

Chapter 10

Spectroscopic Methods

Chapter Overview

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An early example of a colorimetric analysis is Nessler's method for ammonia, which was introduced in 1856. Nessler found that adding an alkaline solution of HgI_2 and KI to a dilute solution of ammonia produces a yellow to reddish brown colloid, with the colloid's color depending on the concentration of ammonia. By visually comparing the color of a sample to the colors of a series of standards, Nessler was able to determine the concentration of ammonia.

Colorimetry, in which a sample absorbs visible light, is one example of a spectroscopic method of analysis. At the end of the nineteenth century, spectroscopy was limited to the absorption, emission, and scattering of visible, ultraviolet, and infrared electromagnetic radiation. Since its introduction, spectroscopy has expanded to include other forms of electromagnetic radiation—such as X-rays, microwaves, and radio waves—and other energetic particles—such as electrons and ions.

10A Overview of Spectroscopy

The focus of this chapter is on the interaction of ultraviolet, visible, and infrared radiation with matter. Because these techniques use optical materials to disperse and focus the radiation, they often are identified as optical spectroscopies. For convenience we will use the simpler term **SPECTROSCOPY** in place of optical spectroscopy; however, you should understand that we are considering only a limited part of a much broader area of analytical techniques.

Despite the difference in instrumentation, all spectroscopic techniques share several common features. Before we consider individual examples in greater detail, let's take a moment to consider some of these similarities. As you work through the chapter, this overview will help you focus on similarities between different spectroscopic methods of analysis. You will find it easier to understand a new analytical method when you can see its relationship to other similar methods.



Figure 10.1 The Golden Gate bridge as seen through rain drops. Refraction of light by the rain drops produces the distorted images. Source: [Mila Zinkova](#) (commons.wikipedia.org).

10A.1 What is Electromagnetic Radiation

ELECTROMAGNETIC RADIATION—light—is a form of energy whose behavior is described by the properties of both waves and particles. Some properties of electromagnetic radiation, such as its refraction when it passes from one medium to another (Figure 10.1), are explained best by describing light as a wave. Other properties, such as absorption and emission, are better described by treating light as a particle. The exact nature of electromagnetic radiation remains unclear, as it has since the development of quantum mechanics in the first quarter of the 20th century.¹ Nevertheless, the dual models of wave and particle behavior provide a useful description for electromagnetic radiation.

WAVE PROPERTIES OF ELECTROMAGNETIC RADIATION

Electromagnetic radiation consists of oscillating electric and magnetic fields that propagate through space along a linear path and with a constant velocity. In a vacuum electromagnetic radiation travels at the speed of light, c , which is $2.997\,92 \times 10^8$ m/s. When electromagnetic radiation moves through a medium other than a vacuum its velocity, v , is less than the speed of light in a vacuum. The difference between v and c is sufficiently small ($<0.1\%$) that the speed of light to three significant figures, 3.00×10^8 m/s, is accurate enough for most purposes.

The oscillations in the electric and magnetic fields are perpendicular to each other, and to the direction of the wave's propagation. [Figure 10.2](#) shows an example of plane-polarized electromagnetic radiation, consisting of a single oscillating electric field and a single oscillating magnetic field.

An electromagnetic wave is characterized by several fundamental properties, including its velocity, amplitude, frequency, phase angle, polariza-

¹ Home, D.; Gribbin, J. *New Scientist* 1991, 2 Nov. 30–33.

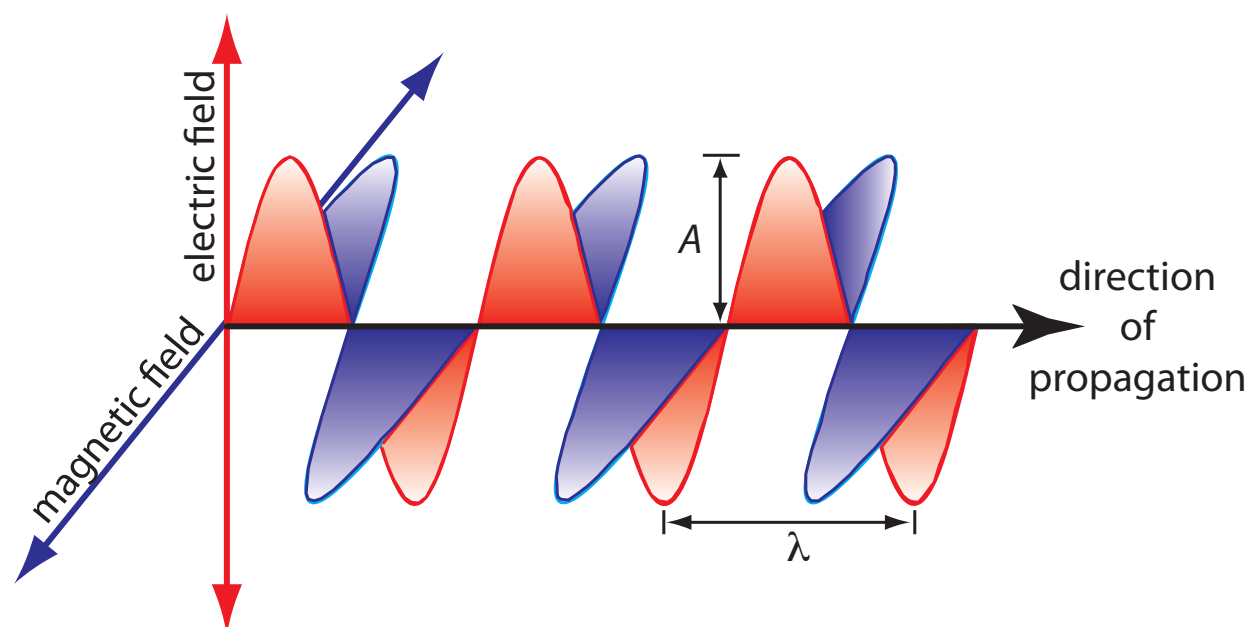


Figure 10.2 Plane-polarized electromagnetic radiation showing the oscillating electric field in red and the oscillating magnetic field in blue. The radiation's amplitude, A , and its wavelength, λ , are shown. Normally, electromagnetic radiation is unpolarized, with oscillating electric and magnetic fields present in all possible planes perpendicular to the direction of propagation.

tion, and direction of propagation.² For example, the amplitude of the oscillating electric field at any point along the propagating wave is

$$A_t = A_e \sin(2\pi\nu t + \Phi)$$

where A_t is the magnitude of the electric field at time t , A_e is the electric field's maximum **AMPLITUDE**, ν is the wave's **FREQUENCY**—the number of oscillations in the electric field per unit time—and Φ is a **PHASE ANGLE**, which accounts for the fact that A_t need not have a value of zero at $t=0$. The identical equation for the magnetic field is

$$A_t = A_m \sin(2\pi\nu t + \Phi)$$

where A_m is the magnetic field's maximum amplitude.

Other properties also are useful for characterizing the wave behavior of electromagnetic radiation. The **WAVELENGTH**, λ , is defined as the distance between successive maxima (see Figure 10.2). For ultraviolet and visible electromagnetic radiation the wavelength is usually expressed in nanometers ($1 \text{ nm} = 10^{-9} \text{ m}$), and for infrared radiation it is given in microns ($1 \mu\text{m} = 10^{-6} \text{ m}$). The relationship between wavelength and frequency is

$$\lambda = \frac{c}{\nu}$$

Another useful unit is the **WAVENUMBER**, $\bar{\nu}$, which is the reciprocal of wavelength

$$\bar{\nu} = \frac{1}{\lambda}$$

Wavenumbers are frequently used to characterize infrared radiation, with the units given in cm^{-1} .

² Ball, D. W. *Spectroscopy* **1994**, 9(5), 24–25.

When electromagnetic radiation moves between different media—for example, when it moves from air into water—its frequency, ν , remains constant. Because its velocity depends upon the medium in which it is traveling, the electromagnetic radiation's wavelength, λ , changes. If we replace the speed of light in a vacuum, c , with its speed in the medium, v , then the wavelength is

$$\lambda = \frac{v}{\nu}$$

This change in wavelength as light passes between two media explains the refraction of electromagnetic radiation shown in [Figure 10.1](#).

Example 10.1

In 1817, Josef Fraunhofer studied the spectrum of solar radiation, observing a continuous spectrum with numerous dark lines. Fraunhofer labeled the most prominent of the dark lines with letters. In 1859, Gustav Kirchhoff showed that the D line in the sun's spectrum was due to the absorption of solar radiation by sodium atoms. The wavelength of the sodium D line is 589 nm. What are the frequency and the wavenumber for this line?

SOLUTION

The frequency and wavenumber of the sodium D line are

$$\nu = \frac{c}{\lambda} = \frac{3.00 \times 10^8 \text{ m/s}}{589 \times 10^{-9} \text{ m}} = 5.09 \times 10^{14} \text{ s}^{-1}$$

$$\bar{\nu} = \frac{1}{\lambda} = \frac{1}{589 \times 10^{-9} \text{ m}} \times \frac{1 \text{ m}}{100 \text{ cm}} = 1.70 \times 10^4 \text{ cm}^{-1}$$

Practice Exercise 10.1

Another historically important series of spectral lines is the Balmer series of emission lines from hydrogen. One of the lines has a wavelength of 656.3 nm. What are the frequency and the wavenumber for this line?

Click [here](#) to review your answer to this exercise.

PARTICLE PROPERTIES OF ELECTROMAGNETIC RADIATION

When matter absorbs electromagnetic radiation it undergoes a change in energy. The interaction between matter and electromagnetic radiation is easiest to understand if we assume that radiation consists of a beam of energetic particles called photons. When a **PHOTON** is absorbed by a sample it is “destroyed,” and its energy acquired by the sample.³ The energy of a photon, in joules, is related to its frequency, wavelength, and wavenumber by the following equalities

$$E = h\nu = \frac{hc}{\lambda} = hc\bar{\nu}$$

where h is Planck's constant, which has a value of $6.626 \times 10^{-34} \text{ J}\cdot\text{s}$.

Example 10.2

What is the energy of a photon from the sodium D line at 589 nm?

SOLUTION

The photon's energy is



³ Ball, D. W. *Spectroscopy* **1994**, 9(6) 20–21.

$$E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} \text{ J} \cdot \text{s})(3.00 \times 10^8 \text{ m/s})}{589 \times 10^{-9} \text{ m}} = 3.37 \times 10^{-19} \text{ J}$$

Practice Exercise 10.2

What is the energy of a photon for the Balmer line at a wavelength of 656.3 nm?

Click [here](#) to review your answer to this exercise.

THE ELECTROMAGNETIC SPECTRUM

The frequency and wavelength of electromagnetic radiation vary over many orders of magnitude. For convenience, we divide electromagnetic radiation into different regions—the **ELECTROMAGNETIC SPECTRUM**—based on the type of atomic or molecular transition that gives rise to the absorption or emission of photons (Figure 10.3). The boundaries between the regions of the electromagnetic spectrum are not rigid, and overlap between spectral regions is possible.

10A.2 Photons as a Signal Source

In the previous section we defined several characteristic properties of electromagnetic radiation, including its energy, velocity, amplitude, frequency, phase angle, polarization, and direction of propagation. A spectroscopic measurement is possible only if the photon's interaction with the sample leads to a change in one or more of these characteristic properties.

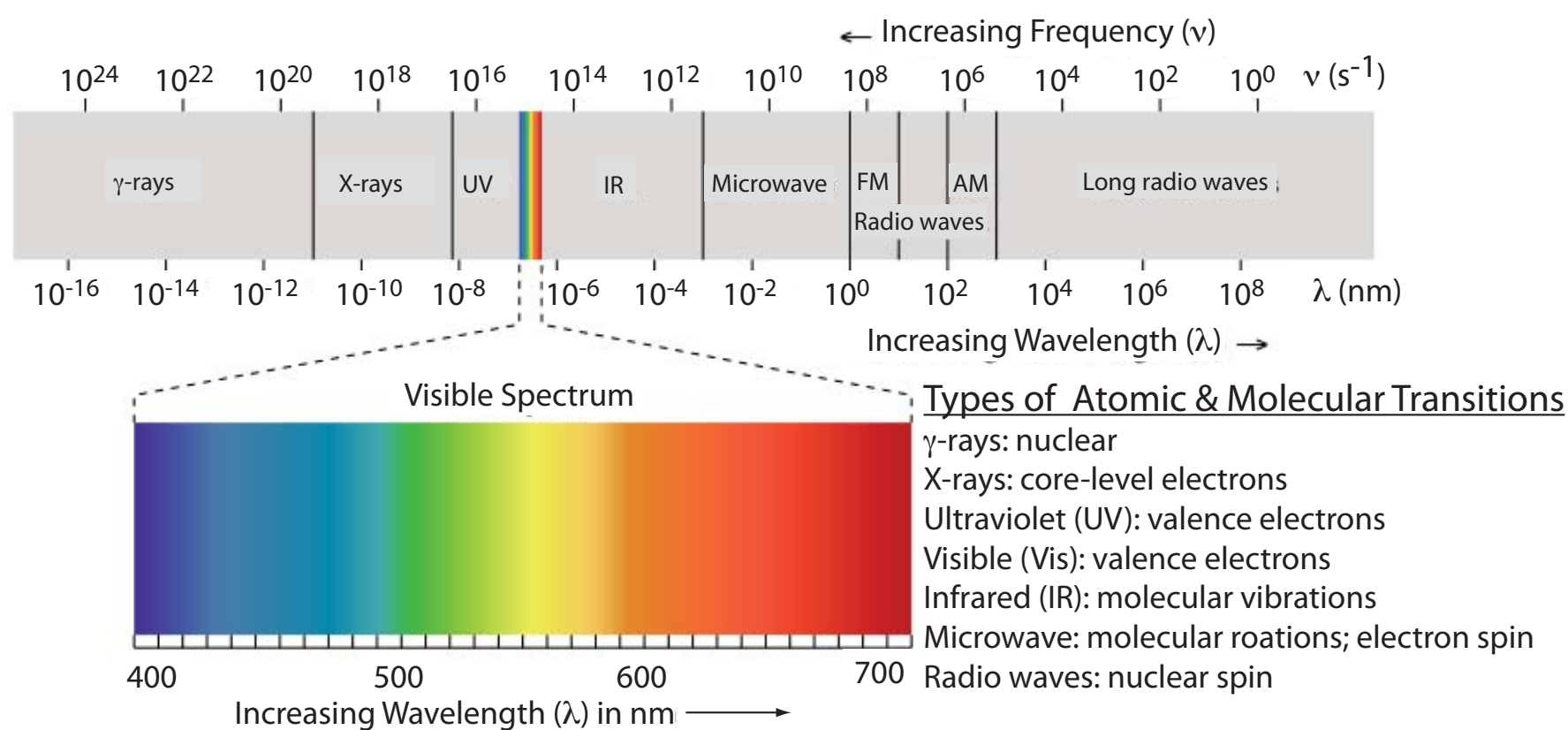


Figure 10.3 The electromagnetic spectrum showing the boundaries between different regions and the type of atomic or molecular transition responsible for the change in energy. The colored inset shows the visible spectrum. Source: modified from [Zedh](http://www.commonswiki.org) (www.commonswiki.org).

Table 10.1 Examples of Spectroscopic Techniques Involving an Exchange of Energy Between a Photon and the Sample

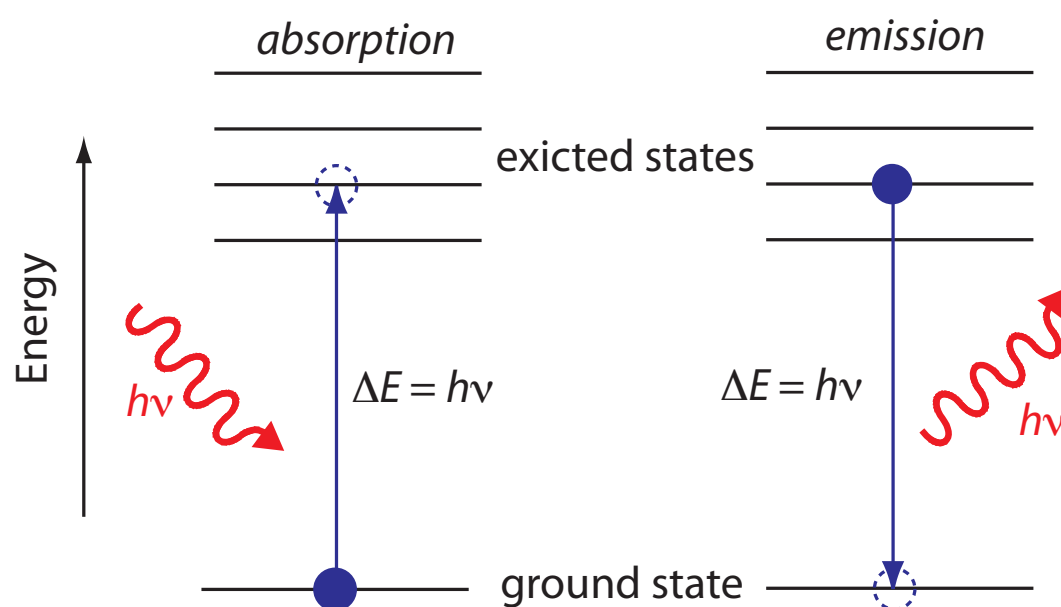
Type of Energy Transfer	Region of Electromagnetic Spectrum	Spectroscopic Technique ^a
absorption	γ -ray	Mossbauer spectroscopy
	X-ray	X-ray absorption spectroscopy
	UV/Vis	<i>UV/Vis spectroscopy</i> <i>atomic absorption spectroscopy</i>
	IR	<i>infrared spectroscopy</i> Raman spectroscopy
	Microwave	microwave spectroscopy
	Radio wave	electron spin resonance spectroscopy nuclear magnetic resonance spectroscopy
	UV/Vis	<i>atomic emission spectroscopy</i>
emission (thermal excitation)		
photoluminescence	X-ray	X-ray fluorescence
	UV/Vis	<i>fluorescence spectroscopy</i> <i>phosphorescence spectroscopy</i> atomic fluorescence spectroscopy
chemiluminescence	UV/Vis	chemiluminescence spectroscopy

^a Techniques discussed in this text are shown in *italics*.

We can divide spectroscopy into two broad classes of techniques. In one class of techniques there is a transfer of energy between the photon and the sample. Table 10.1 provides a list of several representative examples.

In absorption spectroscopy a photon is absorbed by an atom or molecule, which undergoes a transition from a lower-energy state to a higher-energy, or excited state (Figure 10.4). The type of transition depends on the photon's energy. The electromagnetic spectrum in Figure 10.3, for example, shows that absorbing a photon of visible light promotes one of the atom's or molecule's valence electrons to a higher-energy level. When an molecule absorbs infrared radiation, on the other hand, one of its chemical bonds experiences a change in vibrational energy.

Figure 10.4 Simplified energy diagram showing the absorption and emission of a photon by an atom or a molecule. When a photon of energy $h\nu$ strikes the atom or molecule, absorption may occur if the difference in energy, ΔE , between the ground state and the excited state is equal to the photon's energy. An atom or molecule in an excited state may emit a photon and return to the ground state. The photon's energy, $h\nu$, equals the difference in energy, ΔE , between the two states.



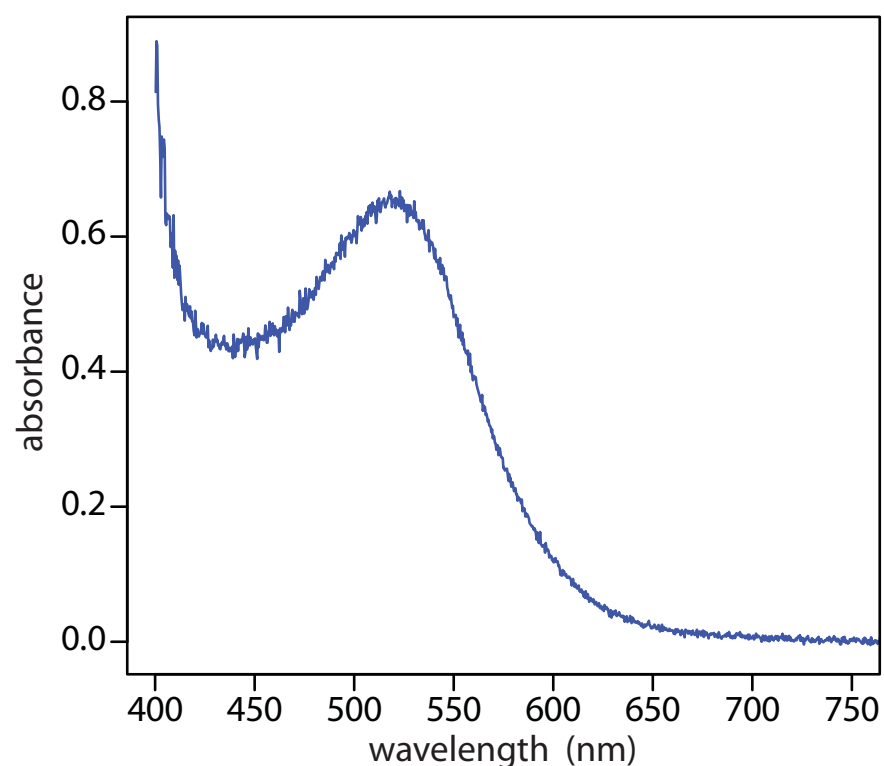


Figure 10.5 Visible absorbance spectrum for cranberry juice. The anthocyanin dyes in cranberry juice absorb visible light with blue, green, and yellow wavelengths (see [Figure 10.3](#)). As a result, the juice appears red.

When it absorbs electromagnetic radiation the number of photons passing through a sample decreases. The measurement of this decrease in photons, which we call **ABSORBANCE**, is a useful analytical signal. Note that each of the energy levels in [Figure 10.4](#) has a well-defined value because they are quantized. Absorption occurs only when the photon's energy, $h\nu$, matches the difference in energy, ΔE , between two energy levels. A plot of absorbance as a function of the photon's energy is called an **ABSORBANCE SPECTRUM**. [Figure 10.5](#), for example, shows the absorbance spectrum of cranberry juice.

When an atom or molecule in an excited state returns to a lower energy state, the excess energy often is released as a photon, a process we call **EMISSION** ([Figure 10.4](#)). There are several ways in which an atom or molecule may end up in an excited state, including thermal energy, absorption of a photon, or by a chemical reaction. Emission following the absorption of a photon is also called **PHOTOLUMINESCENCE**, and that following a chemical reaction is called **CHEMILUMINESCENCE**. A typical emission spectrum is shown in [Figure 10.6](#).

Molecules also can release energy in the form of heat. We will return to this point later in the chapter.

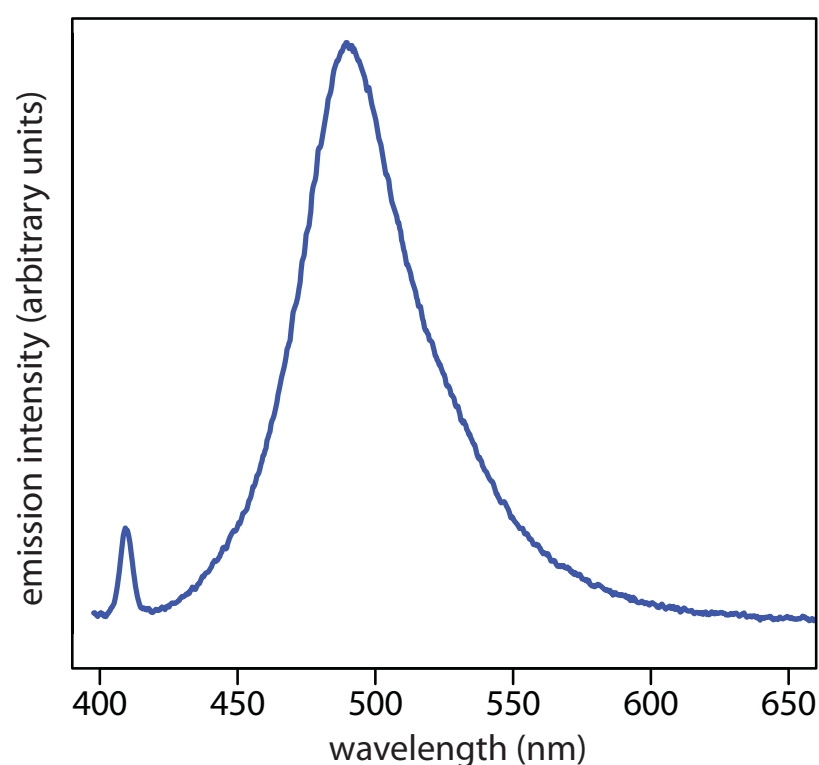


Figure 10.6 Photoluminescence spectrum of the dye coumarin 343, which is incorporated in a reverse micelle suspended in cyclohexanol. The dye's absorbance spectrum (not shown) has a broad peak around 400 nm. The sharp peak at 409 nm is from the laser source used to excite coumarin 343. The broad band centered at approximately 500 nm is the dye's emission band. Because the dye absorbs blue light, a solution of coumarin 343 appears yellow in the absence of photoluminescence. Its photoluminescent emission is blue-green. Source: data from Bridget Gourley, Department of Chemistry & Biochemistry, DePauw University).

Table 10.2 Examples of Spectroscopic Techniques That Do Not Involve an Exchange of Energy Between a Photon and the Sample

Region of Electromagnetic Spectrum	Type of Interaction	Spectroscopic Technique ^a
X-ray	diffraction	X-ray diffraction
UV/Vis	refraction	refractometry
	scattering	<i>nephelometry</i> <i>turbidimetry</i>
	dispersion	optical rotary dispersion

^a Techniques discussed in this text are shown in *italics*.

In the second broad class of spectroscopic techniques, the electromagnetic radiation undergoes a change in amplitude, phase angle, polarization, or direction of propagation as a result of its refraction, reflection, scattering, diffraction, or dispersion by the sample. Several representative spectroscopic techniques are listed in Table 10.2.

10A.3 Basic Components of Spectroscopic Instruments

The spectroscopic techniques in [Table 10.1](#) and Table 10.2 use instruments that share several common basic components, including a source of energy, a means for isolating a narrow range of wavelengths, a detector for measuring the signal, and a signal processor that displays the signal in a form convenient for the analyst. In this section we introduce these basic components. Specific instrument designs are considered in later sections.

SOURCES OF ENERGY

All forms of spectroscopy require a source of energy. In absorption and scattering spectroscopy this energy is supplied by photons. Emission and photoluminescence spectroscopy use thermal, radiant (photon), or chemical energy to promote the analyte to a suitable excited state.

Sources of Electromagnetic Radiation. A source of electromagnetic radiation must provide an output that is both intense and stable. Sources of electromagnetic radiation are classified as either continuum or line sources. A **CONTINUUM SOURCE** emits radiation over a broad range of wavelengths, with a relatively smooth variation in intensity (Figure 10.7). A **LINE SOURCE**, on the other hand, emits radiation at selected wavelengths ([Figure 10.8](#)). [Table 10.3](#) provides a list of the most common sources of electromagnetic radiation.

Sources of Thermal Energy. The most common sources of thermal energy are flames and plasmas. Flames sources use the combustion of a fuel and an oxidant to achieve temperatures of 2000–3400 K. Plasmas, which are hot, ionized gases, provide temperatures of 6000–10 000 K.

You will find a more detailed treatment of these components in the [additional resources](#) for this chapter.

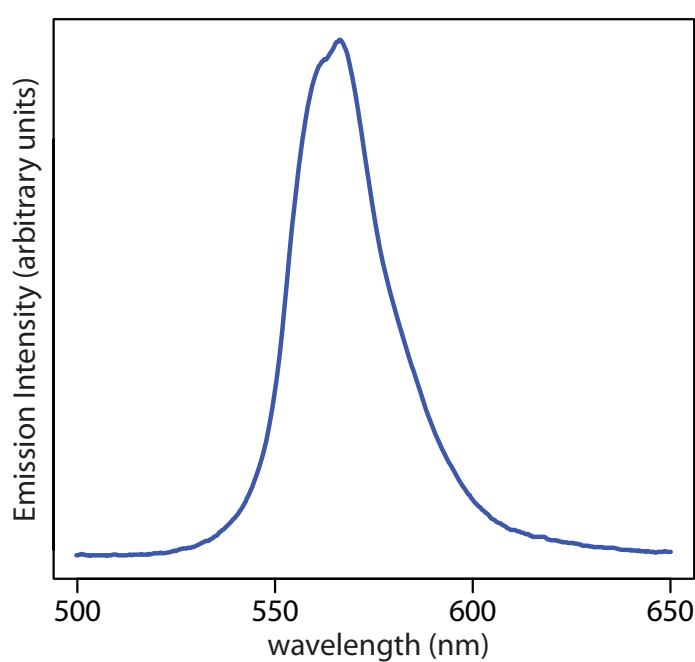


Figure 10.7 Spectrum showing the emission from a green LED, which provides continuous emission over a wavelength range of approximately 530–640 nm.

Table 10.3 Common Sources of Electromagnetic Radiation

Source	Wavelength Region	Useful for...
H ₂ and D ₂ lamp	continuum source from 160–380 nm	molecular absorption
tungsten lamp	continuum source from 320–2400 nm	molecular absorption
Xe arc lamp	continuum source from 200–1000 nm	molecular fluorescence
nernst glower	continuum source from 0.4–20 μm	molecular absorption
globar	continuum source from 1–40 μm	molecular absorption
nichrome wire	continuum source from 0.75–20 μm	molecular absorption
hollow cathode lamp	line source in UV/Visible	atomic absorption
Hg vapor lamp	line source in UV/Visible	molecular fluorescence
laser	line source in UV/Visible/IR	atomic and molecular absorption, fluorescence, and scattering

Chemical Sources of Energy Exothermic reactions also may serve as a source of energy. In chemiluminescence the analyte is raised to a higher-energy state by means of a chemical reaction, emitting characteristic radiation when it returns to a lower-energy state. When the chemical reaction results from a biological or enzymatic reaction, the emission of radiation is called bioluminescence. Commercially available “light sticks” and the flash of light from a firefly are examples of chemiluminescence and bioluminescence.

WAVELENGTH SELECTION

In Nessler’s original colorimetric method for ammonia, described at the beginning of the chapter, the sample and several standard solutions of ammonia are placed in separate tall, flat-bottomed tubes. As shown in [Figure 10.9](#), after adding the reagents and allowing the color to develop, the analyst evaluates the color by passing natural, ambient light through the bottom of the tubes and looking down through the solutions. By matching the sample’s color to that of a standard, the analyst is able to determine the concentration of ammonia in the sample.

In [Figure 10.9](#) every wavelength of light from the source passes through the sample. If there is only one absorbing species, this is not a problem. If two components in the sample absorb different wavelengths of light, then a quantitative analysis using Nessler’s original method becomes impossible. Ideally we want to select a wavelength that only the analyte absorbs. Unfortunately, we can not isolate a single wavelength of radiation from a continuum source. As shown in [Figure 10.10](#), a wavelength selector passes a narrow band of radiation characterized by a **NOMINAL WAVELENGTH**, an **EFFECTIVE BANDWIDTH** and a maximum throughput of radiation. The effective bandwidth is defined as the width of the radiation at half of its maximum throughput.

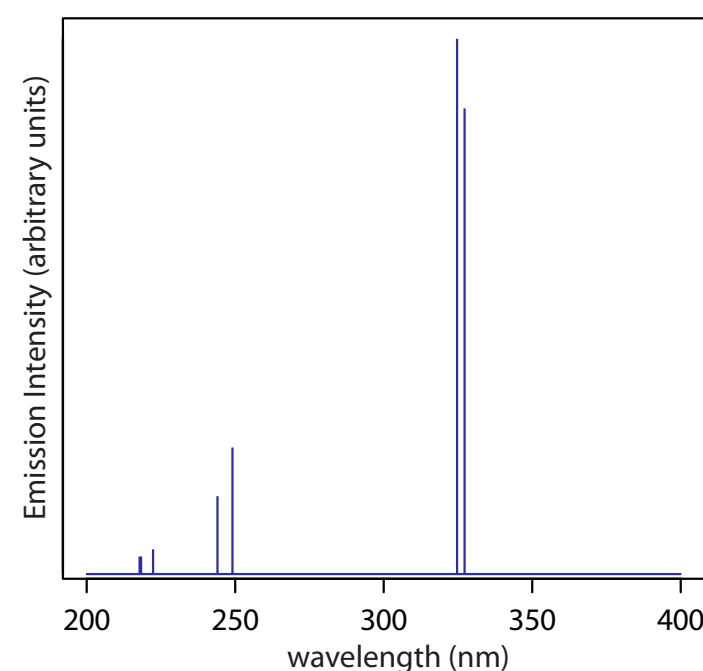


Figure 10.8 Emission spectrum from a Cu hollow cathode lamp. This spectrum consists of seven distinct emission lines (the first two differ by only 0.4 nm and are not resolved in this spectrum). Each emission line has a width of approximately 0.01 nm at ½ of its maximum intensity.

Figure 10.9 Nessler's original method for comparing the color of two solutions. Natural light passes upwards through the samples and standards and the analyst views the solutions by looking down toward the light source. The top view, shown on the right, is what the analyst sees. To determine the analyte's concentration, the analyst exchanges standards until the two colors match.

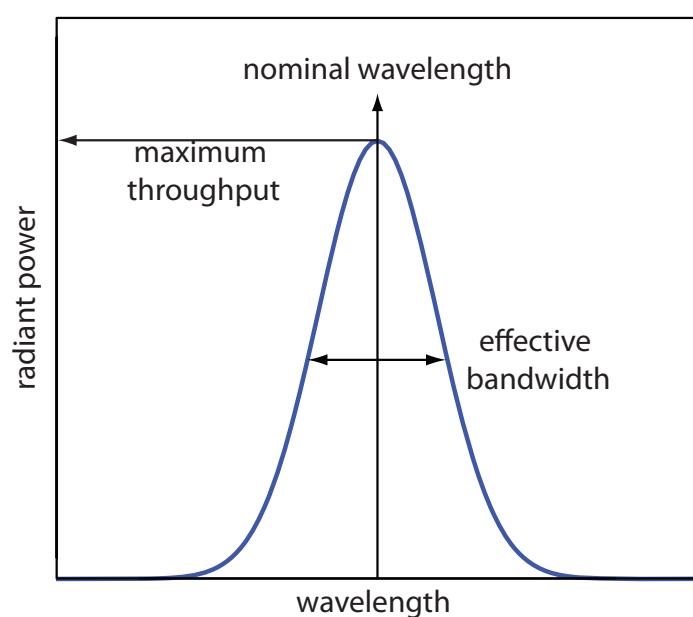
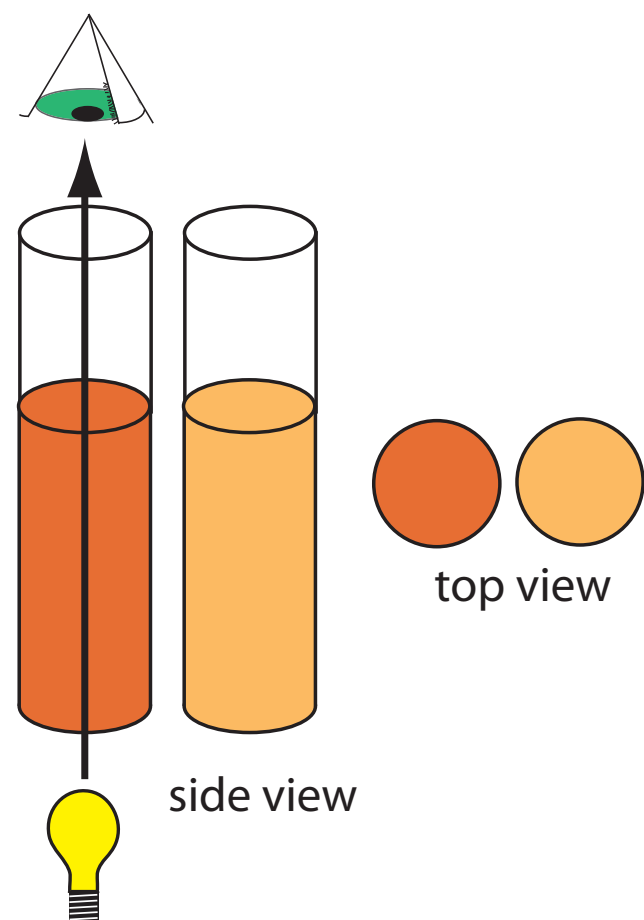


Figure 10.10 Radiation exiting a wavelength selector showing the band's nominal wavelength and its effective bandwidth.

The ideal wavelength selector has a high throughput of radiation and a narrow effective bandwidth. A high throughput is desirable because more photons pass through the wavelength selector, giving a stronger signal with less background noise. A narrow effective bandwidth provides a higher **RESOLUTION**, with spectral features separated by more than twice the effective bandwidth being resolved. As shown in [Figure 10.11](#), these two features of a wavelength selector generally are in opposition. Conditions favoring a higher throughput of radiation usually provide less resolution. Decreasing the effective bandwidth improves resolution, but at the cost of a noisier signal.⁴ For a qualitative analysis, resolution is usually more important than noise, and a smaller effective bandwidth is desirable. In a quantitative analysis less noise is usually desirable.

Wavelength Selection Using Filters. The simplest method for isolating a narrow band of radiation is to use an absorption or interference **FILTER**. Absorption filters work by selectively absorbing radiation from a narrow region of the electromagnetic spectrum. Interference filters use constructive and destructive interference to isolate a narrow range of wavelengths. A simple example of an absorption filter is a piece of colored glass. A purple filter, for example, removes the complementary color green from 500–560 nm. Commercially available absorption filters provide effective bandwidths of 30–250 nm, although the throughput may be only 10% of the source's emission intensity at the low end of this range. Interference filters are more expensive than absorption filters, but have narrower effective bandwidths, typically 10–20 nm, with maximum throughputs of at least 40%.

Wavelength Selection Using Monochromators. A filter has one significant limitation—because a filter has a fixed nominal wavelength, if you need to make measurements at two different wavelengths, then you need to use two

⁴ Jiang, S.; Parker, G. A. *Am. Lab.* **1981**, October, 38–43.

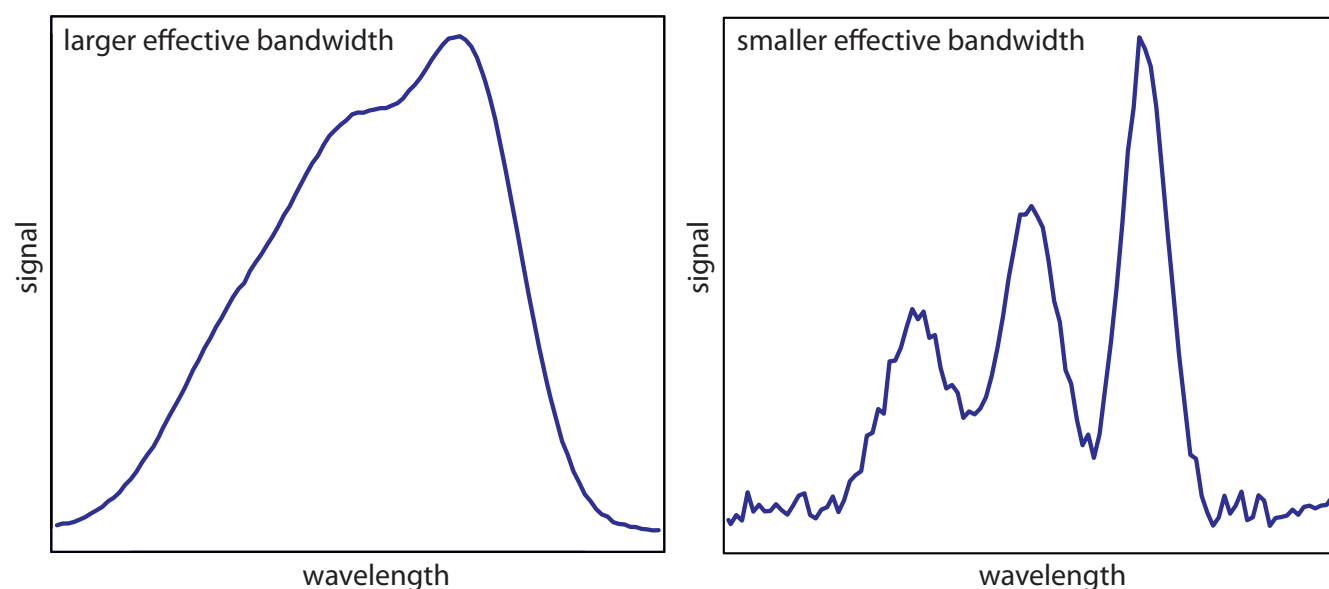


Figure 10.11 Example showing the effect of the wavelength selector's effective bandwidth on resolution and noise. The spectrum with the smaller effective bandwidth has a better resolution, allowing us to see the presence of three peaks, but at the expense of a noisier signal. The spectrum with the larger effective bandwidth has less noise, but at the expense of less resolution between the three peaks.

different filters. A **MONOCHROMATOR** is an alternative method for selecting a narrow band of radiation that also allows us to continuously adjust the band's nominal wavelength.

The construction of a typical monochromator is shown in Figure 10.12. Radiation from the source enters the monochromator through an entrance slit. The radiation is collected by a collimating mirror, which reflects a parallel beam of radiation to a diffraction grating. The diffraction grating is an optically reflecting surface with a large number of parallel grooves (see insert to Figure 10.12). The diffraction grating disperses the radiation and a second mirror focuses the radiation onto a planar surface containing an exit slit. In some monochromators a prism is used in place of the diffraction grating.

Radiation exits the monochromator and passes to the detector. As shown in Figure 10.12, a monochromator converts a **POLYCHROMATIC** source of

Polychromatic means many colored. Polychromatic radiation contains many different wavelengths of light.

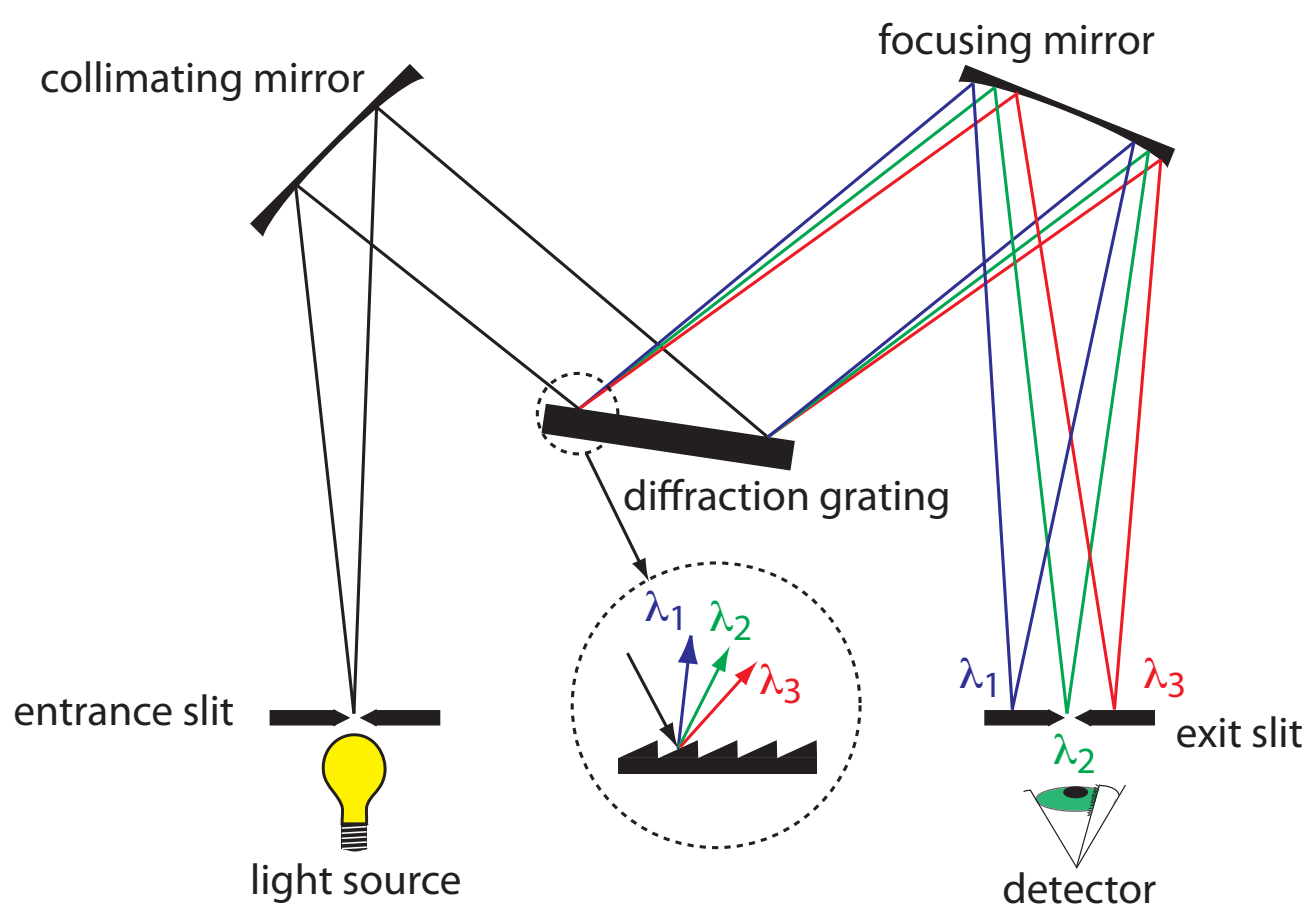


Figure 10.12 Schematic diagram of a monochromator that uses a diffraction grating to disperse the radiation.

Monochromatic means one color, or one wavelength. Although the light exiting a monochromator is not strictly of a single wavelength, its narrow effective bandwidth allows us to think of it as monochromatic.

radiation at the entrance slit to a **MONOCHROMATIC** source of finite effective bandwidth at the exit slit. The choice of which wavelength exits the monochromator is determined by rotating the diffraction grating. A narrower exit slit provides a smaller effective bandwidth and better resolution, but allows a smaller throughput of radiation.

Monochromators are classified as either fixed-wavelength or scanning. In a fixed-wavelength monochromator we select the wavelength by manually rotating the grating. Normally a fixed-wavelength monochromator is used for a quantitative analysis where measurements are made at one or two wavelengths. A scanning monochromator includes a drive mechanism that continuously rotates the grating, allowing successive wavelengths to exit from the monochromator. Scanning monochromators are used to acquire spectra, and, when operated in a fixed-wavelength mode, for a quantitative analysis.

Interferometers. An **INTERFEROMETER** provides an alternative approach for wavelength selection. Instead of filtering or dispersing the electromagnetic radiation, an interferometer allows source radiation of all wavelengths to reach the detector simultaneously (Figure 10.13). Radiation from the source is focused on a beam splitter that reflects half of the radiation to a fixed mirror and transmits the other half to a movable mirror. The radiation recombines at the beam splitter, where constructive and destructive interference determines, for each wavelength, the intensity of light reaching the detector. As the moving mirror changes position, the wavelengths of light experiencing maximum constructive interference and maximum destructive interference also changes. The signal at the detector shows intensity as a function of the moving mirror's position, expressed in units of distance or time. The result is called an **INTERFEROGRAM**, or a time domain spectrum.

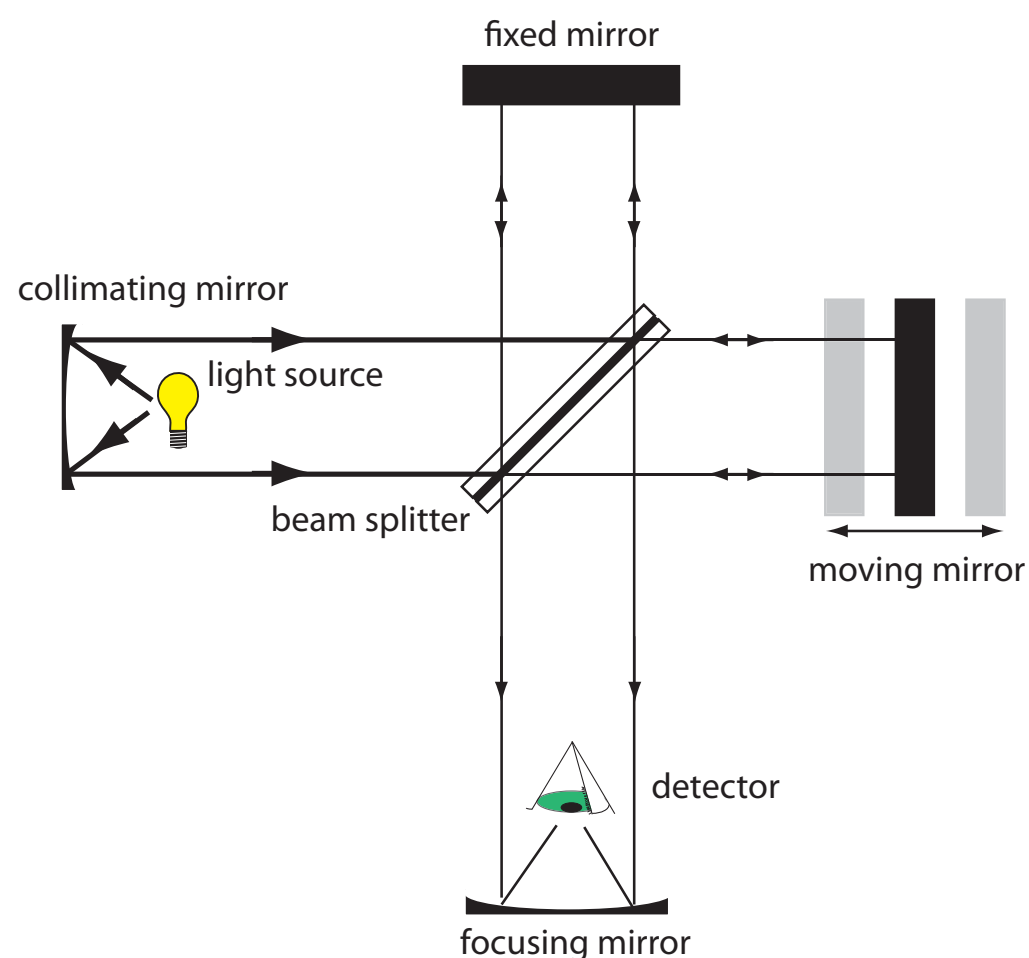


Figure 10.13 Schematic diagram of an interferometer.

The time domain spectrum is converted mathematically, by a process called a Fourier transform, to a spectrum (also called a frequency domain spectrum) showing intensity as a function of the radiation's energy.

In comparison to a monochromator, an interferometer has two significant advantages. The first advantage, which is termed **JACQUINOT'S ADVANTAGE**, is the higher throughput of source radiation. Because an interferometer does not use slits and has fewer optical components from which radiation can be scattered and lost, the throughput of radiation reaching the detector is 80–200× greater than that for a monochromator. The result is less noise. The second advantage, which is called **FELLGETT'S ADVANTAGE**, is a savings in the time needed to obtain a spectrum. Because the detector monitors all frequencies simultaneously, an entire spectrum takes approximately one second to record, as compared to 10–15 minutes with a scanning monochromator.

DETECTORS

In Nessler's original method for determining ammonia ([Figure 10.9](#)) the analyst's eye serves as the detector, matching the sample's color to that of a standard. The human eye, of course, has a poor range—responding only to visible light—nor is it particularly sensitive or accurate. Modern detectors use a sensitive **TRANSDUCER** to convert a signal consisting of photons into an easily measured electrical signal. Ideally the detector's signal, S , is a linear function of the electromagnetic radiation's power, P ,

$$S = kP + D$$

where k is the detector's sensitivity, and D is the detector's **DARK CURRENT**, or the background current when we prevent the source's radiation from reaching the detector.

There are two broad classes of spectroscopic transducers: thermal transducers and photon transducers. Table 10.4 provides several representative examples of each class of transducers.

The mathematical details of the Fourier transform are beyond the level of this textbook. You can consult the chapter's additional resources for additional information.

Transducer is a general term that refers to any device that converts a chemical or physical property into an easily measured electrical signal. The retina in your eye, for example, is a transducer that converts photons into an electrical nerve impulse.

