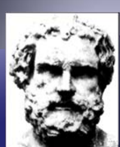
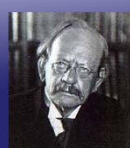
ATOMIC THEORY

Anyone who expects a source of power from the transformation of the atom is talking moonshine.

-- Ernest Rutherford
British Scientist

Who are these men?

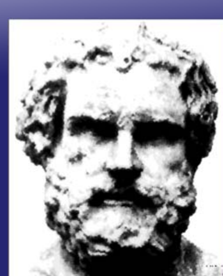




Their quest for knowledge about the fundamental nature of the universe helped define our modern view of the atom.

Democritus (400 BC)

- Greek philosopher who began the search for a description of matter more than <u>2400</u> years ago.
- He asked: "Could matter be divided into smaller and smaller pieces forever, or was there a <u>limit</u> to the number of times a piece of matter could be <u>divided</u>?"

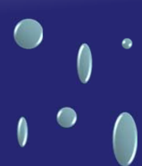


Democritus (400 BC)

- His theory: Matter could NOT be divided forever into smaller and smaller pieces, but eventually a smallest possible piece would be obtained.
- He called this smallest, indivisible piece of matter "atomos," meaning "not to be cut".

Democritus (400 BC)

- To Democritus, atoms were small, hard particles that were all made of the same material but were different shapes and sizes.
- Atoms were infinite in number, always moving and capable of joining together.



Democritus (400 BC)

- This theory was *ignored and forgotten* for more than *2000* years!

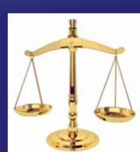
- WHY?

Democritus (400 BC)

- The eminent philosophers of the time, Aristotle and Plato, had a more respected (*yet ultimately wrong*) theory.
- Remember This?

Important Discovery (1789)

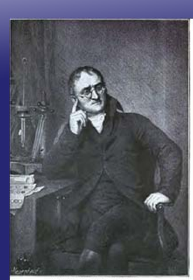
- Conservation of Mass (Antoine Lavoisier)
  - The amount of mass before a chemical reaction is the same as after (*but in different forms*)



- Lavoisier also defined an element as a basic substance that could not be further broken down by chemical processes

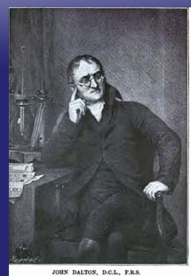
Dalton's Model (1803)

- In the early 1800's, the English Chemist John Dalton performed a number of experiments that eventually led to the acceptance of the idea of atoms.



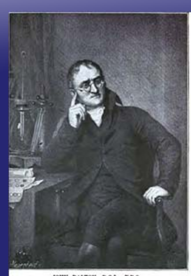
Dalton's Model (1803)

- All elements are composed of atoms (*indivisible and indestructible particles*).
- Atoms of the same element are **exactly** alike (*size, shape, volume…*).
- Atoms of different elements are different.

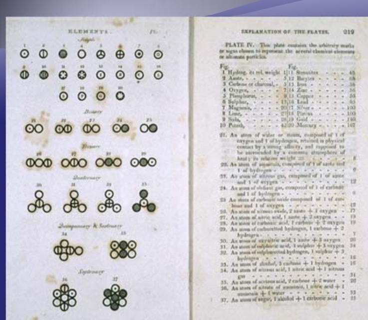


Dalton's Model (1803)

- Compounds are formed by the joining of *two or more* atoms of the same or different elements.
- Elements in compounds always combine in ratios of small whole numbers (*Law of Multiple Proportions*)



Dalton's Model (1803)

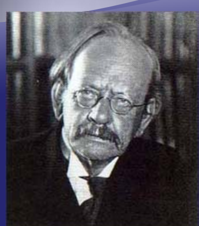


Various Atoms and molecules depicted by John Dalton

Dalton's Model (1803)

- Dalton's Model became **THE** model for atomic theory for the next 100 years and serves as the foundation for our modern view of the atom.
- Still needed some corrections and improvements.

