Chapter 4: The Electron

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Based on a PowerPoint presentation by Sarah Temple

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Section 4-1
Electromagnetic Spectrum

Essential Questions

1. What is electromagnetic radiation, and what is the dual particle-wave nature of light?

2. Using the formulas $c = \lambda \nu$ and $E = h\nu$ how can frequency, wavelength and energy of electromagnetic radiation be calculated?
1. Electromagnetic radiation/spectrum
2. Wavelength
3. Frequency
4. The speed of light: $3.00 \times 10^8$ m/s
5. Quantum (of energy)
6. Photon
7. Planck’s constant $6.626 \times 10^{-26}$ J*s
8. Absorption
9. Emission-line spectrum
10. Heisenberg Uncertainty Principle
11. Schrödinger wave equation
12. Quantum theory
The Wave Description of Light

Electromagnetic radiation is a form of energy that exhibits wavelike behavior as it travels through space.
Blast from the Past: Waves

- **Wavelength** = $\lambda$ (m) distance between corresponding points on adjacent waves.

- **Frequency** = $\nu$ (Hz or s$^{-1}$) number of waves that pass a given point in a specific time, usually one second.
Electromagnetic Spectrum

All forms of electromagnetic radiation, arranged by wavelength
Am I an X-ray photon...? Or a radio photon? Or visible?

Oh hell...! Why worry about all that again...? I'm not even sure if I'm a wave or a particle!

PHOTON SELF-IDENTITY PROBLEMS
Louis de Broglie

Explained that electrons have a wave aspect, like light

1. They exist as particle-waves
2. They can exist at specific frequencies
3. These frequencies correspond to specific energies
What do all of these waves have in common?

1. They act like particles (photons), since they can travel through space
2. They travel at the speed of light: $3 \times 10^8$ m/s
3. They act like waves: each has a unique frequency and wavelength

For more information about this topic, see http://www.colorado.edu/physics/2000/waves_particles/wavpart2.html (if you don’t want to write this down, look under the “Chemistry” topic in the links on my webpage)
What do electrons have to do with light and waves and frequencies???

- When have contact with a wave from the EM spectrum, they absorb energy (absorption).
- The electron can’t hold the energy
- The electron spits it back out (emission).
- This ‘rejected’ energy is in the form of EM radiation – it has a frequency and wavelength as well as energy.
- If its wavelength is within the visible light spectrum, you can even see the wave as a color!!!
- If it’s not in the visible spectrum, you can still see it using a “spectroscope”
Electrons absorb energy from EM radiation, and emit energy as EM radiation.
Section 4-2
Quantum Theory

Essential Questions

1. What is the quantum model of the atom?

2. How are quantum numbers used to describe orbitals and electrons?
Section 4-2 Vocabulary

1. Quantum theory
   1. Heisenberg Uncertainty Principle
   2. Schrödinger

2. Quantum numbers
   1. Principal quantum number (n)
   2. Angular momentum quantum number (l)
   3. Magnetic quantum number (m)
   4. Spin quantum number

3. Orbitals
   1. s-orbital
   2. p-orbital
   3. d-orbital
   4. f-orbital
“It is impossible to determine simultaneously both the position and momentum of an electron or any other particle.”

- Schrödinger Wave Equation
  - Developed an equation to explain electrons as waves
  - Reconciled wave-particle duality of the electron
Quantum Model

1. Mathematically describes the wave properties of electrons and other very small particles

2. Wave functions only give you the probability of finding an electron in a given place
Electrons travel in orbitals: a 3-dimensional region around the nucleus that indicates the probable location of an electron.
Atomic Theory marches on...

- Let’s bring it down, and see where atomic theory has gone from Thomson to Heisenberg:
- F:\Chemistry Resource Disk\Powerpoints\Ch04\75063.html
Quantum Numbers

Quantum numbers specify properties of atomic orbitals and the properties of electrons in orbitals.
The Four Defined Quantum Numbers

1\textsuperscript{st} three describe the orbital in which an electron is located

- Principal main energy level (n)
- Angular momentum shape (l)
- Magnetic orientation (m)

4\textsuperscript{th} describes a state of the electron

- Spin
Principal Quantum Number \((n)\)

Main energy level of electron

\[
\begin{array}{ccccccc}
  n & 1 & 2 & 3 & 4 & O & P & Q \\
  2n^2 & 2 & 8 & 18 & 32 & 18 & 18 & 2 \\
  \text{observed fill} & 2 & 8 & 18 & 32 & 18 & 18 & 2 \\
\end{array}
\]
How many electrons are possible per energy level?

