IDX G9 Chemistry H
Study Guide S1 Monthly 2
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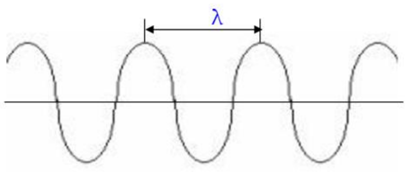
Contents:

6.1 The wave nature of light

- Light: EM radiation, with speed=$3x10^8$m/s



The distance between corresponding points on adjacent waves is the **wavelength (λ)**.

Wavelength unit: m

Unit conversion reminder: nm=10^{-9}m

Frequency(v): how many waves per second going through a point—>s^{-1} is the unit, or Hz.

- Higher frequency indicates this kind of light has greater energy.
- The speed of light, $c=\lambda v$ Thus λ and v is inversely proportional
  - a light with higher energy has high frequency, and thus small wavelength
  - a light with lower energy has low frequency, and thus large wavelength(relatively).

- Range of visible light:400-750nm

Different colors have different wavelengths(λ) and different frequency(v).
Compare the λ and$v$ from violet to red light.

violet, indigo, blue, green, yellow, orange, red.

400 500 600 750 nm

Wavelength: violet < indigo < blue < green < yellow < orange < red.

frequency: violet > indigo > blue > green > yellow > orange > red.

energy: violet > indigo > blue > green > yellow > orange > red.

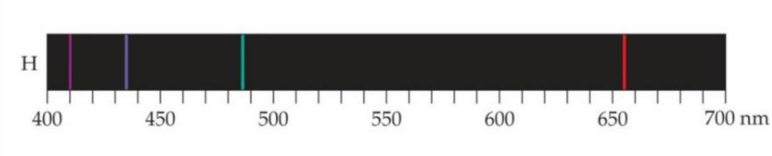
6.2 Quantized Energy and Photons

- Energy come in the smallest unit called **quantum**. Note that quantum is not a single, constant value. Quantum simply indicates the amount of energy that is contained/carried by each individual photon.
  - Energy of single quantum: $E=hv$ h(Planck's constant)=$6.626x10^{-34}$ Joules times second (Js)
- In the **photoelectric effect**, electrons are ejected when light shines on a metal. There are 2 key factors:
  - 1. Frequency. When the frequency of the light reached a certain level, $v_{threshold}$, the corresponding energy reached a certain level $E_{threshold}$ or $E_{minimum}$. Any energy greater or equal to can cause the ejection of electrons, and higher the energy, **higher the speed of the ejected electron**.
    - Kinetic Energy(correspond to speed of electron). $KE=E_{light} - E_{min}$ note that $KE=1/2mv^2$.
  - 2. Amplitude(light intensity): How bright is the light? This determines how many photons is in the light. Higher amplitude indicates higher amount of electrons. The amplitude only matters if $v_{threshold}$ is met. Only then, **the amplitude decides how many electrons are ejected.**

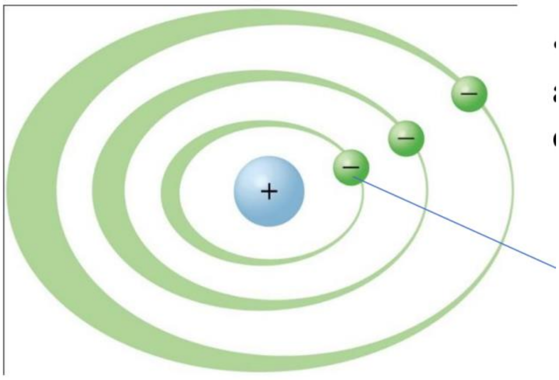
6.3 Line Spectra and the Bohr Model

- Continuous spectrum vs. line spectrum
  - Continuous spectrum: one color fades into another with no discontinuation in between
    - example: sunlight, incandescent light
    - red,yellow,orange,green... this is called the visible spectrum.
  - line spectrum: discrete lines of colors in the spectrum.

- The Atomic Emission Spectrum is a line spectrum.



  - From the figure above, 4 lines here indicate 4 possible frequencies of the light emitted by the atom, after exciting of electrons.
  - Can only explain the electron of hydrogen, but failed to explain other complex atoms.
- The Bohr Model



- This is the Bohr Model. Identification factors: circular orbits, electrons, a positively charged nucleus.
- The proposal of electrons orbiting the nucleus like planets, the planetary system, is **not** by the Bohr Model. Instead, it was first established by the preceding Nuclear(Rutherford) Model.

