Electron Configuration

The arrangement of electrons in an atom.

Why care about where electrons are located?

- Allows us to understand how atoms bond together to form chemicals and compounds.

Electrons are found in orbitals

**Orbital:**
A region in an atom where there is a high probability of finding electrons. (Electrons clouds)

Orbitals are filled from low to high energy or starting closest to the nucleus

**ORBITS ARE NOT LIKE A PLANETS!!!!

Placing Electron

1. Each orbital can only hold a specific number of electrons.
2. Each line can only hold 2 electrons
3. Number of electrons
   1. S = 2 e-
   2. P = 6 e-
   3. D = 10 e-
4. Electrons are placed in orbitals starting with the lowest energy

Aufbau Principle
Examples

Hydrogen: ___ e-
1s^1
Helium: ___ e-
1s^2
Lithium: ___ e-
1s^2 2s^1
Beryllium: ___ e-
1s^2 2s^2

Filling s and p orbitals

Neon: ___ e-
2s^2 2p^6 3s^1

Home work

Determine electron configuration

- Bromine
- Strontium
- Antimony
- Rhenium
- Terbium
- Titanium

Electron configurations get long

- All Nobel gases fill their orbitals completely
  - He 1s^2
  - Ne 1s^2 2s^2 2p^6
  - Ar 1s^2 2s^2 2p^6 3s^2 3p^6
- The last orbital has all possible electrons

Examples

Boron: ___ e-
1s^2 2s^2 2p^1

Nitrogen: ___ e-
1s^2 2s^2 2p^3

Magnesium: ___ e-
1s^2 2s^2 2p^6 3s^2

Sulfur: ___ e-
1s^2 2s^2 2p^6 3s^2 3p^4

Sulfur is only #16 of 114 elements.
What problem do you see when writing electron configuration?

Electron Configurations for Elements in Period Three

<table>
<thead>
<tr>
<th>Element</th>
<th>Atomic number</th>
<th>Complete electron configuration</th>
<th>Electron configuration using noble-gas notation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium</td>
<td>11</td>
<td>1s^2 2s^2 2p^6 3s^1</td>
<td>[Ne]3s^1</td>
</tr>
<tr>
<td>Magnesium</td>
<td>12</td>
<td>1s^2 2s^2 2p^6 3s^2</td>
<td>[Ne]3s^2</td>
</tr>
<tr>
<td>Aluminum</td>
<td>13</td>
<td>1s^2 2s^2 2p^6 3s^2 3p^1</td>
<td>[Ne]3s^2 3p^1</td>
</tr>
<tr>
<td>Silicon</td>
<td>14</td>
<td>1s^2 2s^2 2p^6 3s^2 3p^2 3d^1</td>
<td>[Ne]3s^2 3p^2</td>
</tr>
<tr>
<td>Phosphorus</td>
<td>15</td>
<td>1s^2 2s^2 2p^6 3s^2 3p^3</td>
<td>[Ne]3s^2 3p^3</td>
</tr>
<tr>
<td>Sulfur</td>
<td>16</td>
<td>1s^2 2s^2 2p^6 3s^2 3p^4</td>
<td>[Ne]3s^2 3p^4</td>
</tr>
<tr>
<td>Chlorine</td>
<td>17</td>
<td>1s^2 2s^2 2p^6 3s^2 3p^5 4s^1</td>
<td>[Ne]3s^2 3p^5 4s^1</td>
</tr>
<tr>
<td>Argon</td>
<td>18</td>
<td>1s^2 2s^2 2p^6 3s^2 3p^5 4s^2</td>
<td>[Ne]3s^2 3p^5 4s^2 or [Ar]</td>
</tr>
</tbody>
</table>
The Pauli Exclusion Principle

- The maximum of two electrons can occupy a single atomic orbital
- Electrons move in 2 directions

Hunds Rule

- Single electrons with same spin must occupy equal energy orbital before additional electrons can be added in the opposite spin.
- Ex. 2p x___ y___ z___

Orbital Shapes

- Orbits are not planets, they are energy levels
- Not the true shape, much more complex in nature

S orbitals

P orbitals

D orbital
S1, s2 and p2 orbitals

- Ne 10e-
- 2p x__ y__ x__
- 2s __
- 1s __
- 1s^22s^22p^6

Electron Orbital Animations


Using The Periodic Table of Elements to Determine Electron Configuration

To determine the f orbital being filled, subtract two from the period number. (example: period 5 - 2 = 3) The level 4 f orbitals are being filled.)