Chapter 6
The structure of atoms

Light: its interaction with atoms has provided a lot of information

→ electromagnetic wave.
Wave: a periodic oscillation traveling in space
E.g. sound: wave of air density

\[
\text{speed } v \quad \text{wavelength, } \lambda \quad \text{period: time required for a complete cycle}
\]
\( \frac{1}{\text{period}} = \text{frequency}, \nu \) \( \left[ \text{s}^{-1} = \frac{1}{\text{Hz}} \right] \) Hertz

Speed: \( \nu = \frac{\text{wavelength}}{\text{period}} = \lambda \cdot \nu \)

\[ \nu = \lambda \cdot \nu \]

Greek letter "nu"

Speed of light:

\[ c = 3 \times 10^8 \text{ m/s} \]

(speed of sound in air: 340 m/s)
Convert between wavelengths & frequencies.

6x. Red light corresponds to
\[ \lambda = 6417 - 700 \text{ nm} \]
what frequencies correspond to this range?

\[ C = \lambda \cdot v \Rightarrow v = \frac{C}{\lambda} = \frac{3 \cdot 10^8 \text{ m/s}}{647 \cdot 10^{-9} \text{ m}} \]
\[ v = \frac{3 \cdot 10^8 \text{ m/s}}{700 \cdot 10^{-9} \text{ m}} = 4.6 \cdot 10^{14} \text{ s}^{-1} \]
\[ v = 4.3 \cdot 10^{14} \text{ s}^{-1} \]

So, red light corresponds to
\[ v = 4.3 - 4.6 \cdot 10^{14} \text{ s}^{-1} \]
(\text{Hz})
electro magnetic spectrum

\[ 10^{-16} \quad 10^{-14} \quad 10^{-6} \quad 10^{6} \]

\( f \text{ rays} \quad x \text{-rays} \quad \text{microwaves} \quad \text{Radio} \)

\( UV \quad 400 \text{nm} \quad 700 \text{nm} \quad \text{IR} \)

\( \text{violet} \quad \text{blue} \quad \text{green} \quad \text{yellow} \quad \text{infrared} \)

\( \text{ultraviolet} \)

Classical theory of waves does not work for light.

1) Black body radiation
2) Photoelectric effect
1) Black body: absorbs all frequencies. When heated, it emits radiation.

\[ I \text{ intensity} \]
\[ \text{wavelength} \]

- **Experimental**
- Classical theory

1900: Planck proposed Quantum theory:

energy is quantized (discrete values) & can be emitted in packets of \( h \cdot \nu \)

\( h \): Planck's constant
\( \nu \): frequency

\text{“quanta”}
2) Photoelectric effect:
Shining light on a metal causes emission of electrons. We measure their kinetic energy.

Energy of light = W + K.E.

\[ \text{energy that holds electrons in the metal} \]
\[ \text{Kinetic energy of emitted electron} \]

K.E. \( \frac{\text{exponent}}{\text{frequency of light}} \)
But in classical wave theory, the energy depends on the intensity, not the frequency.

1905: Einstein proposed:
Light consists of particles. Call photons.
Energy of a photon: $h \cdot \nu$

$$h \nu = W + K.E. \Rightarrow$$
$$\Rightarrow K.E. = h \nu - W$$

$\nu$ must be $> W$

Light: dual nature
Spectroscopy = study of the absorption & emission of radiation by matter.

Absorption spectrum: frequencies absorbed
Emission spectrum: frequencies emitted

E.g. Alkali metals in a flame
Na: 2 yellow lines
K: 2 red lines
Li: 1 red, 1 yellow
Spectrum of H

electrons collide with H₂, H goes to a higher energy state, then it drops back to the 'ground state' emitting radiation

This explained by the Bohr model (1913)

+ (++)

n=1, n=2, n=3

electrons move in "orbits". Energy increases from n=1 and up
- Normally, electron is in $n=1$
- It can absorb a photon and go to a higher orbit
- It can drop to a lower orbit by emitting a photon.

$$E_n = -\frac{2.179 \cdot 10^{-18}}{n^2} \text{ J}$$

\[ \begin{align*}
0 & \quad n=1 \\
\equiv & \quad n=2 \\
\equiv & \quad n=3 \\
\equiv & \quad n=4
\end{align*} \]

$$r = 0.529 \times n^2 \text{ Å}$$

\[ \text{Distance from nucleus} \]

(1 Å = 0.1 nm)

\[ \text{Ångstrom} \]
When an electron changes orbit

\[ \Delta E = E_{n_2} - E_{n_1} = -2.179 \times 10^{-18} J \left( \frac{1}{n_2^2} - \frac{1}{n_1^2} \right) \]

= \hbar \cdot \nu \quad \text{(energy of the absorbed or emitted photon)}

Set \ n_1 = 2

\[ n_2 = 3, 4, 5, 6 \]

Calculate \( \Delta E \) & then \( \hbar \nu \rightarrow \lambda \)

For \( n_2 = 3 \), we get \( \lambda = 656.3 \text{ nm} \)

\( n_2 = 4 \), \( \lambda = 486.1 \text{ nm} \)

\( n_2 = 5 \), \( \lambda = 434.0 \text{ nm} \)

\( n_2 = 6 \), \( \lambda = 410.2 \text{ nm} \)
An electron in H falls from the 4th to the 2nd energy level. What is the λ & E of the photon emitted?

\[ n_1 = 2 \]
\[ n_2 = 1 \]
\[ \Delta E = E_4 - E_2 \]
ΔE = -2.179 \times 10^{-19} \left( \frac{1}{\nu^2} - \frac{1}{\nu_c^2} \right)

= 4.086 \times 10^{-19} \text{ J}

= h \cdot ν

h = 6.626 \times 10^{-34} \text{ J s}

\Rightarrow λ = \frac{ν}{c} = \frac{4.086 \times 10^{19} \text{ J}}{6.626 \times 10^{-34} \text{ J s}} = 486.2 \text{ nm}

ΔE = hν

\nu must always be positive
Bohr model: excellent for it but not so good for other atoms.

Matter has a dual nature too. (particles & waves)

1924: de Broglie hypothesis

\[ \lambda = \frac{h}{m \cdot v} \]

Planck's constant

wavelength, mass, velocity

(most important for particles of low m)
Heisenberg Uncertainty Principle

\[ \Delta x \Delta p \geq \frac{\hbar}{4\pi} \]

uncertainty in position  uncertainty in momentum