

Calculate the molarity of the Na₂S₂O₃ solution.
Note
Do not sniff vapors directly from the flask; instead, waft them toward your nose with a waving motion of your hand.

22H-4 The Determination of Copper in Brass

Discussion
The standardization procedure described in Section 22H-3 is readily adapted to the determination of copper in brass, an alloy that also contains appreciable amounts of tin, lead, and zinc (and perhaps minor amounts of nickel and iron). The method is relatively simple and applicable to brasses with less than 2% iron. A weighed sample is treated with nitric acid, which causes the tin to precipitate as a hydrated oxide of uncertain composition. Evaporation with sulfuric acid to the appearance of sulfur trioxide eliminates the excess nitrate, redissolves the tin compound, and possibly causes the formation of lead sulfate. The pH is adjusted through the addition of ammonia, followed by acidification with a measured amount of phosphoric acid. An excess of potassium iodide is added, and the liberated iodine is titrated with standard thiosulfate. See Section 22H-3 for additional discussion.

Procedure
If so directed, free the metal of oils by treatment with an organic solvent; briefly heat in an oven to drive off the solvent. Weigh (to the nearest 0.1 mg) 0.3-g samples into 250-mL conical flasks, and introduce 5 mL of 6 M HNO₃ into each; warm (HOOD) until solution is complete. Add 10 mL of concentrated H₂SO₄, and evaporate (HOOD) until copious white fumes of SO₃ are given off. Allow the mixture to cool. Cautiously add 30 mL of distilled water, boil for 1 to 2 min, and again cool.
Follow the instructions in the third and fourth paragraphs of the procedure in Section 22H-3.
Report the percentage of Cu in the sample.

22I TITRATIONS WITH POTASSIUM BROMATE

Applications of standard bromate solutions to the determination of organic functional groups are described in Section 12H-4. Directions follow for the determination of ascorbic acid in vitamin C tablets.

22I-1 Preparation of Solutions

Procedure
1. Potassium bromate, 0.015 M. Transfer about 1.5 g of reagent-grade potassium bromate to a weighing bottle, and dry at 110°C for at least 1 hr. Cool in a desiccator. Weigh approximately 1.3 g (to the nearest 0.1 mg) into a 500-mL volumetric flask; use a powder funnel to ensure quantitative transfer of the solid. Rinse the
funnel well, and dissolve the KBrO₃ in about 200 mL of distilled water. Dilute to the mark, and mix thoroughly.

*Solid potassium bromate can cause a fire if it comes into contact with damp organic material (such as paper towels in a waste container).* Consult with the instructor concerning the disposal of any excess.

2. Sodium thiosulfate, 0.05 M. Follow the directions in Section 22H-1; use about 12.5 g of Na₂S₂O₃·5H₂O per liter of solution.

3. Starch indicator. See Section 22G-1(b).

### 22I-2 Standardizing Sodium Thiosulfate against Potassium Bromate

**Discussion**

Iodine is generated by the reaction between a known volume of standard potassium bromate and an unmeasured excess of potassium iodide:

\[
\text{BrO}_3^- + 6I^- + 6H^+ \rightarrow \text{Br}^- + 3\text{I}_2 + 3\text{H}_2\text{O}
\]

The iodine produced is titrated with the sodium thiosulfate solution.

**Procedure**

Pipet 25.00-mL aliquots of the KBrO₃ solution into 250-mL conical flasks and rinse the interior wall with distilled water. *Treat each sample individually beyond this point.* Introduce 2 to 3 g of KI and about 5 mL of 3 M H₂SO₄. Immediately titrate with Na₂S₂O₃ until the solution is pale yellow. Add 5 mL of starch indicator, and titrate to the disappearance of the blue color.

Calculate the concentration of the thiosulfate solution.

### 22I-3 The Determination of Ascorbic Acid in Vitamin C Tablets by Titration with Potassium Bromate

**Discussion**

Ascorbic acid, C₆H₈O₆, is cleanly oxidized to dehydroascorbic acid by bromine:

\[
\text{C}_6\text{H}_8\text{O}_6 + \text{Br}_2 \rightarrow \text{C}_6\text{H}_7\text{O}_5\text{CHO} + 2\text{Br}^- + 2\text{H}^+
\]
An unmeasured excess of potassium bromide is added to an acidified solution of the sample. The solution is titrated with standard potassium bromate to the first permanent appearance of excess bromine; this excess is then determined iodometrically with standard sodium thiosulfate. The entire titration must be performed without delay to prevent air oxidation of the ascorbic acid.

**Procedure**

Weigh (to the nearest milligram) 3 to 5 vitamin C tablets (note 1). Pulverize them thoroughly in a mortar, and transfer the powder to a dry weighing bottle. Weigh individual 0.40- to 0.50-g samples (to the nearest 0.1 mg) into dry 250-mL conical flasks. Treat each sample individually beyond this point. Dissolve the sample (note 2) in 50 mL of 1.5 M H$_2$SO$_4$; then add about 5 g of KBr. Titrate immediately with standard KBrO$_3$ to the first faint yellow due to excess Br$_2$. Record the volume of KBrO$_3$ used. Add 3 g of KI and 5 mL of starch indicator; back-titrature (note 3) with standard 0.05 M Na$_2$S$_2$O$_3$.

Calculate the average mass (in milligrams) of ascorbic acid (176.12 g/mol) in each tablet.

**Notes**

1. This method is not applicable to chewable vitamin C tablets.
2. The binder in many vitamin C tablets remains in suspension throughout the analysis. If the binder is starch, the characteristic color of the complex with iodine appears upon the addition of KI.
3. The volume of thiosulfate needed for the back-titration seldom exceeds a few milliliters.