Table 10.4 Examples of Transducers for Spectroscopy

Transducer	Class	Wavelength Range	Output Signal
phototube	photon	200–1000 nm	current
photomultiplier	photon	110–1000 nm	current
Si photodiode	photon	250–1100 nm	current
photoconductor	photon	750–6000 nm	change in resistance
photovoltaic cell	photon	400–5000 nm	current or voltage
thermocouple	thermal	0.8–40 μm	voltage
thermistor	thermal	0.8–40 μm	change in resistance
pneumatic	thermal	0.8–1000 μm	membrane displacement
pyroelectric	thermal	0.3–1000 μm	current

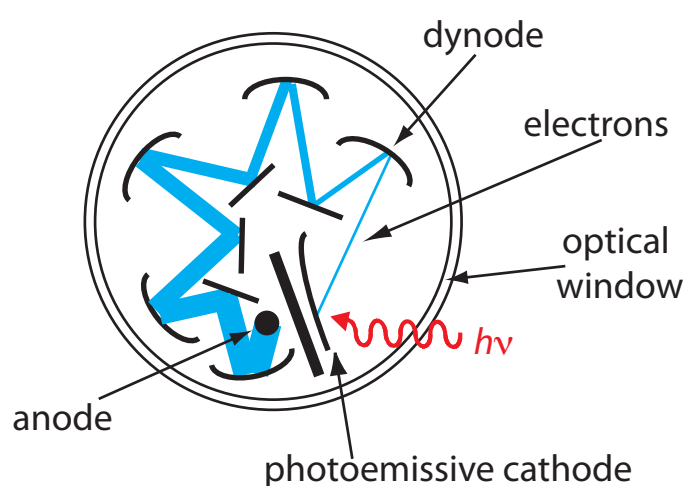


Figure 10.14 Schematic of a photomultiplier. A **photon** strikes the photoemissive cathode producing **electrons**, which accelerate toward a positively charged dynode. Collision of these electrons with the dynode generates additional electrons, which accelerate toward the next dynode. A total of 10^6 – 10^7 electrons per photon eventually reach the anode, generating an electrical current.

Photon Transducers. Phototubes and photomultipliers contain a photosensitive surface that absorbs radiation in the ultraviolet, visible, or near IR, producing an electrical current proportional to the number of photons reaching the transducer (Figure 10.14). Other photon detectors use a semiconductor as the photosensitive surface. When the semiconductor absorbs photons, valence electrons move to the semiconductor's conduction band, producing a measurable current. One advantage of the Si photodiode is that it is easy to miniaturize. Groups of photodiodes may be gathered together in a linear array containing from 64–4096 individual photodiodes. With a width of $25\ \mu\text{m}$ per diode, for example, a linear array of 2048 photodiodes requires only 51.2 mm of linear space. By placing a **PHOTODIODE ARRAY** along the monochromator's focal plane, it is possible to monitor simultaneously an entire range of wavelengths.

Thermal Transducers. Infrared photons do not have enough energy to produce a measurable current with a photon transducer. A thermal transducer, therefore, is used for infrared spectroscopy. The absorption of infrared photons by a thermal transducer increases its temperature, changing one or more of its characteristic properties. A pneumatic transducer, for example, is a small tube of xenon gas with an IR transparent window at one end and a flexible membrane at the other end. Photons enter the tube and are absorbed by a blackened surface, increasing the temperature of the gas. As the temperature inside the tube fluctuates, the gas expands and contracts and the flexible membrane moves in and out. Monitoring the membrane's displacement produces an electrical signal.

Signal Processors

A transducer's electrical signal is sent to a **SIGNAL PROCESSOR** where it is displayed in a form that is more convenient for the analyst. Examples of signal processors include analog or digital meters, recorders, and computers equipped with digital acquisition boards. A signal processor also is used to calibrate the detector's response, to amplify the transducer's signal, to remove noise by filtering, or to mathematically transform the signal.

10B Spectroscopy Based on Absorption

In absorption spectroscopy a beam of electromagnetic radiation passes through a sample. Much of the radiation passes through the sample without a loss in intensity. At selected wavelengths, however, the radiation's intensity is attenuated. This process of attenuation is called absorption.

10B.1 Absorbance Spectra

There are two general requirements for an analyte's absorption of electromagnetic radiation. First, there must be a mechanism by which the radiation's electric field or magnetic field interacts with the analyte. For ultra-

If the retina in your eye is a transducer, then your brain is a signal processor.

violet and visible radiation, absorption of a photon changes the energy of the analyte's valence electrons. A bond's vibrational energy is altered by the absorption of infrared radiation.

The second requirement is that the photon's energy, $h\nu$, must exactly equal the difference in energy, ΔE , between two of the analyte's quantized energy states. [Figure 10.4](#) shows a simplified view of a photon's absorption, which is useful because it emphasizes that the photon's energy must match the difference in energy between a lower-energy state and a higher-energy state. What is missing, however, is information about what types of energy states are involved, which transitions between energy states are likely to occur, and the appearance of the resulting spectrum.

We can use the energy level diagram in [Figure 10.15](#) to explain an absorbance spectrum. The lines labeled E_0 and E_1 represent the analyte's ground (lowest) electronic state and its first electronic excited state. Superimposed on each electronic energy level is a series of lines representing vibrational energy levels.

INFRARED SPECTRA FOR MOLECULES AND POLYATOMIC IONS

The energy of infrared radiation produces a change in a molecule's or a polyatomic ion's vibrational energy, but is not sufficient to effect a change in its electronic energy. As shown in [Figure 10.15](#), vibrational energy levels are quantized; that is, a molecule may have only certain, discrete vibrational energies. The energy for an allowed vibrational mode, E_ν , is

$$E_\nu = \nu + \frac{1}{2} h\nu_0$$

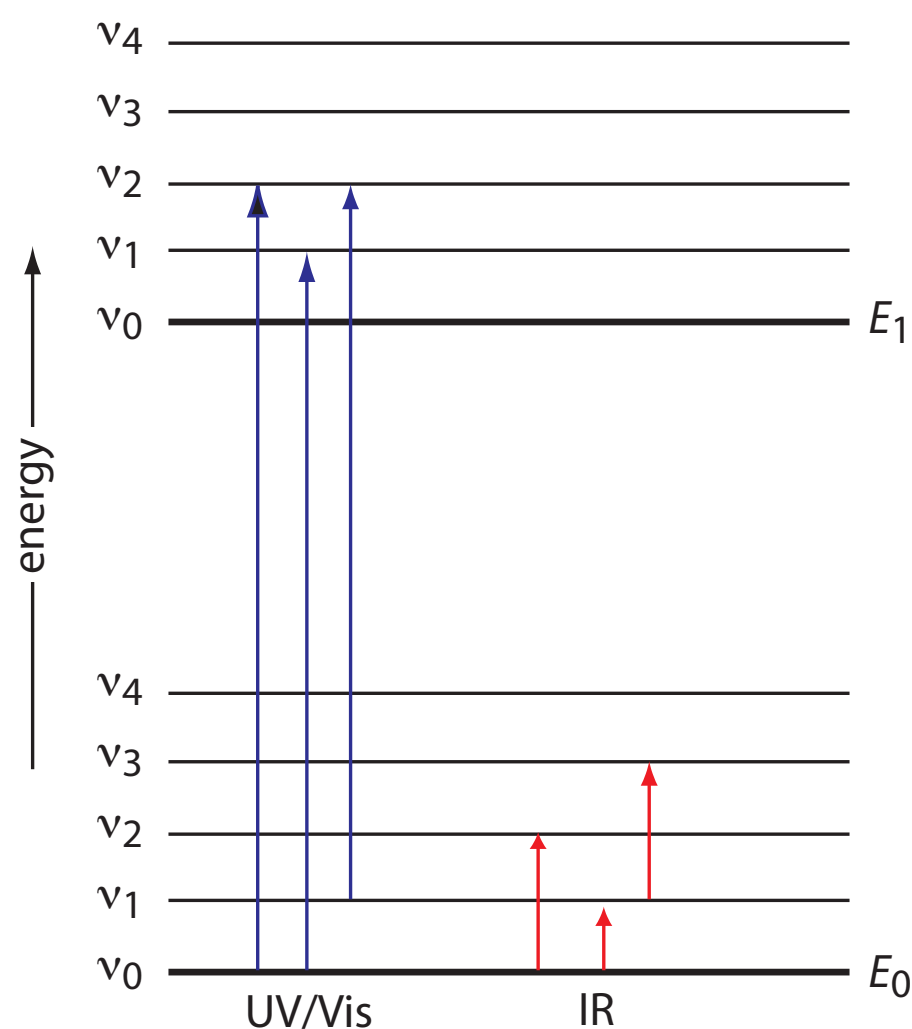


Figure 10.15 Diagram showing two electronic energy levels (E_0 and E_1), each with five vibrational energy levels (ν_0 – ν_4). Absorption of ultraviolet and visible radiation leads to a change in the analyte's electronic energy levels and, possibly, a change in vibrational energy as well. A change in vibrational energy without a change in electronic energy levels occurs with the absorption of infrared radiation.

[Figure 10.3](#) provides a list of the types of atomic and molecular transitions associated with different types of electromagnetic radiation.

Why does a non-linear molecule have $3N-6$ vibrational modes? Consider a molecule of methane, CH_4 . Each of the five atoms in methane can move in one of three directions (x , y , and z) for a total of $5 \times 3 = 15$ different ways in which the molecule can move. A molecule can move in three ways: it can move from one place to another, what we call translational motion; it can rotate around an axis; and its bonds can stretch and bend, what we call vibrational motion.

Because the entire molecule can move in the x , y , and z directions, three of methane's 15 different motions are translational. In addition, the molecule can rotate about its x , y , and z axes, accounting for three additional forms of motion. This leaves $15 - 3 - 3 = 9$ vibrational modes.

A linear molecule, such as CO_2 , has $3N-5$ vibrational modes because it can rotate around only two axes.

where ν is the vibrational quantum number, which has values of 0, 1, 2, ..., and ν_0 is the bond's fundamental vibrational frequency. The value of ν_0 , which is determined by the bond's strength and by the mass at each end of the bond, is a characteristic property of a bond. For example, a carbon-carbon single bond ($\text{C}-\text{C}$) absorbs infrared radiation at a lower energy than a carbon-carbon double bond ($\text{C}=\text{C}$) because a single bond is weaker than a double bond.

At room temperature most molecules are in their ground vibrational state ($\nu = 0$). A transition from the ground vibrational state to the first vibrational excited state ($\nu = 1$) requires absorption of a photon with an energy of $h\nu_0$. Transitions in which $\Delta\nu$ is ± 1 give rise to the fundamental absorption lines. Weaker absorption lines, called overtones, result from transitions in which $\Delta\nu$ is ± 2 or ± 3 . The number of possible normal vibrational modes for a linear molecule is $3N-5$, and for a non-linear molecule is $3N-6$, where N is the number of atoms in the molecule. Not surprisingly, infrared spectra often show a considerable number of absorption bands. Even a relatively simple molecule, such as ethanol ($\text{C}_2\text{H}_6\text{O}$), for example, has $3 \times 9 - 6$, or 21 possible normal modes of vibration, although not all of these vibrational modes give rise to an absorption. The IR spectrum for ethanol is shown in Figure 10.16.

UV/Vis SPECTRA FOR MOLECULES AND IONS

The valence electrons in organic molecules and polyatomic ions, such as CO_3^{2-} , occupy quantized sigma bonding, σ , pi bonding, π , and non-bonding, n , molecular orbitals (MOs). Unoccupied sigma antibonding, σ^* , and pi antibonding, π^* , molecular orbitals are slightly higher in energy. Because the difference in energy between the highest-energy occupied MOs and

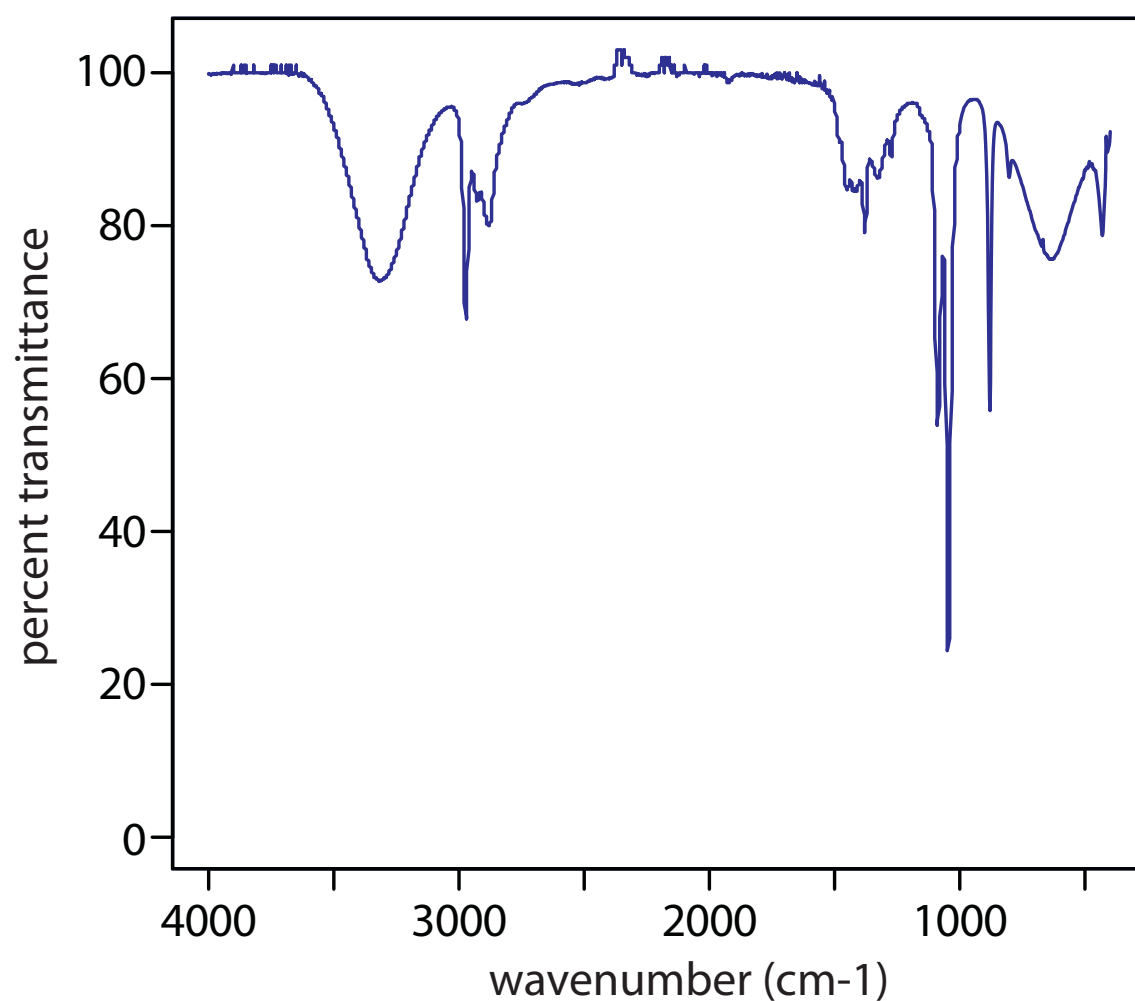


Figure 10.16 Infrared spectrum of ethanol.

Table 10.5 Electronic Transitions Involving n , σ , and π Molecular Orbitals

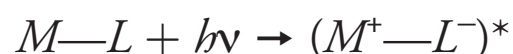
Transition	Wavelength Range	Examples
$\sigma \rightarrow \sigma^*$	<200 nm	C–C, C–H
$n \rightarrow \sigma^*$	160–260 nm	H ₂ O, CH ₃ OH, CH ₃ Cl
$\pi \rightarrow \pi^*$	200–500 nm	C=C, C=O, C=N, C≡C
$n \rightarrow \pi^*$	250–600 nm	C=O, C=N, N=N, N=O

the lowest-energy unoccupied MOs corresponds to ultraviolet and visible radiation, absorption of a photon is possible.

Four types of transitions between quantized energy levels account for most molecular UV/Vis spectra. Table 10.5 lists the approximate wavelength ranges for these transitions, as well as a partial list of bonds, functional groups, or molecules responsible for these transitions. Of these transitions, the most important are $n \rightarrow \pi^*$ and $\pi \rightarrow \pi^*$ because they involve important functional groups that are characteristic of many analytes and because the wavelengths are easily accessible. The bonds and functional groups that give rise to the absorption of ultraviolet and visible radiation are called **CHROMOPHORES**.

Many transition metal ions, such as Cu²⁺ and Co²⁺, form colorful solutions because the metal ion absorbs visible light. The transitions giving rise to this absorption are valence electrons in the metal ion's d -orbitals. For a free metal ion, the five d -orbitals are of equal energy. In the presence of a complexing ligand or solvent molecule, however, the d -orbitals split into two or more groups that differ in energy. For example, in an octahedral complex of Cu(H₂O)₆²⁺ the six water molecules perturb the d -orbitals into two groups, as shown in Figure 10.17. The resulting d – d transitions for transition metal ions are relatively weak.

A more important source of UV/Vis absorption for inorganic metal–ligand complexes is charge transfer, in which absorption of a photon produces an excited state in which there is transfer of an electron from the metal, M , to the ligand, L .



Charge-transfer absorption is important because it produces very large absorbances. One important example of a charge-transfer complex is that of *o*-phenanthroline with Fe²⁺, the UV/Vis spectrum for which is shown in Figure 10.18. Charge-transfer absorption in which an electron moves from the ligand to the metal also is possible.

Comparing the IR spectrum in Figure 10.16 to the UV/Vis spectrum in Figure 10.18 shows us that UV/Vis absorption bands are often significantly broader than those for IR absorption. We can use Figure 10.15 to explain

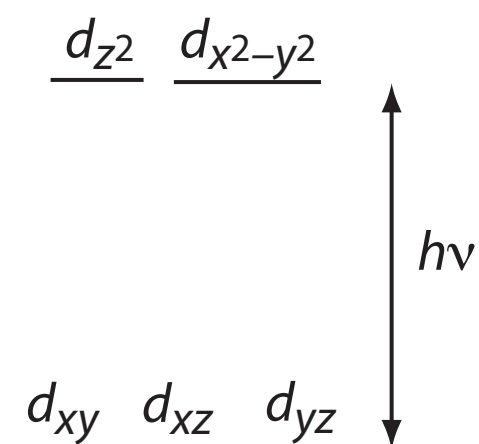
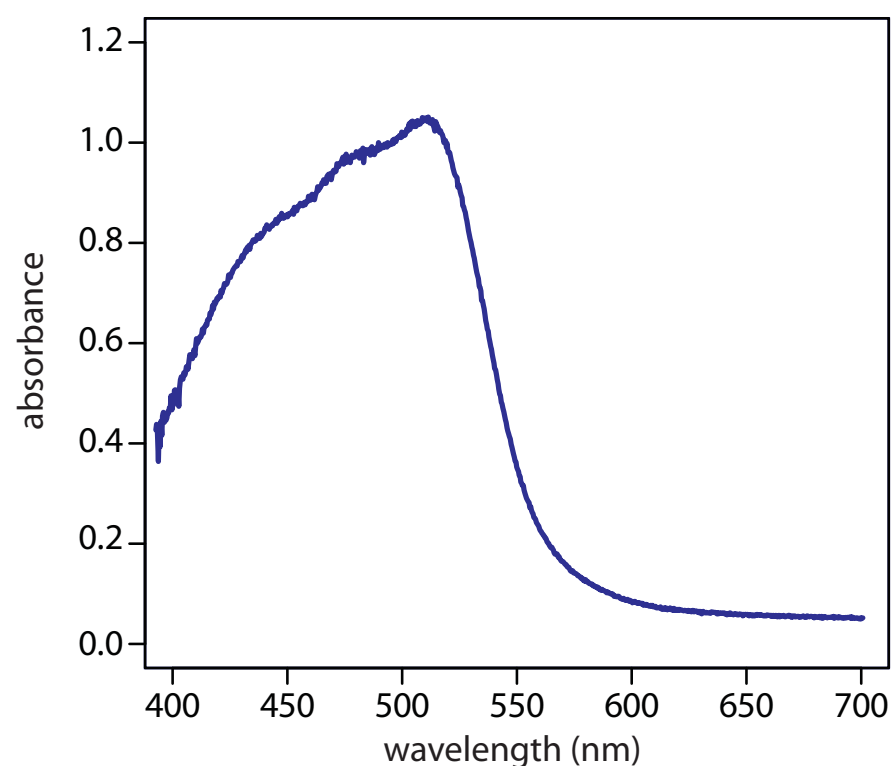


Figure 10.17 Splitting of the d -orbitals in an octahedral field.

Why is a larger absorbance desirable? An analytical method is more sensitive if a smaller concentration of analyte gives a larger signal.

Figure 10.18 UV/Vis spectrum for the metal–ligand complex $\text{Fe}(\text{phen})_3^{2+}$, where phen is the ligand *o*-phenanthroline.



why this is true. When a species absorbs UV/Vis radiation, the transition between electronic energy levels may also include a transition between vibrational energy levels. The result is a number of closely spaced absorption bands that merge together to form a single broad absorption band.

UV/Vis SPECTRA FOR ATOMS

The valence shell energy level diagram in Figure 10.19 might strike you as odd because it shows that the $3p$ orbitals are split into two groups of slightly different energy. The reasons for this splitting are unimportant in the context of our treatment of atomic absorption. For further information about the reasons for this splitting, consult the chapter's [additional resources](#).

The energy of ultraviolet and visible electromagnetic radiation is sufficient to cause a change in an atom's valence electron configuration. Sodium, for example, has a single valence electron in its $3s$ atomic orbital. As shown in Figure 10.19, unoccupied, higher energy atomic orbitals also exist.

Absorption of a photon is accompanied by the excitation of an electron from a lower-energy atomic orbital to an orbital of higher energy. Not all possible transitions between atomic orbitals are allowed. For sodium the only allowed transitions are those in which there is a change of ± 1 in the orbital quantum number (l); thus transitions from $s \rightarrow p$ orbitals are allowed, and transitions from $s \rightarrow d$ orbitals are forbidden.

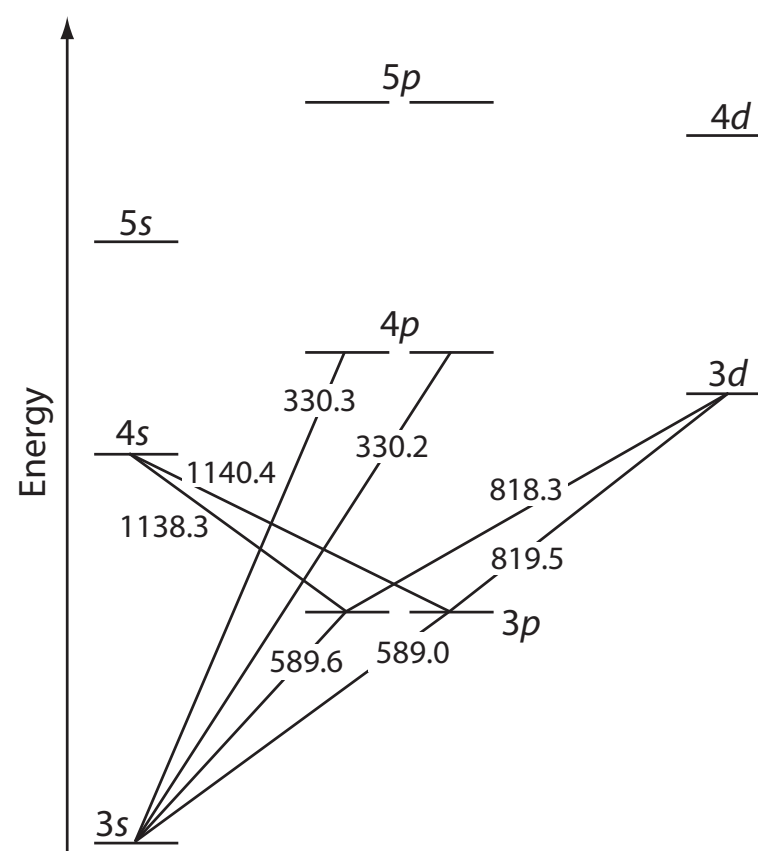


Figure 10.19 Valence shell energy level diagram for sodium. The wavelengths (in wavenumbers) corresponding to several transitions are shown.

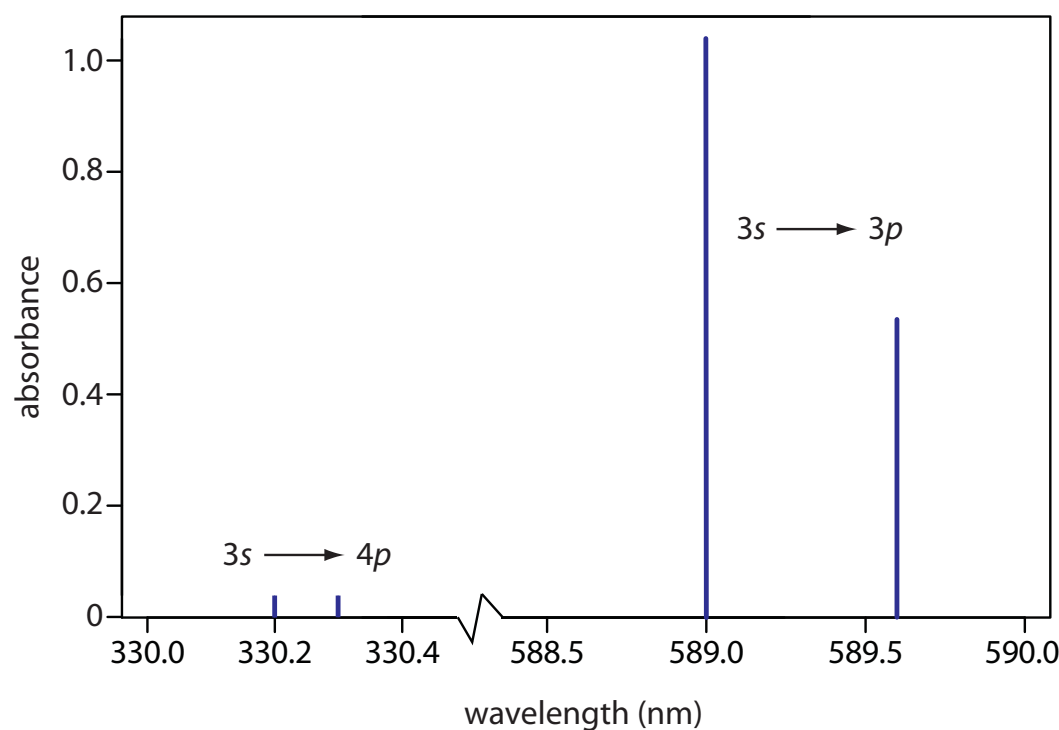


Figure 10.20 Atomic absorption spectrum for sodium. Note that the scale on the x -axis includes a break.

The atomic absorption spectrum for Na is shown in Figure 10.20, and is typical of that found for most atoms. The most obvious feature of this spectrum is that it consists of a small number of discrete absorption lines corresponding to transitions between the ground state (the $3s$ atomic orbital) and the $3p$ and $4p$ atomic orbitals. Absorption from excited states, such as the $3p \rightarrow 4s$ and the $3p \rightarrow 3d$ transitions included in Figure 10.19, are too weak to detect. Because an excited state's lifetime is short—typically an excited state atom takes 10^{-7} to 10^{-8} s to return to a lower energy state—an atom in the excited state is likely to return to the ground state before it has an opportunity to absorb a photon.

Another feature of the atomic absorption spectrum in Figure 10.20 is the narrow width of the absorption lines, which is a consequence of the fixed difference in energy between the ground and excited states. Natural line widths for atomic absorption, which are governed by the uncertainty principle, are approximately 10^{-5} nm. Other contributions to broadening increase this line width to approximately 10^{-3} nm.

10B.2 Transmittance and Absorbance

As light passes through a sample, its power decreases as some of it is absorbed. This attenuation of radiation is described quantitatively by two separate, but related terms: transmittance and absorbance. As shown in Figure 10.21a, **TRANSMITTANCE** is the ratio of the source radiation's power exiting the sample, P_T , to that incident on the sample, P_0 .

$$T = \frac{P_T}{P_0} \quad 10.1$$

Multiplying the transmittance by 100 gives the percent transmittance, $\%T$, which varies between 100% (no absorption) and 0% (complete absorption). All methods of detecting photons—including the human eye and modern photoelectric transducers—measure the transmittance of electromagnetic radiation.

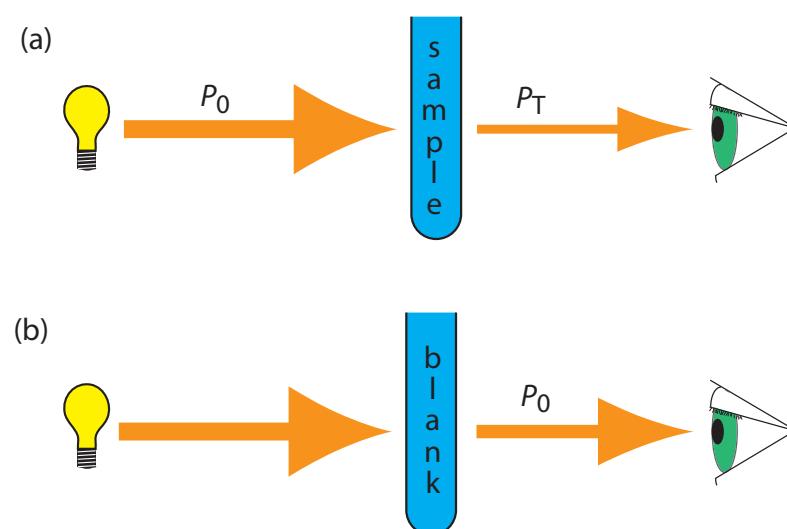


Figure 10.21 (a) Schematic diagram showing the attenuation of radiation passing through a sample; P_0 is the radiant power from the source and P_T is the radiant power transmitted by the sample. (b) Schematic diagram showing how we redefine P_0 as the radiant power transmitted by the blank. Redefining P_0 in this way corrects the transmittance in (a) for the loss of radiation due to scattering, reflection, or absorption by the sample's container and absorption by the sample's matrix.

[Equation 10.1](#) does not distinguish between different mechanisms that prevent a photon emitted by the source from reaching the detector. In addition to absorption by the analyte, several additional phenomena contribute to the attenuation of radiation, including reflection and absorption by the sample's container, absorption by other components in the sample's matrix, and the scattering of radiation. To compensate for this loss of the radiation's power, we use a method blank. As shown in Figure 10.21b, we redefine P_0 as the power exiting the method blank.

An alternative method for expressing the attenuation of electromagnetic radiation is absorbance, A , which we define as

$$A = -\log T = -\log \frac{P_T}{P_0} \quad 10.2$$

Absorbance is the more common unit for expressing the attenuation of radiation because it is a linear function of the analyte's concentration.

Example 10.3

A sample has a percent transmittance of 50%. What is its absorbance?

SOLUTION

A percent transmittance of 50.0% is the same as a transmittance of 0.500. Substituting into equation 10.2 gives

$$A = -\log T = -\log(0.500) = 0.301$$

Practice Exercise 10.3

What is the % T for a sample if its absorbance is 1.27?

Click [here](#) to review your answer to this exercise.

We will show that this is true in [Section 10B.3](#).

[Equation 10.1](#) has an important consequence for atomic absorption. As we saw in [Figure 10.20](#), atomic absorption lines are very narrow. Even with a high quality monochromator, the effective bandwidth for a continuum source is $100\text{--}1000\times$ greater than the width of an atomic absorption line. As a result, little of the radiation from a continuum source is absorbed ($P_0 \approx P_T$), and the measured absorbance is effectively zero. For this reason, atomic absorption requires a line source instead of a continuum source.

10B.3 Absorbance and Concentration: Beer's Law

When monochromatic electromagnetic radiation passes through an infinitesimally thin layer of sample of thickness dx , it experiences a decrease in its power of dP (Figure 10.22). The fractional decrease in power is proportional to the sample's thickness and the analyte's concentration, C ; thus

$$-\frac{dP}{P} = \alpha C dx \quad 10.3$$

where P is the power incident on the thin layer of sample, and α is a proportionality constant. Integrating the left side of equation 10.3 over the entire sample

$$-\int_{P_0}^{P_T} \frac{dP}{P} = \alpha C \int_0^b dx$$

$$\ln \frac{P_0}{P_T} = \alpha b C$$

converting from \ln to \log , and substituting in [equation 10.2](#), gives

$$A = abC \quad 10.4$$

where a is the analyte's **ABSORPTIVITY** with units of $\text{cm}^{-1} \text{conc}^{-1}$. If we express the concentration using molarity, then we replace a with the **MOLAR ABSORPTIVITY**, ϵ , which has units of $\text{cm}^{-1} \text{M}^{-1}$.

$$A = \epsilon b C \quad 10.5$$

The absorptivity and molar absorptivity are proportional to the probability that the analyte absorbs a photon of a given energy. As a result, values for both a and ϵ depend on the wavelength of the absorbed photon.

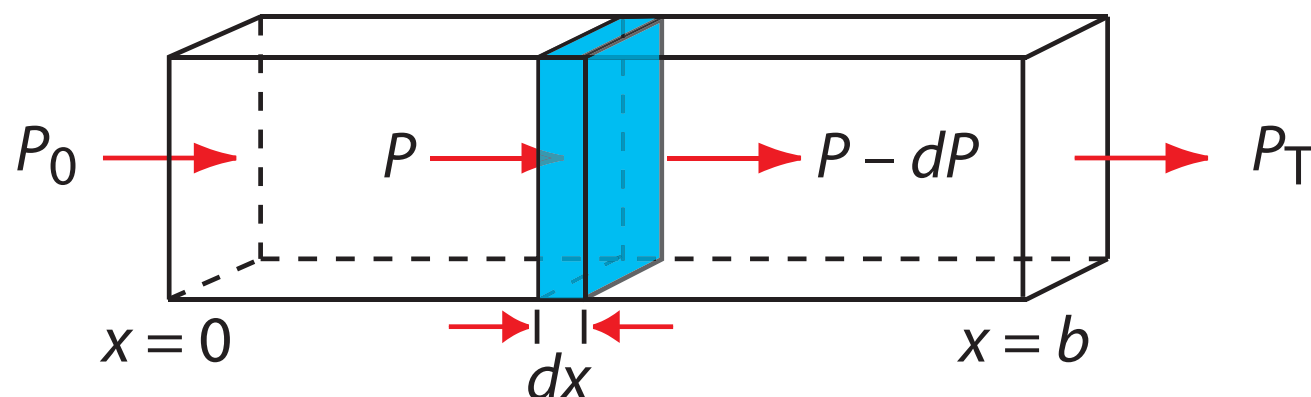


Figure 10.22 Factors used in deriving the Beer-Lambert law.

Example 10.4

A 5.00×10^{-4} M solution of an analyte is placed in a sample cell with a pathlength of 1.00 cm. When measured at a wavelength of 490 nm, the solution's absorbance is 0.338. What is the analyte's molar absorptivity at this wavelength?

SOLUTION

Solving [equation 10.5](#) for ϵ and making appropriate substitutions gives

$$\epsilon = \frac{A}{bC} = \frac{0.338}{(1.00 \text{ cm})(5.00 \times 10^{-4} \text{ M})} = 676 \text{ cm}^{-1} \text{ M}^{-1}$$

Practice Exercise 10.4

A solution of the analyte from Example 10.4 has an absorbance of 0.228 in a 1.00-cm sample cell. What is the analyte's concentration?

Click [here](#) to review your answer to this exercise.

[Equation 10.4](#) and [equation 10.5](#), which establish the linear relationship between absorbance and concentration, are known as the Beer-Lambert law, or more commonly, as **BEER'S LAW**. Calibration curves based on Beer's law are common in quantitative analyses.

10B.4 Beer's Law and Multicomponent Samples

We can extend Beer's law to a sample containing several absorbing components. If there are no interactions between the components, the individual absorbances, A_i , are additive. For a two-component mixture of analyte's X and Y, the total absorbance, A_{tot} , is

$$A_{\text{tot}} = A_X + A_Y = \epsilon_X b C_X + \epsilon_Y b C_Y$$

Generalizing, the absorbance for a mixture of n components, A_{mix} , is

$$A_{\text{mix}} = \sum_{i=1}^n A_i = \sum_{i=1}^n \epsilon_i b C_i \quad 10.6$$

10B.5 Limitations to Beer's Law

Beer's law suggests that a calibration curve is a straight line with a y -intercept of zero and a slope of ab or ϵb . In many cases a calibration curve deviates from this ideal behavior (Figure 10.23). Deviations from linearity are divided into three categories: fundamental, chemical, and instrumental.

FUNDAMENTAL LIMITATIONS TO BEER'S LAW

Beer's law is a limiting law that is valid only for low concentrations of analyte. There are two contributions to this fundamental limitation to Beer's law. At

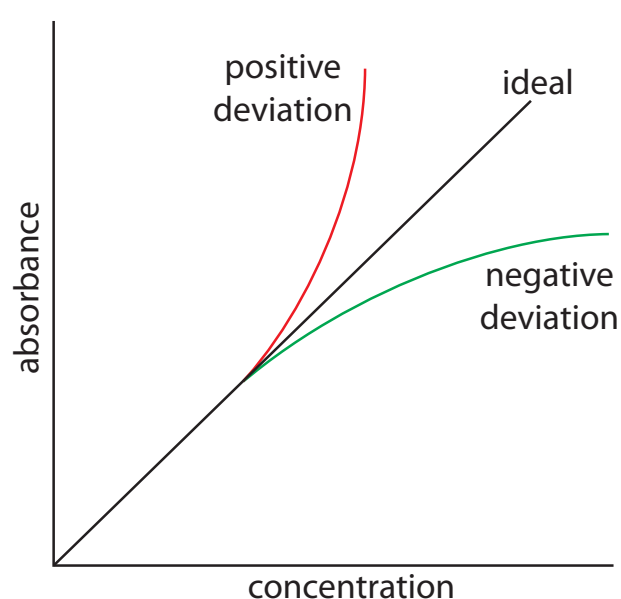


Figure 10.23 Calibration curves showing positive and negative deviations from the ideal Beer's law calibration curve, which is a straight line.

higher concentrations the individual particles of analyte no longer behave independently of each other. The resulting interaction between particles of analyte may change the analyte's absorptivity. A second contribution is that the analyte's absorptivity depends on the sample's refractive index. Because the refractive index varies with the analyte's concentration, the values of a and ϵ may change. For sufficiently low concentrations of analyte, the refractive index is essentially constant and the calibration curve is linear.

CHEMICAL LIMITATIONS TO BEER'S LAW

A chemical deviation from Beer's law may occur if the analyte is involved in an equilibrium reaction. Consider, as an example, an analysis for the weak acid, HA. To construct a Beer's law calibration curve we prepare a series of standards—each containing a known total concentration of HA—and measure each standard's absorbance at the same wavelength. Because HA is a weak acid, it is in equilibrium with its conjugate weak base, A^- .



If both HA and A^- absorb at the chosen wavelength, then Beer's law is

$$A = \epsilon_{\text{HA}} bC_{\text{HA}} + \epsilon_{\text{A}} bC_{\text{A}} \quad 10.7$$

where C_{HA} and C_{A} are the equilibrium concentrations of HA and A^- . Because the weak acid's total concentration, C_{total} , is

$$C_{\text{total}} = C_{\text{HA}} + C_{\text{A}}$$

the concentrations of HA and A^- can be written as

$$C_{\text{HA}} = \alpha_{\text{HA}} C_{\text{total}} \quad 10.8$$

$$C_{\text{A}} = (1 - \alpha_{\text{HA}}) C_{\text{total}} \quad 10.9$$

where α_{HA} is the fraction of weak acid present as HA. Substituting equation 10.8 and equation 10.9 into equation 10.7 and rearranging, gives

$$A = (\epsilon_{\text{HA}} \alpha_{\text{HA}} + \epsilon_{\text{A}} - \epsilon_{\text{A}} \alpha_{\text{HA}}) bC_{\text{total}} \quad 10.10$$

To obtain a linear Beer's law calibration curve, one of two conditions must be met. If α_{HA} and α_{A} have the same value at the chosen wavelength, then equation 10.10 simplifies to

$$A = \epsilon_{\text{A}} bC_{\text{total}}$$

Alternatively, if α_{HA} has the same value for all standards, then each term within the parentheses of equation 10.10 is constant—which we replace with k —and a linear calibration curve is obtained at any wavelength.

$$A = kbC_{\text{total}}$$

For a monoprotic weak acid, the equation for α_{HA} is

$$\alpha_{\text{HA}} = \frac{[\text{H}_3\text{O}^+]}{[\text{H}_3\text{O}^+] + K_a}$$

Problem 10.6 in the end of chapter problems asks you to explore this chemical limitation to Beer's law.

Problem 10.7 in the end of chapter problems ask you to explore the effect of polychromatic radiation on the linearity of Beer's law.

Another reason for measuring absorbance at the top of an absorbance peak is that it provides for a more sensitive analysis. Note that the green calibration curve in Figure 10.24 has a steeper slope—a greater sensitivity—than the red calibration curve.

Because HA is a weak acid, the value of α_{HA} varies with pH. To hold α_{HA} constant we buffer each standard to the same pH. Depending on the relative values of α_{HA} and α_{A} , the calibration curve has a positive or a negative deviation from Beer's law if we do not buffer the standards to the same pH.

INSTRUMENTAL LIMITATIONS TO BEER'S LAW

There are two principal instrumental limitations to Beer's law. The first limitation is that Beer's law assumes that the radiation reaching the sample is of a single wavelength—that is, that the radiation is purely monochromatic. As shown in [Figure 10.10](#), however, even the best wavelength selector passes radiation with a small, but finite effective bandwidth. Polychromatic radiation always gives a negative deviation from Beer's law, but the effect is smaller if the value of ϵ is essentially constant over the wavelength range passed by the wavelength selector. For this reason, as shown in [Figure 10.24](#), it is better to make absorbance measurements at the top of a broad absorption peak. In addition, the deviation from Beer's law is less serious if the source's effective bandwidth is less than one-tenth of the natural bandwidth of the absorbing species.⁵ When measurements must be made on a slope, linearity is improved by using a narrower effective bandwidth.

STRAY RADIATION is the second contribution to instrumental deviations from Beer's law. Stray radiation arises from imperfections in the wavelength selector that allow light to enter the instrument and reach the detector without passing through the sample. Stray radiation adds an additional contribution, P_{stray} , to the radiant power reaching the detector; thus

$$A = -\log \frac{P_{\text{T}} + P_{\text{stray}}}{P_0 + P_{\text{stray}}}$$

For a small concentration of analyte, P_{stray} is significantly smaller than P_0 and P_{T} , and the absorbance is unaffected by the stray radiation. At a higher concentration of analyte, however, less light passes through the sample and

5 (a) Strong, F. C., III *Anal. Chem.* **1984**, *56*, 16A–34A; Gilbert, D. D. *J. Chem. Educ.* **1991**, *68*, A278–A281.

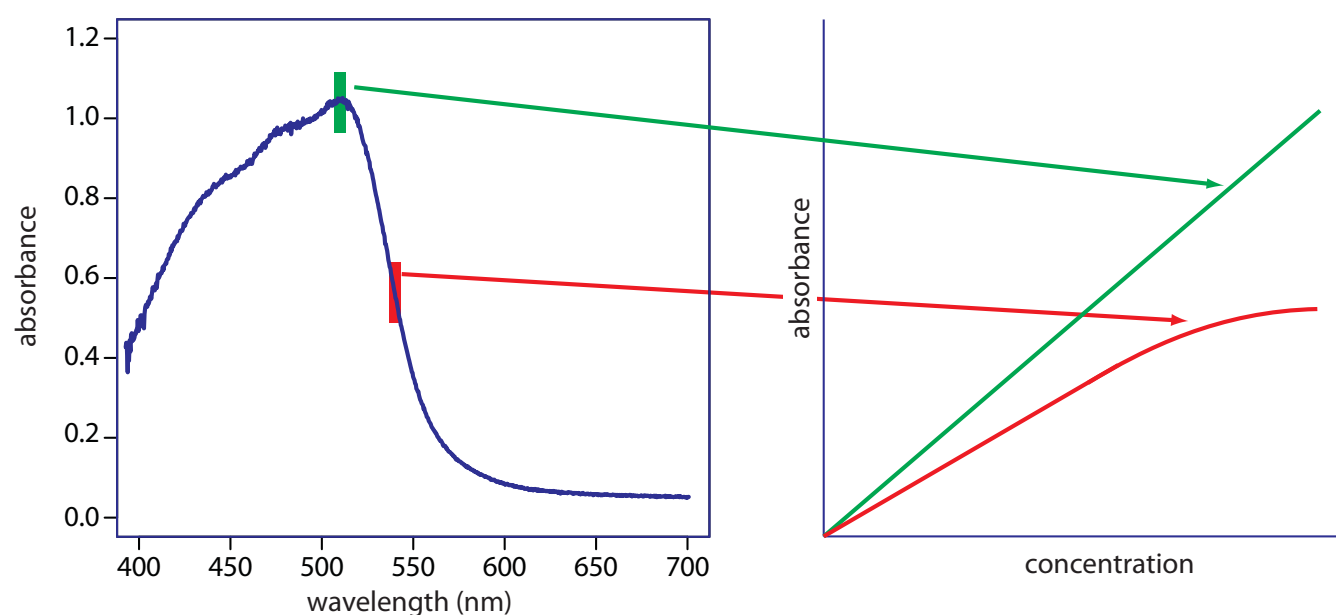


Figure 10.24 Effect of the choice of wavelength on the linearity of a Beer's law calibration curve.

P_T and P_{stray} may be similar in magnitude. This results in an absorbance that is smaller than expected, and a negative deviation from Beer's law.

[Problem 10.8](#) in the end of chapter problems ask you to explore the effect of stray radiation on the linearity of Beer's law.

10C UV/Vis and IR Spectroscopy

In [Figure 10.9](#) we examined Nessler's original method for matching the color of a sample to the color of a standard. Matching the colors was a labor intensive process for the analyst. Not surprisingly, spectroscopic methods of analysis were slow to develop. The 1930s and 1940s saw the introduction of photoelectric transducers for ultraviolet and visible radiation, and thermocouples for infrared radiation. As a result, modern instrumentation for absorption spectroscopy became routinely available in the 1940s—progress has been rapid ever since.

10C.1 Instrumentation

Frequently an analyst must select—from among several instruments of different design—the one instrument best suited for a particular analysis. In this section we examine several different instruments for molecular absorption spectroscopy, emphasizing their advantages and limitations. Methods of sample introduction are also covered in this section.

INSTRUMENT DESIGNS FOR MOLECULAR UV/VIS ABSORPTION

Filter Photometer. The simplest instrument for molecular UV/Vis absorption is a **FILTER PHOTOMETER** (Figure 10.25), which uses an absorption or interference filter to isolate a band of radiation. The filter is placed between the source and the sample to prevent the sample from decomposing when exposed to higher energy radiation. A filter photometer has a single optical path between the source and detector, and is called a **SINGLE-BEAM** instrument. The instrument is calibrated to 0% T while using a shutter to block the source radiation from the detector. After opening the shutter, the in-

The percent transmittance varies between 0% and 100%. As we learned in [Figure 10.21](#), we use a blank to determine P_0 , which corresponds to 100% T. Even in the absence of light the detector records a signal. Closing the shutter allows us to assign 0% T to this signal. Together, setting 0% T and 100% T calibrates the instrument. The amount of light passing through a sample produces a signal that is greater than or equal to that for 0% T and smaller than or equal to that for 100% T.

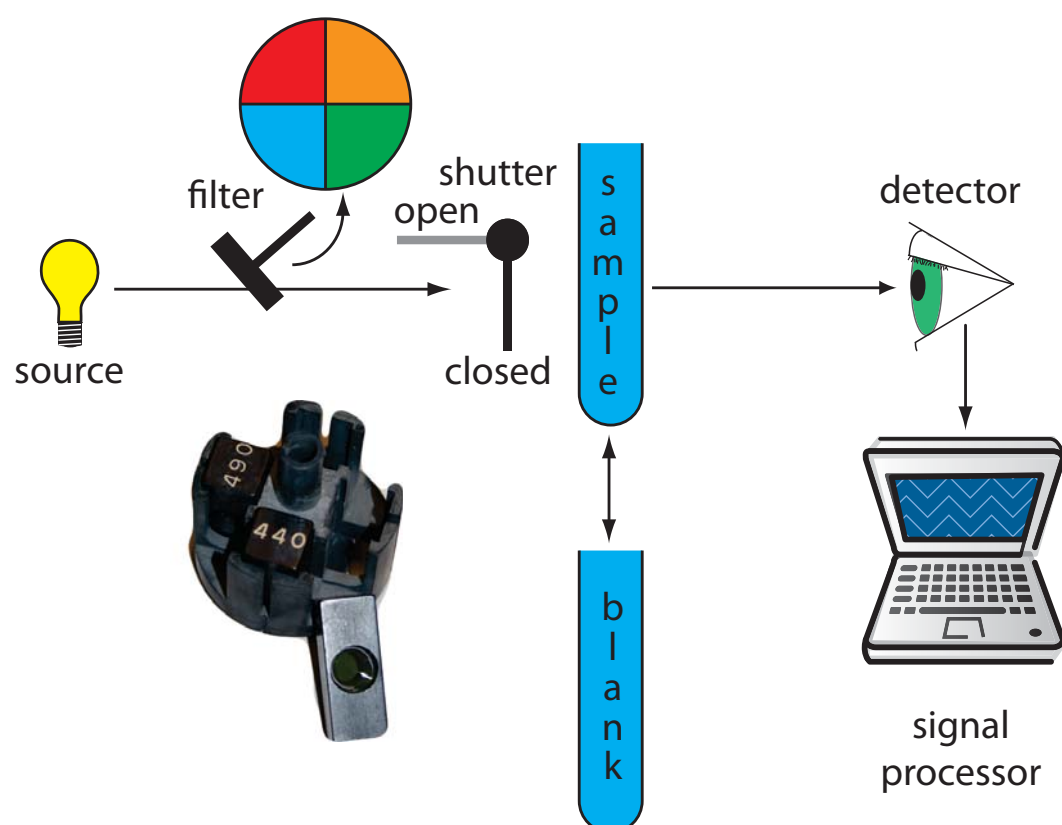


Figure 10.25 Schematic diagram of a filter photometer. The analyst either inserts a removable filter or the filters are placed in a carousel, an example of which is shown in the photographic inset. The analyst selects a filter by rotating it into place.

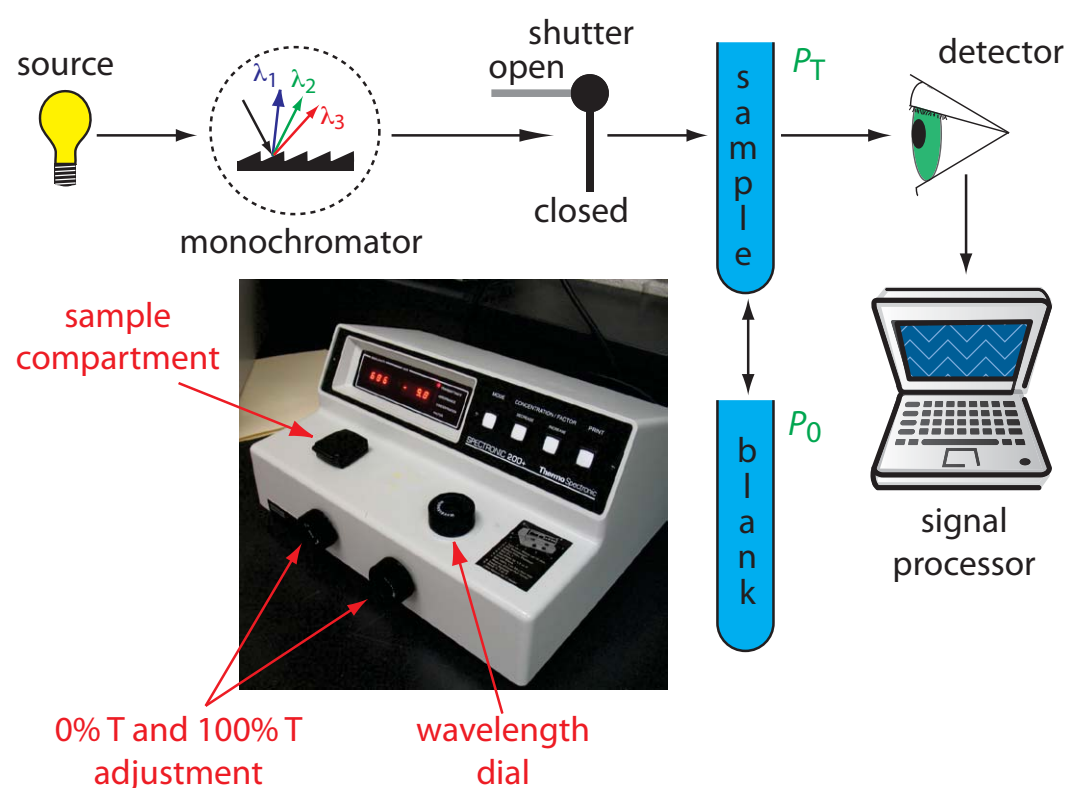


Figure 10.26 Schematic diagram of a fixed-wavelength single-beam spectrophotometer. The photographic inset shows a typical instrument. The shutter remains closed until the sample or blank is placed in the sample compartment. The analyst manually selects the wavelength by adjusting the wavelength dial. Inset photo modified from: [Adi](http://www.commonswikipedia.org) (www.commonswikipedia.org).

strument is calibrated to 100% T using an appropriate blank. The blank is then replaced with the sample and its transmittance measured. Because the source's incident power and the sensitivity of the detector vary with wavelength, the photometer must be recalibrated whenever the filter is changed. Photometers have the advantage of being relatively inexpensive, rugged, and easy to maintain. Another advantage of a photometer is its portability, making it easy to take into the field. Disadvantages of a photometer include the inability to record an absorption spectrum and the source's relatively large effective bandwidth, which limits the calibration curve's linearity.

