Unit 1 – Atomic Structure

4.1 Defining the Atom

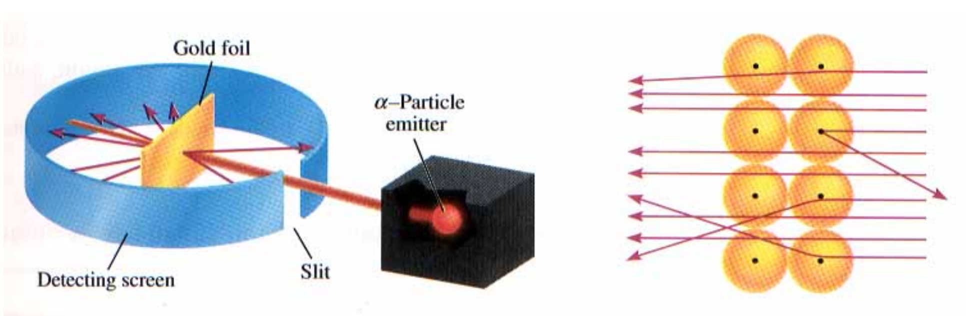
I. Atomic Theory
- A. Modern Atomic Theory
  1. All matter is made up of very tiny particles called atoms
  2. Atoms of the same element are chemically alike
  3. Individual atoms of an element may not all have the same mass. However, the atoms of an element have a definite average mass that is characteristic of the element
  4. Atoms of different elements have different average masses
  5. Atoms are not subdivided, created, or destroyed in chemical rxns

II. Sizes of Atoms
- A. Atomic radius
  1. 40 to 270 picometers (pm)
     - a. 1 pm = 10^{-12}m
  2. Most of the atomic radius is due to the electron cloud
- B. Nuclear radius
  1. 0.001 pm
  2. density is $2x10^8$ metric tons/cm^3
     - a. 1 metric ton = 1000kg

4.2 Structure of the Nuclear Atom

I. The Electron
- A. Discovery
  1. Joseph John Thomson (1897)
     - a. Cathode ray tube produces a ray with a constant charge to mass ratio
     - b. All cathode rays are composed of identical negatively charged particles (electrons)
- B. Inferences from the properties of electrons
  1. Atoms are neutral, so there must be positive charges to balance the negatives
  2. Electrons have little mass, so atoms must contain other particles that account for most of the mass

II. The Nucleus
- A. The Rutherford Experiment (1911)
  1. Alpha particles (helium nuclei) fired at a thin sheet of gold
     - a. Assumed that the positively charged particles were bounced back if they approached a positively charged atomic nucleus head-on (Like charges repel one another)



  2. Very few particles were greatly deflected back from the gold sheet
     - a. nucleus is very small, dense and positively charged
     - b. most of the atom is empty space

III. Structure of the Nucleus

1. Protons
    a. Positive charge, mass of 1.673×10^{-27} kg
    b. The number of protons in the nucleus determines the atom's identity and is called the atomic number
2. Neutrons
    a. No charge, mass of 1.675×10^{-27} kg
3. Nuclear Forces
    a. Short range attractive forces:
        neutron-to-neutron, proton-to-proton, proton-to-neutron

Particle	Symbols	Relative charge	Mass Number
Electron	e^- ${}^{0}_{-1}e$	-1	0
Proton	p^+ ${}^{1}_{1}\mathrm{H}$	+1	1
Neutron	n° ${}^{1}_{0}n$	0	1

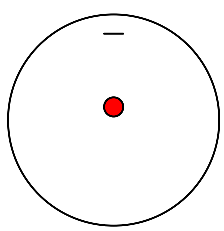
4.3 Distinguishing Among Atoms

I. Atomic Number, Mass Number, and Isotopes

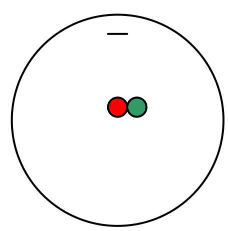
A. Atomic Number (Z)
1. The number of protons in the nucleus of each atom of that element
2. Atoms are identified by their atomic number
3. Because atoms are neutral,
    \# protons = # electrons
4. Periodic Table is in order of increasing atomic number

