The Modern Atom

Not long after Rutherford established that the positive charge in an atom resided in the nucleus, scientists began putting together a new model of the atom using some other discoveries that had previously not been explainable.

The main discovery that got them rethinking the atom was the realization that the energy of light is quantized.

Some background information:

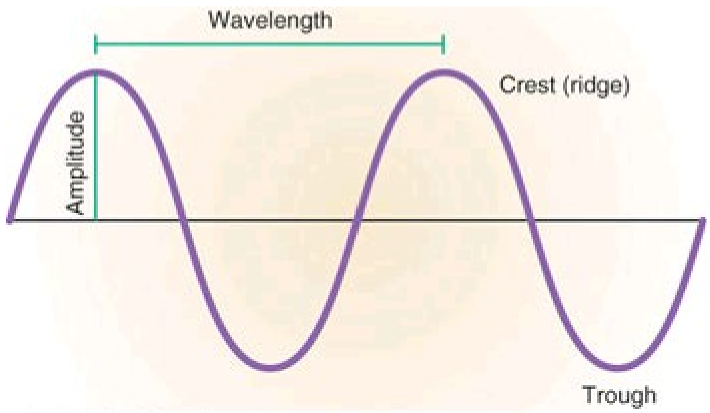
What is light?

- Light is how we refer to **electromagnetic radiation**.
- In a generic sense, electromagnetic radiation not only applies to the light that we can see, but also infrared light, UV light, microwaves, radio waves, X-Rays, and so forth.

What are the properties of light?

- **Light is a wave.**
    - Normally, we think of light as being made of little particles that leave a light source in the same way that a bullet leaves a gun.
    - In actuality, light (like all electromagnetic radiation) takes the form of a wave.
        - An analogy: The waves in the ocean can transmit huge amounts of energy, even though they aren't concrete objects.

- **Light is characterized by its wavelength and frequency.**
  - What is wavelength? (λ)



  - What is frequency? (v)
    - Frequency is the number of times that light "waves" each second.
    - It is measured in Hertz (Hz), also known as /sec.

  - How are frequency and wavelength related?
    - **λv = c**, where c is the speed of light in a vacuum (3.00×10^8 m/s).

- **The energy of light is related to its frequency (and hence, it's wavelength).**
  - E = hv, where…
    - E = the energy of light (in J).
    - h = Planck's constant = 6.626×10^{-34} J·s
    - v = frequency (in Hz or /sec)
  - For example, light with a frequency of 450 G2Hz has an energy of 3.0×10^{-22} J.

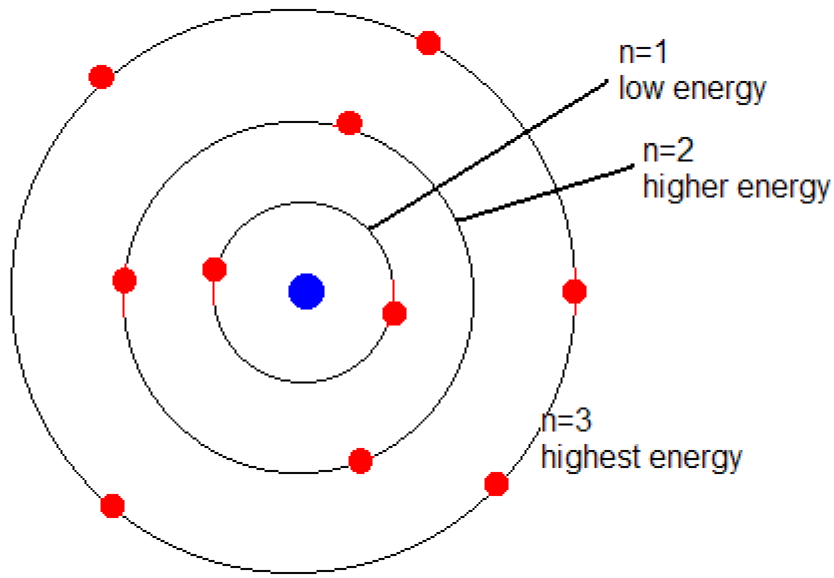
- **Light is quantized**
  - When something is quantized, this means that it can exist only at certain points.
  - An example: Increasing in altitude
    - If you walk up a ramp, you can stop and rest at any altitude you want.
    - If you walk up some stairs, you can only stop at the altitudes that correspond to height of one of the stairs.
  - This explains why the different types of light have different energies:
    - Radio waves have much less energy than visible light, which in turn has much less energy than X-rays.
    - The different colors of visible light have different energies, too – for example, blue light is far more energetic than red light.

