Chapter 4-5. Quantum Theory and Atomic Structure

6.1 The Nature of Light

6.2 Atomic Spectra

6.3 The Wave-Particle Duality of Matter and Energy

6.4 The Quantum-Mechanical Model of the Atom
Visible light is a type of *electromagnetic radiation*. The wave properties of electromagnetic radiation are described by three variables:

- **frequency** \( (\nu) \), cycles per second
- **wavelength** \( (\lambda) \), the distance a wave travels in one cycle
- **amplitude**, the height of a wave crest or depth of a trough.

The speed of light is a constant:

\[
\text{speed of light} = c = \nu \times \lambda
\]

\[
= 3.00 \times 10^8 \text{ m/s in a vacuum}
\]
**Figure 6.1** The reciprocal relationship of frequency and wavelength.

Wavelength = distance per cycle

\[ \lambda_A = 2\lambda_B = 4\lambda_C \]

Wavelength

A

\[ \lambda_A \]

B

\[ \lambda_B \]

C

\[ \lambda_C \]

Frequency = cycles per second

\[ \nu_A = \frac{1}{2}\nu_B = \frac{1}{4}\nu_C \]
**Figure 6.4** Different behaviors of waves and particles.

<table>
<thead>
<tr>
<th>Wave</th>
<th>Particle</th>
</tr>
</thead>
<tbody>
<tr>
<td>Air</td>
<td>Trajectory of a pebble</td>
</tr>
<tr>
<td>Water</td>
<td></td>
</tr>
<tr>
<td>Angle of refraction</td>
<td></td>
</tr>
<tr>
<td>Cretsts of waves</td>
<td>Beam of particles</td>
</tr>
</tbody>
</table>

Direction of light wave
Figure 6.2  Differing amplitude (brightness, or intensity) of a wave.
Energy and frequency

A solid object emits visible light when it is heated to about 1000 K. This is called blackbody radiation.

The color (and the intensity ) of the light changes as the temperature changes. Color is related to wavelength and frequency, while temperature is related to energy.

Energy is therefore related to frequency and wavelength:

\[ E = n \ h \nu \]

- \( E \) = energy
- \( n \) is a positive integer
- \( h \) is Planck’s constant
- \( \nu \) is frequency
The Quantum Theory of Energy

Any object (including atoms) can emit or absorb only certain quantities of energy.

Energy is quantized; it occurs in fixed quantities, rather than being continuous. Each fixed quantity of energy is called a quantum.

An atom changes its energy state by emitting or absorbing one or more quanta of energy.

\[ \Delta E = n\hbar \nu \] where \( n \) can only be a whole number.

Energy of one photon (quantum), \( \Delta E = h\nu \)
Figure 6.8 The photoelectric effect.

Incoming light strikes the metal surface.

Freed electrons travel to the electrode and produce a current.

Battery

Evacuated tube

Metal plate

Current meter

Positve electrode

$h\nu$
Figure 6.9A  The line spectrum of atomic hydrogen.

Electrically excited H atoms emit a beam of light that is narrowed by a slit and dispersed by a prism to create a line spectrum.
Figure 6.9B  The line spectra of Hg and Sr.
Rydberg equation

\[ \frac{1}{\lambda} = R \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right) \]

*R is the Rydberg constant = 1.096776 \times 10^7 \text{ m}^{-1}*

for the visible series, \( n_1 = 2 \) and \( n_2 = 3, 4, 5, ... \)
Bohr’s atomic model postulated the following:

- The H atom has only certain energy levels, which Bohr called *stationary states*.
  - Each state is associated with a fixed circular orbit of the electron around the nucleus.
  - The higher the energy level, the farther the orbit is from the nucleus.
  - When the H electron is in the first orbit, the atom is in its lowest energy state, called the *ground state*.
• The atom does not radiate energy while in one of its stationary states.
• The atom changes to another stationary state only by absorbing or emitting a photon.
  – The energy of the photon \((h\nu)\) equals the difference between the energies of the two energy states.
  – When the E electron is in any orbit higher than \(n = 1\), the atom is in an \textit{excited state}. 
Figure 6.12  The Bohr explanation of three series of spectral lines emitted by the H atom.
Figure 6.5  Formation of a diffraction pattern.
Matter and Energy are alternate forms of the same entity.

\[
E = mc^2
\]

All matter exhibits properties of both particles and waves. Electrons have wave-like motion and therefore have only certain allowable frequencies and energies.