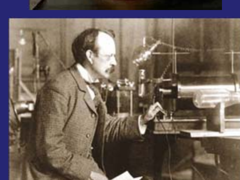
Thomson's Plum Pudding Model (1897)



- J.J. Thomson provided the first hint that an atom is made of even **smaller** particles.

Thomson's Plum Pudding Model (1904)

- In 1897, Thomson studied the passage of an electric current through a gas.
- As the current passed through the gas, the gas gave off rays of negatively charged particles that he "corpuscles", today known as electrons.



Thomson's Plum Pudding Model (1904)

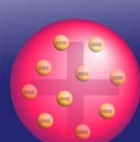
- This was a surprise, because the atoms of a gas were known to be uncharged (*experiments*) and thought to be solid, like a billiard ball (Dalton's Model).
- So where did the negative charges come from?

Thomson's Plum Pudding Model (1904)

- Thomson concluded that the negative charges had to came from *within* the 'solid' atom.
- This meant that a particle smaller than an atom **had to exist**.
- Thus, the atom was divisible! (*Contrary to the claim of Dalton's Model*)
- Also, since the initial gas was known to be neutral, he reasoned that there must be equal amount of positive charge within the atom.

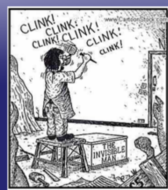
Thomson's Plum Pudding Model (1904)

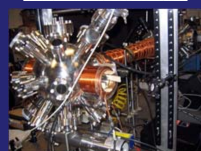
- In 1904, Thompson proposed that atoms consisted of a uniform, positively charged sphere with negatively charged electrons scattered throughout, like raisins in a pudding.
- This view of the atom lead his contemporaries to refer to his model as the "Plum Pudding" Model.



Thomson's Plum Pudding Model (1904)

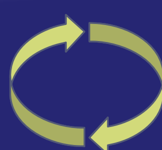
- Thompson's model only lasted for about 7 years before being replaced by another, more accurate model.
- WHY?
- Technology was advancing to the point where much more accurate and detailed experiments could be made.





General Rule of Scientific Advancement

- Better technology, better data
- Better data, better models
- Better models, better understanding
- Better understanding, better experiments
- Better experiments require better technology.

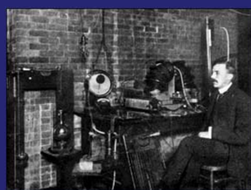


Rutherford's Model (1911)

- In 1909, the English physicist Ernest Rutherford tried to validate Thompson's Plum Pudding model.

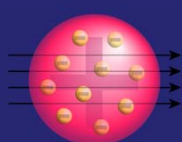
Rutherford's Model (1911)

- Rutherford's experiment involved firing a stream of tiny positively charged particles at a thin sheet of gold foil (2000 atoms thick)



Rutherford's Model (1911)

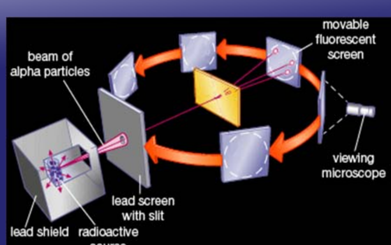
- If the Plum Pudding model was correct, the charged particles would pass through the atom undisturbed or only slightly deflected.



Expected Result

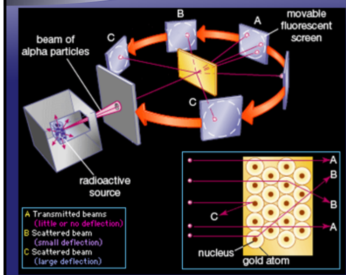
Rutherford's Model (1911)

- Experimental Setup & Expected Result



Rutherford's Model (1911)

- Experimental Results

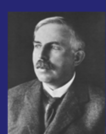


What he found was that MOST of the charged particles would pass through the atom undisturbed or only slightly deflected, but there were some that were deflected at VERY LARGE angles.

Rutherford's Model (1911)

- Rutherford's comment regarding backward scattered charges:

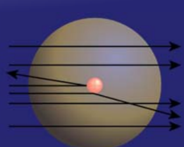
"It was quite the most incredible event that has ever happened to me in my life. It was almost as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you. On consideration, I realized that this scattering backward must be the result of a single collision, and when I made calculations I saw that it was impossible to get anything of that order of magnitude unless you took a system in which the greater part of the mass of the atom was concentrated in a minute nucleus. It was then that I had the idea of an atom with a minute massive center, carrying a charge."



