A p sublevel consists of three p orbitals.
Energy Levels

Energy levels
- are assigned quantum numbers \( n = 1, 2, 3, 4, \) and so on
- increase in energy as the value of \( n \) increases
- have a maximum number of electrons equal to \( 2n^2 \)

<table>
<thead>
<tr>
<th>Energy Level ((n))</th>
<th>1</th>
<th>2</th>
<th>3</th>
<th>4</th>
</tr>
</thead>
<tbody>
<tr>
<td>( 2n^2 )</td>
<td>2(1)^2</td>
<td>2(2)^2</td>
<td>2(3)^2</td>
<td>2(4)^2</td>
</tr>
<tr>
<td>Maximum Number of Electrons</td>
<td>2</td>
<td>8</td>
<td>18</td>
<td>32</td>
</tr>
</tbody>
</table>
Sublevels

A sublevel
- contains electrons with the same energy
- has the same shape but increases in volume at higher energy levels
- is found within each energy level
- is designated by the letters $s$, $p$, $d$, or $f$
Energy of Sublevels

In any energy level
- the $s$ sublevel has the lowest energy
- the $s$ sublevel is followed by the $p$, $d$, $f$ sublevels
- higher sublevels are possible, but only $s$, $p$, $d$, $f$ sublevels are needed to hold the number of electrons in the atoms known today
Number of Sublevels

**Principal energy level**

- \( n = 4 \)
- \( n = 3 \)
- \( n = 2 \)
- \( n = 1 \)

**Types of sublevels**

- \( s \)
- \( p \)
- \( d \)
- \( f \)

The number of sublevels in an energy level is the same as the principal quantum number, \( n \).
Orbitals

An **orbital**

- is a three-dimensional space around a nucleus where an electron is found most of the time
- has a shape that represents electron density (*not a path the electron follows*)
- can hold up to two electrons
- contains two electrons that spin in opposite directions

Opposite spins of electrons in an orbital

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s Orbitals

An *s* orbital

- has a spherical shape around the nucleus
- increases in size around the nucleus as the energy level $n$ value increases
- is a single orbital found in each *s* sublevel

All *s* orbitals have spherical shapes that increase in volume at higher energy levels.
*p* Orbitals

A *p* orbital

- has a two-lobed shape
- is one of three *p* orbitals that make up each *p* sublevel, each aligned along a different axis
- increases in size as the value of *n* increases
Sublevels and Orbitals

Each sublevel consists of a specific number of orbitals.
- an s sublevel contains one s orbital
- a p sublevel contains three p orbitals
- a d sublevel contains five d orbitals
- an f sublevel contains seven f orbitals
The total number of electrons in all the sublevels adds up to give the maximum number of electrons \((2n^2)\) allowed in an energy level.
Learning Check

Indicate the number and type of orbitals in each of the following:

A. 4s sublevel

B. 3d sublevel

C. \( n = 3 \)
Solution

Indicate the number and type of orbitals in each of the following:

A. 4s sublevel
   one 4s orbital
B. 3d sublevel
   five 3d orbitals
C. $n = 3$
   one 3s orbital, three 3p orbitals, and five 3d orbitals
Learning Check

The number of

A. electrons that can occupy a $p$ orbital is
   1) 1          2) 2          3) 3

B. $p$ orbitals in the $2p$ sublevel is
   1) 1          2) 2          3) 3

C. $d$ orbitals in the $n = 4$ energy level is
   1) 1          2) 3          3) 5

D. electrons that can occupy the $4f$ sublevel is
   1) 2          2) 6          3) 14
Solution

The number of

A. electrons that can occupy a \( p \) orbital is
   2) 2

B. \( p \) orbitals in the \( 2p \) sublevel is
   3) 3

C. \( d \) orbitals in the \( n = 4 \) energy level is
   3) 5

D. electrons that can occupy the \( 4f \) sublevel is
   3) 14
In the orbital diagram of carbon, two electrons occupy the 1s orbital, two electrons occupy the 2s orbital, and two electrons each occupy a 2p orbital in the 2p sublevel.
Order of Filling

