Unit 2: Structure of Matter
Content Outline: The Periodic Table & Electron Configuration Models (2.8)

I. **Dmitri Mendeleev (1869)**
   A. He created a table of the known elements based upon *similar chemical properties* and known atomic masses.
   B. **Periodic Law**
      1. The *physical* and *chemical* properties of the element are *periodic* functions of their atomic number.
      2. When elements are *arranged* by increasing atomic number, elements with *similar properties* appear at *regular/periodic* intervals.

II. **Henry Mosely (1911)**
   A. He was a British Scientist.
   B. He modified Mendeleev's table to make it a function of increasing *atomic number*, not increasing atomic mass. (Remember, Atomic number is the number of *protons* in an atom; Atomic Mass is the number of *protons +neutrons.*)
      1. This makes the organizational pattern of the table more *reliable* and *predictable for newly discovered elements*.

III. **John William Strutt & Sir William Ramsay (1894)**
   A. John William Strutt was a British Physicist.
   B. Sir William Ramsay was a Scottish chemist.
   C. They discovered the first **noble gases**.
      1. These are the *least reactive* group of elements on the Periodic Table.
         a. The have a *full* **valence** (outer-most) electron energy shell.
         b. They are in column 18 of the table.

IV. **The Periodic Table *Reading* Format**
   A. The Periodic table appears as **rows** and **columns**.
      1. The **rows** are called **periods.** (Think of it "like" a chemical *sentence;* a *period* goes at the end.)
         a. Each period is associated with an *energy level* and *various orbitals/ sub-orbitals.*
            i. The first period <u>only</u> has an "s" orbital.
            ii. The second period has an "s" and "p" orbitals. Each "p" block has *3 sub-orbitals*. (Remember, orbitals can <u>only</u> hold 2 electrons.)
            iii. The third period is just like the second.
            iv. The fourth period has an "s", "d" and "p" orbitals. Each "d" block has *5 sub-orbitals.*
            v. The fifth period is the same as the forth row.
            vi. The sixth period has an "s", "f", "d", and "p" orbitals. Each "f" block has *7 sub-orbitals.*
               α. The portion of the "f" block in the sixth period is called the **Lanthanide series.**
               b. The portion of the "f" block in the seventh period is called the **Actinide series.**
                  * Most of these are *synthetic elements* (man-made in the laboratory) and radioactive (gives off energy).
      2. The vertical **columns** are called **families**.
         a. Each **family** of elements has *very similar* chemical properties.
         b. There exist 18 families on the periodic table.
            i. Family 1 is called the **Alkali Metals family.**
            ii. Family 2 is called the **Alkaline Earth metals family.**
               α. Families 1 and 2 are the "s" block/orbital.
               b. They are *highly reactive* with air, moisture, or most other elements.
                  * This is because family 1 has 1 *valence electron;* family 2 has *2 valence electrons.*
            iii. Families 3 – 12 are called the **Transitional Metals elements** and make up the "d" block/orbitals.
            iv. Families 13 -18 are collectively known as the **"p" block elements.**
               α. These elements are a collection of *metals, metalloids, non-metals, and gases.*

* The metalloids are *next to the zigzag line* running down through the "p" block in the table.

b. The "s" and "p" blocks *collectively together* are referred to as the **main-group elements.**

c. The family in column 17 is known as the **Halogen family**.

* They are *very reactive* elements as they are *missing*1 electron from having a *full* valence shell.

d. The family in column 18 is known as the **Noble gases**.

* These are the *least reactive* elements on the table, as they possess a *full* valence shell known as an "**octet**" or 8 electrons in a shell.

3. **Metals**
   a. These make up the majority of the Periodic Table.
   b. Metals are defined as *good conductors of electricity,* have a high *luster (shine)* and are malleable (capable of being beat, using hammers or rollers, into wires).
4. **Metalloids** ("oid" means "like a")
   a. These are "like" metals as some are able to conduct *some electricity. (A.K.A. as* **semi-conductors.**)
   b. They are located *between* the metals and the gases along the zigzag line of "p" block.
5. **Non-metals**
   a. These are mostly powders that are non-conductors of electricity.
6. **Gases**
   a. These floating elements that have no *definite shape or volume.* They take the shape and volume of their *container.*
   c. Most can conduct electricity. (Remember the Cathode-Ray tubes and neon signs.)