Single-Beam Spectrophotometer. An instrument that uses a monochromator for wavelength selection is called a **SPECTROPHOTOMETER**. The simplest spectrophotometer is a single-beam instrument equipped with a fixed-wavelength monochromator (Figure 10.26). Single-beam spectrophotometers are calibrated and used in the same manner as a photometer. One example of a single-beam spectrophotometer is Thermo Scientific's Spectronic 20D+, which is shown in the photographic insert to Figure 10.26. The Spectronic 20D+ has a range of 340–625 nm (950 nm when using a red-sensitive detector), and a fixed effective bandwidth of 20 nm. Battery-operated, hand-held single-beam spectrophotometers are available, which are easy to transport into the field. Other single-beam spectrophotometers also are available with effective bandwidths of 2–8 nm. Fixed wavelength single-beam spectrophotometers are not practical for recording spectra because manually adjusting the wavelength and recalibrating the spectrophotometer is awkward and time-consuming. The accuracy of a single-beam spectrophotometer is limited by the stability of its source and detector over time.

Double-Beam Spectrophotometer. The limitations of fixed-wavelength, single-beam spectrophotometers are minimized by using a **DOUBLE-BEAM** spectrophotometer (Figure 10.27). A chopper controls the radiation's path,

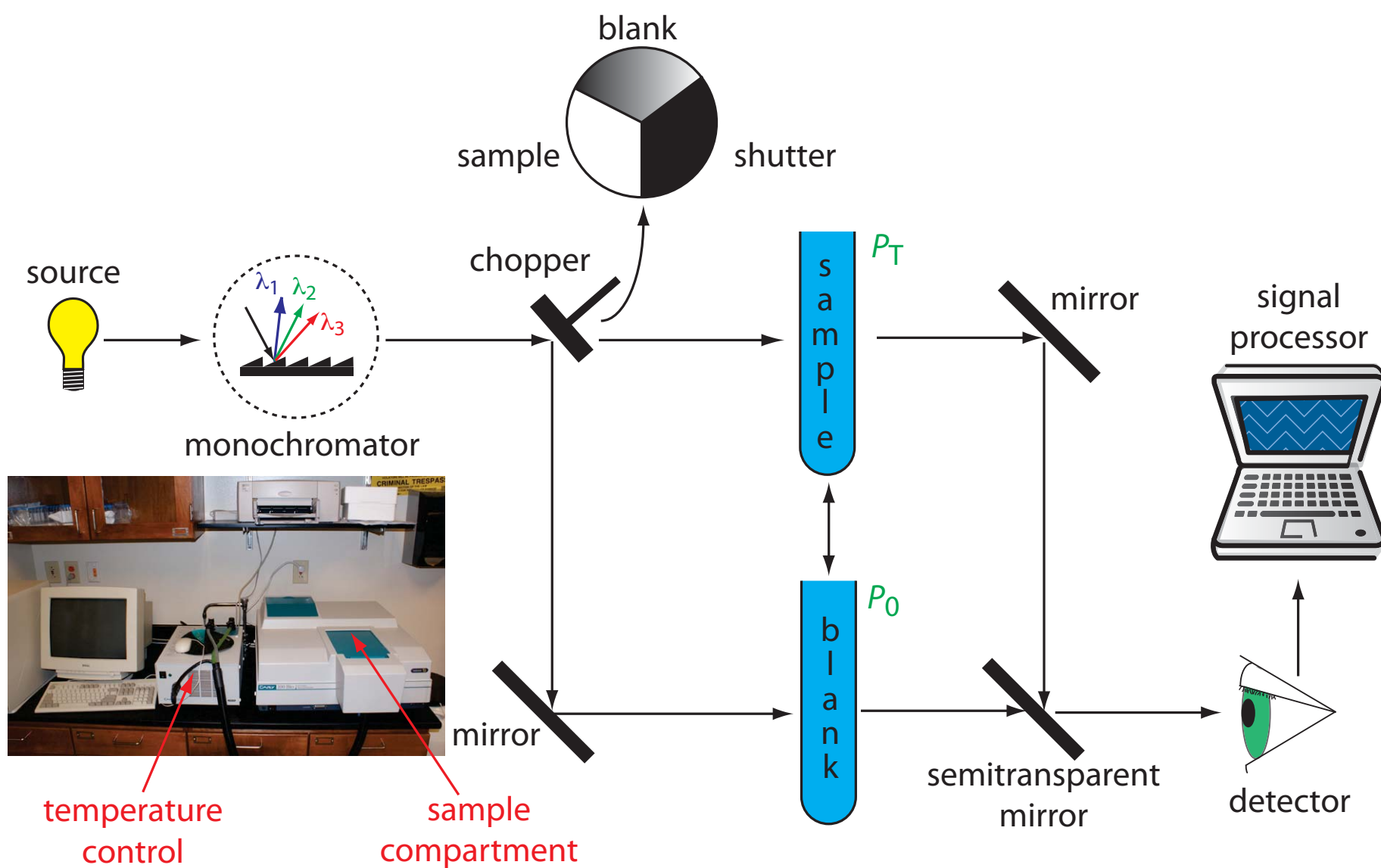


Figure 10.27 Schematic diagram of a scanning, double-beam spectrophotometer. A chopper directs the source's radiation, using a transparent window to pass radiation to the sample and a mirror to reflect radiation to the blank. The chopper's opaque surface serves as a shutter, which allows for a constant adjustment of the spectrophotometer's 0% T. The photographic insert shows a typical instrument. The unit in the middle of the photo is a temperature control unit that allows the sample to be heated or cooled.

alternating it between the sample, the blank, and a shutter. The signal processor uses the chopper's known speed of rotation to resolve the signal reaching the detector into the transmission of the blank, P_0 , and the sample, P_T . By including an opaque surface as a shutter, it is possible to continuously adjust 0% T. The effective bandwidth of a double-beam spectrophotometer is controlled by adjusting the monochromator's entrance and exit slits. Effective bandwidths of 0.2–3.0 nm are common. A scanning monochromator allows for the automated recording of spectra. Double-beam instruments are more versatile than single-beam instruments, being useful for both quantitative and qualitative analyses, but also are more expensive.

Diode Array Spectrometer. An instrument with a single detector can monitor only one wavelength at a time. If we replace a single photomultiplier with many photodiodes, we can use the resulting array of detectors to record an entire spectrum simultaneously in as little as 0.1 s. In a diode array spectrometer the source radiation passes through the sample and is dispersed by a grating (Figure 10.28). The photodiode array is situated at the grating's focal plane, with each diode recording the radiant power over a narrow range of wavelengths. Because we replace a full monochromator with just a grating, a diode array spectrometer is small and compact.

Figure 10.28 Schematic diagram of a diode array spectrophotometer. The photographic insert shows a typical instrument. Note that the 50-mL beaker provides a sense of scale.

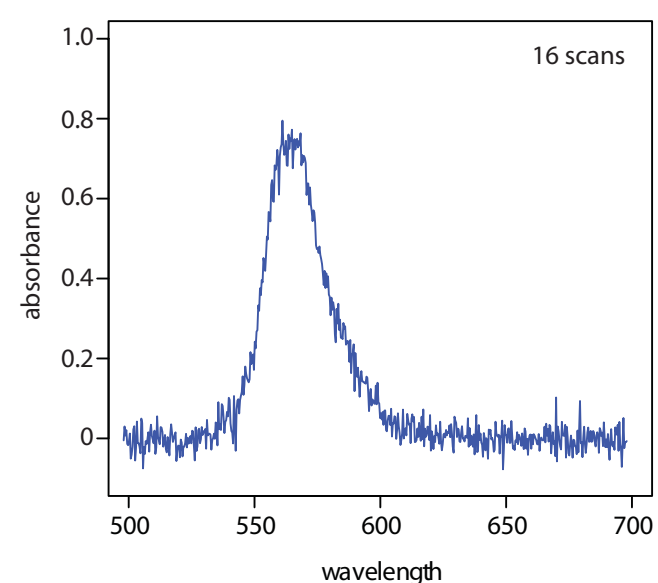
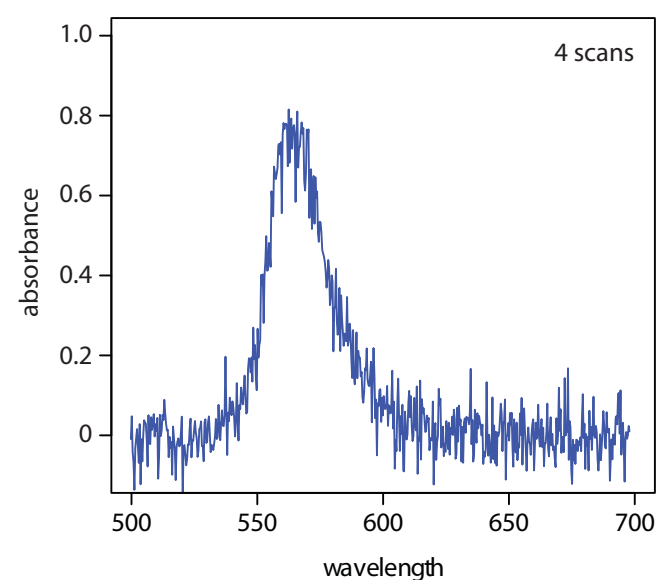
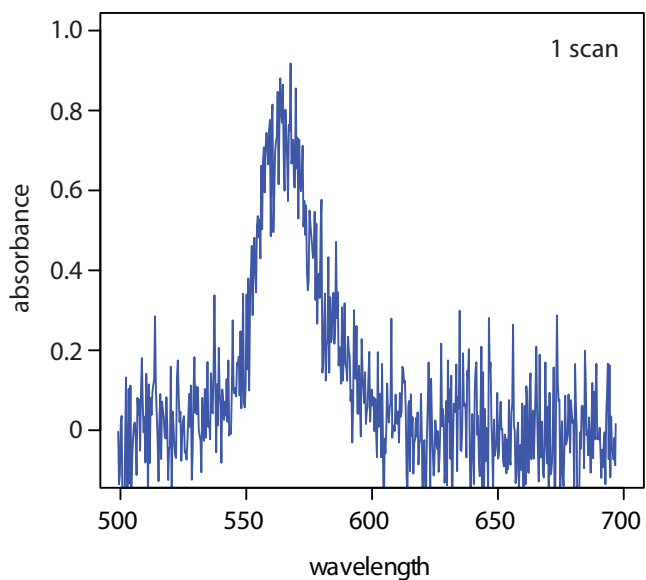
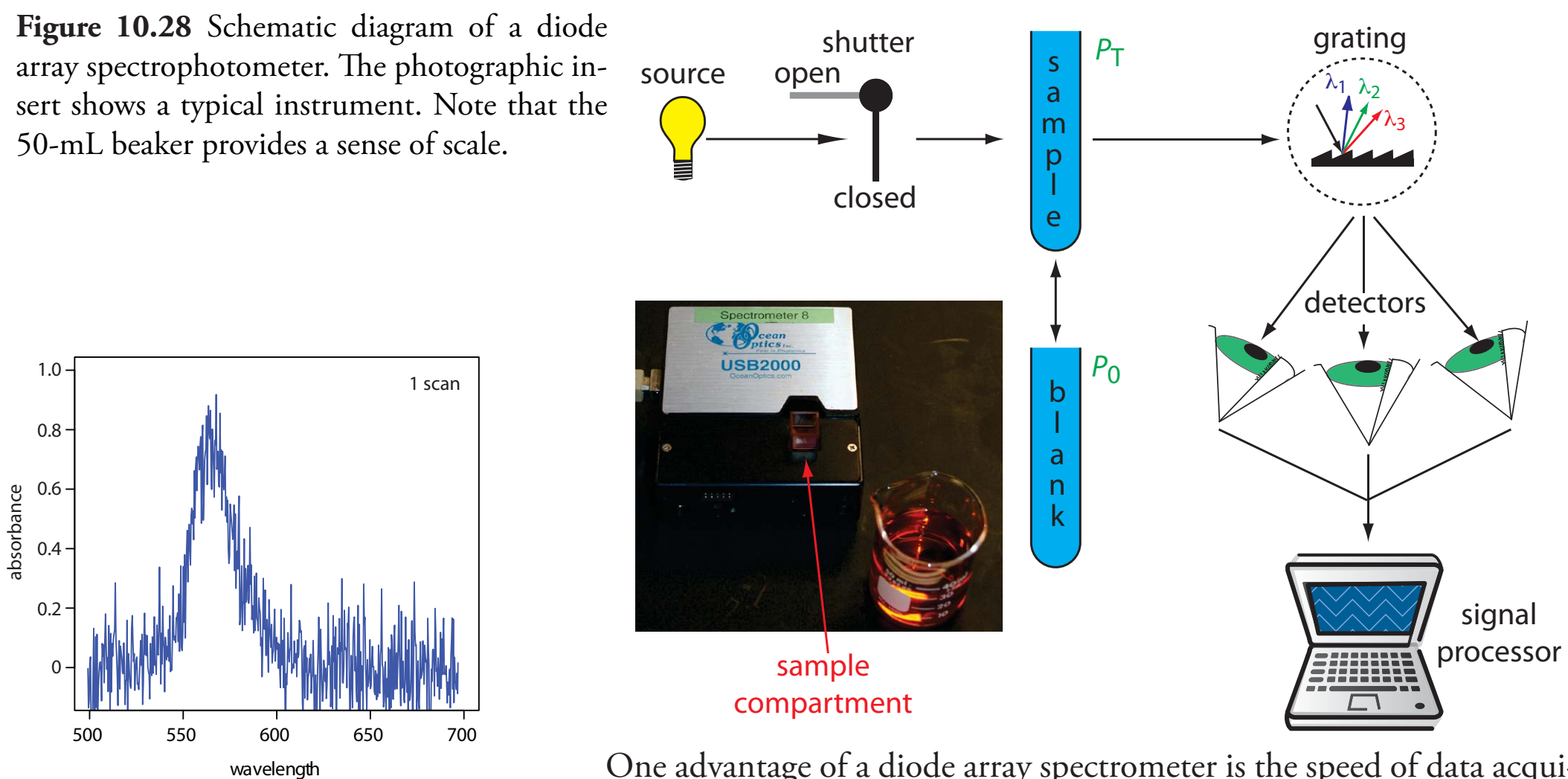


Figure 10.29 Effect of signal averaging on a spectrum's signal-to-noise ratio. From top to bottom: spectrum for a single scan; average spectrum after four scans; and average spectrum after adding 16 scans.

One advantage of a diode array spectrometer is the speed of data acquisition, which allows to collect several spectra for a single sample. Individual spectra are added and averaged to obtain the final spectrum. This **SIGNAL AVERAGING** improves a spectrum's signal-to-noise ratio. If we add together n spectra, the sum of the signal at any point, x , increases as nS_x , where S_x is the signal. The noise at any point, N_x , is a random event, which increases as $\sqrt{n}N_x$ when we add together n spectra. The **SIGNAL-TO-NOISE RATIO** (S/N) after n scans is

$$\frac{S}{N} = \frac{nS_x}{\sqrt{n}N_x} = \sqrt{n} \frac{S_x}{N_x}$$

where S_x/N_x is the signal-to-noise ratio for a single scan. The impact of signal averaging is shown in Figure 10.29. The first spectrum shows the signal for a single scan, which consists of a single, noisy peak. Signal averaging using 4 scans and 16 scans decreases the noise and improves the signal-to-noise ratio. One disadvantage of a photodiode array is that the effective bandwidth per diode is roughly an order of magnitude larger than that for a high quality monochromator.

Sample Cells. The sample compartment provides a light-tight environment that limits the addition of stray radiation. Samples are normally in the liquid or solution state, and are placed in cells constructed with UV/Vis transparent materials, such as quartz, glass, and plastic (Figure 10.30). A quartz or fused-silica cell is required when working at a wavelength <300 nm where other materials show a significant absorption. The most common pathlength is 1 cm (10 mm), although cells with shorter (as little as 0.1 cm) and longer pathlengths (up to 10 cm) are available. Longer pathlength cells



Figure 10.30 Examples of sample cells for UV/Vis spectroscopy. From left to right (with path lengths in parentheses): rectangular plastic cuvette (10.0 mm), rectangular quartz cuvette (5.000 mm), rectangular quartz cuvette (1.000 mm), cylindrical quartz cuvette (10.00 mm), cylindrical quartz cuvette (100.0 mm). Cells often are available as a matched pair, which is important when using a double-beam instrument.

are useful when analyzing a very dilute solution, or for gas samples. The highest quality cells allow the radiation to strike a flat surface at a 90° angle, minimizing the loss of radiation to reflection. A test tube is often used as a sample cell with simple, single-beam instruments, although differences in the cell's pathlength and optical properties add an additional source of error to the analysis.

If we need to monitor an analyte's concentration over time, it may not be possible to physically remove samples for analysis. This is often the case, for example, when monitoring industrial production lines or waste lines, when monitoring a patient's blood, or when monitoring environmental systems. With a **FIBER-OPTIC PROBE** we can analyze samples in situ. An example of a remote sensing fiber-optic probe is shown in Figure 10.31. The probe consists of two bundles of fiber-optic cable. One bundle transmits radiation from the source to the probe's tip, which is designed to allow the sample to flow through the sample cell. Radiation from the source passes through the solution and is reflected back by a mirror. The second bundle

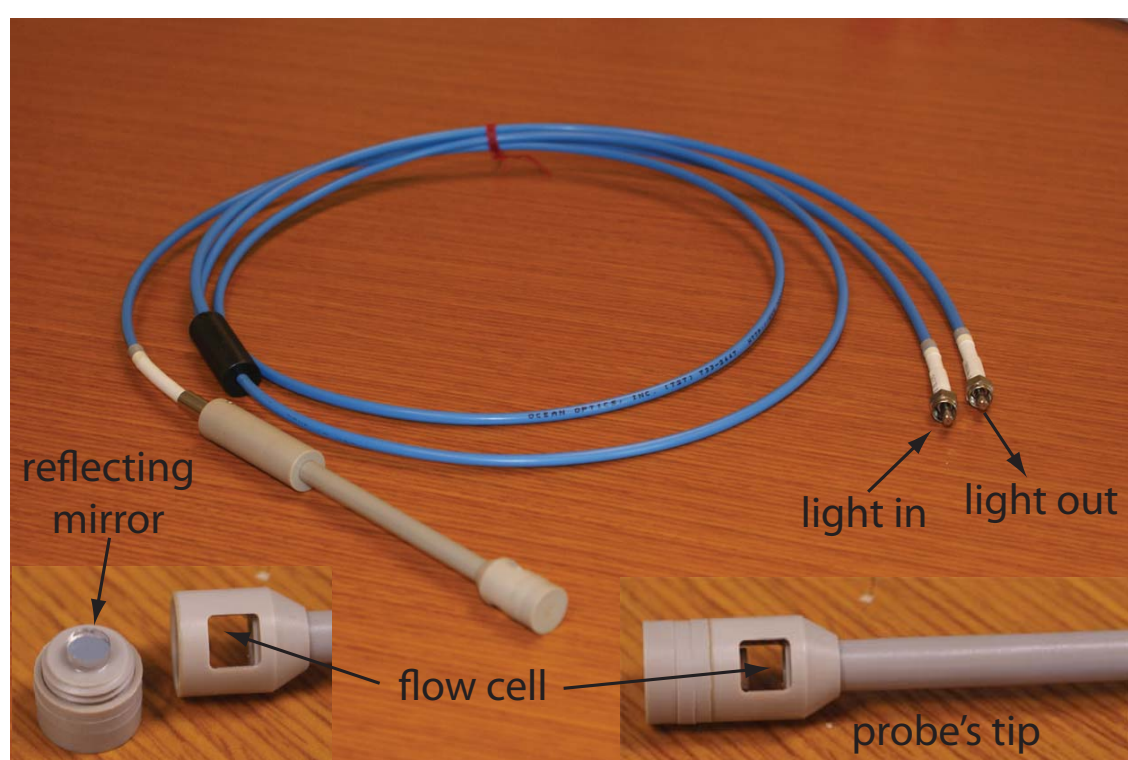


Figure 10.31 Example of a fiber-optic probe. The inset photographs provide a close-up look at the probe's flow cell and the reflecting mirror

of fiber-optic cable transmits the nonabsorbed radiation to the wavelength selector. Another design replaces the flow cell shown in [Figure 10.31](#) with a membrane containing a reagent that reacts with the analyte. When the analyte diffuses across the membrane it reacts with the reagent, producing a product that absorbs UV or visible radiation. The nonabsorbed radiation from the source is reflected or scattered back to the detector. Fiber optic probes that show chemical selectivity are called optrodes.⁶

INSTRUMENT DESIGNS FOR INFRARED ABSORPTION

Filter Photometer. The simplest instrument for IR absorption spectroscopy is a filter photometer similar to that shown in [Figure 10.25](#) for UV/Vis absorption. These instruments have the advantage of portability, and typically are used as dedicated analyzers for gases such as HCN and CO.

Double-beam spectrophotometer. Infrared instruments using a monochromator for wavelength selection use double-beam optics similar to that shown in [Figure 10.27](#). Double-beam optics are preferred over single-beam optics because the sources and detectors for infrared radiation are less stable than those for UV/Vis radiation. In addition, it is easier to correct for the absorption of infrared radiation by atmospheric CO₂ and H₂O vapor when using double-beam optics. Resolutions of 1–3 cm⁻¹ are typical for most instruments.

Fourier transform spectrometer. In a Fourier transform infrared spectrometer, or FT-IR, the monochromator is replaced with an interferometer ([Figure 10.13](#)). Because an FT-IR includes only a single optical path, it is necessary to collect a separate spectrum to compensate for the absorbance of atmospheric CO₂ and H₂O vapor. This is done by collecting a background spectrum without the sample and storing the result in the instrument's computer memory. The background spectrum is removed from the sample's spectrum by ratioing the two signals. In comparison to other instrument designs, an FT-IR provides for rapid data acquisition, allowing an enhancement in signal-to-noise ratio through signal-averaging.

Sample Cells. Infrared spectroscopy is routinely used to analyze gas, liquid, and solid samples. Sample cells are made from materials, such as NaCl and KBr, that are transparent to infrared radiation. Gases are analyzed using a cell with a pathlength of approximately 10 cm. Longer pathlengths are obtained by using mirrors to pass the beam of radiation through the sample several times.

A liquid samples may be analyzed using a variety of different sample cells ([Figure 10.32](#)). For non-volatile liquids a suitable sample can be prepared by placing a drop of the liquid between two NaCl plates, forming a thin film that typically is less than 0.01 mm thick. Volatile liquids must be placed in a sealed cell to prevent their evaporation.

⁶ (a) Seitz, W. R. *Anal. Chem.* **1984**, *56*, 16A–34A; (b) Angel, S. M. *Spectroscopy* **1987**, *2*(2), 38–48.

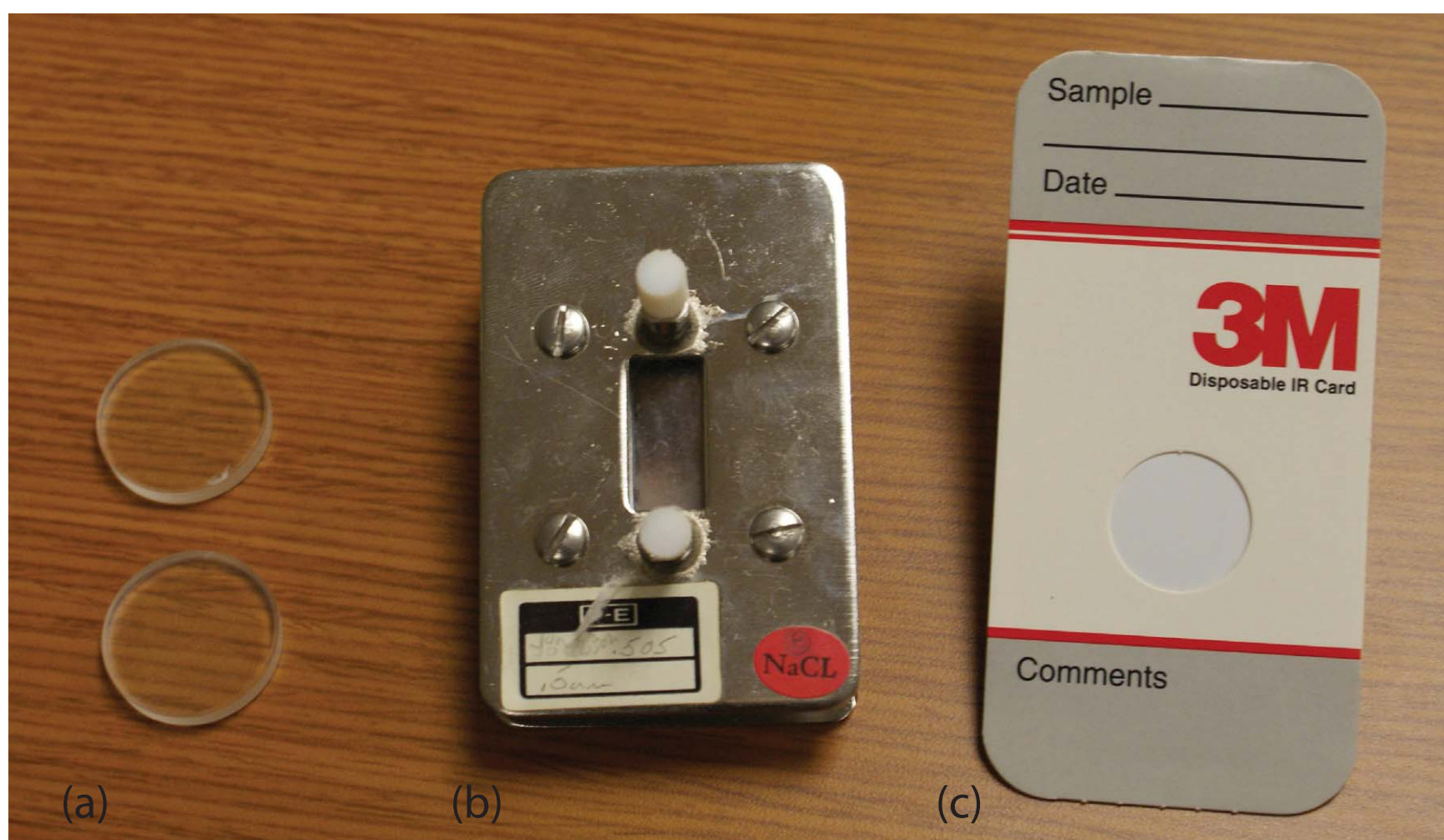


Figure 10.32 Three examples of IR sample cells: (a) NaCl salts plates; (b) fixed pathlength (0.5 mm) sample cell with NaCl windows; (c) disposable card with a polyethylene window that is IR transparent with the exception of strong absorption bands at 2918 cm^{-1} and 2849 cm^{-1} .

The analysis of solution samples is limited by the solvent's IR absorbing properties, with CCl_4 , CS_2 , and CHCl_3 being the most common solvents. Solutions are placed in cells containing two NaCl windows separated by a Teflon spacer. By changing the Teflon spacer, pathlengths from 0.015–1.0 mm can be obtained.

Transparent solid samples can be analyzed directly by placing them in the IR beam. Most solid samples, however, are opaque, and must be dispersed in a more transparent medium before recording the IR spectrum. If a suitable solvent is available, then the solid can be analyzed by preparing a solution and analyzing as described above. When a suitable solvent is not available, solid samples may be analyzed by preparing a mull of the finely powdered sample with a suitable oil. Alternatively, the powdered sample can be mixed with KBr and pressed into an optically transparent pellet.

The analysis of an aqueous sample is complicated by the solubility of the NaCl cell window in water. One approach to obtaining infrared spectra on aqueous solutions is to use **ATTENUATED TOTAL REFLECTANCE** instead of transmission. [Figure 10.33](#) shows a diagram of a typical attenuated total reflectance (ATR) FT-IR instrument. The ATR cell consists of a high refractive index material, such as ZnSe or diamond, sandwiched between a low refractive index substrate and a lower refractive index sample. Radiation from the source enters the ATR crystal where it undergoes a series of total internal reflections before exiting the crystal. During each reflection the radiation penetrates into the sample to a depth of a few microns. The result is a selective attenuation of the radiation at those wavelengths where

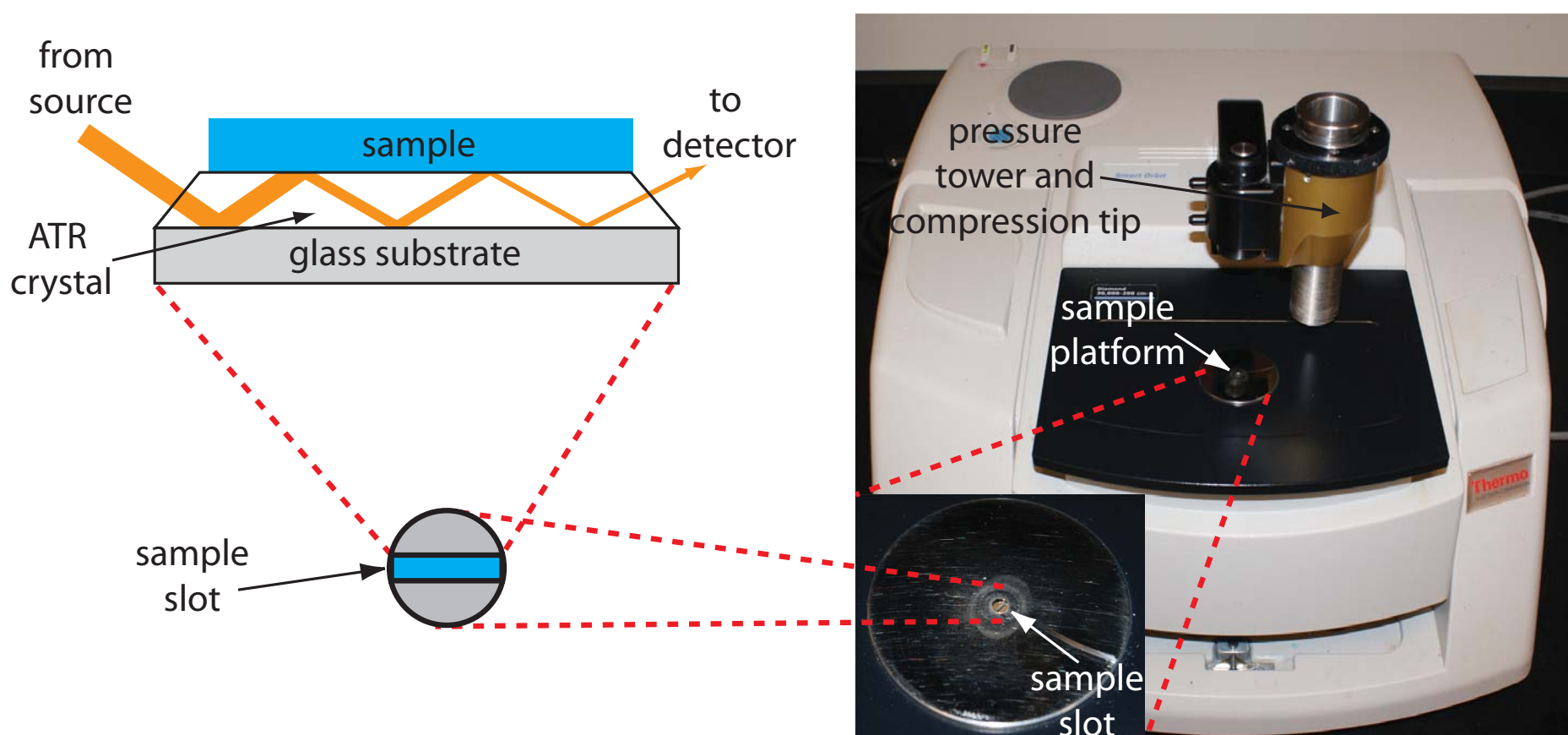


Figure 10.33 FT-IR spectrometer equipped with a diamond ATR sample cell. The inserts show a close-up photo of the sample platform, a sketch of the ATR's sample slot, and a schematic showing how the source's radiation interacts with the sample. The pressure tower is used to ensure the contact of solid samples with the ATR crystal.

the sample absorbs. ATR spectra are similar, but not identical, to those obtained by measuring the transmission of radiation.

Solid samples also can be analyzed using an ATR sample cell. After placing the solid in the sample slot, a compression tip ensures that it is in contact with the ATR crystal. Examples of solids that have been analyzed by ATR include polymers, fibers, fabrics, powders, and biological tissue samples. Another reflectance method is diffuse reflectance, in which radiation is reflected from a rough surface, such as a powder. Powdered samples are mixed with a non-absorbing material, such as powdered KBr, and the reflected light is collected and analyzed. As with ATR, the resulting spectrum is similar to that obtained by conventional transmission methods.

10C.2 Quantitative Applications

The determination of an analyte's concentration based on its absorption of ultraviolet or visible radiation is one of the most frequently encountered quantitative analytical methods. One reason for its popularity is that many organic and inorganic compounds have strong absorption bands in the UV/Vis region of the electromagnetic spectrum. In addition, if an analyte does not absorb UV/Vis radiation—or if its absorbance is too weak—we often can react it with another species that is strongly absorbing. For example, a dilute solution of Fe^{2+} does not absorb visible light. Reacting Fe^{2+} with *o*-phenanthroline, however, forms an orange-red complex of $\text{Fe}(\text{phen})_3^{2+}$ that has a strong, broad absorbance band near 500 nm. An additional advantage to UV/Vis absorption is that in most cases it is relatively easy to adjust experimental and instrumental conditions so that Beer's law is obeyed.

Further details about these, and other methods for preparing solids for infrared analysis can be found in this chapter's [additional resources](#).

[Figure 10.18](#) shows the visible spectrum for $\text{Fe}(\text{phen})_3^{2+}$.

A quantitative analysis based on the absorption of infrared radiation, although important, is less frequently encountered than those for UV/Vis absorption. One reason is the greater tendency for instrumental deviations from Beer's law when using infrared radiation. Because an infrared absorption band is relatively narrow, any deviation due to the lack of monochromatic radiation is more pronounced. In addition, infrared sources are less intense than UV/Vis sources, making stray radiation more of a problem. Differences in pathlength for samples and standards when using thin liquid films or KBr pellets are a problem, although an internal standard can be used to correct for any difference in pathlength. Finally, establishing a 100% T ($A = 0$) baseline is often difficult because the optical properties of NaCl sample cells may change significantly with wavelength due to contamination and degradation. We can minimize this problem by measuring absorbance relative to a baseline established for the absorption band. Figure 10.34 shows how this is accomplished.

ENVIRONMENTAL APPLICATIONS

The analysis of waters and wastewaters often relies on the absorption of ultraviolet and visible radiation. Many of these methods are outlined in [Table 10.6](#). Several of these methods are described here in more detail.

Although the quantitative analysis of metals in waters and wastewaters is accomplished primarily by atomic absorption or atomic emission spectroscopy, many metals also can be analyzed following the formation of a colored metal–ligand complex. One advantage to these spectroscopic methods is that they are easily adapted to the analysis of samples in the field using a filter photometer. One ligand that is used in the analysis of several metals is diphenylthiocarbazone, also known as dithizone. Dithizone is not soluble in water, but when a solution of dithizone in CHCl_3 is shaken

Another approach is to use a cell with a fixed pathlength, such as that shown in [Figure 10.32b](#).

Atomic absorption is the subject of [Section 10D](#) and atomic emission is the subject of [Section 10G](#).

The structure of dithizone is shown to the right. See Chapter 7 for a discussion of extracting metal ions using dithizone.

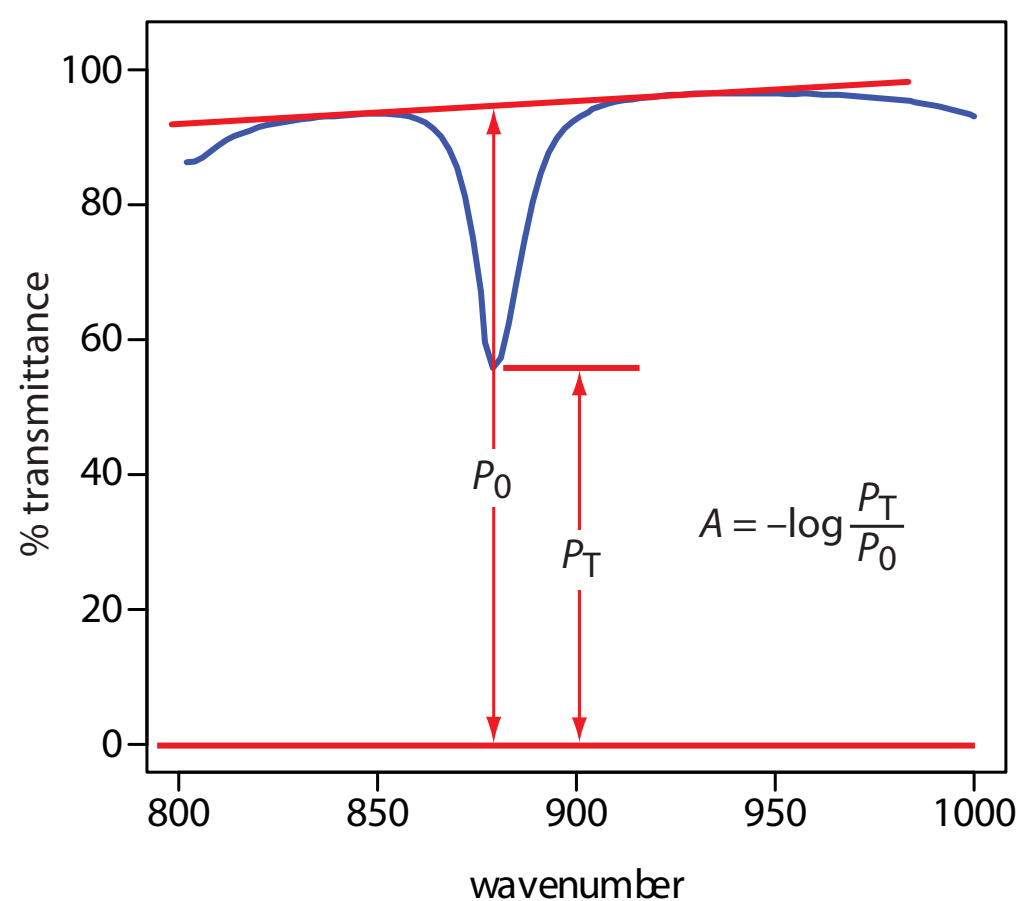
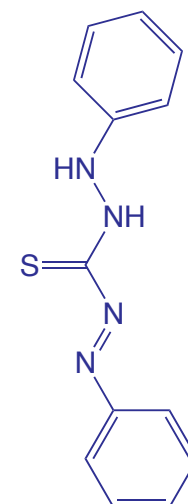


Figure 10.34 Method for determining absorbance from an IR spectrum.

Table 10.6 Examples of the Molecular UV/Vis Analysis of Waters and Wastewaters

Analyte	Method	λ (nm)
<i>Trace Metals</i>		
aluminum	react with Eriochrome cyanide R dye at pH 6; forms red to pink complex	535
arsenic	reduce to AsH ₃ using Zn and react with silver diethyldithiocarbamate; forms red complex	535
cadmium	extract into CHCl ₃ containing dithizone from a sample made basic with NaOH; forms pink to red complex	518
chromium	oxidize to Cr(VI) and react with diphenylcarbazide; forms red-violet product	540
copper	react with neocuprine in neutral to slightly acid solution and extract into CHCl ₃ /CH ₃ OH; forms yellow complex	457
iron	reduce to Fe ²⁺ and react with <i>o</i> -phenanthroline; forms orange-red complex	510
lead	extract into CHCl ₃ containing dithizone from sample made basic with NH ₃ /NH ₄ ⁺ buffer; forms cherry red complex	510
manganese	oxidize to MnO ₄ ⁻ with persulfate; forms purple solution	525
mercury	extract into CHCl ₃ containing dithizone from acidic sample; forms orange complex	492
zinc	react with zincon at pH 9; forms blue complex	620
<i>Inorganic Nonmetals</i>		
ammonia	reaction with hypochlorite and phenol using a manganous salt catalyst; forms blue indophenol as product	630
cyanide	react with chloroamine-T to form CNCl and then with a pyridine-barbituric acid; forms a red-blue dye	578
fluoride	react with red Zr-SPADNS lake; formation of ZrF ₆ ²⁻ decreases color of the red lake	570
chlorine (residual)	react with leuco crystal violet; forms blue product	592
nitrate	react with Cd to form NO ₂ ⁻ and then react with sulfanilamide and <i>N</i> -(1-naphthyl)-ethylenediamine; forms red azo dye	543
phosphate	react with ammonium molybdate and then reduce with SnCl ₂ ; forms molybdenum blue	690
<i>Organics</i>		
phenol	react with 4-aminoantipyrine and K ₃ Fe(CN) ₆ ; forms yellow antipyrine dye	460
anionic surfactant	react with cationic methylene blue dye and extract into CHCl ₃ ; forms blue ion pair	652

with an aqueous solution containing an appropriate metal ion, a colored metal–dithizonate complex forms that is soluble in CHCl_3 . The selectivity of dithizone is controlled by adjusting the sample's pH. For example, Cd^{2+} is extracted from solutions that are made strongly basic with NaOH , Pb^{2+} from solutions that are made basic with an $\text{NH}_3/\text{NH}_4^+$ buffer, and Hg^{2+} from solutions that are slightly acidic.

When chlorine is added to water the portion available for disinfection is called the chlorine residual. There are two forms of chlorine residual. The free chlorine residual includes Cl_2 , HOCl , and OCl^- . The combined chlorine residual, which forms from the reaction of NH_3 with HOCl , consists of monochloramine, NH_2Cl , dichloramine, NHCl_2 , and trichloramine, NCl_3 . Because the free chlorine residual is more efficient at disinfection, there is an interest in methods that can distinguish between the different forms of the total chlorine residual. One such method is the leuco crystal violet method. The free residual chlorine is determined by adding leuco crystal violet to the sample, which instantaneously oxidizes to give a blue colored compound that is monitored at 592 nm. Completing the analysis in less than five minutes prevents a possible interference from the combined chlorine residual. The total chlorine residual (free + combined) is determined by reacting a separate sample with iodide, which reacts with both chlorine residuals to form HOI . When the reaction is complete, leuco crystal violet is added and oxidized by HOI , giving the same blue colored product. The combined chlorine residual is determined by difference.

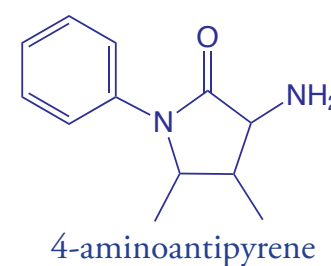
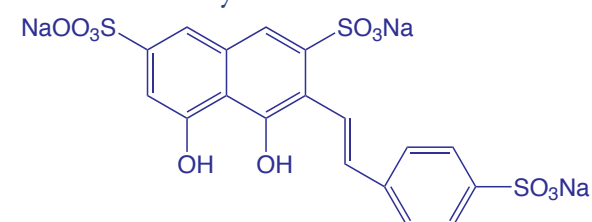
The concentration of fluoride in drinking water may be determined indirectly by its ability to form a complex with zirconium. In the presence of the dye SPADNS, solutions of zirconium form a red colored compound, called a lake, that absorbs at 570 nm. When fluoride is added, the formation of the stable ZrF_6^{2-} complex causes a portion of the lake to dissociate, decreasing the absorbance. A plot of absorbance versus the concentration of fluoride, therefore, has a negative slope.

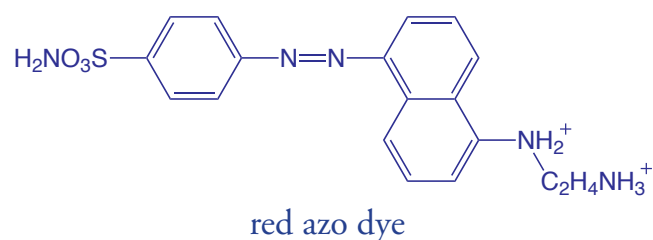
Spectroscopic methods also are used to determine organic constituents in water. For example, the combined concentrations of phenol, and ortho- and meta- substituted phenols are determined by using steam distillation to separate the phenols from nonvolatile impurities. The distillate reacts with 4-aminoantipyrine at $\text{pH } 7.9 \pm 0.1$ in the presence of $\text{K}_3\text{Fe}(\text{CN})_6$, forming a yellow colored antipyrine dye. After extracting the dye into CHCl_3 , its absorbance is monitored at 460 nm. A calibration curve is prepared using only the unsubstituted phenol, $\text{C}_6\text{H}_5\text{OH}$. Because the molar absorptivity of substituted phenols are generally less than that for phenol, the reported concentration represents the minimum concentration of phenolic compounds.

Molecular absorption also can be used for the analysis of environmentally significant airborne pollutants. In many cases the analysis is carried out by collecting the sample in water, converting the analyte to an aqueous form that can be analyzed by methods such as those described in [Table](#)

In Chapter 9 we explored how the total chlorine residual can be determined by a redox titration; see [Representative Method 9.3](#) for further details. The method described here allows us to divide the total chlorine residual into its component parts.

SPADNS, which is shown below, is an abbreviation for the sodium salt of 2-(4-sulfophenylazo)-1,8-dihydroxy-3,6-naphthalenedisulfonic acid, which is a mouthful to say.





10.6. For example, the concentration of NO_2 can be determined by oxidizing NO_2 to NO_3^- . The concentration of NO_3^- is then determined by first reducing it to NO_2^- with Cd, and then reacting NO_2^- with sulfanilamide and *N*-(1-naphthyl)-ethylenediamine to form a red azo dye. Another important application is the analysis for SO_2 , which is determined by collecting the sample in an aqueous solution of HgCl_4^{2-} where it reacts to form $\text{Hg}(\text{SO}_3)_2^{2-}$. Addition of *p*-rosaniline and formaldehyde produces a purple complex that is monitored at 569 nm. Infrared absorption is useful for the analysis of organic vapors, including HCN, SO_2 , nitrobenzene, methyl mercaptan, and vinyl chloride. Frequently, these analyses are accomplished using portable, dedicated infrared photometers.

CLINICAL APPLICATIONS

The analysis of clinical samples is often complicated by the complexity of the sample matrix, which may contribute a significant background absorption at the desired wavelength. The determination of serum barbiturates provides one example of how this problem is overcome. The barbiturates are first extracted from a sample of serum with CHCl_3 and then extracted from the CHCl_3 into 0.45 M NaOH ($\text{pH} \approx 13$). The absorbance of the aqueous extract is measured at 260 nm, and includes contributions from the barbiturates as well as other components extracted from the serum sample. The pH of the sample is then lowered to approximately 10 by adding NH_4Cl and the absorbance remeasured. Because the barbiturates do not absorb at this pH, we can use the absorbance at pH 10, $A_{\text{pH } 10}$, to correct the absorbance at pH 13, $A_{\text{pH } 13}$

$$A_{\text{barb}} = A_{\text{pH } 13} - \frac{V_{\text{samp}} + V_{\text{NH}_4\text{Cl}}}{V_{\text{samp}}} \times A_{\text{pH } 10}$$

where A_{barb} is the absorbance due to the serum barbiturates, and V_{samp} and $V_{\text{NH}_4\text{Cl}}$ are the volumes of sample and NH_4Cl , respectively. [Table 10.7](#) provides a summary of several other methods for analyzing clinical samples.

INDUSTRIAL ANALYSIS

UV/Vis molecular absorption is used for the analysis of a diverse array of industrial samples including pharmaceuticals, food, paint, glass, and metals. In many cases the methods are similar to those described in [Table 10.6](#) and [Table 10.7](#). For example, the amount of iron in food can be determined by bringing the iron into solution and analyzing using the *o*-phenanthroline method listed in [Table 10.6](#).

Many pharmaceutical compounds contain chromophores that make them suitable for analysis by UV/Vis absorption. Products that have been analyzed in this fashion include antibiotics, hormones, vitamins, and analgesics. One example of the use of UV absorption is in determining the

Table 10.7 Examples of the Molecular UV/Vis Analysis of Clinical Samples

Analyte	Method	λ (nm)
total serum protein	react with NaOH and Cu^{2+} ; forms blue-violet complex	540
serum cholesterol	react with Fe^{3+} in presence of isopropanol, acetic acid, and H_2SO_4 ; forms blue-violet complex	540
uric acid	react with phosphotungstic acid; forms tungsten blue	710
serum barbiturates	extract into CHCl_3 to isolate from interferences and then extract into 0.45 M NaOH	260
glucose	react with <i>o</i> -toluidine at 100°C ; forms blue-green complex	630
protein-bound iodine	decompose protein to release iodide, which catalyzes redox reaction between Ce^{3+} and As^{3+} ; forms yellow colored Ce^{4+}	420

purity of aspirin tablets, for which the active ingredient is acetylsalicylic acid. Salicylic acid, which is produced by the hydrolysis of acetylsalicylic acid, is an undesirable impurity in aspirin tablets, and should not be present at more than 0.01% w/w. Samples can be screened for unacceptable levels of salicylic acid by monitoring the absorbance at a wavelength of 312 nm. Acetylsalicylic acid absorbs at 280 nm, but absorbs poorly at 312 nm. Conditions for preparing the sample are chosen such that an absorbance of greater than 0.02 signifies an unacceptable level of salicylic acid.

FORENSIC APPLICATIONS

UV/Vis molecular absorption is routinely used for the analysis of narcotics and for drug testing. One interesting forensic application is the determination of blood alcohol using the Breathalyzer test. In this test a 52.5-mL breath sample is bubbled through an acidified solution of $\text{K}_2\text{Cr}_2\text{O}_7$, which oxidizes ethanol to acetic acid. The concentration of ethanol in the breath sample is determined by the decrease in absorbance at 440 nm where the dichromate ion absorbs. A blood alcohol content of 0.10%, which is above the legal limit, corresponds to 0.025 mg of ethanol in the breath sample.

DEVELOPING A QUANTITATIVE METHOD FOR A SINGLE COMPONENT

In developing a quantitative analytical method, the conditions under which Beer's law is obeyed must be established. First, the most appropriate wavelength for the analysis is determined from an absorption spectrum. In most cases the best wavelength corresponds to an absorption maximum because it provides greater sensitivity and is less susceptible to instrumental limitations. Second, if an instrument with adjustable slits is being used, then an appropriate slit width needs to be chosen. The absorption spectrum also aids in selecting a slit width. Usually we set the slits to be as wide as possible because this increases the throughput of source radiation, while also being narrow enough to avoid instrumental limitations to Beer's law. Finally, a calibration curve is constructed to determine the range of concentrations for which Beer's law is valid. Additional considerations that are important

The best way to appreciate the theoretical and practical details discussed in this section is to carefully examine a typical analytical method. Although each method is unique, the following description of the determination of iron in water and wastewater provides an instructive example of a typical procedure. The description here is based on Method 3500-Fe B as published in *Standard Methods for the Examination of Water and Wastewater*, 20th Ed., American Public Health Association: Washington, D. C., 1998.

Figure 10.18 shows the visible spectrum for $\text{Fe}(\text{phen})_3^{2+}$.

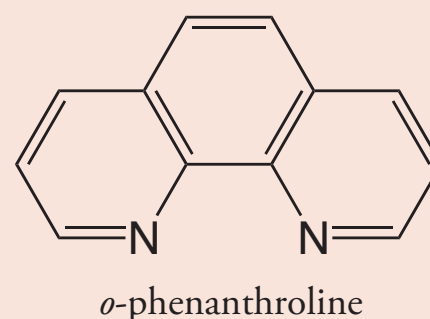
in any quantitative method are the effect of potential interferences and establishing an appropriate blank.

Representative Method 10.1

Determination of Iron in Water and Wastewater

DESCRIPTION OF METHOD

Iron in the +2 oxidation state reacts with *o*-phenanthroline to form the orange-red $\text{Fe}(\text{phen})_3^{2+}$ complex. The intensity of the complex's color is independent of solution acidity between a pH of 3 and 9. Because the complex forms more rapidly at lower pH levels, the reaction is usually carried out within a pH range of 3.0–3.5. Any iron present in the +3 oxidation state is reduced with hydroxylamine before adding *o*-phenanthroline. The most important interferences are strong oxidizing agents, polyphosphates, and metal ions such as Cu^{2+} , Zn^{2+} , Ni^{2+} , and Cd^{2+} . An interference from oxidizing agents is minimized by adding an excess of hydroxylamine, and an interference from polyphosphate is minimized by boiling the sample in the presence of acid. The absorbance of samples and standards are measured at a wavelength of 510 nm using a 1-cm cell (longer pathlength cells also may be used). Beer's law is obeyed for concentrations of within the range of 0.2–4.0 mg Fe/L



PROCEDURE

For samples containing less than 2 mg Fe/L, directly transfer a 50-mL portion to a 125-mL Erlenmeyer flask. Samples containing more than 2 mg Fe/L must be diluted before acquiring the 50-mL portion. Add 2 mL of concentrated HCl and 1 mL of hydroxylamine to the sample. Heat the solution to boiling and continue boiling until the solution's volume is reduced to between 15 and 20 mL. After cooling to room temperature, transfer the solution to a 50-mL volumetric flask, add 10 mL of an ammonium acetate buffer, 2 mL of a 1000 ppm solution of *o*-phenanthroline, and dilute to volume. Allow 10–15 minutes for color development before measuring the absorbance, using distilled water to set 100% T. Calibration standards, including a blank, are prepared by the same procedure using a stock solution containing a known concentration of Fe^{2+} .