Energy levels fill with electrons
• in order of increasing energy
• beginning with quantum number $n = 1$
• beginning with $s$ followed by $p$, $d$, and $f$
The orbitals of an atom fill in order of increasing energy of the sublevels beginning with 1s.
Orbital Diagrams

An **orbital diagram** shows
- orbitals as boxes in each sublevel
- electrons in orbitals as vertical arrows
- electrons in the same orbital with opposite spins (up and down vertical arrows)

<table>
<thead>
<tr>
<th>Atomic Number</th>
<th>Element</th>
<th>Orbital Diagram</th>
</tr>
</thead>
<tbody>
<tr>
<td>3</td>
<td>Li</td>
<td><img src="image" alt="orbital_diagram" /></td>
</tr>
</tbody>
</table>
Order of Filling

Electrons in an atom

- fill each orbital in a sublevel with one electron until half full
- then pair up with an electron of opposite spin

<table>
<thead>
<tr>
<th>Atomic Number</th>
<th>Element</th>
<th>Orbital Diagram</th>
</tr>
</thead>
<tbody>
<tr>
<td>7</td>
<td>N</td>
<td></td>
</tr>
<tr>
<td>8</td>
<td>O</td>
<td></td>
</tr>
<tr>
<td>9</td>
<td>F</td>
<td></td>
</tr>
<tr>
<td>10</td>
<td>Ne</td>
<td></td>
</tr>
</tbody>
</table>

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Writing Orbital Diagrams

The **orbital diagram** for carbon consists of:

- two electrons in the 1s orbital
- two electrons in the 2s orbital
- one electron each in two of the 2p orbitals

Electron configuration for carbon: \(1s^22s^22p^2\)
Learning Check

Write the orbital diagrams for

A. nitrogen

B. oxygen

C. magnesium
Solution

Write the orbital diagrams for

<table>
<thead>
<tr>
<th></th>
<th>1s</th>
<th>2s</th>
<th>2p</th>
<th>3s</th>
</tr>
</thead>
<tbody>
<tr>
<td>A. nitrogen</td>
<td><img src="nitrogen_1s_2s_2p_3s.png" alt="Diagram" /></td>
<td><img src="nitrogen_1s_2s_2p_3s.png" alt="Diagram" /></td>
<td><img src="nitrogen_1s_2s_2p_3s.png" alt="Diagram" /></td>
<td><img src="nitrogen_1s_2s_2p_3s.png" alt="Diagram" /></td>
</tr>
<tr>
<td>B. oxygen</td>
<td><img src="oxygen_1s_2s_2p_3s.png" alt="Diagram" /></td>
<td><img src="oxygen_1s_2s_2p_3s.png" alt="Diagram" /></td>
<td><img src="oxygen_1s_2s_2p_3s.png" alt="Diagram" /></td>
<td><img src="oxygen_1s_2s_2p_3s.png" alt="Diagram" /></td>
</tr>
<tr>
<td>C. magnesium</td>
<td><img src="magnesium_1s_2s_2p_3s.png" alt="Diagram" /></td>
<td><img src="magnesium_1s_2s_2p_3s.png" alt="Diagram" /></td>
<td><img src="magnesium_1s_2s_2p_3s.png" alt="Diagram" /></td>
<td><img src="magnesium_1s_2s_2p_3s.png" alt="Diagram" /></td>
</tr>
</tbody>
</table>
Electron Configuration

An electron configuration

- lists the sublevels filling with electrons in order of increasing energy
- uses superscripts to show the number of electrons in each sublevel
- for carbon is as follows:

\[ 1s^22s^22p^2 \]

Read as “one s two, two s two, two p two”
In Period 1, the first two electrons go into the 1s orbital.

