Early Concepts of the Atom

Leucippus and Democritus (400 BC)
- All matter could only be subdivided so far
- Eventually have an indivisible particle
- Atomos (indivisible) becomes atom
- ‘Thought experiment’
  - No testing

John Dalton
- First evidence that matter is made up of discrete particles.
- Each chemical element is made up of indivisible particles called atoms, which are identical for that element but differ from the atoms of other elements.
- ‘Billiard Ball’ model
Early Concepts of the Atom

- J.J. Thomson
  - discovered electrons
  - CRT
    - Deflected by electric/magnetic fields
  - All electrons were determined to be identical.
    - Not dependent on material that the cathode was made from.
  - Number of electrons in an atom is equal to the atomic number.

Early Concepts of the Atom

- Led to the ‘plum pudding’ model
- Ernest Rutherford – 8 years later!
  - Gold foil experiment
  - Particles were knocked backwards!
  - Proposed the ‘nuclear’ atom
    - 99.9% of mass of the atom is contained in a small core called the nucleus.
    - If an atom was the size of the Astrodome, the nucleus would be about the size of a housefly!

Early Concepts of the Atom

- Learning Goal
  - Describe the atomic models of Dalton, Thomson, and Rutherford
- Questions: 1-3
The Dual Nature of Light

- Visible light of all frequencies is emitted from hot objects
  - ‘Black Body’ radiation
- Classical theory predicts that the intensity of the light should increase rapidly with an increase in the frequency of light.
- Not what happens!

The Dual Nature of Light

- Resolved by Max Planck
  - Energy is ‘quantized’
    - Starsteps
  - \( E = hf \)
  - Planck’s Constant (\( h \))
    - \( h = 6.626 \times 10^{-34} \) J s

The Photoelectric Effect

- Some metals will eject an electron when struck by light.
  - Time to eject an electron almost instantaneous
    - Too short according to classical theory
    - Requires ‘filling’ of energy bucket
  - Certain wavelengths do not eject, no matter how long you shine the light on them.
  - Bucket ‘should’ fill even if flow is slower
The Photoelectric Effect

- Einstein
  - Proposed light exists as ‘quanta’
    - Packets of energy
    - ‘Photons’
  - $E = hf$
    - $E$ is the energy of this photon!

Example

- Find the energy in Joules of the photons of red light of frequency $5.00 \times 10^{14}$ Hz.
- What do we know?
  - $E = hf$
  - $f = 5.00 \times 10^{14}$ s$^{-1}$
  - $h = 6.626 \times 10^{-34}$ Js
  - $E = (5.00 \times 10^{14})/(6.626 \times 10^{-34}) \text{ J}$
  - $E = 3.32 \times 10^{-19}$ J

The Dual Nature of Light

- Chapter 6
  - Light is a wave
- Chapter 9
  - Light is a particle
- ‘Sometimes you feel like a nut; sometimes you don’t!’
  - Travels as a wave
  - Interacts with matter as a particle!
The Dual Nature of Light

- Learning Goals
  - State Planck’s hypothesis and apply the equation for it.
  - Describe and explain the photoelectric effect.
  - Explain the meaning of the dual nature of light.
- Questions: 4-11
- Problems 1, 3

Bohr Theory of the Hydrogen Atom

- Light from an incandescent bulb analyzed we have a continuous set of wavelengths.

Bohr Theory of the Hydrogen Atom

- Light from a gas-discharge tube gives a series of bright lines
  - Line emission spectrum
Bohr Theory of the Hydrogen Atom

- Light of all wavelengths passed through a cool gas, loses some of the lines.
  - Correspond with emission lines

Bohr Atom

- Electrons in ‘quantized’ orbits
- Lines correspond with the energy to move from one orbit to another.
  - Emission spectra - lines from electrons falling from an excited state to a ground state.
  - Absorption spectra - light absorbed to cause electron to go from a ground state to an excited state.
Atomic Size

- Size of an atom depends on the ‘state’ of the electron
- **Ground State**
  - ‘natural’ state
  - \( n = 1 \)
- **Excited State**
  - Requires input of energy
  - \( n > 1 \)
  - \( r_n = 0.053nm \times n^2 \)

Example

- Calculate the size of a hydrogen atom when \( n = 1 \) and when \( n = 2 \)
- **When \( n = 1 \)**
  - \( 0.053 \times 1^2 = 0.053 \text{ nm} \)
- **When \( n = 2 \)**
  - \( 0.053 \text{ nm} \times 2^2 = 0.21 \text{ nm} \)

Bohr Atom
Electron Energies

- Orbit also determines energy of an electron
  - ‘Free electron’ has zero energy
  - Electrons closer to the nucleus have less energy than a free electron
  - Electron energies in orbits are therefore negative!
- \[ E_n = \left(-\frac{13.60}{n^2}\right) \text{ eV} \]

Example

- Determine the energy of an electron in the first and second orbits of a hydrogen atom
- For \( n = 1 \)
  - Energy = \(-13.60 / 1^2 = -13.60 \text{ eV} \)
- For \( n = 2 \)
  - Energy = \(-13.60 / 2^2 = -3.40 \text{ eV} \)