Matter behaves as though it moves in a wave, and the de Broglie wavelength for any particle is given by:

\[
\lambda = \frac{h}{mu}
\]

\(m = \text{mass}\)

\(u = \text{speed in m/s}\)
<table>
<thead>
<tr>
<th>Substance</th>
<th>Mass (kg)</th>
<th>Speed (m/s)</th>
<th>λ (m)</th>
</tr>
</thead>
<tbody>
<tr>
<td>slow electron</td>
<td>9x10^{-31}</td>
<td>1.0</td>
<td>7x10^{-4}</td>
</tr>
<tr>
<td>fast electron</td>
<td>9x10^{-31}</td>
<td>5.9x10^6</td>
<td>1x10^{-10}</td>
</tr>
<tr>
<td>alpha particle</td>
<td>6.6x10^{-27}</td>
<td>1.5x10^7</td>
<td>7x10^{-15}</td>
</tr>
<tr>
<td>one-gram mass</td>
<td>1.09x10^{-3 }</td>
<td>0.01</td>
<td>7x10^{-29}</td>
</tr>
<tr>
<td>baseball</td>
<td>0.142</td>
<td>25.0</td>
<td>2x10^{-34}</td>
</tr>
<tr>
<td>Earth</td>
<td>6.0x10^{24}</td>
<td>3.0x10^4</td>
<td>4x10^{-63}</td>
</tr>
</tbody>
</table>
Sample Problem 6.5 Calculating the de Broglie Wavelength of an Electron

PROBLEM: Find the deBroglie wavelength of an electron with a speed of 1.00x10^6 m/s (electron mass = 9.11x10^{-31} kg; \(h = 6.626\times10^{-34} \text{ kg}\cdot\text{m}^2/\text{s}\)).

PLAN: We know the speed and mass of the electron, so we substitute these into Equation 7.5 to find \(\lambda\).

SOLUTION: \[
\lambda = \frac{h}{mu}
\]
\[
\lambda = \frac{6.626\times10^{-34} \text{ kg}\cdot\text{m}^2/\text{s}}{9.109 \times10^{-31} \text{ kg} \times 1.00\times10^6 \text{ m/s}} = 7.27\times10^{-10} \text{ m}
\]
Heisenberg’s Uncertainty Principle states that it is not possible to know both the position \textit{and} momentum of a moving particle at the same time.

\[ \Delta x \cdot \Delta p \geq \frac{h}{4\pi} \]

\[ \Delta x \cdot m \Delta u \geq \frac{h}{4\pi} \]

\(x = \text{position}\)
\(u = \text{speed}\)
\(p = \text{momentum}\)

The more accurately we know the speed, the less accurately we know the position, and vice versa.
Sample Problem 6.6 Applying the Uncertainty Principle

PROBLEM: An electron moving near an atomic nucleus has a speed $6 \times 10^6$ m/s ± 1%. What is the uncertainty in its position ($\Delta x$)?

PLAN: The uncertainty in the speed ($\Delta u$) is given as ±1% (0.01) of $6 \times 10^6$ m/s. We multiply $u$ by 0.01 and substitute this value into Equation 7.6 to solve for $\Delta x$.

SOLUTION: $\Delta u = (0.01)(6 \times 10^6 \text{ m/s}) = 6 \times 10^4 \text{ m/s}$

$$\Delta x \cdot m \Delta u \geq \frac{h}{4\pi}$$

$$\Delta x \geq \frac{h}{4\pi m \Delta u} \geq \frac{6.626 \times 10^{-34} \text{ kg} \cdot \text{m}^2/\text{s}}{4\pi (9.11 \times 10^{-31} \text{ kg})(6 \times 10^4 \text{ m/s})} \geq 1 \times 10^{-9} \text{ m}$$
\[ E = n \hbar \nu \quad E_{\text{tot}} = E_{\text{kin}} + E_{\text{pot}} + E_{\text{magn}} \quad \Delta(mv)\Delta(x) = \hbar \quad \lambda = \hbar/mv \]

\[ E\psi = H\psi \quad \rightarrow \quad E = -2.178 \times 10^{-28}(1/n^2) \]

\[ \psi = R x \Theta x \Phi \]

\[ n = 1, \ l = 0, \ m = 0 \quad 2\exp^{-r/a}/a^{3/2} \quad \text{const} \quad \text{const} \]