QUESTIONS

1. Explain why strong oxidizing agents are interferences, and why an excess of hydroxylamine prevents the interference.

A strong oxidizing agent oxidizes some Fe^{2+} to Fe^{3+} . Because $\text{Fe}(\text{phen})_3^{3+}$ does not absorb as strongly as $\text{Fe}(\text{phen})_3^{2+}$, the absorbance decreases, producing a negative determinate error. The excess hydroxylamine reacts with the oxidizing agents, removing them from the solution.

2. The color of the complex is stable between pH levels of 3 and 9. What are some possible complications at more acidic or more basic pH's?

Because *o*-phenanthroline is a weak base, its conditional formation constant for $\text{Fe}(\text{phen})_3^{2+}$ is less favorable at more acidic pH levels, where *o*-phenanthroline is protonated. The result is a decrease in absorbance and a less sensitive analytical method. When the pH is greater than 9, competition between OH^- and *o*-phenanthroline for Fe^{2+} also decreased the absorbance. In addition, if the pH is sufficiently basic there is a risk that the iron will precipitate as $\text{Fe}(\text{OH})_2$.

3. Cadmium is an interferent because it forms a precipitate with *o*-phenanthroline. What effect would the formation of precipitate have on the determination of iron?

Because *o*-phenanthroline is present in large excess (2000 μg of *o*-phenanthroline for 100 μg of Fe^{2+}), it is not likely that the interference is due to an insufficient amount of *o*-phenanthroline being available to react with the Fe^{2+} . The presence of a precipitate in the sample cell results in the scattering of radiation, which causes an apparent increase in absorbance. Because the measured absorbance increases, the reported concentration is too high.

4. Even high quality ammonium acetate contains a significant amount of iron. Why is this source of iron not a problem?

Because all samples and standards are prepared using the same volume of ammonium acetate buffer, the contribution of this source of iron is accounted for by the calibration curve's reagent blank.

In [Chapter 9](#) we saw the same effect of pH on the complexation reactions between EDTA and metal ions.

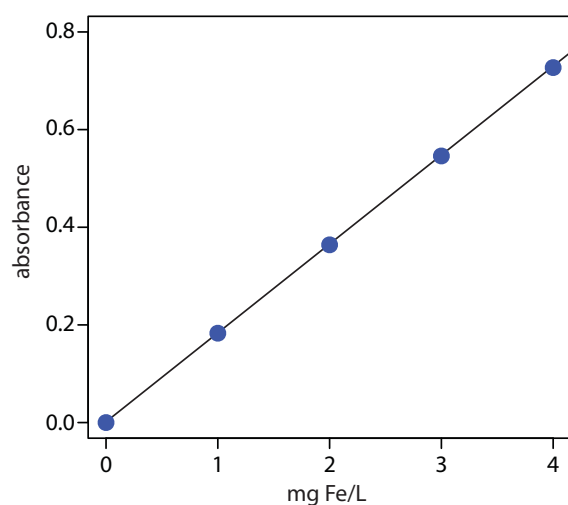
Although scattering is a problem here, it can serve as the basis of a useful analytical method. See [Section 10H](#) for further details.

QUANTITATIVE ANALYSIS FOR A SINGLE ANALYTE

To determine the concentration of a an analyte we measure its absorbance and apply Beer's law using any of the standardization methods described in Chapter 5. The most common methods are a normal calibration curve using external standards and the method of standard additions. A single point standardization is also possible, although we must first verify that Beer's law holds for the concentration of analyte in the samples and the standard.

Example 10.5

The determination of Fe in an industrial waste stream was carried out by the *o*-phenanthroline described in [Representative Method 10.1](#). Using the data in the following table, determine the mg Fe/L in the waste stream.



mg Fe/L	absorbance
0.00	0.000
1.00	0.183
2.00	0.364
3.00	0.546
4.00	0.727
sample	0.269

SOLUTION

Linear regression of absorbance versus the concentration of Fe in the standards gives a calibration curve with the following equation.

$$A = 0.0006 + 0.1817 \times (\text{mg Fe/L})$$

Substituting the sample's absorbance into the calibration expression gives the concentration of Fe in the waste stream as 1.48 mg Fe/L

Practice Exercise 10.5

The concentration of Cu^{2+} in a sample can be determined by reacting it with the ligand cuprizone and measuring its absorbance at 606 nm in a 1.00-cm cell. When a 5.00-mL sample is treated with cuprizone and diluted to 10.00 mL, the resulting solution has an absorbance of 0.118. A second 5.00-mL sample is mixed with 1.00 mL of a 20.00 mg/L standard of Cu^{2+} , treated with cuprizone and diluted to 10.00 mL, giving an absorbance of 0.162. Report the mg Cu^{2+} /L in the sample.

Click [here](#) to review your answer to this exercise.

QUANTITATIVE ANALYSIS OF MIXTURES

Suppose we need to determine the concentration of two analytes, X and Y, in a sample. If each analyte has a wavelength where the other analyte does not absorb, then we can proceed using the approach in [Example 10.5](#). Unfortunately, UV/Vis absorption bands are so broad that it frequently is not possible to find suitable wavelengths. Because Beer's law is additive the mixture's absorbance, A_{mix} , is

$$(A_{\text{mix}})_{\lambda_1} = (\epsilon_X)_{\lambda_1} b C_X + (\epsilon_Y)_{\lambda_1} b C_Y \quad 10.11$$

where λ_1 is the wavelength at which we measure the absorbance. Because equation 10.11 includes terms for the concentration of both X and Y, the absorbance at one wavelength does not provide enough information to determine either C_X or C_Y . If we measure the absorbance at a second wavelength

$$(A_{\text{mix}})_{\lambda_2} = (\epsilon_X)_{\lambda_2} bC_X + (\epsilon_Y)_{\lambda_2} bC_Y \quad 10.12$$

then C_X and C_Y can be determined by solving simultaneously equation 10.11 and equation 10.12. Of course, we also must determine the value for ϵ_X and ϵ_Y at each wavelength. For a mixture of n components, we must measure the absorbance at n different wavelengths.

Example 10.6

The concentrations of Fe^{3+} and Cu^{2+} in a mixture can be determined following their reaction with hexacyanoruthenate (II), $\text{Ru}(\text{CN})_6^{4-}$, which forms a purple-blue complex with Fe^{3+} ($\lambda_{\text{max}} = 550 \text{ nm}$) and a pale-green complex with Cu^{2+} ($\lambda_{\text{max}} = 396 \text{ nm}$).⁷ The molar absorptivities ($\text{M}^{-1} \text{cm}^{-1}$) for the metal complexes at the two wavelengths are summarized in the following table.

	ϵ_{550}	ϵ_{396}
Fe^{3+}	9970	84
Cu^{2+}	34	856

When a sample containing Fe^{3+} and Cu^{2+} is analyzed in a cell with a path-length of 1.00 cm, the absorbance at 550 nm is 0.183 and the absorbance at 396 nm is 0.109. What are the molar concentrations of Fe^{3+} and Cu^{2+} in the sample?

SOLUTION

Substituting known values into equations 10.11 and 10.12 gives

$$A_{550} = 0.183 = 9970C_{\text{Fe}} + 34C_{\text{Cu}}$$

$$A_{396} = 0.109 = 84C_{\text{Fe}} + 856C_{\text{Cu}}$$

To determine C_{Fe} and C_{Cu} we solve the first equation for C_{Cu}

$$C_{\text{Cu}} = \frac{0.183 - 9970C_{\text{Fe}}}{34}$$

and substitute the result into the second equation.

$$0.109 = 84C_{\text{Fe}} + 856 \times \frac{0.183 - 9970C_{\text{Fe}}}{34} = 4.607 - (2.51 \times 10^5)C_{\text{Fe}}$$

Solving for C_{Fe} gives the concentration of Fe^{3+} as $1.79 \times 10^{-5} \text{ M}$. Substituting this concentration back into the equation for the mixture's absorbance at 396 nm gives the concentration of Cu^{2+} as $1.26 \times 10^{-4} \text{ M}$.

Another approach is to multiply the first equation by 856/34 giving

$$4.607 = 251009C_{\text{Fe}} + 856C_{\text{Cu}}$$

Subtracting the second equation from this equation

$$4.607 = 251009C_{\text{Fe}} + 856C_{\text{Cu}}$$

$$-0.109 = 84C_{\text{Fe}} + 856C_{\text{Cu}}$$

$$4.498 = 250925C_{\text{Fe}}$$

we find that C_{Fe} is 1.79×10^{-5} . Having determined C_{Fe} we can substitute back into one of the other equations to solve for C_{Cu} , which is 1.26×10^{-4} .

⁷ DiTusa, M. R.; Schlit, A. A. *J. Chem. Educ.* **1985**, *62*, 541–542.

Practice Exercise 10.6

The absorbance spectra for Cr^{3+} and Co^{2+} overlap significantly. To determine the concentration of these analytes in a mixture, its absorbance was measured at 400 nm and at 505 nm, yielding values of 0.336 and 0.187, respectively. The individual molar absorptivities ($\text{M}^{-1} \text{cm}^{-1}$) are

	ϵ_{400}	ϵ_{505}
Cr^{3+}	15.2	0.533
Co^{2+}	5.60	5.07

Click [here](#) to review your answer to this exercise.

For example, in [Example 10.6](#) the molar absorptivity for Fe^{3+} at 550 nm is $119\times$ that for Cu^{2+} , and the molar absorptivity for Cu^{2+} at 396 nm is $10.2\times$ that for Fe^{3+} .

To obtain results with good accuracy and precision the two wavelengths should be selected so that $\epsilon_X > \epsilon_Y$ at one wavelength and $\epsilon_X < \epsilon_Y$ at the other wavelength. It is easy to appreciate why this is true. Because the absorbance at each wavelength is dominated by one analyte, any uncertainty in the concentration of the other analyte has less of an impact. Figure 10.35 shows that the choice of wavelengths for Practice Exercise 10.6 are reasonable. When the choice of wavelengths is not obvious, one method for locating the optimum wavelengths is to plot ϵ_X/ϵ_Y as function of wavelength, and determine the wavelengths where ϵ_X/ϵ_Y reaches maximum and minimum values.⁸

When the analyte's spectra overlap severely, such that $\epsilon_X \approx \epsilon_Y$ at all wavelength, other computational methods may provide better accuracy and precision. In a multiwavelength linear regression analysis, for example, a mixture's absorbance is compared to that for a set of standard solutions at several wavelengths.⁹ If A_{SX} and A_{SY} are the absorbance values for standard solutions of components X and Y at any wavelength, then

⁸ Mehra, M. C.; Rioux, J. J. *Chem. Educ.* **1982**, *59*, 688–689.

⁹ Blanco, M.; Iturriaga, H.; Maspocho, S.; Tarin, P. J. *Chem. Educ.* **1989**, *66*, 178–180.

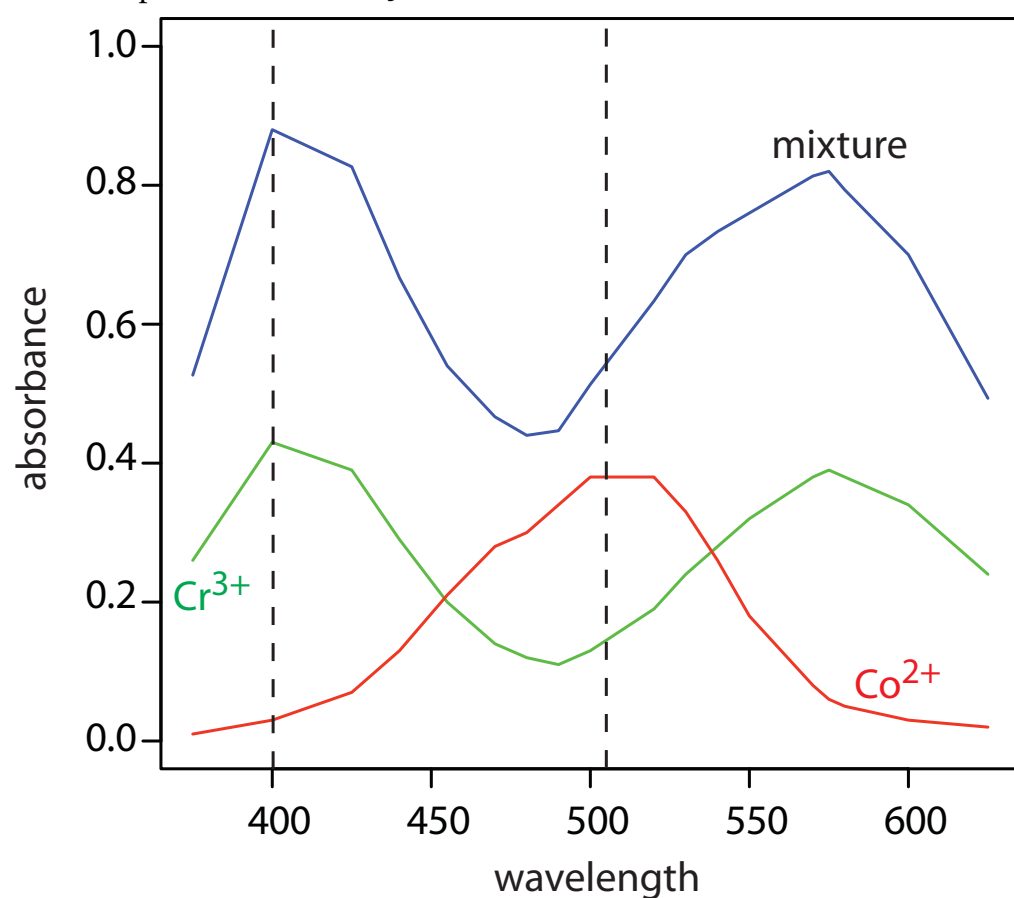


Figure 10.35 Visible absorption spectra for 0.0250 M Cr^{3+} , 0.0750 M Co^{2+} , and for a mixture of Cr^{3+} and Co^{2+} . The two wavelengths used for analyzing the mixture of Cr^{3+} and Co^{2+} are shown by the dashed lines. The data for the two standard solutions are from reference 7.

$$A_{SX} = \epsilon_X b C_{SX} \quad 10.13$$

$$A_{SY} = \epsilon_Y b C_{SY} \quad 10.14$$

where C_{SX} and C_{SY} are the known concentrations of X and Y in the standard solutions. Solving equation 10.13 and equation 10.14 for ϵ_X and ϵ_Y , substituting into [equation 10.11](#), and rearranging, gives

$$\frac{A_{\text{mix}}}{A_{SX}} = \frac{C_X}{C_{SX}} + \frac{C_Y}{C_{SY}} \times \frac{A_{SY}}{A_{SX}}$$

To determine C_X and C_Y the mixture's absorbance and the absorbances of the standard solutions are measured at several wavelengths. Graphing A_{mix}/A_{SX} versus A_{SY}/A_{SX} gives a straight line with a slope of C_Y/C_{SY} and a y -intercept of C_X/C_{SX} . This approach is particularly helpful when it is not possible to find wavelengths where $\epsilon_X > \epsilon_Y$ and $\epsilon_X < \epsilon_Y$.

The approach outlined here for a multi-wavelength linear regression uses a single standard solution for each analyte. A more rigorous approach uses multiple standards for each analyte. The math behind the analysis of this data—what we call a multiple linear regression—is beyond the level of this text. For more details about multiple linear regression see Brereton, R. G. *Chemometrics: Data Analysis for the Laboratory and Chemical Plant*, Wiley: Chichester, England, 2003.

Example 10.7

[Figure 10.35](#) shows visible absorbance spectra for a standard solution of 0.0250 M Cr^{3+} , a standard solution of 0.0750 M Co^{2+} , and a mixture containing unknown concentrations of each ion. The data for these spectra are shown here.¹⁰

λ (nm)	A_{Cr}	A_{Co}	A_{mix}	λ (nm)	A_{Cr}	A_{Co}	A_{mix}
375	0.26	0.01	0.53	520	0.19	0.38	0.63
400	0.43	0.03	0.88	530	0.24	0.33	0.70
425	0.39	0.07	0.83	540	0.28	0.26	0.73
440	0.29	0.13	0.67	550	0.32	0.18	0.76
455	0.20	0.21	0.54	570	0.38	0.08	0.81
470	0.14	0.28	0.47	575	0.39	0.06	0.82
480	0.12	0.30	0.44	580	0.38	0.05	0.79
490	0.11	0.34	0.45	600	0.34	0.03	0.70
500	0.13	0.38	0.51	625	0.24	0.02	0.49

Use a multiwavelength regression analysis to determine the composition of the unknown.

SOLUTION

First we need to calculate values for A_{mix}/A_{SX} and for A_{SY}/A_{SX} . Let's define X as Co^{2+} and Y as Cr^{3+} . For example, at a wavelength of 375 nm A_{mix}/A_{SX} is 0.53/0.01, or 53 and A_{SY}/A_{SX} is 0.26/0.01, or 26. Completing the calculation for all wavelengths and graphing A_{mix}/A_{SX} versus A_{SY}/A_{SX}

¹⁰ The data for the two standards are from Brewer, S. *Solving Problems in Analytical Chemistry*, John Wiley & Sons: New York, 1980.

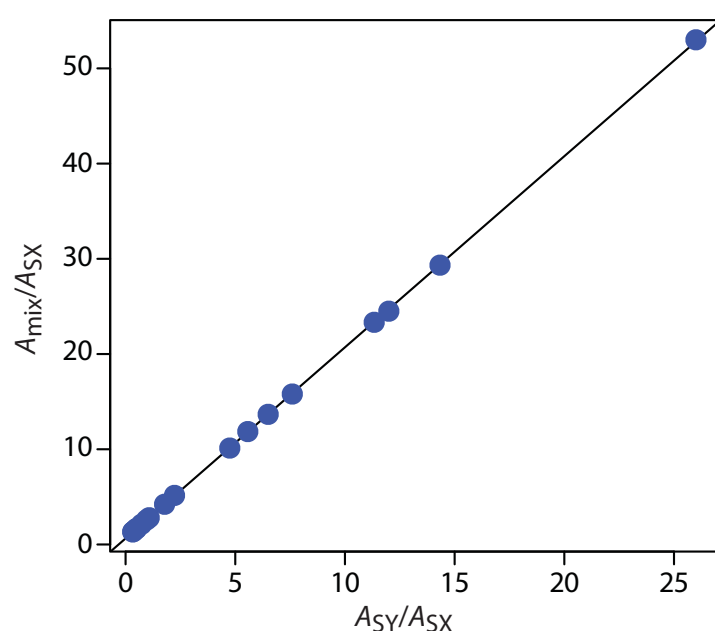


Figure 10.36 Multiwavelength linear regression analysis for the data in Example 10.7.

gives the result shown in Figure 10.36. Fitting a straight-line to the data gives a regression model of

$$\frac{A_{\text{mix}}}{A_{\text{SX}}} = 0.636 + 2.01 \times \frac{A_{\text{SY}}}{A_{\text{SX}}}$$

Using the y -intercept, the concentration of Co^{2+} is

$$\frac{C_{\text{X}}}{C_{\text{SX}}} = \frac{C_{\text{Co}}}{0.0750 \text{ M}} = 0.636$$

or $C_{\text{Co}} = 0.048 \text{ M}$, and using the slope the concentration of Cr^{3+} is

$$\frac{C_{\text{Y}}}{C_{\text{SY}}} = \frac{C_{\text{Cr}}}{0.0250 \text{ M}} = 2.01$$

or $C_{\text{Cr}} = 0.050 \text{ M}$.

Practice Exercise 10.7

A mixture of MnO_4^- and $\text{Cr}_2\text{O}_7^{2-}$, and standards of 0.10 mM KMnO_4 and of $0.10 \text{ mM K}_2\text{Cr}_2\text{O}_7$ gives the results shown in the following table. Determine the composition of the mixture. The data for this problem is from Blanco, M. C.; Iturriaga, H.; Maspocho, S.; Tarin, P. J. *Chem. Educ.* **1989**, *66*, 178–180.

λ (nm)	A_{Mn}	A_{Cr}	A_{mix}
266	0.042	0.410	0.766
288	0.082	0.283	0.571
320	0.168	0.158	0.422
350	0.125	0.318	0.672
360	0.056	0.181	0.366

Click [here](#) to review your answer to this exercise.

10C.3 Qualitative Applications

As discussed earlier in [Section 10B.1](#), ultraviolet, visible, and infrared absorption bands result from the absorption of electromagnetic radiation by specific valence electrons or bonds. The energy at which the absorption occurs, and its intensity is determined by the chemical environment of the absorbing moiety. For example, benzene has several ultraviolet absorption bands due to $\pi \rightarrow \pi^*$ transitions. The position and intensity of two of these bands, 203.5 nm ($\epsilon = 7400 \text{ M}^{-1} \text{ cm}^{-1}$) and 254 nm ($\epsilon = 204 \text{ M}^{-1} \text{ cm}^{-1}$), are sensitive to substitution. For benzoic acid, in which a carboxylic acid group replaces one of the aromatic hydrogens, the two bands shift to 230 nm ($\epsilon = 11600 \text{ M}^{-1} \text{ cm}^{-1}$) and 273 nm ($\epsilon = 970 \text{ M}^{-1} \text{ cm}^{-1}$). A variety of

There are many additional ways to analyze mixtures spectrophotometrically, including generalized standard additions, H-point standard additions, principal component regression to name a few. Consult the chapter's [additional resources](#) for further information.

rules have been developed to aid in correlating UV/Vis absorption bands to chemical structure. Similar correlations have been developed for infrared absorption bands. For example a carbonyl's C=O stretch is sensitive to adjacent functional groups, occurring at 1650 cm^{-1} for acids, 1700 cm^{-1} for ketones, and 1800 cm^{-1} for acid chlorides. The interpretation of UV/Vis and IR spectra receives adequate coverage elsewhere in the chemistry curriculum, notably in organic chemistry, and is not considered further in this text.

With the availability of computerized data acquisition and storage it is possible to build digital libraries of standard reference spectra. The identity of an unknown compound can often be determined by comparing its spectrum against a library of reference spectra, a process is known as **SPECTRAL SEARCHING**. Comparisons are made using an algorithm that calculates the cumulative difference between the sample's spectrum and a reference spectrum. For example, one simple algorithm uses the following equation

$$D = \sum_{i=1}^n |(A_{\text{sample}})_i - (A_{\text{reference}})_i|$$

where D is the cumulative difference, A_{sample} is the sample's absorbance at wavelength or wavenumber i , $A_{\text{reference}}$ is the absorbance of the reference compound at the same wavelength or wavenumber, and n is the number of digitized points in the spectra. The cumulative difference is calculated for each reference spectrum. The reference compound with the smallest value of D provides the closest match to the unknown compound. The accuracy of spectral searching is limited by the number and type of compounds included in the library, and by the effect of the sample's matrix on the spectrum.

Another advantage of computerized data acquisition is the ability to subtract one spectrum from another. When coupled with spectral searching it may be possible, by repeatedly searching and subtracting reference spectra, to determine the identity of several components in a sample without the need of a prior separation step. An example is shown in [Figure 10.37](#) in which the composition of a two-component mixture is determined by successive searching and subtraction. Figure 10.37a shows the spectrum of the mixture. A search of the spectral library selects cocaine·HCl (Figure 10.37b) as a likely component of the mixture. Subtracting the reference spectrum for cocaine·HCl from the mixture's spectrum leaves a result (Figure 10.37c) that closely matches mannitol's reference spectrum (Figure 10.37d). Subtracting the reference spectrum for leaves only a small residual signal (Figure 10.37e).

10C.4 Characterization Applications

Molecular absorption, particularly in the UV/Vis range, has been used for a variety of different characterization studies, including determining the

IR spectra traditionally are displayed using percent transmittance, %T, along the y -axis (for example, see Figure 10.16). Because absorbance—not percent transmittance—is a linear function of concentration, spectral searching and spectral subtraction, is easier to do when displaying absorbance on the y -axis.

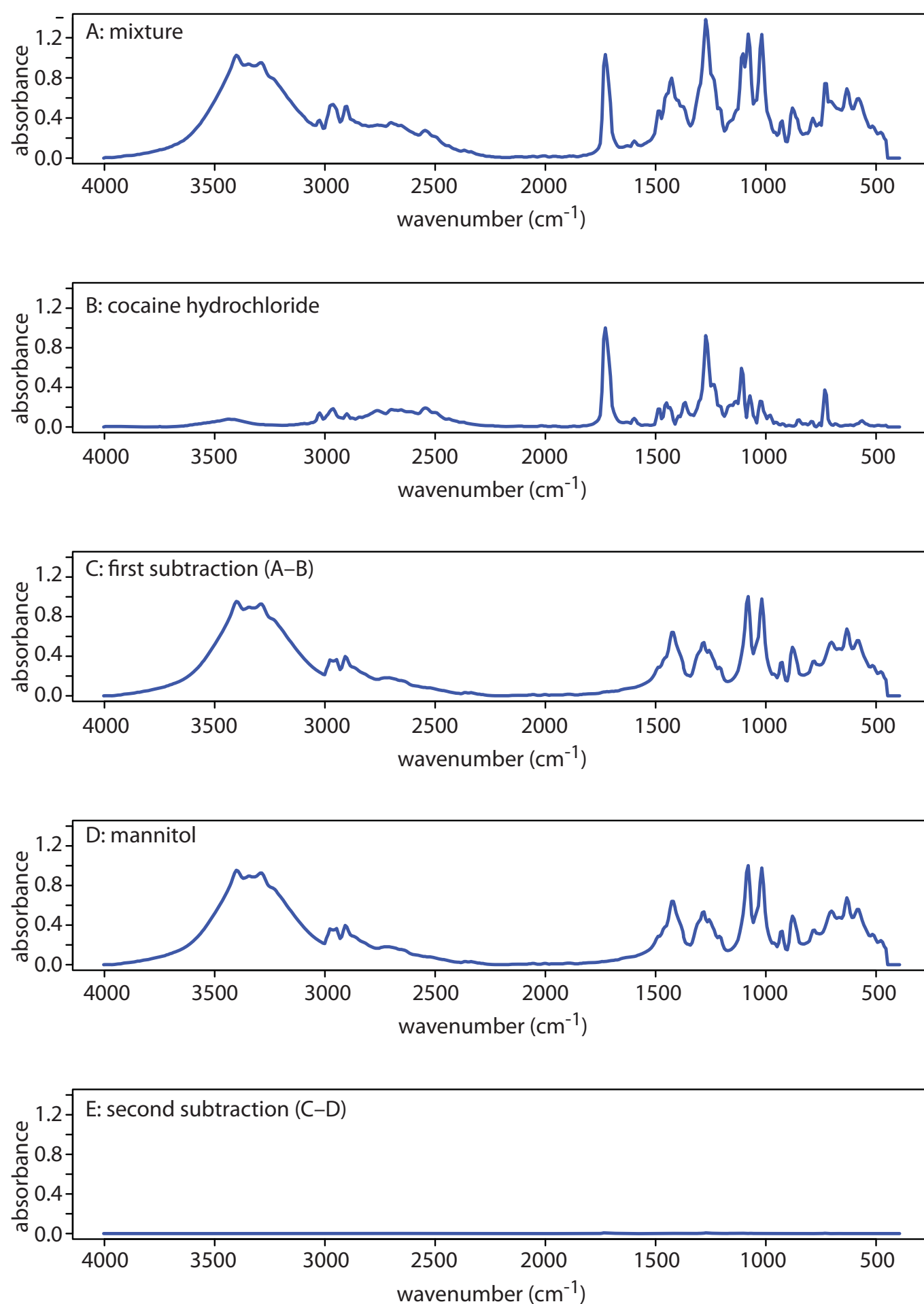
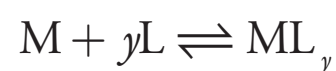


Figure 10.37 Identifying the components of a mixture by spectral searching and subtracting. (a) IR spectrum of the mixture; (b) Reference IR spectrum of cocaine · HCl; (c) Result of subtracting the spectrum of cocaine · HCl from the mixture's spectrum; (d) Reference IR spectrum of mannitol; and (e) The residual spectrum after removing mannitol's contribution to the mixture's spectrum.

stoichiometry of metal–ligand complexes and determining equilibrium constants. Both of these examples are examined in this section.

STOICHIOMETRY OF A METAL-LIGAND COMPLEX

We can determine the stoichiometry of a metal–ligand complexation reaction



using one of three methods: the method of continuous variations, the mole-ratio method, and the slope-ratio method. Of these approaches, the

METHOD OF CONTINUOUS VARIATIONS, also called Job's method, is the most popular. In this method a series of solutions is prepared such that the total moles of metal and ligand, n_{total} , in each solution is the same. If $(n_{\text{M}})_i$ and $(n_{\text{L}})_i$ are, respectively, the moles of metal and ligand in solution i , then

$$n_{\text{total}} = (n_{\text{M}})_i + (n_{\text{L}})_i$$

The relative amount of ligand and metal in each solution is expressed as the mole fraction of ligand, $(X_{\text{L}})_i$, and the mole fraction of metal, $(X_{\text{M}})_i$,

$$(X_{\text{L}})_i = \frac{(n_{\text{L}})_i}{n_{\text{total}}}$$

$$(X_{\text{M}})_i = 1 - \frac{(n_{\text{L}})_i}{n_{\text{total}}} = \frac{(n_{\text{M}})_i}{n_{\text{total}}}$$

The concentration of the metal–ligand complex in any solution is determined by the limiting reagent, with the greatest concentration occurring when the metal and the ligand are mixed stoichiometrically. If we monitor the complexation reaction at a wavelength where only the metal–ligand complex absorbs, a graph of absorbance versus the mole fraction of ligand will have two linear branches—one when the ligand is the limiting reagent and a second when the metal is the limiting reagent. The intersection of these two branches represents a stoichiometric mixing of the metal and the ligand. We can use the mole fraction of ligand at the intersection to determine the value of y for the metal–ligand complex ML_y .

$$y = \frac{n_{\text{L}}}{n_{\text{M}}} = \frac{X_{\text{L}}}{X_{\text{M}}} = \frac{X_{\text{L}}}{1 - X_{\text{L}}}$$

You also can plot the data as absorbance versus the mole fraction of metal. In this case, y is equal to $(1 - X_{\text{M}})/X_{\text{M}}$.

Example 10.8

To determine the formula for the complex between Fe^{2+} and *o*-phenanthroline, a series of solutions is prepared in which the total concentration of metal and ligand is held constant at 3.15×10^{-4} M. The absorbance of each solution is measured at a wavelength of 510 nm. Using the following data, determine the formula for the complex.

X_{L}	absorbance	X_{L}	absorbance
0.000	0.000	0.600	0.693
0.100	0.116	0.700	0.809
0.200	0.231	0.800	0.693
0.300	0.347	0.900	0.347
0.400	0.462	1.000	0.000
0.500	0.578		

To prepare the solutions for this example I first prepared a solution of 3.15×10^{-4} M Fe^{2+} and a solution of 3.15×10^{-4} M *o*-phenanthroline. Because the two stock solutions are of equal concentration, diluting a portion of one solution with the other solution gives a mixture in which the combined concentration of *o*-phenanthroline and Fe^{2+} is 3.15×10^{-4} M. If each solution has the same volume, then each solution contains the same total moles of metal and ligand.

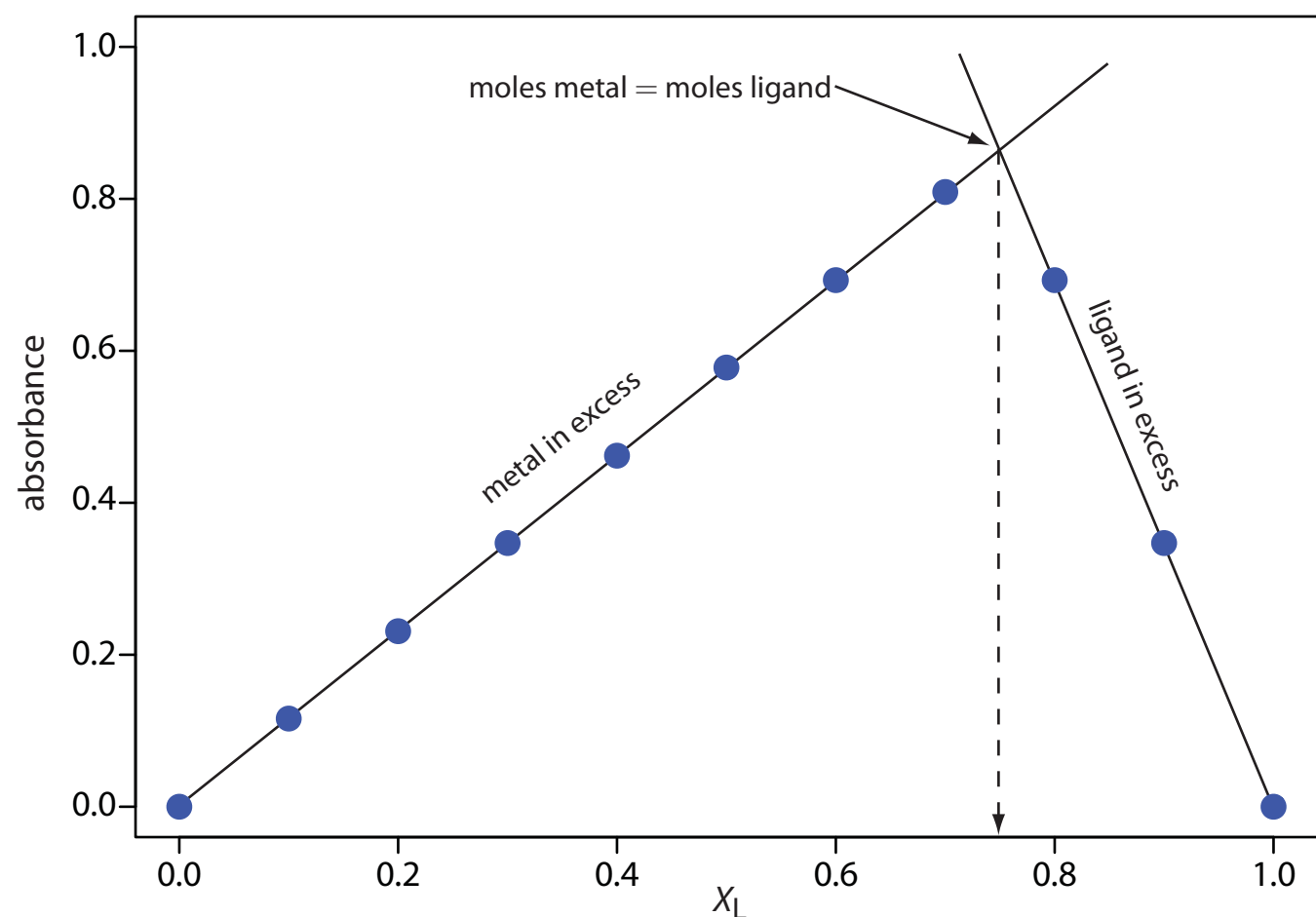


Figure 10.38 Continuous variations plot for Example 10.8. The photo shows the solutions used in gathering the data. Each solution is displayed directly below its corresponding point on the continuous variations plot.



SOLUTION

A plot of absorbance versus the mole fraction of ligand is shown in Figure 10.38. To find the maximum absorbance, we extrapolate the two linear portions of the plot. The two lines intersect at a mole fraction of ligand of 0.75. Solving for y gives

$$y = \frac{X_L}{1 - X_L} = \frac{0.75}{1 - 0.75} = 3$$

The formula for the metal–ligand complex is $\text{Fe}(o\text{-phenanthroline})_3^{2+}$.

Practice Exercise 10.8

Use the continuous variations data in the following table to determine the formula for the complex between Fe^{2+} and SCN^- . The data for this problem is adapted from Meloun, M.; Havel, J.; Högfeltdt, E. *Computation of Solution Equilibria*, Ellis Horwood: Chichester, England, 1988, p. 236.

X_L	absorbance	X_L	absorbance	X_L	absorbance	X_L	absorbance
0.0200	0.068	0.2951	0.670	0.5811	0.790	0.8923	0.325
0.0870	0.262	0.3887	0.767	0.6860	0.701	0.9787	0.071
0.1792	0.471	0.4964	0.807	0.7885	0.540		

Click [here](#) to review your answer to this exercise.

Several precautions are necessary when using the method of continuous variations. First, the metal and the ligand must form only one metal–ligand complex. To determine if this condition is true, plots of absorbance versus X_L are constructed at several different wavelengths and for several different values of n_{total} . If the maximum absorbance does not occur at the same value of X_L for each set of conditions, then more than one metal–ligand complex must be present. A second precaution is that the metal–ligand complex's absorbance must obey Beer's law. Third, if the metal–ligand complex's formation constant is relatively small, a plot of absorbance versus X_L may show significant curvature. In this case it is often difficult to determine the stoichiometry by extrapolation. Finally, because the stability of a metal–ligand complex may be influenced by solution conditions, the composition of the solutions must be carefully controlled. When the ligand is a weak base, for example, the solutions must be buffered to the same pH.

In the **MOLE-RATIO METHOD** the amount of one reactant, usually the moles of metal, is held constant, while the amount of the other reactant is varied. The absorbance is monitored at a wavelength where the metal–ligand complex absorbs. A plot of absorbance as a function of the ligand-to-metal mole ratio, n_L/n_M , has two linear branches, which intersect at a mole–ratio corresponding to the complex's formula. Figure 10.39a shows a mole-ratio plot for the formation of a 1:1 complex in which the absorbance is monitored at a wavelength where only the complex absorbs. Figure 10.39b shows a mole-ratio plot for a 1:2 complex in which all three species—the metal, the ligand, and the complex—absorb at the selected wavelength. Unlike the method of continuous variations, the mole-ratio method can be used for complexation reactions that occur in a stepwise fashion if there is a difference in the molar absorptivities of the metal–ligand complexes, and if the formation constants are sufficiently different. A typical mole-ratio plot for the step-wise formation of ML and ML_2 is shown in Figure 10.39c.

In both the method of continuous variations and the mole-ratio method we determine the complex's stoichiometry by extrapolating absorbance data

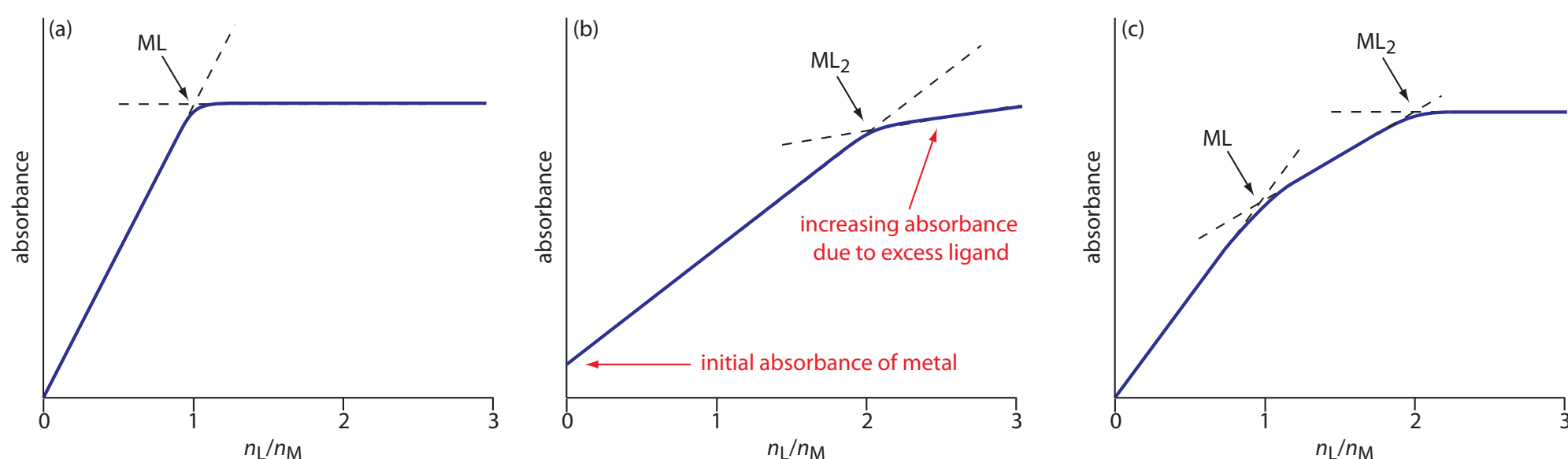


Figure 10.39 Mole-ratio plots for: (a) a 1:1 metal–ligand complex in which only the complex absorbs; (b) a 1:2 metal–ligand complex in which the metal, the ligand, and the complex absorb; and (c) the stepwise formation of a 1:1 and a 1:2 metal–ligand complex.

from conditions in which there is a linear relationship between absorbance and the relative amounts of metal and ligand. If a metal–ligand complex is very weak, a plot of absorbance versus X_L or n_L/n_M may be so curved that it is impossible to determine the stoichiometry by extrapolation. In this case the slope-ratio may be used.

In the **SLOPE-RATIO METHOD** two sets of solutions are prepared. The first set of solutions contains a constant amount of metal and a variable amount of ligand, chosen such that the total concentration of metal, C_M , is much larger than the total concentration of ligand, C_L . Under these conditions we may assume that essentially all the ligand reacts in forming the metal–ligand complex. The concentration of the complex, which has the general form M_xL_y , is

$$[M_xL_y] = \frac{C_L}{y}$$

If we monitor the absorbance at a wavelength where only M_xL_y absorbs, then

$$A = \varepsilon b [M_xL_y] = \frac{\varepsilon b C_L}{y}$$

and a plot of absorbance versus C_L is linear with a slope, s_L , of

$$s_L = \frac{\varepsilon b}{y}$$

A second set of solutions is prepared with a fixed concentration of ligand that is much greater than a variable concentration of metal; thus

$$[M_xL_y] = \frac{C_M}{x}$$

$$A = \varepsilon b [M_xL_y] = \frac{\varepsilon b C_M}{x}$$

$$s_M = \frac{\varepsilon b}{x}$$

A ratio of the slopes provides the relative values of x and y .

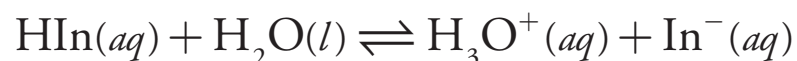
$$\frac{s_M}{s_L} = \frac{\varepsilon b / x}{\varepsilon b / y} = \frac{y}{x}$$

An important assumption in the slope-ratio method is that the complexation reaction continues to completion in the presence of a sufficiently large excess of metal or ligand. The slope-ratio method also is limited to

systems in which only a single complex is formed and for which Beer's law is obeyed.

DETERMINATION OF EQUILIBRIUM CONSTANTS

Another important application of molecular absorption spectroscopy is the determination of equilibrium constants. Let's consider, as a simple example, an acid–base reaction of the general form



where HIn and In[−] are the conjugate weak acid and weak base forms of an acid–base indicator. The equilibrium constant for this reaction is

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{In}^-]}{[\text{HIn}]}$$

To determine the equilibrium constant's value, we prepare a solution in which the reaction is in a state of equilibrium and determine the equilibrium concentration of H₃O⁺, HIn, and In[−]. The concentration of H₃O⁺ is easy to determine by simply measuring the solution's pH. To determine the concentration of HIn and In[−] we can measure the solution's absorbance.

If both HIn and In[−] absorb at the selected wavelength, then, from [equation 10.6](#), we know that

$$A = \epsilon_{\text{HIn}} b[\text{HIn}] + \epsilon_{\text{In}^-} b[\text{In}^-] \quad 10.15$$

where ϵ_{HIn} and ϵ_{In^-} are the molar absorptivities for HIn and In[−]. The total concentration of indicator, C , is given by a mass balance equation

$$C = [\text{HIn}] + [\text{In}^-] \quad 10.16$$

Solving equation 10.16 for [HIn] and substituting into equation 10.15 gives

$$A = \epsilon_{\text{HIn}} b(C - [\text{In}^-]) + \epsilon_{\text{In}^-} b[\text{In}^-]$$

which we simplify to

$$A = \epsilon_{\text{HIn}} bC - \epsilon_{\text{HIn}} b[\text{In}^-] + \epsilon_{\text{In}^-} b[\text{In}^-]$$

$$A = A_{\text{HIn}} + b[\text{In}^-](\epsilon_{\text{In}^-} - \epsilon_{\text{HIn}}) \quad 10.17$$

where A_{HIn} , which is equal to $\epsilon_{\text{HIn}} bC$, is the absorbance when the pH is acidic enough that essentially all the indicator is present as HIn. Solving equation 10.17 for the concentration of In[−] gives

$$[\text{In}^-] = \frac{A - A_{\text{HIn}}}{b(\epsilon_{\text{In}^-} - \epsilon_{\text{HIn}})} \quad 10.18$$

Proceeding in the same fashion, we can derive a similar equation for the concentration of HIn

$$[\text{HIn}] = \frac{A_{\text{In}} - A}{b(\epsilon_{\text{In}} - \epsilon_{\text{HIn}})} \quad 10.19$$

where A_{In} , which is equal to $\epsilon_{\text{In}}bC$, is the absorbance when the pH is basic enough that only In^- contributes to the absorbance. Substituting [equation 10.18](#) and equation 10.19 into the equilibrium constant expression for HIn gives

$$K_a = [\text{H}_3\text{O}^+] \frac{A - A_{\text{HIn}}}{A_{\text{In}} - A} \quad 10.20$$

We can use equation 10.20 to determine the value of K_a in one of two ways. The simplest approach is to prepare three solutions, each of which contains the same amount, C , of indicator. The pH of one solution is made sufficiently acidic such that $[\text{HIn}] \gg [\text{In}^-]$. The absorbance of this solution gives A_{HIn} . The value of A_{In} is determined by adjusting the pH of the second solution such that $[\text{In}^-] \gg [\text{HIn}]$. Finally, the pH of the third solution is adjusted to an intermediate value, and the pH and absorbance, A , recorded. The value of K_a is calculated using equation 10.20.

Example 10.9

The acidity constant for an acid–base indicator is determined by preparing three solutions, each of which has a total indicator concentration of 5.00×10^{-5} M. The first solution is made strongly acidic with HCl and has an absorbance of 0.250. The second solution was made strongly basic and has an absorbance of 1.40. The pH of the third solution is 2.91 and has an absorbance of 0.662. What is the value of K_a for the indicator?

SOLUTION

The value of K_a is determined by making appropriate substitutions into 10.20; thus

$$K_a = (1.23 \times 10^{-3}) \times \frac{0.662 - 0.250}{1.40 - 0.662} = 6.87 \times 10^{-4}$$

Practice Exercise 10.9

To determine the K_a of a merocyanine dye, the absorbance of a solution of 3.5×10^{-4} M dye was measured at a pH of 2.00, a pH of 6.00, and a pH of 12.00, yielding absorbances of 0.000, 0.225, and 0.680, respectively. What is the value of K_a for this dye? The data for this problem is adapted from Lu, H.; Rutan, S. C. *Anal. Chem.*, **1996**, 68, 1381–1386.

Click [here](#) to review your answer to this exercise.

A second approach for determining K_a is to prepare a series of solutions, each containing the same amount of indicator. Two solutions are used to determine values for A_{HIn} and A_{In^-} . Taking the log of both sides of [equation 10.20](#) and rearranging leave us with the following equation.

$$\log \frac{A - A_{\text{HIn}}}{A_{\text{In}^-} - A} = \text{pH} - \text{p}K_a \quad 10.21$$

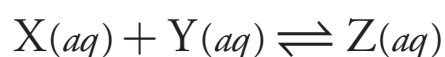
A plot of $\log[(A - A_{\text{HIn}})/(A_{\text{In}^-} - A)]$ versus pH is a straight-line with a slope of +1 and a y -intercept of $-\text{p}K_a$.

Practice Exercise 10.10

To determine the K_a of the indicator bromothymol blue, the absorbance of a series of solutions containing the same concentration of the indicator was measured at pH levels of 3.35, 3.65, 3.94, 4.30, and 4.64, yielding absorbances of 0.170, 0.287, 0.411, 0.562, and 0.670, respectively. Acidifying the first solution to a pH of 2 changes its absorbance to 0.006, and adjusting the pH of the last solution to 12 changes its absorbance to 0.818. What is the value of K_a for this day? The data for this problem is from Patterson, G. S. *J. Chem. Educ.*, **1999**, 76, 395–398.

Click [here](#) to review your answer to this exercise.

In developing these approaches for determining K_a we considered a relatively simple system in which the absorbance of HIn and In^- are easy to measure and for which it is easy to determine the concentration of H_3O^+ . In addition to acid–base reactions, we can adapt these approaches to any reaction of the general form



including metal–ligand complexation reactions and redox reactions, provided that we can determine spectrophotometrically the concentration of the product, Z, and one of the reactants, and that the concentration of the other reactant can be measured by another method. With appropriate modifications, more complicated systems, in which one or more of these parameters can not be measured, also can be treated.¹¹

10C.5 Evaluation of UV/Vis and IR Spectroscopy

SCALE OF OPERATION

Molecular UV/Vis absorption is routinely used for the analysis of trace analytes in macro and meso samples. Major and minor analytes can be determined by diluting the sample before analysis, while concentrating a sample may allow for the analysis of ultratrace analytes. The scale of operations for infrared absorption is generally poorer than that for UV/Vis absorption.

See [Figure 3.5](#) to review the meaning of macro and meso for describing samples, and the meaning of major, minor, and ultratrace for describing analytes.

¹¹ Ramette, R. W. *Chemical Equilibrium and Analysis*, Addison-Wesley: Reading, MA, 1981, Chapter 13.

ACCURACY

Under normal conditions a relative error of 1–5% is easy to obtain with UV/Vis absorption. Accuracy is usually limited by the quality of the blank. Examples of the type of problems that may be encountered include the presence of particulates in a sample that scatter radiation and interferents that react with analytical reagents. In the latter case the interferent may react to form an absorbing species, giving rise to a positive determinate error. Interferents also may prevent the analyte from reacting, leading to a negative determinate error. With care, it may be possible to improve the accuracy of an analysis by as much as an order of magnitude.

PRECISION

In absorption spectroscopy, precision is limited by indeterminate errors—primarily instrumental noise—introduced when measuring absorbance. Precision is generally worse for low absorbances where $P_0 \approx P_T$, and for high absorbances when P_T approaches 0. We might expect, therefore, that precision will vary with transmittance.

We can derive an expression between precision and transmittance by applying the propagation of uncertainty as described in Chapter 4. To do so we rewrite Beer's law as

$$C = -\frac{1}{\epsilon b} \log T \quad 10.22$$

[Table 4.10](#) in Chapter 4 helps us in completing the propagation of uncertainty for equation 10.22, giving the absolute uncertainty in the concentration, s_C , as

$$s_C = -\frac{0.4343}{\epsilon b} \times \frac{s_T}{T} \quad 10.23$$

where s_T is the absolute uncertainty in the transmittance. Dividing equation 10.23 by equation 10.22 gives the relative uncertainty in concentration, s_C/C , as

$$\frac{s_C}{C} = \frac{0.4343 s_T}{T \log T}$$

If we know the absolute uncertainty in transmittance, we can determine the relative uncertainty in concentration for any transmittance.

Determining the relative uncertainty in concentration is complicated because s_T may be a function of the transmittance. As shown in [Table 10.8](#), three categories of indeterminate instrumental error have been observed.¹² A constant s_T is observed for the uncertainty associated with reading %T on a meter's analog or digital scale. Typical values are ± 0.2 – 0.3% (a k_1 of ± 0.002 – 0.003) for an analog scale, and $\pm 0.001\%$ (a k_1 of ± 0.00001) for

¹² Rothman, L. D.; Crouch, S. R.; Ingle, J. D. Jr. *Anal. Chem.* **1975**, *47*, 1226–1233.