- Principal Quantum number (n):1,2,3,... describes energy levels
  - A higher n means higher energy **of the electron**.
  - Moving from higher n to lower n results the **emission of energy from the electron**. Which means, the electron now have LESS energy.
  - As electrons absorbed energy, they are excited. Their potential energy increases, which correlates to their distance to the nucleus. You can understand this as gravitational potential energy, which an object gains as it reached a certain height. The gravity, g is virtually analogized to the nuclear charge attracting the electrons towards the nucleus. Thus, with a higher PE, the electrons are able to reach "higher height" which in this case, further from the nucleus.

6.4 The Wave Behavior of Matter

- 1. Louis de Broglie: light behave like matter, matter behave like waves.
  - Note that sound waves do not behave like matter, so not all waves have matter properties. BUT, all matter have partial wave properties, in this case, **particle wavelength λ.**
  - **λ=h/mv** h is planck's constant. mv is momentum, mass x speed. Macroscopic objects usually have a very small wavelength because their mass is too big.
  - "matter waves"
- 2. Heisenberg Uncertainty Principle: The position and the velocity/momentum of a moving electron cannot simultaneously be measured and known exactly.

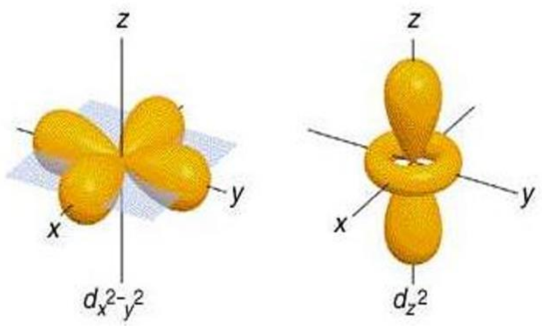
- Only for microscopic/subatomic level.
- Strike electron with photon can measure velocity and position. However this process of striking will deviate the electron's original pathway and state.

6.5 Quantum mechanics and atomic orbitals

- Schrodinger made the quantum mechanical model
- 3 Defining factors:
  a. Quantized energy: electrons only exist in certain amounts of energy levels.
  b. Electron exhibit wavelike behavior: electrons, also have associated wavelengths. They move fastly around the nucleus like a wave. They interact with themselves to form constructive and destructive interference, which is like the diffraction of light. This links to the determination of orbitals and the formation of certain energy levels. You do not have to know these properties, only the highlighted part is tested. But if you are interested, explore more on https://www.youtube.com/watch?v=kpa_HO6yy18
  c. Heisenberg uncertainty principle especially for electrons.
- Atomic Orbitals: a region of space that describes the probability of finding an electron. The denser the graph is dotted, it means more likely electron is going to be found there.
  - 90% of electrons found within the orbital's surface.

6.6 Representations of Orbitals

- **spdf orbitals: corresponding to 1,3,5,7**
- **n=1 ->s**
- **n=2->sp**
- **n=3->spd**
- **n=4->spdf**
- not limited to spdf orbitals only. There's new ones in n=5 and so on, but no need to know that.
- s: sphereical p:dumbell d:four-leaf clover
- Special orbitals(special shape) needed for attention:

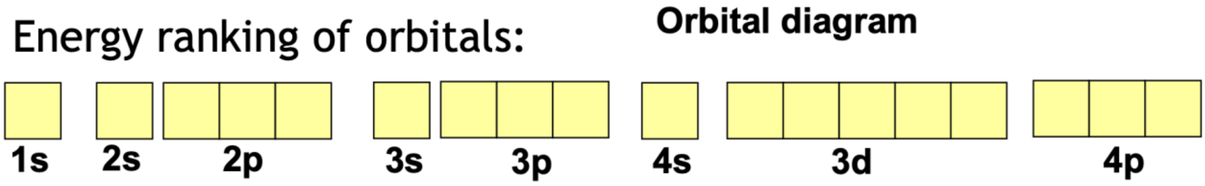


- General pattern for orbitals and electron numbers:
- #of orbitals in a given n=n^2

