**Chapter 7: The Quantum Mechanical Model of the Atom**

**Objectives:**

1. Become familiar with properties of waves

2. Learn about electromagnetic radiation

3. Understand Planck's quantum theory

4. Become familiar with the photoelectric effect

5. Explain Bohr's theory of the hydrogen atom

6. Understand the origin and characteristics of atomic emission spectra

7. Learn about particle-wave duality and the de Broglie equation

8. Understand the Heisenberg uncertainty principle and the Schrodinger equation

9. Understand the meaning of quantum numbers

10. Understand the shape of s, p and d orbitals

**Essential Vocabulary**

* Quantum-mechanical model
* Electromagnetic Radiation
* Wavelength ( λ )
* Frequency ( ν )
* Electromagnetic Spectrum
* Photon
* Photoelectric effect
* Constructive interference
* Destructive interference
* Diffraction
* Photon
* Emission Spectrum
* De Broglie relation
* Heisenberg uncertainty principle
* Orbital
* Quantum number
* Principal Level
* Sublevel
* Probability Density

## Helpful Hints:

## Atomic Orbitals:

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Number | Name | Description | Meaning | Mathematics |
| *n*  | Principle Quantum Number | energy level(i.e.; Bohr’s Theory) | *n* = 1,2,3,4,5,6,7 | *n*= 1,2,3,4,… |
| *l* | Angular Momentum Quantum Number | general (number) shape of the orbitals | *l* = s, p, d, f(s =0, p =1d =2, f =3) | *l* = n-1 |
| *ml* | Magnetic Quantum Number | 3-D orientation of the orbital | s=1, p=3, d=5, f=7 | *ml*= 2*l* + 1 |
| *ms* | Spin Quantum Number | spin of the electron | clockwise or counter-clockwise | *ms* = +1/2 or –1/2 |

## Quantum Numbers:

1. , 1 , 0 , +1/2) Meaning … \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
2. , 3 , 3 , -1/2) Meaning … \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

(2 , 2 , 1 , +1/2) non-existent Why? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**Quantum Theory:** *Max Planck* determined the energy in the electromagnetic spectrum is quantized. For any given frequency of light (radiation), all possible energies are multiples of a certain unit of energy called a quantum. So, energy changes occur in small and specific steps.

The energy of an electron: E= -2.178 X 10 Joules

Energy and Electromagnetic Radiation (light): When atoms absorb energy in the form of electromagnetic radiation, electrons jump to higher energy levels. When electrons drop from higher to lower energy levels, atoms give off energy in the form of electromagnetic radiation. The energy level changes for the electrons of a particular atom are always the same, so atoms can be identified by their emission and absorption spectra.

E = h ν

C = λ ν

E = Energy (Joules)

λ = wavelength (m)

ν = frequency (hz)

c = speed of light (m/s) = 3.0 x108 m/s in a vacuum

**The Bohr Model:** Neils Bohr took the quantum theory and used it to predict that electrons orbit the nucleus at specific fixed radii. This only works for hydrogen and small ions.

**The Heisenberg Uncertainty Priniciple:** Wemer Heisenbarg claimed it was impossible to know the position and momentum of an electron at any moment. Therefore, the Bohr model orbitals are actually a probability function describing the possibility that an electron will be found in a region of space.

**The de Broglie hypothesis:** All matter has wave characteristics. There is a simple relationship between an electron's wave and particle characteristics.

**Chapter 8: Periodic Properties of the Elements**

**Objectives:**

1. Learn to determine the electron configuration an atom or an ion.
2. Learn about the magnetic properties of atoms.
3. Learn about writing the ground-state electron configuration of atoms
4. Learn about the classification of the elements.
5. Learn about properties such as atomic radius, ionization energy, and electron affinity.
6. Understand periodic trends in physical and chemical properties.

**Essential Vocabulary**

* Ground state
* Pauli Exclusion Principle
* Degenerate
* Coulomb’s Law
* Shielding
* Effective nuclear charge
* aufbau principle
* Hund’s Rule
* Valence electrons
* Core electrons
* Van der Waals radius
* Covalent radius
* Atomic radius
* Paramagnetic
* Diamagnetic
* Ionization energy
* Electron affinity

## Helpful Hints:

\*The horizontal rows on the periodic table are called PERIODS

\*The vertical columns on the periodic table are called GROUPS or FAMILIES

**Filling Order of the Orbitals**: use the diagonal rule (some exceptions do occur).

 **1s**

 2s **2p** Start with 1s and follow the arrows.

 3s **3p** 3d Maximums: s = 2, p = 6, d = 10, f = 14

 4s **4p** 4d 4f

 5s **5p** 5d 5f Example: Bromine has 35 e-‘s

 1s22s22p63s23p64s23d104p5

 6s **6p** 6d or … [Ar]4s23d104p5

 *This is known as electron configuration.*

 7s **7p**

1. Pauli Exclusion Principle: no two electrons within an atom may have the same set of quantum numbers therefore; electrons in the same orbitals must have opposite spins.

2. Hund's Rule: Electrons will occupy orbitals within the same sublevel as single electrons until all orbitals contain one electron. At that time pairing will begin with opposite spins.

3. The Aufbau Principle: Electrons enter orbitals with the lowest energy possible

## Orbital Notation

 1s 2s 2p 3s 3p

 Phosphorus (17 e-s)

 Three unpaired electrons in the valence shell.

## Diamagnetism and Paramagnetism:

 Paramagnetic, attracted by a magnet Diamagnetism, repelled by a magnet

Ex) He, Be, Ne

General rule: Any element with an odd-electron number *must be* paramagnetic.