• \# e- = 2n^2
• Where n=principal quantum #
Angular Momentum Quantum Number

Symbol: $l$
Indicates type of sublevel (sub shells) I.e. shape or orbital (s, p, d, f)
Angular Momentum Quantum Number

- Each energy level can contain a different number or sub-shells (orbital shapes)

\[
\begin{align*}
n &= 1 \quad &s \text{ orbital} \\
n &= 2 \quad &s, p \text{ orbitals} \\
n &= 3 \quad &s, p, d \text{ orbitals} \\
n &= 4 \quad &s, p, d, f \text{ orbitals}
\end{align*}
\]
• Magnetic Quantum Number

Symbol = m

Orientation of an orbital around the nucleus

Refers to xyz axis
• Magnetic Quantum Number

- Each orientation adds to the number of subshells available.
• Magnetic Quantum Number
• Magnetic Quantum Number

- Each subshell can hold 2 e-

- So...

<table>
<thead>
<tr>
<th></th>
<th>s orbital</th>
<th>p orbital</th>
<th>d orbital</th>
<th>f orbital</th>
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</thead>
<tbody>
<tr>
<td>Orientations</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Can hold...</td>
<td></td>
<td></td>
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</tr>
</tbody>
</table>
### Relationship Among Values of $n$, $l$, and $m_l$ Through $n = 4$

<table>
<thead>
<tr>
<th>$n$</th>
<th>Possible Values of $l$</th>
<th>Subshell Designation</th>
<th>Possible Values of $m_l$</th>
<th>Number of Orbitals in Subshell</th>
<th>Total Number of Orbitals in Shell</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0</td>
<td>$1s$</td>
<td>0</td>
<td>1</td>
<td>1</td>
</tr>
<tr>
<td></td>
<td>1</td>
<td>$2s$</td>
<td>0</td>
<td>1</td>
<td>4</td>
</tr>
<tr>
<td>2</td>
<td>2</td>
<td>$2p$</td>
<td>1, 0, -1</td>
<td>3</td>
<td>4</td>
</tr>
<tr>
<td></td>
<td>1</td>
<td>$3s$</td>
<td>0</td>
<td>1</td>
<td>9</td>
</tr>
<tr>
<td></td>
<td>1</td>
<td>$3p$</td>
<td>1, 0, -1</td>
<td>3</td>
<td>9</td>
</tr>
<tr>
<td></td>
<td>2</td>
<td>$3d$</td>
<td>2, 1, 0, -1, -2</td>
<td>5</td>
<td>9</td>
</tr>
<tr>
<td>3</td>
<td>4</td>
<td>$4s$</td>
<td>0</td>
<td>1</td>
<td>16</td>
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<tr>
<td></td>
<td>1</td>
<td>$4p$</td>
<td>1, 0, -1</td>
<td>3</td>
<td></td>
</tr>
<tr>
<td></td>
<td>2</td>
<td>$4d$</td>
<td>2, 1, 0, -1, -2</td>
<td>5</td>
<td></td>
</tr>
<tr>
<td></td>
<td>3</td>
<td>$4f$</td>
<td>3, 2, 1, 0, -1, -2, -3</td>
<td>7</td>
<td></td>
</tr>
</tbody>
</table>
Spin Quantum Number

How the electron spins on an axis creating a magnetic field

2 spin states of an e−
Section 4-3
Electron Configuration

Essential Questions

1. How are the Aufbau Principle, the Pauli Exclusion Principle, and Hund’s rule used in determining an element’s ground-state electron configuration?

2. How does one indicate an element’s ground-state electron configuration using orbital notation, electron-configuration notation, and noble-gas notation?
Section 4-3 Vocabulary

1. Electron configuration
2. Ground state
3. Aufbau Principle
4. Paulli Exclusion Principle
5. Hund’s Rule
6. Orbital notation
7. Electron-configuration notation
8. Noble-gas notation
Electron Configuration

The arrangement of electrons in an atom

Lowest-energy arrangement is called the element’s ground-state electron configuration

http://physics.aps.org/assets/3c8aab1848eb192d/video-v1.mp4
Rules Governing Electron Configurations
There are three rules for electron configuration:

1. **Aufbau principle**
2. **Pauli Exclusion Principle**
3. **Hund’s Rule**
1st Rule Governing Electron Configurations

Aufbau principle

The rule of placing electrons within atomic orbitals (within subshells)

1. Electrons are placed in the lowest energetically available subshell
2. An orbital can hold at most 2 electrons
The electronic energy level 'filling' order for the first 56 electrons for s, p & d orbitals.

