

Chem 1B, SP09
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Outline of Lectures from Chapter 12

What atoms really look like AND Building up to e⁻ configurations & periodic trends

12.1 Electromagnetic Radiation and 12.2 Nature of Matter

The overview of these two sections:

1. Light behaves as both particles and waves
2. Matter behaves as both particles and waves
 - Know how to go between frequency, energy, and wavelength
 - The smaller the mass, the more wave-like the behavior ($\lambda = h/mv$)
 - Electrons are small enough to behave as waves

we're about to see that only light of discrete energies can go into or come out of an H atom... working up to quantum mechanics

12.3 Atomic Spectrum of Hydrogen Atom

Overview of 12.3:

1. We can measure energy of transitions (e.g. n=1 to n=2) by looking at *color* of light that is absorbed or emitted over that transition
2. Transitions occur at discrete wavelengths (only certain colors)
 - energy of electron transitions (e.g. n=1 to n=2) is QUANTIZED!

*transitions between ground and excited states are quantized, so what does this mean an atom looks like?
Here's the "best" model before quantum mechanics:*

12.4 The Bohr Model

1. Isn't really correct, but says that electrons exist in "stationary orbits".
2. Does work for finding energies of hydrogen in different excited states
 - $E = -2.18 \times 10^{-18} \text{J} (Z^2/n^2)$

Bohr model is a good attempt, but it doesn't work for anything other than hydrogen... need a new model:

12.5 Quantum Mechanical Description of the Atom

1. Heisenberg's uncertainty principles says, the *uncertainty* in a particle's position is greater when the momentum is known to higher accuracy, and the *uncertainty* in a particles momentum is greater when the particle's position is known to a higher accuracy. This means that the electron in an atom must exist as a *probability density distribution* around the nucleus. The electron is like a cloud.
 - now we know that electrons in atoms are like clouds, but what are the shapes of the clouds?*
2. The standing-wave model is used to come up with the "wavefunctions". We plug these into the Schrodinger equation, and out comes the quantum numbers.
 - From here we get *energy levels* and *shapes* of orbitals

Almost every quantum mechanics equation is solved in exactly the same way, this model is the most simple one that shows how to do it. Don't worry too much about this section unless you're interested...

12.6 Particle in a Box – Don't worry about this, it's just a useful model

12.7 Skip this

12.8 Physical meaning of a wavefunction

→ All you need to know about this is that we use a wavefunction (ψ) to describe an electrons because they're *wave-like*. The square of the wavefunction (ψ^2) gives the *probability density* of an electron. This is plotted versus radius (r) to get the *probability density distribution*.

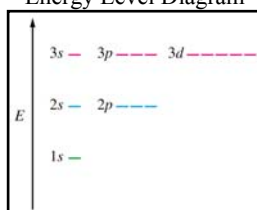
→ All this tells us the geometry of the electron distribution (geometry is important for bonding!)

Under what circumstances do electrons in atoms exist in different orbitals?! Let's start with hydrogen...

12.9 Characteristics of Hydrogen Orbitals and 12.10 e⁻ spin and Pauli

1. From the Schrodinger Equation we get quantum numbers... n , l , m_l , and m_s
 n tells us the energy of the electron
 l tells us which shapes of orbitals are occupied: s , p , d , or f
 m_l tells us how many orbitals there are at each l value. For example, at $l = 1$ there are 3 m_l values, thus 3 p -orbitals.
 m_s tells us the spin of the electron – for two electrons to be in the same orbital they must have opposite spins (Pauli)
2. *IMPORTANT* for an H-atom, n specifies the energy, so $2s$, $2p_x$, $2p_y$, and $2p_z$ all have the same energy (we call them “degenerate” orbitals if they have the same energy). When an H atom is excited into the $n=2$ state its electron takes on the shape of ALL 4 degenerate orbitals simultaneously.
3. This 4-fold degeneracy of $n=2$ is NOT the case for atoms other than hydrogen.

H Atom orbitals for $n=1$, $n=2$, and $n=3$
Energy Level Diagram

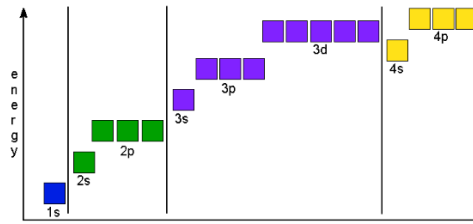


What about other atoms?

12.13 Aufbau Principle and the Periodic Table

1. When we talk about atoms other than hydrogen we can take the orbitals in their energy level diagram and fill in the electrons according to the Pauli Exclusion Principle. The $n=2$ orbitals are no longer degenerate because of the $n=1$ orbitals which remain filled (we're not talking about electron excitation here, we're building up multi-electron atoms).

Spacing of orbital energy levels (Aufbau Principle)



2. So now we add the electrons from lowest energy up. If we have 3 electrons, we'll have 2 in 1s orbital and 1 in 2s, so the electron configuration will be $1s^2 2s^1$.

WE WILL GO MORE IN DEPTH ON ELECTRON CONFIGURATIONS ON MONDAY