Back to our story – what does this all have to do with the atom?

- People had begun noticing that different elements gave off different colors of light when they are heated (called their **spectra).**
  - They had no idea why this happened, but it was clear that this was reproducible and could even be used to tell different elements apart from one another, a process known as **spectroscopy**.
  - This led to the…

Bohr planetary model of the atom (1920's):

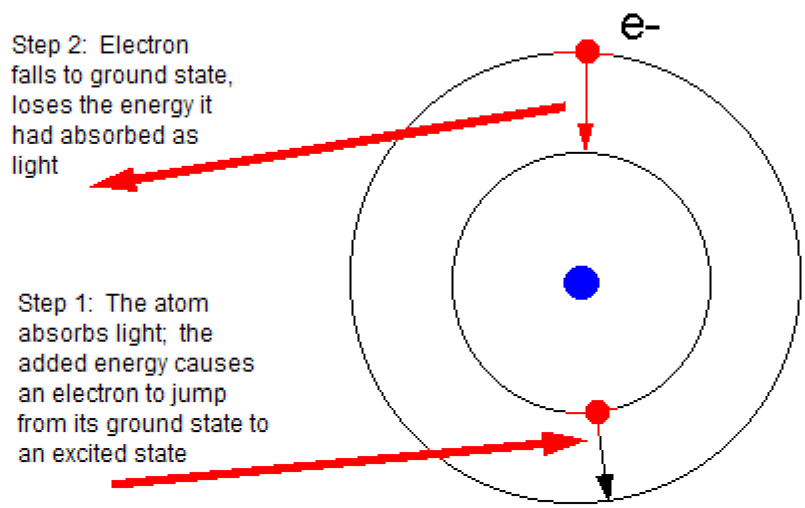
- Bohr hypothesized the following about the structure of the atom:
  - The electrons circle the nucleus in circular orbitals.
  - The orbitals closest to the nucleus have the lowest amount of energy and the orbitals gain energy the further they get from the nucleus.

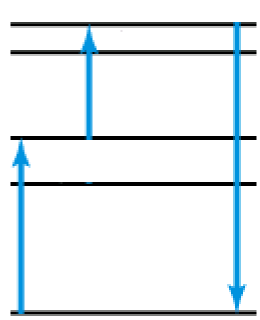


  - The variable "n" is used in the equations that determine both the distance of the orbitals from the nucleus and their energies.

How this results in the emission of different colors of light:

- When an atom is at rest, each electron lies in its lowest possible energy orbital. This orbital is the **ground state.**
- When energy is added to an atom the electrons absorb it and jump to higher energy orbitals called **excited states**.
- After some time has passed, the electrons fall back down to their original ground state orbitals, giving off the energy they have absorbed in the form of light.
- Because there are only so many orbitals, there are only certain energies of light that can be given off by an atom.
- Because the orbitals of all atoms have different energies, no two atomic spectra will ever be the same.





Explain the process using this diagram, so they'll be familiar with it, too!

- **Spectroscopy: The method by which we identify different elements using their spectra.**

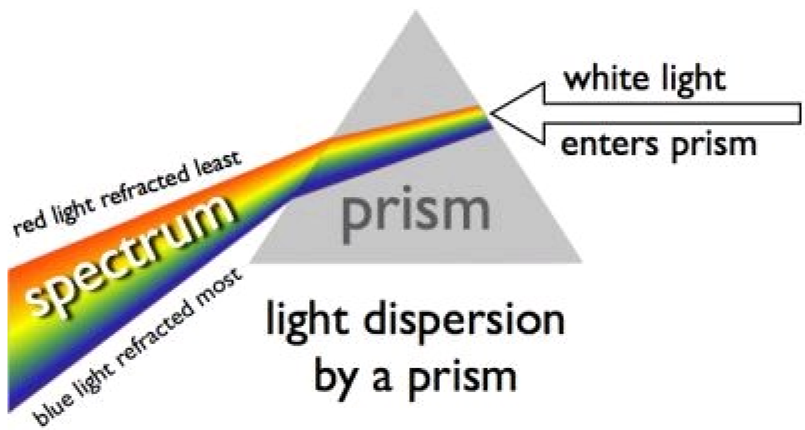
    - Show students what atomic spectra using the poster at the back of the room and explain that the difference in energies between different orbitals in each element gives rise to the observed spectra.

