The Wave Nature of Light

Electromagnetic radiation (defined on next slide) can be described as light energy and/or a wave. A wave is a vibrating disturbance by which energy is transmitted.

Waves are characterized by their length, height and by the number of waves that pass a certain point in one second.

Wavelength, $\lambda$, in meters or nanometers.

Frequency, $\nu$, in $\frac{1}{s}$ or $s^{-1}$, also Hz.

These two waves have the same amplitude; but the top wave has longer wavelength, and the lower wave has higher frequency.

Wavelength is the distance between identical points on successive waves.

Frequency is the number of waves that pass a point in one second.

Amplitude is the vertical distance from the midline of the wave to the peak (or trough).
The speed of a wave is also important, and depends on the type of wave, and the medium of travel.

\[
\text{Speed of a wave} = \text{wavelength} \times \text{frequency} = \lambda \cdot \nu
\]

UNIT CHECK \[m \cdot s^{-1} = m \cdot s^{-1}\]

There are many types of waves; here we are concerned with waves that have systematic fluctuations in intensities of electrical and magnetic components, which Maxwell in 1873 called Electromagnetic waves, of which light is an example.

The transmission (and emission) of energy in the form of electromagnetic waves is called Electromagnetic radiation.

The speed of electromagnetic waves (e.g. light), is a constant in a given medium.

The speed of light, \(c\), is approx. \(3.00 \times 10^8 \text{ m/s}\) in a vacuum.

Often use the eqn: \(c = \lambda \cdot \nu\)

The wavelength of electromagnetic radiation is often reported in nanometers (nm) which is \(10^{-9} \text{ m}\).
The Electromagnetic Spectrum

NOTICE:  Energy increases ⇐  Wavelength increases ⇒  Frequency increases ⇐
**Problem:** Calculate the wavelength, in meters, of radiation with a frequency of $1.18 \times 10^{14}$ s$^{-1}$.

What region of the electromagnetic spectrum is this?

$$\lambda = \frac{c}{\nu}$$

$$\lambda = \frac{3.00 \times 10^8 \text{ m}}{1.18 \times 10^{14} \text{ s}^{-1}} = 2.54 \times 10^{-6} \text{ m} \quad 1 \text{ nm} = 10^{-9} \text{ m}$$

$$= 2540 \text{ nm}$$

This is in the infrared region.
Particle-like Properties of Electromagnetic Radiation: The Planck Equation

Quantized Energy Levels

Pre-1900, classical physics assumed that atoms and molecules could emit (or absorb) any arbitrary amount of radiant energy – but this did not explain some experimental observations concerning radiation emission from some heated solids.

Planck in 1900 said that atoms and molecules could emit (or absorb) energy only in discrete quantities, like small packages or bundles.

Planck gave the name quantum to the smallest quantity of energy that can be emitted (or absorbed) in the form of electromagnetic radiation.

The energy $E$ of a single quantum of energy is given by

$$E = h\nu = \frac{hc}{\lambda}$$

Planck’s constant, $h = 6.626 \times 10^{-34}$ Js

According to quantum theory, energy is always emitted in integral multiples of $h\nu$; ($h\nu$, $2h\nu$, $3h\nu$ . . .), but never, for example, $1.67h\nu$ or $4.98h\nu$.

At the time Planck presented his theory, he could not explain why energies should be fixed or quantized in this manner.
**Problem:** Determine the energy emitted in the previous problem, of radiation with a frequency of $1.18 \times 10^{14} \text{ s}^{-1}$.

$$E = h\nu = (6.626 \times 10^{-34} \text{ J} \cdot \text{s}) (1.18 \times 10^{14} \text{ s}^{-1})$$

$$= 7.82 \times 10^{-20} \text{ J}$$

Can also express energies in $\text{kJ} \cdot \text{mol}^{-1}$

$$7.82 \times 10^{-20} \text{ J} \times \frac{6.022 \times 10^{23}}{\text{mol}} \times \frac{\text{kJ}}{1000 \text{ J}} = 47.1 \frac{\text{kJ}}{\text{mol}}$$
**Photoelectric effect**