### 22J Potentiometric Methods

Potentiometric measurements provide a highly selective method for the quantitative determination of numerous cations and anions. A discussion of the principles and applications of potentiometric measurements is found in Chapter 13. Detailed instructions are given in this section on the use of potentiometric measurements to locate end points in volumetric titrations. In addition, a procedure for the direct potentiometric determination of fluoride ion in drinking water and in toothpaste is described.

#### 22J-1 General Directions for Performing a Potentiometric Titration

The procedure that follows is applicable to the titrimetric methods described in this section. With the proper choice of indicator electrode, it can also be applied to most of the volumetric methods given in Sections 22C through 22I.

**Procedure**

1. Dissolve the sample in 50 to 250 mL of water. Rinse a suitable pair of electrodes with deionized water, and immerse them in the sample solution. Provide magnetic
(or mechanical) stirring. Position the buret so that reagent can be delivered without splashing.

2. Connect the electrodes to the meter, commence stirring, and record the initial buret volume and the initial potential (or pH).

3. Record the meter reading and buret volume after each addition of titrant. Introduce fairly large volumes (about 5 mL) at the outset. Withhold a succeeding addition until the meter reading remains constant within 1 to 2 mV (or 0.05 pH unit) for at least 30 s (note). Judge the volume of reagent to be added by estimating a value for \( \Delta E/\Delta V \) after each addition. In the immediate vicinity of the equivalence point, introduce the reagent in 0.1-mL increments. Continue the titration 2 to 3 mL beyond the equivalence point, increasing the volume increments as \( \Delta E/\Delta V \) again becomes smaller.

Note

Stirring motors occasionally cause erratic meter readings; it may be advisable to turn off the motor while meter readings are being made.

**22J-2 The Potentiometric Titration of Chloride and Iodide in a Mixture**

**Discussion**

Figure 22-2 is a theoretical argentometric titration curve for a mixture of iodide and chloride ions. Initial additions of silver nitrate result in formation of silver iodide exclusively because the solubility of that salt is only about \( 5 \times 10^{-7} \) that of silver chloride. It can be shown that this solubility difference is great enough so that formation of silver chloride is delayed until all but \( 7 \times 10^{-5}\% \) of the iodide has precipitated. Thus, short of the equivalence point, the curve is essentially indistinguishable from that for iodide alone (Figure 22-2). Just beyond the iodide equivalence point, the silver ion concentration is determined by the concentration of chloride ion in the solution, and the titration curve becomes essentially identical to that for chloride ion by itself.

Curves resembling Figure 22-2 can be obtained experimentally by measuring the potential of a silver electrode immersed in the analyte solution. Hence, a chloride/iodide mixture can be analyzed for each of its components. This technique is not as applicable to analyzing iodide/bromide or bromide/chloride mixtures, however, because the solubility differences between the silver salts are not great enough. Thus, the more soluble salt begins to form in significant amounts before precipitation of the less soluble salt is complete. The silver indicator electrode can be a commercial billet type or simply a polished wire. A calomel electrode can be used as reference, although diffusion of chloride ion from the salt bridge may cause the results of the titration to be measurably high. This source of error can be eliminated by placing the calomel electrode in a potassium nitrate solution that is in contact with the analyte solution by means of a KNO₃ salt bridge. Alternatively, the analyte solution can be made slightly acidic with several drops of nitric acid; a glass electrode can then serve as the reference electrode because the pH of the solution and thus its potential will remain essentially constant throughout the titration.
The titration of I\(^{-}\)/\(\text{H}_2\text{O}_2\)/Cl\(^{-}\)/\(\text{H}_2\text{O}_2\) mixtures demonstrates how a potentiometric titration can have multiple end points. The potential of the silver electrode is proportional to pAg. Thus, a plot of \(E_{Ag}\) against titrant volume yields an experimental curve with the same shape as the theoretical curve shown in Figure 22-2 (the ordinate units will be different, of course).

Experimental curves for the titration of I\(^{-}\)/\(\text{H}_2\text{O}_2\)/Cl\(^{-}\)/\(\text{H}_2\text{O}_2\) mixtures do not show the sharp discontinuity that occurs at the first equivalence point of the theoretical curve (Figure 22-2). More important, the volume of silver nitrate needed to reach the I\(^{-}\) end point is generally somewhat greater than theoretical. This effect is the result of coprecipitation of the more soluble AgCl during formation of the less soluble AgI. An overconsumption of reagent thus occurs in the first part of the titration. The total volume closely approaches the correct amount.

Despite this coprecipitation error, the potentiometric method is useful for the analysis of halide mixtures. With approximately equal quantities of iodide and chloride, relative errors can be kept within 2% relative.

**Preparation of Reagents**

1. **Silver nitrate, 0.05 M.** Follow the instructions in Section 22D-1.
2. **Potassium nitrate salt bridge.** Bend an 8-mm glass tube into a U-shape with arms that are long enough to extend nearly to the bottom of two 100-mL beakers. Heat 50 mL of water to boiling, and stir in 1.8 g of powdered agar; continue to heat
and stir until a uniform suspension is formed. Dissolve 12 g of KNO_3 in the hot suspension. Allow the mixture to cool somewhat. Clamp the U-tube with the openings facing up, and use a medicine dropper to fill it with the warm agar suspension. Cool the tube under a cold-water tap to form the gel. When the bridge is not in use, immerse the ends in 2.5 M KNO_3.

Procedure

Obtain the unknown in a clean 250-mL volumetric flask; dilute to the mark with water, and mix well.

