Chapter 5 - Electronic Structure and Periodic Trends

• Electromagnetic Radiation
• Atomic Spectra and Energy Levels
• Energy Levels, Sublevels, & Orbitals
• Orbital Diagrams & Electron Configurations
• Electron Configurations / Periodic Table
• Periodic Trends of the Elements
Electromagnetic Radiation

Electromagnetic radiation
• Is energy that travels as waves through space.
• Is described in terms of wavelength and frequency.
• Moves at the speed of light in a vacuum.
  speed of light = 3.0 x 10^8 m/s
Wavelength

Wavelength ($\lambda$, lamda)

- Is the distance between the top of a wave to the top of the next wave.
- Is expressed in meters (m).
Frequency

Frequency \((\nu, \text{nu})\)

- Is the number of waves that pass by each second.
- Is expressed in cycles per sec = cycles/s = \((s^{-1})\)
- Also uses hertz (Hz) = one wave/s = 1 s\(^{-1}\)
Inverse Relationship of $\gamma$ and $\nu$

The inverse relationship of wavelength and frequency means that:

- Longer wavelengths have lower frequencies.
- Shorter wavelengths have higher frequencies.
- Different types of electromagnetic radiation have different wavelengths and frequencies.

\[ \lambda = \frac{c}{\nu} = 3.0 \times 10^8 \text{ m/s} \]
Learning Check

The short wavelengths of the color blue are dispersed more by the molecules in the atmosphere than longer wavelengths of visible light, which is why we say the sky is blue.
Electromagnetic Spectrum

The electromagnetic spectrum

- Arranges forms of energy from lower to higher.
- Arranges energy from longer to shorter wavelengths.
- Shows visible light with wavelengths from 700-400 nm.
1. Which of the following has the shortest wavelength?
   A. microwaves   B. blue light   C. UV light

2. Which of the following has the lowest energy?
   A. red light   B. blue light   C. green light

3. Which of the following has the highest frequency?
   A. radio waves   B. infrared   C. X-rays
Atomic Spectra and Energy Levels
Spectrum

White light that passes through a prism

• is separated into all colors called a continuous spectrum.
• Gives the colors of a rainbow.
Atomic Spectrum

An atomic spectrum consists of lines

- Of different colors formed when light from a heated element passes through a prism.
- That correspond to photons of energy emitted by electrons going to lower energy levels.
Light Energy and Photons

Light is a stream of small particles called photons that have:

- Energy related to their frequency. Using Plank’s constant ($h$)
  \[ E = h \nu \]
- High energy with a high frequency.
- Low energy with a low frequency.
Electron Energy Levels

Electrons are arranged in specific discrete energy levels that

• Are labeled $n = 1$, $n = 2$, $n = 3$, and so on.

• Increase in energy as $n$ increases.

• Have the electrons with the lowest energy in the first energy level ($n=1$) closest to the nucleus.
Energy Level Changes

An electron

- Absorbs energy to “jump” to a higher energy level.
- Falls to a lower energy level by emitting energy.
- Emits color in the visible range.
Changes in Energy Levels

Example:
An electron in the $n = 3$ energy level falls to the $n = 2$ energy level by emitting photons with energy equal to the energy difference of the two energy levels.

Example:
An electron in the $n = 5$ energy level moves to the $n = 2$ energy level by releasing the energy equal to the energy difference of the fifth and second energy levels.
Energy Emitted

\[ n = 5 \rightarrow n = 2 \]

\[ n = 3 \rightarrow n = 2 \]
Learning Check

In each of the following energy level changes, indicate if energy is
A) absorbed  B) emitted  C) not changed

1. An electron moves from the first energy level ($n = 1$) to the third energy level ($n = 3$). **absorbed**
2. An electron falls from the third energy level to the second energy level. **emitted**
3. An electron moves within the third energy level. **not changed**
Energy Levels, Sublevels, and Orbitals
Energy Levels

Energy levels

- Are assigned quantum numbers $n = 1, 2, 3, 4$ and so on. Aka principal quantum numbers
- 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</th>
<th>Maximum number of electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>$n = 1$</td>
<td>$2(1)^2 = 2(1) = 2$</td>
</tr>
<tr>
<td>$n = 2$</td>
<td>$2(2)^2 = 2(4) = 8$</td>
</tr>
<tr>
<td>$n = 3$</td>
<td>$2(3)^2 = 2(9) = 18$</td>
</tr>
</tbody>
</table>
Sublevels

- Contain electrons with the same energy.
- Are found within each energy level.
- Are designated by the letters $s, p, d, f$.