<table>
<thead>
<tr>
<th>Atomic Number</th>
<th>Element</th>
<th>Orbital Diagram</th>
<th>Electron Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>H</td>
<td>![1s Diagram]</td>
<td>1s^1</td>
</tr>
<tr>
<td>2</td>
<td>He</td>
<td>![1s Diagram]</td>
<td>1s^2</td>
</tr>
</tbody>
</table>
Abbreviated Configurations

An abbreviated configuration shows

- the symbol of the noble gas in brackets that represents completely filled sublevels
- the remaining electrons in order of their sublevels

Example: Fluorine has a configuration and abbreviated electron configuration of

<table>
<thead>
<tr>
<th>Element</th>
<th>Orbital Diagram</th>
<th>Electron Configuration</th>
<th>Abbreviated Electron Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>F</td>
<td><img src="image" alt="Orbital Diagram" /></td>
<td>$1s^22s^22p^5$</td>
<td>[He]$2s^22p^5$</td>
</tr>
</tbody>
</table>
## Period 2 Configurations

<table>
<thead>
<tr>
<th>Atomic Number</th>
<th>Element</th>
<th>Orbital Diagram</th>
<th>Electron Configuration</th>
<th>Abbreviated Electron Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>3</td>
<td>Li</td>
<td><img src="image" alt="1s 2s" /></td>
<td>$1s^22s^1$</td>
<td>[He]$2s^1$</td>
</tr>
<tr>
<td>4</td>
<td>Be</td>
<td><img src="image" alt="1s 2s" /></td>
<td>$1s^22s^2$</td>
<td>[He]$2s^2$</td>
</tr>
<tr>
<td>5</td>
<td>B</td>
<td><img src="image" alt="1s 2s 2p" /></td>
<td>$1s^22s^22p^1$</td>
<td>[He]$2s^22p^1$</td>
</tr>
<tr>
<td>6</td>
<td>C</td>
<td><img src="image" alt="1s 2s 2p" /></td>
<td>$1s^22s^22p^2$</td>
<td>[He]$2s^22p^2$</td>
</tr>
<tr>
<td>7</td>
<td>N</td>
<td><img src="image" alt="1s 2s 2p" /></td>
<td>$1s^22s^22p^3$</td>
<td>[He]$2s^22p^3$</td>
</tr>
<tr>
<td>8</td>
<td>O</td>
<td><img src="image" alt="1s 2s 2p" /></td>
<td>$1s^22s^22p^4$</td>
<td>[He]$2s^22p^4$</td>
</tr>
<tr>
<td>9</td>
<td>F</td>
<td><img src="image" alt="1s 2s 2p" /></td>
<td>$1s^22s^22p^5$</td>
<td>[He]$2s^22p^5$</td>
</tr>
<tr>
<td>10</td>
<td>Ne</td>
<td><img src="image" alt="1s 2s 2p" /></td>
<td>$1s^22s^22p^6$</td>
<td>[He]$2s^22p^6$</td>
</tr>
</tbody>
</table>
# Period 3 Configurations