    - There are two kinds of spectrum:

        - **Continuous spectrum:** Looks like a rainbow and contains all energies of light. This is caused by different processes than line spectra – for more information, you can research something called “black body radiation” and Planck’s law.

        - **Line spectrum:** The spectra given off by an element, called such because the energies given off create visible lines of different colors when the light is broken up in a spectrometer.

- How to do spectroscopy:
  - **Add energy to a sample**. This energy can be in the form of heat, electricity, light, sound. This causes the electrons to absorb energy and jump to higher energy orbitals.
  - **Measure the light given off when the electrons fall back down to their ground state orbitals.** This is done with a prism or diffraction grating.



  - **Compare the emitted light to the spectra of known elements.** The one that matches is what your unknown is.

- The different types of spectroscopy vary in the colors of light they use.

    - **UV-visible light spectroscopy** (UV-Vis) measures the jumping of electrons in the way we've discussed using visible and UV light.

    - **Infrared spectroscopy** (IR) measures molecular bond vibrations and movement using IR light by an analogous process to what we discussed.

    - **Nuclear magnetic resonance** (NMR) **spectroscopy** uses radio waves and magnets to measure nuclear spin changes.

Spectroscopy – flame test lab

- **General chemistry:** Give them the flame test lab from the book.

- **Honors chemistry:** Explain the idea behind the flame test and then set them loose in the lab with the proper equipment and the labeled knowns and unlabeled unknowns.

Spectroscopy homework sheets (2)

Good things about the Bohr model:

- **It explains the existence of line spectra.**
    - No other model of the atom could give a physical basis for this well-known phenomenon.
    - It was completely in agreement with the observed data for the hydrogen atom.

- **It explains the findings from the gold foil and cathode ray tube experiments.**
    - Cathode ray tube experiment: The electrons came flying off the atom because enough energy was added to completely pull them loose.
    - Gold foil experiment: Bohr's model agreed that the nucleus, which contained all the positive charge, was present in the atom.

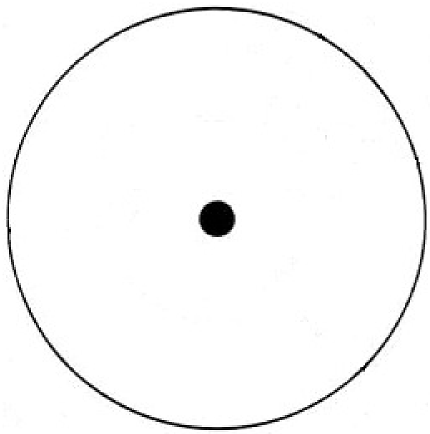
The bad thing about the Bohr model:

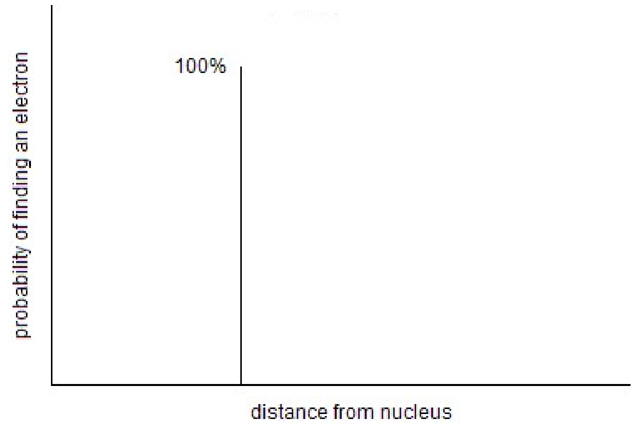
- The orbital energy predictions *only* worked for hydrogen, and not for any other elements.