Unit 2: Structure of Matter
Content Outline: Basic Atomic Structure and Mass (2.2)

I. **Atom**
   A. The *smallest* particle of an element that *still retains* the chemical properties of that element.
   B. Atoms are composed of 3 *sub-atomic* particles:
      1. **Electrons** (Thomson proposed.)
         a. Electrons possess *negative electrical charges*.
         b. Electrons are found *orbiting* the nucleus of an atom, in what is referred to as the **electron cloud.** (They move at the speed of light and "create" a cloud-like appearance.)
         c. Electrons are $1/1837^{th}$ the mass of a single proton or single neutron.
      2. **Protons** (Rutherford, Geiger, and Marsden proposed.)
         a. Protons possess a *positive electrical charge.*
         b. Protons are found clumped together *within the nucleus* of an atom.
         c. Each proton has a mass of 1 atomic mass unit (AMU) or 1 Dalton (Named after John Dalton.)
      3. **Neutrons** (Rutherford proposed.)
         a. Neutrons possess <u>no</u> *electrical charged* and are therefore referred to as *neutral.*
         b. Neutrons are also found clumped together *within the nucleus* of an atom.
         c. Each neutron has a mass of 1 AMU or 1 Dalton.
   C. **Nuclear Forces**
      1. These are short-range proton-to-neutron OR proton-to-proton OR neutron-to-neutron *attractive forces* that help *hold together* the nucleus of an atom.
      2. These forces are *greater* than the repulsive *same charge electrical forces* exhibit by protons.
   D. **Atomic Radii**
      1. This term refers to the *relative size* of an individual atom of an element.
      2. It is measured from the *center of the nucleus to the outermost electron cloud*.
      3. It is measured in **picometers (pm).**
         a. A picometer is 1.0×10^{-12} meters. (So it is very, very small.)
      4. Charles-Augustin de Coulomb (1785)

a. He proposed **Coulomb Forces –** *attractions* that exist between *oppositely* electrically charged particles (protons & electrons) within a single atom.
b. The forces *directly affect* the atomic radii of an atom.
   i. *More protons* than electrons = radii *shrinking* (getting smaller) because the positive charge is greater than the smaller negative charges and pulls them in toward the nucleus.
   ii. *More electrons* than protons = radii *increases* (getting larger) because the electrons are *farther away* from the positive nucleus.
   iii. The *Natural state* of atoms has protons = electrons; so atomic radii are stable (not changing) for each element.

5. Atomic radii *can have an effect* on the *chemical properties* of an element.

II. **Atomic Number**

A. This term refers to the *number of protons* found within the nucleus of an atom for that element.
B. Each element has a *unique and identifying* number of protons.
C. The atomic number for each element led to the creation of the **Periodic Table.**
   1. The Periodic Table was originally created by Dmitri Mendeleev in 1869.
      a. He was a Russian Chemist.
D. The atomic number is usually written as *superscript* (above) the Elements Chemical symbol.
   1. Some of the symbols use the Latin term, instead of the English word like Iron, its symbol is Fe for "Ferrum".
      a. Latin is used because it is a "dead" language (will <u>not</u> change over time) and was the original language of science.
E. The Periodic Table was created based upon *increasing* Atomic Number.

III. **Atomic Mass Units (AMU)**

A. Also known as the **Mass Number.**
B. This term refers to the *total mass* of an atom of that element.
C. It is found by *adding* the number of *protons and neutrons* together.
   1. Each proton OR each neutron has a mass of 1 AMU or 1 Dalton.
   2. The Electrons' mass is *insignificant* as they are so small ($1/1837^{th}$ that of protons/neutrons).
D. The Atomic Mass is usually written as a *subscript* (below) the Element symbol.
E. This was *based on Carbon-12* as the standard element of measure. It has 12.0 AMU.

IV. **Isotopes**

A. This term refers to atoms of an element that have *different* masses (AMUs) because they have *different numbers of neutrons* within the atom; even though it is the same element because they have the *same number of protons*. (Remember, *protons* identify the element.)
   1. The isotopes behave relatively the same as the natural atom in terms of chemical properties.
   2. Some isotopes are *radioactive* (the nucleus is "breaking apart").
B. To find the number of neutrons:
   Start with AMU, subtract the # of protons (atomic number), and that leaves the number neutrons.

