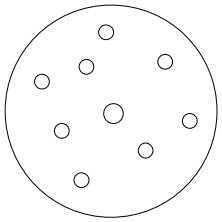
Chapter 5

Electrons in Atoms

Rutherford's Model

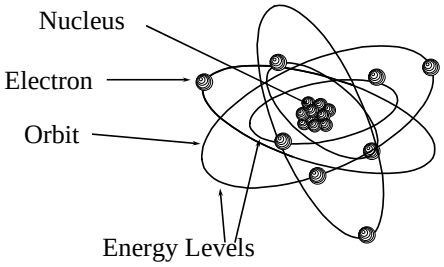
- Discovered the nucleus
- Small dense and positive
- Electrons moved around in Electron cloud



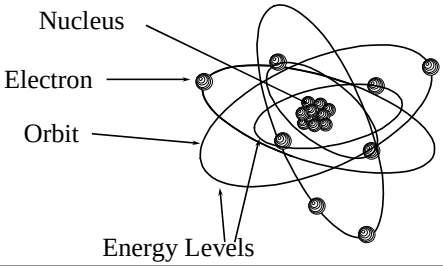
Bohr's Model

- Why don't the electrons fall into the nucleus?
- Move like planets around the sun.
- In circular orbits at different levels.
- Energy separates one level from another.

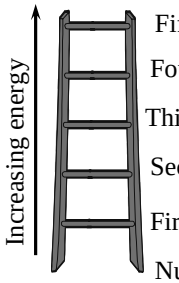
Bohr's Model



Bohr's Model



Bohr's Model



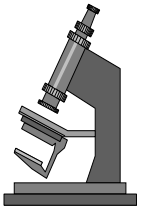
- Further away from the nucleus means more energy.
- There is no "in between" energy
- Energy Levels

The Quantum Mechanical Model

- Energy is quantized. It comes in chunks.
- Quanta - the amount of energy needed to move from one energy level to another.
- Quantum leap in energy.
- Schrödinger derived an equation that described the energy and position of the electrons in an atom
- Treated electrons as waves

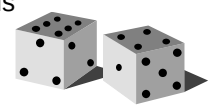
The Quantum Mechanical Model

- a mathematical solution
- It is not like anything you can see.



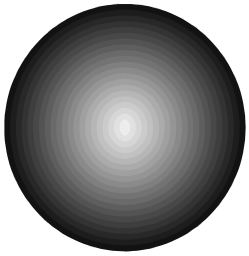
The Quantum Mechanical Model

- Does have energy levels for electrons.
- Orbits are not circular.
- It can only tell us the probability of finding an electron a certain distance from the nucleus.



The Quantum Mechanical Model

- The electron is found inside a blurry "electron cloud"
- An area where there is a chance of finding an electron.
- Draw a line at 90 %

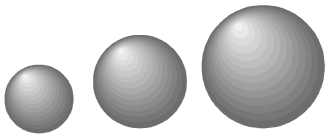


Atomic Orbitals

- Principal Quantum Number (n) = the energy level of the electron.
- Within each energy level the complex math of Schrödinger's equation describes several shapes.
- These are called atomic orbitals
- Regions where there is a high probability of finding an electron.

S orbitals

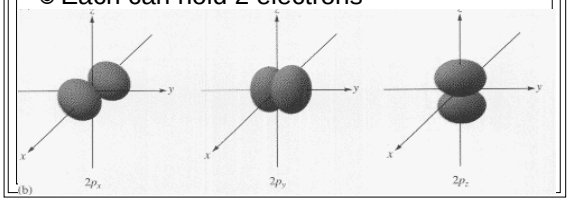
- 1 s orbital for every energy level
- Spherical shaped



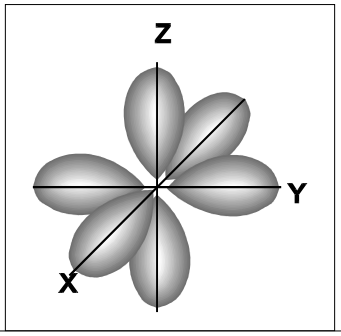
- Each s orbital can hold 2 electrons
- Called the 1s, 2s, 3s, etc.. orbitals.

