Chapter 6
Electronic Structure of Atoms
How do We get From Here..... To here

Dalton, indivisible

Schroedinger/Heisenberg Quantum mechanical object
Experiments

Atomic spectra
- Bunsen, Kirchhoff, 1860
  - 1st spectroscope
  - 1st line spectrum
- Lockyer, 1868
  - He in solar system
- Balmer, 1885
  - H line spectrum

Quantum/em theory
- Maxwell, 1861
  - Relate Electricity and magnetism
- Plank, 1900
  - Black body radiation
- Einstein, 1905
  - Photoelectric effect
- Bohr, 1913
  - Applied to atom structure

Atomic structure
- Dalton, 1803
  - atomic nature
- Thompson, 1807
  - electrons e/m
- Faraday, 1834
  - Electricity & Mag.
- Millikan, 1911
  - oil drop
- Rutherford, 1911
  - gold foil/nucleus
Electro-magnetic radiation (light)

• Light is a wave
  ➢ The nature of waves
    • What is a wave?
    • What is waving?
Waves

• Wave: some sort of periodic function
  ➢ something that periodically changes vs. time.

• **wavelength** \((\lambda)\): distance between equivalent points

• **Amplitude**: “height” of wave, maximum displacement of periodic function.
Waves

- The number of waves passing a given point per unit of time is the frequency \( (ν) \).
- For waves traveling at the same velocity, the longer the wavelength, the smaller the frequency.
Waves

\[ v = \text{wavelength} \times \text{frequency} \]

\[ \text{meters} \times \left(\frac{1}{\text{sec}}\right) = \text{m/sec} \]

\[ v = \lambda \nu \]
Waves

Major question:

• What is waving?
• water wave:
  ➢ water height(pressure)
• Sound wave:
  ➢ air pressure
• Light?
Light waves.

- What is waving? Electric field, and perpendicular magnetic field.
- Faraday thought this, Maxwell proved it.
Electromagnetic Radiation

All electromagnetic radiation travels at the speed of light ($c$), $3.00 \times 10^8$ m/s (in a vacuum).

Therefore: $c = \lambda \nu$
Electromagnetic Radiation

• All electromagnetic radiation travels the speed of light ($c$), 3.00 \times 10^8 \, \text{m/s} \, (\text{in a vacuum}).

• Therefore: \quad \mathbf{c} = \mathbf{\lambda} \mathbf{v}
Index of refraction is:
The speed of light is only a constant in a vacuum.

\[ n = \frac{c}{v} \]

The index of refraction of some common materials are given below.

<table>
<thead>
<tr>
<th>Material</th>
<th>( n )</th>
<th>Material</th>
<th>( n )</th>
</tr>
</thead>
<tbody>
<tr>
<td>Vacuum</td>
<td>1</td>
<td>Crown Glass</td>
<td>1.52</td>
</tr>
<tr>
<td>Air</td>
<td>1.0003</td>
<td>Salt</td>
<td>1.54</td>
</tr>
<tr>
<td>Water</td>
<td>1.33</td>
<td>Asphalt</td>
<td>1.635</td>
</tr>
<tr>
<td>Ethyl Alcohol</td>
<td>1.36</td>
<td>Heavy Flint Glass</td>
<td>1.65</td>
</tr>
<tr>
<td>Fused Quartz</td>
<td>1.4585</td>
<td>Diamond</td>
<td>2.42</td>
</tr>
<tr>
<td>Whale Oil</td>
<td>1.460</td>
<td>Lead</td>
<td>2.6</td>
</tr>
</tbody>
</table>

Values of \( n \) come from the CRC Handbook of Chemistry and Physics.
The major issue of late 19\textsuperscript{th} century physics

- What is light?
- What is the relationship between light and energy?
- How does light interact with matter?
Mystery #1: Blackbody radiation

- Why does metal glow when heated?
- Heating, so K.E. of electrons
- What light is given off?
Black Body Radiation

Spectral output of a black body. Black shows that predicted from classical electricity & magnetism. Colored curves are what you actually get.