B. Mass Number
1. The total number of protons and neutrons in the nucleus of an isotope

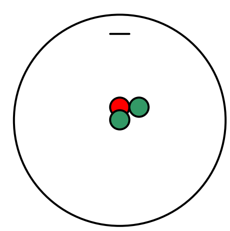
C. Isotopes
1. Atoms of the same element that have different masses
2. All elements of the same element have the same # of protons, but may vary in the number of neutrons
3. Although isotopes have different masses, they do not differ significantly in their chemical behavior
4. Hydrogen as an example:



Protium (hydrogen-1)



Deuterium (hydrogen-2)



Tritium (hydrogen-3)

D. Designating Isotopes
1. Hyphen notation
    a. Mass number is written after the name of the element
        (1) hydrogen-2, helium-4
2. Nuclear Symbol
    a. Composition of the nucleus using the element's symbol
        (1) ${}^{2}_{1}\mathrm{H}$ Mass number =2 Atomic number = 1
        (2) ${}^{4}_{2}\mathrm{He}$ Mass number = 4 Atomic number = 2

II. Using Atomic Mass

A. Average Atomic Masses

1. The weighted average of the atomic masses of the naturally occurring isotopes of an element
    a. Atomic masses on the periodic table are average masses
    b. In calculations using atomic mass, we will round the masses to two decimal places before doing calculations

Examples:		
	Mg = 24.3050	we use: 24.31
	O = 15.9994	we use: 16.00
	N = 14.00674	we use: 14.01

III. Avogadro's Number and the Mole

A. The Mole

1. The amount of substance that contains as many particles are there are in exactly 12 grams of carbon-12
2. The amount of substance that contains the Avogadro number of particles

B. Avogadro's Number

1. The number of particles in exactly one mole of a pure substance
2. Avogadro's number = 6.022×10^{23}

C. Molar Mass

1. The mass of one mole of a pure substance
    a. Units are grams/mole (or g/mol)
    b. Molar mass of an element equals the average atomic mass in gram units

25.1 Nuclear Radiation

I. Introduction

A. Nucleons

1. Neutrons and protons

B. Nuclides

1. Atoms identified by the number of protons and neutrons in the nucleus
    a. radium-228 or $^{228}_{88}Ra$

II. Radioactivity

A. Radioactive Decay

1. The spontaneous disintegration of a nucleus into a slightly lighter and more stable nucleus, accompanied by emission of particles, electromagnetic radiation, or both

B. Nuclear Radiation

1. Particles or electromagnetic radiation emitted from the nucleus during radioactive decay

C. Unstable Nuclides

1. All nuclides beyond atomic # 83 are unstable and radioactive

III. Types of Radioactive Decay

A. Alpha Emission

1. Alpha particle (α) is a helium nucleus ($^{4}_{2}He$), so it has a 2+ charge.
2. Alpha emission is restricted almost entirely to very heavy nuclei

$$^{210}_{84}Po \; + \; \rightarrow ^{206}_{82}Pb \; + \; ^{4}_{2}He$$

B. Beta Emission

1. Beta particle (β) is an electron emitted from the nucleus during nuclear decay

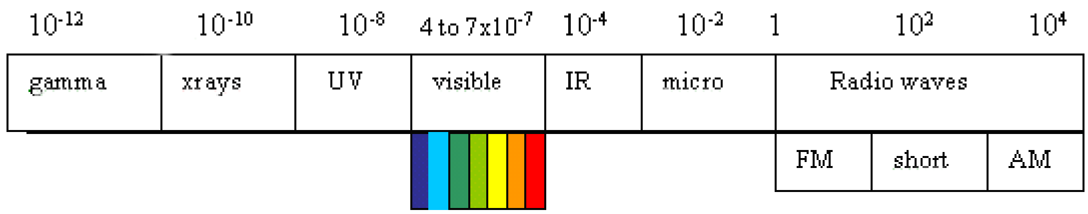
$$^{1}_{0}n \rightarrow ^{1}_{1}p + ^{0}_{-1}\beta$$