<table>
<thead>
<tr>
<th>Atomic Number</th>
<th>Element</th>
<th>Orbital Diagram (3s and 3p orbitals only)</th>
<th>Electron Configuration</th>
<th>Abbreviated Electron Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>11</td>
<td>Na</td>
<td><img src="image" alt="Orbital Diagram for Na" /></td>
<td>1s²2s²2p⁶3s¹</td>
<td>[Ne]3s¹</td>
</tr>
<tr>
<td>12</td>
<td>Mg</td>
<td><img src="image" alt="Orbital Diagram for Mg" /></td>
<td>1s²2s²2p⁶3s²</td>
<td>[Ne]3s²</td>
</tr>
<tr>
<td>13</td>
<td>Al</td>
<td><img src="image" alt="Orbital Diagram for Al" /></td>
<td>1s²2s²2p⁶3s²3p¹</td>
<td>[Ne]3s²3p¹</td>
</tr>
<tr>
<td>14</td>
<td>Si</td>
<td><img src="image" alt="Orbital Diagram for Si" /></td>
<td>1s²2s²2p⁶3s²3p²</td>
<td>[Ne]3s²3p²</td>
</tr>
<tr>
<td>15</td>
<td>P</td>
<td><img src="image" alt="Orbital Diagram for P" /></td>
<td>1s²2s²2p⁶3s²3p³</td>
<td>[Ne]3s²3p³</td>
</tr>
<tr>
<td>16</td>
<td>S</td>
<td><img src="image" alt="Orbital Diagram for S" /></td>
<td>1s²2s²2p⁶3s²3p⁴</td>
<td>[Ne]3s²3p⁴</td>
</tr>
<tr>
<td>17</td>
<td>Cl</td>
<td><img src="image" alt="Orbital Diagram for Cl" /></td>
<td>1s²2s²2p⁶3s²3p⁵</td>
<td>[Ne]3s²3p⁵</td>
</tr>
<tr>
<td>18</td>
<td>Ar</td>
<td><img src="image" alt="Orbital Diagram for Ar" /></td>
<td>1s²2s²2p⁶3s²3p⁶</td>
<td>[Ne]3s²3p⁶</td>
</tr>
</tbody>
</table>
Learning Check

A. The correct electron configuration for nitrogen is
   1) $1s^22p^5$    2) $1s^22s^22p^6$    3) $1s^22s^22p^3$

B. The correct electron configuration for oxygen is
   1) $1s^22p^6$    2) $1s^22s^22p^4$    3) $1s^22s^22p^6$

C. The correct electron configuration for calcium is
   1) $1s^22s^22p^63s^23p^63d^2$
   2) $1s^22s^22p^63s^23p^64s^2$
   3) $1s^22s^22p^63s^23p^8$
Solution

A. The correct electron configuration for nitrogen is
   3) \(1s^22s^22p^3\)

B. The correct electron configuration for oxygen is
   2) \(1s^22s^22p^4\)

C. The correct electron configuration for calcium
   2) \(1s^22s^22p^63s^23p^64s^2\)
Learning Check

Write the electron configuration and abbreviated configuration for each of the following elements:
A. Cl
B. S
C. K
Solution

A. Cl
   \(1s^22s^22p^63s^23p^5\)
   [Ne]3s\(^2\)3p\(^5\)

B. S
   \(1s^22s^22p^63s^23p^4\)
   [Ne]3s\(^2\)3p\(^4\)

C. K
   \(1s^22s^22p^63s^23p^64s^1\)
   [Ar]4s\(^1\)
5.5
Electron Configurations and the Periodic Table

\[ \text{d block} \]

- 3d
- 4d
- 5d
- 6d
Sublevel Blocks on the Periodic Table

The periodic table consists of sublevel blocks arranged in order of increasing energy.

- Groups 1A(1)-2A(2) = s level
- Groups 3A(13)-8A(18) = p level
- Groups 3B(3) to 2B(12) = d level
- Lanthanides/Actinides = f level
Sublevel Blocks

Electron configurations follow the order of sublevels on the periodic table.
Using Sublevel Blocks

To write an electron configuration using **Sublevel blocks**,  
• locate the element on the periodic table  
• starting with H in 1s, write each sublevel block in order going from left to right across each period  
• write the number of electrons in each block

Guide to Writing Electron Configurations with Sublevel Blocks

**STEP 1**  
Locate the element on the periodic table.

**STEP 2**  
Write the filled sublevels in order going across each period.

**STEP 3**  
Complete the configuration by counting the electrons in the unfilled block.
## Writing Electron Configurations

Using the periodic table, write the electron configuration for silicon.

### Solution

<table>
<thead>
<tr>
<th>Period</th>
<th>Sublevel Configuration</th>
<th>Electron Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>Period 1</td>
<td>1s block</td>
<td>1s²</td>
</tr>
<tr>
<td>Period 2</td>
<td>2s → 2p blocks</td>
<td>2s² 2p⁶</td>
</tr>
<tr>
<td>Period 3</td>
<td>3s → 3p blocks</td>
<td>3s² 3p² (Si)</td>
</tr>
</tbody>
</table>

Writing all the sublevel blocks in order gives

\[ 1s² 2s² 2p⁶ 3s² 3p² \]
The 4s orbital has a lower energy than the 3d orbitals.

