CHAPTER 11

MODERN ATOMIC THEORY
**DALTON’S ATOMS**

- No internal structure
- Hard spheres – like marbles or billiard balls
- What holds them together?
THOMSON MODEL

- “Plum Pudding Model”
  - Chocolate Chip Cookie Dough Model

![Image of Thomson with diagram showing spherical cloud of positive charge and electrons.](image.png)
RUTHERFORD’S MODEL

- Dense, positive nucleus
- Electrons outside of nucleus
- Is this stable?
  - Why don’t electrons fall into the nucleus?
ATOMS AND LIGHT

• Elements give off certain colors of light when energized
• Spectroscopy = breaking down this light into its colors
  • Emission spectrum or bright line spectrum
• ID elements + Discover new elements
• Ex. Hydrogen

• Perhaps there is a connection between the colors of light emitted and the structure of the atom
FLAME TESTS

- Heating salts in a burner flame causes the metal ions to color the flame = flame test
  - DEMO: flame tests
- Can also be done with spectrum tubes
  - Glass tubes filled with a gas
  - Gas glows when tube is electrified
  - DEMO: spectrum tubes
- Light viewed through a spectroscope to see emission spectrum
  - DEMO: diffraction grating
EMISSION SPECTRA

Bright Line Spectra of Helium and Neon
BUILD-YOUR-OWN SPECTROSCOPE

• Tomorrow you will build your own spectroscope to use in lab and take home
• Constructed of cardboard and tape
• Diffraction grating on one end; small, narrow opening on the other end
WAVE BASICS

- Wavelength = length of 1 wave cycle
- Frequency = # waves per second (unit = Hertz)
- Amplitude = height of the wave, related to energy
• Light is a form of electromagnetic radiation – part of the electromagnetic spectrum
• Red light has longer waves with lower frequency
• Violet light has shorter waves with higher frequency
PHOTOELECTRIC EFFECT

• 1887 Heinrich Hertz discovers the photoelectric effect
• Hertz found that whether or not electrons were ejected depended on the frequency of the light and not the amplitude.
  • Blue light works but red light does not
PHOTOELECTRIC EFFECT

• In 1905 Albert Einstein explained the photoelectric effect
  • Wins Nobel Prize for this
• EM radiation is a stream of particles called photons
• EM radiation of different frequencies have photons with different amounts of energy
LIGHT AS PARTICLES

- Instead of waves, you can think of light as a stream of particles – photons
- The higher the frequency – the greater the energy of the photons
- Energy of photons demonstration
  - Which color has greater energy?
THE BOHR MODEL

- Electrons occupy specific orbits around the nucleus.
- The greater the energy of the electron, the larger the orbit.
BOHR MODEL

- Electrons can absorb energy and move to an “excited” state
- When they return to the “ground” state – energy is released in the form of photons
- Frequency (color) of photons depends on the amount of energy
  - Greater energy = higher frequency
BOHR MODEL

• The energy changes are not continuous but quantized
  • Not all energies are possible – only specific energies are available
  • Ex: a ramp is a continuous change in height but a staircase is quantized

• Each element has its own specific energy changes and therefore produces its own specific colors of light in its emission spectrum
BOHR MODEL

- Only works for hydrogen
- Exact path of electrons is unknown
- Concept of quantized energy levels is useful but need to be modified to work with all elements
- Modern atomic theory – move away from physical model towards a mathematical model
WAVE MECHANICAL MODEL

- Louis Victor de Broglie (France) and Erwin Schrödinger (Austria)
- If light can behave as a wave and a particle, perhaps an electron can act like a particle and a wave
The double slit experiment shows that electrons can behave as both a particle and a wave.
HEISENBERG UNCERTAINTY

- TED on Heisenberg Uncertainty Principle
WAVE MECHANICAL MODEL

• Wave mechanical model works for hydrogen as well as the other elements
• Electrons are described as being found in orbitals not orbits
Instead of orbits, we use orbitals to describe the behavior of electrons. Orbitals are nothing like orbits – no set path. Orbitals are probability maps or graphs that show where the electron is likely to be found. To understand what an orbital is, think about the following experiment.
THE FIREFLY EXPERIMENT

- Nectar (food) suspended in the middle of a completely dark room
- Camera in the corner with the shutter open
- Every time the firefly lights up, it shows up on the picture
- We will see where the firefly is when it lights up, but not the path that it travels
- What would the picture look like after an hour has elapsed?
  - Where would you expect to find the firefly most frequently?
THE FIREFLY EXPERIMENT