Five years after Planck’s quantum theory, Einstein in 1905 described the photoelectric effect (or photoemission). This is the production (or not) of electrons when light is shone onto a material. Electrons emitted in this manner can be called photoelectrons.

\[ E_{\text{photon}} = h\nu \]

- 700 nm, 1.77 eV
- 550 nm, 2.25 eV
- 400 nm, 3.1 eV

\( v_{\text{max}} = 6.22 \times 10^5 \text{ m/s} \)

\( v_{\text{max}} = 2.96 \times 10^5 \text{ m/s} \)

Potassium - 2.0 eV needed to eject electron
According to *classical electromagnetic (wave)* theory, this effect can be attributed to the transfer of energy from the light to an electron. An alteration in either the *intensity* or *wavelength* of light would induce changes in the *rate* of *emission* of electrons from the metal. Furthermore, according to this theory, a sufficiently *dim* light would be expected to show a *time lag* between the *initial* shining of its light and the subsequent *emission* of an electron.

However, the *experimental* results did **not** correlate with either of the two predictions made by classical theory.

In reality, electrons are *only* dislodged by the impingement of photons that *reach* (or exceed) a *threshold frequency* (*energy*). Below that threshold, no electrons are emitted from the metal regardless of the light intensity or the length of time of exposure to the light.

To make sense of the *experimental fact* that light can eject electrons even if its *intensity* is *low*, Einstein proposed that a beam of *light is not a wave* propagating through space, but rather a *collection of discrete wave packets* (*photons*), each with energy \( h \nu \). This shed light on Planck’s previous discovery of the Planck relation \( E = h \nu \) linking energy \( E \) and frequency \( \nu \) as arising from quantization of energy.

The (kinetic) *energy* of the *ejected electron* is equal to the *energy in*, minus however much *energy* it takes to get the electron away from the material.

\[
KE = h \nu - \Phi
\]

Where \( \Phi \) is the *work function* of the metal (the *binding energy* of the *electron* that must be *overcome* in order to *liberate* the electron).

The main conclusion is that *light* travels in *packets* or as *particles* called *photons*, with energy \( h \nu \).

But electromagnetic radiation (*light*) is a *wave*…

\[
\Rightarrow \text{ the wave-particle duality of light.}
\]
Bohr’s model

The Rutherford–Bohr model or Bohr model, introduced by Niels Bohr in 1913, depicts the atom as a small, positively charged nucleus surrounded by electrons that travel in circular orbits around the nucleus—similar in structure to the solar system, but with attraction provided by electrostatic forces rather than gravity.

The electron (fixed) orbits can only have certain energy values (quantized).

\[ E_n = (-R_H) \frac{1}{n^2} \quad n = 1, 2, 3, 4, \ldots \]

R\textsubscript{H} is the Rydberg constant for Hydrogen = 2.18x10\(^{-18}\) J.

As written, the energy is zero when the electron is away from nucleus; and is more negative the closer the electron is to the attractive nucleus.

\[ E_\infty = (-2.18 \times 10^{-18} \text{ J}) \frac{1}{\infty^2} = 0 \]

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

So \( \nu = \frac{\Delta E}{h} \quad \text{and} \quad \lambda = \frac{hc}{\Delta E} \)
Electronic Structure of an Atom

Atoms can absorb light.

More precisely, the electrons receive the energy and it boosts them from a lower energy level to a higher energy level. The lowest energy state is called the Ground State.

Any higher level is called an Excited state.

Excited states can emit light.

As the electron drops back to the lower energy level, there is an emission of a photon.
Some electromagnetic radiation sources generate continuous (unbroken) spectra, for example sunlight, incandescent light bulb, very hot metal rod, etc.

Whereas passing electricity through a (simple) gas under vacuum produces a line (discontinuous) spectrum, for example sodium vapor, or hydrogen gas.
**Problem:** Calculate the wavelength of light emitted when an electron falls from the \( n = 3 \) to the \( n = 2 \) level in a hydrogen atom.

