THE NATURE OF LIGHT

- Visible light is one type of electromagnetic radiation (electromagnetic radiation).
- The electromagnetic radiation has the wave properties:
  - Frequency ($\nu$): the number of cycles the wave undergoes per second $\Rightarrow 1/$second $\approx$ Hz
  - Wavelength ($\lambda$): the distance between any point on a wave and the corresponding point on the next crest of the wave (the distance the wave travels during one cycle) $\Rightarrow m$, nm, Å ($10^{-10}$ m)
Speed of light (c): $3 \times 10^8$ m/s == $c = \nu \lambda$

where $c$ is constant

That's mean, radiation with a high frequency has a short wavelength, and vice versa

Another characteristic of a wave is: AMPLITUDE

Amplitude: the height of the crest of each wave or intensity of the wave/radiation

The two waves shown have the same wavelength (color) but different amplitudes, and therefore different brightnesses (intensities)

The Electromagnetic Spectrum

Visible light represents a small portion of the continuum of radiant energy, known as electromagnetic spectrum

All the waves in the spectrum travel at the same speed through a vacuum but differ in frequency and therefore wavelength

Wavelength of visible light as different colors: from red ($\lambda=750$ nm) to violet ($\lambda=750$ nm)

Light of single wavelength is called MONOCHROMATIC

Light of many wavelength is called POLYCHROMATIC
**Distinction Between Energy and Matter**

- When light wave passes from air to water, the speed of the wave changes ==> **refraction**
- After strikes the boundary, the light continues at a different angle, therefore change in speed and direction
- The new angle depends on the materials on either side of the boundary and the wavelength of the light
- White light is dispersed into its component colors when pass through a prism, because each incoming wavelength is refracted at a slightly different angle.

---

If waves of light pass through two adjacent slits, the emerging circular waves interact with each other through the process of **interference**.

- If the crests of the waves coincide (in phase), they interfere **constructively** and the amplitudes add together.
- If the crests coincide with troughs (out of phase), they interfere **destructively** and the amplitudes cancel.
- The result is a diffraction pattern of **brighter** and **darker** regions.
The Particle Nature of Light

1. Blackbody Radiation
2. The Photoelectric effect
3. Atomic Spectra

Blackbody Radiation

- When a coal is heated to 1000 K ==> emit visible light (red glow)
- At 1500 K, the light is brighter and more orange, like that from an electric heating coil (elemen pemanas listrik)
- These changes in intensity and wavelength of emitted light as an object is heated are characteristic of blackbody radiation.
- In 1900, Max Planck ==> hot or glowing object could emit or absorb only certain quantities of energy:
  \[ E = nh\nu \]

- \( E \): energy of radiation (J); \( \nu \): frequency (s^{-1}); \( n \): positive integer of a quantum number; \( h \): proportionality constant (Planck's constant in J.s)
- \( h = 6.62606876 \times 10^{-34} \text{ J.s} = 6.626 \times 10^{-34} \text{ J.s} \)
Quantum Energy

- If an atom itself can emit only certain quantities of energy ==> the atom itself can have only certain quantities of energy
- Thus, the energy of an atom is quantized
- Each change in the atom's energy results from the gain or loss of one or more “packet” (amount) of energy.
- Each energy packet is called a quantum (= hν)
- An atom changes its energy state by emitting (or absorbing) one or more quanta, and the energy of the emitted (or absorbed) radiation is equal to the difference in the atom's energy states:
  \[ ∆E_{atom} = E_{emitted \text{ or absorbed} \text{ radiation}} = ∆n hν \]
- The atom can change its energy only by integer multiples of hν
  ==> the smallest change occurs when an atom in a given energy state changes to an adjacent state when ∆n=1
  \[ ΔE = hν \]

Photoelectric Effect and Photon Theory of Light

- When monochromatical light of high enough frequency strikes the metal plate, electrons are freed from the plate and travel to the positive electrode, creating a current
- Planck's idea of quantized energy ==> Einstein: “light itself is particulate, that is quantized into small bundles of electromagnetic energy” ==> PHOTONS
- Planck ==> each atom changes its energy whenever it absorbs or emits one photon, one particle of light, whose energy is fixed by its frequency:
  \[ E_{photon} = hν = ∆E_{atom} \]
How Einstein's Photon Theory Explains the Photoelectric Effect?