Clearly, some improvements in this model were called for. Fortunately, a bunch of really smart guys were on the job and came up with…

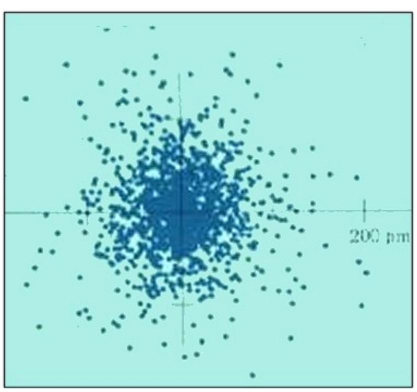
The Quantum Mechanical Model of the Atom

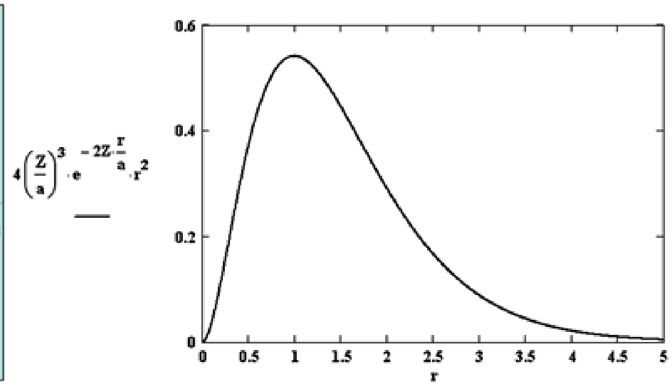
- In the Bohr atom, you know that the electrons will be traveling in circular orbits at known distances from the nucleus.





- When the very simple equations that governed the Bohr model were beefed up a bit, the equations no longer resulted in circular paths for the electrons. Instead, “orbitals” are regions of space in which electrons are likely to be found:





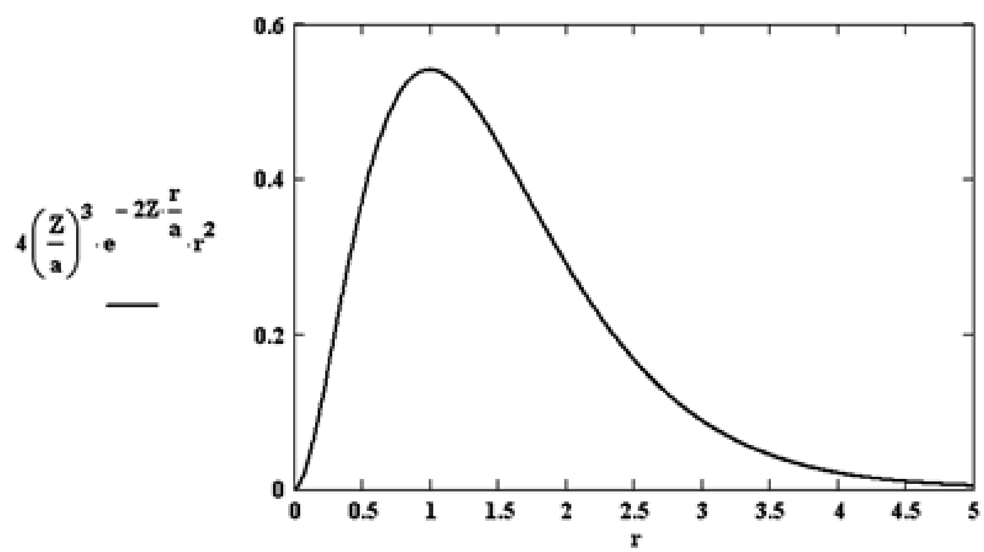
- In short, orbitals are not 2-dimensional places where an electron travels. Rather, electrons are 3-dimensional waves and the space they take up (or can potentially take up) is called an orbital.

It’s a meaningless question to ask where an electron will be in an atom, because the very question implies that an electron is a very small point that exists in one place or another.

- Electrons are smeared out all over the atom in 3-dimensional wave shapes. The shapes that electrons take (or can potentially take) are called orbitals.

- This is similar to any wave. Where **exactly** is a wave? You can’t say where a wave is because a wave is not a single point. You can, however, graph what a wave looks like.

- The graph of where an electron can be found around an atom is called a “radial probability function.”
    - This is a fancy way of graphing the intensity of an electron’s presence as it moves further away from the nucleus:

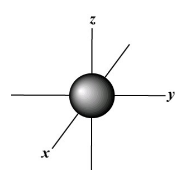


- All orbitals can hold a maximum of two electrons (they can hold 0, 1, or 2 electrons). We’ll see what that has to do with anything in a second.

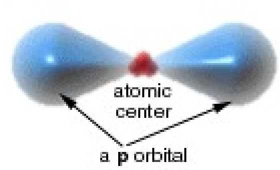
In order to make the energies work out, the Schrodinger equation (the equation that determines the probability densities of the orbitals) has four variables rather than the one variable (n) that's in the Bohr model.