P orbitals

- Start at the second energy level
- 3 different directions
- 3 different shapes (dumbell)
- Each can hold 2 electrons

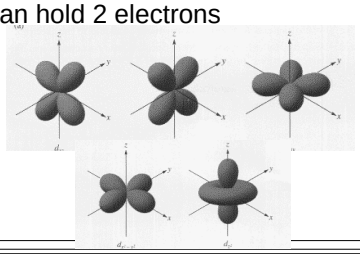


P Orbitals



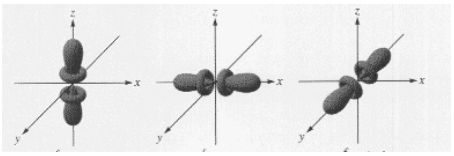
D orbitals

- Start at the third energy level
- 5 different shapes
- Each can hold 2 electrons

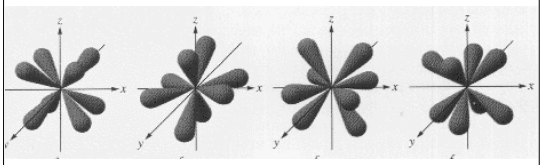


F orbitals

- Start at the fourth energy level
- Have seven different shapes
- 2 electrons per shape



F orbitals



Images

J mol

Summary

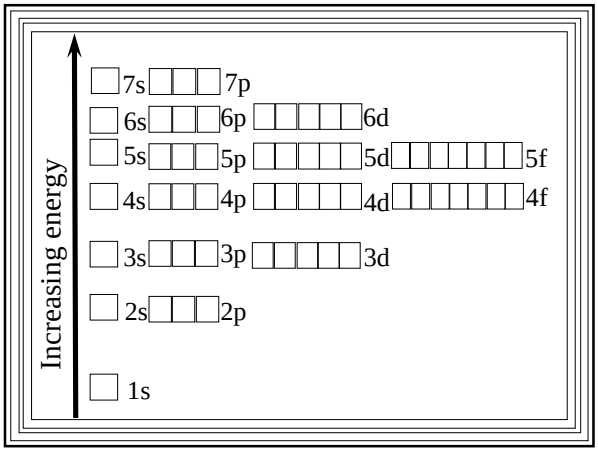
	# of shapes	Max electrons	Starts at energy level
s	1	2	1
p	3	6	2
d	5	10	3
f	7	14	4

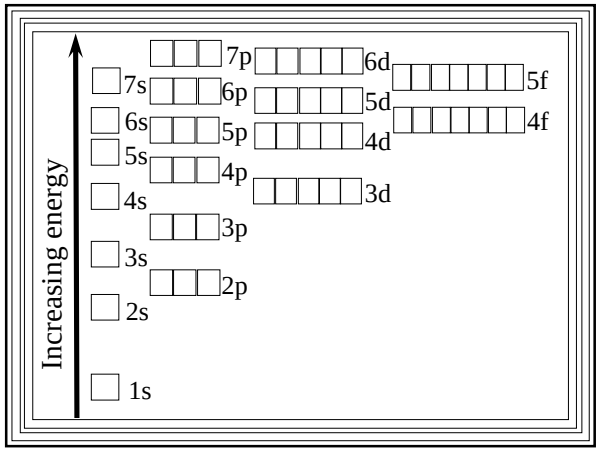
By Energy Level

- First Energy Level
- only s orbital
- only 2 electrons
- $1s^2$
- Second Energy Level
- s and p orbitals are available
- 2 in s, 6 in p
- $2s^2 2p^6$
- 8 total electrons

Filling order

- Lowest energy fill first.
- The energy levels overlap
- The orbitals do not fill up order of energy level.
- Counting system
  - Each box is an orbital shape
  - Room for two electrons