Orbital Shape
Figure 6.17

Electron probability density in the ground-state H atom.
\[ \Psi = R \times \Theta \times \Phi \]

\( n = 2, \ l = 0, \ m = 0 \)
\[ = 2 \exp \frac{-r}{a} \frac{a^3}{2} \]
\[ \text{const} \]
\[ \text{const} \]

\( l = 1, \ m = 0, 1, -1 \)
\[ = \text{const} \left( \frac{1}{a} \right) ^{3/2} ( \frac{r}{a} \exp ^{-r/2a} \cos \theta \right) \]
\[ \text{const} \]
\[ \text{const} \ (x, y, z) \]
p Orbitals
p Orbitals
Three orientations

- $p_x$
- $p_y$
- $p_z$
Table 6.2  The Hierarchy of Quantum Numbers for Atomic Orbitals

<table>
<thead>
<tr>
<th>Name, Symbol (Property)</th>
<th>Allowed Values</th>
<th>Quantum Numbers</th>
</tr>
</thead>
<tbody>
<tr>
<td>Principal, $n$ (size, energy)</td>
<td>Positive integer $(1, 2, 3, \ldots)$</td>
<td>1, 2, 3</td>
</tr>
<tr>
<td>Angular momentum, $l$ (shape)</td>
<td>0 to $n - 1$</td>
<td>0, 0, 1</td>
</tr>
<tr>
<td>Magnetic, $m_l$ (orientation)</td>
<td>$-l, \ldots, 0, \ldots, +l$</td>
<td>0, 0, -1, 0, +1, 0, -1, 0, +1, -2, -1, 0, +1, +2</td>
</tr>
</tbody>
</table>
Sample Problem 6.8  Determining Sublevel Names and Orbital Quantum Numbers

PROBLEM: Give the name, magnetic quantum numbers, and number of orbitals for each sublevel with the following quantum numbers:

(a) \( n = 3, l = 2 \)  (b) \( n = 2, l = 0 \)  (c) \( n = 5, l = 1 \)  (d) \( n = 4, l = 3 \)

PLAN: Combine the \( n \) value and \( l \) designation to name the sublevel. Knowing \( l \), we can find \( m_l \) and the number of orbitals.

SOLUTION:

<table>
<thead>
<tr>
<th>( n )</th>
<th>( l )</th>
<th>sublevel name</th>
<th>possible ( m_l ) values</th>
<th># of orbitals</th>
</tr>
</thead>
<tbody>
<tr>
<td>(a)</td>
<td>3</td>
<td>2</td>
<td>3d</td>
<td>-2, -1, 0, 1, 2</td>
</tr>
<tr>
<td>(b)</td>
<td>2</td>
<td>0</td>
<td>2s</td>
<td>0</td>
</tr>
<tr>
<td>(c)</td>
<td>5</td>
<td>1</td>
<td>5p</td>
<td>-1, 0, 1</td>
</tr>
<tr>
<td>(d)</td>
<td>4</td>
<td>3</td>
<td>4f</td>
<td>-3, -2, -1, 0, 1, 2, 3</td>
</tr>
</tbody>
</table>

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Sample Problem 6.9  Identifying Incorrect Quantum Numbers

PROBLEM: What is wrong with each of the following quantum numbers designations and/or sublevel names?

<table>
<thead>
<tr>
<th>n</th>
<th>l</th>
<th>m_l</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>(a)</td>
<td>1</td>
<td>1</td>
<td>0</td>
</tr>
<tr>
<td>(b)</td>
<td>4</td>
<td>3</td>
<td>+1</td>
</tr>
<tr>
<td>(c)</td>
<td>3</td>
<td>1</td>
<td>-2</td>
</tr>
</tbody>
</table>

SOLUTION:

(a) A sublevel with \( n = 1 \) can only have \( l = 0 \), not \( l = 1 \). The only possible sublevel name is 1s.

(b) A sublevel with \( l = 3 \) is an \( f \) sublevel, to a \( d \) sublevel. The name should be 4f.