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

\[
= (2.18 \times 10^{-18} \text{ J}) \left( \frac{1}{3^2} - \frac{1}{2^2} \right)
\]

\[
= -3.03 \times 10^{-19} \text{ J} \quad \text{meaning energy is emitted}
\]

\[
E = h\nu = \frac{hc}{\lambda}
\]

so

\[
\lambda = \frac{hc}{\Delta E} = \frac{(6.62 \times 10^{-34} \text{ J s}) (3.00 \times 10^8 \frac{\text{m}}{\text{s}})}{3.03 \times 10^{-19} \text{ J}}
\]

\[
= 6.56 \times 10^{-7} \text{ m} = 656 \text{ nm (visible region, red)}
\]

*I am going to assume you know about PEDMAS (or BODMAS) orders of operations.*
Electromagnetic Radiation and Atomic Spectra

Line spectra, also called atomic spectra are a series of discrete lines (or wavelengths) separated by blank areas.

Line spectra are light emissions only at specific wavelengths.

Every element has a unique emission spectrum.

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The experimental fact that there are only certain lines suggests quantization. Rydberg’s equation explained these lines as transitions from one state/level/orbital to another.

Emission lines of Hydrogen

Hydrogen emits several groups of lines (wavelengths).

This experiment suggests that the energy of (an atom’s) electrons in orbitals can only take certain values (quantized).
Wave-like Properties of Matter: the de Broglie Equation

The Wave Behavior of Matter

For the next 10 years scientists struggled with this (seemingly correct) idea that electrons seemed to be restricted to orbiting a nucleus at certain fixed distances.

In 1924, de Broglie suggested that if light (waves) can behave as a stream of particles, then perhaps particles can have wave-like properties.

The de Broglie’s equation relates the mass and speed of the particle to its wavelength.

\[ E = mc^2 \quad \quad E = h\nu = \frac{hc}{\lambda} \quad \quad \text{So} \quad mc^2 = \frac{hc}{\lambda} \]

\[ \lambda = \frac{hc}{mc^2} = \frac{h}{mc} \]

The de Broglie wavelength of something with mass \( m \), and velocity \( v \):

\[ \lambda = \frac{h}{mv} \quad \quad \text{or} \quad \quad v = \frac{h}{m\lambda} \]

This is the (mathematical) connection between waves and particles.

(You get longer wavelengths for smaller masses, so this becomes increasingly important for smaller entities).
He proposed that an electron bound to a nucleus can be described as a standing wave.

For a standing wave (in a box), it must fit in the space exactly, which requires length \( l = n \frac{\lambda}{2} \) where \( n \) is an integer.

Electron orbits are circles, the circumference of an allowed orbit is \( 2\pi r \), then since the circumference of the orbit must be equal to an integral number of wavelengths:

\[
2\pi r = n\lambda.
\]
Diffraction

Shortly after de Broglie introduced his equation, in 1937 Davisson and Germer in the United States, and Thomson in England demonstrated that electrons do indeed possess wave-like properties.

By directing a beam of electrons through a thin piece of gold foil, Thomson obtained a set of concentric rings on a screen, similar to the pattern observed when X-rays (which are waves) were used.

X-ray and electron diffraction patterns

These X-ray and electron diffraction patterns from aluminum foil are similar since they both have wave properties.

One of the most important properties of waves is that they “interfere” with other waves…
Waves interfere with other waves.

Such as colored lights, or electrons in the double slit experiment.

Waves can interfere \textit{constructively} (in phase)… …or \textit{destructively} (out of phase).
**Problem:** What is the de Broglie wavelength of a 145 g baseball with a velocity of 90 mi/hr (40 m/s)?

\[
\lambda = \frac{h}{mv} = \frac{6.626 \times 10^{-34} \text{ J s}}{(145 \text{ g})(40 \text{ m/s})} \times \frac{1000 \text{ g}}{1 \text{ kg}}
\]

\[
= 1.1 \times 10^{-34} \text{ m} \quad \text{(undetectable size)}
\]

**Note:** J = \( \frac{\text{kg m}^2}{\text{s}^2} \)

**Problem:** What is the de Broglie wavelength of an electron (mass = 9.11 \( \times 10^{-31} \text{ kg} \)) traveling at 1/1000 of the speed of light (3.00 \( \times 10^8 \text{ m/s} \))?  