Transfer 50.00 mL of the sample to a clean 100-mL beaker, and add 1 or 2 drops of concentrated HNO_3. Place about 25 mL of 2.5 M KNO_3 in a second 100-mL beaker, and make contact between the two solutions with the agar salt bridge. Immerse a silver electrode in the analyte solution and a calomel reference electrode in the second beaker. Titrate with AgNO_3 as described in Section 22J-1. Use small increments of titrant in the vicinity of the two end points.

Plot the data, and establish end points for the two analyte ions. Plot a theoretical titration curve, assuming the measured concentrations of the two constituents to be correct.

Report the number of milligrams of I^- and Cl^- in the sample or as otherwise instructed.

22J-3 The Potentiometric Determination of Solute Species in a Carbonate Mixture

Discussion

A glass/calomel electrode system can be used to locate end points in neutralization titrations and to estimate dissociation constants. As a preliminary step to the titrations, the electrode system is standardized against a buffer of known pH.

The unknown is issued as an aqueous solution prepared from one or perhaps two adjacent members of the following series: NaHCO_3, Na_2CO_3, and NaOH (see Section 9B-2). The object is to determine which of these components were used to prepare the unknown as well as the weight percent of each solute.

Most unknowns require a titration with either standard acid or standard base. A few may require separate titrations, one with acid and one with base. The initial pH of the unknown provides guidance concerning the appropriate titrant(s); a study of Figure 9-3 and Table 9-2 may be helpful in interpreting the data.

Preparation of Solutions

*Standardized 0.1 M HCl and/or 0.1 M NaOH.* Follow the directions in Sections 22C-3 through 22C-7.

Procedure

Obtain the unknown in a clean 250-mL volumetric flask. Dilute to the mark, and mix well. Transfer a small amount of the diluted unknown to a beaker, and determine its
pH. Titrates a 50.00-mL aliquot with standard acid or standard base (or perhaps both). Use the resulting titration curves to select indicator(s) suitable for end-point detection, and perform duplicate titrations with these indicators.

Identify the solute species in the unknown, and report the mass/volume percent of each. Calculate the approximate dissociation constant that can be obtained for any carbonate-containing species from the titration data. Estimate the ionic strength of the solution and correct the calculated constant to give an approximate thermodynamic dissociation constant.

22J-4 The Direct Potentiometric Determination of Fluoride Ion

Discussion
The solid-state fluoride electrode (Section 13D-6) has found extensive use in the determination of fluoride in a variety of materials. Directions follow for the determination of this ion in drinking water and in toothpaste. A total ionic strength adjustment buffer (TISAB) is used to adjust all unknowns and standards to essentially the same ionic strength; when this reagent is used, the concentration of fluoride, rather than its activity, is measured. The pH of the buffer is about 5, a level at which $\text{F}^-$ is the predominant fluorine-containing species. The buffer also contains cyclohexylaminodinitrilotetraacetic acid, which forms stable chelates with iron(III) and aluminum(III), thus freeing fluoride ion from its complexes with these cations.

Review Sections 13D and 13F before undertaking these experiments.

Preparation of Solutions

1. **Total ionic strength adjustment buffer (TISAB).** This solution is marketed commercially under the trade name TISAB. Sufficient buffer for 15 to 20 determinations can be prepared by mixing (with stirring) 57 mL of glacial acetic acid, 58 g of NaCl, 4 g of cyclohexylaminodinitrilotetraacetic acid, and 500 mL of distilled water in a 1-L beaker. Cool the contents in a water or ice bath, and carefully add 6 M NaOH to a pH of 5.0 to 5.5. Dilute to 1 L with water, and store in a plastic bottle.

2. **Standard fluoride solution, 100 ppm.** Dry a quantity of NaF at 110°C for 2 hr. Cool in a desiccator; then weigh (to the nearest milligram) 0.22 g into a 1-L volumetric flask. (Caution! NaF is highly toxic. Immediately wash any skin touched by this compound with copious quantities of water.) Dissolve in water, dilute to the mark, mix well, and store in a plastic bottle. Calculate the exact concentration of fluoride in parts per million.

A standard $\text{F}^-$ solution can be purchased from commercial sources.

Procedure

The apparatus for this experiment consists of a solid-state fluoride electrode, a saturated calomel electrode, and a pH meter. A sleeve-type calomel electrode is needed for
the determination because the measurement is made on a suspension that tends to clog the liquid junction. The sleeve must be loosened momentarily to renew the interface after each series of measurements.

Determining Fluoride in Drinking Water

Transfer 50.00-mL portions of the water to 100-mL volumetric flasks, and dilute to the mark with TISAB solution.

Prepare a 5-ppm $\text{F}^{-}/\text{H}_2\text{SO}_4$ solution by diluting 25.0 mL of the 100-ppm standard to 500 mL in a volumetric flask. Transfer 5.00-, 10.0-, 25.0-, and 50.0-mL aliquots of the 5-ppm solution to 100-mL volumetric flasks, add 50 mL of TISAB solution, and dilute to the mark. (These solutions correspond to 0.5, 1.0, 2.5, and 5.0 ppm $\text{F}^{-}$ in the sample.) After thorough rinsing and drying with paper tissue, immerse the electrodes in the 0.5-ppm standard. Stir mechanically for 3 min; then record the potential. Repeat with the remaining standards and samples.

Plot the measured potential against the log of the concentration of the standards. Use this plot to determine the concentration in parts per million of fluoride in the unknown.

Determining Fluoride in Toothpaste

Weigh (to the nearest milligram) 0.2 g of toothpaste into a 250-mL beaker. Add 50 mL of TISAB solution, and boil for 2 min with good mixing. Cool and then transfer the suspension quantitatively to a 100-mL volumetric flask, dilute to the mark with distilled water, and mix well. Follow the directions for the analysis of drinking water, beginning with the second paragraph.