(c) A sublevel with \( l = 1 \) can only have \( m_l \) values of -1, 0, or +1, not -2.
Figure 6.18  Representations of the 1s, 2s, and 3s orbitals.
Figure 6.19  The 2p orbitals.

A. Radial probability distribution

B. Cross section of electron cloud depiction

C. Accurate probability contour

D. The three p orbitals
Figure 6.20  The 3d orbitals.
Figure 6.22  Energy levels of the H atom.
Chapter 7

Electron Configuration and Chemical Periodicity

7.1 Characteristics of Many-Electron Atoms

7.2 The Quantum-Mechanical Model and the Periodic Table.

7.3 Periodic Trends

7.4 Atomic Properties and Chemical Reactivity
Figure 7.1  The effect of electron spin.
Table 7.1 Summary of Quantum Numbers of Electrons in Atoms

<table>
<thead>
<tr>
<th>Name</th>
<th>Symbol</th>
<th>Permitted Values</th>
<th>Property</th>
</tr>
</thead>
<tbody>
<tr>
<td>Principal</td>
<td>$n$</td>
<td>positive integers (1, 2, 3, …)</td>
<td>orbital energy (size)</td>
</tr>
<tr>
<td>Angular momentum</td>
<td>$l$</td>
<td>integers from 0 to $n$-1</td>
<td>orbital shape (The $l$ values 0, 1, 2, and 3 correspond to s, p, d, and f orbitals, respectively.)</td>
</tr>
<tr>
<td>Magnetic</td>
<td>$m_l$</td>
<td>integers from -$l$ to 0 to $+l$</td>
<td>orbital orientation</td>
</tr>
<tr>
<td>Spin</td>
<td>$m_s$</td>
<td>+⅓ or -⅓</td>
<td>direction of e- spin</td>
</tr>
</tbody>
</table>
Quantum Numbers and The Exclusion Principle

Each electron in any atom is described completely by a set of four quantum numbers. The first three quantum numbers describe the orbital, while the fourth quantum number describes electron spin.

Pauli’s exclusion principle states that no two electrons in the same atom can have the same four quantum numbers.

An atomic orbital can hold a maximum of two electrons and they must have opposing spins.
Figure 6.22  Energy levels of the H atom.
In general, energies of subshells increase as $n$ increases ($1 < 2 < 3$, etc.) and as $l$ increases ($s < p < d < f$).

As $n$ increases, some sublevels overlap.

Figure 7.5

The order for filling shells with electrons.
Figure 7.6  A vertical orbital diagram for the Li ground state.
The **Aufbau principle** is applied – electrons are always placed in the lowest energy sublevel available.

\[ \text{H (Z = 1) } 1s^1 \]

\[ \text{He (Z = 2) } 1s^2 \]

The **exclusion principle** states that each orbital may contain a maximum of 2 electrons, which must have opposite spins.
Hund’s rule specifies that when orbitals of equal energy are available, the lowest energy electron configuration has the maximum number of unpaired electrons with parallel spins.

\[
N \ (Z = 7) \ 1s^2 2s^2 2p^3
\]

\[
\begin{array}{c}
2s \\
\end{array}
\begin{array}{ccc}
2p
\end{array}
\]

\[
\begin{array}{c}
\uparrow \downarrow \\
\end{array}
\begin{array}{ccc}
\uparrow & \uparrow & \uparrow
\end{array}
\]
Sample Problem 7.1  
Determining Quantum Numbers from Orbital Diagrams

**PROBLEM:** Use the orbital diagrams, write a set of quantum numbers for (a) the third electron and (b) the eighth electron of the F atom.

**PLAN:** Identify the electron of interest and note its level \( n \), sublevel, \( l \), orbital \( m_l \) and spin \( m_s \). Count the electrons in the order in which they are placed in the diagram.

**SOLUTION:**

\[
\text{F (} Z = 9) \ 1s^2 2s^2 2p^5
\]

\[
\begin{array}{c}
\uparrow \downarrow \\
1s
\end{array} \quad \begin{array}{c}
\uparrow \downarrow \\
2s
\end{array} \quad \begin{array}{c}
\uparrow \uparrow \uparrow \\
2p
\end{array}
\]

- **For the 3\(^{rd}\) electron:** \( n = 2, \ l = 0, \ m_l = 0, \ m_s = +\frac{1}{2} \)

- **For the 8\(^{th}\) electron:** \( n = 2, \ l = 1, \ m_l = -1, \ m_s = -\frac{1}{2} \)

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Figure 7.7  Depicting orbital occupancy for the first 10 elements.
Table 7.2 Partial Orbital Diagrams and Electron Configurations* for the Elements in Period 3.