In potassium, K, the last electron enters the 4s orbital, not the 3d (as shown below).

<table>
<thead>
<tr>
<th></th>
<th>1s</th>
<th>2s 2p</th>
<th>3s 3p 3d</th>
<th>4s</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ar</td>
<td>1s²</td>
<td>2s² 2p⁶</td>
<td>3s² 3p⁶</td>
<td></td>
</tr>
<tr>
<td>K</td>
<td>1s²</td>
<td>2s² 2p⁶</td>
<td>3s² 3p⁶</td>
<td>4s¹</td>
</tr>
<tr>
<td>Ca</td>
<td>1s²</td>
<td>2s² 2p⁶</td>
<td>3s² 3p⁶</td>
<td>4s²</td>
</tr>
<tr>
<td>Sc</td>
<td>1s²</td>
<td>2s² 2p⁶</td>
<td>3s² 3p⁶ 3d¹</td>
<td>4s²</td>
</tr>
<tr>
<td>Ti</td>
<td>1s²</td>
<td>2s² 2p⁶</td>
<td>3s² 3p⁶ 3d²</td>
<td>4s²</td>
</tr>
</tbody>
</table>
Writing Electron Configurations

Using the periodic table, write the electron configuration for manganese.

Solution

Period 1 1s block 1s²
Period 2 2s → 2p blocks 2s² 2p⁶
Period 3 3s → 3p blocks 3s² 3p⁶
Period 4 4s → 3d blocks 4s² 3d⁵ (at Mn)

Writing all the sublevel blocks in order gives
1s² 2s² 2p⁶ 3s² 3p⁶ 4s² 3d⁵
Using the periodic table, write the electron configuration for iodine.

**Solution**

<table>
<thead>
<tr>
<th>Period</th>
<th>Configuration</th>
<th>Electron Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>Period 1</td>
<td>1s block</td>
<td>1s²</td>
</tr>
<tr>
<td>Period 2</td>
<td>2s → 2p blocks</td>
<td>2s² 2p⁶</td>
</tr>
<tr>
<td>Period 3</td>
<td>3s → 3p blocks</td>
<td>3s² 3p⁶</td>
</tr>
<tr>
<td>Period 4</td>
<td>4s → 3d → 3p blocks</td>
<td>4s² 3d¹⁰ 4p⁶</td>
</tr>
<tr>
<td>Period 5</td>
<td>5s → 4d → 5p blocks</td>
<td>5s² 4d¹⁰ 5p⁵</td>
</tr>
</tbody>
</table>

Writing all the sublevel blocks in order gives

1s² 2s² 2p⁶ 3s² 3p⁶ 4s² 3d¹⁰ 4p⁶ 5s² 4d¹⁰ 5p⁵ (iodine)
### 4s Block

<table>
<thead>
<tr>
<th>Period number</th>
<th>s block</th>
<th>Atomic Number</th>
<th>Element</th>
<th>Electron Configuration</th>
<th>Abbreviated Electron Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>H, He</td>
<td>19</td>
<td>K</td>
<td>1s²2s²2p⁶3s²3p⁶4s¹</td>
<td>[Ar]4s¹</td>
</tr>
<tr>
<td>2</td>
<td>2s</td>
<td>20</td>
<td>Ca</td>
<td>1s²2s²2p⁶3s²3p⁶4s²</td>
<td>[Ar]4s²</td>
</tr>
<tr>
<td>3</td>
<td>3s</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>4</td>
<td>4s</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>5</td>
<td>5s</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>6</td>
<td>6s</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>7</td>
<td>7s</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
### 3d Block