Light is emitted when atoms vibrate (or oscillate), but they can only oscillate with an energy given by:

- \[ E = n\hbar\nu \]
Mystery 1: Black body radiation

- Higher T leads to shorter wavelength of light
- More K.E., more E
- Must be relationship between E and wavelength
- Plank concluded that energy is quantized. It comes in packets (like fruit snacks) and is proportional to frequency:

\[ E = h \nu \]

where \( h \) is Planck’s constant, \( 6.63 \times 10^{-34} \) J-s. The minimum packet of E.
What did Einstein get the Nobel Prize for?
Mystery #2: The Photo-electric effect

Note, this is what a photo cell does: turn light into work (current).
What might you expect (from normal waves)  

what do you see?  

Constant $\nu$
**Einstein: Light is both a particle and a wave.**

\[ E_{\text{photon}} = \frac{1}{2}mv^2 + h\nu = E_{\text{electron}} \]

Light comes in **packets of energy**. Each **packet** runs into **one electron**. Each packet must have enough E to break electron loose. The rest of the energy goes into kinetic energy.

**Frequency** tells us the E of each packet. **I** tells us how many **packets/second** we get. More packets, more current (more electrons knocked off).
The Nature of light Energy

- Energy, $\lambda$, $\nu$, related:

\[ c = \lambda \nu \]
\[ E = h \nu \]

$c=$ speed of light in vacuum, constant
Mystery number 3: element line spectrum

Gas discharge tube (full of some elemental gas)
Gives off specific frequencies of light only.
Different elements give off different colors. i.e. different energies.

Hydrogen

Neon
The Nature of Light

- A line spectrum of discrete wavelengths is observed from an element.

White light shows a continuous spectrum.
Johann Balmer, School teacher
Just figured out that the lines fit a simple equation:

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

\(R_H=\text{constant}\)
\(n_1\) and \(n_2\) are integers

But why?
Niels Bohr adopted Planck’s assumption and explained these phenomena in this way:

1. Electrons in an atom can only occupy certain orbits (corresponding to certain energies).
Niels Bohr adopted Planck’s assumption and explained these phenomena in this way:

2. Electrons in permitted orbits have specific, “allowed” energies;
Light and matter

• Niels Bohr adopted Planck’s assumption and explained these phenomena in this way:

3. Energy is only absorbed or emitted in such a way as to move an electron from one “allowed” energy state to another; the energy is defined by

\[ E = h\nu \]
The energy absorbed or emitted from electron promotion or demotion can be calculated by the equation:

\[ \Delta E = -R_H \left( \frac{1}{n_f^2} - \frac{1}{n_i^2} \right) \]

where \( R_H \) is the Rydberg constant, \( 2.18 \times 10^{-18} \) J, and \( n_i \) and \( n_f \) are integers, the initial and final energy levels of the electron.
Bohr.

- Using a model that had electrons orbiting the nucleus like planets, Bohr could explain H, but no other elements.

\[ \Delta E = -R_H \left( \frac{1}{n_f^2} - \frac{1}{n_i^2} \right) \]

\[ R_E = \frac{1}{2mec^2} \left( \frac{k ee^2}{hc} \right)^2 \]
The Wave Nature of Matter

- Louis de Broglie: if light can be a particle, maybe matter can be wave-like.

\[
\begin{align*}
Velocity &= \lambda \nu \\
\nu &= \frac{velocity}{\lambda} \\
E &= m(velocity)^2 = h\nu = h\frac{velocity}{\lambda} \\
\lambda &= \frac{h}{m(velocity)}
\end{align*}
\]
Wave-like nature of matter

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

However, the higher the mass, the smaller the wavelength & \( h=6.63 \times 10^{-34} \) J-s, a really small number.

Example: What is \( \lambda \) for a 1 g ball?

\[
\lambda = \frac{6.63 \times 10^{-34} \text{kgm}^2/\text{s}}{.001 \text{kg}(1 \text{m/s})} = 6.63 \times 10^{-31} \text{ m}
\]

wavelengths of everyday objects too small to measure.
Wave-like nature of matter

• What about an electron? \( v = 6 \times 10^6 \text{ m/s} \):
• \( m = 9.1 \times 10^{-28} \text{ g} \).

\[
\lambda = \frac{6.63 \times 10^{-34} \text{kgm}^2/\text{s}}{9.1 \times 10^{-28} (6 \times 10^6 \text{ m/s})} = 1.22 \times 10^{-10} \text{ m} = .122 \text{ nm}
\]

Wavelength of X-rays
Electron microscopy

Because electron wavelengths are very small, you can use them to look at very small things.

