

# Flame Tests & Electron Configuration

## INTRODUCTION

Many elements produce colors in the flame when heated. The origin of this phenomenon lies in the arrangement, or “configuration” of the electrons in the atoms of the different elements. In the “solar system” model of the atom first proposed by Ernest Rutherford and Niels Bohr in the early 1900s, the electrons were pictured as moving around the nucleus in circular orbits in a similar manner that the planets in our solar system orbit the sun. As envisioned by Bohr, during heating, one or more electrons may absorb energy in sufficient amounts to “jump” to an orbit farther away from the nucleus. Since the electron has a higher potential energy in its new orbit, the electron is said to be in a higher *energy level*. When the electron has been promoted to a higher energy level, the atom is said to be in an *excited state*.

When the electrons drop from a higher energy level to a lower energy level (in an orbit closer to the nucleus), energy is released. In the flame test, if this energy has the form of visible light, the flame will produce a color characteristic of the element. Different elements have a unique color in its flame which can be used to identify an element.

When the electrons are in their lowest energy orbits (closest to the nucleus), the atom is said to be in its *ground state*.

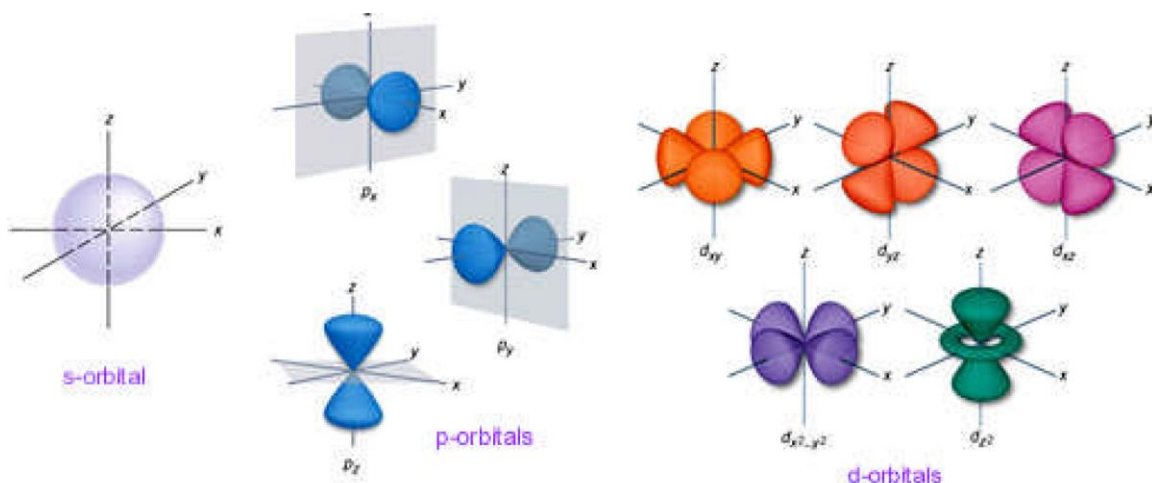
After the Bohr model of the atom, with its electron orbits, came the “wave-mechanical,” or “quantum mechanical,” model in 1926. This model has many similarities to Bohr’s solar system model, such as the concept of energy levels and the ability of electrons to jump to higher energy levels when they absorb energy. However, the quantum-mechanical model takes into account two important aspects of the behavior of electrons: 1) the *wave behavior* of the electron, and 2) the inherent *uncertainty* in knowing the locations of the electrons in the atom.

In our present quantum-mechanical model of the atom, the electrons are still described as being in different energy levels, but now each energy level is pictured as containing *sublevels* within. Finally, in regions within these sublevels are regions called “*orbitals*” (retaining the original though incorrect orbit concept).

The main energy levels are designated by the letter “n,” and they are simply numbered  $n=1$ ,  $n=2$ ,  $n=3$ , etc., as we go further away from the nucleus. The sublevels are designated with the letters “s,” “p,” “d,” “f,” etc. (from the early spectroscopists’ terms “sharp,” “principle,” “diffuse,” and “fundamental”).

The names of the orbital specify the main energy level and sublevel that the electrons occupy. For example, an electron in a “1s” orbital is in the s sublevel of energy level 1. An electron in a “2p” orbital is in the p sublevel of energy level 2. In the p sublevel, there are actually three p-orbitals which, in energy level 2, are more specifically designated as  $2p_x$ ,  $2p_y$ , and  $2p_z$ . *Whatever its type or energy level, an orbital can hold a maximum of only two electrons.*

The regions of space described by the orbitals have interesting shapes:

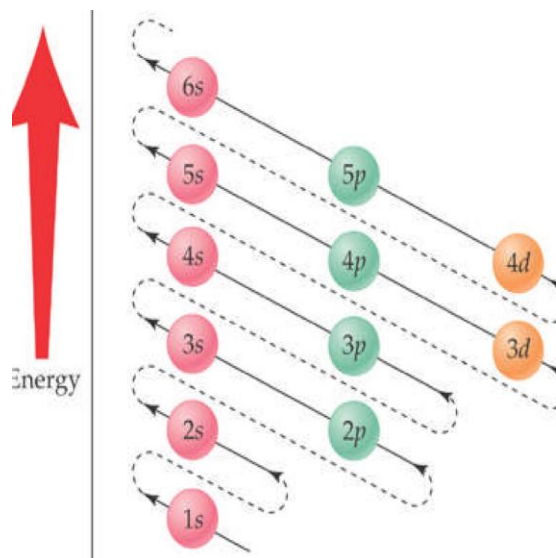


The following tables show the relationship between the main energy levels, sublevels, orbitals, and maximum number of electrons:

Energy Level	Sub Levels
1	s
2	s,p
3	s,p,d
4	s,p,d,f

Sublevel	Number of Orbitals	Maximum Number of Electrons
s	1	2
p	3	6
d	5	10
f	7	14

The order of filling the sublevels with electrons is given by the following illustration. Start with the lowest energy sublevel, which is the 1s sublevel, and then add the electrons, going to the next sublevel in the diagram when the previous level is filled. For example, an atom of carbon has a total of 6 electrons. Starting with the 1s level, the electron configuration will be  $1s^2 2s^2 2p^2$ . The number of electrons in each sublevel is given as superscripts.