The number of sublevels is equal to the value of the principal quantum number ($n$).
Number of Sublevels

<table>
<thead>
<tr>
<th>Principal energy level</th>
<th>s</th>
<th>p</th>
<th>Types of sublevels</th>
</tr>
</thead>
<tbody>
<tr>
<td>$n = 4$</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$n = 3$</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$n = 2$</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$n = 1$</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
## Table 5.2

**Electron Capacity in Sublevels for Principal Energy Levels 1 – 4**

<table>
<thead>
<tr>
<th>Principal Energy Level (Shell)</th>
<th>Number of Sublevels</th>
<th>Type of Sublevels</th>
<th>Maximum Number of Electrons</th>
<th>Maximum Total Electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>4</td>
<td>4</td>
<td>4f</td>
<td>14</td>
<td>32</td>
</tr>
<tr>
<td></td>
<td></td>
<td>4d</td>
<td>10</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>4p</td>
<td>6</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>4s</td>
<td>2</td>
<td></td>
</tr>
<tr>
<td>3</td>
<td>3</td>
<td>3d</td>
<td>10</td>
<td>18</td>
</tr>
<tr>
<td></td>
<td></td>
<td>3p</td>
<td>6</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>3s</td>
<td>2</td>
<td></td>
</tr>
<tr>
<td>2</td>
<td>2</td>
<td>2p</td>
<td>6</td>
<td>8</td>
</tr>
<tr>
<td></td>
<td></td>
<td>2s</td>
<td>2</td>
<td></td>
</tr>
<tr>
<td>1</td>
<td>1</td>
<td>1s</td>
<td>2</td>
<td>2</td>
</tr>
</tbody>
</table>

**Note:** the designation of “n” in each type of sublevel.
Energy of Sublevels

In any energy level,

- The $s$ sublevel has the lowest energy.
- The $s$ sublevel is followed by the $p$, $d$, and $f$ sublevels in order of increasing energy.
- Higher sublevels are possible, but only $s$, $p$, $d$, and $f$ sublevels are needed to hold the electrons in the atoms known today.
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 2 electrons.
- Contains two electrons that must spin in opposite directions. Why?
Orbitals

The shape of the orbital represents the electron density or the probability of finding an electron within that space.

Quantum Mechanics provides a means to mathematically describe the probability or shape of the orbital as well as other parameters of interest.

Quantum Chemistry or Quantum Physics are topics of advanced courses and make use of advanced calculus and differential equations.
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.
**$p$ Orbitals**

A $p$ orbital

- Has a two-lobed shape.
- Is one of three $p$ orbitals that make up each $p$ sublevel.
- 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.
Learning Check

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

1. 4s sublevel
   one 4s orbital
2. 3d sublevel
   five 3d orbitals
3. $n = 3$
   one 3s orbital, three 3p orbitals, and five 3d orbitals
Learning Check

The number of

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

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

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

4. Electrons that can occupy the $4f$ sublevel are
   A) 2    B) 6    C) 14
Writing Orbital Diagrams and Electron Configurations

Orbital diagram of carbon

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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 $f$. 
Energy Diagram for Sublevels
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 to indicate spin).

Orbital diagram for Li

\[
\begin{array}{c}
1s \quad 2s \\
\text{filled} \quad \text{half-filled}
\end{array}
\]

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Order of Filling

Electrons in an atom

- Fill orbitals in sublevels of the same type with one electron until half full,
- Then pair up in the orbitals using opposite spins.