Table 10.8 Effect of Indeterminate Errors on Relative Uncertainty in Concentration

Category	Sources of Indeterminate Error	Relative Uncertainty in Concentration
$s_T = k_1$	%T readout resolution noise in thermal detectors	$\frac{s_C}{C} = \frac{0.4343k_1}{T \log T}$
$s_T = k_2 \sqrt{T^2 + T}$	noise in photon detectors	$\frac{s_C}{C} = \frac{0.4343k_2}{\log T} \sqrt{1 + \frac{1}{T}}$
$s_T = k_3 T$	positioning of sample cell fluctuations in source intensity	$\frac{s_C}{C} = \frac{0.4343k_3}{\log T}$

a digital scale. A constant s_T also is observed for the thermal transducers used in infrared spectrophotometers. The effect of a constant s_T on the relative uncertainty in concentration is shown by curve A in Figure 10.40. Note that the relative uncertainty is very large for both high and low absorbances, reaching a minimum when the absorbance is 0.4343. This source of indeterminate error is important for infrared spectrophotometers and for inexpensive UV/Vis spectrophotometers. To obtain a relative uncertainty in concentration of ± 1 –2%, the absorbance must be kept within the range 0.1–1.

Values of s_T are a complex function of transmittance when indeterminate errors are dominated by the noise associated with photon detectors. Curve B in Figure 10.40 shows that the relative uncertainty in concentration is very large for low absorbances, but is less at higher absorbances. Although the relative uncertainty reaches a minimum when the absorbance is 0.963, there is little change in the relative uncertainty for absorbances within the range 0.5–2. This source of indeterminate error generally limits

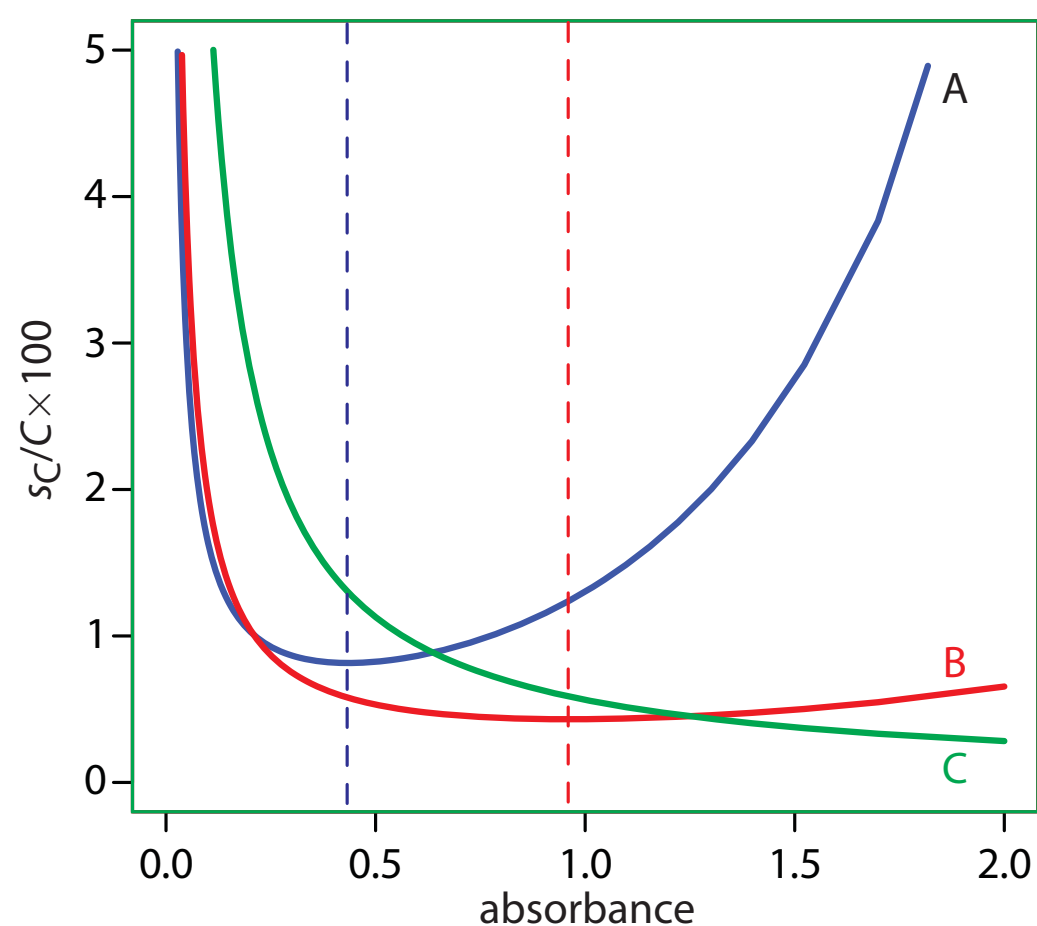


Figure 10.40 Percent relative uncertainty in concentration as a function of absorbance for the categories of indeterminate errors in Table 10.8. A: $k_1 = \pm 0.0030$; B: $k_2 = \pm 0.0030$; C: $k_3 = \pm 0.0130$. The dashed lines correspond to the minimum uncertainty for curve A (absorbance of 0.4343) and for curve B (absorbance of 0.963).

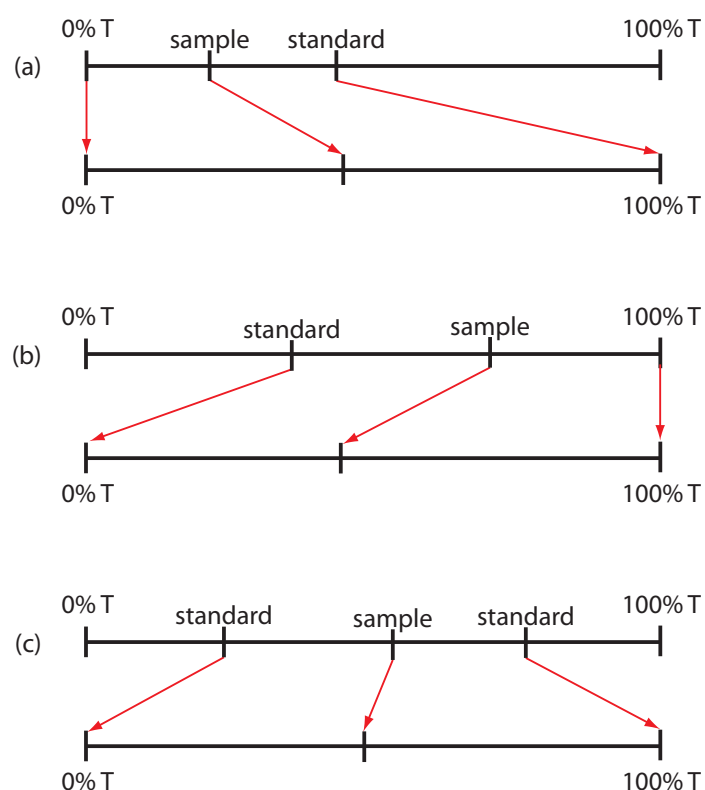


Figure 10.41 Methods for improving the precision of absorption methods: (a) high-absorbance method; (b) low-absorbance method; (c) maximum precision method.

See [Figure 10.24](#) for an example of how the choice of wavelength affects a calibration curve's sensitivity.

the precision of high quality UV/Vis spectrophotometers for mid-to-high absorbances.

Finally, the value of s_T is directly proportional to transmittance for indeterminate errors resulting from fluctuations in the source's intensity and from uncertainty in positioning the sample within the spectrometer. The latter is particularly important because the optical properties of any sample cell are not uniform. As a result, repositioning the sample cell may lead to a change in the intensity of transmitted radiation. As shown by curve C in [Figure 10.40](#), the effect is only important at low absorbances. This source of indeterminate errors is usually the limiting factor for high quality UV/Vis spectrophotometers when the absorbance is relatively small.

When the relative uncertainty in concentration is limited by the %T readout resolution, the precision of the analysis can be improved by redefining 100% T and 0% T. Normally 100% T is established using a blank and 0% T is established while preventing the source's radiation from reaching the detector. If the absorbance is too high, precision can be improved by resetting 100% T using a standard solution of the analyte whose concentration is less than that of the sample ([Figure 10.41a](#)). For a sample whose absorbance is too low, precision can be improved by redefining 0% T using a standard solution of the analyte whose concentration is greater than that of the analyte ([Figure 10.41b](#)). In this case a calibration curve is required because a linear relationship between absorbance and concentration no longer exists. Precision can be further increased by combining these two methods ([Figure 10.41c](#)). Again, a calibration curve is necessary since the relationship between absorbance and concentration is no longer linear.

SENSITIVITY

The sensitivity of a molecular absorption method, which is the slope of a Beer's law calibration curve, is the product of the analyte's absorptivity and the pathlength of the sample cell (ϵb). You can improve a method's sensitivity by selecting a wavelength where absorbance is at a maximum or by increasing the pathlength.

SELECTIVITY

Selectivity is rarely a problem in molecular absorption spectrophotometry. In many cases it is possible to find a wavelength where only the analyte absorbs. When two or more species do contribute to the measured absorbance, a multicomponent analysis is still possible, as shown in [Example 10.6](#) and [Example 10.7](#).

TIME, COST, AND EQUIPMENT

The analysis of a sample by molecular absorption spectroscopy is relatively rapid, although additional time may be required if we need to chemically convert a nonabsorbing analyte into an absorbing form. The cost of UV/

Vis instrumentation ranges from several hundred dollars for a simple filter photometer, to more than \$50,000 for a computer controlled high resolution, double-beam instrument equipped with variable slits, and operating over an extended range of wavelengths. Fourier transform infrared spectrometers can be obtained for as little as \$15,000–\$20,000, although more expensive models are available.

10D Atomic Absorption Spectroscopy

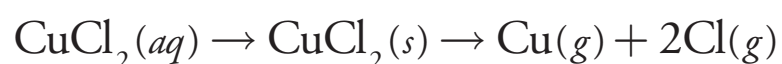
Guystav Kirchoff and Robert Bunsen first used atomic absorption spectroscopy—along with atomic emission—in 1859 and 1860 as a means for identify atoms in flames and hot gases. Although atomic emission continued to develop as an analytical technique, progress in atomic absorption languished for almost a century. Modern atomic absorption spectroscopy has its beginnings in 1955 as a result of the independent work of A. C. Walsh and C. T. J. Alkemade.¹³ Commercial instruments were in place by the early 1960s, and the importance of atomic absorption as an analytical technique was soon evident.

10D.1 Instrumentation

Atomic absorption spectrophotometers use the same single-beam or double-beam optics described earlier for molecular absorption spectrophotometers (see [Figure 10.26](#) and [Figure 10.27](#)). There is, however, an important additional need in atomic absorption spectroscopy—we must convert the analyte into free atoms. In most cases our analyte is in solution form. If our sample is a solid, then we must bring it into solution before the analysis. When analyzing a lake sediment for Cu, Zn, and Fe, for example, we bring the analytes into solution as Cu^{2+} , Zn^{2+} , and Fe^{3+} by extracting them with a suitable reagent. For this reason, only the introduction of solution samples is considered in this text.

ATOMIZATION

The process of converting an analyte to a free gaseous atom is called **ATOMIZATION**. Converting an aqueous analyte into a free atom requires that we strip away the solvent, volatilize the analytes, and, if necessary, dissociate the analyte into free atoms. Desolvating an aqueous solution of CuCl_2 , for example, leaves us with solid particulates of CuCl_2 . Converting the particulate CuCl_2 to gas phases atoms of Cu and Cl requires thermal energy.

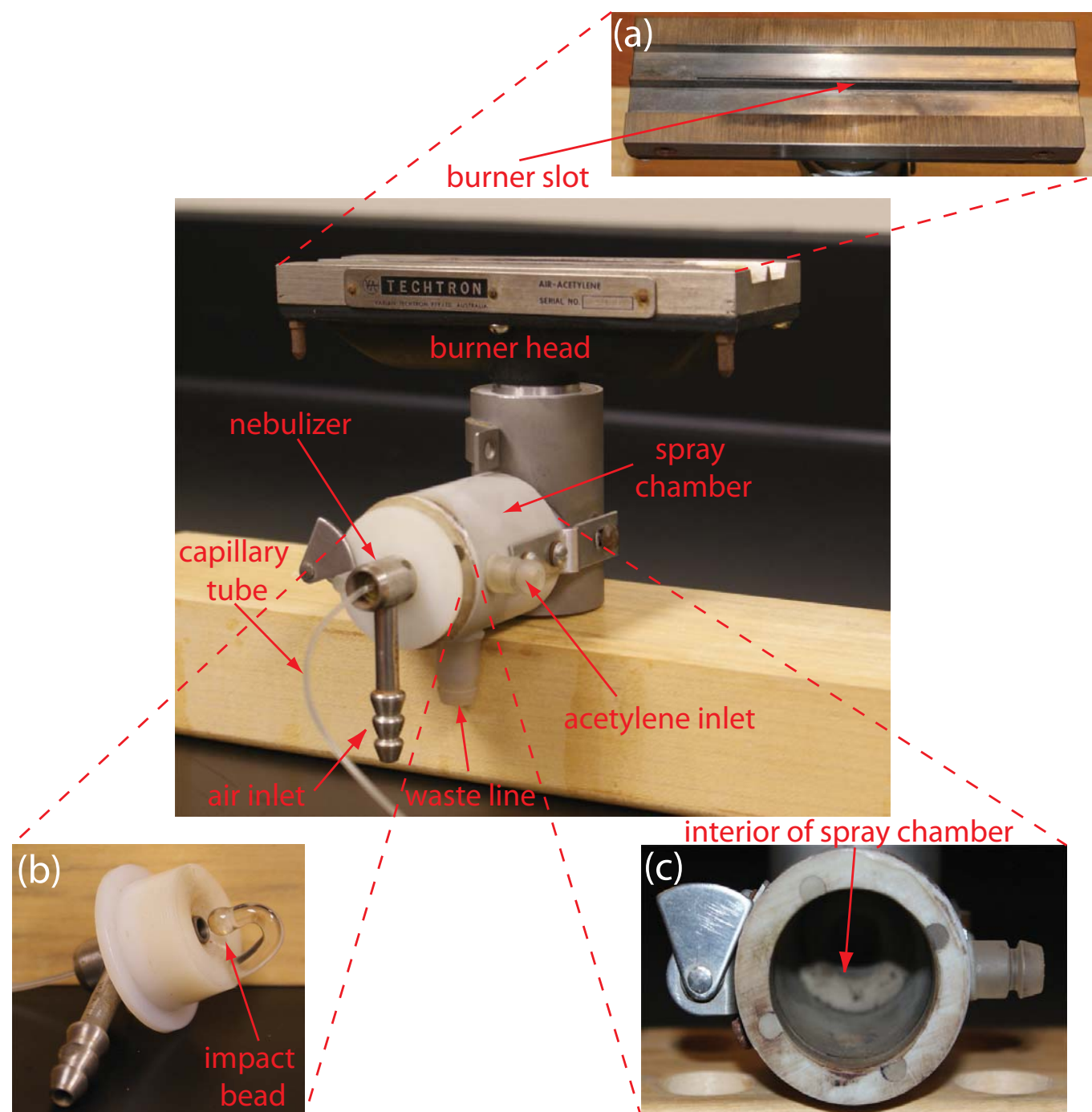


There are two common atomization methods: flame atomization and electrothermal atomization, although a few elements are atomized using other methods.

¹³ (a) Walsh, A. *Anal. Chem.* **1991**, *63*, 933A–941A; (b) Koirtzmann, S. R. *Anal. Chem.* **1991**, *63*, 1024A–1031A; (c) Slavin, W. *Anal. Chem.* **1991**, *63*, 1033A–1038A.

What reagent we choose to use depends on our research goals. If we need to know the total amount of metal in the sediment, then we might use a microwave digestion using a mixture of concentrated acids, such as HNO_3 , HCl , and HF . This destroys the sediment's matrix and brings everything into solution. On the other hand, if our interest is biologically available metals, we might extract the sample under milder conditions, such as a dilute solution of HCl or CH_3COOH at room temperature.

Figure 10.42 Flame atomization assembly with expanded views of (a) the burner head showing the burner slot where the flame is located; (b) the nebulizer's impact bead; and (c) the interior of the spray chamber. Although the unit shown here is from an older instrument, the basic components of a modern flame AA spectrometer are the same.



FLAME ATOMIZER

Figure 10.42 shows a typical flame atomization assembly with close-up views of several key components. In the unit shown here, the aqueous sample is drawn into the assembly by passing a high-pressure stream of compressed air past the end of a capillary tube immersed in the sample. When the sample exits the nebulizer it strikes a glass impact bead, converting it into a fine aerosol mist within the spray chamber. The aerosol mist is swept through the spray chamber by the combustion gases—compressed air and acetylene in this case—to the burner head where the flame's thermal energy desolvates the aerosol mist to a dry aerosol of small, solid particles. The flame's thermal energy then volatilizes the particles, producing a vapor consisting of molecular species, ionic species, and free atoms.

Burner. The slot burner in Figure 10.42a provides a long optical path-length and a stable flame. Because absorbance increases linearly with the path length, a long path length provides greater sensitivity. A stable flame minimizes uncertainty due to fluctuations in the flame.

The burner is mounted on an adjustable stage that allows the entire assembly to move horizontally and vertically. Horizontal adjustments ensure that the flame is aligned with the instrument's optical path. Vertical

Compressed air is one of the two gases whose combustion produces the flame.

adjustments adjust the height within the flame from which absorbance is monitored. This is important because two competing processes affect the concentration of free atoms in the flame. The more time the analyte spends in the flame the greater the atomization efficiency; thus, the production of free atoms increases with height. On the other hand, a longer residence time allows more opportunity for the free atoms to combine with oxygen to form a molecular oxide. For an easily oxidized metal, such as Cr, the concentration of free atoms is greatest just above the burner head. For metals, such as Ag, which are difficult to oxidize, the concentration of free atoms increases steadily with height (Figure 10.43). Other atoms show concentration profiles that maximize at a characteristic height.

Flame. The flame's temperature, which affects the efficiency of atomization, depends on the fuel–oxidant mixture, several examples of which are listed in Table 10.9. Of these, the air–acetylene and the nitrous oxide–acetylene flames are the most popular. Normally the fuel and oxidant are mixed in an approximately stoichiometric ratio; however, a fuel-rich mixture may be necessary for easily oxidized analytes.

Figure 10.44 shows a cross-section through the flame, looking down the source radiation's optical path. The primary combustion zone is usually rich in gas combustion products that emit radiation, limiting its usefulness for atomic absorption. The interzonal region generally is rich in free atoms and provides the best location for measuring atomic absorption. The hottest part of the flame is typically 2–3 cm above the primary combustion zone. As atoms approach the flame's secondary combustion zone, the decrease in temperature allows for formation of stable molecular species.

Sample Introduction. The most common means for introducing samples into a flame atomizer is a continuous aspiration in which the sample flows through the burner while we monitor the absorbance. Continuous aspiration is sample intensive, typically requiring from 2–5 mL of sample.

Flame microsampling allows us to introduce a discrete sample of fixed volume, and is useful when we have a limited amount of sample or when the sample's matrix is incompatible with the flame atomizer. For example, continuously aspirating a sample that has a high concentration of dissolved solids—sea water, for example, comes to mind—may build-up a solid deposit on the burner head that obstructs the flame and that lowers the absorbance. Flame microsampling is accomplished using a micropipet to place

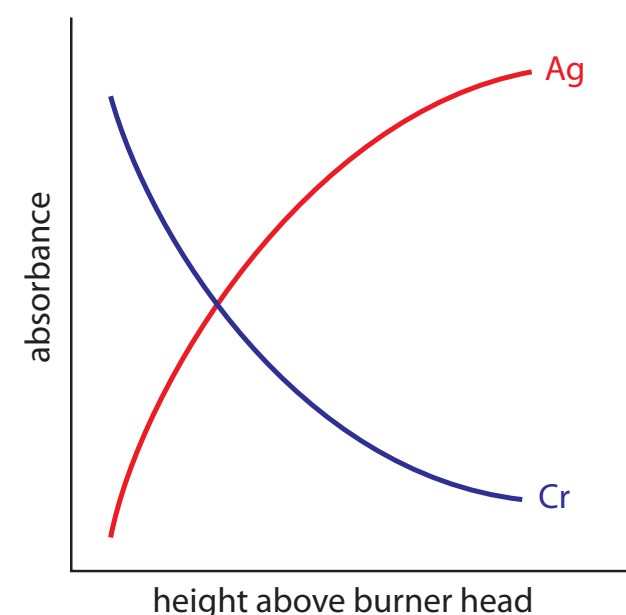


Figure 10.43 Absorbance versus height profiles for Ag and Cr in flame atomic absorption spectroscopy.

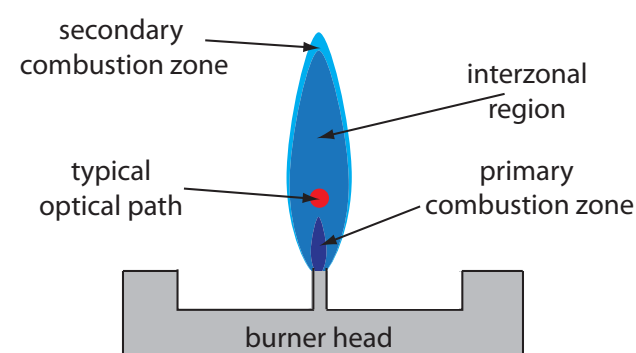


Figure 10.44 Profile of typical flame using a slot burner. The relative size of each zone depends on many factors, including the choice of fuel and oxidant, and their relative proportions.

Table 10.9 Fuels and Oxidants Used for Flame Combustion

fuel	oxidant	temperature range (°C)
natural gas	air	1700–1900
hydrogen	air	2000–2100
acetylene	air	2100–2400
acetylene	nitrous oxide	2600–2800
acetylene	oxygen	3050–3150

50–250 μL of sample in a Teflon funnel connected to the nebulizer, or by dipping the nebulizer tubing into the sample for a short time. Dip sampling is usually accomplished with an automatic sampler. The signal for flame microsampling is a transitory peak whose height or area is proportional to the amount of analyte that is injected.

Advantages and Disadvantages of Flame Atomization. The principal advantage of flame atomization is the reproducibility with which the sample is introduced into the spectrophotometer. A significant disadvantage to flame atomizers is that the efficiency of atomization may be quite poor. There are two reasons for poor atomization efficiency. First, the majority of the aerosol droplets produced during nebulization are too large to be carried to the flame by the combustion gases. Consequently, as much as 95% of the sample never reaches the flame. A second reason for poor atomization efficiency is that the large volume of combustion gases significantly dilutes the sample. Together, these contributions to the efficiency of atomization reduce sensitivity because the analyte's concentration in the flame may be a factor of 2.5×10^{-6} less than that in solution.¹⁴

This is the reason for the waste line shown at the bottom of the spray chamber in [Figure 10.42](#).

ELECTROTHERMAL ATOMIZERS

A significant improvement in sensitivity is achieved by using the resistive heating of a graphite tube in place of a flame. A typical electrothermal atomizer, also known as a **GRAPHITE FURNACE**, consists of a cylindrical graphite tube approximately 1–3 cm in length and 3–8 mm in diameter. As shown in Figure 10.45, the graphite tube is housed in a sealed assembly that has optically transparent windows at each end. A continuous stream of an inert gas is passed through the furnace, protecting the graphite tube from oxidation and removing the gaseous products produced during atomization. A power supply is used to pass a current through the graphite tube, resulting in resistive heating.

Samples of between 5–50 μL are injected into the graphite tube through a small hole at the top of the tube. Atomization is achieved in three stages. In the first stage the sample is dried to a solid residue using a current that raises the temperature of the graphite tube to about 110°C. In the second stage, which is called ashing, the temperature is increased to between

¹⁴ Ingle, J. D.; Crouch, S. R. *Spectrochemical Analysis*, Prentice-Hall: Englewood Cliffs, NJ, 1988; p. 275.

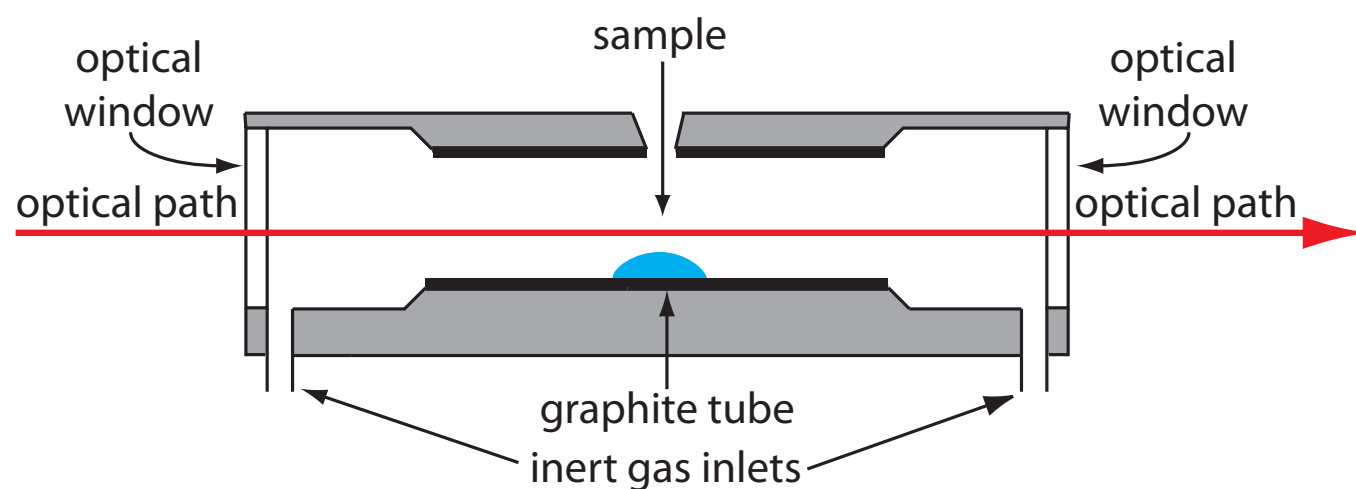


Figure 10.45 Diagram showing a cross-section of an electrothermal analyzer.

350–1200 °C. At these temperatures any organic material in the sample is converted to CO₂ and H₂O, and volatile inorganic materials are vaporized. These gases are removed by the inert gas flow. In the final stage the sample is atomized by rapidly increasing the temperature to between 2000–3000 °C. The result is a transient absorbance peak whose height or area is proportional to the absolute amount of analyte injected into the graphite tube. Together, the three stages take approximately 45–90 s, with most of this time used for drying and ashing the sample.

Electrothermal atomization provides a significant improvement in sensitivity by trapping the gaseous analyte in the small volume within the graphite tube. The analyte's concentration in the resulting vapor phase may be as much as 1000× greater than in a flame atomization.¹⁵ This improvement in sensitivity—and the resulting improvement in detection limits—is offset by a significant decrease in precision. Atomization efficiency is strongly influenced by the sample's contact with the graphite tube, which is difficult to control reproducibly.

MISCELLANEOUS ATOMIZATION METHODS

A few elements may be atomized by a chemical reaction that produces a volatile product. Elements such as As, Se, Sb, Bi, Ge, Sn, Te, and Pb, for example, form volatile hydrides when reacted with NaBH₄ in acid. An inert gas carries the volatile hydrides to either a flame or to a heated quartz observation tube situated in the optical path. Mercury is determined by the cold-vapor method in which it is reduced to elemental mercury with SnCl₂. The volatile Hg is carried by an inert gas to an unheated observation tube situated in the instrument's optical path.

10D.2 Quantitative Applications

Atomic absorption is widely used for the analysis of trace metals in a variety of sample matrices. Using Zn as an example, atomic absorption methods have been developed for its determination in samples as diverse as water and wastewater, air, blood, urine, muscle tissue, hair, milk, breakfast cereals, shampoos, alloys, industrial plating baths, gasoline, oil, sediments, and rocks.

Developing a quantitative atomic absorption method requires several considerations, including choosing a method of atomization, selecting the wavelength and slit width, preparing the sample for analysis, minimizing spectral and chemical interferences, and selecting a method of standardization. Each of these topics is considered in this section.

DEVELOPING A QUANTITATIVE METHOD

Flame or Electrothermal Atomization? The most important factor in choosing a method of atomization is the analyte's concentration. Because

¹⁵ Parsons, M. L.; Major, S.; Forster, A. R. *Appl. Spectrosc.* **1983**, *37*, 411–418.

Table 10.10 Concentration of Analyte Yielding an Absorbance of 0.20

element	Concentration (mg/L) ^a	
	flame atomization	electrothermal atomization
Ag	1.5	0.0035
Al	40	0.015
As	40 ^b	0.050
Ca	0.8	0.003
Cd	0.6	0.001
Co	2.5	0.021
Cr	2.5	0.0075
Cu	1.5	0.012
Fe	2.5	0.006
Hg	70 ^b	0.52
Mg	0.15	0.00075
Mn	1	0.003
Na	0.3	0.00023
Ni	2	0.024
Pb	5	0.080
Pt	70	0.29
Sn	50 ^b	0.023
Zn	0.3	0.00071

^a Source: Varian Cookbook, SpectraAA Software Version 4.00 Pro.

^b As: 10 mg/L by hydride vaporization; Hg: 11.5 mg/L by cold-vapor; and Sn: 18 mg/L by hydride vaporization

of its greater sensitivity, it takes less analyte to achieve a given absorbance when using electrothermal atomization. Table 10.10, which compares the amount of analyte needed to achieve an absorbance of 0.20 when using flame atomization and electrothermal atomization, is useful when selecting an atomization method. For example, flame atomization is the method of choice if our samples contain 1–10 mg Zn²⁺/L, but electrothermal atomization is the best choice for samples containing 1–10 μg Zn²⁺/L.

Selecting the Wavelength and Slit Width. The source for atomic absorption is a hollow cathode lamp consisting of a cathode and anode enclosed within a glass tube filled with a low pressure of Ne or Ar (Figure 10.46). Applying a potential across the electrodes ionizes the filler gas. The positively charged gas ions collide with the negatively charged cathode, sputtering atoms from the cathode's surface. Some of the sputtered atoms are in the excited state and emit radiation characteristic of the metal(s) from which the cathode was manufactured. By fashioning the cathode from the metallic analyte, a hollow cathode lamp provides emission lines that correspond to the analyte's absorption spectrum.

Because atomic absorption lines are narrow, we need to use a line source instead of a continuum source (compare, for example, Figure 10.18 with Figure 10.20). The effective bandwidth when using a continuum source is roughly 1000× larger than an atomic absorption line; thus, $P_T \approx P_0$, $\%T \approx 100$, and $A \approx 0$. Because a hollow cathode lamp is a line source, P_T and P_0 have different values giving a $\%T < 100$ and $A > 0$.

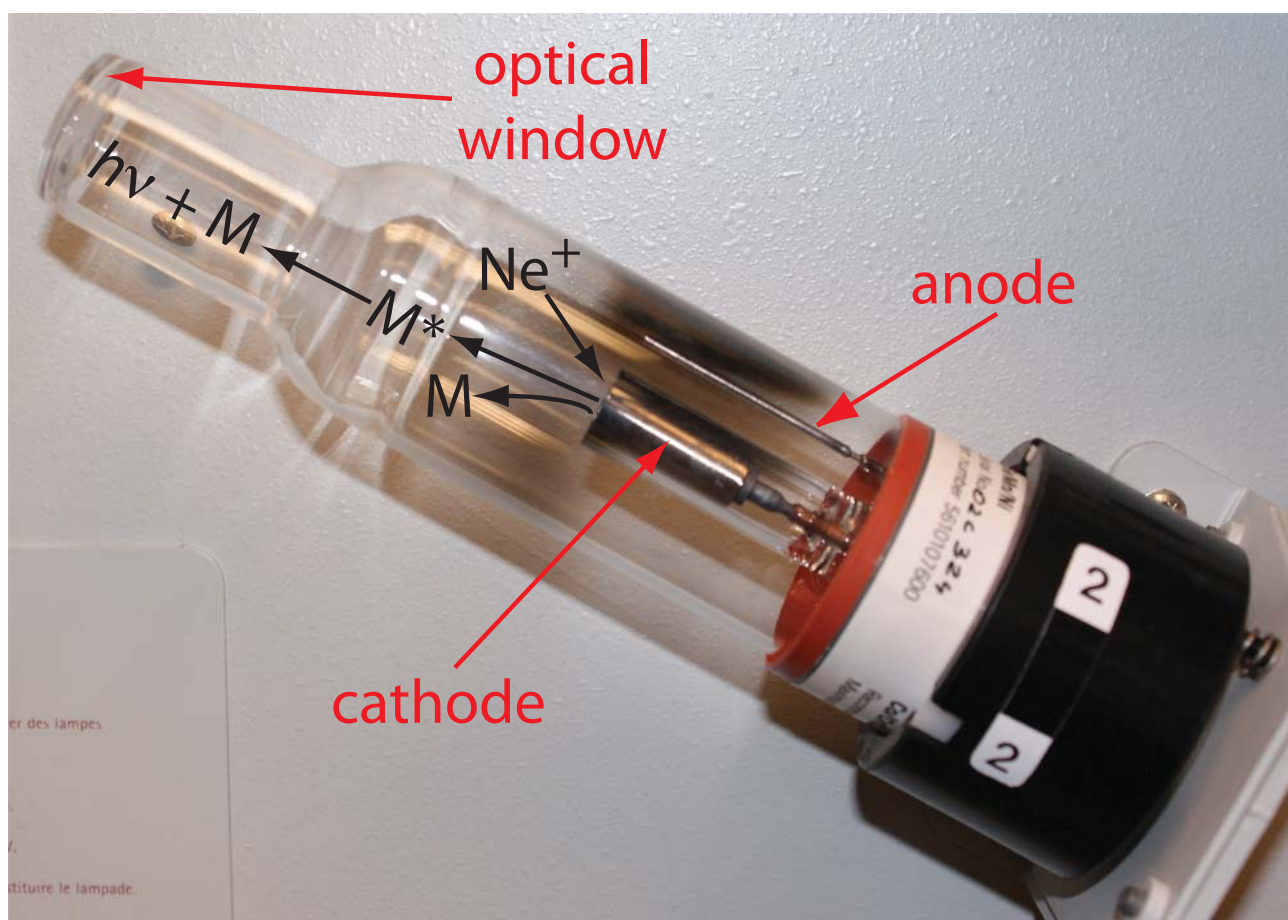


Figure 10.46 Photo of a typical multi-elemental hollow cathode lamp. The cathode in this lamp is fashioned from an alloy containing Co, Cr, Cu, Fe, Mn, and Ni, and is surrounded by a glass shield to isolate it from the anode. The lamp is filled with Ne gas. Also shown is the process leading to atomic emission. See the text for an explanation.

Each element in a hollow cathode lamp provides several atomic emission lines that we can use for atomic absorption. Usually the wavelength that provides the best sensitivity is the one we choose to use, although a less sensitive wavelength may be more appropriate for a larger concentration of analyte. For the Cr hollow cathode lamp in Table 10.11, for example, the best sensitivity is obtained using a wavelength of 357.9 nm.

Another consideration is the intensity of the emission line. If several emission lines meet our need for sensitivity, we may wish to use the emission line with the largest relative P_0 because there is less uncertainty in measuring P_0 and P_T . When analyzing samples containing ≈ 10 mg Cr/L, for example, the first three wavelengths in Table 10.11 provide an appropriate sensitivity. The wavelengths of 425.5 nm and 429.0 nm, however, have a greater P_0 and will provide less uncertainty in the measured absorbance.

The emission spectrum from a hollow cathode lamp includes, besides emission lines for the analyte, additional emission lines for impurities present in the metallic cathode and from the filler gas. These additional lines

Table 10.11 Atomic Emission Lines for a Cr Hollow Cathode Lamp

wavelength (nm)	slit width (nm)	mg Cr/L giving $A = 0.20$	P_0 (relative)
357.9	0.2	2.5	40
425.4	0.2	12	85
429.0	0.5	20	100
520.5	0.2	1500	15
520.8	0.2	500	20

are a source of stray radiation that leads to an instrumental deviation from Beer's law. The monochromator's slit width is set as wide as possible, improving the throughput of radiation, while, at the same time, being narrow enough to eliminate the stray radiation.

Preparing the Sample. Flame and electrothermal atomization require that the sample be in solution. Solid samples are brought into solution by dissolving in an appropriate solvent. If the sample is not soluble it may be digested, either on a hot-plate or by microwave, using HNO_3 , H_2SO_4 , or HClO_4 . Alternatively, we can extract the analyte using a Soxhlet extractor. Liquid samples may be analyzed directly or extracted if the matrix is incompatible with the method of atomization. A serum sample, for instance, is difficult to aspirate when using flame atomization and may produce an unacceptably high background absorbance when using electrothermal atomization. A liquid-liquid extraction using an organic solvent and a chelating agent is frequently used to concentrate analytes. Dilute solutions of Cd^{2+} , Co^{2+} , Cu^{2+} , Fe^{3+} , Pb^{2+} , Ni^{2+} , and Zn^{2+} , for example, can be concentrated by extracting with a solution of ammonium pyrrolidine dithiocarbamate in methyl isobutyl ketone.

Minimizing Spectral Interference. A spectral interference occurs when an analyte's absorption line overlaps with an interferent's absorption line or band. Because they are so narrow, the overlap of two atomic absorption lines is seldom a problem. On the other hand, a molecule's broad absorption band or the scattering of source radiation is a potentially serious spectral interference.

An important consideration when using a flame as an atomization source is its effect on the measured absorbance. Among the products of combustion are molecular species that exhibit broad absorption bands and particulates that scatter radiation from the source. If we fail to compensate for these spectral interference, then the intensity of transmitted radiation decreases. The result is an apparent increase in the sample's absorbance. Fortunately, absorption and scattering of radiation by the flame are corrected by analyzing a blank.

Spectral interferences also occur when components of the sample's matrix other than the analyte react to form molecular species, such as oxides and hydroxides. The resulting absorption and scattering constitutes the sample's background and may present a significant problem, particularly at wavelengths below 300 nm where the scattering of radiation becomes more important. If we know the composition of the sample's matrix, then we can prepare our samples using an identical matrix. In this case the background absorption is the same for both the samples and standards. Alternatively, if the background is due to a known matrix component, then we can add that component in excess to all samples and standards so that the contribution of the naturally occurring interferent is insignificant. Finally, many interferences due to the sample's matrix can be eliminated by increasing the

See [Chapter 7](#) to review different methods for preparing samples for analysis.

atomization temperature. For example, by switching to a higher temperature flame it may be possible to prevent the formation of interfering oxides and hydroxides.

If the identity of the matrix interference is unknown, or if it is not possible to adjust the flame or furnace conditions to eliminate the interference, then we must find another method to compensate for the background interference. Several methods have been developed to compensate for matrix interferences, and most atomic absorption spectrophotometers include one or more of these methods.

One of the most common methods for **BACKGROUND CORRECTION** is to use a continuum source, such as a D₂ lamp. Because a D₂ lamp is a continuum source, absorbance of its radiation by the analyte's narrow absorption line is negligible. Only the background, therefore, absorbs radiation from the D₂ lamp. Both the analyte and the background, on the other hand, absorb the hollow cathode's radiation. Subtracting the absorbance for the D₂ lamp from that for the hollow cathode lamp gives a corrected absorbance that compensates for the background interference. Although this method of background correction may be quite effective, it does assume that the background absorbance is constant over the range of wavelengths passed by the monochromator. If this is not true, subtracting the two absorbances may underestimate or overestimate the background.

Minimizing Chemical Interferences. The quantitative analysis of some elements is complicated by chemical interferences occurring during atomization. The two most common chemical interferences are the formation of nonvolatile compounds containing the analyte and ionization of the analyte.

One example of the formation of nonvolatile compounds is the effect of PO₄³⁻ or Al³⁺ on the flame atomic absorption analysis of Ca²⁺. In one study, for example, adding 100 ppm Al³⁺ to a solution of 5 ppm Ca²⁺ decreased the calcium ion's absorbance from 0.50 to 0.14, while adding 500 ppm PO₄³⁻ to a similar solution of Ca²⁺ decreased the absorbance from 0.50 to 0.38. These interferences were attributed to the formation of nonvolatile particles of Ca₃(PO₄)₂ and an Al–Ca–O oxide.¹⁶

When using flame atomization, we can minimize the formation of nonvolatile compounds by increasing the flame's temperature, either by changing the fuel-to-oxidant ratio or by switching to a different combination of fuel and oxidant. Another approach is to add a releasing agent or a protecting agent to the samples. A **RELEASING AGENT** is a species that reacts with the interferent, releasing the analyte during atomization. Adding Sr²⁺ or La³⁺ to solutions of Ca²⁺, for example, minimizes the effect of PO₄³⁻ and Al³⁺ by reacting in place of the analyte. Thus, adding 2000 ppm SrCl₂ to the Ca²⁺/PO₄³⁻ and Ca²⁺/Al³⁺ mixtures described in the previous paragraph increased the absorbance to 0.48. A **PROTECTING AGENT** reacts with the analyte to form a stable volatile complex. Adding 1% w/w EDTA

¹⁶ Hosking, J. W.; Snell, N. B.; Sturman, B. T. J. *Chem. Educ.* **1977**, *54*, 128–130.

Other methods of background correction have been developed, including Zeeman effect background correction and Smith–Hieftje background correction, both of which are included in some commercially available atomic absorption spectrophotometers. Consult the chapter's [additional resources](#) for additional information.

to the $\text{Ca}^{2+}/\text{PO}_4^{3-}$ solution described in the previous paragraph increased the absorbance to 0.52.

Ionization interferences occur when thermal energy from the flame or the electrothermal atomizer is sufficient to ionize the analyte



where M is the analyte. Because the absorption spectra for M and M^+ are different, the position of the equilibrium in reaction 10.24 affects absorbance at wavelengths where M absorbs. To limit ionization we add a high concentration of an **IONIZATION SUPPRESSOR**, which is simply a species that ionizes more easily than the analyte. If the concentration of the ionization suppressor is sufficient, then the increased concentration of electrons in the flame pushes reaction 10.24 to the left, preventing the analyte's ionization. Potassium and cesium are frequently used as an ionization suppressor because of their low ionization energy.

Standardizing the Method. Because Beer's law also applies to atomic absorption, we might expect atomic absorption calibration curves to be linear. In practice, however, most atomic absorption calibration curves are nonlinear, or linear for only a limited range of concentrations. Nonlinearity in atomic absorption is a consequence of instrumental limitations, including stray radiation from the hollow cathode lamp and the variation in molar absorptivity across the absorption line. Accurate quantitative work, therefore, often requires a suitable means for computing the calibration curve from a set of standards.

When possible, a quantitative analysis is best conducted using external standards. Unfortunately, matrix interferences are a frequent problem, particularly when using electrothermal atomization. For this reason the method of standard additions is often used. One limitation to this method of standardization, however, is the requirement that there be a linear relationship between absorbance and concentration.

Representative Method 10.2

Determination of Cu and Zn in Tissue Samples

DESCRIPTION OF METHOD

Copper and zinc are isolated from tissue samples by digesting the sample with HNO_3 after first removing any fatty tissue. The concentration of copper and zinc in the supernatant are determined by atomic absorption using an air-acetylene flame.

PROCEDURE

Tissue samples are obtained by a muscle needle biopsy and dried for 24–30 h at 105°C to remove all traces of moisture. The fatty tissue in the dried samples is removed by extracting overnight with anhydrous ether. After removing the ether, the sample is dried to obtain the fat-free dry

Most instruments include several different algorithms for computing the calibration curve. The instrument in my lab, for example, includes five algorithms. Three of the algorithms fit absorbance data using linear, quadratic, or cubic polynomial functions of the analyte's concentration. It also includes two algorithms that fit the concentrations of the standards to quadratic functions of the absorbance.

The best way to appreciate the theoretical and practical details discussed in this section is to carefully examine a typical analytical method. Although each method is unique, the following description of the determination of Cu and Zn in biological tissues provides an instructive example of a typical procedure. The description here is based on Bhattacharya, S. K.; Goodwin, T. G.; Crawford, A. J. *Anal. Lett.* **1984**, *17*, 1567–1593, and Crawford, A. J.; Bhattacharya, S. K. Varian Instruments at Work, Number AA–46, April 1985.

tissue weight (FFDT). The sample is digested at 68 °C for 20–24 h using 3 mL of 0.75 M HNO₃. After centrifuging at 2500 rpm for 10 minutes, the supernatant is transferred to a 5-mL volumetric flask. The digestion is repeated two more times, for 2–4 hours each, using 0.9-mL aliquots of 0.75 M HNO₃. These supernatants are added to the 5-mL volumetric flask, which is diluted to volume with 0.75 M HNO₃. The concentrations of Cu and Zn in the diluted supernatant are determined by flame atomic absorption spectroscopy using an air-acetylene flame and external standards. Copper is analyzed at a wavelength of 324.8 nm with a slit width of 0.5 nm, and zinc is analyzed at 213.9 nm with a slit width of 1.0 nm. Background correction using a D₂ lamp is necessary for zinc. Results are reported as µg of Cu or Zn per gram of FFDT.

QUESTIONS

1. Describe the appropriate matrix for the external standards and for the blank?

The matrix for the standards and the blank should match the matrix of the samples; thus, an appropriate matrix is 0.75 M HNO₃. Any interferences from other components of the sample matrix are minimized by background correction.

2. Why is a background correction necessary for the analysis of Zn, but not for the analysis of Cu?

Background correction compensates for background absorption and scattering due to interferences in the sample. Such interferences are most severe when using a wavelength less than 300 nm. This is the case for Zn, but not for Cu.

3. A Cu hollow cathode lamp has several emission lines. Explain why this method uses the line at 324.8 nm.

wavelength (nm)	slit width (nm)	mg Cu/L for $A = 0.20$	P_0 (relative)
217.9	0.2	15	3
218.2	0.2	15	3
222.6	0.2	60	5
244.2	0.2	400	15
249.2	0.5	200	24
324.8	0.5	1.5	100
327.4	0.5	3	87

With 1.5 mg Cu/L giving an absorbance of 0.20, the emission line at 324.8 nm has the best sensitivity. In addition, it is the most intense emission line, which decreases the uncertainty in the measured absorbance.

Example 10.10

To evaluate the method described in [Representative Method 10.2](#), a series of external standard is prepared and analyzed, providing the results shown here.¹⁷

$\mu\text{g Cu/mL}$	absorbance	$\mu\text{g Cu/mL}$	absorbance
0.000	0.000	0.500	0.033
0.100	0.006	0.600	0.039
0.200	0.013	0.700	0.046
0.300	0.020	1.000	0.066
0.400	0.026		

A bovine liver standard reference material was used to evaluate the method's accuracy. After drying and extracting the sample, a 11.23-mg FFDT tissue sample gives an absorbance of 0.023. Report the amount of copper in the sample as $\mu\text{g Cu/g FFDT}$.

SOLUTION

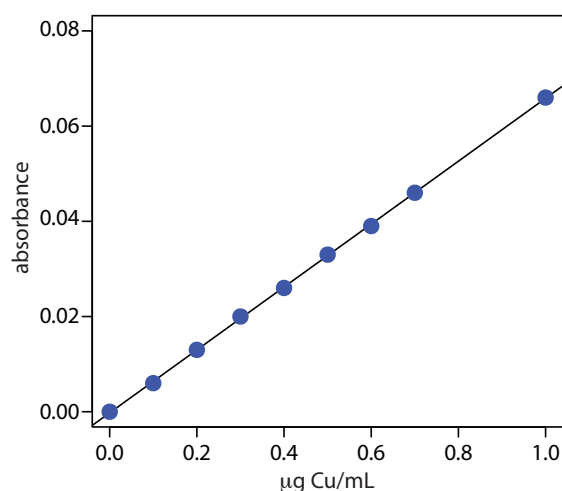
Linear regression of absorbance versus the concentration of Cu in the standards gives a calibration curve with the following equation.

$$A = -0.0002 + 0.0661 \times \frac{\text{g Cu}}{\text{mL}}$$

Substituting the sample's absorbance into the calibration equation gives the concentration of copper as $0.351 \mu\text{g/mL}$. The concentration of copper in the tissue sample, therefore, is

$$\frac{0.351 \text{ g Cu}}{\text{mL}} \times 5.000 \text{ mL} = 1.755 \text{ g Cu}$$

$$\frac{1.755 \text{ g Cu}}{0.01123 \text{ g sample}} = 156 \text{ g Cu/g FFDT}$$



See [Figure 3.5](#) to review the meaning of macro and meso for describing samples, and the meaning of major, minor, and ultratrace for describing analytes.

10D.3 - Evaluation of Atomic Absorption Spectroscopy

SCALE OF OPERATION

Atomic absorption spectroscopy is ideally suited for the analysis of trace and ultratrace analytes, particularly when using electrothermal atomization. For minor and major analyte, sample can be diluted before the analysis. Most analyses use a macro or a meso sample. The small volume requirement for electrothermal atomization or flame microsampling, however, makes practical the analysis micro and ultramicro samples.

¹⁷ Crawford, A. J.; Bhattacharya, S. K. "Microanalysis of Copper and Zinc in Biopsy-Sized Tissue Specimens by Atomic Absorption Spectroscopy Using a Stoichiometric Air-Acetylene Flame," Varian Instruments at Work, Number AA-46, April 1985.

ACCURACY

If spectral and chemical interferences are minimized, an accuracy of 0.5–5% is routinely attainable. When the calibration curve is nonlinear, accuracy may be improved by using a pair of standards whose absorbances closely bracket the sample's absorbance and assuming that the change in absorbance is linear over this limited concentration range. Determinate errors for electrothermal atomization are often greater than that obtained with flame atomization due to more serious matrix interferences.

PRECISION

For absorbance values greater than 0.1–0.2, the relative standard deviation for atomic absorption is 0.3–1% for flame atomization and 1–5% for electrothermal atomization. The principle limitation is the variation in the concentration of free analyte atoms resulting from variations in the rate of aspiration, nebulization, and atomization when using a flame atomizer, and the consistency of injecting samples when using electrothermal atomization.

SENSITIVITY

The sensitivity of a flame atomic absorption analysis is influenced strongly by the flame's composition and by the position in the flame from which we monitor the absorbance. Normally the sensitivity of an analysis is optimized by aspirating a standard solution of the analyte and adjusting operating conditions, such as the fuel-to-oxidant ratio, the nebulizer flow rate, and the height of the burner, to give the greatest absorbance. With electrothermal atomization, sensitivity is influenced by the drying and ashing stages that precede atomization. The temperature and time used for each stage must be optimized for each type of sample.

Sensitivity is also influenced by the sample's matrix. We have already noted, for example, that sensitivity can be decreased by chemical interferences. An increase in sensitivity may be realized by adding a low molecular weight alcohol, ester, or ketone to the solution, or by using an organic solvent.

SELECTIVITY

Due to the narrow width of absorption lines, atomic absorption provides excellent selectivity. Atomic absorption can be used for the analysis of over 60 elements at concentrations at or below the level of $\mu\text{g/L}$.

TIME, COST, AND EQUIPMENT

The analysis time when using flame atomization is short, with sample throughputs of 250–350 determinations per hour when using a fully automated system. Electrothermal atomization requires substantially more time

See [Chapter 14](#) for several strategies for optimizing experiments.

per analysis, with maximum sample throughputs of 20–30 determinations per hour. The cost of a new instrument ranges from between \$10,000–\$50,000 for flame atomization, and from \$18,000–\$70,000 for electrothermal atomization. The more expensive instruments in each price range include double-beam optics, automatic samplers, and can be programmed for multielemental analysis by allowing the wavelength and hollow cathode lamp to be changed automatically.

10E Emission Spectroscopy

An analyte in an excited state possesses an energy, E_2 , that is greater than its energy when it is in a lower energy state, E_1 . When the analyte returns to its lower energy state—a process we call **RELAXATION**—the excess energy, ΔE

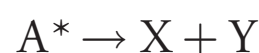
$$\Delta E = E_2 - E_1$$

must be released. [Figure 10.4](#) shows a simplified picture of this process.

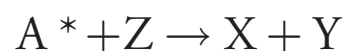
The amount of time the analyte spends in the excited state—its **LIFE-TIME**—is short, typically 10^{-5} – 10^{-9} s for electronic excited states and 10^{-15} s for vibrational excited states. Relaxation of an analyte's excited-state, A^* , occurs through several mechanisms, including collisions with other species in the sample, by photochemical reactions, and by the emission of photons. In the first process, which is called vibrational relaxation, or nonradiative relaxation, the excess energy is released as heat.



Relaxation by a photochemical reaction may involve a decomposition reaction

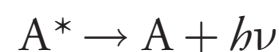


or a reaction between A^* and another species



In both cases the excess energy is used up in the chemical reaction or released as heat.

In the third mechanism, the excess energy is released as a photon of electromagnetic radiation.



The release of a photon following thermal excitation is called emission and that following the absorption of a photon is called photoluminescence. In chemiluminescence and bioluminescence, excitation results from a chemical or biochemical reaction, respectively. Spectroscopic methods based on photoluminescence are the subject of the next section and atomic emission is covered in [Section 10G](#).

