<u>Unit 4: Atomic Structure and Theory</u> (Honors Chemistry)

<u>Homework</u>: *Read Ch. 4* *p86 -107* *<u>Also read and annotate the Unit 4 Article Democritus to Dalton</u>*

<u>A: Defining the Atom</u>

<u>Democritus:</u> Matter is made of "indivisible parts" he called **<u>atomos</u>** "meaning cannot be cut". Because of his ideas we still call these particles atoms.

<u>John Dalton:</u> Proposed that atoms are small indivisible spheres.

<u>Dalton's Model of the Atom:</u> Was called the "**Billiard Ball**" model and was a hard sphere.

<u>Einstein:</u> Proves atoms exist!!!

<u>JJ Thomson:</u> Thompson did experiments with cathode ray tubes. He found that all atoms contain a little particle he called the **<u>electron.</u>** Thompson theorized that since atoms are neutral they must also contain positive charges.

<u>Thomson's Model of the Atom:</u> Was called the "**Plum Pudding or Raisin Bun**" model and was positive dough with negative electrons like raisins scattered throughout.

<u>E. Rutherford:</u> Rutherford did experiments with radioactive Polonium which emitted positive particles called alpha particles. In an experiment he shot the positive alpha particles through a piece of gold foil and then onto a screen. From the screen he determined what happened to the particles as they went through the atoms in the foil. Most of the particles went straight through but a few were deflected and some came straight back. He then concluded **atoms are mostly empty space** with a dense positive center he called the <u>nucleus.</u> He called the positive particles in the nucleus **<u>protons</u>**. Alpha particles are deflected by the nucleus if they get close enough.

<u>Rutherford's Model of the Atom:</u> Was called the "**Planetary Model"** and has a very small compact nucleus made of positively charged protons with electrons around the outside in circular orbits in a lot of empty space.

<u>Chadwick:</u> In 1932 Chadwick discovered the **<u>neutron.</u>**

<u>Bohr:</u> Energy can be used to make electrons go from an inner orbit to a farther out orbit. He theorized electrons are in very specific energy levels around the nucleus. If an electron can gain a precise amount of energy it will jump to a higher level until it is pulled back to its original position. This is called a quantum leap.

<u>Bohr's Model of the Atom:</u> A positive charged very dense nucleus with electrons around the nucleus in a lot of empty space and located in orbitals in specific energy levels.

<u>Schrodinger :</u> Treated an electron as an **energy wave** and proposed that an atom has a positive nucleus and is surrounded by electrons acting as wave patterns in different cloud shapes. He came up with a complicated formula that shows the most probable locations of the electrons around the nucleus.

Born and Heisenberg: Treated an electron as a **particle** and came up with an atom with a positive nucleus, surrounded by probable locations where the electron can be found, and this also made different cloud shapes. Their cloud shapes matched those that Schrodinger came up with when he treated the electrons like energy waves.

It turns out electrons act as both particles and like energy waves.

The Cloud Model of the Atom: Based on Schrodinger, Born and Heisenberg's work, this model places electrons in cloud shape **orbitals** – places where electrons can be found 90% of the time.

Max Plank: In 1900, Max Planck was working on the problem of how the radiation an object emits is related to its temperature. He came up with a formula that agreed very closely with experimental data, but the formula only made sense if he assumed that the energy of a vibrating molecule was **quantized**--that is, it could only take on certain values.

Albert Einstein: Based on Planck's work, Einstein proposed that light also delivers its energy in chunks; light would then consist of little particles, or quanta, called photons. This idea led to Einstein's famous equation $e = mc^2$. This says that energy in anything is equal to it's mass times the speed of light squared!!!

B: Light – Electromagnetic Radiation:

Homework: *Read Chapter 5 pages 116 – 141*

- **Electromagnetic waves:** (radiation) – another name for light waves
- **Light categories-**

1. ______________________________**(low energy)**
2. ______________________________
3. ______________________________
4. ______________________________
5. ______________________________
6. ______________________________
7. ______________________________**(high energy very dangerous!!!!)**

- **Speed of light:** travels at 2.998×10^8 m/sec or 186,000 miles/sec

2 ways we see colors

Emitted Light: ____

Reflected Light: ____

Every element or compound when given a lot of energy, gives off its own set of colors called its **Spectrum**
Spectroscopy – identifying a substance by its spectrum. **Homework:** **Unit 4 Wk 1**

Example: Why does something look blue or orange or any color?

