

# EXPERIMENT 3 – 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*.

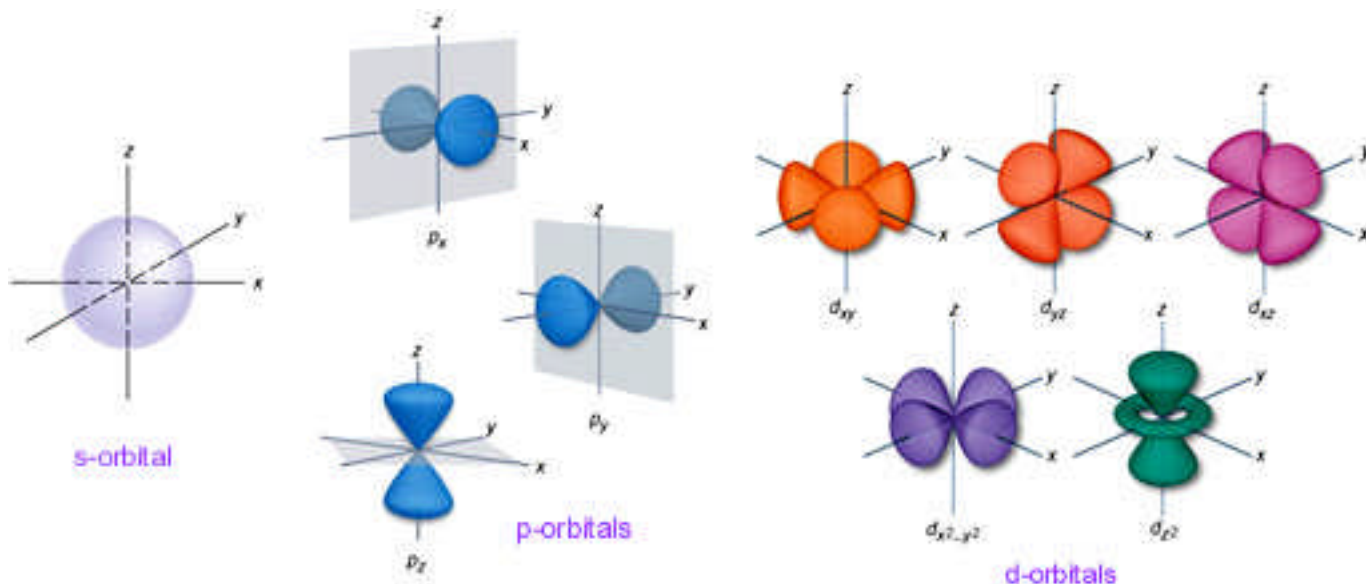
After the Bohr model of the atom, with its electron orbits, came the “wave-mechanical,” or “quantum-mechanical,” model in 1926. This model has many similarities to Bohr’s solar system model, such as the concept of energy levels and the ability of electrons to jump to higher energy levels when they absorb 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:

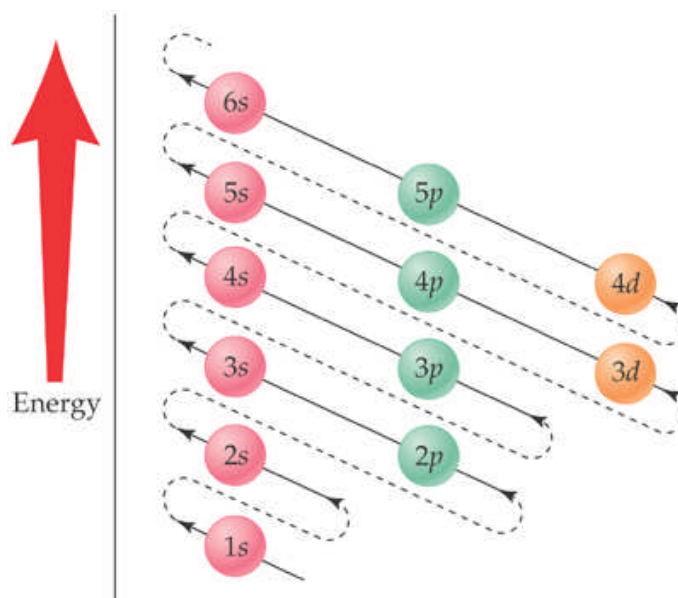
Energy Level	Sublevels
1	s
2	s, p
3	s, p, d
4	s, p, d, f

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.

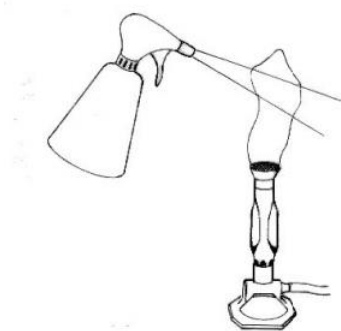
<b>Metal Ion</b>	<b>Color of Flame</b>
Lithium, $\text{Li}^+$	pink-red
Sodium, $\text{Na}^+$	yellow-orange
Potassium, $\text{K}^+$	violet
Calcium, $\text{Ca}^{2+}$	orange-red
Strontium, $\text{Sr}^{2+}$	deep red
Barium, $\text{Ba}^{2+}$	green
Copper(II), $\text{Cu}^{2+}$	blue-green

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.