HIV virus
100 nm, (light microscope limit 400 nm)

T-lymphocyte
Electron microscopy

Because electron wavelengths are very small, you can use them to look at very small things.

HIV virus 100 nm, (light microscope limit 400 nm)

T-lymphocyte

Nobel Prize in Chemistry 2017:
Jacques Dubochet, Joachim Frank, Richard Henderson
Electron Microscopy
Modified picture of the atom
The Uncertainty Principle

- Heisenberg showed that the more precisely the momentum of a particle is known, the less precisely is its position known:

\[(\Delta x) (\Delta mv) \geq \frac{h}{4\pi}\]

- our uncertainty of the whereabouts of an electron can be greater than the size of the atom!

This is a result of the wave/particle duality of matter
“The clues”

1. Plank: E of light is quantized & depends on frequency
2. Einstein/photo-electric effect: Light behaves like a particle when it interacts with matter
3. Emission spectra/Bohr: Potential E. of electrons are quantized in an atom
5. Standing waves: are quantized inherently

Born/Schroedinger/Jordan: use standing wave analogy to explain electron P.E. in atoms.
Quantum Mechanics
Standing waves

\[ l = \left(\frac{1}{2}\right)\lambda \]

1\(v_o\) = frequency
nodes = 2 (gotta have 2)

\[ l = \left(\frac{2}{2}\right)\lambda = \lambda \]

2\(v_o\) = frequency
nodes = 3

\[ l = \left(\frac{3}{2}\right)\lambda \]

3\(v_o\) = frequency
nodes = 4

\[ l = \left(\frac{4}{2}\right)\lambda = 2\lambda \]

4\(v_o\) = frequency
nodes = 5

Allowed \(v\) and \(\lambda\) quantized.

\[ l = \left(\frac{n}{2}\right)\lambda, \quad n \text{ is integer/quantum } \#\]

frequency = \(n v_o\)
Quantum mechanics

- Each electron can be explained using a standing wave equation (wavefunction)
- Quantized frequency corresponds to quantized Energy (Debroglie, Plank, etc.)
- Integer values are critical to this description: quantum numbers.
Quantum mechanics

Examples of wave equations

Propagating wave

Standing wave

$y = \sin x$

$\Psi = \sqrt{\frac{2}{l}} \sin \frac{\pi x}{l}$

$l = 1/2\lambda$

$\nu_0 =$ frequency

nodes = 2
Quantum mechanics

- Using math we do NOT want to deal with, you can do the same thing for an electron in hydrogen:

$$\Psi = \frac{1}{\sqrt{\pi}} e^{-r}$$

But what, physically is $\Psi$? **What is waving?**

Born (1926): $\Psi^2 = $ probability/volume of finding the electron.
Quantum Mechanics

Plot of $\psi^2$ for hydrogen atom.

The closest thing we now have to a physical picture of an electron.

90% contour, will find electron in blue stuff 90% of the time.
Quantum Mechanics

- The wave equation designated with a lower case Greek \( \psi \).
- \( \psi^2 \), gives the probability density of electron location.
- Where it’s “most likely” to be
Quantum Numbers

• Solving the wave equation gives a set of wave functions, or orbitals, and their corresponding energies.
• Each orbital describes a spatial distribution of electron density.
• An orbital is described by a set of three quantum numbers (integers)
• Why three?
Quantum numbers

• 3 dimensions.

• Need three quantum numbers to define a given wavefunction.