   AMU – # protons = # neutrons

C. How Isotopes are written chemically:
   1. **Hyphen notation** – symbol- number, For example: Carbon-14 OR C-14.
   2. **Nuclear notation** – AMU over Atomic Number symbol, for example $^{14}{}_{6}$ C.
D. **Nuclide**
   1. This term is used to refer to the *nucleus only* (no e- cloud) of an Isotope.

V. Average Atomic Mass

A. As some elements have *several* isotopes also present in nature, their masses must also be considered to find the *average* mass for an element. (As seen on the Periodic Table.)

B. How to calculate the *average* Atomic Mass of an element:

Step 1: Multiply the AMU for a single isotope by the % found in nature.

Cu 63 – AMU of 63 = 62.93 AMU; so 62.93 x 69.15% (nature) = 62.93 x .6915 = 43.52 AMU

Cu 65 – AMU of 65 = 64.93AMU; so 64.93 x 30.85% (nature) = 64.93 x .3085 = 20.03 AMU

Step 2: Add the all AMUs *together*.

43.52 + 20.03 = 63.55 AMU

Step 3: Round to two places after the decimal for *each* isotope calculation in Step 1.

Unit 2: Structure of Matter
Content Outline: History of the Atomic Model (2.1)

I. Democritus (400 B. C.)
   A. He was a Greek philosopher of science.
   B. First to use the term "**atom**" to describe the *basic* particle of nature. ("Atom" means "indivisible".)
      1. **Atom** – the *smallest* particle of an element that *still retains* the chemical properties of that element.

II. John Dalton (1808)
   A. He was a British schoolteacher.
   B. He was the first to propose an "Atomic theory" that contains the 5 following statements:
      1. *All matter* is composed of *extremely* small particles called "atoms".
      2. Atoms of a given *element* are *identical in size, mass, and other properties*; atoms of *different elements* differ in size, mass, and properties.
         a. This has since been modified based on discovery of Isotopes and Ions.
      3. Atoms cannot be *subdivided, created, or destroyed.*
         a. This has since been modified based upon current studies in quantum physics. Such examples include muons and quarks.
      4. Atoms of *different* elements combine in simple *whole-number* ratios to form chemical compounds.
      5. In chemical reactions, atoms are *combined, separated, or rearranged.*

III. Joseph John Thomson (1897)
   A. He was a British Physicist.
   B. He worked with glass gas-filled tubes referred to as **Cathode-Ray tubes**.
      1. The glass tubes were filled with a gaseous element under low pressure.
      2. He then passed an electrical current through the gas using a battery and wires.
         a. The electrical current caused the gas within the tube to intensely glow with a beam ("**ray**").
            i. Magnets could make the "ray" move/deflect in various directions.
            ii. The ray is being deflected by the *negative* charge of the magnet.
            iii. Negative charge *repels/deflects* like negative charges, such as electrons possess.
            iv. The ray is made of a negative charge that Thompson called **electrons.** (Since they were associated with the *electrical current.*)
         b. The electrical current came into the chamber (by a wire) at the **cathode** end. (The end where electricity *enters* the tube.)
         c. The electrical current left the tube on the **anode** end. (The end where the electricity goes *back into* the wire.)
         d. Hence the term **Cathode Ray tubes.**
   C. Further investigations using different elements in Cathode-Ray tubes confirmed that *every* element's atoms possess electrons.
   D. He proposed the **"Plum Pudding"** model of atoms.
      1. It stated that *negatively* charged electrons are evenly placed inside a *positively* charged mass.

IV. Robert A. Milliken (1909)
   A. He was an American Physicist.
   B. He was the first to measure the *charge* and *mass* of an electron.
   C. The symbol for an electron is: e-
      1. **Electron charge** = 1.602×10^{-19} Coulombs.
         a. This is an extremely small quantity of energy.
      2. **Electron mass** = 9.11×10^{-31} kg
         a. Electrons are 1/1837th the mass of a *single proton or neutron*.
         b. This is a very, very, very small amount and size.
   C. Milliken's experiments allow for 2 **inferences** (conclusions based upon *evidence* and *reasoning*) to be made:

1. Because atoms, in the *natural* state, are *electrically neutral*, they <u>must</u> also contain an equal amount of positively charged particles.
2. Because electrons have so little mass, atoms <u>must</u> contain other particles with much greater mass (protons & neutrons).