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.



There are two common ways to introduce the metal ions into the flame. First, you can use a wire loop to pick up a drop of solution, and then place the drop into the flame for vaporization. Although this method is simple and inexpensive, it produces only a brief burst of color before the sample evaporates completely. Second, you can introduce a fine mist of sample into the flame by using a spray bottle. This method produces a longer-lived emission that is therefore easier to see. Your instructor will tell you which method to use.



Flame Test by Mist Method

## **Experimental Procedure**

If you use the wire-loop method, obtain a Nichrome wire and about 10 mL of 6 M HCl solution in a 100-ml beaker. Dip the tip of the Nichrome wire into the 6 M HCl solution before every flame test to remove any oxide residue that may be present. Rinse the loop with deionized water using a wash bottle, and then heat the loop in the hottest part of the flame until no color is imparted to the flame by the loop. If the sprayer is used, check to ensure that the sprayer produces a fine mist. Ignite the Bunsen burner and adjust the flame so that it is as hot as possible (blue flame with light-blue inner cone).

### **A. Known Metal Ions**

With either method of introduction (loop or mist), introduce the metal ion solution into the flame and note the color of the flame. The wire loop can be used with solid particles of the compounds by wetting the loop and touching it to the solid (set out for this purpose; do not use the original containers to avoid contaminating the compounds). The solid particles give a more intense and longer-lasting color.

Record your observations on your report form. Repeat the procedure with each metal ion solution. If you use the wire loop method, clean the wire loop with HCl solution and then distilled water before each test. If the loop is clean, it should not give a color when placed in the flame without any sample on it.

### **B. Unknown Metal Ions**

After obtaining the color of each of the known metal ion solutions, repeat the flame test for a single metal unknown and an unknown mixture of two metal ions. From the flame color, identify the metal ions in your unknowns and record your results on your report form.

You may wish to compare the flame color of the unknowns to the color of the flames for the known solutions you observed earlier. For example, if you think your unknown is NaCl, place one wire loop containing NaCl in one side of the flame, and place the wire loop with the unknown in the other side of the flame so that you can see both colors simultaneously.

If available, view the mixture flame through a piece of cobalt glass. Cobalt glass is blue in color because it absorbs yellow light, allowing more blue light to pass through. It makes an effective filter if one component of the mixture, such as sodium, produces a bright yellow flame that obscures another flame color from another ion.

# EXPERIMENT 3 – Flame Tests & Electron Configuration

REPORT FORM

Name \_\_\_\_\_

Instructor \_\_\_\_\_

Date \_\_\_\_\_

## A. Known Metal Ions

<u>Compound</u>	<u>Metal Ion</u>	<u>Color of Flame</u>
LiCl	Li <sup>+</sup>	.....
NaCl	Na <sup>+</sup>	.....
KCl	K <sup>+</sup>	.....
CaCl <sub>2</sub>	Ca <sup>2+</sup>	.....
SrCl <sub>2</sub>	Sr <sup>2+</sup>	.....
BaCl <sub>2</sub>	Ba <sup>2+</sup>	.....
CuCl <sub>2</sub>	Cu <sup>2+</sup>	.....

## B. Unknown Metal Ions

<u>Unknown Number</u>	<u>Color of Flame</u>	<u>Identification of Metal Ion(s)</u>
.....	.....	.....
.....	.....	.....

# EXPERIMENT 3

Name: \_\_\_\_\_

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## Pre-Laboratory Questions and Exercises

Due before lab begins. Answer in the space provided.

1. Write electron configuration for the alkali metals Li, Na, K, and Rb.

Li \_\_\_\_\_

Na \_\_\_\_\_

K \_\_\_\_\_

Rb \_\_\_\_\_

2. Write the electron configuration for strontium and the strontium(II) ion.

Sr \_\_\_\_\_

Sr<sup>2+</sup> \_\_\_\_\_

3. Name the colors in visible light, beginning with that of highest energy (shortest wavelength).

\_\_\_\_\_

4. What is the maximum number of electrons the first four energy levels can hold?

n=1 \_\_\_\_\_      n=2 \_\_\_\_\_      n=3 \_\_\_\_\_      n=4 \_\_\_\_\_

5. What is monochromatic light?

# EXPERIMENT 3

Name: \_\_\_\_\_

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## Post-Laboratory Questions and Exercises

Due after completing the lab. Answer in the space provided.

1. Write the electron configuration of iron and the iron(III) or ferric ion.

Fe \_\_\_\_\_

Fe<sup>3+</sup> \_\_\_\_\_

2. How many electrons are in the outermost, or *valence*, energy level of the following atoms?

Al \_\_\_\_\_      C \_\_\_\_\_      F \_\_\_\_\_      Na \_\_\_\_\_

3. Red light has a longer wavelength than blue light.

Which of these colors has the higher frequency? \_\_\_\_\_

Which of these colors has the higher energy? \_\_\_\_\_

4. In a flame test experiment, sodium gives the brightest and most persistent color in the flame. Do you think that potassium could be detected visually in the presence of sodium by heating this mixture in a flame? Explain your answer.