2. Beta particles are emitted when a neutron is converted into a proton and an electron

$$^{14}_{6}C \rightarrow ^{14}_{7}N + ^{0}_{-1}\beta$$

C. Gamma Emission

1. Gamma rays (γ) are high-energy electromagnetic waves emitted from a nucleus as it changes from an excited state to a ground energy state
2. Gamma rays are produced when nuclear particles undergo transitions in energy levels
3. Gamma emission usually follows other types of decay that leave the nucleus in an excited state

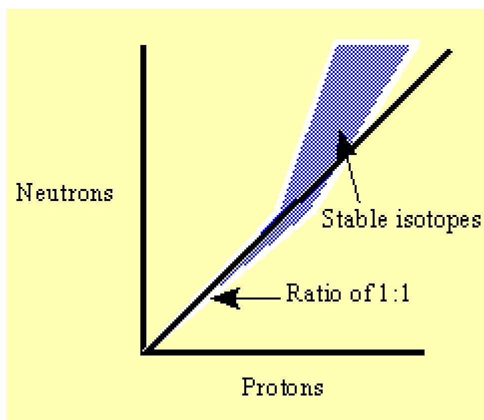


25.2 Radioactive Decay

I. Nuclear Stability and Decay

A. Neutron-to-Proton Ratio determines the type of decay that occurs

1. Band of Stability



II. Half-Life

A. Half-Life ($t_{1/2}$)

1. The time required for half the atoms of a radioactive nuclide to decay
   a. More stable nuclides decay slowly
   b. Less stable nuclides decay rapidly

Table 22-2 **Representative Radioactive Nuclides and Their Half Lives**

Nuclide	*Half-life*	*Nuclide*	*Half-life*
$^{3}_{1}H$	12.32 years	$^{214}_{84}Po$	163.7 µseconds
$^{14}_{6}C$	5715 years	$^{218}_{84}Po$	3.0 minutes
$^{32}_{15}P$	14.28 days	$^{218}_{85}At$	1.6 seconds
$^{40}_{19}K$	1.3 x 109 years	$^{238}_{92}U$	4.46 x 10^{9} years
$^{60}_{27}Co$	10.47 minutes	$^{239}_{94}Pu$	2.41 x 10^{4} years

III. Transmutation Reactions

A. Transmutations

1. A change in the identity of a nucleus as a result of a change in the number of its protons

B. Nuclear Reaction

1. A reaction that affects the nucleus of an atom
2. Small amounts of mass are converted to LARGE amounts of energy
   a. $E = mc^2$

C. Balancing Nuclear Reactions

1. Total atomic numbers and mass numbers must be equal on both sides

$$^{9}_{4}Be + ^{4}_{2}He \rightarrow ^{12}_{6}C + ^{1}_{0}n$$

25.3 Fission and Fusion of Atomic Nuclei

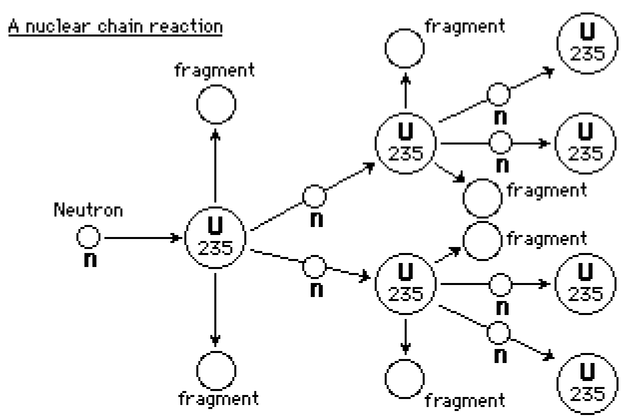
I. Nuclear Fission

A. Nuclear Fission

1. A very heavy nucleus splits into more stable nuclei of intermediate mass
2. The mass of the products is less than the mass of the reactants. Missing mass is converted to energy
   a. Small amounts of missing mass are converted to HUGE amounts of energy ($E = mc^2$)