Report the parts per million of $\text{F}^{-}$ in the sample.

22K ELECTROGRAVIMETRIC METHODS

A convenient example of an electrogravimetric method of analysis is the simultaneous determination of copper and lead in a sample of brass. Additional information concerning electrogravimetric methods is found in Section 20C.

22K-1 The Electrogravimetric Determination of Copper and Lead in Brass

Discussion

This procedure is based on the deposition of metallic copper on a cathode and of lead as $\text{PbO}_2$ on an anode. As a first step, the hydrous oxide of tin ($\text{SnO}_2 \cdot x\text{H}_2\text{O}$) that forms when the sample is treated with nitric acid must be removed by filtration. Lead dioxide is deposited quantitatively at the anode from a solution with a high nitrate ion concentration; copper is only partially deposited on the cathode under these conditions. Therefore, it is necessary to eliminate the excess nitrate after deposition of the $\text{PbO}_2$ is complete. Removal is accomplished through the addition of urea:

22K Electrogravimetric Methods

Copper then deposits quantitatively from the solution after the nitrate ion concentration has been decreased.

**Procedure**

**Preparation of Electrodes**

Immerse the platinum electrodes in hot 6 M HNO₃ for about 5 min (note 1). Wash them thoroughly with distilled water, rinse with several small portions of ethanol, and dry in an oven at 110°C for 2 to 3 min. Cool and weigh both anodes and cathodes to the nearest 0.1 mg.

**Preparation of Samples**

It is not necessary to dry the unknown. Weigh (to the nearest 0.1 mg) 1-g samples into 250-mL beakers. Cover the beakers with watch glasses. Cautiously add about 35 mL of 6 M HNO₃ (HOOD). Digest for at least 30 min; add more acid if necessary. Evaporate to about 5 mL but never to dryness (note 2).

To each sample, add 5 mL of 3 M HNO₃, 25 mL of water, and one-quarter of a tablet of filter paper pulp; digest without boiling for about 45 min. Filter off the SnO₂·H₂O, using a fine-porosity filter paper (note 3); collect the filtrates in tall-form electrolysis beakers. Use many small washes with hot 0.3 M HNO₃ to remove the last traces of copper; test for completeness with a few drops of NH₃. The final volume of filtrate and washings should be between 100 and 125 mL; either add water or evaporate to attain this volume.

**Electrolysis**

With the current switch off, attach the cathode to the negative terminal and the anode to the positive terminal of the electrolysis apparatus. Briefly turn on the stirring motor to be sure the electrodes do not touch. Cover the beakers with split watch glasses, and commence the electrolysis. Maintain a current of 1.3 A for 35 min.

Rinse the cover glasses, and add 10 mL of 3 M H₂SO₄ followed by 5 g of urea to each beaker. Maintain a current of 2 A until the solutions are colorless. To test for completeness of the electrolysis, remove one drop of the solution with a medicine dropper, and mix it with a few drops of NH₃ in a small test tube. If the mixture turns blue, rinse the contents of the tube back into the electrolysis vessel and continue the electrolysis for an additional 10 min. Repeat the test until no blue Cu(NH₃)₂⁺ is produced.

When electrolysis is complete, discontinue stirring but leave the current on. Rinse the electrodes thoroughly with water as they are removed from the solution. After rinsing is complete, turn off the electrolysis apparatus (note 4), disconnect the electrodes, and dip them in acetone. Dry the cathodes for about 3 min and the anodes for about 15 min at 110°C. Allow the electrodes to cool in air, and then weigh them.

Report the percentages of lead (note 5) and copper in the brass.

**Notes**

1. Alternatively, grease and organic materials can be removed by heating platinum electrodes to redness in a flame. Electrode surfaces should not be touched with your fingers after cleaning because grease and oil cause nonadherent deposits that can flake off during washing and weighing.

2. Chloride ion must be totally excluded from this determination because it attacks the platinum anode during electrolysis. This reaction is not only destructive but
also causes positive errors in the analysis by codepositing platinum with copper on the cathode.

3. If desired, the tin content can be determined gravimetrically by ignition of the SnO₂ · H₂O to SnO₂.

4. It is important to maintain a potential between the electrodes until they have been removed from the solution and washed. Some copper may redissolve if this precaution is not observed.

5. Experience has shown that a small amount of moisture is retained by the PbO₂ and that better results are obtained if 0.8643 is used instead of 0.8662, the stoichiometric factor.

22L METHODS BASED ON THE ABSORPTION OF RADIATION

Molecular absorption methods are discussed in Chapter 15. Directions follow for (1) the use of a calibration curve for the determination of iron in water, (2) the use of a standard-addition procedure for the determination of manganese in steel, and (3) a spectrophotometric determination of the pH of a buffer solution.