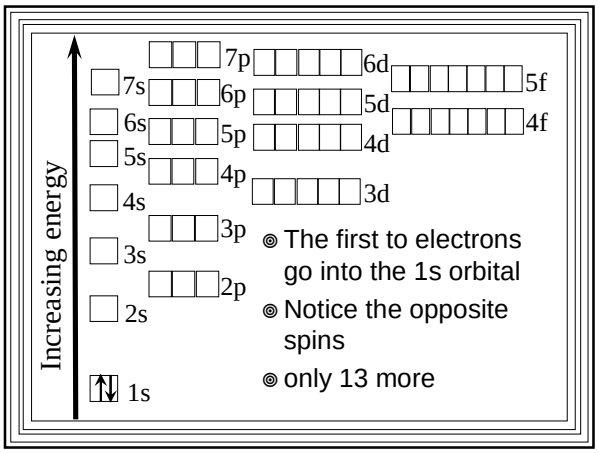


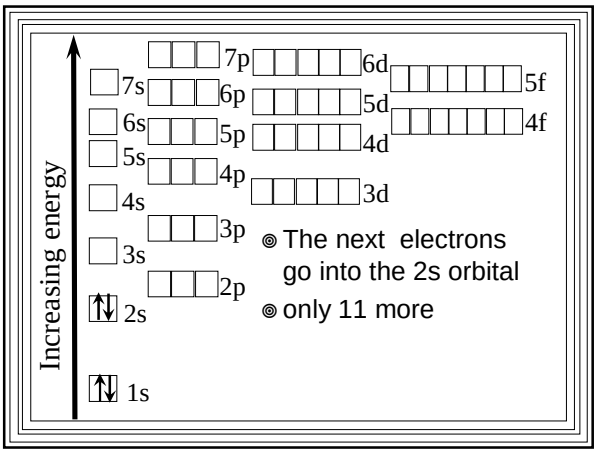
Electron Configurations

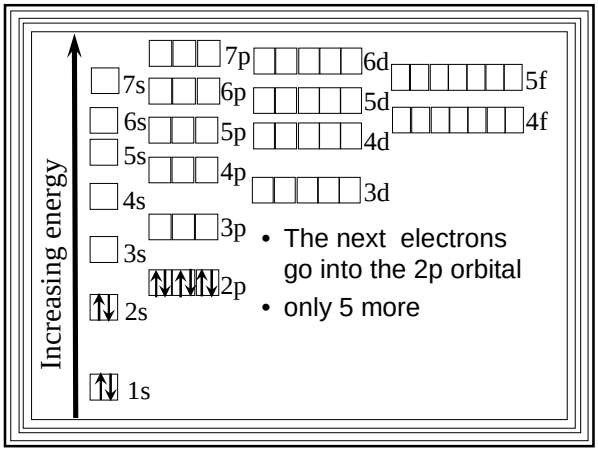
- The way electrons are arranged in atoms.
- Aufbau principle- electrons enter the lowest energy first.
- This causes difficulties because of the overlap of orbitals of different energies.
- Pauli Exclusion Principle- at most 2 electrons per orbital - different spins