Homework: **Unit 4 Wk 1**

Wavelength: (λ) (lambda) – Is the distance between 2 waves, can be measured anywhere on the wave.

Frequency = number of waves per second (ν) **Measured in hertz (Hz) 1 hertz = 1 cycle/second (s^{-1})**

Formula: $c = \lambda\nu$

The product of the wavelength **λ, (lambda)** and the frequency **ν, (nu)** is the speed of light (**c**)

c = speed of light in a vacuum is **2.998×10^8 m/s** (I will give you this on the test)

Examples: Calculate the wavelength of yellow light emitted by a sodium lamp if the frequency is 5.10×10^{14} Hz (5.10×10^{14} s^{-1})?

What is the frequency of orange light with a wavelength of 650 nm? (**λ is in nm and c is in m so convert!!)**

What is the frequency of blue light with a wavelength of 510 nm?

What is the wavelength of light with a frequency 5.89×10^{14} Hz? What color is it? (Look up in your book)

What is the wavelength of X-rays with a frequency 5.162×10^{18} Hz?

- The wavelength and the frequency are inversely proportional.
- This means as one increases the other decreases.

- **So higher energy waves have a smaller wavelength. Compare your answers in the previous questions.**

Particle Nature of Light:

PLANCK:

Said light is absorbed or emitted in packages called quanta or photons and made a formula to calculate the energy in a photon. Energy is measured in joules (J)

A photon is a particle of electromagnetic radiation with no mass that carries a quantum of energy.

Formula: $E = h\nu$

The Energy per photon is equal to the product of Planck's constant (h) and the frequency (ν) in Hz

Planck's constant = 6.626×10^{-34} J•s (I will give you this number on the test)

Examples:
What is the energy of a photon given of when Potassium is heated. The color is violet and has a frequency of $\nu = 7.23 \times 10^{14}$ Hz.

What is the energy of a photon with a $\nu = 6.32 \times 10^{20}$ Hz. What type of radiation is it? (See your notes).

Lasers in CD players have a $\lambda = 780$nm. Get the energy per photon. (Takes 2 steps)

PLANCK:

- Said light is absorbed or emitted in packages called quanta or photons and made a formula to calculate the energy in a photon. Energy is measured in joules (J)
- A photon is a particle of electromagnetic radiation with no mass that carries a quantum of energy.

EINSTEIN:

- Based on Planck's work, Einstein proposed that light also delivers its energy in chunks; light would then consist of little particles, or quanta, called photons. This idea led to Einsteins famous equation **$e = mc^2$**.

Homework: **Unit 4 Wk 2 all**

C: Spectrum Analysis: (Bohr)

- Bohr's theory – electrons can only be in certain areas around the nucleus called orbitals which are in bigger areas called energy levels.

Bohr asked 2 important questions:

1. Why do atoms give off a set of colors (spectrum) when given a lot of energy?

The energy enables the electrons to pull away from the positive charged nucleus and go to farther out energy levels. A fraction of a second later the nucleus pulls the electron back and the extra energy is released as a photon of light. This light is what produces the spectrum.

2. Why do different types of atoms give off a different set of colors (different spectrum)?
Every type of atom has a different number of protons, so they have a different positive pull from the nucleus. Therefore the energy needed to jump to farther out orbitals requires a different amount of energy. So on the way back down this different energy is also released in different size photons (different colors).

D: Quantum Mechanics Theory:

Miscellaneous:

Quantum mechanics theory: Name for modern day theory of the atom.

- Protons (p+) and Neutrons (n) in the nucleus in the center of the atom, the nucleus is very compact
- The nucleus is held together by the **nuclear force** which is supplied by the neutrons.
- Electrons are located outside the nucleus in areas called orbitals which are in larger spherical areas called energy levels.
- The electrons travel at the speed of light in a lot of empty space.
- The lowest energy level is called the **ground state.** It takes energy to move out of the ground state.