22L-1 The Cleaning and Handling of Cells

The accuracy of spectrophotometric measurements is critically dependent on the availability of good-quality matched cells. These cells should be calibrated against one another at regular intervals to detect differences resulting from scratches, etching, and wear. Equally important is the proper cleaning of the exterior sides (the windows) just before the cells are inserted into a photometer or spectrophotometer. The preferred method is to wipe the windows with a lens paper soaked in methanol; the methanol is then allowed to evaporate, leaving the windows free of contaminants. It has been shown that this method is far superior to the usual procedure of wiping the windows with a dry lens paper, which tends to leave a residue of lint and a film on the window.¹⁰

22L-2 The Determination of Iron in a Natural Water

Discussion

The red-orange complex that forms between iron(II) and 1,10-phenanthroline (orthophenanthroline) is useful for determining iron in water supplies. The reagent is a weak base that reacts to form phenanthrolinium ion, PhenH⁺, in acidic media. Complex formation with iron is thus best described by the equation

\[
Fe^{2+} + 3\text{PhenH}^+ \rightleftharpoons Fe(\text{Phen})_3^{2+} + 3\text{H}^+
\]

¹⁰For further information, see J. O. Erickson and T. Surles, *Amer. Lab.*, 1976, 8 (6), 50.
The structure of the complex is shown in Section 12D-1. The formation constant for this equilibrium is $2.5 \times 10^6$ at 25°C. Iron(II) is quantitatively complexed in the pH range between 3 and 9. A pH of about 3.5 is ordinarily recommended to prevent precipitation of iron salts, such as phosphates.

An excess of a reducing reagent, such as hydroxylamine or hydroquinone, is needed to maintain iron in the +2 state. The complex, once formed, is very stable.

This determination can be performed with a spectrophotometer set at 508 nm or with a photometer equipped with a green filter.

### Preparation of Solutions

1. **Standard iron solution, 0.01 mg/mL.** Weigh (to the nearest 0.2 mg) 0.0702 g of reagent-grade Fe(NH$_4$)$_2$(SO$_4$)$_2$·6H$_2$O into a 1-L volumetric flask. Dissolve in 50 mL of water that contains 1 to 2 mL of concentrated sulfuric acid; dilute to the mark, and mix well.

2. **Hydroxylamine hydrochloride** (sufficient for 80 to 90 measurements). Dissolve 10 g of H$_2$NOH·HCl in about 100 mL of distilled water.

3. **Orthophenanthroline solution** (sufficient for 80 to 90 measurements). Dissolve 1.0 g of orthophenanthroline monohydrate in about 1 L of water. Warm slightly if necessary. Each milliliter is sufficient for no more than about 0.09 mg of Fe. Prepare no more reagent than needed; it darkens on standing and must then be discarded.

4. **Sodium acetate, 1.2 M** (sufficient for 80 to 90 measurements). Dissolve 166 g of NaOAc·3H$_2$O in 1 L of distilled water.

### Procedure

#### Preparation of a Calibration Curve

Transfer 25.00 mL of the standard iron solution to a 100-mL volumetric flask and 25 mL of distilled water to a second 100-mL volumetric flask. Add 1 mL of hydroxylamine, 10 mL of sodium acetate, and 10 mL of orthophenanthroline to each flask. Allow the mixtures to stand for 5 min; dilute to the mark, and mix well.

Clean a pair of matched cells for the instrument. Rinse each cell with at least three portions of the solution it is to contain. Determine the absorbance of the standard with respect to the blank.

Repeat this procedure with at least three other volumes of the standard iron solution; attempt to encompass an absorbance range between 0.1 and 1.0. Plot a calibration curve.

#### Determination of Iron

Transfer 10.00 mL of the unknown to a 100-mL volumetric flask; treat in exactly the same way as the standards, measuring the absorbance with respect to the blank. Alter the volume of unknown taken to obtain absorbance measurements for replicate samples that are within the range of the calibration curve.

Report the parts per million of iron in the unknown.
22L-3 The Determination of Manganese in Steel

Discussion

Small quantities of manganese are readily determined photometrically by the oxidation of Mn(II) to the intensely colored permanganate ion. Potassium periodate is an effective oxidizing reagent for this purpose. The reaction is

\[ 5\text{IO}_4^- + 2\text{Mn}^{2+} + 3\text{H}_2\text{O} \rightarrow 5\text{IO}_3^- + 2\text{MnO}_4^- + 6\text{H}^+ \]

Permanganate solutions that contain an excess of periodate are quite stable.

Interferences to the method are few. The presence of most colored ions can be compensated for with a blank. Cerium(III) and chromium(III) are exceptions; these yield oxidation products with periodate that absorb to some extent at the wavelength used for the measurement of permanganate.

The method given here is applicable to steels that do not contain large amounts of chromium. The sample is dissolved in nitric acid. Any carbon in the steel is oxidized with peroxydisulfate. Iron(III) is eliminated as a source of interference by complexation with phosphoric acid. The standard-addition method (page 480) is used to establish the relationship between absorbance and amount of manganese in the sample.

A spectrophotometer set at 525 nm or a photometer with a green filter can be used for the absorbance measurements.

Preparation of Solutions

*Standard manganese(II) solution* (sufficient for several hundred analyses). Weigh 0.1 g (to the nearest 0.1 mg) of manganese into a 50-mL beaker, and dissolve in about 10 mL of 6 M HNO₃ (HOOD). Boil gently to eliminate oxides of nitrogen. Cool; then transfer the solution quantitatively to a 1-L volumetric flask. Dilute to the mark with water, and mix thoroughly. The manganese in 1 mL of the standard solution, after being converted to permanganate, causes a volume of 50 mL to increase in absorbance by about 0.09.