- According to the photon theory, a beam of light consists of an enormous number of photons.
- Light intensity is related to the number of photons striking the surface per unit time, but not to their energy.
- Therefore, a photon of a certain minimum energy must be absorbed for an electron to be freed.
- Since energy depends on frequency ($h\nu$), the theory predicts a threshold frequency.
- An electron can not save up energy from several photons below the minimum energy until it has enough to break free.
- Rather, one electron breaks free the moment it absorbs one photon of enough energy.
- The current is weaker in dim light than in bright light because fewer photons of enough energy are present, so fewer electrons break free per unit time.
- But some current flows the moment photons reach the metal plate.

Examples of Energy Radiation

- A cook uses a microwave oven to heat a metal. The wavelength of the radiation is 1.20 cm. What is energy of one photon of this microwave radiation?

  Solution:
  \[ E = h\nu = \frac{hc}{\lambda} \]
  \[ = \frac{(6.626\times10^{-34} \text{ J.s})(3.00\times10^8 \text{ m/s})}{(1.20 \text{ cm})(10^{-2} \text{ m/cm})} \]
  \[ = 1.66\times10^{-23} \text{ J} \]
ATOMIC SPECTRA

- What happens when an element is vaporized and then electrically excited?
- Light from excited Hydrogen atoms passes through a narrow slit and is then refracted by a prism.
- This light does not create a continuous spectrum, or rainbow, as sunlight does.
- Instead, it creates a **line spectrum**, a series of fine lines of individual colors separated by colorless spaces (black).
- The wavelength of these spectral lines are characteristic of the element producing them.
- Next the figure =>

Example: The line spectra of hydrogen
Spectral Lines of Hydrogen Atom

- Spectroscopist studying the spectrum of atomic hydrogen had identified several series of such lines in different regions of the electromagnetic spectrum.
- Three series of spectral lines of atomic hydrogen:

![](image)

Rydberg Equation

- Rydberg Equation $\Rightarrow$ to predict the position and wavelength of any lines in a given series:

$$\frac{1}{\lambda} = R \cdot \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$$

where $\lambda$ is the wavelength of a spectral line, $n_1$ and $n_2$ are positive integers with $n_2 > n_1$, and $R$ is the Rydberg constant ($1.096776 \times 10^7 \text{ m}^{-1}$).

- For the visible series of lines, $n_1 = 2$:

$$\frac{1}{\lambda} = R \cdot \left( \frac{1}{2^2} - \frac{1}{n_2^2} \right), \text{with } n_2 = 3, 4, 5, ...$$
The Bohr Model of The Hydrogen Atom

- **Niels Bohr** (1885-1962) suggested a model for the H atom that predicted the existence of **line spectra**.
- In this model, Bohr used Planck’s and Einstein’s ideas about quantized energy and proposed three postulates:
  - **The H atom has only certain allowable energy levels** => stationary states. Each of these states is associated with a fixed circular orbit of the electron around the nucleus.
  - **The atom does not radiate energy while in one of its stationary states**. That is, even though it violates the ideas of classical physics, the atom does not change energy while the electron moves within an orbit.
  - **The atom changes to another stationary state (the electron moves to another orbit) only by absorbing or emitting a photon** whose energy equals the difference in energy between the two states:
    \[ E_{\text{photon}} = E_{\text{state } A} - E_{\text{state } B} = h\nu \]

Quantum staircase of Hydrogen Atom

- Analogy for the energy levels of the hydrogen atom
- An electron can absorb a photon and jump to a higher step (stationary state)
- or emit a photon and jump down to a lower one
- But the electron cannot lie between two steps
Bohr's Model

- A spectral line results when a photon of specific energy (and frequency) is emitted as the electron moves from a higher energy state to a lower one.
- Therefore, Bohr's model explains that the atomic spectrum is not continuous because the atom's energy has only certain discrete levels or states.
- In Bohr's model, the quantum number $n$ is associated with the radius of an electron orbit, which is related to the electron's energy.
  
  the lower the $n$ value, the smaller the radius of the orbit, and the lower the energy level.
- When the electron is in the first orbit ($n=1$), the orbit closest to the nucleus, the H atom is in its lowest energy level ==>
  
  GROUND STATE

Cont'd ....

