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 charactistic color of light emitted by calcium, copper, lithium, potassium, sodium and strontium ions will be observed.

Concepts

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

Background

When a substance is heated in a flame, the atoms absorbed 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.

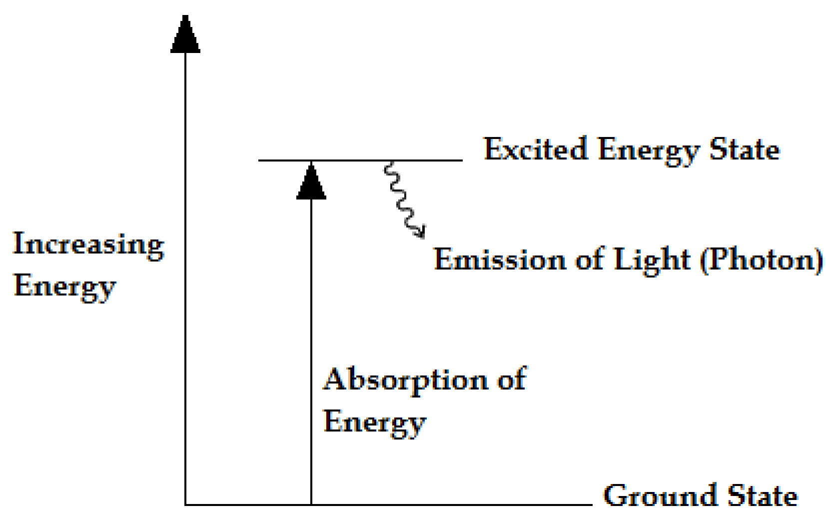


Figure 1. Absorption and emission of energy.

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 emited 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.

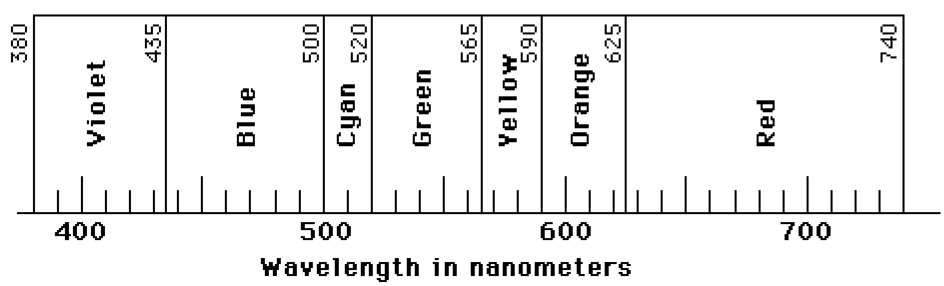
$$\triangle E = \frac{hc}{\lambda} \qquad \text{Equation 1}$$

$\triangle E$ is the difference in energy between the two energy levels in joules (J).
h is Planck's constant (h = 6.626 x 10^{-34} J·s).
c is the speed of light (c = 2.998 x 10^{8} m/sec).
λ (lambda) is the wavelength of light in meters. The wavelength of visible light are given in units of nanometers (1 m = 1 x 10^{9} nm). 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 γ-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 visible.



The visible spectrum spans the wavelength region from about 400 to 700 nm (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 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.

Table 1.

Representative Wavelength, nm	**Wavelength Region, nm**	**Color**
410	400-425	Violet
470	425-480	Blue
490	480-500	Blue-green
520	500-560	Green
565	560-580	Yellow-green
580	580-585	Yellow
600	585-650	Orange
650	650-700	Red

Experiment Overview

The purpose of this experiment is to observe the characteristic flame test colors of different metal compounds and to use this information to identify metal ions by color.

Pre-Lab Questions

1. Fill in the blanks.

   When an atom absorbs energy, the electrons move from their ___________________ state to an ___________________ state. When an atom emits energy, the electrons move from a(n) ___________________ state to their ___________________ state and give off ___________________.

2. Is a flame test a qualitative or quantitative test for the identity of an unknown? Explain.____
   ____
   ____

Materials

Calcium chloride, $CaCl_2$
Copper (II) chloride, $CuCl_2$
Lithium chloride, LiCl
Potassium chloride, KCl
Sodium chloride, NaCl

Beakers, 250 mL - 2
Meker burner
Scoop/spatula
Nichrome wire
Weighing boats

Strontium chloride, $SrCl_2$

Safety Precautions

Copper (II) chloride is highly toxic by ingestion; avoid contact with eyes, skin, and mucous membranes. Lithium chloride is moderately toxic by ingestion and is a body tissue irritant. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Wash hands thoroughly with soap and water before leaving the laboratory.

Procedure

1. Fill a 250 mL beaker about half-full with distilled water.
2. Light the laboratory burner.
3. Dip the nichrome wire in the water.
4. Dip the nichrome wire in one sample of the metal chloride. Place it in the flame. Observe the color of the flame. If necessary, repeat the test with additional salt.
5. Rinse the nichrome thoroughly.
6. Record your observations for the flame color produced by the metal chloride in a data table.
7. Repeat steps 3-6 for each of the metal chlorides. Record your observations for the flame color produced by each metal ion in a data table.

Results and Calculations

1. Determine the wavelength, λ, in nm for each of the metals.
2. Convert the wavelength to meters.
3. Calculate the $\triangle E$ in Joules for each metal using Equation 1.
4. What evidence is there from your results that the characteristic color observed for each compound is due to the metal ion in each case?
5. A glass rod was heated in a burner flame and gave off a bright yellow flame. What metal ion predominates in the glass rod?
6. The alkali metals cesium (Cs) and rubidium (Rb) were discovered based on their characteristic flame colors. Cesium is named after the sky and rubidium after the gem color. What colors of light do you think these metals give off when heated in a flame?