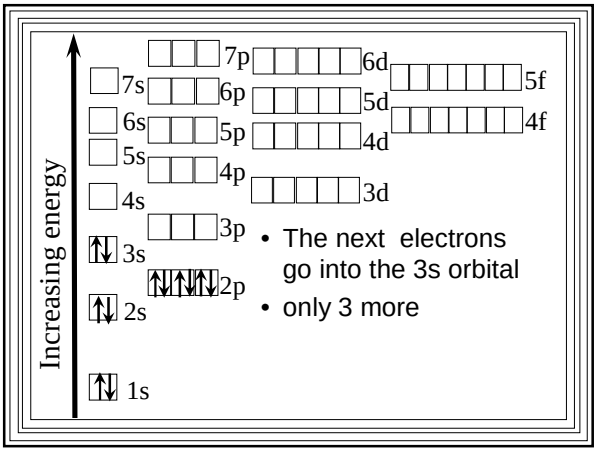
Electron Configuration

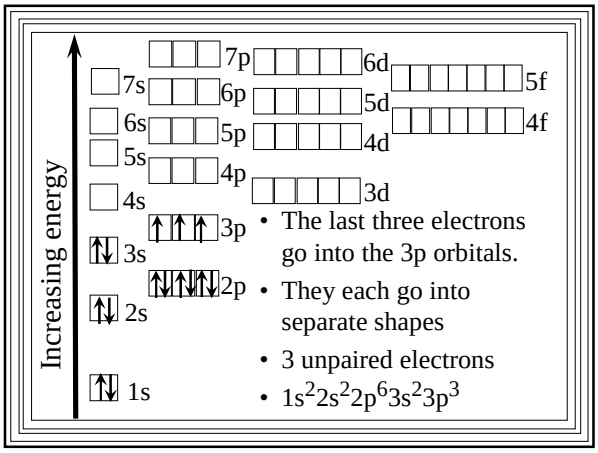
- Hund's Rule- When electrons occupy orbitals of equal energy they don't pair up until they have to .
- Let's determine the electron configuration for Phosphorus
- Need to account for 15 electrons

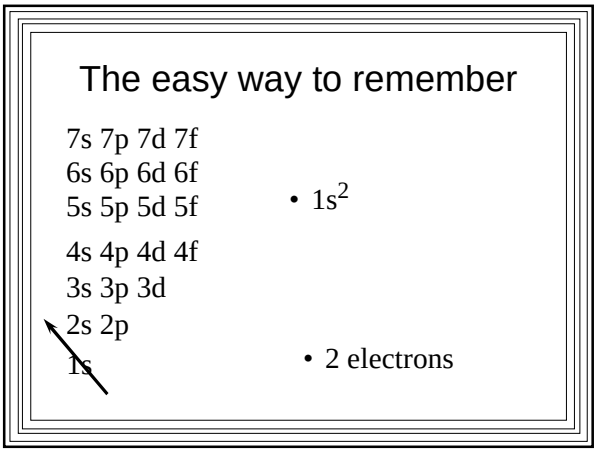

Increasing energy
7p 6d 5f
7s 6p 5d 4f
6s 5p 4d
5s 4p 3d
4s 3p
3s 2p
2s
1s
The first to electrons go into the 1s orbital
Notice the opposite spins
only 13 more


Increasing energy
7p 6d 5f
7s 6p 5d 4f
6s 5p 4d
5s 4p 3d
4s 3p
3s 2p
2s
1s
The next electrons go into the 2s orbital
only 11 more

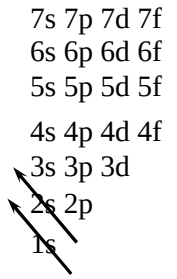

Increasing energy
7p 6d 5f
7s 6p 5d 4f
6s 5p 4d
5s 4p 3d
4s 3p
3s 2p
2s
1s
• The next electrons go into the 2p orbital
• only 5 more


Increasing energy
7p 6d 5f
7s 6p 5d 4f
6s 5p 4d
5s 4p 3d
4s 3p
3s 2p
2s
1s
• The next electrons go into the 3s orbital
• only 3 more


Increasing energy
7p 6d 5f
7s 6p 5d 4f
6s 5p 4d
5s 4p 3d
4s 3p
3s 2p
2s
1s
• The last three electrons go into the 3p orbitals.
• They each go into separate shapes
• 3 unpaired electrons
• $1s^2 2s^2 2p^6 3s^2 3p^3$

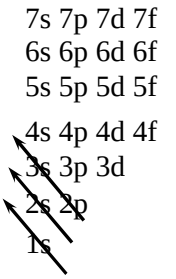

The easy way to remember
7s 7p 7d 7f
6s 6p 6d 6f
5s 5p 5d 5f
4s 4p 4d 4f
3s 3p 3d
2s 2p
1s
• $1s^2$
• 2 electrons

Fill from the bottom up following the arrows



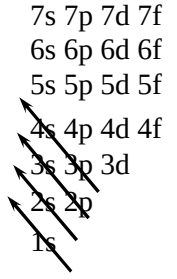
- $1s^2\ 2s^2$
- 4 electrons

Fill from the bottom up following the arrows



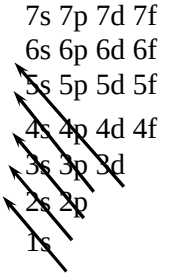
- $1s^2\ 2s^2\ 2p^6\ 3s^2$
- 12 electrons