B. Nuclear Chain Reaction

1. A reaction in which the material that starts the reaction is also one of the products and can start another reaction



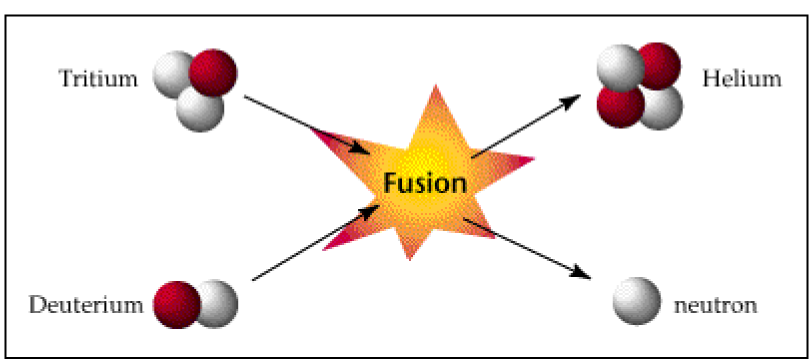
C. Critical Mass

1. The minimum amount of nuclide that provides the number of neutrons needed to sustain a chain reaction

II. Nuclear Fusion

A. Nuclear Fusion

1. Light-mass nuclei combine to form a heavier, more stable nucleus



B. Fusion Reactions
   1. More energetic than fission rxns
   2. Source of energy of the hydrogen bomb
   3. Could produce energy for human use if a way can be found to contain a fusion rxn (magnetic field?)

25.4 Radiation in Your Life

A. Penetrating Ability
   1. Alpha Particles
      a. Least penetrating ability due to large mass and charge
      b. Travel only a few centimeters through air
      c. Cannot penetrate skin
      d. Can cause harm through ingestion or inhalation
   2. Beta Particles
      a. Travel at speeds close to the speed of light
      b. Penetrating ability about 100 times greater than that of alpha particles.
      c. They have a range of a few meters in air.
   3. Gamma rays
      a. Greatest penetrating ability
      b. Protection requires shielding with thick layers of lead, cement, or both

C. Penetrating ability of radiation

Increasing penetrating ability →

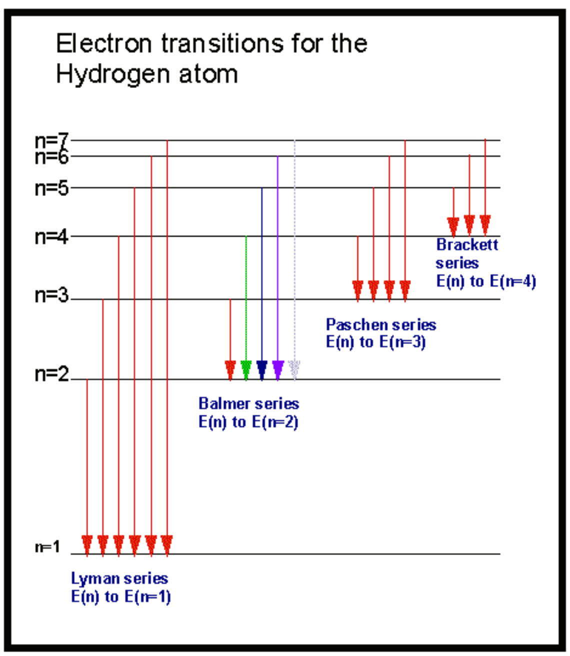
ALPHA → BETA → GAMMA
Least harmful — most harmful

C. Radioactive Elements
   1. All isotopes of all man-made elements are radioactive
   2. Some naturally isotopes are radioactive
      a. All isotopes of all elements beyond bismuth (atomic #83) are radioactive