Rutherford's Model (1911)

Rutherford's Conclusions:

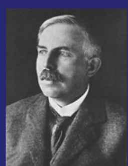
- Gold atoms had to be mostly empty space, NOT a pudding filled with a positively charged material (*since most of the charged particles passed through un-deflected*).
- Atoms had to have a small, dense, positively charged center (*Only positive charge could repel another positive charge AND large scattering angles*)

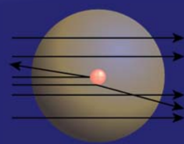


Rutherford's Model (1911)

Rutherford's Conclusions:

- He called the center of the atom the "nucleus"
- The nucleus is tiny compared to the atom as a whole.

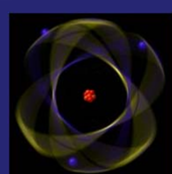




Rutherford's Model (1911)

In 1911, Rutherford presented his own model of the atom.

- The atom is mostly empty space
- Most of the atom's mass is concentrated in a tiny center, the nucleus.
- Electrons are held in orbit around the nucleus by electrostatic attraction.
- The nucleus was around 1/5000 the diameter of an atom.



Additional Comments

- In 1919, addition experiments lead Rutherford to the discovery of the proton.
  - An **integer number** of protons exist in the nucleus (*particles*)
- In 1932, James Chadwick discovered the neutron, which also resides in the nucleus.
  - A neutron has a mass nearly identical to a proton but no charge

Not to scale! →

Comment

- At this point, it would seem that the nature of the atom had been revealed.
  - Time to put our understanding/model to the test!
- How?
  - By using it to predict observed experimental data acquired years earlier.

Atomic Spectra (1855-1885)

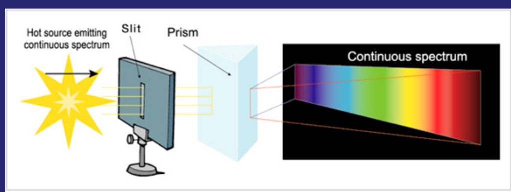
- Scientists discovered that specific colors of light were emitted from the sparks of heated metals
- Each metal had its own unique set of colors
- Later found that **EVERY** element had its own unique color pattern

Atomic Spectra (1855-1885)

- The study and measurement of these colors emitted from elements is known as spectroscopy
- Two Methods
  - Emission Spectrum
  - Absorption Spectrum

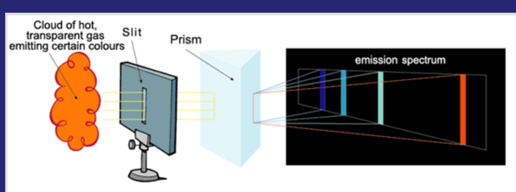
Atomic Spectra (1855-1885)

- Full or Continuous Spectra
  - White Light can be separated out into its component colors by passing it through a prism



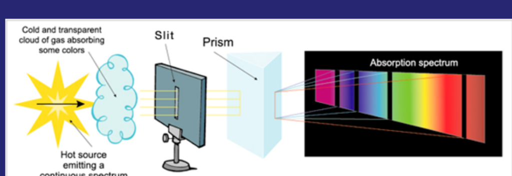
Atomic Spectra (1855-1885)

- Emission Spectra
  - When light from a heated element or gas is passed through a prism, only certain colors are observed. These specific sets of colors are called spectral lines.



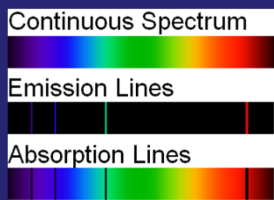
Atomic Spectra (1855-1885)

- Absorption Spectra
  - When white light traveling through a cold gas is passed through a prism, certain colors are absent. These specific sets of absent colors are called absorption lines.