<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>1S 2S 2P</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>
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.
Learning Check

Write the orbital diagrams for

1. Nitrogen

2s

2p

3s

3p

2. Oxygen

2p

3s

3p

3. Magnesium

2p

3s

3p

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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 neon is as follows:

\[
\begin{array}{c|c}
\text{sublevel} & \text{number of electrons} \\
\hline
1s^2 & 2 \\
2s^2 & 2 \\
2p^6 & 8 \\
\end{array}
\]

\[\text{[Ne]}\]
Period 1 Configurations

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 orbital diagram]</td>
<td>1s¹</td>
</tr>
<tr>
<td>2</td>
<td>He</td>
<td>![1s orbital diagram]</td>
<td>1s²</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: Chlorine has a configuration of

\[1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^5\]

\([\text{Ne}]\]

The abbreviated configuration for chlorine is

\([\text{Ne}] \ 3s^2 \ 3p^5\)
### Period 2 Configurations

<table>
<thead>
<tr>
<th>Atomic Number</th>
<th>Element</th>
<th>Orbital Diagram</th>
<th>Electron Configuration</th>
<th>Abbreviated Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>3</td>
<td>Li</td>
<td><img src="image" alt="Orbital Diagram" /></td>
<td>$1s^22s^1$</td>
<td>[He] $2s^1$</td>
</tr>
<tr>
<td>4</td>
<td>Be</td>
<td><img src="image" alt="Orbital Diagram" /></td>
<td>$1s^22s^2$</td>
<td>[He] $2s^2$</td>
</tr>
<tr>
<td>5</td>
<td>B</td>
<td><img src="image" alt="Orbital Diagram" /></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="Orbital Diagram" /></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="Orbital Diagram" /></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="Orbital Diagram" /></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="Orbital Diagram" /></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="Orbital Diagram" /></td>
<td>$1s^22s^22p^6$</td>
<td>[He] $2s^22p^6$</td>
</tr>
</tbody>
</table>

*Unpaired electrons*
## 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 Form</th>
</tr>
</thead>
<tbody>
<tr>
<td>11</td>
<td>Na</td>
<td><img src="image" alt="3s and 3p orbitals diagram" /></td>
<td>1s²2s²2p⁶3s¹</td>
<td>[Ne] 3s¹</td>
</tr>
<tr>
<td>12</td>
<td>Mg</td>
<td><img src="image" alt="3s and 3p orbitals diagram" /></td>
<td>1s²2s²2p⁶3s²</td>
<td>[Ne] 3s²</td>
</tr>
<tr>
<td>13</td>
<td>Al</td>
<td><img src="image" alt="3s and 3p orbitals diagram" /></td>
<td>1s²2s²2p⁶3s²3p¹</td>
<td>[Ne] 3s²3p¹</td>
</tr>
<tr>
<td>14</td>
<td>Si</td>
<td><img src="image" alt="3s and 3p orbitals diagram" /></td>
<td>1s²2s²2p⁶3s²3p²</td>
<td>[Ne] 3s²3p²</td>
</tr>
<tr>
<td>15</td>
<td>P</td>
<td><img src="image" alt="3s and 3p orbitals diagram" /></td>
<td>1s²2s²2p⁶3s²3p³</td>
<td>[Ne] 3s²3p³</td>
</tr>
<tr>
<td>16</td>
<td>S</td>
<td><img src="image" alt="3s and 3p orbitals diagram" /></td>
<td>1s²2s²2p⁶3s²3p⁴</td>
<td>[Ne] 3s²3p⁴</td>
</tr>
<tr>
<td>17</td>
<td>Cl</td>
<td><img src="image" alt="3s and 3p orbitals diagram" /></td>
<td>1s²2s²2p⁶3s²3p⁵</td>
<td>[Ne] 3s²3p⁵</td>
</tr>
<tr>
<td>18</td>
<td>Ar</td>
<td><img src="image" alt="3s and 3p orbitals diagram" /></td>
<td>1s²2s²2p⁶3s²3p⁶</td>
<td>[Ne] 3s²3p⁶</td>
</tr>
</tbody>
</table>

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Learning Check

1. The correct electron configuration for nitrogen is

   A) $1s^2 \ 2p^5$  
   B) $1s^2 \ 2s^2 \ 2p^6$  
   C) $1s^2 \ 2s^2 \ 2p^3$

2. The correct electron configuration for oxygen is

   A) $1s^2 \ 2p^6$  
   B) $1s^2 \ 2s^2 \ 2p^4$  
   C) $1s^2 \ 2s^2 \ 2p^6$

3. The correct electron configuration for calcium

   A) $1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6 \ 3d^2$
   B) $1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6 \ 4s^2$
   C) $1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^8$
Learning Check