• Another name for wavefunction: **Orbital** (because of Bohr).
Principal Quantum Number, $n$

- The principal quantum number, $n$, describes the energy level on which the orbital resides.
- Largest E difference is between E levels.
- The values of $n$ are integers $> 0$.
- $1, 2, 3, \ldots, n$. 
Azimuthal Quantum Number, $l$

- defines **shape** of the orbital.
- Allowed values of $l$ are integers ranging from 0 to $n - 1$.
- We use letter designations to communicate the different values of $l$ and, therefore, the shapes and types of orbitals.
- Note, allowed quantum number possibilities depend on each other.
Azimuthal Quantum Number, \( l \)

\[ l = 0, 1..., n-1 \]

<table>
<thead>
<tr>
<th>Value of ( l )</th>
<th>0</th>
<th>1</th>
<th>2</th>
<th>3</th>
</tr>
</thead>
<tbody>
<tr>
<td>Type of orbital</td>
<td>( s )</td>
<td>( p )</td>
<td>( d )</td>
<td>( f )</td>
</tr>
</tbody>
</table>

So each of these letters corresponds to a shape of orbital.
Magnetic Quantum Number, $m_l$

- Describes, primarily, the **three-dimensional orientation** of the orbital.
- Values are integers ranging from -$l$ to $l$:
  \[-l \leq m_l \leq l.\]
- Therefore, **on any given energy level**, there can be up to:
  - 1 s ($l=0$) orbital ($m_l=0$),
  - 3 p ($l=1$) orbitals, ($m_l=-1,0,1$)
  - 5 d ($l=2$) orbitals, ($m_l=-2,-1,0,1,2$)
  - 7 f ($l=3$) orbitals, ($m_l=-3,-2,-1,0,1,2,3$)
Magnetic Quantum Number, $m_l$

- Orbitals with the same value of $n$ form a shell.
- Different orbital types within a shell are subshells (s, p, d, f).

<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>2</td>
<td>0</td>
<td>2s</td>
<td>0</td>
<td>1</td>
<td>4</td>
</tr>
<tr>
<td></td>
<td>1</td>
<td>2p</td>
<td>1, 0, −1</td>
<td>3</td>
<td>4</td>
</tr>
<tr>
<td>3</td>
<td>0</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>4</td>
<td>0</td>
<td>4s</td>
<td>0</td>
<td>1</td>
<td>16</td>
</tr>
<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>
s Orbitals

- Value of $l = 0$.
- Spherical in shape.
- Radius of sphere increases with increasing value of $n$. 
s Orbitals

$s$ orbitals:

$n-1$ nodes.

Node: 0 amplitude

0 probability

Note: $1s$ orbitals have 0 nodes!
$p$ Orbitals

- Value of $l = 1$.
- Have two lobes with a **nodal plane** between them.

Note: always 3 $p$ orbitals for a given $n$
$d$ Orbitals

- Value of $l$ is 2.
- 2 nodal planes
- Four of the five orbitals have 4 lobes; the other resembles a $p$ orbital with a doughnut around the center.

Note: always 5 $d$ orbitals for a given $n$. 
# Orbitals and nodes

<table>
<thead>
<tr>
<th>Orbital</th>
<th>Symmetry</th>
<th>Node geometry</th>
<th>radial nodes/shell*</th>
<th>Orbitals / E level</th>
</tr>
</thead>
<tbody>
<tr>
<td>s</td>
<td>spherical</td>
<td>spherical</td>
<td>n-1</td>
<td>1</td>
</tr>
<tr>
<td>p</td>
<td>cylindrical</td>
<td>1 planar</td>
<td>n – 2</td>
<td>3</td>
</tr>
<tr>
<td></td>
<td>around x, y, or z axis</td>
<td>remainder spherical</td>
<td></td>
<td></td>
</tr>
<tr>
<td>d</td>
<td>complex</td>
<td>2 planar surfaces</td>
<td>n - 3</td>
<td>5</td>
</tr>
<tr>
<td></td>
<td></td>
<td>diagonal to Cartesian axis; remainder spherical</td>
<td></td>
<td></td>
</tr>
<tr>
<td>f</td>
<td>complex</td>
<td>complex</td>
<td>n - 4</td>
<td>7</td>
</tr>
</tbody>
</table>

- $n = \text{the shell, with } n = 1 \text{ the ground state or lowest possible energy shell. Thus } n \text{ may have integral values from 1 - infinity.}$
- Nodes:
- Total # of nodes = $n - 1$
- # of angular (planar) nodes = $l$
- # radial (spherical) nodes = $n – 1 - l$
Energies of Orbitals