Fill from the bottom up following the arrows



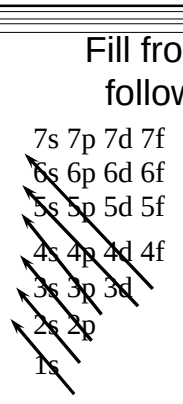
- $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$
- 20 electrons

Fill from the bottom up following the arrows



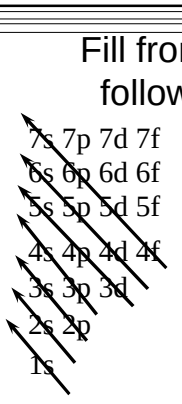
- $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2$
- 38 electrons

Fill from the bottom up following the arrows



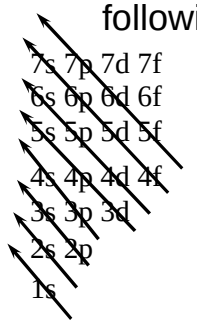
- $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2$
- 56 electrons

Fill from the bottom up following the arrows



- $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^{10}\ 6p^6\ 7s^2$
- 88 electrons

Fill from the bottom up following the arrows



- $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^{10}\ 6p^6\ 7s^2\ 5f^{14}\ 6d^{10}\ 7p^6$
- 118 electrons

Rewrite when done

- $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^{10}\ 6p^6\ 7s^2\ 5f^{14}\ 6d^{10}\ 7p^6$
  - Group the energy levels together
- $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ 4d^{10}\ 4f^{14}\ 5s^2\ 5p^6\ 5d^{10}5f^{14}6s^2\ 6p^6\ 6d^{10}\ 7s^2\ 7p^6$

Exceptions to Electron Configuration

Orbitals fill in order

- Lowest energy to higher energy.
- Adding electrons can change the energy of the orbital.
- Filled and half-filled orbitals have a lower energy.
- Makes them more stable.
- Changes the filling order of d orbitals

Write these electron configurations

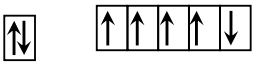
- Titanium - 22 electrons
- $1s^2 2s^2 2p^6 3s^2 3p^6 3d^2 4s^2$
- Vanadium - 23 electrons
  $1s^2 2s^2 2p^6 3s^2 3p^6 3d^3 4s^2$
- Chromium - 24 electrons
- $1s^2 2s^2 2p^6 3s^2 3p^6 3d^4 4s^2$ is expected
- But this is wrong!!

Chromium is actually

- $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$
- Why?
- This gives us two half filled orbitals.

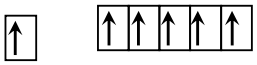
Chromium is actually

- $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$
- Why?
- This gives us two half filled orbitals.



Chromium is actually

- $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$
- Why?
- This gives us two half filled orbitals.



- Slightly lower in energy.
- The same principle applies to copper.

Copper's electron configuration

- Copper has 29 electrons so we expect
- $1s^2 2s^2 2p^6 3s^2 3p^6 3d^9 4s^2$
- But the actual configuration is
- $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$
- This gives one filled orbital and one half filled orbital.
- Remember these exceptions
  - $d^4 s^2 \rightarrow d^5 s^1$
  - $d^9 s^2 \rightarrow d^{10} s^1$

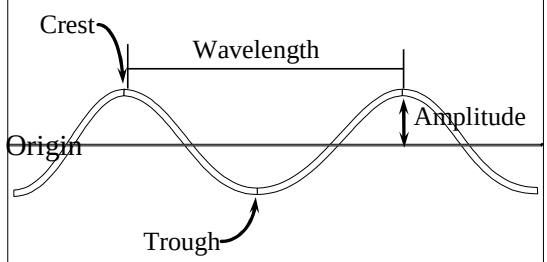
In each energy level

- The number of electrons that can fit in each energy level is calculated with
- Max $e^- = 2n^2$ where n is energy level
- 1^{st}
- 2^{nd}
- 3^{rd}

Light

- The study of light led to the development of the quantum mechanical model.
- Light is a kind of electromagnetic radiation.
- Electromagnetic radiation includes many kinds of waves
- All move at 3.00 x 10^8 m/s (c)

Parts of a wave



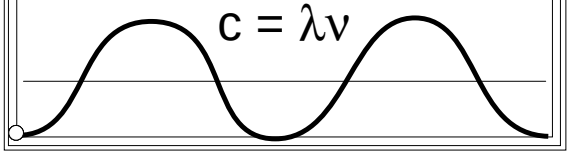
Parts of Wave

- Origin - the base line of the energy.
- Crest - high point on a wave
- Trough - Low point on a wave
- Amplitude - distance from origin to crest
- Wavelength - distance from crest to crest
- Wavelength - is abbreviated λ –Greek letter lambda.