Write the electron configuration and abbreviated configuration for each of the following elements:

1. Cl
   \[
   1s^2 2s^2 2p^6 3s^2 3p^5
   \]
   \[
   [Ne] 3s^2 3p^5
   \]

2. S
   \[
   1s^2 2s^2 2p^6 3s^2 3p^4
   \]
   \[
   [Ne] 3s^2 3p^4
   \]

3. K
   \[
   1s^2 2s^2 2p^6 3s^2 3p^6 4s^1
   \]
   \[
   [Ar] 4s^1
   \]
Electron Configuration and the Periodic Table
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

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Using Sublevel Blocks

To write an electron configuration using sublevel blocks

• Locate the element on the periodic table.
• Write each sublevel block in order going left to right across each period starting with H.
• Write the number of electrons in each block.
Writing Electron Configurations

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

Solution

Period 1  $1s$ block  $1s^2$
Period 2  $2s \rightarrow 2p$ blocks  $2s^2 \ 2p^6$
Period 3  $3s \rightarrow 3p$ blocks  $3s^2 \ 3p^2$ (at Si)

Writing all the sublevel blocks in order gives:

$1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^2$
Note:

4s is lower than 3d
5s is lower than 4d
Electron Configurations $d$ Sublevel

- The $4s$ orbital has a lower energy than the $3d$ orbitals.
- In potassium $^{19}_{\text{K}}$, the last electron enters the $4s$ orbital, not the $3d$.

<table>
<thead>
<tr>
<th></th>
<th>1s</th>
<th>2s $2p$</th>
<th>3s $3p$</th>
<th>3d</th>
<th>4s</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^{18}_{\text{Ar}}$</td>
<td>1s$^2$</td>
<td>2s$^2$ 2p$^6$</td>
<td>3s$^2$ 3p$^6$</td>
<td></td>
<td></td>
</tr>
<tr>
<td>$^{19}_{\text{K}}$</td>
<td>1s$^2$</td>
<td>2s$^2$ 2p$^6$</td>
<td>3s$^2$ 3p$^6$</td>
<td>4s$^1$</td>
<td></td>
</tr>
<tr>
<td>$^{20}_{\text{Ca}}$</td>
<td>1s$^2$</td>
<td>2s$^2$ 2p$^6$</td>
<td>3s$^2$ 3p$^6$</td>
<td>4s$^2$</td>
<td></td>
</tr>
<tr>
<td>$^{21}_{\text{Sc}}$</td>
<td>1s$^2$</td>
<td>2s$^2$ 2p$^6$</td>
<td>3s$^2$ 3p$^6$</td>
<td>3d$^1$ 4s$^2$</td>
<td></td>
</tr>
<tr>
<td>$^{22}_{\text{Ti}}$</td>
<td>1s$^2$</td>
<td>2s$^2$ 2p$^6$</td>
<td>3s$^2$ 3p$^6$</td>
<td>3d$^2$ 4s$^2$</td>
<td></td>
</tr>
</tbody>
</table>
Writing Electron Configurations

Using the periodic table, write the electron configuration for Manganese 7B(7).

Solution

Mn is in Period 4

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

[\text{Mn}] 4s² 3d⁵
Writing Electron Configurations

Using the periodic table, write the electron configuration for Iodine 7A(17).

Solution

I is in Period 5

Period 1 1s block 1s²
Period 2 2s → 2p blocks 2s² 2p⁶
Period 3 3s → 3p blocks 3s² 3p⁶
Period 4 4s → 3d → 3p blocks 4s² 3d¹⁰ 4p⁶
Period 5 5s → 4d → 5p blocks 5s² 4d¹⁰ 5p⁵

Writing all the sublevel blocks in order gives:
1s² 2s² 2p⁶ 3s² 3p⁶ 4s² 3d¹⁰ 4p⁶ 5s² 4d¹⁰ 5p⁵ (iodine)
**4s Block**

