

Flame Tests

Atomic Emission and Electron Energy Levels

Introduction

Just as a fingerprint is unique to each person, the color of light emitted by an element heated in a flame is also unique to each element. In this experiment, the characteristic color of light emitted by calcium, copper, lithium, potassium, sodium and strontium ions will be observed.

Concepts

- Atomic emission
- Wavelength and energy of light
- Excited vs. ground states
- Flame tests

Background

When a substance is heated in a flame, the atoms absorb energy from the flame. This absorbed energy allows the electrons to be promoted to excited energy levels. From these excited energy levels, there is a natural tendency for the electrons to make a transition or drop back down to the ground state. When an electron makes a transition from a higher energy level to a lower energy level, a particle of light called a photon is emitted (see Figure 1). Both the absorption and emission of energy are quantized—only certain energy levels are allowed.

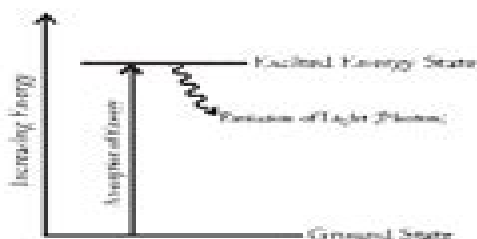


Figure 1. Absorption and emission of energy.

An electron may drop all the way back down to the ground state in a single step, emitting a photon in the process. Alternatively, an electron may drop back down to the ground state in a series of smaller steps, emitting a photon with each step. In either case, the energy of each emitted photon is equal to the difference in energy between the excited state and the state to which the electron relaxes. The energy of the emitted photon determines the color of light observed in the flame. The flame color may be described in terms of its wavelength, and Equation 1 may be used to calculate the energy of the emitted photon.

$$\Delta E = \frac{hc}{\lambda}$$

Equation 1

ΔE is the difference in energy between the two energy levels in joules (J).

h is Planck's constant ($h = 6.626 \times 10^{-34}$ Jsec).

c is the speed of light ($c = 2.998 \times 10^8$ m/sec) and

λ (lambda) is the wavelength of light in meters. The wavelengths of visible light are given in units of nanometers ($1 \text{ nm} = 1 \times 10^{-9}$ m). See Table 1 on the following page.

The color of light observed when a substance is heated in a flame varies from one substance to another. Because each element has a different spacing of electron energy levels, the possible electron transitions for a given substance are unique. Therefore, the difference in energy between energy levels, the exact energy of the emitted photon and the corresponding wavelength and color are unique to each substance. As a result, the colors observed when a substance is heated in a flame may be used as a means of identification.

The Visible Portion of the Electromagnetic Spectrum

Visible light is a form of electromagnetic radiation. Other familiar forms of electromagnetic radiation include gamma rays, X-rays, ultraviolet (UV) radiation, infrared (IR) radiation, microwave radiation and radio waves. Together, all forms of electromagnetic radiation make up the electromagnetic spectrum. The visible portion of the electromagnetic spectrum is the only portion that can be detected by the human eye—all other forms of electromagnetic radiation are invisible.

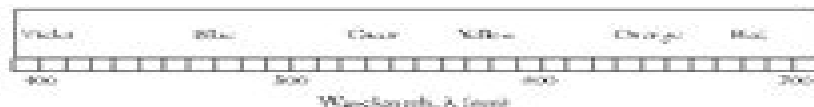


Figure 2. Visible spectrum.

The visible spectrum spans the wavelength region from about 400 to 700 nm (see Figure 2). Light of 400 nm is seen as violet and light of 700 nm is seen as red. According to Equation 1, wavelength is inversely proportional to energy. Therefore, violet light is higher energy light than red light. As the color of light changes, so does the amount of energy it possesses.

Table 1 (page 2) lists the wavelengths associated with each of the colors in the visible spectrum. The representative wavelengths may be used as a benchmark for each color. For example, instead of referring to green as light in the wavelength range 500–560 nm, we may approximate the wavelength of a green light as 520 nm. An infinite number of shades of each color may be observed.

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