V. Ernest Rutherford, Hans Geiger, and Ernest Marsden (1911)
   A. Geiger and Marsden were *students* of Rutherford a New Zealand Physicist.
   B. They performed the **Gold Foil Experiment**.
   C. They used high-energy **alpha particle radiation** (2 protons & 2 neutrons *ejected* from a *decomposing, radioactive* element.) to bombard a piece of gold foil that was surrounded by a fluorescent screen.
      1. As alpha particles struck the fluorescent screen, they would produce a small *detectable burst of light.*
      2. As the experiment was running, they detected light burst mainly *behind* the gold foil, but also occasionally *all around the ring.*
         a. These bursts of light around the ring were because of the *positively charged* alpha particles being *deflected by positively charged particles* in the atoms of the foil.
         b. The positively charged particles of an atom became known as **protons**.
            i. Just as with the electrons, positive charges *repel/deflect* like positive charges.
         c. As most of the bursts of light occurred *behind* the gold foil, they concluded that the *majority* of space in an atom is "empty space" that the alpha particles travelled through and never hit anything.
   D. Rutherford proposes the idea of the neutrally charged **neutron** particle in 1920 too.

VI. Niels Bohr (1913)
   A. He was also a student of Rutherford's.
   B. He proposed the **Bohr model** of an atom.
      1. The electrons move in a circular pattern around the positively charged center. (Much like the planets revolve around the sun.)

VII. Dmitri Ivanenko & Victor Ambartsumian (1930)
   A. These gentlemen were Russian Physicists.
   B. They proposed a model of the nucleus of an atom that is composed of positively charged **protons** and neutral charged particles (neutrons).

VIII. James Chadwick (1932)
   A. He was a British Physicist.
   B. He proved that the nucleus is definitely composed of protons and neutrons through his experiments with **alpha particle radiation.**

Unit 2: Structure of Matter
Content Outline: Bohr Model of atoms and Electron Energy (2.5)

I. Niels Bohr (1913)
   A. He was a Danish Physicist.
   B. Proposed the Bohr Model of Atom structure.
      1. Electrons travel in set paths *around the nucleus,* called **orbits** or **energy levels**.
      2. Each orbit corresponds with an *energy level.*
         a. Electrons have a *natural tendency* to occupy the *lowest (most stable)* energy level *first*.
            i. The *lowest* energy level (**ground state**) is the *closest* to the nucleus.
            ii. This is related to **Coulomb forces –** *opposite* electrical forces *attract*.
            iii. The *farther* away from the nucleus the *greater* the Potential Energy for that electron.
            iv. The *closer* to the nucleus the *less* Potential Energy for that electron.
         b. Electrons can absorb energy (**absorption**) from their *surroundings* from *another energy source*, such as sunlight energy (A.K.A **electromagnetic energy**).

i. Electrons that *gain energy* (**absorption**) are said to be **"excited"**.
ii. Electrons that *lose energy* (**emission**) emit *light* as they *return* to a more stable (less energy) grounded state.
iii. The *unit of light energy* is referred to as a **photon**.
iv. The unit of measurement for the *energy lost OR gained by an atom* is a **quantum**.

II. Electrons as *Particles* & *Waves*
A. Electrons can move as *particles* around the nucleus because they have *mass* (if ever so small).
B. Electrons, as they are moving, move in *wave-like fashion.* (Like waves in the Gulf of Mexico… up, down, up, down as it moves forward.)
C. *Properties* of waves:
1. **Wavelength (λ)**
a. This is defined as the *distance* between two identical points (such as crest –*top* or ebb- *bottom*) on *adjacent waves*.
i. As it is *distance*, some unit of measurement of distance (meter) is used for wavelength, usually nanometer (nm OR 10^{-9}).
2. **Frequency (v)**
a. This is defined as the *number of waves* that pass a given point in a *specified time*, usually seconds.
b. Frequency is expressed in **Hertz (Hz)** or waves/sec.
i. Heinrich Hertz defined 1 wave/sec = 1 Hertz.
3. **Speed of light (c)**
a. Electrons travel at the speed of light.
b. Waves are *measured against* the speed of light (**electromagnetic radiation**).
c. **$C = \lambda v$** is the equation for the speed of light.
i. As light speed never changes, it is considered to be a **constant** at 3.00×10^8 m/sec.
ii. The properties of light are *inversely proportional.*
α. As wavelength *decreases*, frequency *increases.*
b. As wavelength *increases*, frequency *decreases.*
d. As electrons *gain* more energy, they become excited and get *farther* from the nucleus.