Subatomic particles:

Protons (p+) in nucleus (smallest unit of + charge) **Mass of proton = 1.67 x 10^{-24} g**
Neutrons (n) in nucleus function to hold the nucleus together **Mass of neutron = 1.67 x 10^{-24} g**
Electrons (e-) in outer part of the atom in orbitals within energy levels. **Mass of electron = 9.11 x 10^{-28} g**
(Smallest unit of – charge)

6 types of quarks: up, down, top, bottom, charm, and strange **(up and down make up protons and neutrons** and the other 4 are made in high energy collisions**)**

- All protons are identical, all electrons are identical, all neutrons are identical
- Differences between atoms come from different number of protons, electrons, and neutrons **(especially protons)**
- All atoms, before reacting have a neutral charge (same number of protons and electrons)
- All atoms before reacting have a **NEUTRAL CHARGE.** Which means means they have the same number of protons and electrons. (In some types of reactions atoms lose or gain electrons and end up with a + or – charge thus becoming an ion.)

- Atoms are so small, it is difficult to discuss how much they weigh in grams.
- Use atomic mass units (amu).
- An atomic mass unit (amu) is one twelfth the mass of a carbon-12 atom.

Atomic Number – Indicated by a **Z** **The number of p+ in a nucleus**

3 things have the same # value

Atomic Number
Number of protons
Number of electrons

Mass number (Atomic Mass) - Indicated by a **A** **The # of p+ and n in an atom**

2 things have the same # value

Atomic Mass
Molar Mass

Ionic Charge- **The # of p+ and e$^-$ in an atom is not the same and does not add up to ZERO!!!**
Charge = **The # of p+ - # e$^-$**

Isotopes:____

____

Example: Oxygen 16 ____

Oxygen 17 ____

Oxygen 18 ____

Example (Shorthand notation): O O O

Remember these are Isotopes so you CAN't use the Periodic Table Mass!!!!!!

Isotope name	Atomic number	Mass Number	# of protons	# of neutrons	# of electrons	shorthand notation
8 oxygen-16						
47 silver-108						
10 neon-22						
55 cesium-134						
19 potassium- 41						

Remember these are Isotopes so you CAN't use the Periodic Table Mass!!!!!!

Atomic Symbol	Atomic number	Mass Number	# of protons	# of neutrons	# of electrons	Charge
Hg			**80**			**0**
Ag			**47**			**+1**
	7	**15**			**10**	
		132		**54**		**0**

Homework: ***Unit 4 Wk 3 You will have a quiz on this material.***

E: ORGANIZATION OF ELECTRONS

Electrons are inside orbital within energy levels. Energy levels are called _________________________

These are spherical areas around the nucleus. Have the symbol **n,** called the _____________________

n = 1, 2, 3, etc from the nucleus

Orbitals – *area inside energy levels that can hold a maximum number of 2 electrons*

4 main types of orbitals – s, p, d, f

Shapes:

s = ________________**(one of these)** **p =** _______________ **(3 of these)**

d = _______________ **(5 of these)** **f =** shape is too complex **(7 of these)**

Electrons have 4 quantum numbers to distinguish them from other electrons in the atom. (n, ℓ, m_s and m_ℓ)

Principle Quantum Number n indicates energy level

Indicated by an n.
n = 1, 2, 3, 4, etc So for an e- on the 4th energy level **n = 4**

Quantum Number ℓ indicates type of orbital

s = __________
p =__________
d =__________
f =__________

Quantum Number m_s

Indicates the type of spin
_____ = clockwise
_____ = counterclockwise

Quantum Number m_ℓ indicates the direction of the orbital on the x, y, or z axis

$m_\ell = +/- \ell$ if $\boldsymbol{\ell = 2}$, m_ℓ = +2, +1, 0, -1, -2

List the 4 quantum numbers and their symbols :