• Is this what you would expect?
ELECTRON ORBITALS

- In modern atomic theory, we plot the likely location of the electron rather than its exact path.
- An electron orbital is the graph that represents the area in which the electron is likely to be found 90% of the time.
- Orbitals have various shapes, sizes, and orientations.
- The simplest is the hydrogen 1s orbital – it is a sphere.
ENERGY LEVELS

• Recall that Bohr explained emission spectra by describing the electron jumping to and falling from different energy levels.
• In the wave mechanical model, principal energy level describes the relative size of the orbital.
• Electrons in higher energy levels have a greater chance to be farther from the nucleus.
The principal energy level tells us the relative size of the orbitals – the higher the level, the larger the orbital.
Principal energy levels are broken down into sublevels.
SUBLEVELS AND ORBITALS

• The sublevel (s, p, d, or f) tells us the shape of the orbitals that make up the sublevel and how many orbitals are available
  • s sublevels have 1 orbital
  • p sublevels have 3 orbitals
  • d sublevels have 5 orbitals
  • f sublevels have 7 orbitals

• The p orbitals

• The d orbitals
ELECTRONS IN ORBITALS

• The orbitals are where the electrons are found.
• Each orbital can hold up to 2 electrons.
  • An s sublevel has 1 orbital and can hold 2 electrons.
  • A p sublevel has 3 orbitals and can hold 6 electrons.
  • A d sublevel has 5 orbitals and can hold 10 electrons.
  • An f sublevel has 7 orbitals and can hold 14 electrons.
BOX (ORBITAL) DIAGRAM

- Sets of boxes represent the orbitals found in a sublevel
- # of boxes = # of orbitals

- Arrows represent electrons
BOX (ORBITAL) DIAGRAM

- There are 3 rules that need to be followed when filling in the boxes
  - Aufbau Principle
  - Pauli Exclusion Principle
  - Hund’s Rule
Aufbau principle: the lowest energy orbitals fill first until all electrons are placed: 1s, 2s, 2p, 3s, 3p...

After 3p, the pattern is not as simple

There is a simple diagram to help figure out the pattern known as the diagonal rule
DIAGONAL RULE

- This diagram gives the order in which sublevels fill
- Start at the top and follow each arrow
- The pattern repeats itself for each successive element
PAULI EXCLUSION PRINCIPLE

• Pauli exclusion principle: no two electrons can have the same set of quantum numbers – electrons in the same orbital must have opposite spin

• Example:
HUND’S RULE

- Hund's rule: electrons are placed into orbitals so that unpaired electrons are maximized – place 1 electron in each orbital (box) until all are half-filled then double up

- Example:
EXAMPLES

- Hydrogen
  ![Hydrogen](image)

- Helium
  ![Helium](image)

- Lithium
  ![Lithium](image)

- Oxygen
  ![Oxygen](image)

- Chlorine
  ![Chlorine](image)
MORE EXAMPLES

- Calcium

- Iron

- Look for unpaired electrons in the last sublevel to fill
EVEN MORE EXAMPLES

• Arsenic

• Osmium
UNPAIRED ELECTRONS

- Electrons that are alone in a box (orbital) are referred to as unpaired electrons
- Only found at the end
ELECTRON CONFIGURATIONS

• Electron configuration is like the box diagram without the boxes and arrows
• Energy levels and sublevels are listed and superscripts are used to show the number of electrons in each sublevel
• Ex. Chlorine = 1s^2 2s^2 2p^6 3s^2 3p^5
• What is the electron configuration for strontium?
  • 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2
VALENCE ELECTRONS

• Valence electrons = the outermost electrons; the electrons in the highest principal energy level
• These electrons that are most responsible for an element’s chemical properties
• Look for the highest principal energy level number and add up the electrons that are in that level
• In what sublevels will the valence electrons be found?
  • Only electrons in s and p can be valence electrons
EXAMPLES

• Write the electron configuration for silicon and tell how many valence electrons and unpaired electrons it contains
  • e. config. = 1s²2s²2p⁶3s²3p²
  • Valence e. = 4
  • Unpaired e. = 2

• Write the electron configuration for palladium and tell how many valence electrons and unpaired electrons it contains
  • e. config. = 1s²2s²2p⁶3s²3p⁶4s²3d¹⁰4p⁶5s²4d⁸
  • Valence e. = 2
  • Unpaired e. = 2
PERIODIC TABLE

- The order of the elements is based on # of protons (atomic #)
- The arrangement of the elements is based on electron configurations
- The periodic table can be broken down into blocks based on the orbitals that are being filled
BLOCKS ON THE PERIODIC TABLE
BLOCKS ON THE PERIODIC TABLE