III. **Electromagnetic Spectrum** (Light Energy)
A. This term refers to the *whole* spectrum (variations) of **electromagnetic radiation.**
1. **Radiation** is used to define the *wave-like* movement of light particles.
2. Light moves at 3.00×10^8 m/sec.
3. The electromagnetic spectrum includes: x-rays, microwaves, visible (white) light, ultra-violet light, infrared light, and radio waves.

Unit 2: Structure of Matter
Content Outline: Photoelectric Effect & Emission Spectrum (2.6)

I. **Photon**
   A. Albert Einstein proposed the concept of "photon" in 1905.
   B. A **Photon** is a *unit of light energy* having *no mass* and possessing a *single quantum of energy.*
      1. **Quantum**
         a. *Minimum quantity of energy* that can be *lost* or *gained* by an atom.
         b. It was proposed by German physicist Max Planck in 1900.
         c. This sets the field of Quantum Physics (nanoscale physics) in motion.
         d. Planck wins the Nobel Prize in 1918 for this work.

II. **Photoelectric Effect**
   A. The **emission** (ejection) of an electron from a *metal surface* when light shines on the surface.
      1. This shows a direct connection between light and light possessing energy.
      2. This light energy is Einstein's photon.
   B. The light <u>has</u> to be of a *minimum frequency* in order for the effect to take place.
   C. Each metal requires a *different* frequency of light.

III. **Planck's Constant Theory**
   A. This *tries to explain* the **Photoelectric Effect** by proposing a *relationship between a quantum of energy* and *the frequency of radiation*. Remember, light is considered electromagnetic radiation, so the frequency changes with the various forms of light/radiation. This is called the **Electromagnetic Spectrum.**
   B. **E = hv**
      1. E = energy for a *quantum of radiation.*
         a. It is measured in **Joules (J)**.
      2. v = *frequency* of the radiation.
         a. It is measured in **waves/sec. (s) $^{-1}$ or Hertz**
      3. h = **Planck's Constant**
         a. Defined as **6.626 x 10^{-34} J * s** (* = times)
   C. Planck-Einstein Relation
      1. Albert Einstein expanded on Planck's work in 1905.
      2. He proposed that light has a *combination* of wave properties and particle properties.
         a. Each particle of light carries 1 quantum of energy.
      3. **E = hv** (Planck's version) then becomes **E_{photon} = hv** (Einstein's version).
      4. Matter can <u>only</u> absorb Electromagnetic radiation (light) in *whole* number (1, 2, and so forth) quantities of photons.
      5. In order for a *single electron* to be emitted from the metal surface, the electron <u>must</u> be struck by a *single photon* possessing at *least the minimum amount of energy* to eject the electron.
         a. This minimum amount of energy is *directly related to the frequency.*
            i. The *greater* the frequency – the more possible to emit an electron from the surface.
            ii. The *smaller* the frequency – the less likely to emit an electron from the surface.
         b. Different metals *require* different frequencies for the Photoelectric effect to take place.

IV. **Emission Spectrum** of Hydrogen
   A. This is an *expansion* of the Cathode-Ray Tube experiment.
      1. It uses Hydrogen gas (which glows pink) and a glass prism (triangular shaped piece) placed in the path of the light ray.
         a. The light ray *split* into 4 different colors (red, green, blue, and purple).
            i. These become known as the **light emission spectrum**.
            ii. Each color represents a *fixed* quantity of energy for an *excited electron*.
            iii. It is later added for other frequencies of light, such Infrared and Ultra-violet.

V. These experiments set the groundwork for the **Modern Quantum Atomic Model of atoms.**

Unit 2: Structure of Matter
Content Outline: Modern Quantum Model of Atoms (2.7)

I. Louis de Broglie (1924)
  A. He was a French Scientist.
  B. He proposed that electrons (e-) had wave-like properties using Bohr's Atomic model.
    1. He stated that electrons are *confined* to areas around the nucleus, known as **orbits.**
    2. Using the Planck-Einstein Relation, he reasoned that electrons have *specified energies/frequencies.*
  C. **Diffraction** (bending of waves) experiments with electron beams and light beam showed similar results, proving electrons travel like light… in waves.
  D. **Interference** (waves overlapping) experiments with waves and electron beams showed similar results, proving electrons have particle like properties associated with energy.
    1. Some areas increased in energy as a result of overlap; other areas decreased in energy.