\[
\lambda = \frac{h}{mv} = \frac{6.626 \times 10^{-34} \text{ J s}}{(9.11 \times 10^{-31} \text{ kg}) (3.00 \times 10^5 \text{ m/s})}
\]

\[
= 2.42 \times 10^{-9} \text{ m} \quad \text{(Subatomic particles have measurable wavelengths)}
\]
Quantum Mechanics and the Heisenberg Uncertainty Principle

Heisenberg

If you describe an electron as a wave, you run into some problems, such as how do you exactly define the position of the electron (wave).

In 1927, Heisenberg described the problem by stating it is inherently impossible to know simultaneously both the exact momentum ($mv$) of the electron and its exact location ($x$) in space.

$$\Delta x \Delta mv \geq \frac{h}{4\pi}$$

Where $\Delta$ means the “uncertainty” in measuring the location, or momentum.

So the more precisely you know the location, the less precisely you know the momentum, and vice versa.
Wave Functions and Quantum Numbers

Quantum Mechanics and Atomic Orbitals

In 1926, Schrödinger formulated an equation that describes the behavior and energies of submicroscopic particles in general, an equation analogous to Newton’s laws of motion for macroscopic objects.

The Schrödinger equation incorporates both particle behavior, in terms of mass $m$, and wave behavior, in terms of a wave function $\psi$ (psi), which depends on the location in space of the system (such as an electron in an atom).

The wave function itself has no direct physical meaning. However, the probability of finding the electron in a certain region in space is proportional to the square of the wave function, $\psi^2$.

The use of probability density and electron density – likelihood of finding an electron at that location, somewhat sidesteps the problems raised in the Heisenberg Uncertainty Principle.

Schrödinger’s equation began a new era in physics and chemistry, quantum mechanics.

(Schrödinger's wave equation states when the Hamiltonian operator, $H$, acts on a certain wave function $\psi$, and the result is proportional to the same wave function $\psi$, then $\psi$ is a stationary state, and the proportionality constant, $E$, is the energy of the state $\psi$:

$$\hat{H}\psi = E\psi \quad \hat{H} = -\frac{\hbar^2}{2m} \frac{d^2}{dx^2} + V(x)$$

Where $V$ is a potential energy function).
Orbitals and Quantum Numbers

The complete solution to Schrödinger's equation for the hydrogen atom yields a set of wave equations and corresponding energies.

These wavefunctions are called orbitals.

An atomic orbital is a mathematical function that describes the wave-like behavior of an electron in an atom.

Quantum mechanics uses three quantum numbers $n$, $l$, and $m_l$, to describe an orbital.

1. **Principal quantum number**, $n$, can have integer values of 1, 2, 3 and so forth. As $n$ increases the orbital becomes larger and increases in energy.

2. **Azimuthal quantum number**, $l$, can have values from $(n–1)$ to 0, for each value of $n$. This quantum number gives the orbital its shape (the angular momentum quantum number).

<table>
<thead>
<tr>
<th>Value of $l$</th>
<th>0</th>
<th>1</th>
<th>2</th>
<th>3</th>
<th>4</th>
<th>5</th>
</tr>
</thead>
<tbody>
<tr>
<td>Letter used (Name of Orbital)</td>
<td>s</td>
<td>p</td>
<td>d</td>
<td>f</td>
<td>g</td>
<td>h</td>
</tr>
</tbody>
</table>

3. **Magnetic quantum number**, $m_l$, can have integer values between $l$ and $–l$, including zero. This quantum number describes the orientation of the orbital in space.
The collection of orbitals with the *same value* of $n$ is called an electron **shell**.

The set of orbitals with the *same* $n$ and $l$ values is called a **subshell**.

Each **subshell** is designated by a *number* (the value of $n$) and a *letter* (s, p, d, or f, corresponding to the value of $l$)

E.g. 2s subshell has $n = 2$, $l = 0$

4f subshell has $n = 4$, $l = 3$.