### 4s Block Elements

<table>
<thead>
<tr>
<th>Period</th>
<th>Element</th>
<th>Electronic Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>H</td>
<td>1s²</td>
</tr>
<tr>
<td>1</td>
<td>He</td>
<td>1s²</td>
</tr>
<tr>
<td>2</td>
<td>2s</td>
<td>1s²2s²</td>
</tr>
<tr>
<td>3</td>
<td>3s</td>
<td>1s²2s²2p⁶</td>
</tr>
<tr>
<td>4</td>
<td>4s</td>
<td>1s²2s²2p⁶3s²</td>
</tr>
<tr>
<td>5</td>
<td>5s</td>
<td>1s²2s²2p⁶3s²3p⁶</td>
</tr>
<tr>
<td>6</td>
<td>6s</td>
<td>1s²2s²2p⁶3s²3p⁶4s¹</td>
</tr>
<tr>
<td>7</td>
<td>7s</td>
<td>1s²2s²2p⁶3s²3p⁶4s²</td>
</tr>
</tbody>
</table>

19  K  \(1s^22s^22p^63s^23p^64s^1\)

20  Ca \(1s^22s^22p^63s^23p^64s^2\)
The half-filled or fully filled 3d orbitals are more stable than a filled 4s and partially filled 3d.
4p Block

$p$ block

4p Block Elements

31  Ga  $1s^22s^22p^63s^23p^64s^23d^{10}4p^1$
32  Ge  $1s^22s^22p^63s^23p^64s^23d^{10}4p^2$
33  As  $1s^22s^22p^63s^23p^64s^23d^{10}4p^3$
34  Se  $1s^22s^22p^63s^23p^64s^23d^{10}4p^4$
35  Br  $1s^22s^22p^63s^23p^64s^23d^{10}4p^5$
36  Kr  $1s^22s^22p^63s^23p^64s^23d^{10}4p^6$
Learning Check

1. The last two sublevels blocks in the electron configuration for Co 8B(9) Period 4 are
   A) $3p^64s^2$
   B) $4s^24d^7$
   C) $4s^23d^7$

2. The last three sublevel blocks in the electron configuration for Sn 4A(14) Period 5 are
   A) $5s^25p^24d^{10}$
   B) $5s^24d^{10}5p^2$
   C) $5s^25d^{10}5p^2$
Learning Check

Give the symbol of the element that has

1. \([\text{Ar}]4s^2\ 3d^6\)  \(\text{Fe}\)

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

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

4. The element that has the electron configuration

\[1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^2\]  \(\text{Ti}\)
Periodic Trends of the Elements

Li atom

Na atom

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Valence Electrons Review

The valence electrons

- Determine the chemical properties of the elements.
- 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 Review

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 $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2$
## Periodic Table and Valence Electrons

### Table 9.3

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<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>C</td>
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<td>F</td>
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<td>2s²2p²</td>
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<td>2s²2p⁴</td>
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<td>4s²4p⁴</td>
<td>4s²4p⁵</td>
<td>4s²4p⁶</td>
</tr>
</tbody>
</table>
Atomic Size

Atomic radius
• Is the distance from the nucleus to the valence electrons.
Atomic Radius Within A Group

Atomic radius increases

• Going down each group of representative elements.
• As the number of energy levels increases.
• Ionization energy decreases.
Atomic Radius Across a Period

Atomic radius decreases

- Going from left to right across a period.
- As more protons increase nuclear attraction for valence electrons.

![Diagram showing Li, C, and Ne atoms with their atomic radii and electron configurations.](image)
Sizes of Metal Atoms and Ions

A positive ion

• Has lost its valence electrons.
• Is smaller (about half the size) than its corresponding metal atom.
Size of Sodium Ion

The sodium ion $\text{Na}^+$

- Forms when the Na atom loses one electron from the $3^{\text{rd}}$ energy level.
- Is smaller than a Na atom.
Sizes of Nonmetal Atoms and Ions

A negative ion

• Has a complete octet.
• Increases the number of valence electrons.
• Is larger (about twice the size) than its corresponding metal atom.
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 $F$ atom.
Ionization Energy Review

Ionization energy

• Is the energy it takes to remove a valence electron.

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

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Metals have
• 1-3 valence electrons.
• Lower ionization energies.
Nonmetals have

- 5-7 valence electrons.
- Have higher ionization energies.
Ionization Energy

Nobel gases have
• Complete octets (He has two valence electrons.)
• Have the highest ionization energies in each period.