#### d block

<table>
<thead>
<tr>
<th>Block</th>
<th>Elements</th>
</tr>
</thead>
<tbody>
<tr>
<td>3d</td>
<td>Sc, Ti, V, Cr, Mn, Fe, Co, Ni, Cu, Zn</td>
</tr>
<tr>
<td>4d</td>
<td>[Ar]4s²3d⁴ (half-filled d sublevel is stable)</td>
</tr>
<tr>
<td>5d</td>
<td>[Ar]4s²3d⁵</td>
</tr>
<tr>
<td>6d</td>
<td>[Ar]4s²3d⁶</td>
</tr>
</tbody>
</table>

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### 4p Block

#### p block

<table>
<thead>
<tr>
<th>Atomic Number</th>
<th>Element</th>
<th>Electron Configuration</th>
<th>Abbreviated Electron Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>31</td>
<td>Ga</td>
<td>1s²2s²2p⁶3s²3p⁶4s²3d¹⁰4p¹</td>
<td>[Ar]4s²3d¹⁰4p¹</td>
</tr>
<tr>
<td>32</td>
<td>Ge</td>
<td>1s²2s²2p⁶3s²3p⁶4s²3d¹⁰4p²</td>
<td>[Ar]4s²3d¹⁰4p²</td>
</tr>
<tr>
<td>33</td>
<td>As</td>
<td>1s²2s²2p⁶3s²3p⁶4s²3d¹⁰4p³</td>
<td>[Ar]4s²3d¹⁰4p³</td>
</tr>
<tr>
<td>34</td>
<td>Se</td>
<td>1s²2s²2p⁶3s²3p⁶4s²3d¹⁰4p⁴</td>
<td>[Ar]4s²3d¹⁰4p⁴</td>
</tr>
<tr>
<td>35</td>
<td>Br</td>
<td>1s²2s²2p⁶3s²3p⁶4s²3d¹⁰4p⁵</td>
<td>[Ar]4s²3d¹⁰4p⁵</td>
</tr>
<tr>
<td>36</td>
<td>Kr</td>
<td>1s²2s²2p⁶3s²3p⁶4s²3d¹⁰4p⁶</td>
<td>[Ar]4s²3d¹⁰4p⁶</td>
</tr>
</tbody>
</table>

*Exceptions to the order of filling.*
Learning Check

A. The last two sublevel blocks in the electron configuration for Co are
   1) $3p^64s^2$
   2) $4s^24d^7$
   3) $4s^23d^7$

B. The last three sublevel blocks in the electron configuration for Sn are
   1) $5s^25p^24d^{10}$
   2) $5s^24d^{10}5p^2$
   3) $5s^25d^{10}5p^2$
Solutions

A. The last two sublevel blocks in the electron configuration for Co are
3) $4s^23d^7$

B. The last three sublevel blocks in the electron configuration for Sn are
2) $5s^24d^{10}5p^2$
Learning Check

Using the periodic table, write the electron configuration and abbreviated configuration for each of the following elements:

A. Zn

B. Sr

C. I
Solution

A. Zn
\[1s^22s^22p^63s^23p^64s^23d^{10}\]
[Ar] \[4s^23d^{10}\]

B. Sr
\[1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^2\]
[Kr]\[5s^2\]

C. I
\[1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^5\]
[Kr]\[5s^24d^{10}5p^5\]
Learning Check

Give the symbol of the element that has

A.  \([\text{Ar}]4s^23d^6\)

B.  Four 3\(p\) electrons

C.  Two electrons in the 4\(d\) sublevel

D.  Electron configuration

\[1s^22s^22p^63s^23p^64s^23d^2\]
Solution

Give the symbol of the element that has

A. \([\text{Ar}]4s^23d^6\)  \(\text{Fe}\)

B. Four 3p electrons  \(\text{S}\)

C. Two electrons in the 4d sublevel  \(\text{Zr}\)

D. Electron configuration  \(\text{Ti}\)
\[1s^22s^22p^63s^23p^64s^23d^2\]
Chapter 5 Electron Configuration and Periodic Trends

5.6 Periodic Trends of the Elements
Valence Electrons

The valence electrons

• determine the chemical properties of an element
• are the electrons in the s and p sublevels in the highest energy level
• are related to the group number of the element

Example: Phosphorus has 5 valence electrons

P Group 5A(15)  
\[1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^3\]
Group Number and Valence Electrons

All the elements in a group have the same number of valence electrons.

