1 Atomic Structure and Energy of Electrons

1.1 Atomic Theory & Structure

Theories vs. Laws

- A theory is an explanation based on many observations.
- A law is a fact of nature that is observed so often it is accepted as truth.
- Theories EXPLAIN laws
- Both a scientific theory and scientific law are accepted to be true by the scientific community as a whole.
- A theory is like a car. Components of it can be changed or improved upon, without changing the overall truth of the theory as a whole.

What is atomic theory?

- The idea that matter is made up of atoms, the smallest pieces of matter.
- Over the years, atomic theory has evolved and changed to better explain scientific observations about atoms.
- Ancient Greeks believed all matter was made up of four basic elements: fire, earth, water and air.
- Democritus
  - Greek philosopher
  - Idea of 'democracy'
  - Idea of 'atomos'
    - Atomos = 'indivisible'
    - 'Atom' is derived
  - No experiments to support idea

Democritus's model of the atom consisted of a solid and indestructable atom with no protons, electrons, or neutrons.

- Lavoisier 18th century
- Proposed the law of conservation of mass/matter.
- Observed that the mass of the reactants equaled the mass of the products in a chemical reaction.
- Proust - Proposed the law of definite proportions for compounds.

Dalton's Atomic Theory

- All matter is made of tiny indivisible particles called atoms.
- Atoms of the same element are identical, those of different elements are different.
- Atoms of different elements combine in whole number ratios to form compounds.
- Chemical reactions involve the rearrangement of atoms. No new atoms are created or destroyed.

Thomson

- J.J. Thomson - English physicist. 1897

- Made a piece of equipment called a cathode ray tube. It is a vacuum tube - all the air has been pumped out.
- Thomson's Model - Plum Pudding Model (also called Chocolate Chip Cookie Model)
  - Atoms are composed of charged particles (subatomic particles).
  - The particles that were attracted to the positive plate were negative.
    - These were called "electrons"
    - Protons were discovered the same way.

Rutherford 1895

- Experiment: Gold Foil Experiment
- Most particles pass through, but some are bounced back towards the source.
- Model: Rutherford explained that atoms must be mostly empty space with a small, concentrated center of positive charge.

Chadwick

- Discovered the neutron.
  - Neutron is a subatomic particle roughly the size of a proton (large compared to electrons).

Bohr

- Model: proposed the "electron cloud" in which electrons orbit at a given distance from the nucleus.
- Small orbits = low energy
- Big orbits = high energy

Quantum Mechanical Model
Modern atomic theory describes the electronic structure of the atom as the probability of finding electrons within certain regions of space (orbitals).

Modern View

- The atom is mostly empty space
- Two regions
  - Nucleus
    - protons and neutrons
  - Electron cloud
    - region where you might find an electron

1.2 Structure of Atom & Isotopes

Major Parts of the Atom

- Nucleus: dense, central part of the atom
- Protons and neutrons are found in the nucleus
- Electron cloud: large area outside of the nucleus
- Electrons occupy the electron cloud

Protons are located in the nucleus with positive charge and have a large relative size.

Nuetrons are located in the nucleus with 0 charge and with a large relative size.

Electrons are located in the electron cloud, have a negative charge and have a tiny relative size.

Atoms and the Periodic Table

- Atomic Number - the whole number in an element's box on the periodic table.
  - Atomic # = # protons = # electrons
  - The atomic number determines an element's identity!

Exercise - An atom has 24 protons. What element is it? (chromium)

- Mass Number - the sum of the protons and neutrons
- This number isn't on the periodic table, because the number of neutrons can vary (these are called isotopes)
- Atomic Mass - the decimal number on the periodic table. The weighted average mass of all isotopes of that element.
- Isotopes - atoms of the same element that have different mass numbers.
- This means the number of protons is the same, and the number of neutrons is different.

Isotopes of Hydrogen

- Protium - 1 proton, 1 electron, mass number of 1
- Deuterium - 1 proton, 1 neutron, 1 electron, mass number of 2
- Tritium - 1 proton, 2 neutrons, 1 electron, mass number of 3

How to write isotopes

- Method 1: Subscript/Superscript Method
- The atomic # is your subscript (below) and the mass # is the superscript (above), both on the left side of the symbol
- Method 2: Hyphen-notation method
- This symbol is written, then hyphen, then mass #

Exercise - given ruthenium and the super/sub method of $^{101}_{4}4$Ru, write the atomic number, mass number, number of protons, neutrons, and electrons and the hyphen method for this element.

Answer: atomic number - 44, mass number - 101, protons - 44, neutrons - 57, electrons - 44, hyphen method - Ru-101

1.3 Average Atomic Mass

- Atoms can't be easily measured in grams because they are so small.
- Scientists devised "atomic mass units" - a carbon-12 isotope is 12.000000 amu's.

Average Atomic Mass

- A different kind of average - a "weighted" average.
- This means that we take into account the abundance of each isotope found in nature.

Formula to memorize:

$$[(\text{mass})(\text{abundance})+(\text{mass})(\text{abundance})+(\text{mass})(\text{abundance})]/100.000$$

- That's for 3 isotopes. Use the (mass)(abundance) for as many isotopes as there are.
- The 100 won't limit sig figs in your answer. Your answer is limited by whichever mass or abundance has the fewest sig figs.

