ATOMIC STRUCTURE

Development of Atomic Theory:

Law of Conservation Of Mass (Lavoisier): The total mass remains constant during a chemical reaction.

Law of Definite Proportions (Proust): All samples of a compound have the same composition; that is, all samples have the same proportion by mass of the elements present.

Know the experiments and the contributions to atomic theory of the following scientists:

- Dalton
- Thomson
- Millikan
- Rutherford
- Planck
- Einstein
- Bohr
- DeBroglie
- Schrodinger

Atomic Structure:

Review the subatomic particles, their charges, relative masses and location.
Review ions and isotopes.

Mass Spectrometry:

- Mass spectrometer separates positive gaseous ions depending on their mass to charge ratio
- Ions are deflected into circular paths, lighter ions are deflected the most, forming tighter circles.
- Gives a mass spectrum which shows the relative number of each atom of each mass in the sample (used to determine average atomic masses).
- Know how to calculate the average atomic mass of an element.

Ex: A sample of naturally occurring carbon contains 98.892 % carbon-12 and 1.108 % carbon-13. What is the average atomic mass of carbon?

The Wave Nature of Light:

- **Wavelength (λ)** - the distance between two crests (or troughs) - measured in nm (10^{-9} m) or angstroms (1 A = 10^{-10} m)
- **Frequency (ν)** - the number of cycles of a wave that pass a point in a certain time - measured in hertz (Hz) - 1 Hz = $1s^{-1}$.

- **Amplitude** - the height of the wave from origin to crest.
- **Speed (c) -** in a vacuum the speed of light = 3.00×10^8 m/s

  $\mathbf{c = \lambda\nu}$

Ex: Calculate the frequency in Hertz of an X-ray that has a wavelength of 8.21 nm.

- **Electromagnetic spectrum -** shows complete range of wavelengths and frequencies for electromagnetic radiation.
- **Continuous spectrum -** all wavelengths and frequencies (rainbow)
- **Line spectrum -** only certain wavelengths and colors represented.
- **Emission spectra -** line spectra from analyzing light emitted by elements excited by heat or electricity - unique for each element (atomic fingerprint).

Planck's Quantum Hypothesis:

- Energy can only be absorbed or emitted as a quantum (or multiples of a quantum).
- **$E = h\nu$ E = energy (J)   h = Planck's constant (6.63×10^{-34} Js) ν = frequency (s^{-1})**

The Photoelectric Effect:

- When a photon of light with the correct energy hits the surface of a substance, an electron is ejected.
- Energy of photon : $\mathbf{E = h\nu}$
- Energy of one mole of photons: $\mathbf{E = 6.02 \times 10^{23} \times h \times \nu}$

Ex: Calculate the energy in joules of a photon of violet light that has a frequency of $6.15 \times 10^{14}\ s^{-1}$.

Ex: Calculate the energy, in joules per photon, of ultraviolet light with a wavelength of 235 nm.

Ex: A laser produces red light of wavelength 632.8 nm. Calculate the energy, in kilojoules, of one mole of photons of this red light.

Bohr's Hydrogen Model:

- An electron has a distinct energy level (E_n)
- $\mathbf{E_n = \dfrac{-2.178 \times 10^{-18}\ J}{n^2}}$ **n = an integer**

Ex: Calculate the energy of an electron in the second energy level of a hydrogen atom.

- The amount of energy absorbed or emitted by an electron as it changes levels can be found by calculating the energy at each of the two levels then finding the difference.

Ex: Calculate the energy change, in joules, that occurs when an electron falls from the n = 5 to the n = 3 energy level.

- The frequency and wavelength of photons released can then be calculated using $E=h\nu$

Ex: Calculate the frequency of the radiation released by the transition of an electron in a hydrogen atom from the n = 5 energy level to the n = 3 energy level.

- **Ground state** - the electrons of an atom are at the lowest possible energy level.
- **Excited state -** An electron is promoted to a higher energy level.

Quantum Mechanics (Wave Mechanics):

- **DeBroglie's equation: $\lambda = h / m\nu$ m = mass**
- All matter exhibits wavelike behavior
- Led to development of electron microscope.
- **Schrodinger's equation** describes wave functions - gives the **probability** of finding an electron in a particular volume of space in the atom.
- **Heisenberg uncertainty principle:** We cannot determine both the position and momentum of a small particle simultaneously.

Quantum Numbers:

1. The Principle Quantum Number (n)

- Describes the main energy level an electron occupies. It may be any positive integer.
  n = 1, 2, 3, 4, ...
  K, L, M, N (corresponding to Bohr's atom shells)

2. The Orbital Angular Momentum Quantum Number (l)

- Designates a specific kind of atomic orbital (sublevel) an electron occupies.
- Determines the shape of the orbital.

l = 0, 1, 2, 3, ...(n-1)
s, p, d, f

3. The Magnetic Quantum Number (m_l)

- Designates the spatial orientation of an atomic orbital.
- Tells how many atomic orbitals are associated with a sublevel.
  $m_l = -l, \ldots 0, \ldots +l$

4. The Spin Quantum Number (m_s)

- Designates the spin of an electron and the orientation of the electromagnetic field produced by this spin.
  m_s = +1/2 , -1/2
  ↑ ↓
- This means each atomic orbital can accommodate two electrons, each with a different spin.

Pauli Exclusion Principle:

- No two electrons in an atom may have identical sets of four quantum numbers.

Shapes Of Orbitals:

s orbital - 1 type, spherical

p orbitals - 3 types (p_x, p_y, and p_z), dumbell/hourglass

d orbitals - 5 types (d_z2, d_x2-$_y2$, d_{xy}, d_{xz}, d_{yz}), see book for shapes

f orbitals - 7 types (complex shapes and naming)