Iron, with 26 electrons, will be seen to have the configuration  $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$ .

This scheme gives the correct configuration of most elements.

The chemistry of an element strongly depends on the arrangement of its electrons. Since the energies of the electrons in the atoms of different elements are different, the emission spectrum of each element is unique. The emission spectrum may be used to detect the presence of an element in both qualitative and quantitative ways. A number of common metallic elements emit light strongly in the visible region, allowing their detection to be made with a spectroscope. For these elements, the emissions are so intense that the elements may often be recognized simply by the color that they impart to a flame.

<u>Metal Ion</u>	<u>Color of Flame</u>
Lithium, Li <sup>+</sup>	pink-red
Sodium, Na <sup>+</sup>	yellow-orange
Potassium, K <sup>+</sup>	violet
Calcium, Ca <sup>2+</sup>	orange-red
Strontium, Sr <sup>2+</sup>	deep red
Barium, Ba <sup>2+</sup>	green
Copper(II), Cu <sup>2+</sup>	blue-green

If we examine the emission spectra of these ions with a spectroscope, we find that as with mercury and hydrogen, the emission spectra are composed of a series of lines, which are unique for each element. Consequently, a flame into which both lithium and strontium, for example are placed would be red and we could not tell with our eye that both these ions were present. However, with the aid of a sensitive spectroscope, we could detect the presence of both ions.

To do a flame test on a metallic element, the metal is first dissolved in a solution and the solution is then held in the hot, blue flame of a Bunsen burner. This test works well for metal ions, and was perfected by Robert Bunsen (1811 – 1899). Many metallic ions exhibit characteristic colors when vaporized in the burner flame. In this experiment, you will use a Bunsen burner as an excitation source and observe the color imparted to the flame by these ions. With this information you will determine the contents of unknown solutions. The compounds in the unknowns will contain one or two of the metal ions in the table above.

The value of the flame test is limited by interference from other brighter colors and by ambiguities where certain different metals cause the same flame color. Sodium, in particular, is present on your hands, and is a contaminant in many compounds and will color the flame. Sometimes a colored glass is used to filter out light from one metal. Cobalt glass is often used to filter out the yellow of sodium.

### **Prelab Questions:**

1. What particles are found in the chemicals that may be responsible for the production of colored light?
2. Why do different chemicals emit different colors of light, and why is the color emitted specific to that chemical?
3. Why do chemicals have to be heated in the flame first before the colored light is emitted?
4. Colorful light emissions are applicable to everyday life. Where else have you observed colorful light emissions? Are these light emission applications related? Explain.
5. What is the characteristic flame color for Sodium, Potassium, Lithium, Barium, Strontium, and Copper? Explain why.
6. Write electron configuration for the alkali metals Li, Na, K, and Rb.
7. Write the electron configuration for strontium and the strontium(II) ion.
8. Name the colors in visible light, beginning with that of highest energy (shortest wavelength).
9. What is the maximum number of electrons the first four energy levels can hold?  $n=1$  \_\_\_\_\_  $n=2$  \_\_\_\_\_  $n=3$  \_\_\_\_\_  $n=4$  \_\_\_\_\_
10. What is monochromatic light?

### **Materials:**

- Bunsen Burner
- Flame test wire (nichrome) and holder
- 6 M HCl
- ground powders of each of the following:
  - Sodium Chloride
  - Potassium Chloride
  - Lithium Chloride
  - Barium Chloride
  - Strontium Chloride
  - Cupric Chloride
- unknown

**Procedure:**

1. If necessary clean the nichrome wire by alternately heating it and dipping it into the concentrated acid solution. Repeat several times until the flame remains almost colorless when the wire is heated.
2. Dip the wire into the first test sample, and then hold it into the hot burner flame. Record the color of the flame.
3. Repeat step 2, looking through cobalt glass to see the flame. Record any changes.
4. Move to station 2 for the next compound. Repeat steps 2 and 3 for each of the known materials.
5. Perform a flame test with each of the unknown materials.

**Some questions that you should address in your lab report include:**

How did the colors that you observed compare to the reported colors?

When the cobalt glass was used, what differences did you observe, if any, in the flame color?

What do you think that you would see if you observed the flame through a diffraction grating or spectroscope?

Why were some of the solids/solutions colored, while others were colorless?

In your analysis section, please answer the following questions:

Which ion emitted the higher energy photons in the visible region:  $\text{Cu}^{2+}$  or  $\text{Sr}^{2+}$ ?

Which ion emitted photons with the longer wavelength in the visible region:  $\text{Li}^+$  or  $\text{Na}^+$ ?

Which ion emitted the lower frequency photons in the visible region:  $\text{Ba}^{2+}$  or  $\text{K}^+$ ?

The brilliant colors in fireworks are from the emission of light from ions. The red in fireworks is often due to the emission of light from  $\text{Sr}^{2+}$ . If the primary wavelength is 650 nm, what are the frequency and energy of the light?

Based on the data collected, and what you have learned so far, would the anion have a dramatic effect on the color of the light emitted? Explain the observations you made during lab that support your answer.