Emission and Absorption of Photons

- Photon Emission
- Photon Absorption
Electron Energies

- Conservation of Energy
  \[ E_{\text{photon}} = E_{ni} - E_{nf} \]
- Calculate the energy of the photon emitted when an electron falls from \( n = 2 \) to \( n = 1 \) orbit:
  - From previous Example Problem
    - \( n = 2; E = -3.40 \text{ eV} \)
    - \( n = 1; E = -13.60 \text{ eV} \)
    - \( E_{\text{photon}} = 3.40 - (-13.6) = 10.20 \text{ eV} \)

Hydrogen Line Spectra

- Balmer Series
  - Electrons fall from \( n > 2 \) to \( n=2 \)
  - These are the visible lines we saw in the hydrogen spectrum.
- Lyman Series
  - Electrons from \( n > 1 \) to \( n = 1 \)
  - UV
- Paschen Series
  - Electrons from \( n > 3 \) to \( n = 3 \)
  - Infrared
- Predicted by Bohr!
Bohr Theory of the Hydrogen Atom

- Learning Goals
  - Describe Bohr’s model of the hydrogen atom.
  - Explain the formation of line spectra.
- Questions: 12 – 21
- Problems: 5, 7, 9

Applications

- Spectroscopy
  - Study of interaction of light and matter
- Microwave Ovens
  - Quantized rotational / vibrational modes for water
  - Microwaves excite water molecules
  - De-excitation releases heat
  - Surface phenomenon
    - Must let sit for conductive heat transfer

Applications

- X-Rays
  - High energy electromagnetic radiation
  - Emitted by exciting ‘internal’ electrons
  - Interacts with "dense" material while ‘soft tissue’ is transparent
Applications

- Lasers
  - Light Amplification by Stimulated Emission of Radiation
  - Must get ‘metastable state’
    - Population inversion
    - More electrons in an excited state than a ground state
  - Spontaneous emission leads to stimulated emission!

Lasers

- Monochromatic Radiation
  - All light is of the same wavelength
  - Stimulated emission can only be the same energy as spontaneous emission!
  - ‘Coherent’ radiation
    - Light bulb ‘incoherent’
      - Short ‘wavelets’ with mix of phase relationships
    - Stimulated emission coherent
      - Same phase relationship and direction
      - More explanation in Chapter 7!
Applications of Lasers

- Distance measurement
  - Moon orbit
  - Tectonic plate movement
- Communication
  - Optical fibers
  - CD / DVD readers
- Surgery
  - ‘Traditional’ and eye surgeries
- Many other applications!

Learning Goals
- Describe and explain the operation of a microwave oven
- Tell how x-rays are produced, and explain their spectra
- Explain how laser light is produced.
- Questions: 22 - 26

Heisenberg's Uncertainty Principle
- Classical Mechanics
  - Measurements can be refined indefinitely with more reliable tools
- Quantum Mechanics
  - At a point, the measuring device changes the substance being measured!
  - Cannot measure 0K
    - Measuring device transfers heat to the system to cause the temp to be greater than 0K!
Heisenberg's Uncertainty Principle

- IT IS IMPOSSIBLE TO KNOW A PARTICLE’S EXACT POSITION AND VELOCITY SIMULTANEOUSLY!
  - Not very noticeable for very massive object.
  - Very important for electrons
  - $m \left( \Delta v \right) \left( \Delta x \right)$ is very close to Planck’s constant

Heisenberg’s Uncertainty Principle

- Learning Goal
  - Explain the significance of Heisenberg’s Uncertainty Principle.
- Questions: 27, 28

Matter Waves

- de Broglie
  - Hypothesized matter has wave/particle duality
  - wavelength = $h / mv$
    - $h$ = Planck’s constant ($6.63 \times 10^{-34}$ J s)
    - $m$ = mass in kg
    - $v$ = speed of the object
  - Only significant wavelengths for very light objects (electrons)
**Matter Waves**

- Confirmed experimentally in 1927
- Allows viewing of objects that would be impossible to resolve using visible light.

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**Matter Waves**

- Learning Goal
  - Explain the term dual nature of matter.
- Questions: 29 – 32
- Problem: 11

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**The Electron Cloud Model of the Atom**

- Bohr
  - Studied hydrogen because it is the simplest element
  - only one electron
- Multi-electron elements the mathematical model breaks down!
- Wave Mechanics.
The Electron Cloud Model of the Atom

- Erwin Schrodinger
  - Derived the ‘quantum mechanical’ model for the atom
  - Based on wave mechanics
  - Electrons exist somewhere in electron clouds.
  - Orbitals
    - Probability functions for the electrons
    - Three dimensional areas
    - Uncertainty (Heisenberg)

Wave Functions

- Helps to explain why an electron stays out of the nucleus
  - Wave function would collapse!

The Electron Cloud Model of the Atom

- Calculate ‘most probable radius’ for hydrogen
- Matches ‘Bohr radius’ for the atom
  - All energy levels match!
- Helps to explain multi-electron atoms also.
  - Doesn’t have as pretty of a picture as the ‘traditional’ picture of the atom
The Electron Cloud Model of the Atom

- Learning Goal
  - Describe the quantum mechanical model of the atom.
- Questions: 30, 31