Example:
Elements in Group 2A (2) have two (2) valence electrons.

Be \(1s^2\ 2s^2\)
Mg \(1s^2\ 2s^2\ 2p^6\ 3s^2\)
Ca \(1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\)
Sr \(1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\)
### TABLE 5.3 Valence Electrons for Representative Elements in Periods 1–4

<table>
<thead>
<tr>
<th>1A (1)</th>
<th>2A (2)</th>
<th>3A (13)</th>
<th>4A (14)</th>
<th>5A (15)</th>
<th>6A (16)</th>
<th>7A (17)</th>
<th>8A (18)</th>
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<td>As</td>
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<tr>
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<td>36</td>
<td>Kr</td>
<td>4s²4p⁶</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Learning Check

State the number of valence electrons for each:

A.  O
   1)  4       2)  6       3)  8

B.  Al
   1)  13      2)  3       3)  1

C.  Cl
   1)  2       2)  5       3)  7
Solution

State the number of valence electrons for each.

A. O
   2) 6

B. Al
   2) 3

C. Cl
   3) 7
Learning Check

State the number of valence electrons for each.

A. Calcium
   1) 1  2) 2  3) 3

B. Group 6A (16)
   1) 2  2) 4  3) 6

C. Tin
   1) 2  2) 4  3) 14
Solution

State the number of valence electrons for each.

A. Calcium
   2) 2

B. Group 6A (16)
   3) 6

C. Tin
   2) 4
Learning Check

State the number of valence electrons for each.
A. $1s^22s^22p^63s^23p^3$

B. $1s^22s^22p^63s^23p^64s^23d^{10}4p^4$

C. $1s^22s^22p^5$
State the number of valence electrons for each.

A. $1s^22s^22p^63s^23p^3$  
   \[
   5
   \]

B. $1s^22s^22p^63s^23p^64s^23d^{10}4p^4$  
   \[
   6
   \]

C. $1s^22s^22p^5$  
   \[
   7
   \]
Electron-Dot Symbols

An electron-dot symbol
- indicates the valence electrons as dots around the symbol of the element
- for Mg shows two valence electrons placed as single dots on the sides of the symbol Mg

\[ \cdot \text{Mg} \cdot \text{or Mg} \cdot \text{or Mg} \cdot \text{or Mg} \cdot \]
Writing Electron-Dot Symbols

The electron-dot symbols for

- Groups 1A (1) to 4A (14) use single dots
  
  \[
  \begin{align*}
  \text{Na} & \quad \cdot \quad \text{Mg} \\
  \text{Al} & \quad \cdot \quad \text{C}
  \end{align*}
  \]

- Groups 5A (15) to 7A (17) use pairs and single dots
  
  \[
  \begin{align*}
  \cdot & \quad \cdot \quad \cdot \quad \cdot \\
  \cdot & \quad : \quad O \quad : \\
  \end{align*}
  \]
Groups and Electron-Dot Symbols

In a group, all the electron-dot symbols have the same number of valence electrons (dots).

Example: Atoms of elements in Group 2A (2) each have two valence electrons.