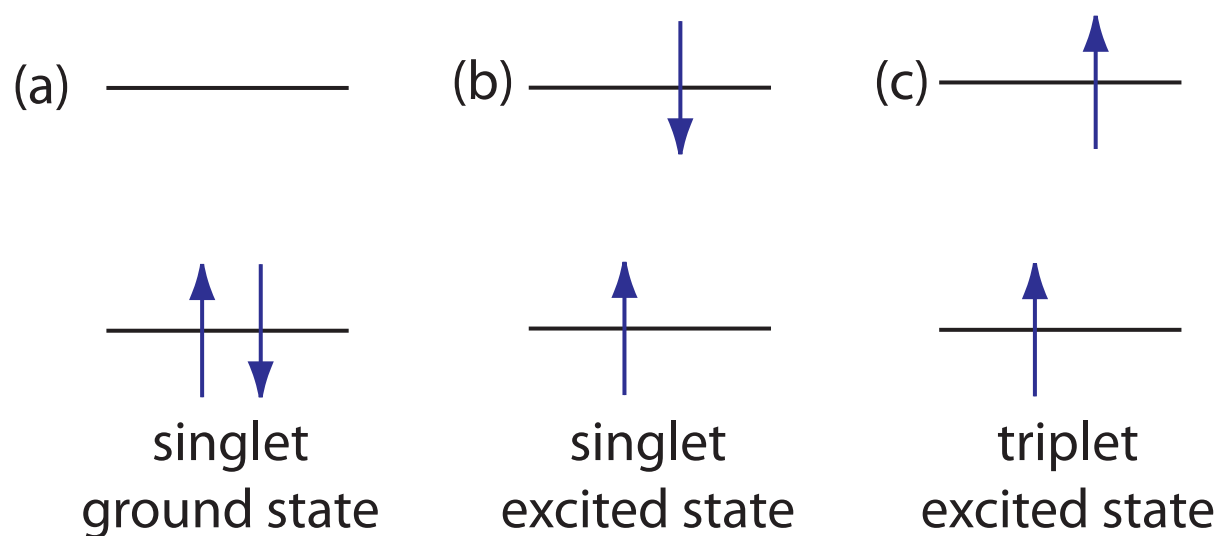


Figure 10.47 Electron configurations for (a) a singlet ground state; (b) a singlet excited state; and (c) a triplet excited state.

10F Photoluminescence Spectroscopy

Photoluminescence is divided into two categories: fluorescence and phosphorescence. A pair of electrons occupying the same electronic ground state have opposite spins and are said to be in a singlet spin state (Figure 10.47a). When an analyte absorbs an ultraviolet or visible photon, one of its valence electrons moves from the ground state to an excited state with a conservation of the electron's spin (Figure 10.47b). Emission of a photon from the **SINGLET EXCITED STATE** to the singlet ground state—or between any two energy levels with the same spin—is called **FLUORESCENCE**. The probability of fluorescence is very high and the average lifetime of an electron in the excited state is only 10^{-5} – 10^{-8} s. Fluorescence, therefore, decays rapidly once the source of excitation is removed.

In some cases an electron in a singlet excited state is transformed to a **TRIPLET EXCITED STATE** (Figure 10.47c) in which its spin is no longer paired with the ground state. Emission between a triplet excited state and a singlet ground state—or between any two energy levels that differ in their respective spin states—is called **PHOSPHORESCENCE**. Because the average lifetime for phosphorescence ranges from 10^{-4} – 10^4 s, phosphorescence may continue for some time after removing the excitation source.

The use of molecular fluorescence for qualitative analysis and semi-quantitative analysis can be traced to the early to mid 1800s, with more accurate quantitative methods appearing in the 1920s. Instrumentation for fluorescence spectroscopy using a filter or a monochromator for wavelength selection appeared in, respectively, the 1930s and 1950s. Although the discovery of phosphorescence preceded that of fluorescence by almost 200 years, qualitative and quantitative applications of molecular phosphorescence did not receive much attention until after the development of fluorescence instrumentation.

10F.1 Fluorescence and Phosphorescence Spectra

To appreciate the origin of fluorescence and phosphorescence we must consider what happens to a molecule following the absorption of a photon. Let's assume that the molecule initially occupies the lowest vibrational energy level of its electronic ground state, which is a singlet state labeled S_0 in

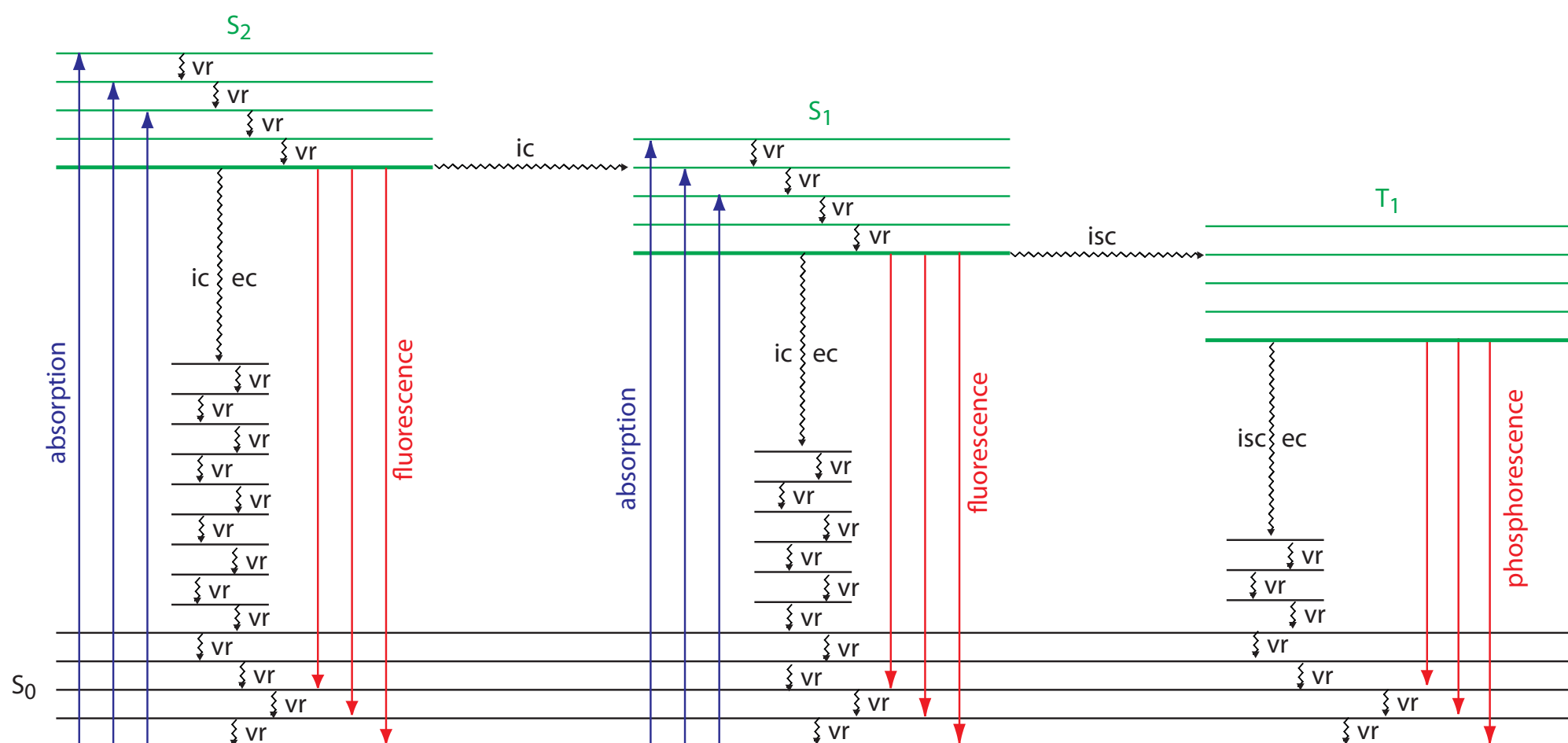


Figure 10.48 Energy level diagram for a molecule showing pathways for the deactivation of an excited state: *vr* is vibrational relaxation; *ic* is internal conversion; *ec* is external conversion; and *isc* is an intersystem crossing. The lowest vibrational energy for each electronic state is indicated by the thicker line. The electronic ground state is shown in black and the three electronic excited states are shown in green. The absorption, fluorescence, and phosphorescence of photons also are shown.

Figure 10.48. Absorption of a photon excites the molecule to one of several vibrational energy levels in the first excited electronic state, S_1 , or the second electronic excited state, S_2 , both of which are singlet states. Relaxation to the ground state occurs by a number of mechanisms, some involving the emission of photons and others occurring without emitting photons. These relaxation mechanisms are shown in Figure 10.48. The most likely relaxation pathway is the one with the shortest lifetime for the excited state.

RADIATIONLESS DEACTIVATION

When a molecule relaxes without emitting a photon we call the process **RADIATIONLESS DEACTIVATION**. One example of radiationless deactivation is **VIBRATIONAL RELAXATION**, in which a molecule in an excited vibrational energy level loses energy by moving to a lower vibrational energy level in the same electronic state. Vibrational relaxation is very rapid, with an average lifetime of $<10^{-12}$ s. Because vibrational relaxation is so efficient, a molecule in one of its excited state's higher vibrational energy levels quickly returns to the excited state's lowest vibrational energy level.

Another form of radiationless deactivation is an **INTERNAL CONVERSION** in which a molecule in the ground vibrational level of an excited state passes directly into a higher vibrational energy level of a lower energy electronic state of the same spin state. By a combination of internal conversions and vibrational relaxations, a molecule in an excited electronic state may return

to the ground electronic state without emitting a photon. A related form of radiationless deactivation is an **EXTERNAL CONVERSION** in which excess energy is transferred to the solvent or to another component of the sample's matrix.

A final form of radiationless deactivation is an **INTERSYSTEM CROSSING** in which a molecule in the ground vibrational energy level of an excited electronic state passes into a higher vibrational energy level of a lower energy electronic state with a different spin state. For example, an intersystem crossing is shown in [Figure 10.48](#) between a singlet excited state, S_1 , and a triplet excited state, T_1 .

RELAXATION BY FLUORESCENCE

Fluorescence occurs when a molecule in an excited state's lowest vibrational energy level returns to a lower energy electronic state by emitting a photon. Because molecules return to their ground state by the fastest mechanism, fluorescence is observed only if it is a more efficient means of relaxation than a combination of internal conversions and vibrational relaxations.

A quantitative expression of fluorescence efficiency is the **FLUORESCENT QUANTUM YIELD**, Φ_f , which is the fraction of excited state molecules returning to the ground state by fluorescence. Fluorescent quantum yields range from 1, when every molecule in an excited state undergoes fluorescence, to 0 when fluorescence does not occur.

The intensity of fluorescence, I_f , is proportional to the amount of radiation absorbed by the sample, $P_0 - P_T$, and the fluorescent quantum yield

$$I_f = k\Phi_f(P_0 - P_T) \quad 10.25$$

where k is a constant accounting for the efficiency of collecting and detecting the fluorescent emission. From Beer's law we know that

$$\frac{P_T}{P_0} = 10^{-\epsilon b C} \quad 10.26$$

where C is the concentration of the fluorescing species. Solving equation 10.26 for P_T and substituting into equation 10.25 gives, after simplifying

$$I_f = k\Phi_f P_0 (1 - 10^{-\epsilon b C}) \quad 10.27$$

When $\epsilon b C < 0.01$, which often is the case when concentration is small, equation 10.27 simplifies to

$$I_f = 2.303k\Phi_f \epsilon b C P_0 = k' P_0 \quad 10.28$$

where k' is a collection of constants. The intensity of fluorescent emission, therefore, increases with an increase in the quantum efficiency, the source's incident power, and the molar absorptivity and the concentration of the fluorescing species.

Fluorescence is generally observed when the molecule's lowest energy absorption is a $\pi \rightarrow \pi^*$ transition, although some $n \rightarrow \pi^*$ transitions show

Let's use [Figure 10.48](#) to illustrate how a molecule can relax back to its ground state without emitting a photon. Suppose our molecule is in the highest vibrational energy level of the second electronic excited state. After a series of vibrational relaxations brings the molecule to the lowest vibrational energy level of S_2 , it undergoes an internal conversion into a higher vibrational energy level of the first excited electronic state. Vibrational relaxations bring the molecule to the lowest vibrational energy level of S_1 . Following an internal conversion into a higher vibrational energy level of the ground state, the molecule continues to undergo vibrational relaxation until it reaches the lowest vibrational energy level of S_0 .

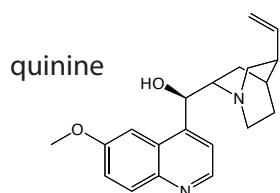


Figure 10.49 Tonic water, which contains quinine, is fluorescent when placed under a UV lamp. Source: [Splarka](https://commons.wikimedia.org/wiki/File:Canada_Dry_Tonic_Water_UV.jpg) (commons.wikimedia.org).

weak fluorescence. Most unsubstituted, nonheterocyclic aromatic compounds have favorable fluorescence quantum yields, although substitutions on the aromatic ring can significantly effect Φ_f . For example, the presence of an electron-withdrawing group, such as $-\text{NO}_2$, decreases Φ_f , while adding an electron-donating group, such as $-\text{OH}$, increases Φ_f . Fluorescence also increases for aromatic ring systems and for aromatic molecules with rigid planar structures. Figure 10.49 shows the fluorescence of quinine under a UV lamp.

A molecule's fluorescent quantum yield is also influenced by external variables, such as temperature and solvent. Increasing the temperature generally decreases Φ_f because more frequent collisions between the molecule and the solvent increases external conversion. A decrease in the solvent's viscosity decreases Φ_f for similar reasons. For an analyte with acidic or basic functional groups, a change in pH may change the analyte's structure and its fluorescent properties.

As shown in [Figure 10.48](#), fluorescence may return the molecule to any of several vibrational energy levels in the ground electronic state. Fluorescence, therefore, occurs over a range of wavelengths. Because the change in energy for fluorescent emission is generally less than that for absorption, a molecule's fluorescence spectrum is shifted to higher wavelengths than its absorption spectrum.

RELAXATION BY PHOSPHORESCENCE

A molecule in a triplet electronic excited state's lowest vibrational energy level normally relaxes to the ground state by an intersystem crossing to a singlet state or by an external conversion. Phosphorescence occurs when the molecule relaxes by emitting a photon. As shown in [Figure 10.48](#), phosphorescence occurs over a range of wavelengths, all of which are at lower energies than the molecule's absorption band. The intensity of phosphorescence, I_p , is given by an equation similar to [equation 10.28](#) for fluorescence

$$I_p = 2.303k\Phi_p\varepsilon bCP_0 = k'P_0 \quad 10.29$$

where Φ_p is the **PHOSPHORESCENT QUANTUM YIELD**.

Phosphorescence is most favorable for molecules with $n \rightarrow \pi^*$ transitions, which have a higher probability for an intersystem crossing than $\pi \rightarrow \pi^*$ transitions. For example, phosphorescence is observed with aromatic molecules containing carbonyl groups or heteroatoms. Aromatic compounds containing halide atoms also have a higher efficiency for phosphorescence. In general, an increase in phosphorescence corresponds to a decrease in fluorescence.

Because the average lifetime for phosphorescence is very long, ranging from 10^{-4} – 10^4 s, the phosphorescent quantum yield is usually quite small. An improvement in Φ_p is realized by decreasing the efficiency of external conversion. This may be accomplished in several ways, including lowering the temperature, using a more viscous solvent, depositing the sample on a

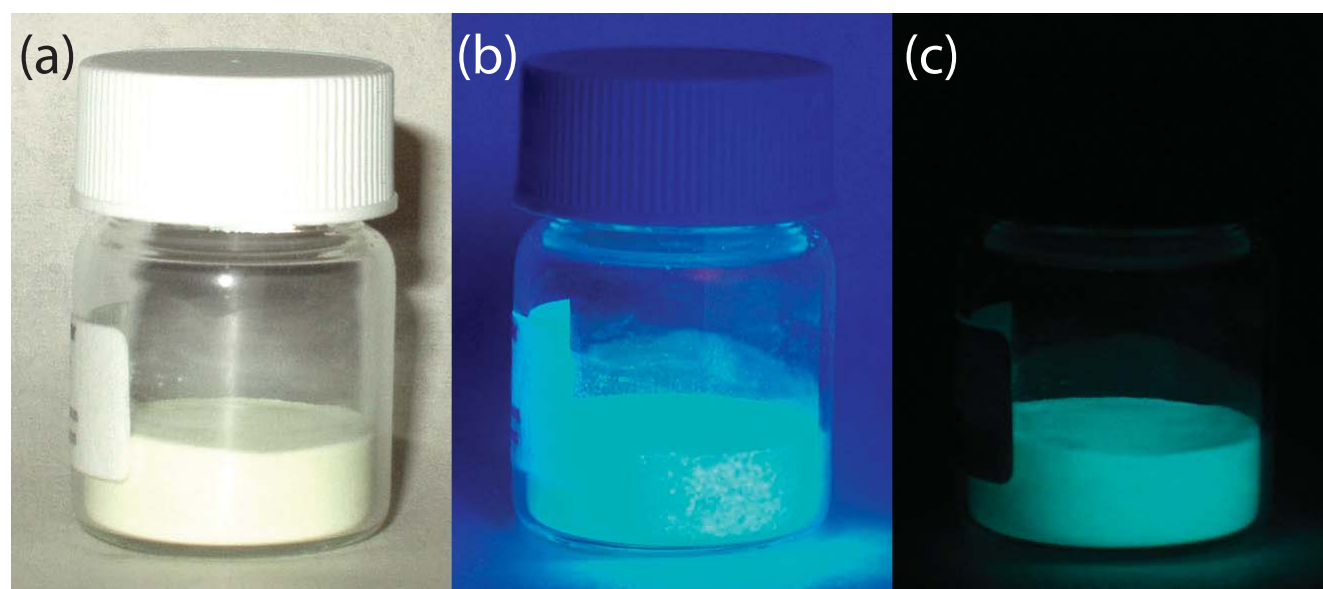


Figure 10.50 An europium doped strontium silicate-aluminum oxide powder under (a) natural light, (b) a long-wave UV lamp, and (c) in total darkness. The photo taken in total darkness shows the phosphorescent emission. Source: modified from [Splarka](#) (commons.wikipedia.org).

solid substrate, or trapping the molecule in solution. Figure 10.50 shows an example of phosphorescence.

EXCITATION VERSUS EMISSION SPECTRA

Photoluminescence spectra are recorded by measuring the intensity of emitted radiation as a function of either the excitation wavelength or the emission wavelength. An **EXCITATION SPECTRUM** is obtained by monitoring emission at a fixed wavelength while varying the excitation wavelength. When corrected for variations in the source's intensity and the detector's response, a sample's excitation spectrum is nearly identical to its absorbance spectrum. The excitation spectrum provides a convenient means for selecting the best excitation wavelength for a quantitative or qualitative analysis.

In an **EMISSION SPECTRUM** a fixed wavelength is used to excite the sample and the intensity of emitted radiation is monitored as function of wavelength. Although a molecule has only a single excitation spectrum, it has two emission spectra, one for fluorescence and one for phosphorescence. Figure 10.51 shows the UV absorption spectrum and the UV fluorescence emission spectrum for tyrosine.

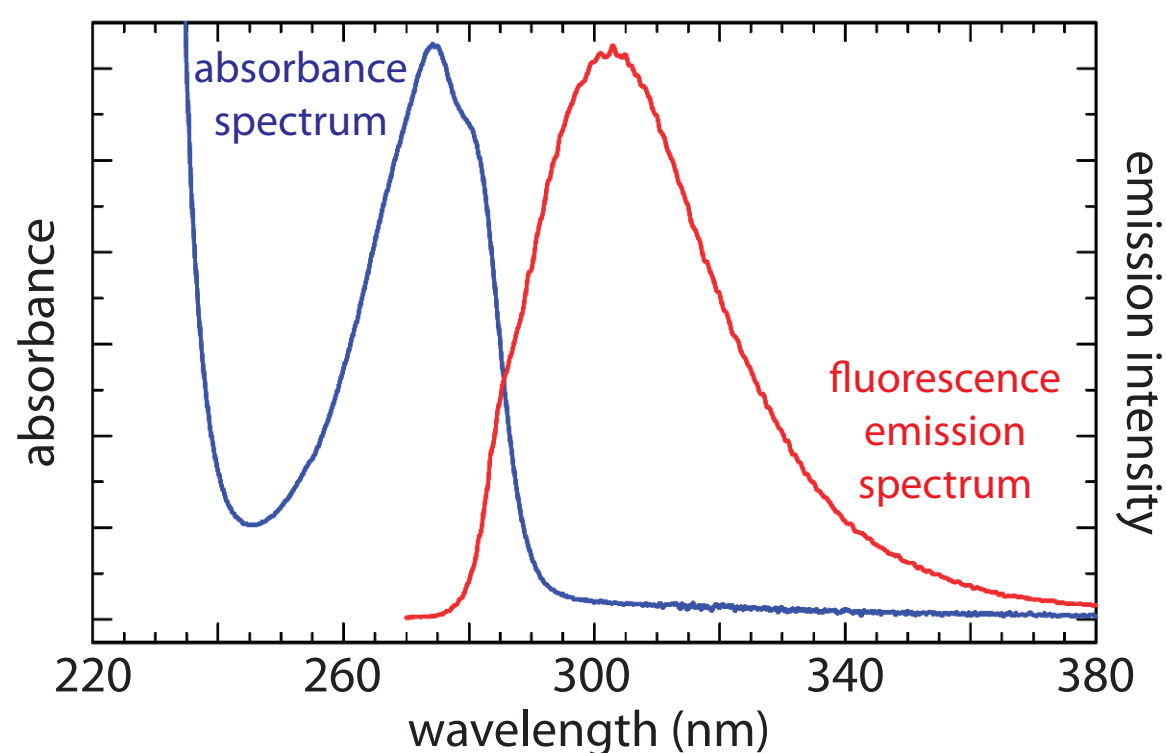


Figure 10.51 Absorbance spectrum and fluorescence emission spectrum for tyrosine in a pH 7, 0.1 M phosphate buffer. The emission spectrum uses an excitation wavelength of 260 nm. Source: modified from [Mark Somoza](#) (commons.wikipedia.org).

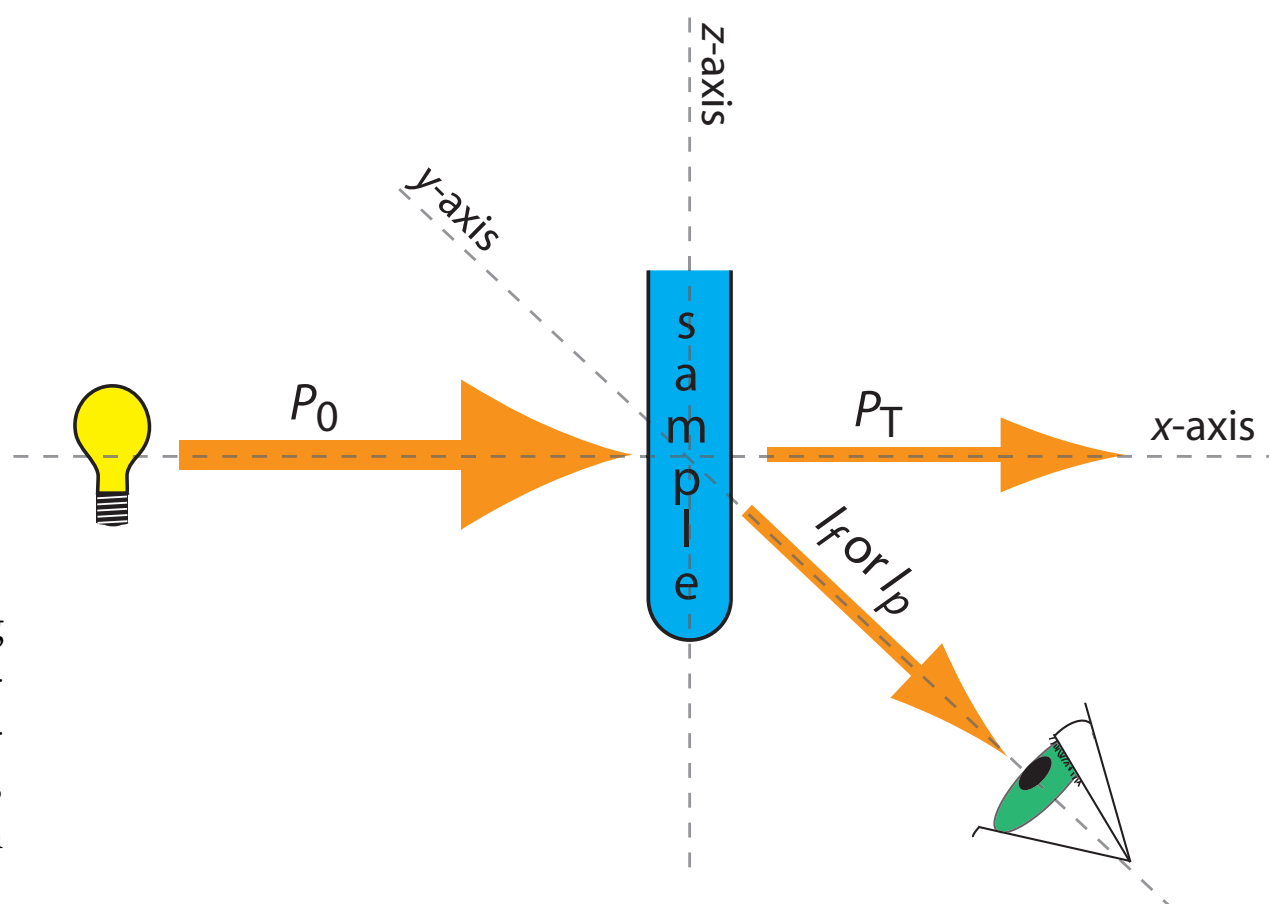


Figure 10.52 Schematic diagram showing the orientation of the source and the detector when measuring fluorescence and phosphorescence. Contrast this to [Figure 10.21](#), which shows the orientation for absorption spectroscopy.

10F.2 Instrumentation

The basic instrumental needs for monitoring fluorescence and phosphorescence—a source of radiation, a means of selecting a narrow band of radiation, and a detector—are the same as those for absorption spectroscopy. The unique demands of both techniques, however, require some modifications to the instrument designs seen earlier in [Figure 10.25](#) (filter photometer), [Figure 10.26](#) (single-beam spectrophotometer), [Figure 10.27](#) (double-beam spectrophotometer), and [Figure 10.28](#) (diode array spectrometer). The most important difference is the detector cannot be placed directly across from the source. [Figure 10.52](#) shows why this is the case. If we place the detector along the source's axis it will receive both the transmitted source radiation, P_T , and the fluorescent, I_f , or phosphorescent, I_p , radiation. Instead, we rotate the detector and place it at 90° to the source.

INSTRUMENTS FOR MEASURING FLUORESCENCE

[Figure 10.53](#) shows the basic design of an instrument for measuring fluorescence, which includes two wavelength selectors, one for selecting an excitation wavelength from the source and one for selecting the emission wavelength from the sample. In a **FLUORIMETER** the excitation and emission wavelengths are selected using absorption or interference filters. The excitation source for a fluorimeter is usually a low-pressure Hg vapor lamp that provides intense emission lines distributed throughout the ultraviolet and visible region (254, 312, 365, 405, 436, 546, 577, 691, and 773 nm). When a monochromator is used to select the excitation and emission wavelengths, the instrument is called a **SPECTROFLUORIMETER**. With a monochromator the excitation source is usually high-pressure Xe arc lamp, which has a continuous emission spectrum. Either instrumental design is

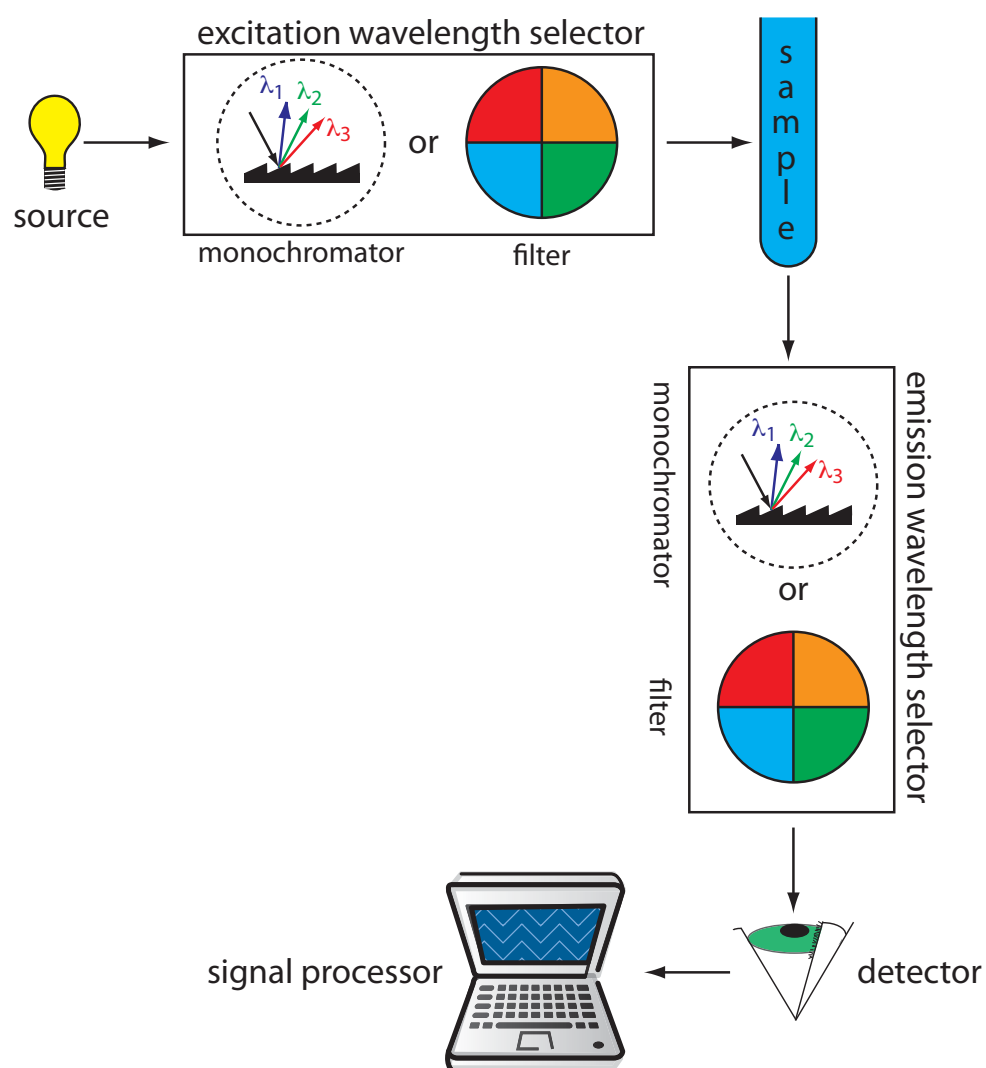


Figure 10.53 Schematic diagram for measuring fluorescence showing the placement of the wavelength selectors for excitation and emission. When a filter is used the instrument is called a fluorimeter, and when a monochromator is used the instrument is called a spectrofluorimeter.

appropriate for quantitative work, although only a spectrofluorimeter can be used to record an excitation or emission spectrum.

The sample cells for molecular fluorescence are similar to those for molecular absorption. Remote sensing with fiber optic probes also can be adapted for use with either a fluorimeter or spectrofluorimeter. An analyte that is fluorescent can be monitored directly. For analytes that are not fluorescent, a suitable fluorescent probe molecule can be incorporated into the tip of the fiber optic probe. The analyte's reaction with the probe molecule leads to an increase or decrease in fluorescence.

INSTRUMENTS FOR MEASURING PHOSPHORESCENCE

Instrumentation for molecular phosphorescence must discriminate between phosphorescence and fluorescence. Because the lifetime for fluorescence is shorter than that for phosphorescence, discrimination is easily achieved by incorporating a delay between exciting the sample and measuring phosphorescent emission. Figure 10.54 shows how two out-of-phase choppers can be used to block emission from reaching the detector when the sample is being excited, and to prevent source radiation from reaching the sample while we are measuring the phosphorescent emission.

Because phosphorescence is such a slow process, we must prevent the excited state from relaxing by external conversion. Traditionally, this has been accomplished by dissolving the sample in a suitable organic solvent, usually a mixture of ethanol, isopentane, and diethylether. The resulting solution is frozen at liquid- N_2 temperatures, forming an optically clear solid. The solid matrix minimizes external conversion due to collisions between the analyte and the solvent. External conversion also is minimized

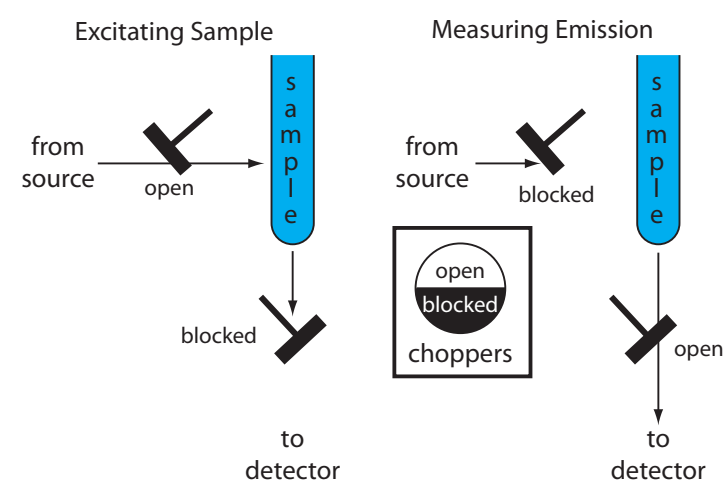


Figure 10.54 Schematic diagram showing how choppers are used to prevent fluorescent emission from interfering with the measurement of phosphorescent emission.

by immobilizing the sample on a solid substrate, making possible room temperature measurements. One approach is to place a drop of the solution containing the analyte on a small disc of filter paper. After drying the sample under a heat lamp, the sample is placed in the spectrofluorimeter for analysis. Other solid surfaces that have been used include silica gel, alumina, sodium acetate, and sucrose. This approach is particularly useful for the analysis of thin layer chromatography plates.

10F.3 Quantitative Applications

Molecular fluorescence and, to a lesser extent, phosphorescence have been used for the direct or indirect quantitative analysis of analytes in a variety of matrices. A direct quantitative analysis is possible when the analyte's fluorescent or phosphorescent quantum yield is favorable. When the analyte is not fluorescent or phosphorescent, or if the quantum yield is unfavorable, then an indirect analysis may be feasible. One approach is to react the analyte with a reagent to form a product with fluorescent or phosphorescent properties. Another approach is to measure a decrease in fluorescence or phosphorescence when the analyte is added to a solution containing a fluorescent or phosphorescent probe molecule. A decrease in emission is observed when the reaction between the analyte and the probe molecule enhances radiationless deactivation, or produces a nonemitting product. The application of fluorescence and phosphorescence to inorganic and organic analytes are considered in this section.

INORGANIC ANALYTES

Except for a few metal ions, most notably UO_2^+ , most inorganic ions are not sufficiently fluorescent for a direct analysis. Many metal ions may be determined indirectly by reacting with an organic ligand to form a fluorescent, or less commonly, a phosphorescent metal–ligand complex. One example is the reaction of Al^{3+} with the sodium salt of 2, 4, 3'-trihydroxyazobenzene-5'-sulfonic acid—also known as alizarin garnet R—which forms a fluorescent metal–ligand complex (Figure 10.55). The analysis is carried out using an excitation wavelength of 470 nm, monitoring fluorescence at 500 nm. [Table 10.12](#) provides additional examples of chelating reagents that form fluorescent metal–ligand complexes with metal ions. A few inorganic nonmetals are determined by their ability to decrease, or quench, the fluorescence of another species. One example is the analysis for F^- based on its ability to quench the fluorescence of the Al^{3+} –alizarin garnet R complex.

ORGANIC ANALYTES

As noted earlier, organic compounds containing aromatic rings generally are fluorescent and aromatic heterocycles are often phosphorescent. As shown in [Table 10.13](#), several important biochemical, pharmaceutical, and environmental compounds may be analyzed quantitatively by fluorimetry or

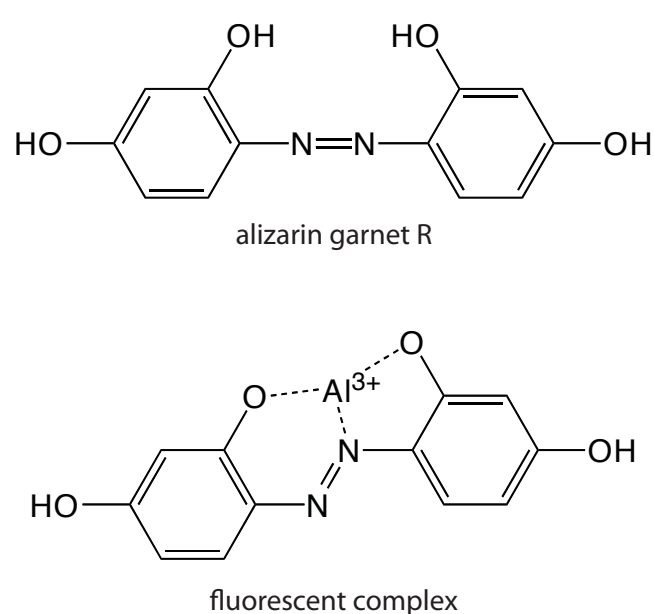


Figure 10.55 Structure of alizarin garnet R and its metal–ligand complex with Al^{3+} .

Table 10.12 Chelating Agents for the Fluorescence Analysis of Metal Ions

chelating agent	metal ions
8-hydroxyquinoline	Al^{3+} , Be^{2+} , Zn^{2+} , Li^+ , Mg^{2+} (and others)
flavonal	Zr^{2+} , Sn^{4+}
benzoin	$\text{B}_4\text{O}_7^{2-}$, Zn^{2+}
2', 3, 4', 5, 7-pentahydroxyflavone	Be^{2+}
2-(<i>o</i> -hydroxyphenyl) benzoxazole	Cd^{2+}

Table 10.13 Examples of Naturally Photoluminescent Organic Analytes

class	compounds (F = fluorescence; P = phosphorescence)
aromatic amino acids	phenylalanine (F) tyrosine (F) tryptophan (F, P)
vitamins	vitamin A (F) vitamin B2 (F) vitamin B6 (F) vitamin B12 (F) vitamin E (F) folic acid (F)
catecholamines	dopamine (F) norepinephrine (F)
pharmaceuticals and drugs	quinine (F) salicylic acid (F, P) morphine (F) barbiturates (F) LSD (F) codeine (P) caffeine (P) sulfanilamide (P)
environmental pollutants	pyrene (F) benzo[a]pyrene (F) organothiophosphorous pesticides (F) carbamate insecticides (F) DDT (P)

phosphorimetry. If an organic analyte is not naturally fluorescent or phosphorescent, it may be possible to incorporate it into a chemical reaction that produces a fluorescent or phosphorescent product. For example, the enzyme creatine phosphokinase can be determined by using it to catalyze the formation of creatine from phosphocreatine. Reacting the creatine with ninhydrin produces a fluorescent product of unknown structure.

STANDARDIZING THE METHOD

From [equation 10.28](#) and [equation 10.29](#) we know that the intensity of fluorescent or phosphorescent emission is a linear function of the analyte's concentration provided that the sample's absorbance of source radiation ($A = \epsilon bC$) is less than approximately 0.01. Calibration curves often are linear over four to six orders of magnitude for fluorescence and over two to four orders of magnitude for phosphorescence. For higher concentrations of analyte the calibration curve becomes nonlinear because the assumptions leading to [equation 10.28](#) and [equation 10.29](#) no longer apply. Nonlinearity may be observed for small concentrations of analyte due to the presence of fluorescent or phosphorescent contaminants. As discussed earlier, quantum efficiency is sensitive to temperature and sample matrix, both of which must be controlled when using external standards. In addition, emission intensity depends on the molar absorptivity of the photoluminescent species, which is sensitive to the sample matrix.

The best way to appreciate the theoretical and practical details discussed in this section is to carefully examine a typical analytical method. Although each method is unique, the following description of the determination of quinine in urine provides an instructive example of a typical procedure. The description here is based on Mule, S. J.; Hushin, P. L. *Anal. Chem.* **1971**, *43*, 708–711, and O'Reilly, J. E.; *J. Chem. Educ.* **1975**, *52*, 610–612.

[Figure 10.49](#) shows the fluorescence of the quinine in tonic water.

Representative Method 10.3

Determination of Quinine in Urine

DESCRIPTION OF METHOD

Quinine is an alkaloid used in treating malaria. It is a strongly fluorescent compound in dilute solutions of H_2SO_4 ($\Phi_f = 0.55$). Quinine's excitation spectrum has absorption bands at 250 nm and 350 nm and its emission spectrum has a single emission band at 450 nm. Quinine is rapidly excreted from the body in urine and is easily determined by measuring its fluorescence following its extraction from the urine sample.

PROCEDURE

Transfer a 2.00-mL sample of urine to a 15-mL test tube and adjust its pH to between 9 and 10 using 3.7 M NaOH. Add 4 mL of a 3:1 (v/v) mixture of chloroform and isopropanol and shake the contents of the test tube for one minute. Allow the organic and the aqueous (urine) layers to separate and transfer the organic phase to a clean test tube. Add 2.00 mL of 0.05 M H_2SO_4 to the organic phase and shake the contents for one minute. Allow the organic and the aqueous layers to separate and transfer the aqueous phase to the sample cell. Measure the fluorescent emission at 450 nm using an excitation wavelength of 350 nm. Determine the concentration of quinine in the urine sample using a calibration curve prepared with a set of external standards in 0.05 M H_2SO_4 , prepared from a 100.0 ppm solution of quinine in 0.05 M H_2SO_4 . Use distilled water as a blank.

QUESTIONS

1. Chloride ion quenches the intensity of quinine's fluorescent emission. For example, in the presence of 100 ppm NaCl (61 ppm Cl^-) quinine's emission intensity is only 83% of its emission intensity in the

absence of chloride. The presence of 1000 ppm NaCl (610 ppm Cl^-) further reduces quinine's fluorescent emission to less than 30% of its emission intensity in the absence of chloride. The concentration of chloride in urine typically ranges from 4600–6700 ppm Cl^- . Explain how this procedure prevents an interference from chloride.

The procedure uses two extractions. In the first of these extractions, quinine is separated from urine by extracting it into a mixture of chloroform and isopropanol, leaving the chloride behind in the original sample.

2. Samples of urine may contain small amounts of other fluorescent compounds, which interfere with the analysis if they are carried through the two extractions. Explain how you can modify the procedure to take this into account?

One approach is to prepare a blank using a sample of urine known to be free of quinine. Subtracting the blank's fluorescent signal from the measured fluorescence from urine samples corrects for the interfering compounds.

3. The fluorescent emission for quinine at 450 nm can be induced using an excitation frequency of either 250 nm or 350 nm. The fluorescent quantum efficiency is the same for either excitation wavelength. Quinine's absorption spectrum shows that ϵ_{250} is greater than ϵ_{350} . Given that quinine has a stronger absorbance at 250 nm, explain why its fluorescent emission intensity is greater when using 350 nm as the excitation wavelength.

From [equation 10.28](#) we know that I_f is a function of the following terms: k , Φ_f , P_0 , ϵ , b , and C . We know that Φ_f , b , and C are the same for both excitation wavelengths and that ϵ is larger for a wavelength of 250 nm; we can, therefore, ignore these terms. The greater emission intensity when using an excitation wavelength of 350 nm must be due to a larger value for P_0 or k . In fact, P_0 at 350 nm for a high-pressure Xe arc lamp is about 170% of that at 250 nm. In addition, the sensitivity of a typical photomultiplier detector (which contributes to the value of k) at 350 nm is about 140% of that at 250 nm.

Example 10.11

To evaluate the method described in [Representative Method 10.3](#), a series of external standard was prepared and analyzed, providing the results shown in the following table. All fluorescent intensities were corrected using a blank prepared from a quinine-free sample of urine. The fluorescent intensities are normalized by setting I_f for the highest concentration standard to 100.

[quinine] ($\mu\text{g/mL}$)	I_f
1.00	10.11
3.00	30.20
5.00	49.84
7.00	69.89
10.00	100.0

After ingesting 10.0 mg of quinine, a volunteer provided a urine sample 24-h later. Analysis of the urine sample gives an relative emission intensity of 28.16. Report the concentration of quinine in the sample in mg/L and the percent recovery for the ingested quinine.

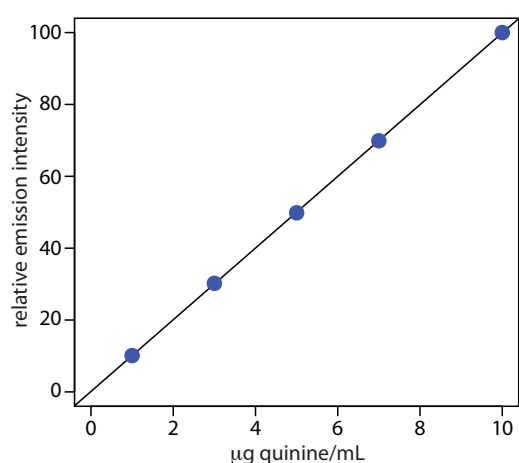
SOLUTION

Linear regression of the relative emission intensity versus the concentration of quinine in the standards gives a calibration curve with the following equation.

$$I_f = 0.124 + 9.978 \times \frac{\text{g quinine}}{\text{mL}}$$

Substituting the sample's relative emission intensity into the calibration equation gives the concentration of quinine as 2.81 $\mu\text{g/mL}$. Because the volume of urine taken, 2.00 mL, is the same as the volume of 0.05 M H_2SO_4 used in extracting quinine, the concentration of quinine in the urine also is 2.81 $\mu\text{g/mL}$. The recovery of the ingested quinine is

$$\frac{\frac{2.81 \text{ g}}{\text{ml urine}} \times 2.00 \text{ mL urine} \times \frac{1 \text{ mg}}{1000 \text{ g}}}{10.0 \text{ mg quinine ingested}} \times 100 = 0.0562\%$$



It can take up 10–11 days for the body to completely excrete quinine.

See [Figure 3.5](#) to review the meaning of macro and meso for describing samples, and the meaning of major, minor, and ultratrace for describing analytes.

10F.4 Evaluation of Photoluminescence Spectroscopy

SCALE OF OPERATION

Photoluminescence spectroscopy is used for the routine analysis of trace and ultratrace analytes in macro and meso samples. Detection limits for fluorescence spectroscopy are strongly influenced by the analyte's quantum yield. For an analyte with $\Phi_f > 0.5$, a picomolar detection limit is possible when using a high quality spectrofluorimeter. For example, the detection limit for quinine sulfate, for which Φ_f is 0.55, is generally between 1 part per billion and 1 part per trillion. Detection limits for phosphorescence are somewhat higher, with typical values in the nanomolar range for low-temperature phosphorimetry, and in the micromolar range for room-temperature phosphorimetry using a solid substrate.

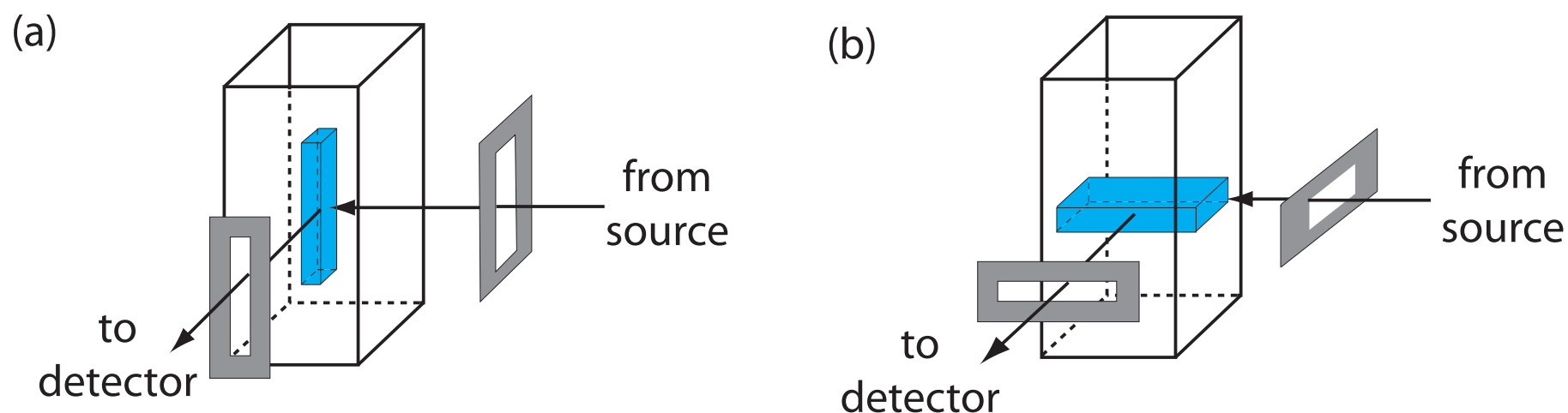


Figure 10.56 Use of slit orientation to change the volume from which fluorescence is measured: (a) vertical slit orientation; (b) horizontal slit orientation. Suppose the slit's dimensions are $0.1 \text{ mm} \times 3 \text{ mm}$. In (a) the dimensions of the **sampling volume** are $0.1 \text{ mm} \times 0.1 \text{ mm} \times 3 \text{ mm}$, or 0.03 mm^3 . For (b) the dimensions of the **sampling volume** are $0.1 \text{ mm} \times 3 \text{ mm} \times 3 \text{ mm}$, or 0.9 mm^3 , a 30-fold increase in the sampling volume.

ACCURACY

The accuracy of a fluorescence method is generally between 1–5% when spectral and chemical interferences are insignificant. Accuracy is limited by the same types of problems affecting other optical spectroscopic methods. In addition, accuracy is affected by interferences influencing the fluorescent quantum yield. The accuracy of phosphorescence is somewhat greater than that for fluorescence.

PRECISION

The relative standard deviation for fluorescence is usually between 0.5–2% when the analyte's concentration is well above its detection limit. Precision is usually limited by the stability of the excitation source. The precision for phosphorescence is often limited by reproducibility in preparing samples for analysis, with relative standard deviations of 5–10% being common.

SENSITIVITY

From [equation 10.28](#) and [equation 10.29](#) we know that the sensitivity of a fluorescent or phosphorescent method is influenced by a number of parameters. The importance of quantum yield and the effect of temperature and solution composition on Φ_f and Φ_p already have been considered. Besides quantum yield, the sensitivity of an analysis can be improved by using an excitation source that has a greater emission intensity, P_0 , at the desired wavelength, and by selecting an excitation wavelength that has a greater absorbance. Another approach for improving sensitivity is to increase the volume in the sample from which emission is monitored. Figure 10.56 shows how rotating a monochromator's slits from their usual vertical orientation to a horizontal orientation increases the sampling volume. The result can increase the emission from the sample by 5–30 \times .

SELECTIVITY

The selectivity of fluorescence and phosphorescence is superior to that of absorption spectrophotometry for two reasons: first, not every compound that absorbs radiation is fluorescent or phosphorescent; and, second, selectivity between an analyte and an interferent is possible if there is a difference in either their excitation or their emission spectra. The total emission intensity is a linear sum of that from each fluorescent or phosphorescent species. The analysis of a sample containing n components, therefore, can be accomplished by measuring the total emission intensity at n wavelengths.

TIME, COST, AND EQUIPMENT

As with other optical spectroscopic methods, fluorescent and phosphorescent methods provide a rapid means for analyzing samples and are capable of automation. Fluorimeters are relatively inexpensive, ranging from several hundred to several thousand dollars, and often are satisfactory for quantitative work. Spectrofluorimeters are more expensive, with models often exceeding \$50,000.

10G Atomic Emission Spectroscopy

The focus of this section is on the emission of ultraviolet and visible radiation following the thermal excitation of atoms. Atomic emission spectroscopy has a long history. Qualitative applications based on the color of flames were used in the smelting of ores as early as 1550 and were more fully developed around 1830 with the observation of atomic spectra generated by flame emission and spark emission.¹⁸ Quantitative applications based on the atomic emission from electric sparks were developed by Lockyer in the early 1870 and quantitative applications based on flame emission were pioneered by Lundegardh in 1930. Atomic emission based on emission from a plasma was introduced in 1964.

10G.1 Atomic Emission Spectra

Atomic emission occurs when a valence electron in a higher energy atomic orbital returns to a lower energy atomic orbital. Figure 10.57 shows a portion of the energy level diagram for sodium, which consists of a series of discrete lines at wavelengths corresponding to the difference in energy between two atomic orbitals.

The intensity of an atomic emission line, I_e , is proportional to the number of atoms, N^* , populating the excited state,

$$I_e = kN^* \quad 10.30$$

where k is a constant accounting for the efficiency of the transition. If a system of atoms is in thermal equilibrium, the population of excited state i is related to the total concentration of atoms, N , by the Boltzmann distribu-

¹⁸ Dawson, J. B. *J. Anal. At. Spectrosc.* **1991**, *6*, 93–98.