Procedure

The unknown does not require drying. If there is evidence of oil, rinse with acetone and dry briefly. Weigh (to the nearest 0.1 mg) duplicate samples (note 1) into 150-mL beakers. Add about 50 mL of 6 M HNO₃, and boil gently (HOOD); heating for 5 to 10 min should suffice. Cautiously add about 1 g of ammonium peroxydisulfate, and boil gently for an additional 10 to 15 min. If the solution is pink or has a deposit of MnO₂, add 1 mL of NH₄HSO₃ (or 0.1 g of NaHSO₃) and heat for 5 min. Cool; transfer quantitatively (note 2) to 250.0-mL volumetric flasks. Dilute to the mark with water, and mix well. Use a 20.0-mL pipet to transfer three aliquots of each sample to individual beakers. Treat as follows:

<table>
<thead>
<tr>
<th>Aliquot</th>
<th>Volume of 85% H₃PO₄, mL</th>
<th>Volume of Standard Mn, mL</th>
<th>Weight of KIO₄, g</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>5</td>
<td>0.00</td>
<td>0.4</td>
</tr>
<tr>
<td>2</td>
<td>5</td>
<td>5.00 (note 3)</td>
<td>0.4</td>
</tr>
<tr>
<td>3</td>
<td>5</td>
<td>0.00</td>
<td>0.0</td>
</tr>
</tbody>
</table>
Boil each solution gently for 5 min, cool, and transfer it quantitatively to a 50-mL volumetric flask. Mix well. Measure the absorbance of aliquots 1 and 2 using aliquot 3 as the blank (note 4).

Report the percentage of manganese in the unknown.

Notes
1. The sample size depends on the manganese content of the unknown; consult with the instructor.
2. If there is evidence of turbidity, filter the solutions as they are transferred to the volumetric flasks.
3. The volume of the standard addition may be dictated by the absorbance of the sample. It is useful to obtain a rough estimate by generating permanganate in about 20 mL of sample, diluting to about 50 mL, and measuring the absorbance.
4. A single blank can be used for all measurements, provided the samples weigh within 50 mg of one another.

## 22L-4 The Spectrophotometric Determination of pH

### Discussion

The pH of an unknown buffer is determined by addition of an acid/base indicator and spectrophotometric measurement of the absorbance of the resulting solution. Because overlap exists between the spectra for the acid and base forms of the indicator, it is necessary to evaluate individual molar absorptivities for each form at two wavelengths. See page 483 for further discussion.

The relationship between the two forms of bromocresol green in an aqueous solution is described by the equilibrium

$HIn + H_2O \leftrightarrow H_3O^+ + In^-$

for which

$$K_a = \frac{[H_3O^+][In^-]}{[HIn]} = 1.6 \times 10^{-5}$$

The spectrophotometric evaluation of $[In^-]$ and $[HIn]$ permits the calculation of $[H_3O^+]$ and thus pH.

### Preparation of Solutions

1. **Bromocresol green**, $1.0 \times 10^{-4} M$ (sufficient for about five determinations). Dissolve 40.0 mg (to the nearest 0.1 mg) of the sodium salt of bromocresol green ($720 \text{ g/mol}$) in water, and dilute to 500 mL in a volumetric flask.
2. **HCl**, $0.5 M$. Dilute about 4 mL of concentrated HCl to approximately 100 mL with water.
3. **NaOH**, $0.4 M$. Dilute about 7 mL of 6 M NaOH to about 100 mL with water.
Procedure

Determination of Individual Absorption Spectra
Transfer 25.00-mL aliquots of the bromocresol green indicator solution to two 100-mL volumetric flasks. To one, add 25 mL of 0.5 M HCl; to the other, add 25 mL of 0.4 M NaOH. Dilute to the mark and mix well.

Obtain the absorption spectra for the acid and conjugate-base forms of the indicator between 400 and 600 nm, using water as a blank. Record absorbance values at 10-nm intervals routinely and at closer intervals as needed to define maxima and minima. Evaluate the molar absorptivity for HIn and In⁻ at wavelengths corresponding to their absorption maxima.

Determination of the pH of an Unknown Buffer
Transfer 25.00 mL of the stock bromocresol green indicator to a 100-mL volumetric flask. Add 50.0 mL of the unknown buffer, dilute to the mark, and mix well. Measure the absorbance of the diluted solution at the wavelengths for which absorptivity data were calculated.

Report the pH of the buffer.

22M ATOMIC SPECTROSCOPY

Several methods of analysis based on atomic spectroscopy are discussed in Section 15F. One such application is atomic absorption, which is demonstrated in the experiment that follows.

22M-1 The Determination of Lead in Brass by Atomic Absorption Spectroscopy

Discussion
Brasses and other copper-based alloys contain from 0 to about 10% lead as well as tin and zinc. Atomic absorption spectroscopy permits the quantitative estimation of these elements. The accuracy of this procedure is not as great as that obtainable with gravimetric or volumetric measurements, but the time needed to acquire the analytical information is considerably less.

A weighed sample is dissolved in a mixture of nitric and hydrochloric acids, the latter being needed to prevent the precipitation of tin as metastannic acid, SnO₂·xH₂O. After suitable dilution, the sample is aspirated into a flame, and the absorption of radiation from a hollow cathode lamp is measured.

Preparation of Solutions

_Standard lead solution, 100 mg/L._ Dry a quantity of reagent-grade Pb(NO₃)₂ for about 1 hr at 110°C. Cool; weigh (to the nearest 0.1 mg) 0.17 g into a 1-L volumetric flask. Dissolve in a solution of 5 mL water and 1 to 3 mL of concentrated HNO₃. Dilute to the mark with distilled water, and mix well.
Procedure

Weigh duplicate samples of the unknown (note 1) into 150-mL beakers. Cover with watch glasses; then dissolve (HOOD) in a mixture consisting of about 4 mL of concentrated HNO₃ and 4 mL of concentrated HCl (note 2). Boil gently to remove oxides of nitrogen. Cool; transfer the solutions quantitatively to a 250-mL volumetric flask, dilute to the mark with water, and mix well.

Use a buret to deliver 0-, 5-, 10-, 15-, and 20-mL portions of the standard lead solution to individual 50-mL volumetric flasks. Add 4 mL of concentrated HNO₃ and 4 mL of concentrated HCl to each, and dilute to the mark with water.