Atomic Spectra (1855-1885)

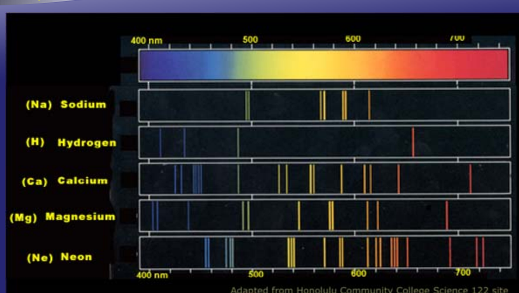
- Spectral and Absorption Lines
  - When using the same element, the spectral lines AND absorption lines occur at the SAME location.



Hydrogen

Atomic Spectra (1855-1885)

- Spectral Lines for Various Elements



Atomic Spectra (1855-1885)

- Question?
  - How are these spectral lines being produced by atoms?
- This question could not be answered using Dalton's solid atom atomic model.
- Best solutions at the time were 'empirical models'
  - Models based solely on fitting data

Atomic Spectra (1855-1885)

- For Hydrogen
  - The empirical model (*Rydberg Formula*) looks like this:

$$\frac{1}{\lambda} = R_H\left(\frac{1}{2^2} - \frac{1}{n^2}\right) \qquad n = 3, 4, 5 \ldots$$

Rydberg Constant

$R_H = 1.09737 \, x \, 10^{-2} \, nm^{-1}$

Atomic Spectra (1855-1885)

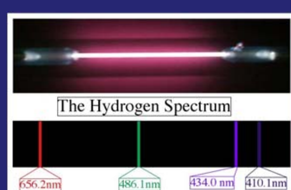
Example

Calculate the wavelength of light emitted from Hydrogen for n = 4 using the Rydberg formula.

$$\frac{1}{\lambda} = R_H\left(\frac{1}{2^2} - \frac{1}{n^2}\right)$$

Rydberg Constant

$R_H = 1.09737 \, x \, 10^{-2} \, nm^{-1}$

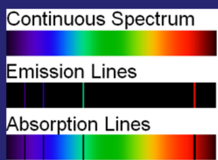


Atomic Spectra (1855-1885)

- Good, but….
  - While the empirical models worked, they provided NO understanding or knowledge about the atom OR how the light was produced
  - For calculation purposes only
  - Need an atomic model that will provide same results as empirical models

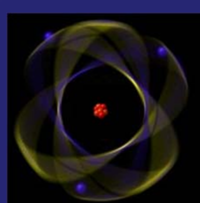
Bohr Model (1913)

- In 1913, the Danish scientist Niels Bohr proposed a model, based largely on Rutherford's conclusions, to explain the emission lines of Hydrogen.



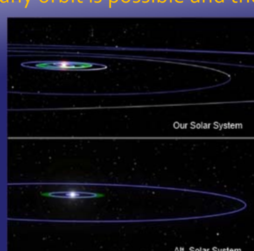
Bohr Model (1913)

- According to the Rutherford view of atomic structure, an electron (e-) traveled about the nucleus in some sort of an orbit.



Bohr Model (1913)

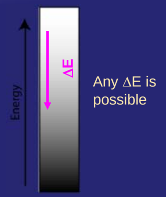
- Classically, any orbit is possible and thus, any energy.



- Planets can be **ANY** distance from the sun if they have the right amount of orbital energy.

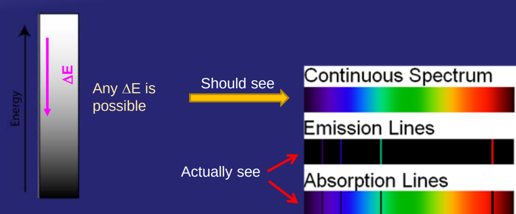
Bohr Model (1913)

- If any orbit or energy is possible, then the number of possibilities for the change in energy (ΔE) between any two orbits is infinite.



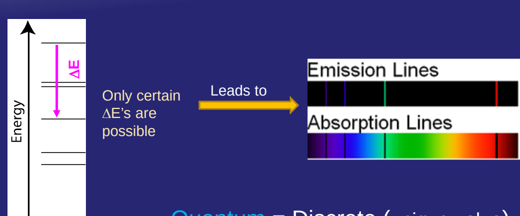
Bohr Model (1913)

- In terms of atoms and light emission, if any orbit or energy is possible, then the infinite number of changes in energy (ΔE) would result in all colors being emitted.



