Chapter 5 : Electrons in Atoms

Section 3 Notes

Ground-State Electron Configuration

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The arrangement of electrons in the atom is called the electron configuration.

The aufbau principle states that each electron occupies the lowest energy orbital available.

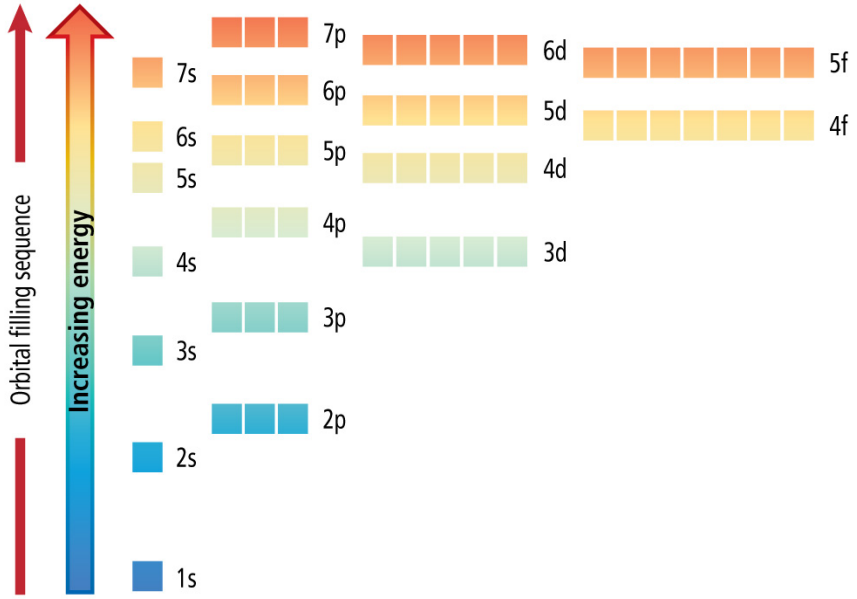

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Table 5.3 Features of the Aufbau Diagram

Feature	Example
All orbitals related to an energy sublevel are of equal energy.	All three 2p orbitals are of equal energy.
In a multi-electron atom, the energy sublevels within a principal energy level have different energies.	The three 2p orbitals are of higher energy than the 2s orbital.
In order of increasing energy, the sequence of energy sublevels within a principal energy level is s, p, d, and f.	If $n = 4$, then the sequence of energy sublevels is 4s, 4p, 4d, and 4f.
Orbitals related to energy sublevels within one principal energy level can overlap orbitals related to energy sublevels within another principal level.	The orbital related to the atom's 4s sublevel has a lower energy than the five orbitals related to the 3d sublevel.

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The Pauli exclusion principle states that a maximum of two electrons can occupy a single orbital, but only if the electrons have opposite spins.

Hund's rule states that single electrons with the same spin must occupy each equal-energy orbital before additional electrons with opposite spins can occupy the same energy level orbitals.

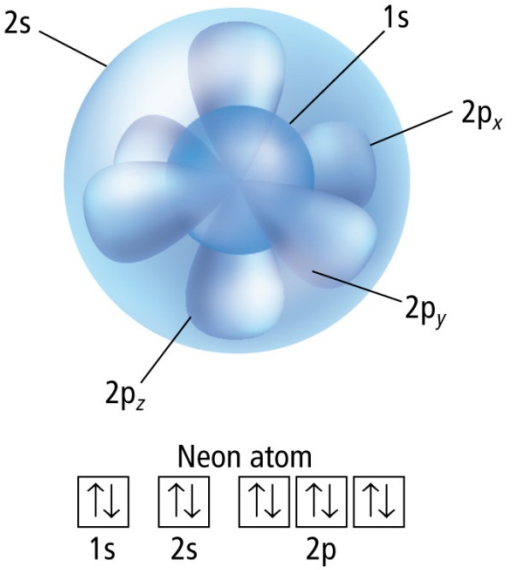

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Table 5.4 Electron Configurations and Orbital Diagrams for Elements 1–10

Element	Atomic Number	Orbital Diagram (1s 2s $2p_x$ $2p_y$ $2p_z$)	Electron Configuration Notation
Hydrogen	1	↑	$1s^1$
Helium	2	↑↓	$1s^2$
Lithium	3	↑↓ ↑	$1s^2\ 2s^1$
Beryllium	4	↑↓ ↑↓	$1s^2\ 2s^2$
Boron	5	↑↓ ↑↓ ↑	$1s^2\ 2s^2\ 2p^1$
Carbon	6	↑↓ ↑↓ ↑ ↑	$1s^2\ 2s^2\ 2p^2$
Nitrogen	7	↑↓ ↑↓ ↑ ↑ ↑	$1s^2\ 2s^2\ 2p^3$
Oxygen	8	↑↓ ↑↓ ↑↓ ↑ ↑	$1s^2\ 2s^2\ 2p^4$
Fluorine	9	↑↓ ↑↓ ↑↓ ↑↓ ↑	$1s^2\ 2s^2\ 2p^5$
Neon	10	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	$1s^2\ 2s^2\ 2p^6$

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Noble gas notation uses noble gas symbols in brackets to shorten inner electron configurations of other elements.

Table 5.5 Electron Configurations for Elements 11–18

Element	Atomic Number	Complete Electron Configuration	Electron Configuration Using Noble Gas
Sodium	11	$1s^2 2s^2 2p^6 3s^1$	$[Ne]3s^1$
Magnesium	12	$1s^2 2s^2 2p^6 3s^2$	$[Ne]3s^2$
Aluminum	13	$1s^2 2s^2 2p^6 3s^2 3p^1$	$[Ne]3s^2 3p^1$
Silicon	14	$1s^2 2s^2 2p^6 3s^2 3p^2$	$[Ne]3s^2 3p^2$
Phosphorus	15	$1s^2 2s^2 2p^6 3s^2 3p^3$	$[Ne]3s^2 3p^3$
Sulfur	16	$1s^2 2s^2 2p^6 3s^2 3p^4$	$[Ne]3s^2 3p^4$
Chlorine	17	$1s^2 2s^2 2p^6 3s^2 3p^5$	$[Ne]3s^2 3p^5$
Argon	18	$1s^2 2s^2 2p^6 3s^2 3p^6$	$[Ne]3s^2 3p^6$ or [Ar]

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The electron configurations (for chromium, copper, and several other elements) reflect the increased stability of half-filled and filled sets of s and d orbitals.

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Valence electrons are defined as electrons in the atom's outermost orbitals—those associated with the atom's highest principal energy level.

Electron-dot structure consists of the element's symbol representing the nucleus, surrounded by dots representing the element's valence electrons.

Valence Electrons (cont.)

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Table 5.6 Electron Configurations and Dot Structures

Element	Atomic Number	Electron Configuration	Electron-Dot Structure
Lithium	3	$1s^2 2s^1$	Li·
Beryllium	4	$1s^2 2s^2$	·Be·
Boron	5	$1s^2 2s^2 2p^1$	·$\dot{B}$·
Carbon	6	$1s^2 2s^2 2p^2$	·$\dot{C}$·
Nitrogen	7	$1s^2 2s^2 2p^3$	·$\ddot{N}$·
Oxygen	8	$1s^2 2s^2 2p^4$	:$\ddot{O}$·
Fluorine	9	$1s^2 2s^2 2p^5$	:$\ddot{F}$·
Neon	10	$1s^2 2s^2 2p^6$	:$\ddot{Ne}$:

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