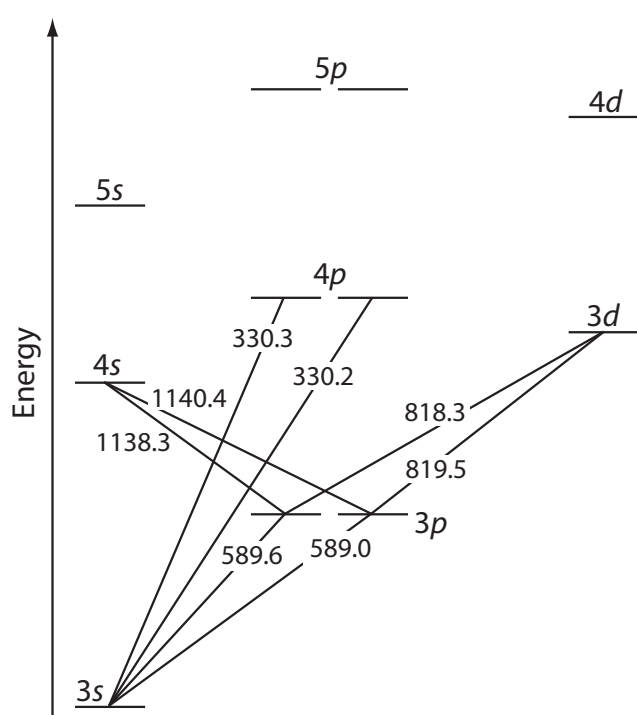


Figure 10.57 Valence shell energy level diagram for sodium. The wavelengths corresponding to several transitions are shown. Note that this is the same energy level diagram as [Figure 10.19](#).

tion. For many elements at temperatures of less than 5000 K the Boltzmann distribution is approximated as

$$N^* = N \left(\frac{g_i}{g_0} \right) e^{-E_i/kT} \quad 10.31$$

where g_i and g_0 are statistical factors that account for the number of equivalent energy levels for the excited state and the ground state, E_i is the energy of the excited state relative to a ground state energy, E_0 , of 0, k is Boltzmann's constant (1.3807×10^{-23} J/K), and T is the temperature in kelvin. From equation 10.31 we expect that excited states with lower energies have larger populations and more intense emission lines. We also expect emission intensity to increase with temperature.

10G.2 Equipment

An atomic emission spectrometer is similar in design to the instrumentation for atomic absorption. In fact, it is easy to adapt most flame atomic absorption spectrometers for atomic emission by turning off the hollow cathode lamp and monitoring the difference in the emission intensity when aspirating the sample and when aspirating a blank. Many atomic emission spectrometers, however, are dedicated instruments designed to take advantage of features unique to atomic emission, including the use of plasmas, arcs, sparks, and lasers as atomization and excitation sources, and an enhanced capability for multielemental analysis.

ATOMIZATION AND EXCITATION

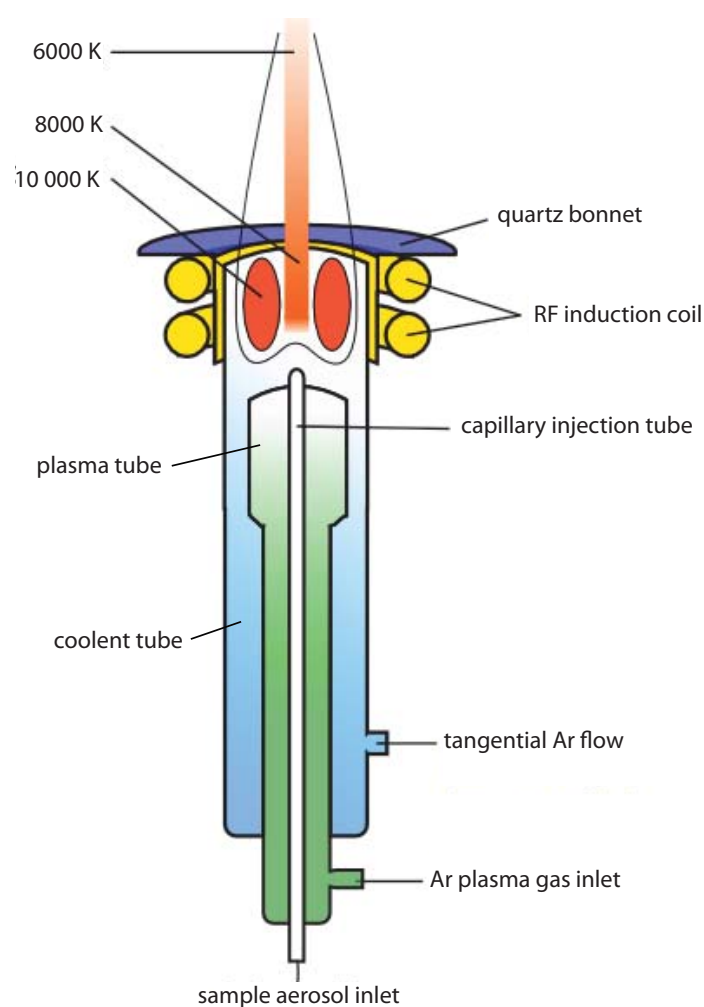
Atomic emission requires a means for converting a solid, liquid, or solution analyte into a free gaseous atom. The same source of thermal energy usually serves as the excitation source. The most common methods are flames and plasmas, both of which are useful for liquid or solution samples. Solid samples may be analyzed by dissolving in a solvent and using a flame or plasma atomizer.

FLAME SOURCES

Atomization and excitation in flame atomic emission is accomplished using the same nebulization and spray chamber assembly used in atomic absorption ([Figure 10.42](#)). The burner head consists of single or multiple slots, or a Meker style burner. Older atomic emission instruments often used a total consumption burner in which the sample is drawn through a capillary tube and injected directly into the flame.

PLASMA SOURCES

A **PLASMA** is a hot, partially ionized gas that contains an abundant concentration of cations and electrons. The plasmas used in atomic emission are formed by ionizing a flowing stream of argon gas, producing argon ions and



electrons. A plasma's high temperature results from resistive heating as the electrons and argon ions move through the gas. Because plasmas operate at much higher temperatures than flames, they provide better atomization and a higher population of excited states.

A schematic diagram of the inductively coupled plasma source (ICP) is shown in Figure 10.58. The ICP torch consists of three concentric quartz tubes, surrounded at the top by a radio-frequency induction coil. The sample is mixed with a stream of Ar using a nebulizer, and is carried to the plasma through the torch's central capillary tube. Plasma formation is initiated by a spark from a Tesla coil. An alternating radio-frequency current in the induction coils creates a fluctuating magnetic field that induces the argon ions and the electrons to move in a circular path. The resulting collisions with the abundant unionized gas give rise to resistive heating, providing temperatures as high as 10 000 K at the base of the plasma, and between 6000 and 8000 K at a height of 15–20 mm above the coil, where emission is usually measured. At these high temperatures the outer quartz tube must be thermally isolated from the plasma. This is accomplished by the tangential flow of argon shown in the schematic diagram.

Figure 10.58 Schematic diagram of an inductively coupled plasma torch. Source: modified from [Xvlun](http://commons.wikimedia.org/wiki/File:Xvlun) (commons.wikimedia.org).

MULTIELEMENTAL ANALYSIS

Atomic emission spectroscopy is ideally suited for multielemental analysis because all analytes in a sample are excited simultaneously. If the instrument includes a scanning monochromator, we can program it to move rapidly to an analyte's desired wavelength, pause to record its emission intensity, and then move to the next analyte's wavelength. This sequential analysis allows for a sampling rate of 3–4 analytes per minute.

Another approach to a multielemental analysis is to use a multichannel instrument that allows us to simultaneously monitor many analytes. A simple design for a multichannel spectrometer couples a monochromator with multiple detectors that can be positioned in a semicircular array around the monochromator at positions corresponding to the wavelengths for the analytes ([Figure 10.59](#)).

10G.3 Quantitative Applications

Atomic emission is widely used for the analysis of trace metals in a variety of sample matrices. The development of a quantitative atomic emission method requires several considerations, including choosing a source for atomization and excitation, selecting a wavelength and slit width, preparing the sample for analysis, minimizing spectral and chemical interferences, and selecting a method of standardization.

CHOICE OF ATOMIZATION AND EXCITATION SOURCE

Except for the alkali metals, detection limits when using an ICP are significantly better than those obtained with flame emission ([Table 10.14](#)).

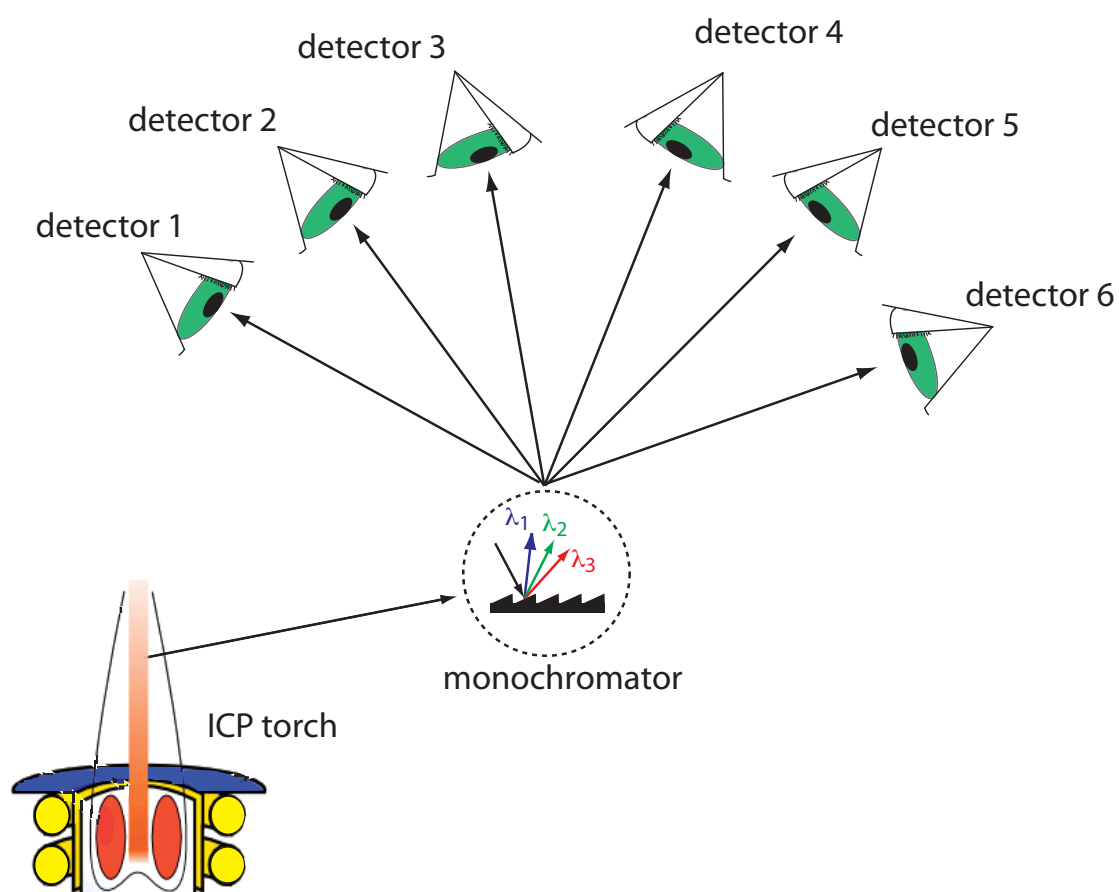


Figure 10.59 Schematic diagram of a multichannel atomic emission spectrometer for the simultaneous analysis of several elements. The ICP torch is modified from [Xvlun](https://commons.wikimedia.org/wiki/File:ICP_torch.jpg) (commons.wikimedia.org). Instruments may contain as many as 48–60 detectors.

Table 10.14 Detection Limits for Atomic Emission^a

element	detection limit in $\mu\text{g/mL}$	
	flame emission	ICP
Ag	2	0.2
Al	3	0.2
As	2000	2
Ca	0.1	0.0001
Cd	300	0.07
Co	5	0.1
Cr	1	0.08
Fe	10	0.09
Hg	150	1
K	0.01	30
Li	0.001	0.02
Mg	1	0.003
Mn	1	0.01
Na	0.01	0.1
Ni	10	0.2
Pb	0.2	1
Pt	2000	0.9
Sn	100	3
Zn	1000	0.1

^a Source: Parsons, M. L.; Major, S.; Forster, A. R.; *App. Spectrosc.* 1983, 37, 411–418.

Plasmas also are subject to fewer spectral and chemical interferences. For these reasons a plasma emission source is usually the better choice.

SELECTING THE WAVELENGTH AND SLIT WIDTH

The choice of wavelength is dictated by the need for sensitivity and the need to avoid interferences from the emission lines of other constituents in the sample. Because an analyte's atomic emission spectrum has an abundance of emission lines—particularly when using a high temperature plasma source—it is inevitable that there will be some overlap between emission lines. For example, an analysis for Ni using the atomic emission line at 349.30 nm is complicated by the atomic emission line for Fe at 349.06 nm. Narrower slit widths provide better resolution, but at the cost of less radiation reaching the detector. The easiest approach to selecting a wavelength is to record the sample's emission spectrum and look for an emission line that provides an intense signal and is resolved from other emission lines.

PREPARING THE SAMPLE

Flame and plasma sources are best suited for samples in solution and liquid form. Although a solid sample can be analyzed by directly inserting it into the flame or plasma, they usually are first brought into solution by digestion or extraction.

MINIMIZING SPECTRAL INTERFERENCES

The most important spectral interference is broad, background emission from the flame or plasma and emission bands from molecular species. This background emission is particularly severe for flames because the temperature is insufficient to break down refractory compounds, such as oxides and hydroxides. Background corrections for flame emission are made by scanning over the emission line and drawing a baseline (Figure 10.60). Because a plasma's temperature is much higher, a background interference due to molecular emission is less of a problem. Although emission from the plasma's core is strong, it is insignificant at a height of 10–30 mm above the core where measurements normally are made.

MINIMIZING CHEMICAL INTERFERENCES

Flame emission is subject to the same types of chemical interferences as atomic absorption. These interferences are minimized by adjusting the flame's composition and adding protecting agents, releasing agents, or ionization suppressors. An additional chemical interference results from **SELF-ABSORPTION**. Because the flame's temperature is greatest at its center, the concentration of analyte atoms in an excited state is greater at the flame's center than at its outer edges. If an excited state atom in the flame's center emits a photon while returning to its ground state, then a ground state atom in the cooler, outer regions of the flame may absorb the photon, decreasing

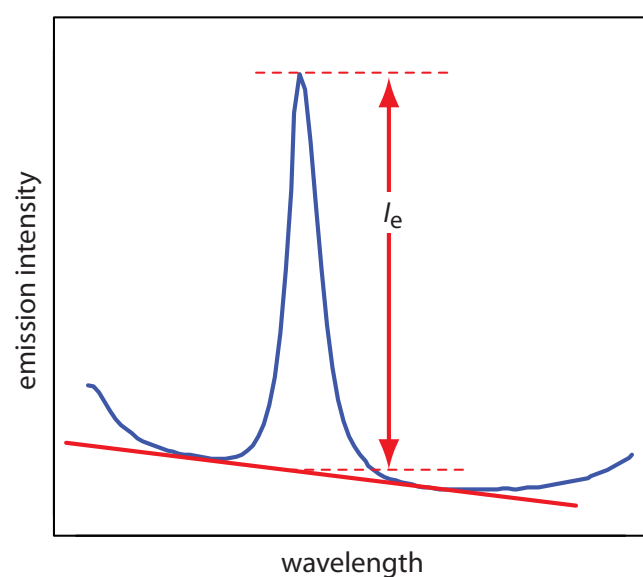


Figure 10.60 Method for correcting an analyte's emission for the flame's background emission.

the emission intensity. For higher concentrations of analyte self-absorption may invert the center of the emission band (Figure 10.61).

Chemical interferences with plasma sources generally are not significant because the plasma's higher temperature limits the formation of nonvolatile species. For example, PO_4^{3-} is a significant interferent when analyzing samples for Ca^{2+} by flame emission, but has a negligible effect when using a plasma source. In addition, the high concentration of electrons from the ionization of argon minimizes ionization interferences.

STANDARDIZING THE METHOD

From [equation 10.30](#) we know that emission intensity is proportional to the population of the analyte's excited state, N^* . If the flame or plasma is in thermal equilibrium, then the excited state population is proportional to the analyte's total population, N , through the Boltzmann distribution ([equation 10.31](#)).

A calibration curve for flame emission is usually linear over two to three orders of magnitude, with ionization limiting linearity when the analyte's concentrations is small and self-absorption limiting linearity for higher concentrations of analyte. When using a plasma, which suffers from fewer chemical interferences, the calibration curve often is linear over four to five orders of magnitude and is not affected significantly by changes in the matrix of the standards.

Emission intensity may be affected significantly by many parameters, including the temperature of the excitation source and the efficiency of atomization. An increase in temperature of 10 K, for example, produces a 4% increase in the fraction of Na atoms occupying the 3p excited state. This is potentially significant uncertainty that may limit the use of external standards. The method of internal standards can be used when variations in source parameters are difficult to control. To compensate for changes in the temperature of the excitation source, the internal standard is selected so that its emission line is close to the analyte's emission line. In addition, the internal standard should be subject to the same chemical interferences to compensate for changes in atomization efficiency. To accurately compensate for these errors the analyte and internal standard emission lines must be monitored simultaneously.

Representative Method 10.4

Determination of Sodium in a Salt Substitute

DESCRIPTION OF METHOD

Salt substitutes, which are used in place of table salt for individuals on low-sodium diets, replaces NaCl with KCl. Depending on the brand, fumaric acid, calcium hydrogen phosphate, or potassium tartrate also may be present. Although intended to be sodium-free, salt substitutes contain small amounts of NaCl as an impurity. Typically, the concentration of

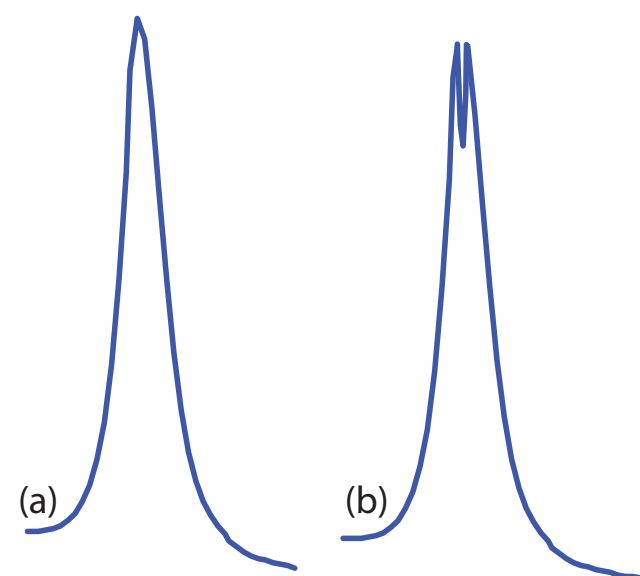


Figure 10.61 Atomic emission lines for (a) a low concentration of analyte, and (b) a high concentration of analyte showing the effect of self-absorption.

The best way to appreciate the theoretical and practical details discussed in this section is to carefully examine a typical analytical method. Although each method is unique, the following description of the determination of sodium in salt substitutes provides an instructive example of a typical procedure. The description here is based on Goodney, D. E. *J. Chem. Educ.* **1982**, *59*, 875–876.

sodium in a salt substitute is about $100 \mu\text{g/g}$. The exact concentration of sodium is easily determined by flame atomic emission. Because it is difficult to match the matrix of the standards to that of the sample, the analysis is accomplished by the method of standard additions.

PROCEDURE

A sample is prepared by placing an approximately 10-g portion of the salt substitute in 10 mL of 3 M HCl and 100 mL of distilled water. After the sample has dissolved, it is transferred to a 250-mL volumetric flask and diluted to volume with distilled water. A series of standard additions is prepared by placing 25-mL portions of the diluted sample into separate 50-mL volumetric flasks, spiking each with a known amount of an approximately 10 mg/L standard solution of Na^+ , and diluting to volume. After zeroing the instrument with an appropriate blank, the instrument is optimized at a wavelength of 589.0 nm while aspirating a standard solution of Na^+ . The emission intensity is measured for each of the standard addition samples and the concentration of sodium in the salt substitute is reported in $\mu\text{g/g}$.

QUESTIONS

1. Potassium ionizes more easily than sodium. What problem might this present if you use external standards prepared from a stock solution of 10 mg Na/L instead of using a set of standard additions?

Because potassium is present at a much higher concentration than sodium, its ionization suppresses the ionization of sodium. Normally suppressing ionization is a good thing because it increases emission intensity. In this case, however, the difference between the matrix of the standards and the sample's matrix means that the sodium in a standard experiences more ionization than an equivalent amount of sodium in a sample. The result is a determinate error.

2. One way to avoid a determinate error when using external standards is to match the matrix of the standards to that of the sample. We could, for example, prepare external standards using reagent grade KCl to match the matrix to that of the sample. Why is this not a good idea for this analysis?

Sodium is a common contaminant, which is found in many chemicals. Reagent grade KCl, for example, may contain 40–50 $\mu\text{g Na/g}$. This is a significant source of sodium, given that the salt substitute contains approximately $100 \mu\text{g Na/g}$.

3. Suppose you decide to use an external standardization. Given the answer to the previous questions, is the result of your analysis likely to underestimate or overestimate the amount of sodium in the salt substitute?

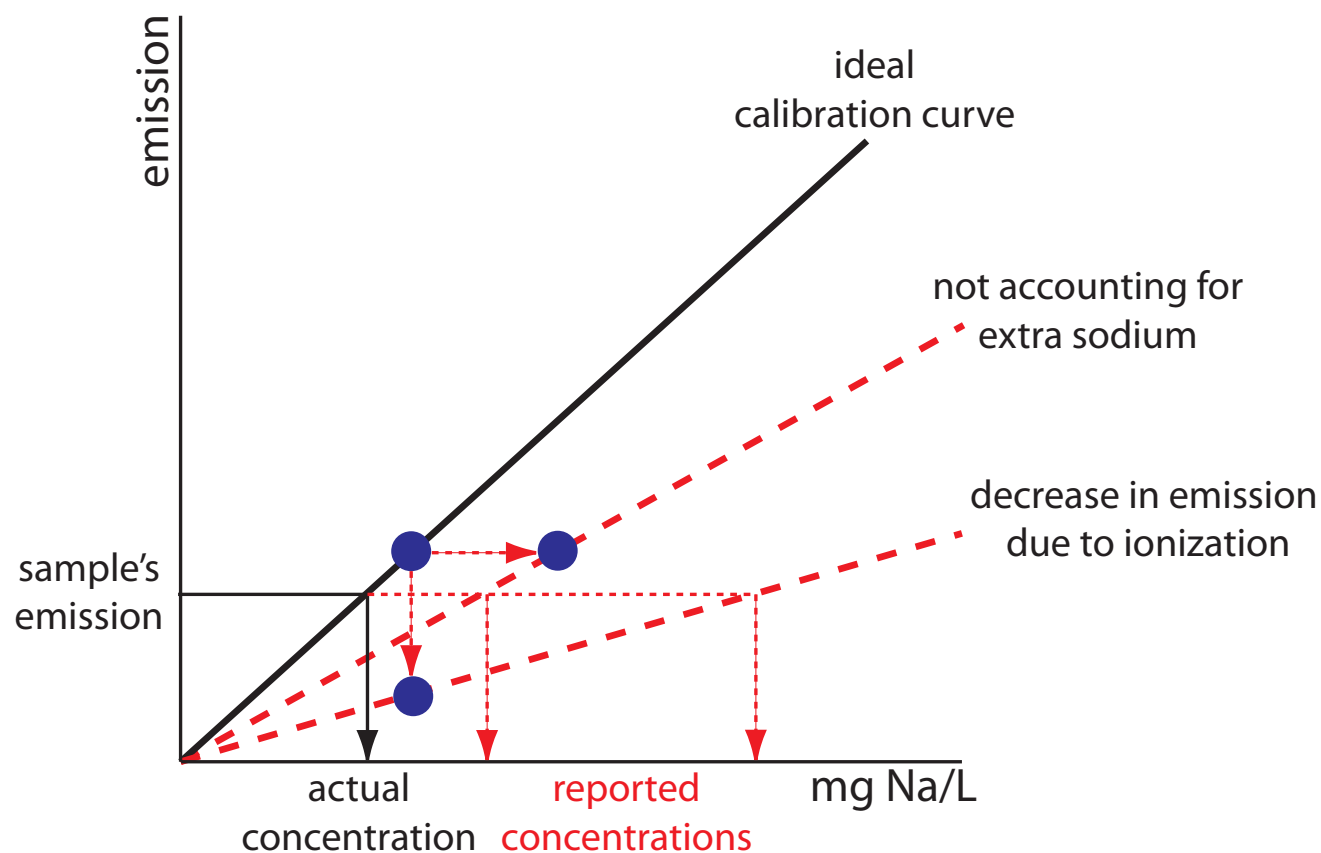


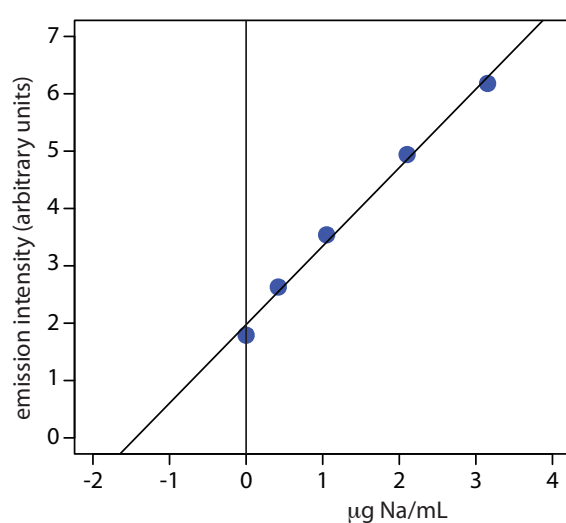
Figure 10.62 External standards calibration curves for the flame atomic emission analysis of Na in a salt substitute. The solid black line shows the ideal calibration curve assuming matrix matching of samples and standards with pure KCl. The lower of the two dashed red lines shows the effect of failing to add KCl to the external standards, which decreases emission. The other dashed red line shows the effect of using KCl that is contaminated with NaCl, which causes us to underestimate the concentration of Na in the standards. In both cases, the result is a positive determinate error in the analysis of samples.

The solid black line in Figure 10.62 shows the ideal calibration curve assuming that we match the matrix of the standards to the sample's matrix, and that we do so without adding an additional sodium. If we prepare the external standards without adding KCl, the emission for each standard decreases due to increased ionization. This is shown by the lower of the two dashed red lines. Preparing the standards by adding reagent grade KCl increases the concentration of sodium due to its contamination. Because we underestimate the actual concentration of sodium in the standards, the resulting calibration curve is shown by the other dashed red line. In both cases, the sample's emission results in our overestimating the concentration of sodium in the sample.

4. One problem with analyzing salt samples is their tendency to clog the aspirator and burner assembly. What effect does this have on the analysis?

Clogging the aspirator and burner assembly decreases the rate of aspiration, which decreases the analyte's concentration in the flame. The result is a decrease in the emission intensity and a negative determinate error.

See [Section 5C.3](#) in Chapter 5 to review the method of standard additions.



See [Figure 3.5](#) to review the meaning of macro and meso for describing samples, and the meaning of major, minor, and ultratrace for describing analytes.

Example 10.12

To evaluate the method described in [Representative Method 10.4](#), a series of standard additions is prepared using a 10.0077-g sample of a salt substitute. The results of a flame atomic emission analysis of the standards is shown here.¹⁹

added Na ($\mu\text{g/mL}$)	I_e (arb. units)
0.000	1.79
0.420	2.63
1.051	3.54
2.102	4.94
3.153	6.18

What is the concentration of sodium, in $\mu\text{g/g}$, in the salt substitute.

SOLUTION

Linear regression of emission intensity versus the concentration of added Na gives a standard additions calibration curve with the following equation.

$$I_e = 1.97 + 1.37 \times \frac{\text{g Na}}{\text{mL}}$$

The concentration of sodium in the sample is equal to the absolute value of the calibration curve's x -intercept. Substituting zero for the emission intensity and solving for sodium's concentration gives a result of $1.44 \mu\text{g Na/mL}$. The concentration of sodium in the salt substitute is

$$\frac{1.44 \text{ g Na}}{\text{mL}} \times \frac{50.00 \text{ mL}}{25.00 \text{ mL}} \times 250.0 \text{ mL} = 71.9 \text{ g Na/g}$$

10G.4 Evaluation of Atomic Emission Spectroscopy

SCALE OF OPERATION

The scale of operations for atomic emission is ideal for the direct analysis of trace and ultratrace analytes in macro and meso samples. With appropriate dilutions, atomic emission also can be applied to major and minor analytes.

ACCURACY

When spectral and chemical interferences are insignificant, atomic emission is capable of producing quantitative results with accuracies of between

¹⁹ Goodney, D. E. *J. Chem. Educ.* **1982**, *59*, 875–876

1–5%. Accuracy frequently is limited by chemical interferences. Because the higher temperature of a plasma source gives rise to more emission lines, the accuracy of using plasma emission often is limited by stray radiation from overlapping emission lines.

PRECISION

For samples and standards in which the analyte's concentration exceeds the detection limit by at least a factor of 50, the relative standard deviation for both flame and plasma emission is about 1–5%. Perhaps the most important factor affecting precision is the stability of the flame's or the plasma's temperature. For example, in a 2500 K flame a temperature fluctuation of ± 2.5 K gives a relative standard deviation of 1% in emission intensity. Significant improvements in precision may be realized when using internal standards.

SENSITIVITY

Sensitivity is strongly influenced by the temperature of the excitation source and the composition of the sample matrix. Sensitivity is optimized by aspirating a standard solution of analyte and maximizing the emission by adjusting the flame's composition and the height from which we monitor the emission. Chemical interferences, when present, decrease the sensitivity of the analysis. The sensitivity of plasma emission is less affected by the sample matrix. In some cases a calibration curve prepared using standards in a matrix of distilled water can be used for samples with more complex matrices.

SELECTIVITY

The selectivity of atomic emission is similar to that of atomic absorption. Atomic emission has the further advantage of rapid sequential or simultaneous analysis.

TIME, COST, AND EQUIPMENT

Sample throughput with atomic emission is very rapid when using automated systems capable of multielemental analysis. For example, sampling rates of 3000 determinations per hour have been achieved using a multi-channel ICP, and 300 determinations per hour with a sequential ICP. Flame emission is often accomplished using an atomic absorption spectrometer, which typically costs between \$10,000–\$50,000. Sequential ICP's range in price from \$55,000–\$150,000, while an ICP capable of simultaneous multielemental analysis costs between \$80,000–\$200,000. Combination ICP's that are capable of both sequential and simultaneous analysis range in price from \$150,000–\$300,000. The cost of Ar, which is consumed in significant quantities, can not be overlooked when considering the expense of operating an ICP.

10H Spectroscopy Based on Scattering

The blue color of the sky during the day and the red color of the sun at sunset result from the scattering of light by small particles of dust, molecules of water, and other gases in the atmosphere. The efficiency of a photon's scattering depends on its wavelength. We see the sky as blue during the day because violet and blue light scatter to a greater extent than other, longer wavelengths of light. For the same reason, the sun appears red at sunset because red light is less efficiently scattered and is more likely to pass through the atmosphere than other wavelengths of light. The scattering of radiation has been studied since the late 1800s, with applications beginning soon thereafter. The earliest quantitative applications of scattering, which date from the early 1900s, used the elastic scattering of light by colloidal suspensions to determine the concentration of colloidal particles.

10H.1 Origin of Scattering

If we send a focused, monochromatic beam of radiation with a wavelength λ through a medium of particles with dimensions $<1.5\lambda$, the radiation scatters in all directions. For example, visible radiation of 500 nm is scattered by particles as large as 750 nm in the longest dimension. Two general categories of scattering are recognized. In elastic scattering, radiation is first absorbed by the particles and then emitted without undergoing change in the radiation's energy. When the radiation emerges with a change in energy, the scattering is said to be inelastic. Only elastic scattering is considered in this text.

Elastic scattering is divided into two types: Rayleigh, or small-particle scattering, and large-particle scattering. Rayleigh scattering occurs when the scattering particle's largest dimension is less than 5% of the radiation's wavelength. The intensity of the scattered radiation is proportional to its frequency to the fourth power, ν^4 —accounting for the greater scattering of blue light than red light—and is symmetrically distributed (Figure 10.63a). For larger particles, scattering increases in the forward direction and decreases in the backwards direction as the result of constructive and destructive interferences (Figure 10.63b).

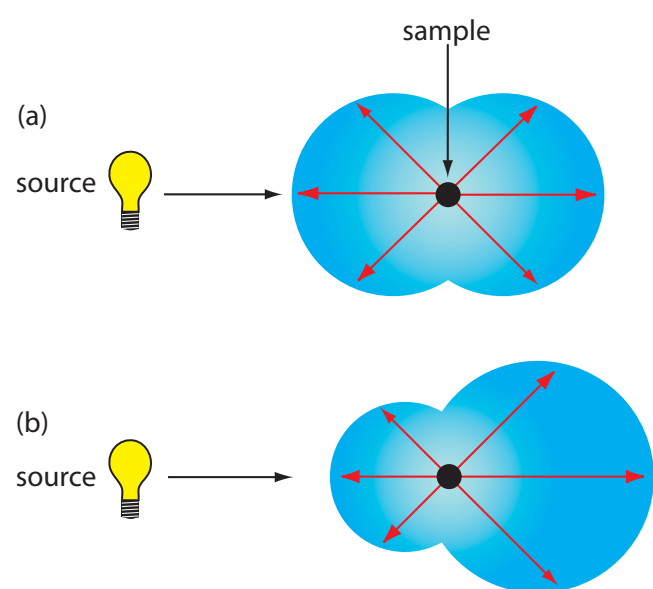


Figure 10.63 Distribution of radiation for (a) Rayleigh, or small-particle scattering, and (b) large-particle scattering.

10H.2 Turbidimetry and Nephelometry

Turbidimetry and nephelometry are two techniques based on the elastic scattering of radiation by a suspension of colloidal particles. In **TURBIDIMETRY** the detector is placed in line with the source and the decrease in the radiation's transmitted power is measured. In **NEPHELOMETRY** scattered radiation is measured at an angle of 90° to the source. The similarity of turbidimetry to absorbance and of nephelometry to fluorescence is evident in the instrumental designs shown in [Figure 10.64](#). In fact, turbidity can be measured using a UV/Vis spectrophotometer and a spectrofluorimeter is suitable for nephelometry.

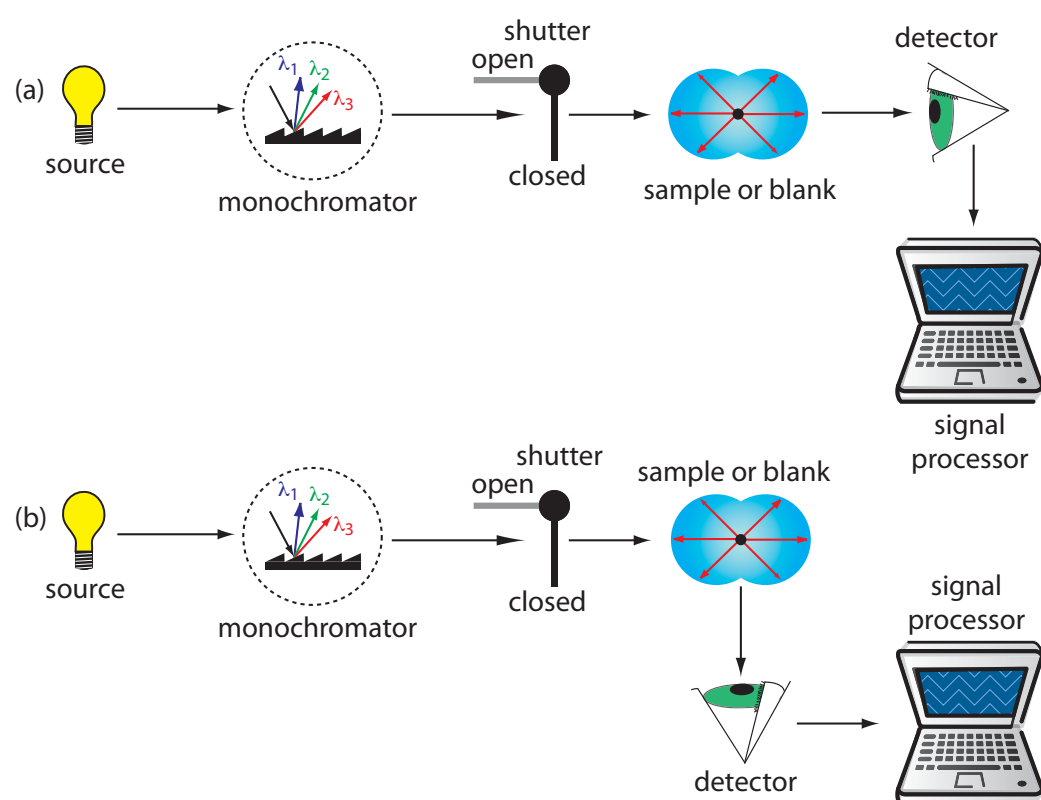


Figure 10.64 Schematic diagrams for (a) a turbidimeter, and (b) a nephelometer.

TURBIDIMETRY OR NEPHELOMETRY?

When developing a scattering method the choice of using turbidimetry or nephelometry is determined by two factors. The most important consideration is the intensity of the scattered radiation relative to the intensity of the source's radiation. If the solution contains a small concentration of scattering particles the intensity of the transmitted radiation, I_T , is approximately the same as the intensity of the source's radiation, I_0 . As we learned earlier in the section on molecular absorption, there is substantial uncertainty in determining a small difference between two intense signals. For this reason, nephelometry is a more appropriate choice for samples containing few scattering particles. Turbidimetry is a better choice when the sample contains a high concentration of scattering particles.

A second consideration in choosing between turbidimetry and nephelometry is the size of the scattering particles. For nephelometry, the intensity of scattered radiation at 90° increases when the particles are small and Rayleigh scattering is in effect. For larger particles, as shown in [Figure 10.63](#), the intensity of scattering is diminished at 90° . When using an ultraviolet or visible source of radiation, the optimum particle size is $0.1\text{--}1\ \mu\text{m}$. The size of the scattering particles is less important for turbidimetry where the signal is the relative decrease in transmitted radiation. In fact, turbidimetric measurements are still feasible even when the size of the scattering particles results in an increase in reflection and refraction, although a linear relationship between the signal and the concentration of scattering particles may no longer hold.

DETERMINING CONCENTRATION BY TURBIDIMETRY

In turbidimetry the measured transmittance, T , is the ratio of the intensity of source radiation transmitted by the sample, I_T , to the intensity of source radiation transmitted by a blank, I_0 .

$$T = \frac{I_T}{I_0}$$

The relationship between transmittance and the concentration of the scattering particles is similar to that given by Beer's law

$$-\log T = kbC \quad 10.32$$

where C is the concentration of the scattering particles in mass per unit volume (w/v), b is the pathlength, and k is a constant that depends on several factors, including the size and shape of the scattering particles and the wavelength of the source radiation. As with Beer's law, equation 10.32 may show appreciable deviations from linearity. The exact relationship is established by a calibration curve prepared using a series of standards containing known concentrations of analyte.

DETERMINING CONCENTRATION BY NEPHELOMETRY

In nephelometry the relationship between the intensity of scattered radiation, I_S , and the concentration of scattering particles is

$$I_S = k_S I_0 C \quad 10.33$$

where k_S is an empirical constant for the system and I_0 is the source radiation's intensity. The value of k_S is determined from a calibration curve prepared using a series of standards containing known concentrations of analyte.

SELECTING A WAVELENGTH FOR THE INCIDENT RADIATION

The choice of wavelength is based primarily on the need to minimize potential interferences. For turbidimetry, where the incident radiation is transmitted through the sample, a monochromator or filter allow us to avoid wavelengths that are absorbed by the sample. For nephelometry, the absorption of incident radiation is not a problem unless it induces fluorescence from the sample. With a nonfluorescent sample there is no need for wavelength selection, and a source of white light may be used as the incident radiation. For both techniques, other considerations in choosing a wavelength including the effect of wavelength on scattering intensity, the transducer's sensitivity, and the source's intensity. For example, many common photon transducers are more sensitive to radiation at 400 nm than at 600 nm.

PREPARING THE SAMPLE FOR ANALYSIS

Although equation 10.32 and equation 10.33 relate scattering to the concentration of scattering particles, the intensity of scattered radiation is also influenced by the particle's size and shape. Samples containing the same number of scattering particles may show significantly different values for $-\log T$ or I_S , depending on the average diameter of the particles. For a quan-

titative analysis, therefore, it is necessary to maintain a uniform distribution of particle sizes throughout the sample and between samples and standards.

Most turbidimetric and nephelometric methods rely on precipitation to form the scattering particles. As we learned in Chapter 8, the properties of a precipitate are determined by the conditions under which it forms. To maintain a reproducible distribution of particle sizes between samples and standards, it is necessary to control parameters such as the concentration of reagents, the order of adding reagents, the pH and temperature, the agitation or stirring rate, the ionic strength, and the time between the precipitate's initial formation and the measurement of transmittance or scattering. In many cases a surface-active agent—such as glycerol, gelatin, or dextrin—is added to stabilize the precipitate in a colloidal state and to prevent the coagulation of the particles.

APPLICATIONS

Turbidimetry and nephelometry are widely used to determine the clarity of water, beverages, and food products. For example, a nephelometric determination of the turbidity of water compares the sample's scattering to the scattering of a set of standards. The primary standard for measuring turbidity is formazin, which is an easily prepared, stable polymer suspension (Figure 10.65).²⁰ Formazin prepared by mixing a 1 g/100 mL solution of hydrazine sulfate, $N_2H_4 \cdot H_2SO_4$, with a 10 g/100 mL solution of hexamethylenetetramine produces a suspension that is defined as 4000 nephelometric turbidity units (NTU). A set of standards with NTUs between 0 and 40 is prepared and used to construct a calibration curve. This method is readily adapted to the analysis of the clarity of orange juice, beer, and maple syrup.

A number of inorganic cations and anions can be determined by precipitating them under well-defined conditions and measuring the transmittance or scattering of radiation from the precipitated particles. The trans-

²⁰ Hach, C. C.; Bryant, M. "Turbidity Standards," Technical Information Series, Booklet No. 12, Hach Company: Loveland, CO, 1995.

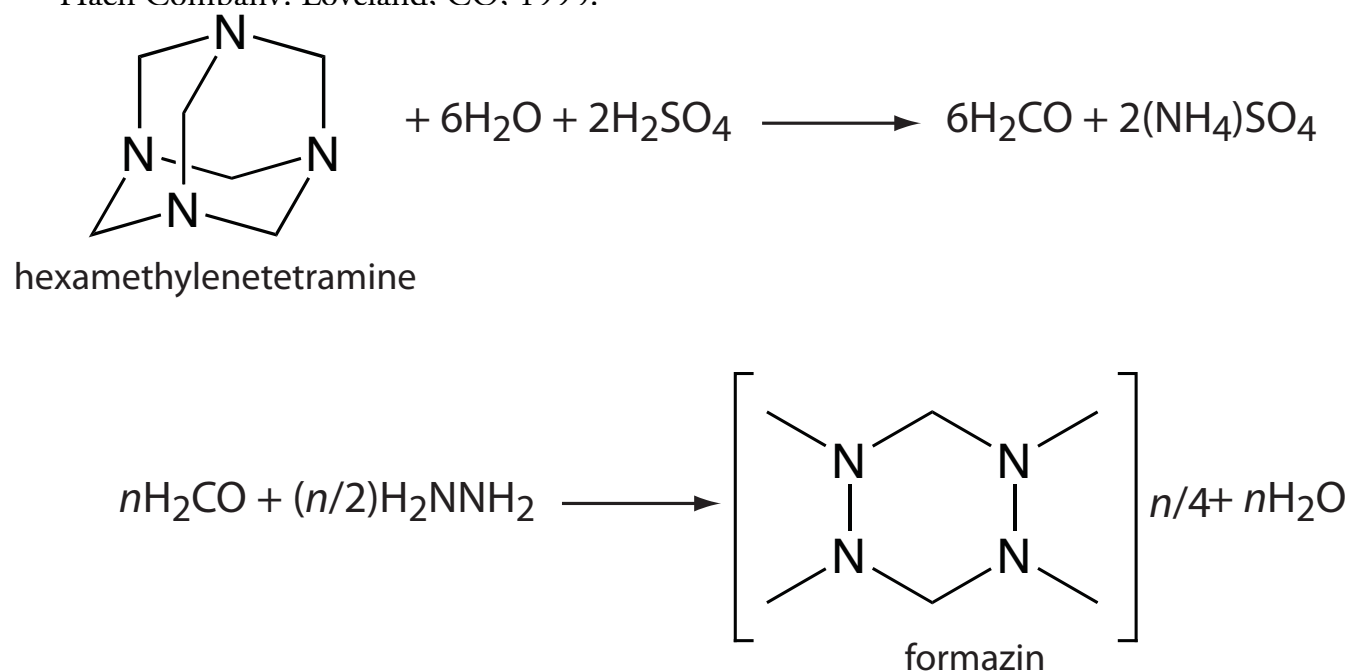


Figure 10.65 Scheme for preparing formazin for use as a turbidity standard.

Table 10.15 Examples of Analytes Determined by Turbidimetry or Nephelometry

analyte	precipitant	precipitate
Ag ⁺	NaCl	AgCl
Ca ²⁺	Na ₂ C ₂ O ₄	CaC ₂ O ₄
Cl ⁻	AgNO ₃	AgCl
CN ⁻	AgNO ₃	AgCN
CO ₃ ²⁻	BaCl ₂	BaCO ₃
F ⁻	CaCl ₂	CaF ₂
SO ₄ ²⁻	BaCl ₂	BaSO ₄

mittance or scattering, as given by [equation 10.32](#) or [equation 10.33](#) is proportional to the concentration of the scattering particles, which, in turn, is related by the stoichiometry of the precipitation reaction to the analyte's concentration. Examples of analytes that have been determined in this way are listed in Table 10.15.

Representative Method 10.5

Turbidimetric Determination of Sulfate in Water

DESCRIPTION OF METHOD

Adding BaCl₂ to an acidified sample precipitates SO₄²⁻ as BaSO₄. The concentration of SO₄²⁻ may be determined either by turbidimetry or by nephelometry using an incident source of radiation of 420 nm. External standards containing known concentrations of SO₄²⁻ are used to standardize the method.

PROCEDURE

Transfer a 100-mL sample to a 250-mL Erlenmeyer flask along with 20.00 mL of a buffer. For samples containing >10 mg SO₄²⁻/L, the buffer contains 30 g of MgCl₂·6H₂O, 5 g of CH₃COONa·3H₂O, 1.0 g of KNO₃, and 20 mL of glacial CH₃COOH per liter. The buffer for samples containing <10 mg SO₄²⁻/L is the same except for the addition of 0.111 g of Na₂SO₄ per L.

Place the sample and the buffer on a magnetic stirrer operated at the same speed for all samples and standards. Add a spoonful of 20–30 mesh BaCl₂, using a measuring spoon with a capacity of 0.2–0.3 mL, to precipitate the SO₄²⁻ as BaSO₄. Begin timing when the BaCl₂ is added and stir the suspension for 60 ± 2 s. Measure the transmittance or scattering intensity 5.0 ± 0.5 min after the end of stirring.

Prepare a calibration curve over the range 0–40 mg SO₄²⁻/L by diluting a 100-mg SO₄²⁻/L standard. Treat the standards using the procedure described above for the sample. Prepare a calibration curve and use it to determine the amount of sulfate in the sample.

The best way to appreciate the theoretical and practical details discussed in this section is to carefully examine a typical analytical method. Although each method is unique, the following description of the determination of sulfate in water provides an instructive example of a typical procedure. The description here is based on Method 4500-SO₄²⁻-C in *Standard Methods for the Analysis of Water and Wastewater*, American Public Health Association: Washington, D. C. 20th Ed., 1998.

QUESTIONS

1. What is the purpose of the buffer?

If the precipitate's particles are too small, I_T may be too small to measure reliably. Because rapid precipitation favors the formation of microcrystalline particles of BaSO_4 , we use conditions that favor precipitate growth over nucleation. The buffer's high ionic strength and its acidity favors precipitate growth and prevents the formation of microcrystalline BaSO_4 .

2. Why is it important to use the same stirring rate and time for the samples and standards?

How fast and how long we stir the sample after adding BaCl_2 influence the size of the precipitate's particles.

3. Many natural waters have a slight color due to the presence of humic and fulvic acids, and may contain suspended matter (Figure 10.66). Explain why these might interfere with the analysis for sulfate. For each interferent, suggest a method for minimizing its effect on the analysis.

Suspended matter in the sample contributes to the scattering, resulting in a positive determinate error. We can eliminate this interference by filtering the sample prior to its analysis. A sample that is colored may absorb some of the source's radiation, leading to a positive determinate error. We can compensate for this interference by taking a sample through the analysis without adding BaCl_2 . Because no precipitate forms, we can use the transmittance of this sample blank to correct for the interference.

4. Why is Na_2SO_4 added to the buffer for samples containing $<10 \text{ mg SO}_4^{2-}/\text{L}$?

The uncertainty in a calibration curve is smallest near its center. If a sample has a high concentration of SO_4^{2-} , we can dilute it so that its concentration falls near the middle of the calibration curve. For a sample with a small concentration of SO_4^{2-} , the buffer increases the concentration of sulfate by

$$\frac{0.111 \text{ g Na}_2\text{SO}_4}{\text{L}} \times \frac{96.06 \text{ g SO}_4^{2-}}{142.04 \text{ g Na}_2\text{SO}_4} \times \frac{1000 \text{ mg}}{\text{g}} \times \frac{20.00 \text{ mL}}{250.0 \text{ mL}} = \frac{6.00 \text{ mg SO}_4^{2-}}{\text{L}}$$

After using the calibration curve to determine the amount of sulfate in the sample, subtracting $6.00 \text{ mg SO}_4^{2-}/\text{L}$ corrects the result.

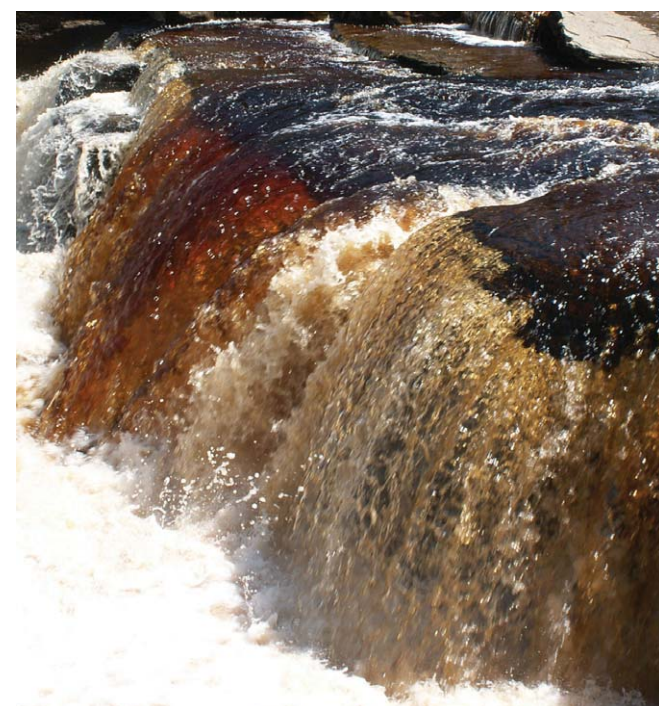


Figure 10.66 Waterfall on the River Swale in Richmond, England. The river, which flows out of the moors in the Yorkshire Dales, is brown in color as the result of organic matter that leaches from the peat found in the moors.

Example 10.13

To evaluate the method described in [Representative Method 10.5](#), a series of external standard was prepared and analyzed, providing the results shown in the following table.

mg SO ₄ ²⁻ /L	transmittance
0.00	1.00
10.00	0.646
20.00	0.417
30.00	0.269
40.00	0.174

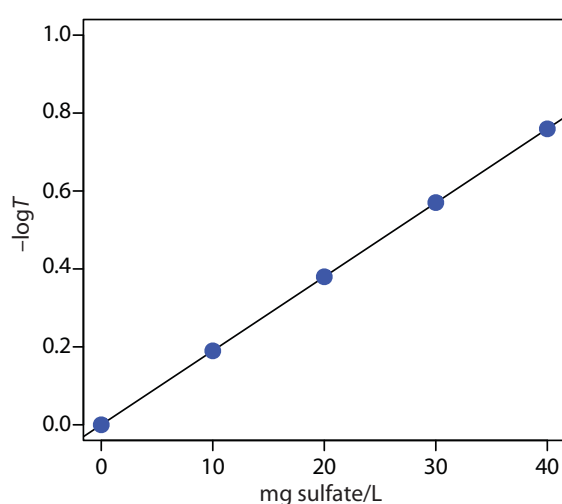
Analysis of a 100.0-mL sample of a surface water gives a transmittance of 0.538. What is the concentration of sulfate in the sample?