Bohr Model (1913)

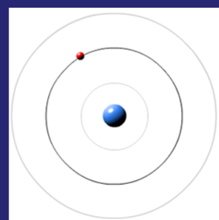
- Because the observed spectral lines are discrete colors, that must mean only certain orbits or changes in energy are allowed.



Quantum = Discrete (unique value)

Bohr Model (1913)

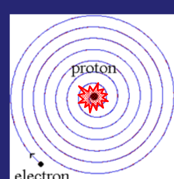
- Bohr Assumption #1:
  - The electrons orbited the nucleus in circular orbits.



  - BIG PROBLEM!!!
    Already?

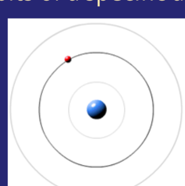
Bohr Model (1913)

- Recall:
  - Electrons moving in a circular orbits are accelerating (*changing direction*)
  - Charges that are accelerating emit EM radiation (*energy*)
- Conclusion:
  - Electrons should continually emit energy and spiral into the nucleus, but they don't !?!?



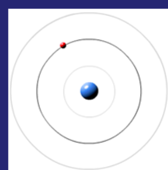
Bohr Model (1913)

- So how does Bohr overcome this issue?
- Bohr Assumption #2:
  - Electrons can only populate certain **STABLE** circular orbits of a specific amount of energy.



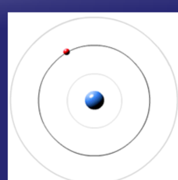
Bohr Model (1913)

- These stable circular orbits are called energy levels, since each requires a unique amount of energy.
- These stable orbits are located at specific distances from the nucleus.
- The smallest possible circular orbit or lowest energy level is called the ground state. All larger/higher levels are called excited states.



Bohr Model (1913)

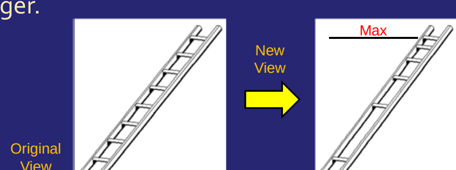
- These energy levels can be viewed much like rungs in a ladder.



- It takes a certain step size to go up or down the ladder.
  - Same idea is assumed for electrons in an atom!

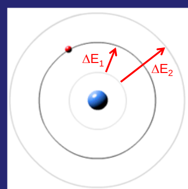
Bohr Model (1913)

- The next assumption was necessary for Bohr to explain the ejection of electrons from atoms.
- Bohr Assumption #3:
  - Energy levels get closer together as they get larger.



Bohr Model (1913)

- Bohr Assumption #4:
  - Electrons can transition (move) up or down energy levels by absorbing or emitting the exact energy difference (ΔE) between two levels *only*!



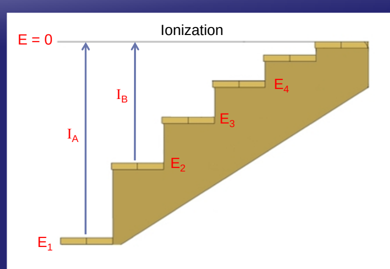
$$|\Delta E_2| > |\Delta E_1|$$

Bohr Model (1913)

- Ionization – when an electron gets ejected from an atom

I_n is the amount of energy required to ionize an electron from energy level E_n

$$I_n = |E_n|$$



Bohr Model (1913)

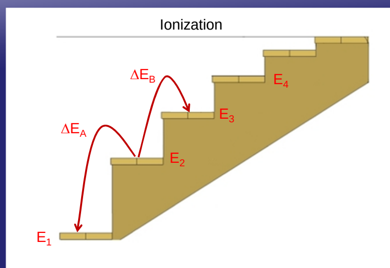
- In general, the energy difference between energy levels is:

$$\Delta E = E_f - E_i$$

  - If ΔE > 0, energy has been absorbed by the electron
  - If ΔE < 0, energy has been emitted by the electron