**(Two) Restrictions on the possible values of the quantum numbers.**

1. The shell with the principal quantum number $n$ will consist of exactly $n$ **subshells**.

   $n = 1$ has 1 subshell

   $n = 2$ has 2 subshells

   $n = 3$ has 3 subshells, etc.
2. Each subshell consists of a specific number of orbitals.

Each orbital corresponds to a different allowed value of $m_l$.

For a given value of $l$ there are $(2l + 1)$ allowed values for $m_l$, ranging from $l$ to $-l$.

- Each s ($l = 0$) subshell consists of one orbital
  \[ 2(0) + 1 = 1 \]

- Each p ($l = 1$) subshell consists of three orbitals
  \[ 2(1) + 1 = 3 \]

- Each d ($l = 2$) subshell consists of five orbitals
  \[ 2(2) + 1 = 5 \]

In summary:

**n** = principal quantum number; all orbitals with same n are in the same shell.

**l** = secondary (azimuthal) quantum number; divides shells into subshells.

**m_l** = magnetic quantum number; divides subshells into individual orbitals.
Orbitals are described by the three quantum numbers \( n \), \( l \) and \( m_l \).

**Summary of Relationships among the Quantum Numbers \( n \), \( \ell \), \( m_\ell \)**

<table>
<thead>
<tr>
<th>Value of ( n )</th>
<th>Value of ( \ell )</th>
<th>Value of ( m_\ell )</th>
<th>Subshell</th>
<th>Number of Orbitals</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0</td>
<td>0</td>
<td>1s</td>
<td>1</td>
</tr>
<tr>
<td>2</td>
<td>0</td>
<td>0</td>
<td>2s</td>
<td>1</td>
</tr>
<tr>
<td></td>
<td>1</td>
<td>−1, 0, 1</td>
<td>2p</td>
<td>3</td>
</tr>
<tr>
<td>3</td>
<td>0</td>
<td>0</td>
<td>3s</td>
<td>1</td>
</tr>
<tr>
<td></td>
<td>1</td>
<td>−1, 0, 1</td>
<td>3p</td>
<td>3</td>
</tr>
<tr>
<td></td>
<td>2</td>
<td>−2, −1, 0, 1, 2</td>
<td>3d</td>
<td>5</td>
</tr>
<tr>
<td>4</td>
<td>0</td>
<td>0</td>
<td>4s</td>
<td>1</td>
</tr>
<tr>
<td></td>
<td>1</td>
<td>−1, 0, 1</td>
<td>4p</td>
<td>3</td>
</tr>
<tr>
<td></td>
<td>2</td>
<td>−2, −1, 0, 1, 2</td>
<td>4d</td>
<td>5</td>
</tr>
<tr>
<td></td>
<td>3</td>
<td>−3, −2, −1, 0, 1, 2, 3</td>
<td>4f</td>
<td>7</td>
</tr>
</tbody>
</table>

\( l \) can have values from \((n-1)\) to 0, for each value of \( n \).

\( m_\ell \) can have integer values between \( l \) and \(-l\), including zero.
The Shapes of Orbitals: Representations of Orbitals

**The s orbitals** (one per \( n \))

- spherically symmetrical.
- probability function approaches zero as the distance from the nucleus increases.
- 2s orbital has a node (region where the probability, \( \psi^2 \), goes to zero).
- 3s orbital has two nodes.
- As \( n \) increases, the electron is more likely to be farther from the nucleus.

The size of the orbital increases as \( n \) increases.

Relative sizes of the s orbitals:

We designate a boundary of probability for the electron, and make that the shape of the orbital (e.g. 90% chance of the electron being found inside).
The p orbitals (3 per n)

- **not** spherically symmetrical
- all p orbitals have **nodes** at the *nucleus*
- on either side of the node there are **lobes**
- \( n = 2 \) has three p orbitals; all the same *size*, but one *oriented* along each axis: \( p_x, p_y, p_z \).
- p orbitals *increase* in size as \( n \) *increases*.