____

____

____

____

An electron is out in the third energy level in a d orbital and has a counterclockwise spin. Answer the following:

_________ = n   _________ = ℓ _________ = m_s

An electron is out in the fifth energy level in an s orbital and has a clockwise spin. Answer the following:

_________ = n   _________ = ℓ _________ = m_s

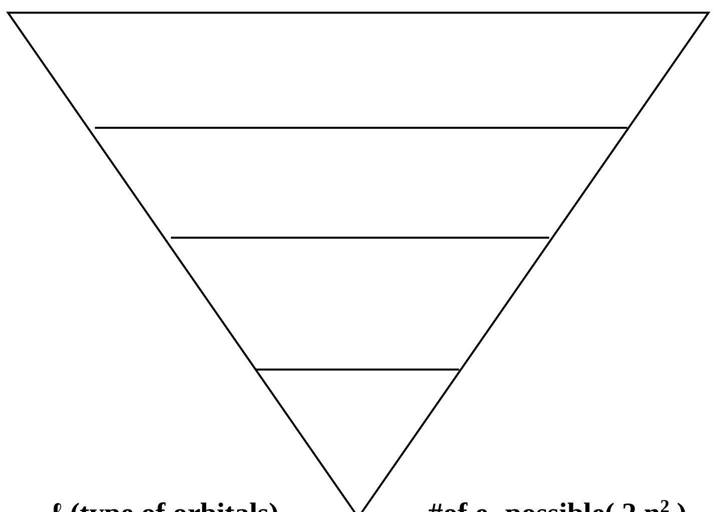
Describe an electron from the follow information. n = 2, ℓ = 2, and m_s = +½

____

There are n^2 orbitals per energy level (n = energy level)

There can only be $2n^2$ electrons in each energy level

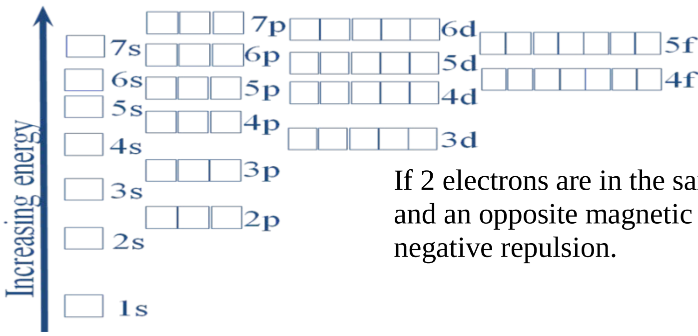
One more orbital type is added per energy level as you go out from the nucleus. (there is more room for orbitals as the farther you go from the nucleus)



n (energy level)	n^2 (# of orbitals)	ℓ (type of orbitals)	#of e- possible($2n^2$)
1	$1^2 = 1$	1s (1)	1 x 2 =2 e-
2	$2^2 = 4$	2s 2p (1)(3) = 4	4 x 2 = 8
3	$3^2 = 9$	3s 3p 3d (1)(3)(5) = 9	9 x 2 = 18 e-
4	$4^2 = 16$	4s 4p 4d 4f (1)(3)(5)(7) = 16	16 x 2 = 32 e-

Electron Configuration:

- **Aufbau Principle: Electrons will occupy the lowest energy level**
- **Hund's Rule: Electrons will only share an orbital if they have to. (singles before doubles)**
- **Pauli Exclusion Principle: Electrons in the same orbital have opposite spins**



If 2 electrons are in the same orbital they have opposite spin and an opposite magnetic fields. This helps overcome some of the negative repulsion.

Rules for Writing the e- configuration for atoms and drawing the e- configuration.