5.1 Models of the Atom

I. The Bohr Model of the Atom

A. Electron Orbits, or Energy Levels
   1. Electrons can circle the nucleus only in allowed paths or orbits
   2. The energy of the electron is greater when it is in orbits farther from the nucleus
   3. The atom achieves the ground state when atoms occupy the closest possible positions around the nucleus
   4. Electromagnetic radiation is emitted when electrons move closer to the nucleus



B. Energy transitions
   1. Energies of atoms are fixed and definite quantities
   2. Energy transitions occur in jumps of discrete amounts of energy
   3. Electrons only lose energy when they move to a lower energy state

C. Shortcomings of the Bohr Model
   1. Doesn't work for atoms larger than hydrogen (more than one electron)
   2. Doesn't explain chemical behavior

II. The Quantum Mechanical Model

A. Probability and the Electron

1. The position and direction of motion of the electron cannot be simultaneously determined

Translated*: "The more certain I am about where it is, the less certain I can be about where it is going. The more certain I am about where it is going, the less certain I can be about where it is."*

B. Regions of probability in which electrons may be found in an atom are determined by mathematical equations. These regions are called orbitals.

III. <u>Atomic Orbitals</u>

A. Atomic orbital

1. A region in space where there is a high probability of finding an electron

B. Energy Levels of electrons (n)

1. Indicates the distance of the energy level from the nucleus
2. Values of n are positive integers
   a. n=1 is closest to the nucleus, and lowest in energy
3. The number of orbitals possible per energy level (or "shell") is equal to n^2

C. Energy Sublevels

1. Indicates the shape of the orbital
2. Number of orbital shapes allowed in an energy level = *n*
   a. Shapes in the first four shells are designated *s, p, d, f*

Energy Level (*n*)	Sublevels in main energy level (*n* sublevels)	Number of orbitals per sublevel	Number of electrons per sublevel	Number of electrons per main energy level ($2n^2$)
1	s	1	2	2
2	s p	1 3	2 6	8
3	s p d	1 3 5	2 6 10	18
4	s p d f	1 3 5 7	2 6 10 14	32

D. Electron Spin

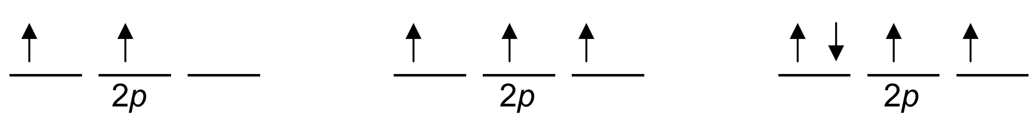
1. A single orbital can contain only two electrons, which must have opposite spins
2. Two possible values for spin, +1/2, -1/2

<u>5.2 Electron Arrangement in Atoms</u>

I. <u>Writing Electrons Configurations</u>

A. Rules

1. Aufbau Principle
   a. An electron occupies the lowest-energy orbital that can receive it
2. Pauli Exclusion Principle
   a. No two electrons in the same atom can have the same set of four quantum numbers
3. Hund's Rule
   a. Orbitals of equal energy are each occupied by one electron before any orbital is occupied by a second electron, and all electrons in singly occupied orbitals must have the same spin



B. Orbital Notation

1. Unoccupied orbitals are represented by a line, _____
   a. Lines are labeled with the principal quantum number and the sublevel letter
2. Arrows are used to represent electrons
   a. Arrows pointing up and down indicate opposite spins