Bohr Model (1913)

- Visualization



$$\Delta E = E_f - E_i$$

$$\Delta E_A = E_1 - E_2$$
$$\Delta E_A < 0$$

$$\Delta E_B = E_3 - E_2$$
$$\Delta E_B > 0$$

Bohr Model (1913)

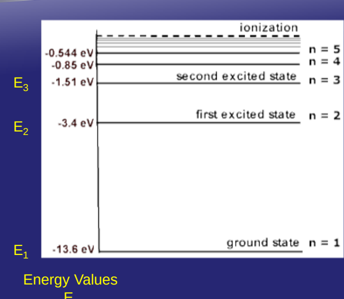
- All the pieces are almost in place for the new model.
- Just need a way to determine E_n for each energy level.
- For Hydrogen, Bohr found that

$$E_n = \frac{-13.6\ eV}{n^2}$$

- n represents the energy level, with $n = 1$ is the ground state, $n = 2$ is the 1st excited state…

Bohr Model (1913)

- This predicts energy levels for hydrogen of …



Bohr Model (1913)

- This is impressive and all, but does it reproduce the spectral lines of hydrogen?
- Recall: The energy of a photon of light is:

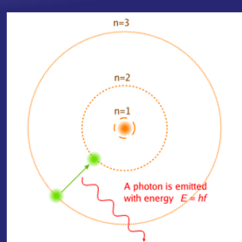
$$E = hf \quad or \quad E = \frac{hc}{\lambda}$$

- Recall: The amount of energy emitted by an electron when it changes energy levels is:

$$\Delta E = E_f - E_i$$

Bohr Model (1913)

- Bohr said that **the energy of the emitted photon is equal to the change in energy level of the electron** as it moves from a high energy level to a lower one.



Bohr Model (1913)

- Equating the two previous expressions yields:

$$f = \frac{1}{h}|\Delta E| \quad or \quad \frac{1}{\lambda} = \frac{|\Delta E|}{hc}$$

- Inserting Bohr's value for E_n:

$$f = \frac{13.6\,eV}{h}\left|\frac{1}{n_f^2} - \frac{1}{n_i^2}\right| \quad or \quad \frac{1}{\lambda} = \frac{13.6\,eV}{hc}\left|\frac{1}{n_f^2} - \frac{1}{n_i^2}\right|$$

- How does this compare to empirical models?

Bohr Model (1913)

- Comparing Bohr's model to the Rydberg Formula:

$$\frac{1}{\lambda} = \frac{13.6\,eV}{hc}\left|\frac{1}{n_f^{\,2}} - \frac{1}{n_i^{\,2}}\right|$$ Bohr

$$\frac{1}{\lambda} = R_H\left(\frac{1}{2^2} - \frac{1}{n^2}\right)$$ Rydberg

- Recall

Rydberg Constant

$R_H = 1.09737\ x\,10^{-2}\ nm^{-1}$ But… $\frac{13.6\,eV}{hc} = 1.09737\ x\,10^{-2}\ nm^{-1}$

Identical if $n_f = 2$

Bohr Model (1913)

- Bohr's model matches perfectly with the empirical model of Rydberg!
- Knowing what assumptions went into Bohr's model, we have a better understanding of HOW light is produced by an atom.

Nobel Prize

Bohr Model (1913)

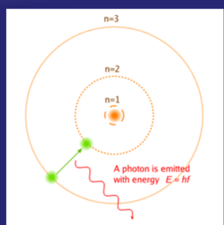
Bohr's Conclusions:

- Electrons move in stable circular orbits called energy levels about the nucleus
- Electrons move up or down energy levels by emitting or absorbing a specific amount energy
- The energy of the emitted photon **is equal to** the change in energy level of the electron as it moves from a high energy level to a lower one

Bohr Model (1913)

Example

An electron jumps from $n = 3$ to $n = 2$ in a hydrogen atom. What is the wavelength of the emitted light?