Frequency

- The number of waves that pass a given point per second.
- Units are cycles/sec or hertz (Hz)
- Abbreviated ν – the Greek letter nu

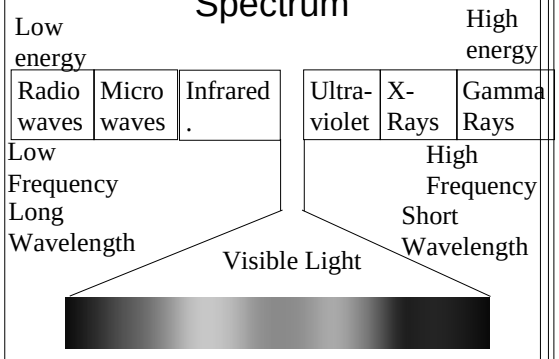
$c = \lambda\nu$



Frequency and wavelength

- Are inversely related
- As one goes up the other goes down.
- Different frequencies of light is different colors of light.
- There is a wide variety of frequencies
- The whole range is called a spectrum

Spectrum



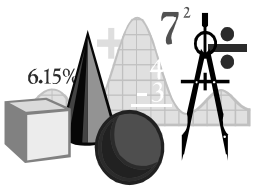
Light is a Particle

- Energy is quantized.
- Light is energy
- Light must be quantized
- These smallest pieces of light are called photons.
- Energy and frequency are directly related.

Energy and frequency

- $E = h \times \nu$
- E is the energy of the photon
- ν is the frequency
- h is Planck's constant
- $h = 6.626 \times 10^{-34}$ Joules sec.

The Math in Chapter 5



- Only 2 equations
- $c = \lambda\nu$
- $E = h\nu$
- c is always 3.00×10^8 m/s
- h is always 6.626×10^{-34} J s

Examples

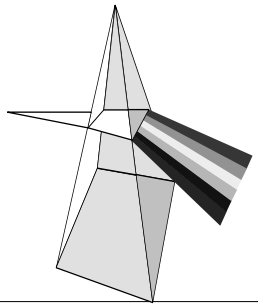
- What is the frequency of red light with a wavelength of 4.2×10^{-5} cm?
- What is the wavelength of KFI, which broadcasts at with a frequency of 640 kHz?
- What is the energy of a photon of each of the above?

Atomic Spectrum

How color tells us about atoms

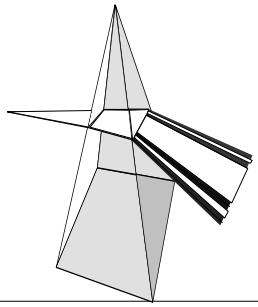
Prism

- White light is made up of all the colors of the visible spectrum.
- Passing it through a prism separates it.



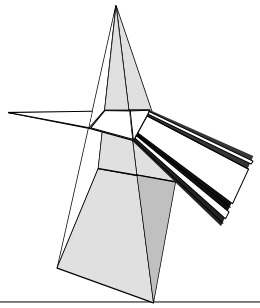
If the light is not white

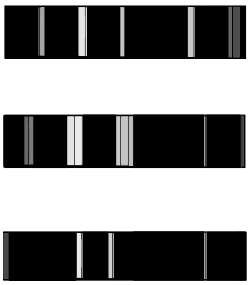
- By heating a gas or with electricity we can get it to give off colors.
- Passing this light through a prism does something different.



Atomic Spectrum

- Each element gives off its own characteristic colors.
- Can be used to identify the atom.
- How we know what stars are made of.