3 individual p orbitals

superimposed
The d orbitals (5 per n)

- four of the d ($l = 2$) orbitals have "four-leaf clover" shapes.
- they lie primarily in a plane: $d_{xy}$, $d_{xz}$, $d_{yz}$, with lobes between axes, whereas the $d_{x^2 - y^2}$ lies in the x-y plane along the axes.
- the $d_{z^2}$ orbital has two lobes along the z axis and a "doughnut" in the xy plane. It has the same energy as the other four d orbitals.

5 individual p orbitals
The f orbitals (7 per n)

The f orbitals ($l = 3$) are even more complicated than the d orbitals.
Relative Energies of Orbitals of (Multielectron) Atoms
Electron Spin and the Pauli Exclusion Principle

Electron Spin Quantum Number

How do electrons populate the available orbitals?

A new (4th) quantum number called the Electron Spin Quantum Number, $m_s$, is introduced.

It has two possible values: $+\frac{1}{2}$ and $-\frac{1}{2}$ (often referred to as spin up and spin down).

The Pauli Exclusion Principle states that no two electrons (in the same atom) can have the same set of four quantum numbers ($n, l, m_l, \text{ and } m_s$).

An orbital can hold a maximum of two electrons, and they must have opposite spins.
Diamagnetism and Paramagnetism

When two electrons in the same orbital have different spins we say they are spin paired = **diamagnetic**

- Sample not attracted to magnetic field (slight repulsion)
- Magnetic effects tend to cancel each other

When two electrons in are different orbitals with the same spin they can be spin unpaired = **paramagnetic**

- Sample attracted to a magnetic field
- *More* unpaired electrons lead to a *stronger* attraction to a magnet

We can **measure** the extent of the attraction, which provides information about the number of **unpaired** spins.
Orbital Energy Levels in Many-Electron Atoms

In *many-electron* atoms, the energy levels differ.

E.g. the 2s subshell is lower in energy than the 2p subshell.

Any electron density between the nucleus and the electron of interest will *reduce* the nuclear charge acting on that electron.

The *net positive charge* attracting the electron is called the *effective nuclear charge*, $Z_{\text{eff}}$.

$$Z_{\text{eff}} = Z - S$$

Where $Z =$ number of protons in nucleus; $S =$ the average number of electrons between the nucleus and the electron in question.
The energy of the orbital (electron) depends on the Effective Nuclear Charge and Shielding.

This figure shows the radial probability of different orbitals, even though the 2s is on average further from the nucleus than the 2p, notice the region close to the nucleus for the 2s.

The electron density closer to the nucleus is greater for the 2s than the 2p.

The 2s experiences less shielding by the 1s than the 2p does – so if 1s has electrons then 2s is lower in energy than 2p.

\[ \text{“penetrating power” } \quad 2s > 2p > d > f \]
Energies of Orbitals

In a many-electron atom, for a given value of \( n \), \( Z_{\text{eff}} \) decreases with increasing value of \( l \).

In a many-electron atom, for a given value of \( n \), the energy of an orbital increases with increasing value of \( l \).

\[ s < p < d \]

Orbitals of a subshell have the same energy.

Orbitals with the same energy are said to be degenerate.

Electron Configurations of Multielectron Atoms

The way in which the electrons are distributed among the various orbitals of an atom is called its electron configuration.

The aufbau principle (1920) says we add electrons one at a time, to the lowest available orbital, until we have added the correct number of electrons.
**Orbital diagram:** each orbital is represented by a box (or circle) and each electron is represented by a half arrow (or arrow).

Electrons having *opposite* spins are paired, when they are in the same orbital.

The most stable, or *ground state*, electron configuration of an atom is that in which the electrons are in the *lowest* possible energy states. (*Higher* energy arrangements are called *excited* states).

**Hund's rule:** (1927) for degenerate orbitals, the *lowest* energy is attained when the number of *electrons* with the *same* spin is maximized.