Transfer 10.00-mL aliquots of each sample to 50-mL volumetric flasks, add 4 mL of concentrated HCl and 4 mL of concentrated HNO₃, and dilute to the mark with water.

Set the monochromator at 283.3 nm, and measure the absorbance for each standard and the sample at that wavelength. Take at least three, and preferably more, readings for each measurement.

Plot the calibration data. Report the percentage of lead in the brass.

Notes

1. The weight of sample depends on the lead content of the brass and on the sensitivity of the instrument used for the absorption measurements. A sample containing 6 to 10 mg of lead is reasonable. Consult with the instructor.

2. Brasses that contain a large percentage of tin require additional HCl to prevent the formation of metastannic acid. The diluted samples may develop some turbidity on prolonged standing; a slight turbidity has no effect on the determination of lead.

22M-2 The Determination of Sodium, Potassium, and Calcium in Mineral Waters by Atomic Emission Spectroscopy

Discussion

A convenient method for the determination of alkali and alkaline-earth metals in water and in blood serum is based on the characteristic spectra these elements emit upon being aspirated into a natural gas/air flame. The accompanying directions are suitable for the analysis of the three elements in water samples. Radiation buffers are used to minimize the effect of each element on the emission intensity of the others.

Preparation of Solutions

1. **Standard calcium solution, approximately 500 ppm.** Dry a quantity of CaCO₃ for about 1 hr at 110°C. Cool in a desiccator; weigh (to the nearest milligram) 1.25 g into a 600-mL beaker. Add about 200 mL of distilled water and about 10 mL of concentrated HCl; cover the beaker with a watch glass during the addition of the acid to avoid loss due to spattering. After reaction is complete, transfer the solution quantitatively to a 1-L volumetric flask, dilute to the mark, and mix well.

2. **Standard potassium solution, approximately 500 ppm.** Dry a quantity of KCl for about 1 hr at 110°C. Cool; weigh (to the nearest milligram) about 0.95 g into a 1-L volumetric flask. Dissolve in distilled water, and dilute to the mark.

A radiation buffer is a substance that is added in large excess to both samples and standards to swamp out the effect of matrix species and thus minimize interferences.
3. Standard sodium solution, approximately 500 ppm. Proceed as in part 2, using 1.25 g (to the nearest milligram) of dried NaCl.
4. Radiation buffer for the determination of calcium. Prepare about 100 mL of a solution that has been saturated with NaCl, KCl, and MgCl₂, in that order.
5. Radiation buffer for the determination of potassium. Prepare about 100 mL of a solution that has been saturated with NaCl, CaCl₂, and MgCl₂, in that order.
6. Radiation buffer for the determination of sodium. Prepare about 100 mL of a solution that has been saturated with CaCl₂, KCl, and MgCl₂, in that order.

Procedure

Preparation of Working Curves

Add 5.00 mL of the appropriate radiation buffer to each of a series of 100-mL volumetric flasks. Add volumes of standard that will produce solutions that range from 0 to 10 ppm in the cation to be determined. Dilute to the mark with water, and mix well.

Measure the emission intensity for each solution, taking at least three readings for each. Aspirate distilled water between each set of measurements. Correct the average values for background luminosity, and prepare a working curve from the data.

Repeat for the other two cations.

Analysis of a Water Sample

Prepare duplicate aliquots of the unknown as directed for measurement with the working curves. If necessary, use a standard to adjust the response of the instrument to the working curve; then measure the emission intensity of the unknown. Correct the data for background. Determine the cation concentration in the unknown by comparison with the working curve.

22N SEPARATING CATIONS BY ION EXCHANGE

The application of ion-exchange resins to the separation of ionic species of opposite charge is discussed in Section 16E. Directions follow for the ion-exchange separation of nickel(II) from zinc(II) based on converting the zinc ions to negatively charged chloro complexes. After separation, each of the cations is determined by EDTA titration.

Discussion

The separation of the two cations is based on differences in their tendency to form anionic complexes. Stable chlorozincate(II) complexes (such as ZnCl₃⁻ and ZnCl₅²⁻) are formed in 2-M hydrochloric acid and retained on an anion-exchange resin. In contrast, nickel(II) is not complexed appreciably in this medium and passes rapidly through such a column. After separation is complete, elution with water effectively decomposes the chloro complexes and permits removal of the zinc.

Both nickel and zinc are determined by titration with standard EDTA at pH 10. Eriochrome Black T is the indicator for the zinc titration. Bromopyrogallol Red or murexide is used for the nickel titration.
Preparation of Solutions

1. **Standard EDTA.** See Section 22E-2.
2. **pH 10 buffer.** See Section 22E-1.
3. **Eriochrome Black T indicator.** See Section 22E-1.
4. **Bromopyrogallol Red indicator** (sufficient for 100 titrations). Dissolve 0.5 g of the solid indicator in 100 mL of 50% (v/v) ethanol.
5. **Murexide indicator.** The solid is approximately 0.2% indicator by weight in NaCl. Approximately 0.2 g is needed for each titration. The solid preparation is used because solutions of the indicator are quite unstable.

Preparation of Ion-Exchange Columns

A typical ion-exchange column is a cylinder 25 to 40 cm in length and 1 to 1.5 cm in diameter. A stopcock at the lower end permits adjustment of liquid flow through the column. A buret makes a convenient column. It is recommended that two columns be prepared to permit the simultaneous treatment of duplicate samples.