- For a one-electron hydrogen atom, orbitals on the same energy level have the same energy.
- That is, they are degenerate.
Energies of Orbitals

- As the number of electrons increases, though, so does the repulsion between them.
- Therefore, in many-electron atoms, orbitals on the same energy level are no longer degenerate.
Energies of Orbitals

- For a given energy level \((n)\):
  - Energy:
    - \(s < p < d < f\) (Always)
    - \(s\) lowest energy, where electrons go first
    - Next p
    - Then d

Why?
The closer to the nucleus, the lower the energy
Spin Quantum Number, $m_s$

- A fourth dimension required. Why?
Spin Quantum Number, $m_s$

- A fourth dimension required. Why?
- Time. Adding time changes $E$
- Another integer (quantum number) needed.
- Time dependent Schroedinger equation.
Spin Quantum Number, $m_s$

- This leads to a fourth quantum number, the spin quantum number $m_s$.
- The spin quantum number has only 2 values $+1/2$ and $-1/2$
- Describes magnetic field vector of electron
Why do we call it “spin”?

Note: apparently only two values for the magnetic field.

Because electrons behave like little magnets.
Spin Quantum Number, $m_s$

- And charges that spin produce magnetic fields
The problem with quantum mechanics

- It’s not hard to solve equations for the various wavefunctions if they are all alone (like H)
- The problem is what happens in the presence of other electrons
- **The electron interaction problem**
- Electron interaction so complex, exact solutions are only possible for H!
- **Why?**
  - Electron probabilities overlap a lot, must interact a lot, repulsion keeps them from ever “touching”
Pauli Exclusion Principle

• No two electrons in the same atom can have exactly the same energy.
• No two electrons in the same atom can have identical sets of quantum numbers.
Electron Configurations  Every electron has a home

- Address of each electron unique
- Address consists of four numbers: \(n, l, m_l, m_s\)
- Example:
  - 578 S Shaw Ln. East Lansing MI
- Note, in an address each part depends on others
- Rome Ga, not the same as Rome Italy.
- QM are the same.
Electron Configurations

• Distribution of all electrons in an atom
• Consist of
  ➢ Number denoting the energy level

$4p^5$
Electron Configurations

• Distribution of all electrons in an atom

• Consist of
  ➢ Number denoting the energy level
  ➢ Letter denoting the l quantum number, “orbital type”
Electron Configurations

- Distribution of all electrons in an atom.
- Consist of
  - Number denoting the energy level.
  - Letter denoting the type of orbital.
  - Superscript denoting the number of electrons in those orbitals.
Orbital Diagrams

- Each box represents one orbital.
- Half-arrows represent the electrons.
- The direction of the arrow represents the spin of the electron.

Li

1s

2s

1s^2 2s^1
Hund’s Rule
(of maximum multiplicity)

“For degenerate orbitals, the lowest energy is attained when the number of electrons with the same spin is maximized.”

\[
\begin{align*}
&\begin{array}{c}
1\downarrow \\
1s
\end{array} \quad \begin{array}{c}
1\downarrow \\
2s
\end{array} \quad \begin{array}{c}
1\downarrow \ 1\downarrow \ 1\downarrow \\
2p
\end{array} \\
\text{NOT:} \quad &\begin{array}{c}
1\uparrow \\
1s
\end{array} \quad \begin{array}{c}
1\uparrow \\
2s
\end{array} \quad \begin{array}{c}
1\uparrow \ 1\uparrow \\
2p
\end{array}
\end{align*}
\]
# Electron Configurations