Exercise - Argon has three isotopes with the following percent abundances: Ar-36 with a mass of 35.968 amu and an abundance of 0.3337%. Ar-38 with a mass of 37.963 amu and an abundance of 0.063%. Ar-40 with a mass of 39.962 amu and an abundance of 99.600%. Calculate the average atomic mass. (40. amu)

Exercise - The atomic weight of gallium is 69.72 amu. The masses of naturally occurring isotopes are 68.92 amu for Ga-69 and 70.92 amu for Ga-71. Calculate the percent abundance of each isotope. (Ga-71: 40%, Ga-69: 60%)

1.4 Moles

- A mole is the amount of substance that contains the same number of atoms as 12 grams of Carbon-12.
- It is a counting unit just like a dozen.
- A mole is 6.02×10^{23} of something.
- 6.02×10^{23} is called "Avogadro's Number" because Amedeo Avogadro discovered it.
- 1 mole of any element has a mass (in grams) equal to its average atomic mass.

Exercise - 1 mole of potassium has a mass of _____ g. (39.10)

Molar Mass

- When we write out the average atomic mass in "grams" we call this the molar mass - it is literally the mass of one mole.

Exercise - What is the molar mass of fluorine? (19.00 grams)

Conversions

1.0000 mole of any substance equals 6.02×10^{23} atoms of that element equals molar mass in grams of that element.

To do a molar conversion problem:

- Do dimensional analysis.
- Start with what you're given.
- Bring that unit down and over.
- Put the unit you want on top.
- Fill in the numbers.
  - Put a "1" in front of moles in a conversion.
  - Put "6.022×10^{23}" in front of atoms in a conversion.
  - Put the molar mass in front of grams in a conversion.

Exercise - How many atoms are in 55.4 grams of lithium? (4.81$\times 10^{24}$ atoms)

Exercise - What is the mass in grams of 3.011×10^{23} atoms of iron? (27.93 g Fe)

Exercise - How many atoms are in 8.43 moles of nickel? (5.07×10^{24} atoms Ni)

Exercise - How many atoms are in 1.00×10^{-10} grams of gold? (3.06×10^{11} atom Au)

1.5 Electron Configuration

Energy Level

- The region surrounding the nucleus where an electron is likely to be found.
- Think of rungs on a ladder, fixed levels with space in between.
- Sublevel - smaller part of an energy level indicated by letters (1s, 2s, 4d, etc.)
- Orbital - smaller part of a sublevel, each orbital holds 2 electrons, moving in opposite direction... (4 possible shapes)
- "Electron configuration" describes the location of electrons in a given atom. This determines how an element behaves chemically, and thus is the core of chemistry.

We'll learn three ways to show electron configuration

- Orbital Notation
- Electron Configuration
- Lewis Dot Structures

Aufbau Principle - electrons enter orbitals of lowest energy first. Low energy orbitals are closer to the nucleus.

Pauli Exclusion Principle - no two electrons can be in the same orbital moving the same way. Each electron is unique.

Hund's Rule - when electrons are filling up orbitals of equal energy (say for instance 3 orbitals, which is 6 electrons), one electron enters each orbital until they're half-filled with electrons spinning in the same direction, then they fill with the opposite-spin electrons

Orbital Notation

- Numbers represent energy levels and letters represent sublevels
- Lines represent 1 orbital each (can also use boxes)

Electron configuration notation

- Write coefficient & letter for each energy sublevel.
- Superscript (number on top) shows # of electrons at that sublevel.
- This method simply takes less space.

Shorthand Notation

- If you had to show the electron configuration for bismuth, it would be long. There is a way to shorten what you have to write.
- Use the symbol for the noble gas before the element you are using and put it in brackets. That represents all the electrons up until that noble gas. Then continue with the rest of the electron configuration.

f-block issues

- Period 6
  - f-block includes elements La to Yb
  - d-block includes elements Lu to Hg
- Period 7
  - f-block includes Ac to No
  - d-block includes Lr to Uub

Lewis Dot Notation

- Lewis Dot diagrams show electrons available for bonding. These are the outermost electrons (valence electrons).
- Valence electrons are the total electrons in the last energy level (highest coefficient).
- Notice that electrons do not pair up until all four sides have one electron already.

Exercise - Halogens have how many valence electrons? (7)

Exercise - Copper is part of which block? (d)

Exercise - Which group contains the alkaline earth metals? (2)

Exercise - Which block do the lanthanide and actinide series belong to? (f)

1.6 Ion Electron Configurations

- How do positive ions (cations) form? Atoms (typically metals) lose electrons.
- How to negative ions (anions) form? Atoms (typically nonmetals) gain electrons.

When representative elements (s and p block) become ions, they take on the electron configuration of the nearest noble gas. This gives them 8 valence electrons.

