Contents:

6.1 Wave Nature of Light

- Light is electromagnetic radiation (EMR): oscillating electric and magnetic fields.

- Important wave properties:

  o **Wavelength (λ)** – distance between peaks (units: m, nm).

  o **Frequency (ν)** – number of waves per second (units: s^{-1} = Hz).

- All EM waves in vacuum travel at: **Speed of light: $c = 3.00 \times 10^8$ m/s.**

- Relationship between λ and ν: **$c = \lambda\nu$.**

- **Visible Spectrum(from shortest λ / highest energy to longest λ / lowest energy): Violet $\rightarrow$ Indigo $\rightarrow$ Blue $\rightarrow$ Green $\rightarrow$ Yellow $\rightarrow$ Orange $\rightarrow$ Red.**

  o So: Wavelength: λ(violet) < … < λ(red); Frequency & energy: ν(violet) > … > ν(red).

- **Typical Question Types**

  o Given λ, find ν (or vice versa) using $c = \lambda\nu$.

  o Compare colors or regions of EM spectrum in terms of λ, ν, and energy.

  o Conceptual questions about how changing λ affects energy.

- **Common Traps**
  - **Forgetting to convert nm → m (Very Important!)** when using formulas.
  - Confusing brightness with energy per photon (**energy per photon depends on frequency**, not brightness).

6.2 Quantized Energy and Photons

- **Planck: Energy is quantized** — it comes in packets called quanta.
  - **Planck's constant: $h = 6.626 \times 10^{-34}$ J·s.**
- The energy of one photon: $E = h\nu$.

- **Photoelectric effect**:
  - Light shining on a metal can eject electrons from the surface.
  - **Ejection only happens if photon energy ≥ threshold energy of the metal (frequency must be ≥ threshold frequency ν_0).**
  - If you **increase intensity (brightness) at the same frequency**: more photons → **more electrons ejected**, but same KE for each electron.
  - If you increase frequency (above ν_0): each photon has more energy → electrons have higher KE.
  - **Remember that brightness doesn't change the frequency!**

- **Important Formulas:**
  - Photon energy: $E\_photon = h\nu = hc/\lambda.$
  - Threshold energy: $E\_threshold = h\nu_0.$
  - Kinetic energy of an ejected electron: $KE = E\_photon - E\_threshold = h\nu - h\nu_0.$

- **Typical Question Types**
  - Calculate the energy of a photon given ν or λ.
  - Find the total energy of N photons.
  - Given λ of light and the metal's threshold frequency ν_0, find the KE of ejected electrons or decide if electrons are ejected.

- **Common Traps**
  - Using $E = h\nu_0$ when the question is asking for KE. Remember: $KE = h(\nu - \nu_0)$.
  - Thinking very intense low-frequency light can eject electrons. If $\nu < \nu_0$, no electrons are ejected, **no matter how bright.**

6.3 Line Spectra and the Bohr Model

- Continuous vs Line Spectra:
  - **Continuous spectrum:** contains all wavelengths in a range (like white light through a prism → smooth rainbow).
  - **Atomic emission spectrum**: only specific discrete lines (certain wavelengths/frequencies), each element has a unique line spectrum.

- **Bohr Model of the Hydrogen Atom**:
  - Electrons occupy specific energy levels (orbits) labeled by $n = 1, 2, 3, \ldots$.
  - $n = 1$: lowest energy level → ground state.
  - $n \geq 2$: higher energy levels → excited states.
  - When electrons absorb energy, they jump from lower n to higher n (excited state).
  - When they fall back to lower n, they emit photons whose energy equals the energy gap between the levels.
    - **Any state n>3 that goes back to n=3 level exhibit infrared rays**
    - **Any state n>2 that goes back to n=2 exhibit visible light**
    - **Any state n>1 that goes back to n=1 (ground state) exhibit UV light**

- Energy change in a transition: **$\Delta E = E\_higher - E\_lower = h\nu.$**
- Because only specific n levels exist, only specific energy differences are allowed → only specific frequencies (lines) appear.

- **Strengths and Limits of Bohr Model**
  - **Strengths**: explains line spectrum of hydrogen and other one-electron ions (He^+, Li^{2+}, etc.); introduces quantized energy levels.
  - **Limits**: doesn't work well for multi-electron atoms; electrons are not actually moving in simple circular orbits; cannot fully explain chemical behavior trends.

- **Typical Question Types**
  - Explain why hydrogen shows a line spectrum instead of a continuous one.
  - Define ground state and excited state in terms of energy levels n.
  - Relate a spectral line to an electron transition (higher n → lower n).

- **Common Traps**
  - Mixing up absorption (electron moves up in n) and emission (electron falls down in n).
  - Forgetting that a specific line = a specific pair of energy levels.