- #of electrons in a given n=$2n^2$

6.7 Many Electorn Atoms

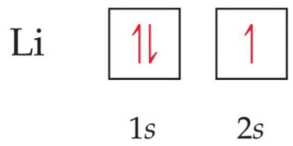
- Orbitals and their energies
  - In one electron atom, subshells all have the same energy. In many electron atoms, electron-electron repulsions cause different subshells at different energy. Energies of subshells follow order ns< np< nd <nf
  - In many electron atoms,orbital sets in the same sublevel have the same energy. Chemists call them degenerate orbitals.
  - The range of energy levels within a principal energy level can overlap the energy levels of another principal level. (e.g., 4s is lower in energy than 3d.)
- Aufbau Principle
  - electrons occupy the orbitals of lowest energy first.
  - Orbitals of the same sublevel always have the same energy.
  - On the same principle energy level, s<p<d<f
  - The range of energy levels within a principal energy level can overlap the energy levels of another principal level.
  - Energy ranking: 1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p 6s 4f 5d 6p 7s 5f 6d 7p
- Pauli Exclusion Principle
  - an atomic orbital can hold a maximum of two electrons and they must have opposite spin(clockwise or counterclockwise).
  - When electrons with opposite spins occupy an orbital, the electrons are said to be paired; a single electron in one orbital is unpaired
  - Place in arrows up and down into the boxes belew when drawing orbital diagrams.



6.8~6.9 Electron configurations and the periodic table

- Orbital Diagrams and Electron Configurations

Orbital Diagrams



- Each box in the diagram represents one orbital.
- Half-arrows represent the electrons.
- The direction of the arrow represents the relative spin of the electron.

Electron Configurations

$1s^2 2s^2$

- A number denoting the energy level;
- A letter denoting the type of orbital;
- A superscript denoting the number of electrons in those orbitals.

- Hund's Rule
  - For degenerate orbitals, the lowest energy is attained when the number of electrons with the same spin is maximized.
  - This means that, for a set of orbitals in the same sublevel, there must be one electron in each orbital before pairing and the electrons have the same spin, as much as possible.
- Abbreviated Electron configurations (Condensed electron configurations, noble gas notation)
  - the electron configuration of the nearest noble-gas element of lower atomic number(core electrons) is represented by its chemical symbol in brackets

1s | 2s 2p | 3s 3p | 4s 3d 4p | 5s 4d 5p | 6s 4f 5d 6p | 7s 5f 6d 7p···

[He] 2 [Ne] 10 [Ar] 18 [Kr] 36 [Xe] 54 [Rn] 86

eg: The noble gas in 2nd period is Ne: $1s^2\ 2s^2\ 2p^6$

Elements in 3rd period like Cl: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

Cl: [Ne] $3s^2\ 3p^5$

- Valence electrons
  - The inner-shell electrons are referred to as the core electrons.
  - The electrons given after the noble-gas core are called the outer-shell electrons. The outer-shell electrons include the electrons involved in chemical bonding, which are called the valence electrons.

- For the elements with atomic number 30 or less, all of the outer-shell electrons are valence electrons. The v.e. number for transition elements may vary. e.g. Fe [Ar]4s23d6
- Valence electrons are involved in chemical bonding and are largely responsible for an atom's chemical behavior. e.g. alkali metals: Li [He]2s1, Na [Ne]3s1, K [Ar]4s1, …

- Electron configuration of excited state
  - The most stable organization is the lowest possible energy, called the ground state. The electron configuration for an excited state electron shows one or more electrons in a higher energy level.
- Electron Configuration for ions
  - Metal atoms → lose electrons → cations
  - Transition metals will lose outmost shell electrons first, then lose the inner d level electrons.
  - Nonmetal atoms → gain electrons → anions
- Exceptional Electron configurations

Incorrect configurations:	Correct configurations:
Cr $1s^2 2s^2 2p^6 3s^2 3p^6 \mathbf{3d^4 4s^2}$	$1s^2 2s^2 2p^6 3s^2 3p^6 \mathbf{3d^5 4s^1}$
Cu $1s^2 2s^2 2p^6 3s^2 3p^6 \mathbf{3d^9 4s^2}$	$1s^2 2s^2 2p^6 3s^2 3p^6 \mathbf{3d^{10} 4s^1}$

- Electron configuration and periodic table
  - Different blocks on the periodic table correspond to different types of orbitals: s and p-block are representative elements; d block are transition elements; f block are lanthanides and actinides, or inner transition elements