II. Werner Heisenberg (1927)
  A. He was a German Physicist.
  B. He calculated that electrons and photons have about the *same amounts of energy.*
    1. Photons are used to help *detect* the presence of electrons.
      a. When they collide, the electron is deflected in a *random* direction.
  C. **Heisenberg's Uncertainty Principle**
    1. This states that it is *impossible* to determine exactly both *position* and *velocity* simultaneously of an electron in an orbit.

III. Erwin Schrödinger (1926)
  A. He was an Austrian Physicist.
  B. He helped develop the **Quantum Theory of Atoms.**
    1. This theory tries to describe, by mathematics, the wave-like properties of electrons and other very small particles.
    2. This re-enforces that electrons travel in **orbitals**.
      a. **Orbital** is defined as a 3- Dimensional *region* around a nucleus that *indicates* the *probable location* of a *single* electron within an orbital.
  C. **Quantum Numbers**
    1. These are *four* numbers, written consecutively, that *specify* the properties of *atomic orbitals* and properties of the *electrons* within an orbital.
       For example: n,l,m,s 1,0,0,1/2 or 2,1,1,-1/2

    2. The *first number:* the **Principle Quantum number.**
      a. Symbolized as "**n**".
      b. This states the *electron energy level* (or shell).
      c. Uses only whole numbers 1-7.
      d. As "n" *increases*, the *distance* from the nucleus also *increases*; therefore the *Potential Energy* of that electron also *increases.*
    3. The *second number:* the **Angular Momentum Quantum number.**
      a. Symbolized as "***l***".
      b. This states the *shape of orbitals* or *sub-orbitals.*
        i. 0 = s (sphere shaped); there is only 1 orbital/energy level.
        ii. 1 = p (dumbbell shaped); there can be up to 3/energy level.
        iii. 2 = d (4 leaf clover shaped with a possible ring); there can be up to 5/energy level.
        iv. 3 = f (too complex to describe); there can be up to 7/energy level.
    4. The *third number:* the **Magnetic Quantum number.**
      a. Symbolized by "**m**".
      b. This states the *orientation* (up/down, left/right, forward/backward) of an orbital around the nucleus.

i. s = 0
ii. p = 1, 0, -1 (are possible)
iii. d = 2, 1, 0, -1, -2 (are possible)
iv. f = 3, 2, 1, 0, -1, -2, -3 (are possible)

5. The *fourth number:* the **Spin Quantum number.**
   a. Symbolized as "**s**"
   b. As each orbital can only hold two *negatively* charged electrons, they must be opposites in spin (motion).
   c. We use either +1/2 (up) or -1/2 (down) for this.

IV. **Orbital Electron Configuration Models**

A. These are used to help *represent* the electron arrangement of an atom; electrons are represented by numbers such as $1s^1$ for Hydrogen or $1s^2$ $2s^2$ $2p^5$ for Fluorine or as arrows, either 🡅 🡇
B. It is important to remember that each orbital can <u>only</u> hold 2 electrons maximum.
C. You "build the model" starting from the **ground state** (lowest energy level) and work upward, starting at $1s^1$; the exponent states the number of electrons present at that level… hint look at the periodic table *ROWS*. Highest is row 7.
D. This uses 3 rules:
   1. **Aufbau Principle** ("Aufbau" is German for "to build".)
      a. Electrons occupy the *lowest* energy level first that can receive the electron. Starting at 1s and work toward row 7.
   2. **Pauli Exclusion Principle**
      a. Contributed by Wolfgang Pauli, an Austrian physicist.
      b. No 2 electrons, *in the same atom*, can have the same 4 Quantum numbers.
   3. **Hund's Rule**
      a. Contributed by Friedrich Hund, a German Physicist.
      b. States that orbitals and sub-orbitals of *equal energy* are each occupied by a solo electron *before* a *second* electron is entered into the orbital or sub-orbital and all electrons in singly occupied orbitals/sub-orbitals must have the *same spin number*. (Start with =1/2 🡅; then go back and use -1/2 🡇)

V. Noble Gas Electron Configuration

A. This method is used *to shorten* the traditional model building.
B. Starting at Neon (Ne), you can begin the notation using a Noble gas first.
   1. The Noble Gas must be in [ ]. Then start with the next energy row.