Insert a plug of glass wool to retain the resin particles. Then introduce sufficient strong-base anion-exchange resin (note) to give a 10- to 15-cm column. Wash the column with about 50 mL of 6 M NH₃, followed by 100 mL of water and 100 mL of 2 M HCl. At the end of this cycle, the flow should be stopped so that the liquid level remains about 1 cm above the resin column. **At no time should the liquid level be allowed to drop below the top of the resin.**

Note

Amberlite® CG 400 or its equivalent can be used.

Procedure

Obtain the unknown, which should contain between 2 and 4 mmol of Ni²⁺ and Zn²⁺, in a clean 100-mL volumetric flask. Add 16 mL of 12 M HCl, dilute to the mark with distilled water, and mix well. The resulting solution is approximately 2 M in acid. Transfer 10.00 mL of the diluted unknown onto the column. Place a 250-mL conical flask beneath the column, and slowly drain until the liquid level is barely above the resin. Rinse the interior of the column with several 2- to 3-mL portions of the 2 M HCl; lower the liquid level to just above the resin surface after each washing. Elute the nickel with about 50 mL of 2 M HCl at a flow rate of 2 to 3 mL/min.

Elute the Zn(II) by passing about 100 mL of water through the column, using the same flow rate; collect the liquid in a 500-mL conical flask.

Titration of Nickel

Evaporate the solution containing the nickel to dryness to eliminate excess HCl. Avoid overheating; the residual NiCl₂ must not be permitted to decompose to NiO. Dissolve the residue in 100 mL of distilled water, and add 10 to 20 mL of pH 10 buffer. Add 15 drops of Bromopyrogallol Red indicator or 0.2 g of murexide. Titrare to the color change (blue to purple for Bromopyrogallol Red, yellow to purple for murexide).

Calculate the milligrams of nickel in the unknown.
22O GAS-LIQUID CHROMATOGRAPHY

As noted in Chapter 17, gas-liquid chromatography enables the analyst to resolve the components of complex mixtures. The accompanying directions are for the determination of ethanol in beverages.

22O-1 The Gas-Chromatographic Determination of Ethanol in Beverages

Discussion

Ethanol is conveniently determined in aqueous solutions by means of gas chromatography. The method is readily extended to measurement of the proof of alcoholic beverages. By definition, the proof of a beverage is two times its volume percent of ethanol at 60°F.

The operating instructions pertain to a 1/4-in. (o.d.) × 0.5-m Poropack column containing 80- to 100-mesh packing. A thermal conductivity detector is needed. (Flame ionization is not satisfactory because of its insensitivity to water.)

The determination is based on a calibration curve in which the ratio of the area under the ethanol peak to the area under the ethanol-plus-water peak is plotted as a function of the volume percent of ethanol:

\[
vol \% \text{ EtOH} = \frac{vol \text{ EtOH}}{vol \text{ soln}} \times 100\%
\]

This relationship is not strictly linear. At least two reasons can be cited to account for the curvature. First, the thermal conductivity detector responds linearly to mass ratios rather than volume ratios. Second, at the high concentrations involved, the volumes of ethanol and water are not strictly additive, as would be required for linearity. That is,

\[vol \text{ EtOH} + vol \text{ H}_2\text{O} \neq vol \text{ soln}\]

Preparation of Standards

Use a buret to deliver 10.00, 20.00, 30.00, and 40.00 mL of absolute ethanol into separate 50-mL volumetric flasks (note). Dilute to volume with distilled water, and mix well.

\[^{11}\text{Adapted from J. J. Leary, J. Chem. Educ., 1983, 60, 675.}\]
Note

The coefficient of thermal expansion for ethanol is approximately five times that for water. It is thus necessary to keep the temperature of the solutions used in this experiment constant to ±1°C during volume measurements.

Procedure

The following operating conditions have yielded satisfactory chromatograms for this determination:

<table>
<thead>
<tr>
<th>Parameter</th>
<th>Setting</th>
</tr>
</thead>
<tbody>
<tr>
<td>Column temperature</td>
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<tr>
<td>Detector temperature</td>
<td>130°C</td>
</tr>
<tr>
<td>Injection-port temperature</td>
<td>120°C</td>
</tr>
<tr>
<td>Bridge current</td>
<td>100 mA</td>
</tr>
<tr>
<td>Flow rate</td>
<td>60 mL/min</td>
</tr>
</tbody>
</table>

Inject a 1 μL sample of the 20% (v/v) standard, and record the chromatogram. Obtain additional chromatograms, adjusting the recorder speed until the water peak has a width of about 2 mm at half-height. Then vary the volume of sample injected and the attenuation until peaks with a height of at least 40 mm are produced. Obtain chromatograms for the remainder of the standards (including pure water and pure ethanol) in the same way. Measure the area under each peak, and plot \( \frac{\text{area}_{\text{EtOH}}}{\text{area}_{\text{EtOH}} + \text{area}_{\text{H}_2\text{O}}} \) as a function of the volume percentage of ethanol.

Obtain chromatograms for the unknown. Report the volume percentage of ethanol.

WEB WORKS

Material Safety Data Sheets (MSDSs) are an important resource for anyone using chemicals. These documents provided essential information on the properties and toxicity of chemicals that are used in the laboratory. Set your browser to http://www.ilpi.com/msds/index.html and find a comprehensive listing of most of the internet sites that present MSDSs. Browse to one of the sites and look up the MSDS for oxalic acid. Browse the entire MSDS, read the reactivity data, note its chemical properties, and find the first aid treatment for ingestion of oxalic acid. Chemical manufacturers are required by law to furnish an MSDS for every chemical that they sell, and many of them can be found on the internet. It is a good idea to examine the MSDS for any substance that you use in the laboratory.