Ground state electron configurations are the foundation for understanding molecular bonding, properties, and structures. From the electrons in an atom, to the differing orbitals and hybridization, the ground state electron configuration sheds light on many different atomic properties. Fundamentally, understanding electron configuration leads to an understanding of the periodic table.

Introduction

In 1913, Niels Bohr proposed that electrons could orbit an atom at a certain distance without collapsing into the atom, and that each orbit distance had its own energy level. He proposed that each orbital's angular momentum, M, was equal to a multiple, n, of Plank's constant, h, divided by 2π. This gives the equation:

\[ M = n\hbar \text{ where } \hbar = \frac{h}{2\pi} \text{ and } n = 1, 2, 3, 4 \]

This model proposed the Bohr atom, which shows circular orbits surrounding the nucleus.

Orbitals

In addition to having different energy levels, orbitals also have different shapes and orientations, and each can be occupied by two electrons. For each principal quantum number, n, there is one s orbital, three p orbitals, five d orbitals and seven f orbitals. Therefore, an s orbital can hold two electrons, a p orbital can hold six electrons, a d orbital can hold ten electrons, and an f orbital can hold 14 electrons.

Ground State Electron Configuration

Quantum numbers

There are four quantum numbers \( n, l, m_l, \text{ and } m_s \). The principal quantum number \( n \) is a positive integer (1,2,3,4) and it represents the energy of the orbital. The angular momentum quantum number \( l \), is from 0 to \( n - 1 \). The \( l \) values of 0, 1, 2,
and 3 correspond to the s, p, d and f orbitals, respectively. The magnetic quantum number $m_l$ ranges from $-l$ to $+l$. This quantum number dictates the orbital orientation, such as $p_x$, $p_y$, or $p_z$. The quantum spin number $m_s$, is either $+1/2$ or $-1/2$ and it dictates the electron spin.

### Aufbau Principle

The Aufbau principle states that electrons must fill lowest energy shells first.

Following the model, electrons fill the 1s orbital with two electrons, then the 2s with two electrons, then the 2p with six electrons, then the 3s with two electrons, etc.

There are some exceptions to the Aufbau Principle. This occurs mainly with electrons in the d orbital where extra stability is obtained from a half filled or fully filled d orbital. Therefore, if there are 4 electrons, or 9 electrons in the d orbital, it will move one electron from the s orbital below it to fill the extra space.

**Example 1: Chromium**

Cr's electron configuration, following the model would be: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4$, but instead it is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$, because there is extra stability gained from the half-filled d orbital.

### Hund's Rule

Hund's rule states that when filled sub-levels other than s orbital, electrons must not be spin paired in the orbitals until each orbital contains one electron, and no orbital can have two electrons with the same spin ($m_s$).

### Pauli Exclusion Principle

Pauli Exclusion Principle states that no two electrons can have the same quantum numbers. An orbital can only hold 0, 1, or 2 electrons. They must have opposite spins if there are 2 electrons in the orbital.

### Periodic Trend

Valence electron shells in the periodic table follow a trend. This can be referred to as the s block, the p block, the d block
and the f block (lanthanides and actinides) meaning that, in its ground state, an element in a certain "block" will have its valence electrons in the s, p, d, or f orbitals depending.

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How to Write Ground State Electron Configurations

**Basics**

Electron configurations are written using the principal quantum number $n$, followed by the orbital (s, p, d, or f) with the total number of electrons written as a superscript. Example: $1s^2$ For writing ground state electron configurations, a few main steps should be followed.

1. Find the amount of electrons in the atom. Example: Na: $11 \text{ e}\text{^{-}}$ Na$^+$: $10 \text{ e}\text{^{-}}$
2. Fill orbitals following the model until all electrons have been accounted for.

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  1s
  2s  2p
  3s  3p  3d
  4s  4p  4d  4f
  5s  5p  5d  5f
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Example: Na: $11 \text{ e}\text{^{-}}$ $1s^2$ $2s^2$ $2p^6$ $3s^1$ or Na$^+$: $1s^2$ $2s^2$ $2p^6$

3. After that, it is important to check for a nearly half-filled or filled d orbital ($d^4$ or $d^9$) and adjust accordingly by removing an electron from the s orbital beneath it.
Example 2: Chromium

Cr: \(1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^4\) half filled orbital, s orbital beneath it

\(1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5\)

Shorthand

Because writing the entire electron configuration can become cumbersome, there is a shorthand option. It is done by using the symbol of the noble gas in the period above the element to represent the electron configuration before it.

Example: Na: \([\text{Ne}]\ 3s^1\)

References


Outside Links

• http://www.jce.divched.org/JCEDLib/L...ter2window.htm
• http://www.jce.divched.org/JCEDLib/L...ter8window.htm

Problems

1. Write the expanded and shortened ground state electron configuration for Cl.
2. Write the expanded and shortened ground state electron configuration for Cr.
3. Write the expanded and shortened ground state electron configuration for Cu.
4. Write the expanded and shortened ground state electron configuration for Co\(^{2+}\).
5. Write the ground state electron configuration for P\(^{3-}\).

Solution 1. Expanded: \(1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5\)

Shorthand: \([\text{Ne}]\ 3s^2\ 3p^5\)
Solution 2. Expanded: 1s² 2s² 2p⁶ 3s² 3p⁶ 4s¹ 3d⁵
Shorthand: [Ar] 4s¹ 3d⁵

Solution 3. Expanded: 1s² 2s² 2p⁶ 3s² 3p⁶ 4s¹ 3d¹⁰
Shorthand: [Ar] 4s¹ 3d¹⁰

Solution 4. Expanded: 1s² 2s² 2p⁶ 3s² 3p⁶ 4s² 3d⁵
Shorthand: [Ar] 4s² 3d⁵

Solution 5. 1s² 2s² 2p⁶ 3s² 3p⁶

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