• Principal energy level of s and p is given by the period number
• d sublevels lag behind by 1 (Ex. 4s → 3d)
• f sublevels lag behind by 2 (Ex. 6s → 4f)
• Electron configurations can be read right off of the periodic table
EXAMPLES

- Carbon
  - $1s^22s^22p^2$
  - 4 valence
  - 2 unpaired
- Manganese
  - $1s^22s^22p^63s^23p^64s^23d^5$
  - 2 valence
  - 5 unpaired
EXAMPLES

- Gold
  - $1s^22s^22p^63s^23p^6$
  - $4s^23d^{10}4p^65s^2$
  - $4d^{10}5p^66s^24f^{14}$
  - $5d^9$
- 2 valence
- 1 unpaired
ABBREVIATED CONFIGURATIONS

• Valence electrons play the most important role in the behavior of an element
• Valence electrons max out in noble gases
• Abbreviated electron configurations use the noble gas from the row above [in brackets] to represent the inner electrons
• Count forward from the noble gas to the desired element
• Example: aluminum = [Ne]3s\(^2\)3p\(^1\)
• Example: tin = [Kr]5s\(^2\)4d\(^{10}\)5p\(^2\)
• Can still determine valence and unpaired electrons
SAMPLE PROBLEM

• Give the abbreviated electron configuration for molybdenum and tell how many valence electrons and unpaired electrons it has.
  - [Kr]5s\(^2\)4d\(^4\)
  - 2 valence electrons; 4 unpaired electrons
SAMPLE PROBLEM

- Give the abbreviated electron configuration for francium and tell how many valence electrons and unpaired electrons it has.
  - \([\text{Rn}]7s^1\)
  - 1 valence electron; 1 unpaired electron
SAMPLE PROBLEM

- What element is represented by the abbreviated electron configuration [Ar]\(4s^2\)3d\(^{10}\)4p\(^4\)

- Selenium
ONE MORE SAMPLE PROBLEM

- Give the abbreviated electron configuration for europium and tell how many valence electrons and unpaired electrons it has.
  - \([\text{Xe}]6s^24f^7\)
  - 2 valence electrons
  - 7 unpaired electrons
A dot diagram is used to represent the valence electrons of an element.

The element symbol represents the nucleus and the core (inner) electrons.

Dots around the symbol represent the valence electrons.

One side represents s, the other 3 sides represent p.

Example: sodium = [Ne]3s¹
  - 1 valence electron
  - Dot diagram:
**DOT DIAGRAM**

- Ex. Iron = [Ar]\(4s^23d^6\) = 2 valence electrons
  - Dot diagram: 

- Ex. Indium = [Kr]\(5s^24d^{10}5p^1\) = 3 valence electrons
  - Dot diagram: 

- Ex. Phosphorus = [Ne]\(3s^23p^3\) = 5 valence electrons
  - Dot diagram: 

- Ex. Oxygen = [He]\(2s^22p^4\) = 6 valence electrons
  - Dot diagram: 

OCTET RULE

• Maximum valence electrons = 8
• This is particularly stable – Ex. Noble Gases
• Octet rule: atoms will gain, lose, or share valence electrons in order to acquire a full set of 8 valence electrons
• Explains why metals lose electrons to form ions
  • Octet < half full – easier to lose a few than gain many
• And nonmetals gain electrons to form ions
  • Octet > half full – easier to gain a few than lose many
• And noble gases don’t lose, gain, or share
  • Full octet (except helium)
PERIODIC TRENDS

- Elements originally organized by similar properties
  - Mendeleev
- Modern periodic table retains this feature
  - Organized by electron configurations
- Elements in same group have similar configurations
- Periodic trend refers to pattern found within the table
- Ex. Valence electrons/charge of ions formed
- Ex. Atomic radius = size of atoms
  - Larger as you go down a column – Why?
  - Smaller as you go across a period – Why?
  - Lower Left Large
PERIODIC TRENDS

• Ex. Ionization energy (IE) = energy required to remove an electron
  • IE increases as you move across a period – Why?
  • IE decreases as you go down a column – Why?
  • Trend for IE is exactly opposite atomic radius
SAMPLE PROBLEMS

- Rank the following from smallest to largest: Cu, Si, Rb, O, K
  - Answer: O, Si, Cu, K, Rb

- Rank the following in order from lowest ionization energy to highest: Zr, Ne, Kr, Sr, Se
  - Answer: Sr, Zr, Se, Kr, Ne