- If the H atom absorbs a photon whose energy equals the difference between the first and second energy levels, the electron moves to the second orbit ($n=2$), the next orbit out from the nucleus.
- When the electron is in the second or any higher orbit, the atom is said to be in an **EXCITED STATE**
- If the H atom in the first excited state (electron in second orbit) emits a photon of that same energy, it returns to the ground state.
- When electron drops from an outer orbit to an inner one, the atom emits a photon of specific energy that give rise to a spectral line ==>
  
  look at the next Figure
Limitations of the Bohr's Model

- The Bohr Model failed to predict the spectrum of any other atom, even that of helium, the next simplest element.
- It is suitable for H atom and for other one-electron species.
- But, it does not work for atoms with more than one electron because in these systems, additional nucleus-electron attractions and electron-electron repulsions are present.
- As a picture of the atom, the Bohr model is incorrect, but we still use the terms “ground state” and “excited state”.
- and retain one of Bohr's central ideas that: “the energy of an atom occurs in discrete levels”
The Energy States of the Hydrogen Atom

- Bohr’s work ==> calculation of energy levels of an atom (which derived from principles of electrostatic attraction and circular motion):
  \[ E = (-2.18 \times 10^{-18} \text{ J}) \left( \frac{Z^2}{n^2} \right) \]
  
  where \( Z \) is the charge of the nucleus.

- For H atom with \( Z=1 \):
  \[ E = (-2.18 \times 10^{-18} \text{ J}) \left( \frac{1}{n^2} \right) = (-2.18 \times 10^{-18} \text{ J}) \left( \frac{1}{n^2} \right) \]

- Therefore, the energy of the ground state (\( n=1 \)) is:
  \[ E = (-2.18 \times 10^{-18} \text{ J}) \left( \frac{1}{1^2} \right) = -2.18 \times 10^{-18} \text{ J} \]

- The negative sign appears because we define the zero point of the atom's energy when the electron is completely removed from the nucleus.

- Thus, \( E=0 \) when \( n=\infty \), so \( E<0 \) for any smaller \( n \).

\[
E = -2.18 \times 10^{-18} \text{ J} \frac{Z^2}{n^2}
\]

- Since \( n \) is in the denominator of the energy equation, as the electron moves closer to the nucleus (\( n \) decreases), the atom becomes more stable (less energetic)

- its energy becomes a larger negative number.

- As the electron moves away from the nucleus (\( n \) increases), the atom's energy increases (becomes a smaller negative number)
The energy difference between any two levels:

\[ \Delta E = E_{\text{final}} - E_{\text{initial}} = -2.18 \times 10^{-18} \text{J} \left( \frac{1}{n_{\text{final}}^2} - \frac{1}{n_{\text{initial}}^2} \right) \]

We can predict the wavelengths of the spectral lines of H atom:

\[ \Delta E = h \nu = \frac{hc}{\lambda} = -2.18 \times 10^{-18} \text{J} \left( \frac{1}{n_{\text{final}}^2} - \frac{1}{n_{\text{initial}}^2} \right) \]

Therefore \((n_{\text{final}} = n_2, n_{\text{initial}} = n_1)\):

\[ \frac{1}{\lambda} = \frac{hc}{\Delta E} = \frac{(-2.18 \times 10^{-18} \text{J}) \left( \frac{1}{n_{\text{final}}^2} - \frac{1}{n_{\text{initial}}^2} \right)}{(6.626 \times 10^{-34} \text{J} \cdot \text{s}) (3 \times 10^8 \text{m/s}) \left( \frac{1}{n_{\text{final}}^2} - \frac{1}{n_{\text{initial}}^2} \right)} \]

\[ = -1.10 \times 10^7 \text{m}^{-1} \left( \frac{1}{n_{\text{final}}^2} - \frac{1}{n_{\text{initial}}^2} \right) \]