Using you periodic chart follow these short cut rules

s and p correspond to the row
d's are 1 behind row
f's are 2 behind row

Whole atom practice: Indicate all of the electrons in the atom

Example**: Fe (26) $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $4s^2$ $3d^6$**

Ca (20) ____

Ag (47)____

U (92)____

Noble gas electron configuration (Shortcut) practice: Go to the row the atom is in and write only those electrons

Example**: Fe (26) [Ar] $4s^2$ $3d^6$**

Li(3) ____________________

Zn(30) ____________________

Cl (17) ____________________

Sn (50) ____________________

Last Orbital Practice: Only the orbital the last electron is in

Example: Fe(26) $3d^6$

Li(3) ____________________ Cl (17) ____________________

Zn(30) ____________________ Sn (50) ____________________

Valence Electrons
Def: electrons on the outer level of an atom
-electrons that may be involved in bonding to another atom.
-usually only **s and p orbitals**

Number of Valence Electrons (read off periodic table)

The dots represent the valance electrons the element has in it's outer shell.
The most an element can have is 8. (Octet rule)

Ca = _______ Pb = ________ Cl=_______ B = __________

Valence Electron Configuration:

Go to the row the atom is in and write only the s and p electrons.

Example**: Fe (26)** $\mathbf{4s^2}$

Ca (20) ________________ Ag (47) _________________ Po (84) __________________

Electrons Dot Structures (Lewis Dot)

The dots represent the valance electrons the element has in its outer shell

Ne Ca Pb Cl B

Name that element

The electrons in the last orbital are given. Write the symbol for the element
Example: $\mathbf{3d^6}$

$7s^2$ ______________________

$5p^4$ ______________________

$4d^1$ ______________________

Drawing the whole atom configuration

Whole atom practice**:** Write out the electron configuration and then draw a picture of all the electrons. Remember there are only 2 electrons in each orbital.
Example**:**

Fluorine (9)

(Writing) $\mathbf{1s^2}$ $\mathbf{2s^2}$ $\mathbf{2p^5}$

(Drawing)

<u>Potassium</u> (19)

(Writing) **$1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $4s^1$**

(Drawing)

<u>Homework:</u> Unit 4 Wk 4 all.

Details and exceptions:

- Elements in the same column have the same electron configuration.
- Put in columns because of similar properties.
- Similar properties because of electron configuration.
- Noble gases have filled energy levels.
- Transition metals are filling the d orbitals
- **Scientists aren't sure of why it happens, We pretend all follow rules**

Exceptions:

- Copper expect [Ar] $4s^2$ $3d^9$
  - However one of the 4s e- jumps into the 3d orbital and the real configuration is [Ar] $4s^1$ $3d^{10}$
  - (This is why Cu can form Cu^{2+} and Cu^+ ions) Remember the 49'ers!!!!
- Chromium expect [Ar] $4s^2$ $3d^4$
  - However one of the 4s e- jumps into the 3d orbital and the really configuration is [Ar] $4s^1$ $3d^5$
  - This is why Cu can form Cu^{2+} and Cu^+ ions Remember the 49'ers!!!!
- Other exceptions: Ti = $4s^2$ $3d^2$, V = $4s^2$ $3d^3$, Cr = $4s^1$ $3d^5$, Mn = $4s^2$ $3d^5$
- Half filled orbitals.
- Silver expect [Kr] $5s^2$ $4d^9$
  - however one of the 5s e- jumps into the 4d orbital and the really configuration is [Kr] $5s^1$ $4d^{10}$
  - This is why Ag is Ag^{1+} not Ag^{2+} Remember the 49'ers!!!!
- **Scientists aren't sure of why it happens, We pretend all follow rules**

<u>Isotope Problems:</u> (Calculating Average Atomic Mass)

<u>Examples:</u>

There are two isotopes of carbon ^{12}C with a mass of 12.00000 amu (98.892%), and ^{13}C with a mass of 13.00335 amu (1.108%). What is the average atomic mass?

1)Multiply the percent in decimal form with the atomic mass for each.

2) Now add the numbers together, this is the average Atomic Mass

Practice:

There are two isotopes of carbon ^{63}Cu with a mass of 62.93 amu (69.09%), and ^{65}Cu with a mass of 64.93 amu (30.91%). What is the average atomic mass?

There are two isotopes of nitrogen, one with an atomic mass of 14.0031 amu and one with a mass of 15.0001 amu. What is the percent abundance of each? (this is a tuffy)

Homework: Unit 4 Wk 5 Isotope sheet and Unit 4 Review