<table>
<thead>
<tr>
<th>Atomic Number</th>
<th>Element</th>
<th>Partial Orbital Diagram (3s and 3p Sublevels Only)</th>
<th>Full Electron Configuration†</th>
<th>Condensed Electron Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>11</td>
<td>Na</td>
<td>$3s$ up $3p$ up up up up up up</td>
<td>$[1s^22s^22p^6] 3s^1$</td>
<td>[Ne] $3s^1$</td>
</tr>
<tr>
<td>12</td>
<td>Mg</td>
<td>$3s$ up $3p$ up up up up up up</td>
<td>$[1s^22s^22p^6] 3s^2$</td>
<td>[Ne] $3s^2$</td>
</tr>
<tr>
<td>13</td>
<td>Al</td>
<td>$3s$ up $3p$ up up up up up up</td>
<td>$[1s^22s^22p^6] 3s^23p^1$</td>
<td>[Ne] $3s^23p^1$</td>
</tr>
<tr>
<td>14</td>
<td>Si</td>
<td>$3s$ up $3p$ up up up up up up</td>
<td>$[1s^22s^22p^6] 3s^23p^2$</td>
<td>[Ne] $3s^23p^2$</td>
</tr>
<tr>
<td>15</td>
<td>P</td>
<td>$3s$ up $3p$ up up up up up up</td>
<td>$[1s^22s^22p^6] 3s^23p^3$</td>
<td>[Ne] $3s^23p^3$</td>
</tr>
<tr>
<td>16</td>
<td>S</td>
<td>$3s$ up $3p$ up up up up up up</td>
<td>$[1s^22s^22p^6] 3s^23p^4$</td>
<td>[Ne] $3s^23p^4$</td>
</tr>
<tr>
<td>17</td>
<td>Cl</td>
<td>$3s$ up $3p$ up up up up up up</td>
<td>$[1s^22s^22p^6] 3s^23p^5$</td>
<td>[Ne] $3s^23p^5$</td>
</tr>
<tr>
<td>18</td>
<td>Ar</td>
<td>$3s$ up $3p$ up up up up up up</td>
<td>$[1s^22s^22p^6] 3s^23p^6$</td>
<td>[Ne] $3s^23p^6$</td>
</tr>
</tbody>
</table>

*Colored type indicates the sublevel to which the last electron is added.

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Electron Configuration and Group

Elements in the same group of the periodic table have the same outer electron configuration.

Elements in the same group of the periodic table exhibit similar chemical behavior.

*Similar outer electron configurations correlate with similar chemical behavior.*
Figure 7.8  Condensed electron configurations in the first three periods.

<table>
<thead>
<tr>
<th>Period</th>
<th>1</th>
<th>2</th>
<th>3</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>1</td>
<td>2</td>
<td>3</td>
</tr>
<tr>
<td>1</td>
<td>1s(^1)</td>
<td>H</td>
<td>1s(^1)</td>
</tr>
<tr>
<td>2</td>
<td>[He] 2s(^2)</td>
<td>Li</td>
<td>2s(^2)</td>
</tr>
<tr>
<td></td>
<td>[He] 2s(^2)</td>
<td>Be</td>
<td>2p(^2)</td>
</tr>
<tr>
<td>3</td>
<td>[Ne] 3s(^1)</td>
<td>Na</td>
<td>3s(^2)</td>
</tr>
<tr>
<td></td>
<td>[Ne] 3s(^2)</td>
<td>Mg</td>
<td>3p(^1)</td>
</tr>
<tr>
<td>13</td>
<td>[He] 2s(^2)</td>
<td>Al</td>
<td>3p(^2)</td>
</tr>
<tr>
<td>14</td>
<td>[Ne] 3s(^2)</td>
<td>Si</td>
<td>3p(^2)</td>
</tr>
<tr>
<td>15</td>
<td>[Ne] 3s(^2)</td>
<td>P</td>
<td>3p(^2)</td>
</tr>
<tr>
<td>16</td>
<td>[Ne] 3s(^2)</td>
<td>S</td>
<td>3p(^2)</td>
</tr>
<tr>
<td>17</td>
<td>[Ne] 3s(^2)</td>
<td>Cl</td>
<td>3p(^2)</td>
</tr>
<tr>
<td>18</td>
<td>[He] 2s(^2)</td>
<td>Ar</td>
<td>3p(^6)</td>
</tr>
</tbody>
</table>
The order in which the orbitals are filled can be obtained directly from the periodic table.
Figure 7.10  A periodic table of partial ground-state electron configurations.
Sample Problem 7.2  Determining Electron Configurations