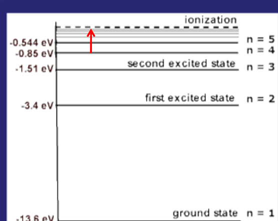
$$\frac{1}{\lambda} = R_H \left| \frac{1}{n_f^2} - \frac{1}{n_i^2} \right|$$

$$R_H = 1.09737 \, x \, 10^{-2} \, nm^{-1}$$

Bohr Model (1913)

Example

An electron is in the $n = 4$ energy level in a hydrogen atom. How much energy is required to ionize the electron?



$$I_n = |E_n|$$

$$I_n = \left| \frac{-13.6 \, eV}{n^2} \right|$$

Where Is This Knowledge Useful?

- Astronomy/Cosmology
  - Red-Shift
  - Blue-Shift
- Lasers

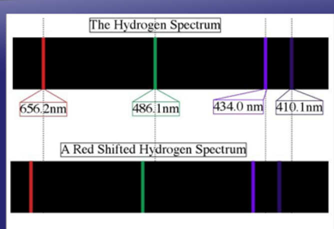
Astronomy/Cosmology

Recall: Doppler Effect

- Wavelengths of emitted or received sound change because of their motion
- The same principle applies to light if viewed as a wave
- By determining the magnitude of the shift, the relative velocity can be determined

Astronomy/Cosmology

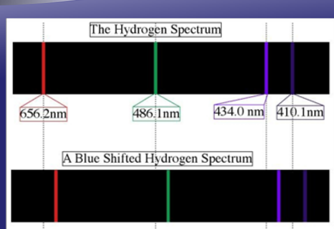
Red-Shift



Red-Shifted objects are moving AWAY from us!

Astronomy/Cosmology

Blue-Shift



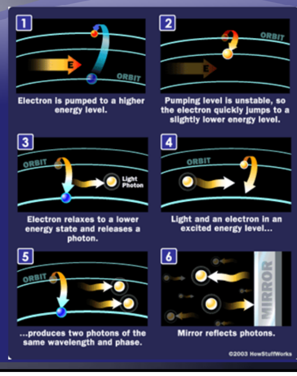
Blue-Shifted objects are moving TOWARD from us!

Lasers

Light
Amplification by the
Stimulated
Emission of
Radiation

Contains only ONE color

White light – ALL colors



Lasers

Red (650 nm)

Orange (593.5 nm)

Green (532 nm)

Blue (473 nm)

Purple (405 nm)

DEMO

Comments

- As good as the Bohr model was, it did NOT correctly predict/match the data for the rest of the elements
- It did tell us something important about nature on very small scales:
  It is NOT continuous, but discrete
- What about the rest of the elements? What model do we use for them?

Quantum Model (1920's)

- The modern Quantum Model is based on three key principles:
  - Uncertainty Principle
  - Electron has Wave Properties
  - Probabilities

Quantum Model (1920's)

- Some aspects of the Bohr model can still be used because they worked:
  - discrete energy levels
  - photon = change in energy between levels
  - changing levels require specific energy amounts

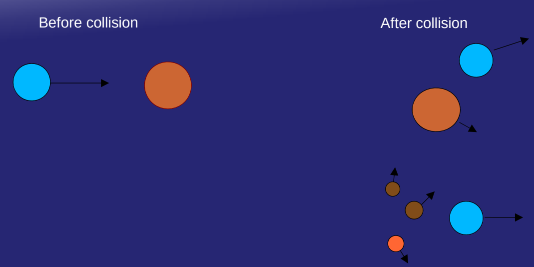
Quantum Model (1920's)

- Classical View

  Objects are classified as Matter or Energy
  - **Matter**
    - Definite Mass
    - Occupies a Specific Location or Position
    - Specific Velocity
    - Made of Particles

Quantum Model (1920's)

When particles collide they cannot pass through each other! They bounce off or shatter.



Quantum Model (1920's)

- Classical View
  Objects are classified as Matter or Energy
  - **Energy**
    - Mass-less
    - Delocalized (Spread out)
    - Non-unique Velocity
    - Wave-like

Quantum Model (1920's)

- Waves can pass through each other!
- As they pass through each other they can enhance or cancel each other.
- Later they regain their original form!