Orbitals of a *subshell* will fill with electrons of the *same* spin first, *then*, with electrons of *opposite* spin.
Electron Configuration of the First 10 Elements

The filling of the 2p subshell is complete at neon, which has a stable configuration with eight electrons (an octet) in the outermost shell.
Sodium has a single electron beyond the stable configuration of neon. We abbreviate the electron configuration of sodium as follows:

\[ \text{Na} = [\text{Ne}] 3s^1 \]

which is the same as \[ 1s^2 2s^2 2p^6 3s^1 \]

Since \([\text{Ne}] = 1s^2 2s^2 2p^6\)

Similarly \[ \text{Li} = [\text{He}] 2s^1 \]

*Electrons in subshells* not occupied in the nearest *noble-gas element* of *lower* atomic number are referred to as *outer-shell* electrons, or *valence* electrons.

The electrons in the *inner shells* are the *core* electrons.

All of the above applies for Periods (rows) 1, 2 and 3; for Period 4 and beyond it gets a little more involved…
All of the above applies for Periods (rows) 1, 2 and 3; for Period 4 and beyond it gets a little more involved...

Ar = 1s$^2$ 2s$^2$2p$^6$ 3s$^2$3p$^6$

So is K 1s$^2$ 2s$^2$2p$^6$ 3s$^2$3p$^6$3d$^1$

Or 1s$^2$ 2s$^2$2p$^6$ 3s$^2$3p$^6$ 4s$^1$

K has properties similar to Na. This suggests that the electron is in an s orbital.

This means that the 4s orbital is lower in energy than the 3d (see earlier diagrams #30 and #33).

This is the order in which the atomic subshells are filled.

Start with the 1s orbital and move downward following the red arrow.

So 1s, 2s, 2p, 3s, 3p, 4s, 3d, …
Electron Configurations and the Periodic Table

The periodic table is also a guide to the order in which orbitals are filled.

The classification of groups of elements in the periodic table according to the type of subshell being filled with electrons.
Notice these periodic occurrences:

All of Gp 1A are \( ns^1 \)

All of 3A are \( ns^2 \text{np}^1 \)

All of 7A are \( ns^2 \text{np}^5 \)

E.g. \( _{17}\text{Cl} = [\text{Ne}] 3s^23p^5 \)

\( _{83}\text{Bi} = [\text{Xe}] 6s^24f^{14}5d^{10}6p^3 \)

Xe = 54

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**Outer electron Configuration for all the elements**

<table>
<thead>
<tr>
<th>1</th>
<th>1A</th>
<th>2</th>
<th>2A</th>
<th>13</th>
<th>3A</th>
<th>14</th>
<th>4A</th>
<th>15</th>
<th>5A</th>
<th>16</th>
<th>6A</th>
<th>17</th>
<th>7A</th>
<th>18</th>
<th>8A</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td><strong>H</strong> 1s(^1)</td>
<td>2</td>
<td><strong>He</strong> 1s(^2)</td>
<td>3</td>
<td><strong>Li</strong> 2s(^1)</td>
<td>4</td>
<td><strong>Be</strong> 2s(^2)</td>
<td>5</td>
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<td><strong>C</strong> 2s(^2)</td>
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<td><strong>Pm</strong> 6s(^1)</td>
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</table>

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Some Anomalous Electron Configurations

These rules apply *almost* all of the time, but you should be aware of some *exceptions*.

E.g. **Cr** should be \([\text{Ar}] \ 4s^2 \ 3d^4\) but actually is \([\text{Ar}] \ 4s^1 \ 3d^5\)

\[\begin{array}{c|c}
\text{[Ar]} & \begin{array}{c}
\uparrow \\
4s^1
\end{array} \quad \begin{array}{c}
\uparrow \uparrow \uparrow \uparrow \uparrow \\
3d^5
\end{array}
\end{array}\]

Also, **Cu** should be \([\text{Ar}] \ 4s^2 \ 3d^9\) but is actually \([\text{Ar}] \ 4s^1 \ 3d^{10}\)

\[\begin{array}{c|c}
\text{[Ar]} & \begin{array}{c}
\uparrow \\
4s^1
\end{array} \quad \begin{array}{c}
\uparrow \uparrow \uparrow \uparrow \uparrow \downarrow \downarrow \downarrow \downarrow \downarrow \\
3d^{10}
\end{array}
\end{array}\]

The difference arises since there is an *enhanced stability* associated with half or fully filled subshells.

A similar effect is observed for Silver and Gold.