Exercise - Write the electron configuration for the nitrogen atom ($1s^2 2s^2 2p^3$)

Exercise - Which noble gas does the nitrogen ion mimic? (neon)

- Transition metals (Groups 3-12) often have variable charges
- Use these guidelines to help figure out their electron configuration
  - Transition elements usually lose their s and p electrons first.
  - Completely full, half-full, or empty sublevels are stable.
  - Electrons can move from s-sublevels to d-sublevels if it makes the atom more stable.

Exercise - What is the electron configuration for a copper atom?

- cuprous, Cu^+ - $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$
- cupric, Cu^{2+} - $1s^2 2s^2 2p^6 3s^2 3p^6 3d^9$

Memorizing Monatomic Ions

- Monatomic cations - attach “ion” to the element name
- Monatomic anions - change the element ending to “-ide”
- The systematic name just uses a Roman numeral to indicate the charge. Used for transition metals (variably charged)

Memorizing Polyatomic Ions

- Help with formulas
  - Does the polyatomic ion contain an element in “the elbow”?
  - If so, the ion “-ate” 3 oxygen atoms
  - If not, the ion “-ate” 4 oxygen atoms
  - “-ite” ions contain 1 less oxygen atom than “-ate”
  - “hypo-x-ite” ions have 1 less oxygen atom than “-ite”
  - “per-x-ate” ions have 1 more oxygen atom than “-ate”
- Help with charges
  - There are only two polyatomic cations; the rest are anions
  - If the polyatomic ion contains oxygen, look at what group the other element is in
    - If it’s an even # group, the ion charge is even
    - If it’s an odd # group, the ion charge is odd

1.7 EM Spectrum

The Wave-Particle Theory

- A theory that attempts to explain how electrons can behave in two different ways
  - as waves (like light)
  - as particles (like a ball)

First, we will look at wave behavior.

Light consists of electromagnetic waves that travel 3.00×10^8 m/s.

- That's 670,616,629 miles per hour!
- This is the "speed of light", also known as "c"

Electromagnetic Waves

- The electromagnetic (EM) spectrum is a series of waves that have different wavelengths.
- Visible light is small portion of the EM spectrum, with mid-energy.
- EM waves are also called radiation.

EM Wave Characteristics

- Amplitude - height from origin to crest
- Frequency - number of cycles that pass a given point in a given amount of time
  - Measured in Hertz (Hz)
  - 1 Hz = 1 wave passes per second
  - 1 Hz = 1/s = s^{-1}
  - Symbol is nu, ν
- Wavelength - distance between crests of a wave
  - Symbol is lambda, λ

All EM Waves move at the speed of light

$$c = \lambda\nu$$

As wavelength increases, frequency decreases. They are inversely proportional.

Exercise - What is the wavelength of a wave with a frequency of 7600 Hz? (39000 m)

Important conversions:

- 1 m = 1×10^9 nm
- 1 MHz = 1×10^6 Hz

Exercise - What is the frequency of a wave with a wavelength of 467 nm? (6.42×10^{14} Hz)

- The visible spectrum of ROYGBIV is continuous; there are no breaks and the colors blend together.
- White light is a combination of all colors of light. A prism breaks up white light into the separate colors so we can see them.
- Each color has a definite frequency and wavelength.
  - The speed these colors of light are traveling never changes; it's always the speed of light, c

Low energy colors have a long wavelength and low frequency, while high energy colors have a short wavelength and high frequency.

- Remember that electrons occupy energy levels.
- When electrons are in the lowest energy level, they are said to be in their ground state
- It is possible for electrons to jump from ground state to a higher energy level (called excited state) by absorbing energy.
- When electrons lose energy they will fall back down to their ground state and release energy, and some of it is released as waves we can see - LIGHT!
- With many electrons jumping to energy levels and falling back, many different shades of light are released and blended.
- We can use a prism to separate the light to see the individual shades.

- This is called an atomic emission spectrum.

Types of Spectra

- Continuous Spectrum - no breaks
- Atomic Emission Spectrum - a lot of black space, aka "bright line" spectrum
- Absorption Spectrum - small dark regions, aka "dark line" spectrum

Spectroscopy is the science of producing atomic spectra and studying them.

Particle Model

- The idea that light can act as a particle
- Particles of light are called photons, or quanta (plural for quantum)
- A quantum behaves like a particle, and can move other matter

The Photoelectric Effect

- The particle model was needed to explain why when you shine a high energy light on some metals, electrons are ejected (moved) from the metal
- Einstein proposed in 1905 that light can behave as both a wave and a particle.
- He defined a photon as a particle of electromagnetic radiation with no mass that carries a quantum of energy.
- For this, he won the Nobel Prize.
- The energy contained in a photon (a quantum) depends on its frequency

$$E_{\text{photon}} = h\nu$$

  $E =$ energy in joules [J]
  $h =$ Planck's constnat $= 6.626 \times 10^{-34}$ J·s $\nu =$ frequency (nu), [Hz]

- According to Planck, matter can emit or absorb energy only in whole quanta ($1h\nu$, $2h\nu$, etc.)

Exercise - Calculate the frequency of a photon with 7.2×10^{-34} J of energy. (1.1 Hz)

Exercise - Calculate the wavelength of a photon with 5.32×10^{-33} J of energy. (3.74×10^{7} m)