<table>
<thead>
<tr>
<th>Element</th>
<th>Total Electrons</th>
<th>Orbital Diagram</th>
<th>Electron Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li</td>
<td>3</td>
<td>![Orbital Diagram for Li]</td>
<td>1s(^2)2s(^1)</td>
</tr>
<tr>
<td>Be</td>
<td>4</td>
<td>![Orbital Diagram for Be]</td>
<td>1s(^2)2s(^2)</td>
</tr>
<tr>
<td>B</td>
<td>5</td>
<td>![Orbital Diagram for B]</td>
<td>1s(^2)2s(^2)2p(^1)</td>
</tr>
<tr>
<td>C</td>
<td>6</td>
<td>![Orbital Diagram for C]</td>
<td>1s(^2)2s(^2)2p(^2)</td>
</tr>
<tr>
<td>N</td>
<td>7</td>
<td>![Orbital Diagram for N]</td>
<td>1s(^2)2s(^2)2p(^3)</td>
</tr>
<tr>
<td>Ne</td>
<td>10</td>
<td>![Orbital Diagram for Ne]</td>
<td>1s(^2)2s(^2)2p(^6)</td>
</tr>
<tr>
<td>Na</td>
<td>11</td>
<td>![Orbital Diagram for Na]</td>
<td>1s(^2)2s(^2)2p(^6)3s(^1)</td>
</tr>
</tbody>
</table>

*TABLE 6.3 Electron Configurations of Several Lighter Elements*
Why do we accept this wacko stuff?

- It must explain all the data
- It should predict things

- Q.M. is consistent with all our data
  - Black body radiation
  - photoelectric effect,
  - emission spectra of elements,
  - dual wave/particle weirdness, etc.

- **One prediction:**
  - elements with similar electron configuration should have similar chemical properties
Why do we accept this wacko stuff?

It predicts the periodicity of the periodic table!!

We fill orbitals in increasing order of energy.
Different blocks on the periodic table, then correspond to different types of orbitals.

1868: mystery
1920’s: QM
Mystery solved.
Why do we accept this wacko stuff?

It predicts the periodicity of the periodic table!!

- Remember: The periodic table was arranged the way it was based on chemical properties.

- Totally empirical, until now. Based only on observation.
Periodic table tells you about the last electron that went in!!!

- Which will be the highest energy electron, most reactive

- Periodic table also makes it easy to do electron configurations.
### Short cut for writing electron configurations

<table>
<thead>
<tr>
<th>TABLE 6.4 Electron Configurations of the Group 2A and 3A Elements</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Group 2A</strong></td>
</tr>
<tr>
<td>Be</td>
</tr>
<tr>
<td>Mg</td>
</tr>
<tr>
<td>Ca</td>
</tr>
<tr>
<td>Sr</td>
</tr>
<tr>
<td>Ba</td>
</tr>
<tr>
<td>Ra</td>
</tr>
<tr>
<td><strong>Group 3A</strong></td>
</tr>
<tr>
<td>B</td>
</tr>
<tr>
<td>Al</td>
</tr>
<tr>
<td>Ga</td>
</tr>
<tr>
<td>In</td>
</tr>
<tr>
<td>Tl</td>
</tr>
</tbody>
</table>

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Electron configurations of the elements

<table>
<thead>
<tr>
<th>1A</th>
<th>2A</th>
<th>3A</th>
<th>4A</th>
<th>5A</th>
<th>6A</th>
<th>7A</th>
<th>8A</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>Li</td>
<td>Be</td>
<td>B</td>
<td>C</td>
<td>N</td>
<td>O</td>
<td>F</td>
</tr>
<tr>
<td>1s^1</td>
<td>2s^2</td>
<td>2p^2</td>
<td>2s^2</td>
<td>2p^2</td>
<td>2p^5</td>
<td>2p^4</td>
<td>2p^5</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Core</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>[He]</th>
</tr>
</thead>
<tbody>
<tr>
<td>Na</td>
</tr>
<tr>
<td>3s^1</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>[Ne]</th>
</tr>
</thead>
<tbody>
<tr>
<td>K</td>
</tr>
<tr>
<td>3d^1</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>[Ar]</th>
</tr>
</thead>
<tbody>
<tr>
<td>Rb</td>
</tr>
<tr>
<td>4d^1</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>[Kr]</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cs</td>
</tr>
<tr>
<td>5d^1</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>[Xe]</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fr</td>
</tr>
<tr>
<td>6d^1</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>[Rn]</th>
</tr>
</thead>
</table>