The ascending order in which atomic orbitals are filled. Each box represents one orbital and can contain a maximum of 2 electrons. There is 1 s orbital, 3 p orbitals and 5 d orbitals for each principal quantum level where an s, p or d orbital can exist.
• Filling order mnemonic
• How do I remember the order of the subshells?

Use your handy-dandy periodic table!!!
2nd Rule Governing Electron Configurations

Pauli Exclusion Principle

Define:
No two electrons in the same atom can have the same set of four quantum numbers

Each orbital can only contain 2 electrons with opposite spin states

NOT

\[ \text{NOT} \]

\[ 1s^2 \]
3rd Rule Governing Electron Configurations

**Hund’s Rule**

1. Every orbital in a subshell is singly occupied (i.e. with only one electron) before any of the orbital get two electrons
2. All electrons in singly occupied orbitals have the same spin

**Nitrogen**

\[
\begin{array}{cccc}
\uparrow \downarrow & \uparrow \downarrow & \uparrow & \uparrow \\
1s & 2s & 2p & \\
\end{array}
\]
3rd Rule Governing Electron Configurations

**Yes**

- Nitrogen
  - 1s
  - 2s
  - 2p

**No**

- Nitrogen
  - 1s
  - 2s
  - 2p

- Oxygen
  - 1s
  - 2s
  - 2p

- Carbon
  - 1s
  - 2s
  - 2p
• What are the orbital notations for hydrogen and helium?

• Carbon?

• Oxygen?
More examples
Please take out your “Orbital Notation Practice” worksheet, and let’s make it happen!

• PRACTICE TIME!!!
• Writing Electron Configurations

**Orbital Notation**

\[ 1s^2 \]

**Electron-Configuration Notation**

Example = Al

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

The superscripts add up to the atomic number of the atom.
<table>
<thead>
<tr>
<th>Element</th>
<th>Configuration</th>
<th>Orbital Notation</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>1s(^1)</td>
<td>1</td>
</tr>
<tr>
<td>He</td>
<td>1s(^2)</td>
<td>1\rotatebox{90}{(\uparrow)}</td>
</tr>
<tr>
<td>Li</td>
<td>1s(^2)2s(^1)</td>
<td>1\rotatebox{90}{(\uparrow)} 1</td>
</tr>
<tr>
<td>Be</td>
<td>1s(^2)2s(^2)</td>
<td>1\rotatebox{90}{(\uparrow)} 1\rotatebox{90}{(\uparrow)}</td>
</tr>
<tr>
<td>B</td>
<td>1s(^2)2s(^2)2p(^1)</td>
<td>1\rotatebox{90}{(\uparrow)} 1\rotatebox{90}{(\uparrow)} 1</td>
</tr>
<tr>
<td>C</td>
<td>1s(^2)2s(^2)2p(^2)</td>
<td>1\rotatebox{90}{(\uparrow)} 1\rotatebox{90}{(\uparrow)}</td>
</tr>
<tr>
<td>N</td>
<td>1s(^2)2s(^2)2p(^3)</td>
<td>1\rotatebox{90}{(\uparrow)} 1\rotatebox{90}{(\uparrow)} 1\rotatebox{90}{(\uparrow)} 1\rotatebox{90}{(\uparrow)}</td>
</tr>
<tr>
<td>O</td>
<td>1s(^2)2s(^2)2p(^4)</td>
<td>1\rotatebox{90}{(\uparrow)} 1\rotatebox{90}{(\uparrow)} 1\rotatebox{90}{(\uparrow)} 1\rotatebox{90}{(\uparrow)}</td>
</tr>
</tbody>
</table>
| Cr      | [Ar]4s\(^1\)3d\(^5\) | 1\rotatebox{90}{\(\uparrow\)} 1\rotatebox{90}{\(\uparrow\)} 1\rotatebox{90}{\(\uparrow\)} 1\rotatebox{90}{\(\uparrow\)} 1\rotatebox{90}{\(\uparrow\)} 1\rotatebox{90}{\(\uparrow\)} 1\rotatebox{90}{\(\uparrow\)} 1\rotatebox{90}{\(\uparrow\)}
Writing Electron Configurations

**Noble-Gas Notation**

Condensing electron-configuration notation

\[[\text{Ne}] \ 3s^23p^1\]

where \[[\text{Ne}] = 1s^22s^22p^6\]

So normal electron-configuration:

\[1s^22s^22p^63s^23p^1\]
Electron Configuration of K (19)

\[ 1s^22s^22p^63s^23p^64s \]
Filling Order:

1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p 6s 4f 5d 6p 7s 5f 6d