- These are called line spectra
- unique to each element.
- These are emission spectra
- Mirror images are absorption spectra
- Light with black missing

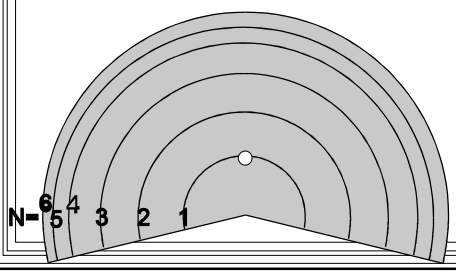
An explanation of Atomic Spectra

Where the electron starts

- When we write electron configurations we are writing the lowest energy.
- The energy level an electron starts from is called its ground state.

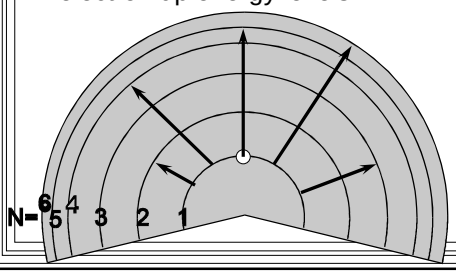
Changing the energy

- Let's look at a hydrogen atom



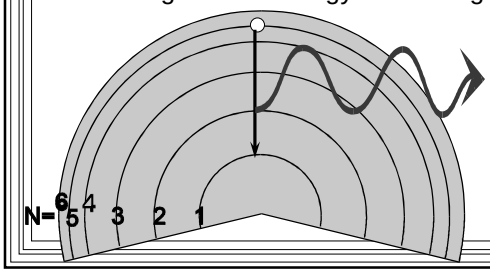
Changing the energy

- Heat or electricity or light can move the electron up energy levels



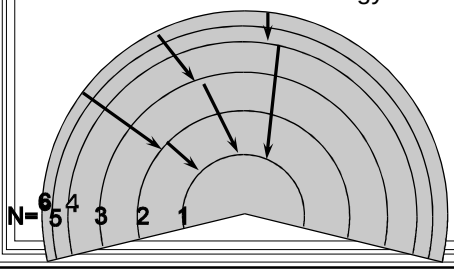
Changing the energy

- As the electron falls back to ground state it gives the energy back as light

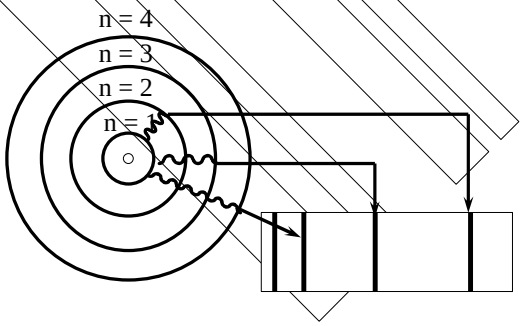


Changing the energy

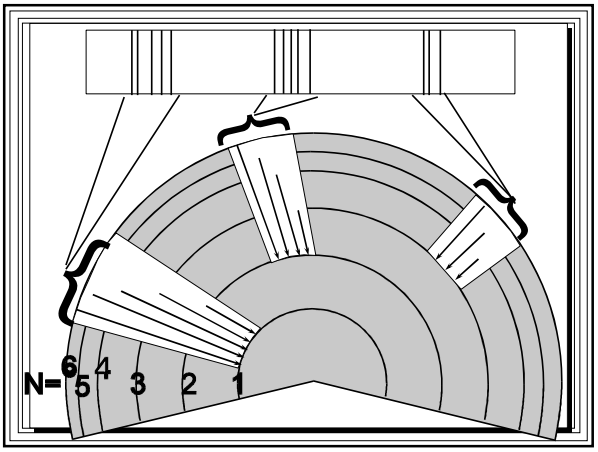
- May fall down in steps
- Each with a different energy

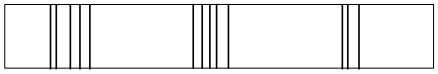


The Bohr Ring Atom



69





Ultraviolet Visible Infrared

- Further they fall, more energy, higher frequency.
- This is simplified
- the orbitals also have different energies inside energy levels
- All the electrons can move around.

What is light?