PROBLEM: Using the periodic table on the inside cover of the text (not Figure 7.10 or Table 7.3), give the full and condensed electron configurations, partial orbital diagrams showing valence electrons only, and number of inner electrons for the following elements:

(a) potassium (K; Z = 19)  
(b) technetium (Tc; Z = 43) 
(c) lead (Pb; Z = 82)

PLAN: The atomic number gives the number of electrons, and the periodic table shows the order for filling orbitals. The partial orbital diagram includes all electrons added after the previous noble gas except those in filled inner sublevels.
Sample Problem 7.2

SOLUTION:

(a) For K (Z = 19)

full configuration: $1s^22s^22p^63s^23p^64s^1$

condensed configuration: [Ar] $4s^1$

partial orbital diagram:

\[
\begin{array}{c}
\uparrow \\
4s \\
3d \\
4p
\end{array}
\]

There are 18 core electrons.
Sample Problem 7.2

SOLUTION:

(b) For Tc \( (Z = 43) \)

full configuration \( 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^5 \)

condensed configuration \([\text{Kr}]5s^2 4d^5\)

partial orbital diagram

\[
\begin{array}{c}
\uparrow \downarrow \\
5s \\
\end{array} \quad \begin{array}{cccc}
\uparrow & \uparrow & \uparrow & \uparrow \\
4d \\
\end{array} \quad \begin{array}{c}
\text{ } \\
5p \\
\end{array}
\]

There are 36 core electrons.
Sample Problem 7.2

SOLUTION:

(a) For Pb (Z = 82)

full configuration \[ 1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^{14}5d^{10}6p^2 \]

condensed configuration \[ [\text{Xe}] \ 6s^24f^{14}5d^{10}6p^2 \]

partial orbital diagram

\[ \begin{array}{c}
\uparrow \downarrow \\
6s \\
\end{array} \quad \begin{array}{ccc}
\uparrow \\
6p \\
\end{array} \]

There are 78 core electrons.
**Trends in Ionization Energy**

*Ionization energy (IE)* is the energy required for the complete removal of 1 mol of electrons from 1 mol of gaseous atoms or ions.

Atoms with a *low IE* tend to form *cations*. Atoms with a *high IE* tend to form *anions* (except the noble gases).

Ionization energy tends to *decrease* down a group and *increase* across a period.
Figure 7.15  Periodicity of first ionization energy (IE$_1$).
Beryllium has 2 valence electrons, so $\text{IE}_3$ is much larger than $\text{IE}_2$. 
**Trends in Electron Affinity**

*Electron Affinity (EA)* is the energy change that occurs when 1 mol of electrons is *added* to 1 mol of gaseous atoms or ions.

Atoms with a *low EA* tend to form *cations*. Atoms with a *high EA* tend to form *anions*.

The trends in electron affinity are not as regular as those for atomic size or *IE*.
Figure 7.18  Electron affinities of the main-group elements (in kJ/mol).
Figure 7.19  Trends in three atomic properties.
Electron configurations of Monatomic Ions

Elements at either end of a period gain or lose electrons to attain a filled outer level. The resulting ion will have a noble gas electron configuration and is said to be isoelectronic with that noble gas.

\[
\text{Na}(1s^22s^22p^63s^1) \rightarrow e^- + \text{Na}^+([\text{He}]2s^22p^6) \\
\text{[isoelectronic with Ne]}
\]

\[
\text{Br}([\text{Ar}]4s^23d^{10}4p^5) + e^- \rightarrow \text{Br}^- ([\text{Ar}]4s^23d^{10}4p^6) \\
\text{[isoelectronic with Kr]}
\]
Figure 7.22 Main-group elements whose ions have noble gas electron configurations.