- Principal Quantum Number (n): Determines energy
    - Allowable values are 1, 2, 3, …, n
    - This number determines the main energy level of an atom – it was the only variable in the Bohr model.
    - The number of electrons present in an energy level can be described by the equation $2n^2$.

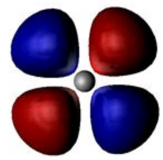
- Angular Momentum Quantum Number (l): Determines type
    - Allowable values are 0, 1, 2, …, (n-1)
    - The angular momentum quantum number determines the shape and relative energies of the electrons in an energy level.
        - l=0 defines an "s-orbital" with a spherical probability density:



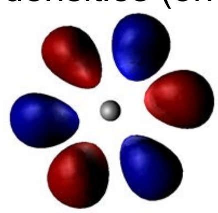
        - l=1 defines a "p-orbital" with a barbell-shaped probability density:



        - l=2 defines a "d-orbital" with variously-shaped probability densities (one of which is shown here):



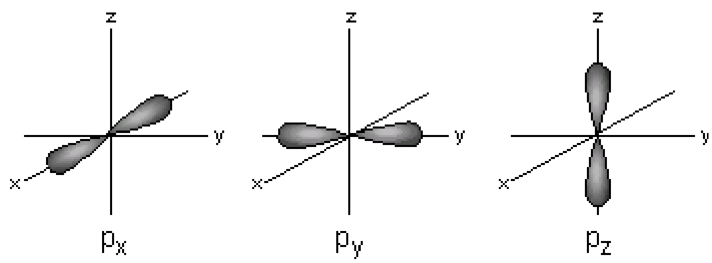
        - l=3 defines an "f-orbital" with variously-shaped probability densities (one of which is shown here):



- Magnetic quantum number (m_l): Determines spatial orientation
    - Possible values: $-l, -l+1, \ldots, -1, 0, 1, \ldots, l-1, l$
        - Example: If l=2, m_l can equal -2, -1, 0, 1, 2

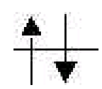
- The magnetic quantum number determines the orientation of the orbitals in space.

  - Example: p-orbitals
    - $n=2$, $l=1$, $m_l = -1, 0, 1$
    - This means that there are 3 p-orbitals and that they are oriented in three directions (x, y, z axes):



- Orbitals of the same type in the same energy level are said to be “degenerate”.

- Spin Quantum Number (m_s): Identifies which electron in each orbital we're talking about.
    - Possible values: +1/2, -1/2.
    - Why we need them:
        - The Pauli Exclusion Principle: All of the electrons in an orbital have to have unique sets of quantum numbers.
            - This is important because if they had the same quantum numbers, the wavefunctions would be identical and you couldn't fit two electrons in an orbital.
            - We have the spin quantum number so we can distinguish between the two electrons in an orbital.
            - These two electrons are said to be "spin up" and "spin down", and are shown as pointing up or down when you draw them in an orbital:



Sample problem: What are all the possible quantum numbers for the n=2 energy level?

n=2 l=0, 1 m_l = -1, 0, 1 m_s = 1/2, -1/2

So, what *really* happens during electronic transitions (as during the flame test)?

- Electrons absorb energy. This causes the electron to change its waveform such that it has different quantum numbers and different properties.

- When the electron reverts back to its original form, the energy it absorbed is given off as light.

- Important: The electron doesn't move – it just changes shape and energy!

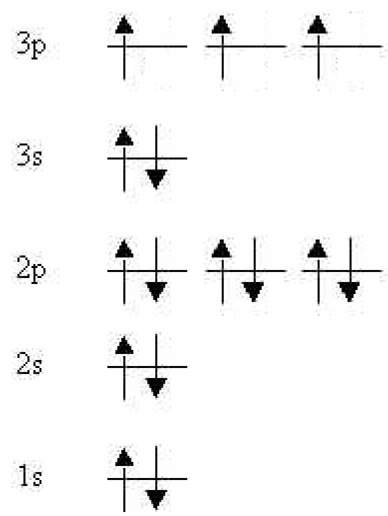
    - When we drew the ladder before to describe what happened when electrons jumped in energy, this showed *only* the energy, not where the electrons went!

Electron configurations:

- Show them how to do electron configurations, both the long and the short versions.

- Show them orbital filling diagrams.

    - Explain Hund's rule: Electrons like to stay unpaired whenever possible because of mutual repulsion.



Practice worksheet

Homework sheets