Introduction

In section 6.5.2.4 we introduced the spin quantum number and we now want to take a deeper look into why an electron needs one more quantum number to describe its wave function.

Stern-Gerlach Experiment

When scientists analyzed the emission and absorption spectra of the elements more closely, they saw that for elements having more than one electron, nearly all the lines in the spectra were actually pairs of very closely spaced lines. Because each line represents an energy level available to electrons in the atom, there are twice as many energy levels available as would be predicted solely based on the quantum numbers $n$, $l$, and $m_l$. Scientists also discovered that applying a magnetic field caused the lines in the pairs to split farther apart, as is demonstrated by the Stern-Gerlach apparatus of Figure \((\PageIndex{1})\).

Figure \((\PageIndex{1})\): The Stern-Gerlach apparatus. Image used with permission from Wikipedia

In 1925, two graduate students in physics in the Netherlands, George Uhlenbeck (1900–1988) and Samuel Goudsmit (1902–1978), proposed that the splittings were caused by an electron spinning about its axis, much as Earth spins about its axis. When an electrically charged object spins, it produces a magnetic moment parallel to the axis of rotation, making it behave like a magnet. Although the electron cannot be viewed solely as a particle, spinning or otherwise, it is indisputable that it does have a magnetic moment. This magnetic moment is called electron spin.
In an external magnetic field, the electron has two possible orientations (Figure Figure \(\PageIndex{2}\)). These are described by a fourth quantum number \((m_s)\), which for any electron can have only two possible values, designated \(+\frac{1}{2}\) (up) and \(-\frac{1}{2}\) (down) to indicate that the two orientations are opposites; the subscript \(s\) is for spin. An electron behaves like a magnet that has one of two possible orientations, aligned either with the magnetic field or against it.

**Pauli Exclusion Principle**

The implications of electron spin for chemistry were recognized almost immediately by an Austrian physicist, Wolfgang Pauli (1900–1958; Nobel Prize in Physics, 1945), who determined that each orbital can contain no more than two electrons. He developed the Pauli exclusion principle: *No two electrons in an atom can have the same values of all four quantum numbers \((n, l, m_l, m_s)\).*

Note: Pauli Exclusion Principle

*No two electrons in an atom can have the same values of all four quantum numbers \((n, l, m_l, m_s)\).*

By giving the values of \(n\), \(l\), and \(m_l\), we also specify a particular orbital (e.g., 1s with \(n = 1, l = 0, m_l = 0\)). Because \(m_s\) has only two possible values (+½ or −½), two electrons, *and only two electrons*, can occupy any given orbital, one with spin up and one with spin down. With this information, we can proceed to construct the entire periodic table, which was originally based on the physical and chemical properties of the known elements.

**Example \(\PageIndex{1}\)**

List all the allowed combinations of the four quantum numbers \((n, l, m_l, m_s)\) for electrons in a 2p orbital and predict the maximum number of electrons the 2p subshell can accommodate.

**Solution**

**Given:** orbital
**Asked for:** allowed quantum numbers and maximum number of electrons in orbital

**Strategy:**

A. List the quantum numbers \((n, l, m_l)\) that correspond to an \(n = 2p\) orbital. List all allowed combinations of \((n, l, m_l)\).

B. Build on these combinations to list all the allowed combinations of \((n, l, m_l, m_s)\).

C. Add together the number of combinations to predict the maximum number of electrons the 2\(p\) subshell can accommodate.

**Solution:**

A. For a 2\(p\) orbital, we know that \(n = 2, l = n - 1 = 1\), and \(m_l = -l, (-l + 1), \ldots, (l - 1), l\). There are only three possible combinations of \((n, l, m_l)\): \((2, 1, 1)\), \((2, 1, 0)\), and \((2, 1, -1)\).

B. Because \(m_s\) is independent of the other quantum numbers and can have values of only \(+\frac{1}{2}\) and \(-\frac{1}{2}\), there are six possible combinations of \((n, l, m_l, m_s)\): \((2, 1, 1, +\frac{1}{2})\), \((2, 1, 1, -\frac{1}{2})\), \((2, 1, 0, +\frac{1}{2})\), \((2, 1, 0, -\frac{1}{2})\), \((2, 1, -1, +\frac{1}{2})\), and \((2, 1, -1, -\frac{1}{2})\).

C. Hence the 2\(p\) subshell, which consists of three 2\(p\) orbitals (2\(px\), 2\(py\), and 2\(pz\)), can contain a total of six electrons, two in each orbital.

**Exercise \(\PageIndex{1}\)**

List all the allowed combinations of the four quantum numbers \((n, l, m_l, m_s)\) for a 6\(s\) orbital, and predict the total number of electrons it can contain.

**Answer**

\((6, 0, 0, +\frac{1}{2}), (6, 0, 0, -\frac{1}{2})\); two electrons