- Light is a particle - it comes in chunks.
- Light is a wave- we can measure its wave length and it behaves as a wave
- If we combine $E=mc^2$, $c=\lambda\nu$, $E = 1/2\ mv^2$ and $E = h\nu$
- We can get $\lambda = h/mv$
- The wavelength of a particle.

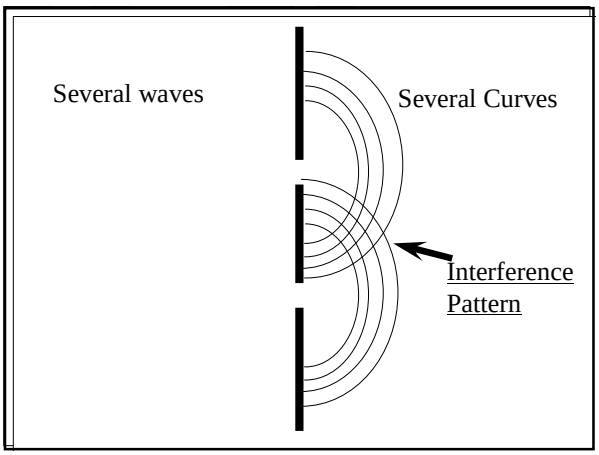
Matter is a Wave

- Does not apply to large objects
- Things bigger than an atom
- A baseball has a wavelength of about 10^{-32} m when moving 30 m/s
- An electron at the same speed has a wavelength of 10^{-3} cm
- Big enough to measure.

Diffraction

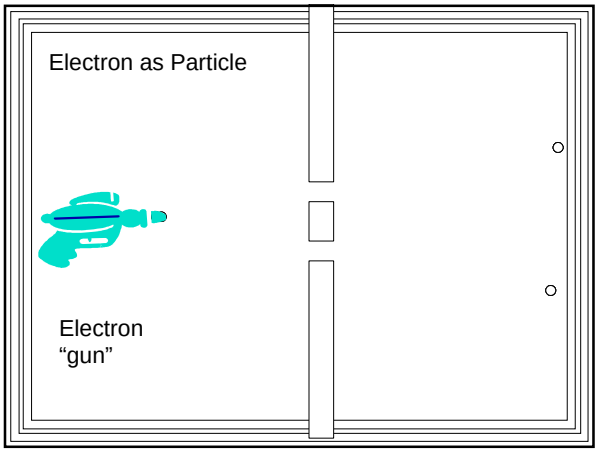
- **When light passes through, or reflects off, a series of thinly spaced lines, it creates a rainbow effect**
- **because the waves interfere with each other.**

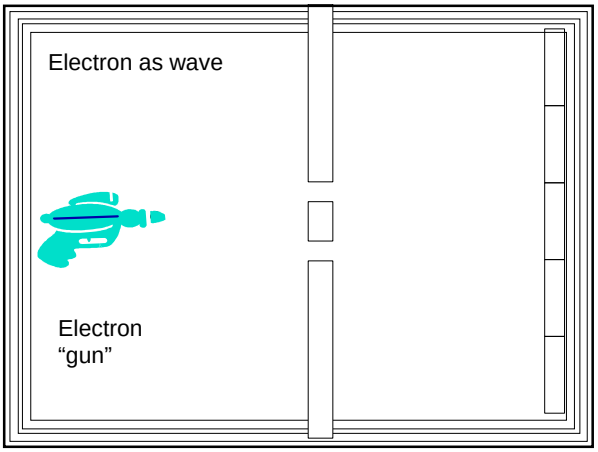
74



Diffraction

- Light shows interference patterns
- Light is a wave
- What will an electron do when going through two slits?
  - Go through one slit or the other and make two spots
  - Go through both and make a interference pattern





Which did it do?

- It made the diffraction pattern
- The electron is a wave
- Led to Schrödingers equation

The physics of the very small

- Quantum mechanics explains how the very small behaves.
- Quantum mechanics is based on probability because

Heisenberg Uncertainty Principle

- It is impossible to know exactly the speed and position of a particle.
- The better we know one, the less we know the other.
- The act of measuring changes the properties.

More obvious with the very small

- To measure where a electron is, we use light.
- But the light moves the electron
- And hitting the electron changes the frequency of the light.