**Diamagnetic elements**

\*have all electrons paired within the orbitals

\*have all of their subshells completed

\*Are not affected by magnetic fields

**Paramagnetic elements**

\*do not have all their electrons paired

\*are strongly affected by magnetic field

## Shielding:

\*Inner spherical orbitals (s-orbitals) "shield" electrons on higher levels from the pull of the nucleus partially since they completely surround it. This is most readily observed in the great difference between the ionization energies of the noble gases (high) and the alkali metals (low) that immediately follow them

\*s-orbitals shield p and d orbital electrons on the same energy level to a much smaller extent. This effect is only really obvious when the ionization energy of an element like Mg is compared to that of Al. The poorer shielding as electrons are added on the same energy level and more protons are added in the nucleus accounts for the decrease in size across a period since the electrons feel the pull of the nucleus more.

Example: A 2s electron spends more of its time (on average) around the nucleus than a 2p electron due to the size of the orbitals. The 2s is said to be more “penetrating” than the 2p, and is less shielded by the 1s electrons.

\*\*\* Periodic properties are best understood in terms of shielding

## Periodic Table Trends

To understand atoms and their electrons, follow these three basic periodic trends:

1. Electrons are attracted to the protons in the nucleus.

* + The closer an electron is to the nucleus, the more strongly it is attracted to the nucleus.
	+ The more protons in the nucleus, the more strongly electrons are attracted to the nucleus.

2. In an atom, electrons repel other electrons. If there are electrons between the nucleus and the valence electrons, the valence electrons will be less attracted to the nucleus due to **electron shielding** by the internal electrons.

3. Completed principle energy shells (and to a lesser extent, completed subshells) are very stable. Atoms prefer to add or subtract electrons to create complete shells if possible.

## The BIG picture: properties such as atomic size (or radius), ionization energy, electronegativity, metallic character, etc. repeat regularly throughout the table because the electron configurations repeat regularly

##

Atomic Radius - Increases as you move LEFT (poor shielding) or DOWN (good shielding)

atoms shrink as you advance across a period because added protons in the nucleus pull the added electrons more tightly in the absence of effective shielding

atoms in the same column increase in size as you go down the family since electrons are being added on successively higher energy levels and are well-shielded by the previous completed electron energy levels

Ionic Radius – Cations are smaller, Anions are larger (than the neutral atom)

Ionization Energy - Decreases as you move DOWN, High in full or ½ full orbitals

\*The energy required to remove a valence electron from an atom is called the first ionization energy. Once an electron has been removed, the atom is a positively charged ion. The energy required to remove the next electron from the ion is called the second ionization energy, and so on. The successive ionization energies increase due to the decreased electron-electron repulsion which allows the electrons to move closer to the nucleus thereby, increasing the attractive forces of the nucleus and the ionization energy.

Ionization energy increases as you move right because the outer-most electron becomes increasingly more difficult to remove as you advance across a period due to poor shielding and the added protons continue to exert a stronger attraction.

As you go down a group, ionization energy decreases, therefore the outer-most electron is easier to remove as you go down a column since each time you are one energy level farther from the nucleus, benefiting from that much more shielding

**Electronegativity: (FONClBrISCH)** The attractive force exerted by the nucleus on its valence electrons. Electronegativity increases up and across from left to right on the period table.

Electron Affinity – Increases as you move right, Decreases as you move DOWN, High in almost full orbitals, low in full orbitals

\* This is a measure of the change of energy of an atom when an electron is added to it. If the addition of an electron yields a more stable atom, energy is released. If the addition of an electron yields a less stable atom, energy is absorbed because the electron must be placed in a higher energy level.

\*The ability to attract additional electrons increases as you advance across a period since shielding is largely ineffective and the pull of the increased protons in the nucleus is sufficient to influence electrons which approach the outer limits of the atom from other sources

\*Any attractive force of the nucleus is so diminished by effective shielding of electrons in lower levels that additional electrons become more difficult to hold onto as you move down a column

Metallic Character - Increases as you move LEFT or DOWN

\*metallic properties are related to the freedom of the outer electrons to move and thus these properties are most pronounced at the lower left side of the table where shielding is at a maximum

\*Example: The most reactive metal would be Fr. The most reactive non-metal would be F.

In reacting, metals typically form + ions. For this to happen electrons must be lost. This will be favored by excellent shielding.

In reacting, non-metals typically form - ions. For this to happen, electrons must be gained and held. This will be favored by poor shielding.

## Chemical Properties:

Hydrogen Metal or a non-metal (oxidation # of +1 and –1)

Group 1A Metal w/oxidation # of +1, increase reactivity down table, react with water to form H2, forms oxides or peroxides when reacted with O2 *lone, well-shielded outer electron is removed and the remaining electrons are tightly held*

Group 2A Metal w/oxidation # of +2, react with acid to form H2, may react with O2 , *well-shielded outer electrons removed and the remaining electrons are tightly held*

Group 3A Metal w/oxidation # of +3, [why is +1 observed in some Group IIIA elements?] reacts with acid and O2,

Group 4A Metalloids w/oxidation # of +4 or –4, will not react with H2O, may react with acid

Group 5A First 2 are nonmetals, later are metalloids. Wide variation with oxidation #

Group 6A First 3 are nonmetals, later are metalloids. Commonly –2 oxidation #

Very reactive with numerous compounds possible. E*lectrons are too tightly held to remove, but shielding is so poor that high electronegativity is able to attract enough electrons to fill level (2 more)*

Group 7A Most reactive non-metals. Commonly, -1 oxidation #. React with H to form hydrogen halides.

Group 8A Noble gases. Reactive only in extreme situations. NOT INERT.