SOLUTION

Linear regression of $-\log T$ versus concentration of SO₄²⁻ gives a standardization equation of

$$-\log T = -1.04 \times 10^{-5} + 0.0190 \times \frac{\text{mg SO}_4^{2-}}{\text{L}}$$

Substituting the sample's transmittance into the calibration curve's equation gives the concentration of sulfate in sample as 14.2 mg SO₄²⁻/L.



As you review this chapter, try to define a key term in your own words. Check your answer by clicking on the key term, which will take you to the page where it was first introduced. Clicking on the **KEY TERM** there, will bring you back to this page so that you can continue with another key term.

101 Key Terms

absorbance	absorbance spectrum	absorptivity
amplitude	attenuated total reflectance	atomization
background correction	Beer's law	chemiluminescence
chromophore	continuum source	dark current
double-beam	effective bandwidth	electromagnetic radiation
electromagnetic spectrum	emission	emission spectrum
excitation spectrum	external conversion	Fellgett's advantage
fiber-optic probe	filter	filter photometer
fluorescence	fluorescent quantum yield	fluorimeter
frequency	graphite furnace	interferogram
interferometer	internal conversion	intersystem crossing
ionization suppressor	Jacquinot's advantage	lifetime
line source	method of continuous variations	molar absorptivity
mole-ratio method	monochromatic	monochromator
nephelometry	nominal wavelength	phase angle

phosphorescence	phosphorescent quantum yield	photodiode array
photoluminescence	photon	plasma
polychromatic	protecting agent	radiationless deactivation
relaxation	releasing agent	resolution
self-absorption	signal averaging	signal processor
signal-to-noise ratio	single-beam	singlet excited state
slope-ratio method	spectral searching	spectrofluorimeter
spectrophotometer	spectroscopy	stray radiation
transducer	transmittance	triplet excited state
turbidimetry	vibrational relaxation	wavelength
wavenumber		

10J Chapter Summary

The spectrophotometric methods of analysis covered in this chapter include those based on the absorption, emission, or scattering of electromagnetic radiation. When a molecule absorbs UV/Vis radiation it undergoes a change in its valence shell configuration. A change in vibrational energy results from the absorption of IR radiation. Experimentally we measure the fraction of radiation transmitted, T , by the sample. Instrumentation for measuring absorption requires a source of electromagnetic radiation, a means for selecting a wavelength, and a detector for measuring transmittance. Beer's law relates absorbance to both transmittance and to the concentration of the absorbing species ($A = -\log T = \epsilon b C$).

In atomic absorption we measure the absorption of radiation by gas phase atoms. Samples are atomized using thermal energy from either a flame or a graphite furnace. Because the width of an atom's absorption band is so narrow, the continuum sources common for molecular absorption can not be used. Instead, a hollow cathode lamp provides the necessary line source of radiation. Atomic absorption suffers from a number of spectral and chemical interferences. The absorption or scattering of radiation from the sample's matrix are important spectral interferences that may be minimized by background correction. Chemical interferences include the formation of nonvolatile forms of the analyte and ionization of the analyte. The former interference is minimized by using a releasing agent or a protecting agent, and an ionization suppressor helps minimize the latter interference.

When a molecule absorbs radiation it moves from a lower energy state to a higher energy state. In returning to the lower energy state the molecule may emit radiation. This process is called photoluminescence. One form of photoluminescence is fluorescence in which the analyte emits a photon without undergoing a change in its spin state. In phosphorescence, emission occurs with a change in the analyte's spin state. For low concentrations

of analyte, both fluorescent and phosphorescent emission intensities are a linear function of the analyte's concentration. Thermally excited atoms also emit radiation, forming the basis for atomic emission spectroscopy. Thermal excitation is achieved using either a flame or a plasma.

Spectroscopic measurements may also involve the scattering of light by a particulate form of the analyte. In turbidimetry, the decrease in the radiation's transmission through the sample is measured and related to the analyte's concentration through an equation similar to Beer's law. In nephelometry we measure the intensity of scattered radiation, which varies linearly with the analyte's concentration.

10K Problems

1. Provide the missing information in the following table.

wavelength (m)	frequency (s ⁻¹)	wavenumber (cm ⁻¹)	energy (J)
4.50×10^{-9}			
	1.33×10^{15}		
		3215	
			7.20×10^{-19}

2. Provide the missing information in the following table.

[analyte] (M)	absorbance	%T	molar absorptivity (M ⁻¹ cm ⁻¹)	pathlength (cm)
1.40×10^{-4}			1120	1.00
	0.563		750	1.00
2.56×10^{-4}	0.225		440	
1.55×10^{-3}	0.167			5.00
		33.3	565	1.00
4.35×10^{-3}		21.2	1550	
1.20×10^{-4}		81.3		10.00

3. A solution's transmittance is 35.0%. What is the transmittance if you dilute 25.0 mL of the solution to 50.0 mL?
4. A solution's transmittance is 85.0% when measured in a cell with a pathlength of 1.00 cm. What is the %T if you increase the pathlength to 10.00 cm?

5. The accuracy of a spectrophotometer can be evaluated by preparing a solution of 60.06 ppm $\text{K}_2\text{Cr}_2\text{O}_7$ in 0.0050 M H_2SO_4 , and measuring its absorbance at a wavelength of 350 nm in a cell with a pathlength of 1.00 cm. The absorbance should be 0.640. What is the molar absorptivity of $\text{K}_2\text{Cr}_2\text{O}_7$ at this wavelength?
6. A chemical deviation to Beer's law may occur if the concentration of an absorbing species is affected by the position of an equilibrium reaction. Consider a weak acid, HA, for which K_a is 2×10^{-5} . Construct Beer's law calibration curves of absorbance versus the total concentration of weak acid ($C_{\text{total}} = [\text{HA}] + [\text{A}^-]$), using values for C_{total} of 1×10^{-5} , 3×10^{-5} , 5×10^{-5} , 9×10^{-5} , 11×10^{-5} , and 13×10^{-5} M for the following sets of conditions:
- $\epsilon_{\text{HA}} = \epsilon_{\text{A}^-} = 2000 \text{ M}^{-1} \text{ cm}^{-1}$; unbuffered solution.
 - $\epsilon_{\text{HA}} = 2000 \text{ M}^{-1} \text{ cm}^{-1}$ and $\epsilon_{\text{A}^-} = 500 \text{ M}^{-1} \text{ cm}^{-1}$; unbuffered solution.
 - $\epsilon_{\text{HA}} = 2000 \text{ M}^{-1} \text{ cm}^{-1}$ and $\epsilon_{\text{A}^-} = 500 \text{ M}^{-1} \text{ cm}^{-1}$; solution is buffered to a pH of 4.5.

Assume a constant pathlength of 1.00 cm for all samples.

7. One instrumental limitation to Beer's law is the effect of polychromatic radiation. Consider a line source that emits radiation at two wavelengths, λ' and λ'' . When treated separately, the absorbances at these wavelengths, A' and A'' , are

$$A' = -\log \frac{P_{\text{T}}'}{P_0'} = \epsilon' b C$$

$$A'' = -\log \frac{P_{\text{T}}''}{P_0''} = \epsilon'' b C$$

If both wavelengths are measured simultaneously the absorbance is

$$A = -\log \frac{(P_{\text{T}}' + P_{\text{T}}'')}{(P_0' + P_0'')}$$

- (a) Show that if the molar absorptivity at λ' and λ'' are the same ($\epsilon' = \epsilon'' = \epsilon$), the absorbance is equivalent to

$$A = \epsilon b C$$

- (b) Construct Beer's law calibration curves over the concentration range of zero to 1×10^{-4} M using $\epsilon' = 1000$ and $\epsilon'' = 1000$, and $\epsilon' = 1000$ and $\epsilon'' = 100$. Assume a value of 1.00 cm for the pathlength. Explain the difference between the two curves.

8. A second instrumental limitation to Beer's law is stray radiation. The following data were obtained using a cell with a pathlength of 1.00 cm when stray light is insignificant ($P_{\text{stray}} = 0$).

[analyte] (mM)	absorbance
0.00	0.00
2.00	0.40
4.00	0.80
6.00	1.20
8.00	1.60
10.00	2.00

Calculate the absorbance of each solution when P_{stray} is 5% of P_0 , and plot Beer's law calibration curves for both sets of data. Explain any differences between the two curves. (*Hint*: Assume that P_0 is 100).

9. In the process of performing a spectrophotometric determination of Fe, an analyst prepares a calibration curve using a single-beam spectrophotometer similar to that shown in [Figure 10.26](#). After preparing the calibration curve, the analyst drops and breaks the cuvette. The analyst acquires a new cuvette, measures the absorbance of his sample, and determines the %w/w Fe in the sample. Will the change in cuvette lead to a determinate error in the analysis? Explain.
10. The spectrophotometric methods for determining Mn in steel and for determining glucose use a chemical reaction to produce a colored species whose absorbance we can monitor. In the analysis of Mn in steel, colorless Mn^{2+} is oxidized to give the purple MnO_4^- ion. To analyze for glucose, which is colorless, we react it with a yellow colored solution of the $\text{Fe}(\text{CN})_6^{3-}$, forming the colorless $\text{Fe}(\text{CN})_6^{4-}$ ion. The directions for the analysis of Mn do not specify precise reaction conditions, and samples and standards may be treated separately. The conditions for the analysis of glucose, however, require that the samples and standards be treated simultaneously at exactly the same temperature and for exactly the same length of time. Explain why these two experimental procedures are so different.
11. One method for the analysis of Fe^{3+} , which can be used with a variety of sample matrices, is to form the highly colored Fe^{3+} -thioglycolic acid complex. The complex absorbs strongly at 535 nm. Standardizing the method is accomplished using external standards. A 10.00 ppm Fe^{3+} working standard is prepared by transferring a 10-mL aliquot of a 100.0 ppm stock solution of Fe^{3+} to a 100-mL volumetric flask and diluting to volume. Calibration standards of 1.00, 2.00, 3.00, 4.00, and 5.00 ppm are prepared by transferring appropriate amounts of the 10.0 ppm

working solution into separate 50-mL volumetric flasks, each containing 5 mL of thioglycolic acid, 2 mL of 20% w/v ammonium citrate, and 5 mL of 0.22 M NH_3 . After diluting to volume and mixing, the absorbances of the external standards are measured against an appropriate blank. Samples are prepared for analysis by taking a portion known to contain approximately 0.1 g of Fe^{3+} , dissolving in a minimum amount of HNO_3 , and diluting to volume in a 1-L volumetric flask. A 1.00-mL aliquot of this solution is transferred to a 50-mL volumetric flask, along with 5 mL of thioglycolic acid, 2 mL of 20% w/v ammonium citrate, and 5 mL of 0.22 M NH_3 and diluted to volume. The absorbance of this solution is used to determine the concentration of Fe^{3+} in the sample.

- (a) What is an appropriate blank for this procedure?
 - (b) Ammonium citrate is added to prevent the precipitation of Al^{3+} . What is the effect on the reported concentration of iron in the sample if there is a trace impurity of Fe^{3+} in the ammonium citrate?
 - (c) Why does the procedure specify that the sample contain approximately 0.1 g of Fe^{3+} ?
 - (d) Unbeknownst to the analyst, the 100-mL volumetric flask used to prepare the 10.00 ppm working standard of Fe^{3+} has a volume that is significantly smaller than 100.0 mL. What effect will this have on the reported concentration of iron in the sample?
12. A spectrophotometric method for the analysis of iron has a linear calibration curve for standards of 0.00, 5.00, 10.00, 15.00, and 20.00 mg Fe/L. An iron ore sample that is 40–60% w/w is to be analyzed by this method. An approximately 0.5-g sample is taken, dissolved in a minimum of concentrated HCl, and diluted to 1 L in a volumetric flask using distilled water. A 5.00 mL aliquot is removed with a pipet. To what volume—10, 25, 50, 100, 250, 500, or 1000 mL—should it be diluted to minimize the uncertainty in the analysis? Explain.
13. Lozano-Calero and colleagues describe a method for the quantitative analysis of phosphorous in cola beverages based on the formation of the intensely blue-colored phosphomolybdate complex, $(\text{NH}_4)_3[\text{PO}_4(\text{MoO}_3)_2]$.²¹ The complex is formed by adding $(\text{NH}_4)_6\text{Mo}_7\text{O}_{24}$ to the sample in the presence of a reducing agent, such as ascorbic acid. The concentration of the complex is determined spectrophotometrically at a wavelength of 830 nm, using a normal calibration curve as a method of standardization.

21 Lozano-Calero, D.; Martín-Palomeque, P.; Madueño-Loriguillo, S. *J. Chem. Educ.* **1996**, *73*, 1173–1174.

In a typical analysis, a set of standard solutions containing known amounts of phosphorous was prepared by placing appropriate volumes of a 4.00 ppm solution of P_2O_5 in a 5-mL volumetric flask, adding 2 mL of an ascorbic acid reducing solution, and diluting to volume with distilled water. Cola beverages were prepared for analysis by pouring a sample into a beaker and allowing it to stand for 24 h to expel the dissolved CO_2 . A 2.50-mL sample of the degassed sample was transferred to a 50-mL volumetric flask and diluted to volume. A 250- μ L aliquot of the diluted sample was then transferred to a 5-mL volumetric flask, treated with 2 mL of the ascorbic acid reducing solution, and diluted to volume with distilled water.

- The authors note that this method can be applied only to noncolored cola beverages. Explain why this is true.
- How might you modify this method so that it could be applied to any cola beverage?
- Why is it necessary to remove the dissolved gases?
- Suggest an appropriate blank for this method?
- The author's report a calibration curve of

$$A = -0.02 + 0.72 \times \text{ppm } P_2O_5$$

A sample of Crystal Pepsi, analyzed as described above, yields an absorbance of 0.565. What is the concentration of phosphorous, reported as ppm P, in the original sample of Crystal Pepsi?

14. EDTA forms colored complexes with a variety of metal ions that may serve as the basis for a quantitative spectrophotometric method of analysis. The molar absorptivities of the EDTA complexes of Cu^{2+} , Co^{2+} , and Ni^{2+} at three wavelengths are summarized in the following table (all values of ϵ are in $M^{-1} \text{ cm}^{-1}$).

metal	$\epsilon_{462.9}$	$\epsilon_{732.0}$	$\epsilon_{378.7}$
Co^{2+}	15.8	2.11	3.11
Cu^{2+}	2.32	95.2	7.73
Ni^{2+}	1.79	3.03	13.5

Using this information determine the following:

- The concentration of Cu^{2+} in a solution that has an absorbance of 0.338 at a wavelength of 732.0 nm.
- The concentrations of Cu^{2+} and Co^{2+} in a solution that has an absorbance of 0.453 at a wavelength of 732.0 nm and 0.107 at a wavelength of 462.9 nm.

Crystal Pepsi was a colorless, caffeine-free soda produced by PepsiCo. It was available in the United States from 1992 to 1993.

- (c) The concentrations of Cu^{2+} , Co^{2+} , and Ni^{2+} in a sample that has an absorbance of 0.423 at a wavelength of 732.0 nm, 0.184 at a wavelength of 462.9 nm, and 0.291 at a wavelength of 378.7 nm.

The pathlength, b , is 1.00 cm for all measurements.

15. The concentration of phenol in a water sample is determined by separating the phenol from non-volatile impurities by steam distillation, followed by reacting with 4-aminoantipyrine and $\text{K}_3\text{Fe}(\text{CN})_6$ at pH 7.9 to form a colored antipyrine dye. A phenol standard with a concentration of 4.00 ppm has an absorbance of 0.424 at a wavelength of 460 nm using a 1.00 cm cell. A water sample is steam distilled and a 50.00-mL aliquot of the distillate is placed in a 100-mL volumetric flask and diluted to volume with distilled water. The absorbance of this solution is found to be 0.394. What is the concentration of phenol (in parts per million) in the water sample?
16. Saito describes a quantitative spectrophotometric procedure for iron based on a solid-phase extraction using bathophenanthroline in a poly(vinyl chloride) membrane.²² In the absence of Fe^{2+} the membrane is colorless, but when immersed in a solution of Fe^{2+} and I^- , the membrane develops a red color as a result of the formation of an Fe^{2+} –bathophenanthroline complex. A calibration curve determined using a set of external standards with known concentrations of Fe^{2+} gave a standardization relationship of

$$A = (8.60 \times 10^3 \text{ M}^{-1}) \times [\text{Fe}^{2+}]$$

What is the concentration of iron, in mg Fe/L, for a sample with an absorbance of 0.100?

17. In the DPD colorimetric method for the free chlorine residual, which is reported as mg Cl_2/L , the oxidizing power of free chlorine converts the colorless amine N,N-diethyl-*p*-phenylenediamine to a colored dye that absorbs strongly over the wavelength range of 440–580 nm. Analysis of a set of calibration standards gave the following results.

mg Cl_2/L	absorbance
0.00	0.000
0.50	0.270
1.00	0.543
1.50	0.813
2.00	1.084

22 Saito, T. *Anal. Chim. Acta* **1992**, 268, 351–355.

A sample from a public water supply is analyzed to determine the free chlorine residual, giving an absorbance of 0.113. What is the free chlorine residual for the sample in mg Cl_2/L ?

18. Lin and Brown described a quantitative method for methanol based on its effect on the visible spectrum of methylene blue.²³ In the absence of methanol, the visible spectrum for methylene blue shows two prominent absorption bands centered at approximately 610 nm and 660 nm, corresponding to the monomer and dimer, respectively. In the presence of methanol, the intensity of the dimer's absorption band decreases, while that of the monomer increases. For concentrations of methanol between 0 and 30% v/v, the ratio of the absorbance at 663 nm, A_{663} , to that at 610 nm, A_{610} , is a linear function of the amount of methanol. Using the following standardization data, determine the %v/v methanol in a sample for which A_{610} is 0.75 and A_{663} is 1.07.

%v/v methanol	A_{663}/A_{610}
0.0	1.21
5.0	1.29
10.0	1.42
15.0	1.52
20.0	1.62
25.0	1.74
30.0	1.84

19. The concentration of the barbiturate barbital in a blood sample was determined by extracting 3.00 mL of blood with 15 mL of CHCl_3 . The chloroform, which now contains the barbital, is extracted with 10.0 mL of 0.45 M NaOH (pH \approx 13). A 3.00-mL sample of the aqueous extract is placed in a 1.00-cm cell and an absorbance of 0.115 is measured. The pH of the sample in the absorption cell is then adjusted to approximately 10 by adding 0.5 mL of 16% w/v NH_4Cl , giving an absorbance of 0.023. When 3.00 mL of a standard barbital solution with a concentration of 3 mg/100 mL is taken through the same procedure, the absorbance at pH 13 is 0.295 and the absorbance at a pH of 10 is 0.002. Report the mg barbital/100 mL in the sample.
20. Jones and Thatcher developed a spectrophotometric method for analyzing analgesic tablets containing aspirin, phenacetin, and caffeine.²⁴ The sample is dissolved in CHCl_3 and extracted with an aqueous solution of NaHCO_3 to remove the aspirin. After the extraction is complete, the chloroform is transferred to a 250-mL volumetric flask and diluted to volume with CHCl_3 . A 2.00-mL portion of this solution is diluted to volume in a 200-mL volumetric flask with CHCl_3 . The absorbance of

²³ Lin, J.; Brown, C. W. *Spectroscopy* **1995**, 10(5), 48–51.

²⁴ Jones, M.; Thatcher, R. L. *Anal. Chem.* **1951**, 23, 957–960.

the final solution is measured at wavelengths of 250 nm and 275 nm, at which the absorptivities, in $\text{ppm}^{-1} \text{cm}^{-1}$, for caffeine and phenacetin are

	a_{250}	a_{275}
caffeine	0.0131	0.0485
phenacetin	0.0702	0.0159

Aspirin is determined by neutralizing the NaHCO_3 in the aqueous solution and extracting the aspirin into CHCl_3 . The combined extracts are diluted to 500 mL in a volumetric flask. A 20.00-mL portion of the solution is placed in a 100-mL volumetric flask and diluted to volume with CHCl_3 . The absorbance of this solution is measured at 277 nm, where the absorptivity of aspirin is $0.00682 \text{ ppm}^{-1} \text{cm}^{-1}$. An analgesic tablet treated by this procedure is found to have absorbances of 0.466 at 250 nm, 0.164 at 275 nm, and 0.600 at 277 nm when using a cell with a 1.00 cm pathlength. Report the milligrams of aspirin, caffeine, and phenacetin in the analgesic tablet.

21. The concentration of SO_2 in a sample of air was determined by the *p*-rosaniline method. The SO_2 was collected in a 10.00-mL solution of HgCl_4^{2-} , where it reacts to form $\text{Hg}(\text{SO}_3)_2^{2-}$, by pulling the air through the solution for 75 min at a rate of 1.6 L/min. After adding *p*-rosaniline and formaldehyde, the colored solution was diluted to 25 mL in a volumetric flask. The absorbance was measured at 569 nm in a 1-cm cell, yielding a value of 0.485. A standard sample was prepared by substituting a 1.00-mL sample of a standard solution containing the equivalent of 15.00 ppm SO_2 for the air sample. The absorbance of the standard was found to be 0.181. Report the concentration of SO_2 in the air in $\text{mg SO}_2/\text{L}$. The density of air is 1.18 g/liter.
22. Seaholtz and colleagues described a method for the quantitative analysis of CO in automobile exhaust based on the measurement of infrared radiation at 2170 cm^{-1} .²⁵ A calibration curve was prepared by filling a 10-cm IR gas cell with a known pressure of CO and measuring the absorbance using an FT-IR. The standardization relationship was found to be

$$A = -1.1 \times 10^{-4} + (9.9 \times 10^{-4}) \times P_{\text{CO}}$$

Samples were prepared by using a vacuum manifold to fill the gas cell. After measuring the total pressure, the absorbance of the sample at 2170 cm^{-1} was measured. Results are reported as %CO ($P_{\text{CO}}/P_{\text{total}}$). The analysis of five exhaust samples from a 1973 coupe give the following results.

25 Seaholtz, M. B.; Pence, L. E.; Moe, O. A. Jr. *J. Chem. Educ.* **1988**, *65*, 820–823.

P_{total} (torr)	absorbance
595	0.1146
354	0.0642
332	0.0591
233	0.0412
143	0.0254

Determine the %CO for each sample, and report the mean value and the 95% confidence interval.

23. [Figure 10.32](#) shows an example of a disposable IR sample card made using a thin sheet of polyethylene. To prepare an analyte for analysis, it is dissolved in a suitable solvent and a portion of the sample placed on the IR card. After the solvent evaporates, leaving the analyte behind as a thin film, the sample's IR spectrum is obtained. Because the thickness of the polyethylene film is not uniform, the primary application of IR cards is for a qualitative analysis. Zhao and Malinowski reported how an internal standardization with KSCN can be used for a quantitative IR analysis of polystyrene.²⁶ Polystyrene was monitored at 1494 cm^{-1} and KSCN at 2064 cm^{-1} . Standard solutions were prepared by placing weighed portions of polystyrene in a 10-mL volumetric flask and diluting to volume with a solution of 10 g/L KSCN in methyl isobutyl ketone. A typical set of results is shown here.

g polystyrene	0.1609	0.3290	0.4842	0.6402	0.8006
A_{1494}	0.0452	0.1138	0.1820	0.3275	0.3195
A_{2064}	0.1948	0.2274	0.2525	0.3580	0.2703

When a 0.8006-g sample of a poly(styrene/maleic anhydride) copolymer was analyzed, the following results were obtained.

replicate	A_{1494}	A_{2064}
1	0.2729	0.3582
2	0.2074	0.2820
3	0.2785	0.3642

What is the %w/w polystyrene in the copolymer? Given that the reported %w/w polystyrene is 67%, is there any evidence for a determinate error at $\alpha = 0.05$?

24. The following table lists molar absorptivities for the Arsenazo complexes of copper and barium.²⁷ Suggest appropriate wavelengths for analyzing mixtures of copper and barium using their Arsenazo complexes.

²⁶ Zhao, Z.; Malinowski, E. R. *Spectroscopy* **1996**, 11(7), 44–49.

²⁷ Grossman, O.; Turanov, A. N. *Anal. Chim. Acta* **1992**, 257, 195–202.

wavelength (nm)	ϵ_{Cu} ($\text{M}^{-1} \text{cm}^{-1}$)	ϵ_{Ba} ($\text{M}^{-1} \text{cm}^{-1}$)
595	11900	7100
600	15500	7200
607	18300	7400
611	19300	6900
614	19300	7000
620	17800	7100
626	16300	8400
635	10900	9900
641	7500	10500
645	5300	10000
650	3500	8600
655	2200	6600
658	1900	6500
665	1500	3900
670	1500	2800
680	1800	1500

25. Blanco and colleagues report several applications of multiwavelength linear regression analysis for the simultaneous determination of two-component mixtures.⁹ For each of the following, determine the molar concentration of each analyte in the mixture.

- (a) Titanium and vanadium were determined by forming complexes with H_2O_2 . Results for a mixture of Ti(IV) and V(V) and for standards of 63.1 ppm Ti(IV) and 96.4 ppm V(V) are listed in the following table.

wavelength (nm)	absorbance		
	Ti(V) standard	V(V) standard	mixture
390	0.895	0.326	0.651
430	0.884	0.497	0.743
450	0.694	0.528	0.665
470	0.481	0.512	0.547
510	0.173	0.374	0.314

- (b) Copper and zinc were determined by forming colored complexes with 2-pyridyl-azo-resorcinol (PAR). The absorbances for PAR, a mixture of Cu^{2+} and Zn^{2+} , and standards of 1.00 ppm Cu^{2+} and 1.00 ppm Zn^{2+} are listed in the following table. Note that you must correct the absorbances for the metal for the contribution from PAR.

wavelength (nm)	PAR	absorbance		
		Cu standard	Zn standard	mixture
480	0.211	0.698	0.971	0.656
496	0.137	0.732	1.018	0.668
510	0.100	0.732	0.891	0.627
526	0.072	0.602	0.672	0.498
540	0.056	0.387	0.306	0.290

26. The stoichiometry of a metal–ligand complex, ML_n , was determined by the method of continuous variations. A series of solutions was prepared in which the combined concentrations of M and L were held constant at 5.15×10^{-4} M. The absorbances of these solutions were measured at a wavelength where only the metal–ligand complex absorbs. Using the following data, determine the formula of the metal–ligand complex.

mole fraction M	mole fraction L	absorbance
1.0	0.0	0.001
0.9	0.1	0.126
0.8	0.2	0.260
0.7	0.3	0.389
0.6	0.4	0.515
0.5	0.5	0.642
0.4	0.6	0.775
0.3	0.7	0.771
0.2	0.8	0.513
0.1	0.9	0.253
0.0	1.0	0.000

27. The stoichiometry of a metal–ligand complex, ML_n , was determined by the mole-ratio method. A series of solutions was prepared in which the concentration of metal was held constant at 3.65×10^{-4} M, and the ligand's concentration was varied from 1×10^{-4} M to 1×10^{-3} M. Using the following data, determine the stoichiometry of the metal–ligand complex.

[ligand] (M)	absorbance
1.0×10^{-4}	0.122
2.0×10^{-4}	0.251
3.0×10^{-4}	0.376
4.0×10^{-4}	0.496
5.0×10^{-4}	0.625
6.0×10^{-4}	0.752

7.0×10^{-4}	0.873
8.0×10^{-4}	0.937
9.0×10^{-4}	0.962
1.0×10^{-3}	1.002

28. The stoichiometry of a metal–ligand complex, ML_n , was determined by the slope-ratio method. Two sets of solutions were prepared. For the first set of solutions the concentration of the metal was held constant at 0.010 M and the concentration of the ligand was varied. When the absorbance of these solutions was measured at a wavelength where only the metal–ligand complex absorbs, the following data were obtained.

[ligand] (M)	absorbance
1.0×10^{-5}	0.012
2.0×10^{-5}	0.029
3.0×10^{-5}	0.042
4.0×10^{-5}	0.055
5.0×10^{-5}	0.069

For the second set of solutions the concentration of the ligand was held constant at 0.010 M, and the concentration of the metal was varied, yielding the following absorbances.

[metal] (M)	absorbance
1.0×10^{-5}	0.040
2.0×10^{-5}	0.085
3.0×10^{-5}	0.125
4.0×10^{-5}	0.162
5.0×10^{-5}	0.206

Using this data, determine the stoichiometry of the metal–ligand complex.

29. Kawakami and Igarashi developed a spectrophotometric method for nitrite based on its reaction with 5, 10, 15, 20-tetrakis(4-aminophenyl) porphyrine (TAPP). As part of their study they investigated the stoichiometry of the reaction between TAPP and NO_2^- . The following data are derived from a figure in their paper.²⁸

[TAPP] (M)	$[\text{NO}_2^-]$ (M)	absorbance
8.0×10^{-7}	1.6×10^{-7}	0.227
8.0×10^{-7}	3.2×10^{-7}	0.192

²⁸ Kawakami, T.; Igarashi, S. *Anal. Chim. Acta* **1996**, *33*, 175–180.

8.0×10^{-7}	4.8×10^{-7}	0.158
8.0×10^{-7}	8.0×10^{-7}	0.126
8.0×10^{-7}	1.6×10^{-6}	0.065
8.0×10^{-7}	2.4×10^{-6}	0.047
8.0×10^{-7}	3.2×10^{-6}	0.042
8.0×10^{-7}	4.0×10^{-6}	0.042

What is the stoichiometry of the reaction?

30. The equilibrium constant for an acid–base indicator is determined by preparing three solutions, each of which has a total indicator concentration of 1.35×10^{-5} M. The pH of the first solution is adjusted until it is acidic enough to ensure that only the acid form of the indicator is present, yielding an absorbance of 0.673. The absorbance of the second solution, whose pH is adjusted to give only the base form of the indicator, is 0.118. The pH of the third solution is adjusted to 4.17 and has an absorbance of 0.439. What is the acidity constant for the acid–base indicator?
31. The acidity constant for an organic weak acid was determined by measuring its absorbance as a function of pH while maintaining a constant total concentration of the acid. Using the data in the following table, determine the acidity constant for the organic weak acid.

pH	absorbance
1.53	0.010
2.20	0.010
3.66	0.035
4.11	0.072
4.35	0.103
4.75	0.169
4.88	0.193
5.09	0.227
5.69	0.288
7.20	0.317
7.78	0.317

32. Suppose you need to prepare a set of calibration standards for the spectrophotometric analysis of an analyte that has a molar absorptivity of $1138 \text{ M}^{-1} \text{ cm}^{-1}$ at a wavelength of 625 nm. To maintain an acceptable precision for the analysis, the %T for the standards should be between 15% and 85%.

- (a) What is the concentration of the most concentrated and the least concentrated standard you should prepare, assuming a 1.00-cm sample cell.
- (b) Explain how you will analyze samples with concentrations that are 10 μM , 0.1 mM, and 1.0 mM in the analyte.
33. When using a spectrophotometer whose precision is limited by the uncertainty of reading %T, the analysis of highly absorbing solutions can lead to an unacceptable level of indeterminate errors. Consider the analysis of a sample for which the molar absorptivity is $1.0 \times 10^4 \text{ M}^{-1} \text{ cm}^{-1}$ and the pathlength is 1.00 cm.
- (a) What is the relative uncertainty in concentration for an analyte whose concentration is $2.0 \times 10^{-4} \text{ M}$ if s_T is ± 0.002 ?
- (b) What is the relative uncertainty in the concentration if the spectrophotometer is calibrated using a blank consisting of a $1.0 \times 10^{-4} \text{ M}$ solution of the analyte?
34. Hobbins reported the following calibration data for the flame atomic absorption analysis for phosphorous.²⁹

mg P/L	absorbance
2130	0.048
4260	0.110
6400	0.173
8530	0.230

To determine the purity of a sample of Na_2HPO_4 , a 2.469-g sample is dissolved and diluted to volume in a 100-mL volumetric flask. Analysis of the resulting solution gives an absorbance of 0.135. What is the purity of the Na_2HPO_4 ?

35. Bonert and Pohl reported results for the atomic absorption analysis of several metals in the caustic suspensions produced during the manufacture of soda by the ammonia-soda process.³⁰
- (a) The concentration of Cu was determined by acidifying a 200-mL sample of the caustic solution with 20 mL of concentrated HNO_3 , adding 1 mL of 27% w/v H_2O_2 , and boiling for 30 min. The resulting solution was diluted to 500 mL, filtered, and analyzed by flame atomic absorption using matrix matched standards. The results for a typical analysis are shown in the following table.

29 Hobbins, W. B. "Direct Determination of Phosphorous in Aqueous Matricies by Atomic Absorption," Varian Instruments at Work, Number AA-19, February 1982.

30 Bonert, K.; Pohl, B. "The Determination of Cd, Cr, Cu, Ni, and Pb in Concentrated $\text{CaCl}_2/\text{NaCl}$ solutions by AAS," AA Instruments at Work (Varian) Number 98, November, 1990.

solution	mg Cu/L	absorbance
blank	0.000	0.007
standard 1	0.200	0.014
standard 2	0.500	0.036
standard 3	1.000	0.072
standard 4	2.000	0.146
sample		0.027

Determine the concentration of Cu in the caustic suspension.

- (b) The determination of Cr was accomplished by acidifying a 200-mL sample of the caustic solution with 20 mL of concentrated HNO_3 , adding 0.2 g of Na_2SO_3 and boiling for 30 min. The Cr was isolated from the sample by adding 20 mL of NH_3 , producing a precipitate that includes the chromium as well as other oxides. The precipitate was isolated by filtration, washed, and transferred to a beaker. After acidifying with 10 mL of HNO_3 , the solution was evaporated to dryness. The residue was redissolved in a combination of HNO_3 and HCl and evaporated to dryness. Finally, the residue was dissolved in 5 mL of HCl , filtered, diluted to volume in a 50-mL volumetric flask, and analyzed by atomic absorption using the method of standard additions. The atomic absorption results are summarized in the following table.

sample	mg Cr_{added} /L	absorbance
blank		0.001
sample		0.045
standard addition 1	0.200	0.083
standard addition 2	0.500	0.118
standard addition 3	1.000	0.192

Report the concentration of Cr in the caustic suspension.

36. Quigley and Vernon report results for the determination of trace metals in seawater using a graphite furnace atomic absorption spectrophotometer and the method of standard additions.³¹ The trace metals were first separated from their complex, high-salt matrix by coprecipitating with Fe^{3+} . In a typical analysis a 5.00-mL portion of 2000 ppm Fe^{3+} was added to 1.00 L of seawater. The pH was adjusted to 9 using NH_4OH , and the precipitate of $\text{Fe}(\text{OH})_3$ allowed to stand overnight. After isolating and rinsing the precipitate, the $\text{Fe}(\text{OH})_3$ and coprecipitated metals were dissolved in 2 mL of concentrated HNO_3 and diluted to volume in a 50-mL volumetric flask. To analyze for Mn^{2+} , a 1.00-mL sample of this solution was diluted to 100 mL in a volumetric flask. The following samples were injected into the graphite furnace and analyzed.

³¹ Quigley, M. N.; Vernon, F. J. *Chem. Educ.* **1996**, *73*, 671–673.

sample	absorbance
2.5- μ L sample + 2.5 μ L of 0 ppb Mn^{2+}	0.223
2.5- μ L sample + 2.5 μ L of 2.5 ppb Mn^{2+}	0.294
2.5- μ L sample + 2.5 μ L of 5.0 ppb Mn^{2+}	0.361

Report the parts per billion Mn^{2+} in the sample of seawater.

37. The concentration of Na in plant materials may be determined by flame atomic emission. The material to be analyzed is prepared by grinding, homogenizing, and drying at 103 °C. A sample of approximately 4 g is transferred to a quartz crucible and heated on a hot plate to char the organic material. The sample is heated in a muffle furnace at 550 °C for several hours. After cooling to room temperature the residue is dissolved by adding 2 mL of 1:1 HNO_3 and evaporated to dryness. The residue is redissolved in 10 mL of 1:9 HNO_3 , filtered and diluted to 50 mL in a volumetric flask. The following data were obtained during a typical analysis for the concentration of Na in a 4.0264-g sample of oat bran.

sample	mg Na/L	emission (arbitrary units)
blank	0.00	0.0
standard 1	2.00	90.3
standard 2	4.00	181
standard 3	6.00	272
standard 4	8.00	363
standard 5	10.00	448
sample		238

Determine the mg Na/L in the sample of oat bran.

38. Gluodenis describes the use of ICP atomic emission to analyze samples of brass for Pb and Ni.³² The analysis for Pb uses external standards prepared from brass samples containing known amounts of lead. Results are shown in the following table.

%w/w Pb	emission intensity
0.000	4.29×10^4
0.0100	1.87×10^5
0.0200	3.20×10^5
0.0650	1.28×10^6
0.350	6.22×10^6
0.700	1.26×10^7

³² Gluodenis, T. J. Jr. *Am. Lab.* November 1998, 245–275.

1.04	1.77×10^7
2.24	3.88×10^7
3.15	5.61×10^7
9.25	1.64×10^8

What is the %w/w Pb in a sample of brass that gives an emission intensity of 9.25×10^4 ?

The analysis for Ni uses an internal standard. Results for a typical calibration are shown in the following table.

% w/w Ni	emission intensity ratio
0.000	0.00267
0.0140	0.00154
0.0330	0.00312
0.130	0.120
0.280	0.246
0.280	0.247
0.560	0.533
1.30	1.20
4.82	4.44

What is the %w/w Ni in a sample for which the ratio of emission intensity is 1.10×10^{-3} ?

39. Yan and colleagues developed a method for the analysis of iron based on the formation of a fluorescent metal–ligand complex with the ligand 5-(4-methylphenylazo)-8-aminoquinoline.³³ In the presence of the surfactant cetyltrimethyl ammonium bromide the analysis is carried out using an excitation wavelength of 316 nm with emission monitored at 528 nm. Standardization with external standards gives the following calibration curve.

$$I_f = -0.03 + 1.594 \times \frac{\text{mg Fe}^{3+}}{\text{L}}$$

A 0.5113-g sample of dry dog food was ashed to remove organic materials, and the residue dissolved in a small amount of HCl and diluted to volume in a 50-mL volumetric flask. Analysis of the resulting solution gave a fluorescent emission intensity of 5.72. Determine the mg Fe/L in the sample of dog food.

40. A solution of 5.00×10^{-5} M 1,3-dihydroxynaphthelene in 2 M NaOH has a fluorescence intensity of 4.85 at a wavelength of 459 nm. What

³³ Yan, G.; Shi, G.; Liu, Y. *Anal. Chim. Acta* **1992**, 264, 121–124.

is the concentration of 1,3-dihydroxynaphthelene in a solution with a fluorescence intensity of 3.74 under identical conditions?

41. The following data was recorded for the phosphorescence intensity for several standard solutions of benzo[a]pyrene.

[benzo[a]pyrene] (M)	emission intensity
0	0.00
1.00×10^{-5}	0.98
3.00×10^{-5}	3.22
6.00×10^{-5}	6.25
1.00×10^{-4}	10.21

What is the concentration of benzo[a]pyrene in a sample yielding a phosphorescent emission intensity of 4.97?

42. The concentration of acetylsalicylic acid, $C_9H_8O_4$, in aspirin tablets can be determined by hydrolyzing to the salicylate ion, $C_7H_5O_2^-$, and determining the concentration of the salicylate ion spectrofluorometrically. A stock standard solution is prepared by weighing 0.0774 g of salicylic acid, $C_7H_6O_2$, into a 1-L volumetric flask and diluting to volume with distilled water. A set of calibration standards is prepared by pipeting 0, 2.00, 4.00, 6.00, 8.00, and 10.00 mL of the stock solution into separate 100-mL volumetric flasks containing 2.00 mL of 4 M NaOH and diluting to volume with distilled water. The fluorescence of the calibration standards was measured at an emission wavelength of 400 nm using an excitation wavelength of 310 nm; results are listed in the following table.

mL of stock solution	emission intensity
0.00	0.00
2.00	3.02
4.00	5.98
6.00	9.18
8.00	12.13
10.00	14.96

Several aspirin tablets are ground to a fine powder in a mortar and pestle. A 0.1013-g portion of the powder is placed in a 1-L volumetric flask and diluted to volume with distilled water. A portion of this solution is filtered to remove insoluble binders and a 10.00-mL aliquot transferred to a 100-mL volumetric flask containing 2.00 mL of 4 M NaOH. After diluting to volume the fluorescence of the resulting solution is found to be 8.69. What is the %w/w acetylsalicylic acid in the aspirin tablets?

43. Selenium (IV) in natural waters can be determined by complexing with ammonium pyrrolidine dithiocarbamate and extracting into CHCl_3 . This step serves to concentrate the Se(IV) and to separate it from Se(VI). The Se(IV) is then extracted back into an aqueous matrix using HNO_3 . After complexing with 2,3-diaminonaphthalene, the complex is extracted into cyclohexane. Fluorescence is measured at 520 nm following its excitation at 380 nm. Calibration is achieved by adding known amounts of Se(IV) to the water sample before beginning the analysis. Given the following results what is the concentration of Se(IV) in the sample.

[Se (IV)] added (nM)	emission intensity
0.00	323
2.00	597
4.00	862
6.00	1123

44. Fibrinogen is a protein that is produced by the liver and found in human plasma. Its concentration in plasma is clinically important. Many of the analytical methods used to determine the concentration of fibrinogen in plasma are based on light scattering following its precipitation. For example, da Silva and colleagues describe a method in which fibrinogen precipitates in the presence of ammonium sulfate in a guanidine hydrochloride buffer.³⁴ Light scattering is measured nephelometrically at a wavelength of 340 nm. Analysis of a set of external calibration standards gives the following calibration equation

$$I_s = -4.66 + 9907.63 \times C$$

where I_s is the intensity of scattered light and C is the concentration of fibrinogen in g/L. A 9.00-mL sample of plasma was collected from a patient and mixed with 1.00 mL of an anticoagulating agent. A 1.00-mL aliquot of this solution was then diluted to 250 mL in a volumetric flask. Analysis of the resulting solution gave a scattering intensity of 44.70. What is the concentration of fibrinogen, in gram per liter, in the plasma sample?

10L Solutions to Practice Exercises

Practice Exercise 10.1

The frequency and wavenumber for the line are

$$\nu = \frac{c}{\lambda} = \frac{3.00 \times 10^8 \text{ m/s}}{656.3 \times 10^{-9} \text{ m}} = 4.57 \times 10^{14} \text{ s}^{-1}$$

³⁴ da Silva, M. P.; Fernandez-Romero, J. M.; Luque de Castro, M. D. *Anal. Chim. Acta* **1996**, 327, 101–106.

$$\bar{\nu} = \frac{1}{\lambda} = \frac{1}{656.3 \times 10^{-9} \text{ m}} \times \frac{1 \text{ m}}{100 \text{ cm}} = 1.524 \times 10^4 \text{ cm}^{-1}$$

Click [here](#) to return to the chapter.

Practice Exercise 10.2

The photon's energy is

$$E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{656.3 \times 10^{-9} \text{ m}} = 3.03 \times 10^{-19} \text{ J}$$

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Practice Exercise 10.3

To find the transmittance, T , we begin by noting that

$$A = 1.27 = -\log T$$

Solving for T

$$-1.27 = \log T$$

$$10^{-1.27} = T$$

gives a transmittance of 0.054, or a %T of 5.4%.

Click [here](#) to return to the chapter.

Practice Exercise 10.4

Making appropriate substitutions into Beer's law

$$A = 0.228 = \epsilon b C = (676 \text{ M}^{-1} \text{ cm}^{-1})(1 \text{ cm}) C$$

and solving for C gives a concentration of $3.37 \times 10^{-4} \text{ M}$.

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Practice Exercise 10.5

For this standard addition we can write the following equations relating absorbance to the concentration of Cu^{2+} in the sample. First, for the sample, we have

$$0.118 = \epsilon b C_{\text{Cu}}$$

and for the standard addition we have

$$0.162 = \epsilon b \left(C_{\text{Cu}} + \frac{20.00 \text{ mg Cu}}{\text{L}} \times \frac{1.00 \text{ mL}}{10.00 \text{ mL}} \right)$$

The value of ϵb is the same in both equation. Solving each equation for ϵb and equating

$$\frac{0.162}{C_{\text{Cu}} + \frac{20.00 \text{ mg Cu}}{\text{L}} \times \frac{1.00 \text{ mL}}{10.00 \text{ mL}}} = \frac{0.118}{C_{\text{Cu}}}$$

leaves us with an equation in which C_{Cu} is the only variable. Solving for C_{Cu} gives its value as

$$\frac{0.162}{C_{\text{Cu}} + 2.00 \text{ mg Cu/L}} = \frac{0.118}{C_{\text{Cu}}}$$

$$0.162C_{\text{Cu}} = 0.118C_{\text{Cu}} + 0.236 \text{ mg Cu/L}$$

$$0.044C_{\text{Cu}} = 0.236 \text{ mg Cu/L}$$

$$C_{\text{Cu}} = 5.4 \text{ mg Cu/L}$$

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Practice Exercise 10.6

Substituting into [equation 10.11](#) and [equation 10.12](#) gives

$$A_{400} = 0.336 = 15.2C_{\text{Cr}} + 5.60C_{\text{Co}}$$

$$A_{505} = 0.187 = 0.533C_{\text{Cr}} + 5.07C_{\text{Co}}$$

To determine C_{Cr} and C_{Co} we solve the first equation for C_{Co}

$$C_{\text{Co}} = \frac{0.336 - 15.2C_{\text{Cr}}}{5.60}$$

and substitute the result into the second equation.

$$0.187 = 0.533C_{\text{Cr}} + 5.07 \times \frac{0.336 - 15.2C_{\text{Cr}}}{5.60} = 0.3042 - 13.23C_{\text{Cr}}$$

Solving for C_{Cr} gives the concentration of Cr^{3+} as 8.86×10^{-3} M. Substituting this concentration back into the equation for the mixture's absorbance at 400 nm gives the concentration of Co^{2+} as 3.60×10^{-2} M.

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Practice Exercise 10.7

Letting X represent MnO_4^- and Y represent $\text{Cr}_2\text{O}_7^{2-}$, we plot the equation

$$\frac{A_{\text{mix}}}{A_{\text{SX}}} = \frac{C_{\text{X}}}{C_{\text{SX}}} + \frac{C_{\text{Y}}}{C_{\text{SY}}} \times \frac{A_{\text{SY}}}{A_{\text{SX}}}$$

placing $A_{\text{mix}}/A_{\text{SX}}$ on the y -axis and $A_{\text{SY}}/A_{\text{SX}}$ on the x -axis. For example, at a wavelength of 266 nm the value $A_{\text{mix}}/A_{\text{SX}}$ of is 0.766/0.042, or 18.2, and the value of $A_{\text{SY}}/A_{\text{SX}}$ is 0.410/0.042, or 9.76. Completing the calculations for all wavelengths and plotting the data gives the result shown in Figure 10.67. Fitting a straight-line to the data gives a regression model of

$$\frac{A_{\text{mix}}}{A_{\text{SX}}} = 0.8147 + 1.7839 \times \frac{A_{\text{SY}}}{A_{\text{SX}}}$$

Using the y -intercept, the concentration of MnO_4^- is

$$\frac{C_{\text{X}}}{C_{\text{SX}}} = \frac{[\text{MnO}_4^-]}{1.0 \times 10^{-4} \text{ M MnO}_4^-} = 0.8147$$

or $8.15 \times 10^{-5} \text{ M MnO}_4^-$, and using the slope, the concentration of $\text{Cr}_2\text{O}_7^{2-}$ is

$$\frac{C_{\text{Y}}}{C_{\text{SY}}} = \frac{[\text{Cr}_2\text{O}_7^{2-}]}{1.0 \times 10^{-4} \text{ M Cr}_2\text{O}_7^{2-}} = 1.7839$$

or $1.78 \times 10^{-4} \text{ M Cr}_2\text{O}_7^{2-}$.

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Practice Exercise 10.8

Figure 10.68 shows a continuous variations plot for the data in this exercise. Although the individual data points show substantial curvature—enough curvature that there is little point in trying to draw linear branches for excess metal and excess ligand—the maximum absorbance clearly occurs at an X_{L} of approximately 0.5. The complex's stoichiometry, therefore, is $\text{Fe}(\text{SCN})^{2+}$.

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Practice Exercise 10.9

The value of K_{a} is

$$K_{\text{a}} = (1.00 \times 10^{-6}) \times \frac{0.225 - 0.000}{0.680 - 0.225} = 4.95 \times 10^{-7}$$

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Practice Exercise 10.10

To determine K_{a} we use [equation 10.21](#), plotting $\log[(A - A_{\text{HIn}})/(A_{\text{In}} - A)]$ versus pH, as shown in Figure 10.69. Fitting a straight-line to the data gives a regression model of

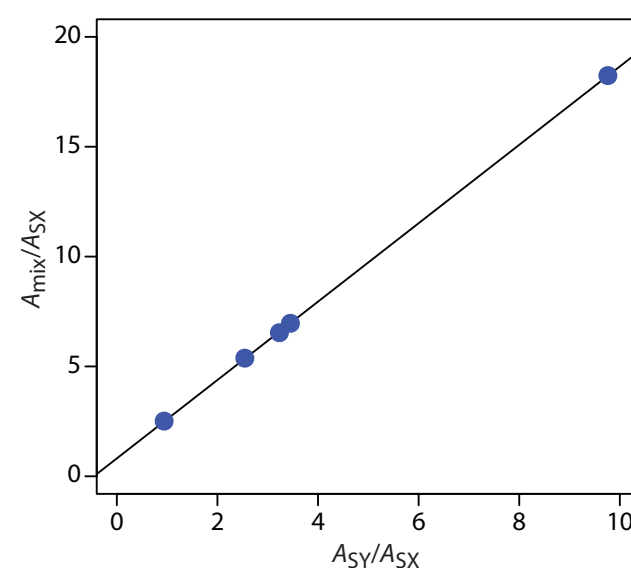


Figure 10.67 Multiwavelength linear regression analysis for the data in Practice Exercise 10.7.

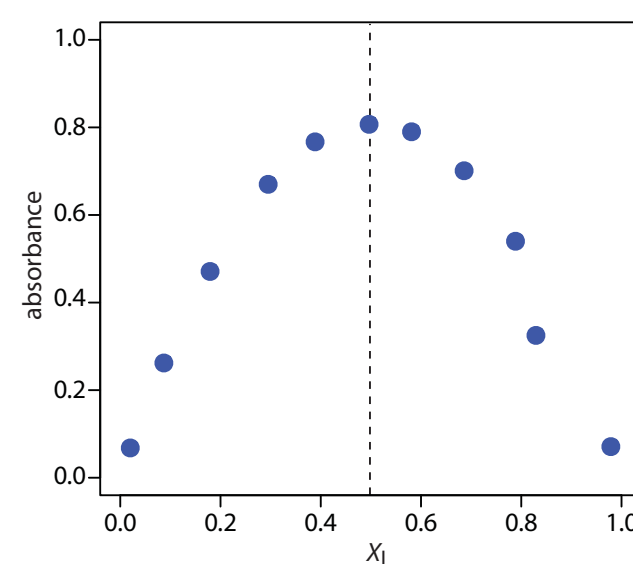


Figure 10.68 Continuous variations plot for the data in Practice Exercise 10.8.

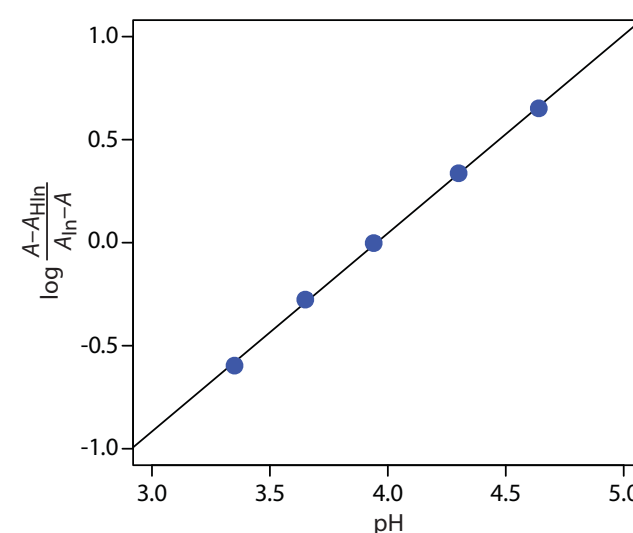


Figure 10.69 Determining the pKa of bromothymol blue using the data in Practice Exercise 10.10.

$$\log \frac{A - A_{\text{HIn}}}{A_{\text{In}} - A} = -3.80 + 0.962\text{pH}$$

The y -intercept is $-\text{p}K_{\text{a}}$; thus, the $\text{p}K_{\text{a}}$ is 3.80 and the K_{a} is 1.58×10^{-4} .

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