6.4 The Wave Behavior of Matter
- **de Broglie Matter Waves**:
  - Particles of matter have **wave properties.**
  - Wavelength of a particle: $\lambda = $ **h/(mv)**.
  - **For large objects** (like a baseball or a car), mass is huge → λ is extremely small → **wave behavior is negligible.**
  - **For very small particles** (like electrons), mass is tiny → λ can be comparable to atomic dimensions → **wave nature is important.**

- **Heisenberg Uncertainty Principle**:
  - Fundamentally impossible to know both the exact position and exact momentum/velocity of an electron at the same time.
  - Measuring an electron (for example, by hitting it with a photon) disturbs it; this is a built-in limit of nature, not just bad equipment.

- o de Broglie wavelength: $\lambda = h/(mv)$.

- **Typical Question Types**

  - o Calculate or compare de Broglie wavelengths of particles with different masses and speeds.
  - o Explain why we see wave behavior for electrons but not for macroscopic objects like soccer balls.
  - o Explain the uncertainty principle qualitatively.

- **Common Traps**
  - o Thinking the uncertainty principle is just due to bad measurements instead of a fundamental physical limit.
  - o Forgetting that matter waves matter most for very small, light particles (like electrons).

6.5 Quantum Mechanics and Atomic Orbitals

- **From Bohr to Quantum Mechanics**:
  - o The quantum mechanical model combines quantized energies, wave nature of matter, and the uncertainty principle.
  - o Schrödinger developed a wave equation that describes electrons in atoms.

- **Schrödinger's Wave functions and Orbitals**:
  - o Solutions to Schrödinger's equation are wave functions that describe atomic orbitals.
- **Atomic orbital**: a 3D region of space around the nucleus where there is a high probability of finding an electron (often drawn as enclosing about 90% of the electron's probability).

- **Orbit vs Orbital:**

- o **Bohr orbit**: fixed circular path with defined radius (like a mini planet orbiting the nucleus).
  - o **Quantum orbital**: no fixed path, a probability cloud showing where the electron is likely to be found.
- Energy levels are still quantized, but electrons are not in neat circular paths.

- **Typical Question Types**
  - o Define atomic orbital (must mention region of space and probability).
  - o **Compare orbit vs orbital.(Remember they are different!)**
  - o Explain why we can't state the exact location of an electron in an atom.

- **Common Traps**
  - o Saying that an orbital is a "path" or "track" instead of a region of high probability.
  - o Forgetting that even though we use probability clouds, the energy levels are still quantized.

6.6 Representations of Orbitals
- Different atomic orbitals denoted by **s, p, d, f**
  - o s: Spherical shape
  - o p: Dumbell shape
  - o d: Four leaf clover shape
- For a given principle energy greater than 1, there's **1s orbital, 3p, 5d, 7f**, due to the different orientations of the orbitals.
- The principle quantum number n always equals the number of sublevels within that principle energy level.

- **For example:** 1s 2s 3s orbitals are identical in size and orientation, **but their sizes are increasing due to electrons with increasing energy.**

6.7 Many Electron Atoms

- 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** (e.g. n=3 orbitals increase in energy in the order 3s<3p<3d.)

- 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.

- **Aufbao Principle**
  - According to the 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. (e.g. **4s and 3d**)
    - **When writing electron configurations, look out for s and d! For example, when 4s (2) and 3d (4), 4s should give one electron to 3d to make it more stable, making it 4s (1) 3d (5)**

- **Energy Ranking of electrons : 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 h**old a maximum of two electrons and they must have opposite spin.**
  - According to the Pauli exclusion principle, when electrons with **opposite spins** occupy an orbital, the electrons are said to be **paired**; a **single electron** in one orbital is **unpaired**.

6.8 - 6.9 Electron Configurations and the Periodic Table

- **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**
    - 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)**
    - In writing the Abbreviated electron configuration of an element, **the electron configuration of the nearest noble-gas element of lower atomic number(core electrons) is represented by its chemical symbol in brackets.** Eg: [Ne]

- **Inner shell electrons, Outer shell electrons and 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)

- o **Valence electrons are involved in chemical bonding** and are largely responsible for an atom's chemical behavior.

- **Electron Configuration of Excited State**
  - o 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**
  - o **Transition metals will lose outmost shell electrons first, then lose the inner d level electrons.**
- **Exceptional Electron Configurations**
  - o Some actual electron configurations differ from those assigned using the aufbau principle because although half-filled sublevels are not as stable as filled sublevels, they are more stable than other configurations.
- **Electron Configuration and the Periodic Table**
  - o Chemists **fill orbitals in increasing order of energy.**
  - o 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**

s,p,d,f blocks in the periodic table.