2A (2)
- Be
- Mg
- Ca
- Sr
- Ba
## Periodic Table and Electron-Dot Symbols

**TABLE 5.4 Electron-Dot Symbols for Selected Elements in Periods 1–4**

<table>
<thead>
<tr>
<th>Number of Valence Electrons</th>
<th>Group Number</th>
<th>Electron-Dot Symbols</th>
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<td>F·</td>
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<tr>
<td></td>
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<td>Ne·</td>
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</table>
Learning Check

A. \( \cdot \) is the electron-dot symbol for
   1) Na 2) K 3) Al

B. \( \cdot \cdot \cdot \) is the electron-dot symbol of
   1) B 2) N 3) P
Solution

A. \( X \) is the electron-dot symbol for
   1) Na        2) K

B. \( X \) is the electron-dot symbol of
   2) N        3) P
The **atomic radius**

*is the distance from the nucleus to the valence electrons*
Atomic Radius within a Group

The **atomic radius** increases

- going down each group of representative elements
- as the number of energy levels increases
Atomic Radius across a Period

The **atomic radius** decreases

- going from left to right across a period
- as more protons increase the nuclear attraction for valence electrons
Learning Check

Select the element in each pair with the larger atomic radius.

A. Li or K
B. K or Br
C. P or Cl
Solution

Select the element in each pair with the larger atomic radius.

A. K is larger than Li
B. K is larger than Br
C. P is larger than Cl
Ionization Energy

Ionization energy
• is the energy it takes to remove a valence electron

\[ \text{Na}(g) + \text{energy (ionization)} \rightarrow \text{Na}^+ + e^- \]
Ionization Energy

**Metals** have

- 1-3 valence electrons
- lower ionization energies
Nonmetals have
• 5-7 valence electrons
• higher ionization energies
Noble gases have

- complete octets (He has two valence electrons)
- the highest ionization energies in each period
Learning Check

Select the element in each pair with the higher ionization energy.

A. Li or K
B. K or Br
C. P or Cl
Solution

Select the element in each pair with the higher ionization energy.

A. Li  
B. Br  
C. Cl
Sizes of Metal Atoms and Ions

A positive ion
- has lost its valence electrons
- is smaller than the corresponding metal atom (about half the size)
Size of Sodium Ion

The sodium ion \( \text{Na}^+ \)

- forms when the Na atom loses one electron from the third energy level
- is smaller than a Na atom

![Na atom and Na⁺ ion comparison](image)

<table>
<thead>
<tr>
<th>1s</th>
<th>2s</th>
<th>2p</th>
<th>3s</th>
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(1 pm = 10^{-12} m)
Sizes of Nonmetal Atoms and Ions

A negative ion
- has a complete octet
- increases the number of valence electrons
- is larger than the corresponding nonmetal atom (about twice the size)
Size of Fluoride Ion

The fluoride ion $F^-$
- forms when a valence electron is added
- has increased repulsions due to the added valence electron
- is larger than a F atom

$F$ atom
- 1s: 2 electrons
- 2s: 2 electrons
- 2p: 5 electrons

$F^-$ ion
- 1s: 2 electrons
- 2s: 2 electrons
- 2p: 7 electrons

Radius increases from 72 pm to 133 pm

(1 pm = $10^{-12}$ m)
Learning Check

1. Which is larger in each of the following?
   A. K or K$^+$
   B. Al or Al$^{3+}$
   C. S$^{2-}$ or S

2. Which is smaller in each of the following?
   A. N$^{3-}$ or N
   B. Cl or Cl$^-$
   C. Sr$^{2+}$ or Sr
Solution

1. Which is larger in each of the following?
   A. $K \ > \ K^+$
   B. $Al \ > \ Al^{3+}$
   C. $S^{2-} \ > \ S$

2. Which is smaller in each of the following?
   A. $N \ < \ N^{3-}$
   B. $Cl \ < \ Cl^-$
   C. $Sr^{2+} \ < \ Sr$
Concept Map

Electronic Structure and Periodic Trends

- Electromagnetic radiation
  - relates
    - Wavelength, $\lambda$
    - Frequency, $\nu$
    - Energy, $E$

Absorbed go to higher
Emitted go to lower

- Electrons
  - are arranged in
    - Energy levels
      - containing

- Sublevels
- Orbitals
  - shown as

- Electron configurations
  - determine
    - Atomic radius
    - Valence electrons
      - shown as
        - Electron-dot symbols
    - Ionization energy

- Orbital diagrams

Consists of
- Photons