Lanthanide series

<table>
<thead>
<tr>
<th>57 La</th>
<th>58 Ce</th>
<th>59 Pr</th>
<th>60 Nd</th>
<th>61 Pm</th>
<th>62 Sm</th>
<th>63 Eu</th>
<th>64 Gd</th>
</tr>
</thead>
<tbody>
<tr>
<td>5d^1</td>
<td>5d^2</td>
<td>5d^3</td>
<td>5d^4</td>
<td>5d^5</td>
<td>5d^6</td>
<td>5d^7</td>
<td>5d^8</td>
</tr>
</tbody>
</table>

Actinide series

<table>
<thead>
<tr>
<th>89 Ac</th>
<th>90 Th</th>
<th>91 Pa</th>
<th>92 U</th>
<th>93 Np</th>
<th>94 Pu</th>
<th>95 Am</th>
<th>96 Cm</th>
</tr>
</thead>
<tbody>
<tr>
<td>5f^1</td>
<td>5f^2</td>
<td>5f^3</td>
<td>5f^4</td>
<td>5f^5</td>
<td>5f^6</td>
<td>5f^7</td>
<td>5f^8</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Metals</th>
<th>Metalloids</th>
<th>Nonmetals</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Some Anomalies

Some irregularities occur when there are enough electrons to half-fill s and d orbitals on a given row.
Some Anomalies

For instance, the electron configuration for Chromium, is $[\text{Ar}]\ 4s^1\ 3d^5$ rather than the expected $[\text{Ar}]\ 4s^2\ 3d^4$. 
Some Anomalies

- This occurs because the 4s and 3d orbitals are very close in energy.
- These anomalies occur in f-block atoms, as well.
What’s on the exam?

- Chapter 4, (everything)
- Chapter 5 (everything)
- Chapter 6 (everything)
Chapter 4.

Solution stoichiometry

Strong vs. weak electrolytes
Know strong electrolytes
  strong acids
  soluble salts

precipitation reactions
Ionic equation
net ionic equation
Neutralization reactions
gas forming reactions
  \( \text{H}_2\text{CO}_3 \rightarrow \text{CO}_2 + \text{H}_2\text{O} \)

Molarity
Dilution
Titration
Chapter 4.

Solution stoichiometry
Molarity
Dilution
Titration
Oxidation reduction
assigning oxidation numbers
who is oxidizing and reducing?
activity series
Chapter 5

• Heat and work
• $E = q + w$
• $H$ is $q$ at constant $P$
• Hess’ s law problems
• Heat of formation problems
• Calorimetry problems
Chapter 6

- History of Light
- Electromagnetic radiation order
- Blackbody radiation
- Photo electric effect
- \( E = h \nu \)
- Quantum numbers
  - What are they for?
  - Why are there 3 or 4?
  - What does each stand for?
  - What is their relationship (\( l = 0, 1, 2 \ldots n-1 \))
- Electron configurations using the periodic table
Test breakdown:

Ch. 4:
- str./weak electrolyte (1)
- Net ionic equation (1)
- Assign/change in Ox Numbers (1)
- Neutralization/titration (1)
- Solution stoich. (1)
- Act. Series (1)

Ch. 5
- Work, system/surroundings (1)
- Enthalpy of rxn calc/Hess Law (1)
- Calorimetry (1)
- Enthalpy of phase change (1)

CH 6.
- Theory/experiments (1)
- E=hν (1)
- ν = λν (1)
- EM radiation (1)
- Orbital energies (1)
- QM numb (2)
- Electron config (3)
- Orbital shapes (1)
The Nobel Prize in Chemistry 2008

Roger Kornberg

X-ray crystallography

How does RNA Pol II decode DNA into RNA?
Nobel Prize in Chemistry

- Green Fluorescent protein

Osamu Shimonura

Marty Chalfie
GFP

Roger Tsien

Cerebral cortex
Tsien/Chalfie
Lichtman/Sanes
Gunpowder

$$10 \text{KNO}_3 + 3 \text{S} + 8 \text{C} \rightarrow 2 \text{K}_2\text{CO}_3 + 3 \text{K}_2\text{SO}_4 + 6 \text{CO}_2 + 5 \text{N}_2.$$