Modern Atomic Theory

Chapter 10

10.1 Rutherford’s Atom

Rutherford showed:
- Atomic nucleus is composed of protons (positive) and neutrons (neutral).
- The nucleus is very small compared to the size of the entire atom.

Questions left unanswered:
- How are e’s arranged and how do they move?

10.2 Electromagnetic Radiation

- Electromagnetic radiation is radiant energy, both visible and invisible.
- Electromagnetic radiation given off by atoms when they have been excited by any form of energy
  - light bulbs
Electromagnetic radiation is radiant energy, both visible and invisible.

Electromagnetic radiation given off by atoms when they have been excited by any form of energy:
- light bulbs
- flame tests

Electromagnetic radiation travels in waves.

All waves are characterized by:
- Wavelength, \( \lambda \): distance between two consecutive peaks or troughs in a wave
  - measured in nanometers (1 nm = 10^{-9} m)
  - same distance for troughs

Wavelength, \( \lambda \): distance between two consecutive peaks or troughs in a wave.

Frequency, \( \nu \): the number of waves that pass a point in space in one second.

Hertz (Hz) = wave/sec

Hz = sec^{-1}
10.2 Electromagnetic Radiation

- Electromagnetic radiation travels in waves.
- All waves are characterized by:
  - Wavelength, \( \lambda \)
  - Frequency, \( \nu \)
  - Amplitude, \( A \): measure of the intensity of the wave, "brightness"
  - Height of the waves
  - Velocity, \( c \): speed of light
    \[ c = 2.997925 \times 10^8 \text{ m/s} \]
- All types of light energy travel at the same speed.

\[ c = \lambda \times \nu \]

10.2 Electromagnetic Radiation

- Radiowaves
- Microwaves
- Infrared (IR)
- Visible
- Ultraviolet (UV)
- X-rays
- Gamma rays

Light can have wave and particle-like nature.
- Photons: tiny particle-like packets of energy.
- Photons of light are quantized:
  - have fixed amounts of energy
  - \( E_{\text{photon}} \propto \nu \)
  - Higher frequency = More energy in photons.
10.3 Emission of Energy by Atoms

- Atoms which have gained extra energy release that energy in the form of light.

- The light given off or gained is of very specific wavelengths called a line spectrum.
  - Light given off = emission spectrum
  - Light energy gained = absorption spectrum

- Each element has its own line spectrum which can be used to identify it.
  - H₂ emission spectrum:
    - 410 nm, 434 nm, 486 nm, 656 nm

- All samples of an element give the exact same pattern of lines:
  - Every atom of that element must have identical energy states.

- The energy of atoms is quantized:
  - If the atom could have all possible energies, then the result would be a continuous spectrum instead of lines.
10.4 The Energy Levels of Hydrogen

- All samples of an element give the exact same pattern of lines.
- The energy of atoms is quantized.

10.5 Bohr’s Model of the Atom

- Energy of atom is related to the distance electrons are from the nucleus.
- Energy of the atom is quantized.
  - Atom can only have certain specific energy states called quantum levels or energy levels.
  - When atom gains energy, electron “moves” to a higher quantum level.
  - When atom loses energy, electron “moves” to a lower energy level.
  - Lines in spectrum correspond to the difference in energy between levels.

Ground state:

- The ground state of H corresponds to having its electron in an energy level that is closest to the nucleus.

Excited states:

- The farther an electron is from the nucleus, the higher its energy.
**10.5 Bohr’s Model of the Atom**

- To put an e⁻ in an excited state requires the addition of energy to the atom.
- Bringing the e⁻ back to the ground state releases energy in the form of light.
- Distances between energy levels decreases as the energy increases.
  - Light given off in a transition from the 2nd energy level to the 1st has a higher energy than light given off in a transition from the 3rd to the 2nd, etc.

**10.5 Bohr’s Model of the Atom**

- Electrons "orbit" the nucleus much like planets orbiting the sun.
- Each energy level can hold $2n^2$ e⁻.
  - 1st: 2 e⁻
  - 2nd: 8 e⁻
  - 3rd: 18 e⁻, etc.
- Farther from nucleus = more space = less repulsion.
- **Valence shell:**

**10.5 Bohr’s Model of the Atom**

- The problems with Bohr’s Model:
  - Only explains hydrogen atom spectrum and other 1 e⁻ systems.
  - Neglects interactions between electrons.
  - Assumes circular or elliptical orbits for e⁻, which is not true.
10.6 Wave Mechanical Model
- Experiments later showed that $e^-$ could be treated as waves
  - just as light energy could be treated as particles
- Schrödinger Wave Equation: uses wave mathematics to calculate probability densities of finding the $e^-$ in a particular region in the atom
  - can only be solved for simple systems, but approximated for others

10.7 Hydrogen Orbitals
- Orbitals:
  - usually use 90% probability to set the limit
  - three-dimensional
- Quantum numbers:

10.7 Hydrogen Orbitals
- Principal energy levels, $n$,
  - higher values mean orbital has higher energy
  - higher values mean orbital has farther average distance from the nucleus
10.7 Hydrogen Orbitals

- Each principal energy level contains one or more sublevels
  - there are \( n \) sublevels in each principal energy level
  - each type of sublevel has a different shape & energy
    - \( s < p < d < f \)

10.7 Hydrogen Orbitals

- Each sublevel contains one or more orbitals
  - \( s = 1 \) orbital
  - \( p = 3 \) orbitals
  - \( d = 5 \) orbitals
  - \( f = 7 \) orbitals

10.8 More Wave Mechanical Model

Pauli Exclusion Principle

- No orbital may have more than 2 \( e^- \)
- **Degenerate:**
  - each \( p \) sublevel has 3 degenerate \( p \) orbitals
  - each \( d \) sublevel has 5 degenerate \( d \) orbitals
  - each \( f \) sublevel has 7 degenerate \( f \) orbitals
  - \( s \) sublevel holds 2 \( e^- \)
  - \( p \) sublevel holds 6 \( e^- \)
  - \( d \) sublevel holds 10 \( e^- \)
  - \( f \) sublevel holds 14 \( e^- \)
- \( e^- \) in the same orbital must have opposite spins
10.9 Electron Arrangements

Electronic configuration

H: 1s^1 
He: 1s^2 
Li: 1s^2 2s^1

Orbital diagram:

H: \[ \uparrow \]
He: \[ \uparrow \downarrow \]
Li: \[ \uparrow \downarrow \] \[ \uparrow \downarrow \]

10.9 Electron Arrangements

Orbital diagrams:

H: 
He: 
Li: 
Be: 
B: 

Hund’s Rule:
10.9 Electron Arrangements

- Write the electronic configuration & orbital diagram for sodium and magnesium

- If we look at the orbital of the highest energy for each element, we see a periodic trend:

10.10 Periodic Trends
10.10 Periodic Trends

Principle energy level
- For s & p block, n = row number
- For d block, n = row number - 1
- For f block, n = row number - 2

10.10 Periodic Trends
- Write the electronic configuration of zinc

10.10 Periodic Trends
- Write the electronic configuration of palladium
10.10 Periodic Trends

- Write the electronic configuration of mercury

Valence shell:
- Core e\(^{-}\)s
- Noble gas core:
  - Ne: 1s\(^2\)2s\(^2\)2p\(^6\)
  - Mg: [Ne]3s\(^2\)

Write the short-hand elec. config. for:
- Manganese
- Iodine

10.10 Periodic Trends

- Elements in the same column on the Periodic Table have
  - Similar chemical and physical properties
  - Similar valence shell electron configurations
  - Same numbers of valence electrons
  - Same orbital types
  - Different energy levels
10.11 Atomic Properties

Metals
- malleable & ductile
- shiny, lustrous
- conduct heat & electricity
- most oxides basic & ionic
- form cations & oxidation

Metalloids
- also known as semi-metals
- show some metal & some nonmetal properties

Nonmetals
- solids are brittle
- dull
- electrical & thermal insulators
- most oxides are acidic & molecular
- form anions & polyatomic anions
- reduced rxns

10.11 Atomic Properties

Ionization Energy

- The lower the IE, the easier it is to remove the e-
- metals have low IE
- IE decreases down the group
- valence electron farther from nucleus

10.11 Atomic Properties

Ionization Energy:

- IE increases left to right across the period
  - Na: [Ne]3s¹ → Na⁺: [Ne] + e⁻
  - Cl: [Ne]3s²3p⁵ → Cl⁻: [Ne]3s²3p⁶ + e⁻
- Reactivity of metals increases to the left on the Period and down in the column
- follows ease of losing an e⁻
  - Na: [Ne]3s¹ → Na⁺: [Ne] + e⁻
  - Mg: [Ne]3s² → Mg²⁺: [Ne]3s² + e⁻ → Mg²⁺: [Ne] + 2e⁻
10.11 Atomic Properties

Periodic trend in atomic size
- Increases down column
  - valence shell farther from nucleus
- Decreases left to right across period
  - adding electrons to same valence shell
  - valence shell held closer because more protons in nucleus

10.11 Atomic Properties

Electron Affinity:
- Reactivity of nonmetals (excluding the noble gases) increases to the right on the Period and up in the column
  - Follows ease of attracting \( e^- \)
    - \( \text{Cl}: [\text{Ne}]3s^23p^5 + e^- \rightarrow [\text{Ne}]3s^23p^6 \rightarrow \text{Cl}^- : [\text{Ar}] \)
    - \( \text{S}: [\text{Ne}]3s^23p^5 + e^- \rightarrow [\text{Ne}]3s^23p^6 \)

Diagram showing the periodic table and electron configurations.