Energy needed to completely remove the electron from an H atom:

\[ \text{H(g)} \rightarrow \text{H}^+(\text{g}) + e^- \quad \Delta E = ? \]

\(n_{\text{final}} = \infty\) and \(n_{\text{initial}} = 1\), and obtain:

\[ \Delta E = E_{\text{final}} - E_{\text{initial}} = -2.18 \times 10^{-18} \text{J} \left( \frac{1}{\infty^2} - \frac{1}{1^2} \right) \]

\[ \rightarrow \Delta E = (-2.18 \times 10^{-18} \text{J})(0 - 1) = 2.18 \times 10^3 \text{kJ/mol} \]

\(\Delta E\) is positive because energy is absorbed to remove the electron from the vicinity of the nucleus.

For 1 mol of H atoms:

\[ \Delta E = (-2.18 \times 10^{-18} \text{J/atom}) (6.022 \times 10^{23} \text{atoms/mo1}) \left( \frac{1 \text{kJ}}{10^3 \text{J}} \right) \]

\[ \Delta E = 1.31 \times 10^{-3} \text{kJ/mol} \]

This is the ionization energy of the H atom, the energy required to form 1 mol of gaseous H\(^+\) ions from 1 mol of gaseous H atoms.
de Broglie Wavelength

- From $E=mc^2$ and $E=hc/\lambda$ ==> *de Broglie*: wavelength of any particles:
  \[ \lambda = \frac{h}{mu} \]

- Example: Find the de Broglie wavelength of an electron with a speed of $1.00 \times 10^6$ m/s (electron mass $= 9.11 \times 10^{-31}$ kg; $h=6.626 \times 10^{-34}$ kg.m$^2$/s)
  \[ \lambda = \frac{h}{mu} = \frac{6.626 \times 10^{-34} \text{ kg.m}^2/\text{s}}{(9.11 \times 10^{-31} \text{ kg})(1.00 \times 10^6 \text{ m/s})} = 7.27 \times 10^{-10} \text{ m} \]

The Heisenberg Uncertainty Principle

- Werner Heisenberg (1927) postulated the uncertainty principle.
  \[ \Delta x \Delta \mu \geq \frac{h}{4\pi} \]

- $\Delta x$: the uncertainty in position; $\Delta \mu$: the uncertainty in speed.

- Example: An electron moving near an atomic nucleus has a speed of $6 \times 10^6$ m/s ± 1%. What is uncertainty in its position ($\Delta x$)?

- Solution: Finding uncertainty in speed, $\Delta \mu$:
  \[ \Delta \mu = 1\% \text{ of } u = (0.01)(6 \times 10^6 \text{ m/s}) = 6 \times 10^4 \text{ m/s} \]

- Calculating the uncertainty in position, $\Delta x$:
  \[ \Delta x \Delta \mu \geq \frac{h}{4\pi} \]
  \[ \Delta x \geq \frac{h}{4\pi \Delta \mu} \geq \frac{6.626 \times 10^{-34} \text{ kg.m}^2/\text{s}}{4\pi (9.11 \times 10^{-31} \text{ kg})(6 \times 10^4 \text{ m/s})} \geq 1 \times 10^{-9} \text{ m} \]
Summary

- **Blackbody radiation** ==> Planck: Energy is quantized; only certain values allowed
- **Photoelectric effect** ==> Einstein: Light has particulate behavior (photons)
- **Atomic line spectra** ==> Bohr: Energy of atoms is quantized; photon emitted when electron changes orbit
- **de Broglie**: All matter travels in waves: energy of atom is quantized due to wave motion of electrons
- According to the **uncertainty principle**, we cannot know simultaneously the exact position and speed of an electron