C. Configuration Notation

1. The number of electrons in a sublevel is indicated by adding a superscript to the sublevel designation

   Hydrogen = $1s^1$
   Helium = $1s^2$
   Lithium = $1s^2 2s^1$

D. Exceptional Electron Configurations

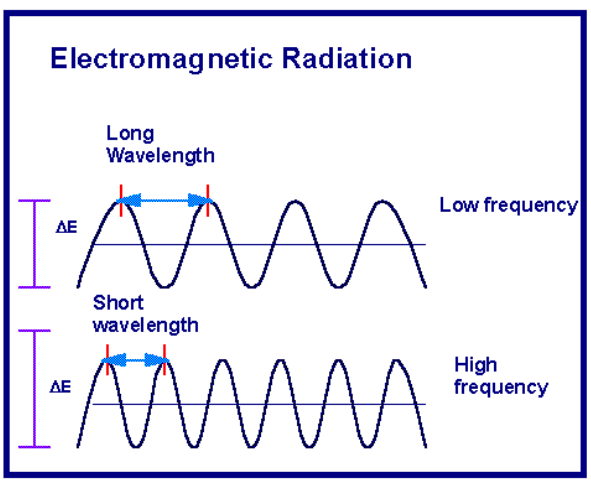
1. Irregularity of Chromium
   a. Expected: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4$
   b. Actual: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$
2. Irregularity of Copper
   c. Expected: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9$
   d. Actual: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$
3. Numerous transition and rare-earth elements transfer electrons from smaller sublevels in order to half-fill, or fill, larger sublevels

5.3 Physics and the Quantum Mechanical Model

I. Properties of Light

A. Electromagnetic Radiation

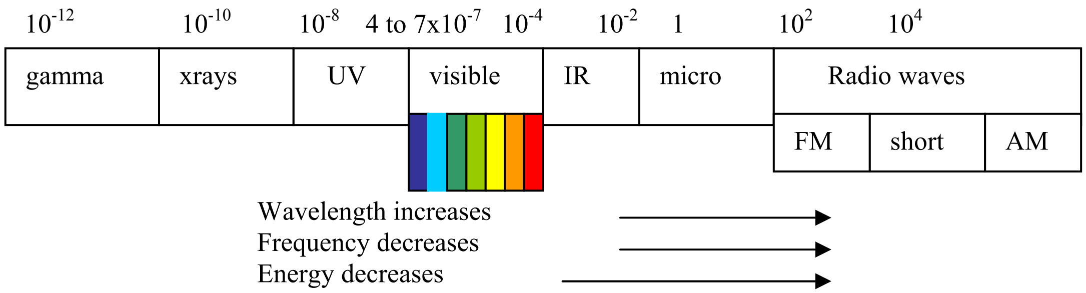
1. Many types of EM waves
   a. visible light
   b. x-rays
   c. ultraviolet light
   d. infrared light
   e. radio waves
2. EM radiation are forms of energy which move through space as waves
   a. Move at speed of light
      (1). 3.00×10^8 m/s
   b. Speed is equal to the frequency times the wavelength $c = \nu\lambda$
      (1). Freqency (ν) is the number of waves passing a given point in one second
      (2). Wavelength (λ) is the distance between peaks of adjacent waves



   c. Speed of light is a constant, so $\nu\lambda$ is also a constant
      (1) ν and λ must be inversely proportional

B. Light and Energy

1. Radiant energy is transferred in units (or quanta) of energy called <u>photons</u>
    a. A photon is a particle of energy having a rest mass of zero and carrying a quantum of energy
    b. A quantum is the minimum amount of energy that can be lost or gained by an atom
2. Energy of a photon is directly proportional to the frequency of radiation
    a. $E = h\nu$ (h is Planck's constant, 6.62554×10^{-27} erg sec)

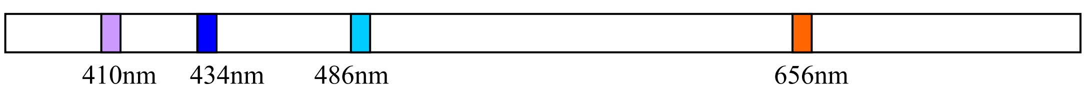


II. <u>Atomic Spectra</u>

A. Ground State
    1. The lowest energy state of an atom

B. Excited State
    1. A state in which an atom has a higher potential energy than in its ground state

C. Bright line spectrum
    1. Light is given off by excited atoms as they return to lower energy states
    2. Light is given off in very definite wavelengths
    3. A <u>spectroscope</u> reveals lines of particular colors



a. Definite frequency
b. Definite wavelength