UNIT 2: ATOMS, MOLES AND THE PERIODIC TABLE

Chapter 3: Atoms and Moles

3.1: Substances Are Made of Atoms

Early Fundamental Chemical Laws

1. **Law of Conservation of Mass**: - mass is neither created nor destroyed in a chemical reaction. (*Lavoisier*)

Example 1: 23.04 g of ethanol is burned in an alcohol burner. The carbon dioxide gas and water vapour were collected and found to have masses of 44.0 g and 27.0 g respectively. Determine the mass of oxygen gas used in this combustion reaction?

Ethanol + Oxygen → Carbon Dioxide + Water Vapour
23.04 g ? g 44.0 g 27.0 g

Mass of Ethanol + Mass of Oxygen = Mass of Carbon Dioxide + Mass of Water Vapour
23.04 g + Mass of Oxygen = 44.0 g + 27.0 g
Mass of Oxygen = (44.0 g + 27.0 g) − 23.04 g

Mass of Oxygen = 48.0 g

2. **Law of Definite Proportion**: - the same compound always contains exactly the same proportion of elements by mass. (*Proust*)

Example 2: Water contains about 8 parts oxygen to 1 part hydrogen by mass. A 192 g of unknown liquid compose of hydrogen and oxygen contains 12 g of hydrogen. Is the unknown liquid water? Justify your response.

192 g total – 12 g of hydrogen = 180 g of oxygen

$$\frac{180\text{ g oxygen}}{12\text{ g hydrogen}} = \frac{15\text{ parts oxygen}}{1\text{ part hydrogen}}$$

Since the ratio between oxygen and hydrogen is 8:1 in water, the unknown liquid is NOT water.

3. **Law of Multiple Proportion**: - when two elements form a series of compounds, the ratios of the masses of the second element that combine with the first element can always be reduced to small whole numbers. (*Dalton*)

Example 3: State the ratios of hydrogen between the following hydrocarbon compounds.

Hydrocarbons	Mass of Carbon per 1 g of Hydrogen
Compound A	2.973 g
Compound B	3.963 g
Compound C	4.459 g

$$\frac{A}{B} = \frac{2.973\text{ g}}{3.963\text{ g}} = 0.7501892506 \approx 0.75 \qquad \boldsymbol{\frac{A}{B} = \frac{3}{4}}$$

$$\frac{B}{C} = \frac{3.963\text{ g}}{4.459\text{ g}} = 0.8887642969 \approx 0.888\ldots \qquad \boldsymbol{\frac{B}{C} = \frac{8}{9}}$$

$$\frac{C}{A} = \frac{4.459\text{ g}}{2.973\text{ g}} = 1.49983182 \approx 1.5 \qquad \boldsymbol{\frac{C}{A} = \frac{3}{2}}$$

In 1808, a British scientist by the name of ***John Dalton*** published his theory of atoms that would have profound effects on the philosophy on chemistry and physics. (The word atom comes from the Greek word *atomos*, which means indivisible. A Greek philosopher ***Democritus*** in 5^{th}-century BC first suggested this concept). The **Dalton's Atomic Theory** can be summarized as:

1. **All elements are made up of tiny particles called atoms that cannot be subdivided, created or destroyed.**
2. **The atoms of a particular element are identical and have the same physical and chemical properties.**
3. **Atoms of different elements have different physical and chemical properties.**
4. **Chemical compounds are formed when different kinds of atoms combine together in simple, whole number ratios. A particular compound always has the same relative numbers and types of atoms.**
5. **Chemical reactions deal with the rearrangement of the atom, which changes the way they are combined together. There is no change to the atoms themselves (they cannot be created or destroyed) in a chemical reaction.**

Assignment
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3.2: The Structure of Atoms

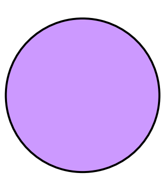
The practice of using symbols to represent elements can be traced back to the ancient Greek alchemists. Their purpose was to find a chemical recipe to make gold from other less valuable metals. (We now know that it is only possible now if we can change the number of protons in the nucleus).

Early Chemistry Models

1. **Aristotle and the Four Basic Elements**: - Aristotle beloved that all matters are made up of Fire, Air, Earth and Water in different proportions with four qualities of Hot. Dry, Moist and Cold.
2. **Democritus and the Indivisible Atoms**: - Democritus proposed that all matter are made up of indivisible particles called "atomos" (Greek for indivisible) that are characterized by the physical properties of the substance.

Alchemy: - using the Aristotle's idea, matters can be transformed by adjusting the portions of the four elements with the four qualities.
- led to a quest to turn common metal like lead into precious metal like gold.

Since the time of Dalton's Atomic Theory, scientists had improved upon his model to better explain the structure of an atom. The following is a summary of the different atomic models.

	**Dalton's Atomic Model:** In 1808, John Dalton proposed that all matter is made up of tiny particles called atoms. Atoms cannot be divided, created or destroyed. Different elements have different kinds of atoms. The difference is mainly due to the different atomic masses.

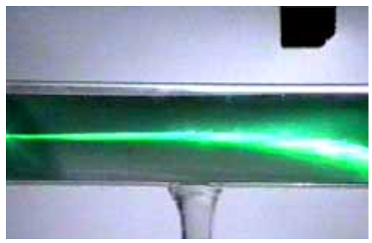
Benjamin Franklin: - discover the flow of electricity in the atmosphere in lightning.

Cathode Ray Tube: - a vacuum tube with negative and positive charged plates (cathode and anode) on either end to facilitate the flow of negative charged particles from one end of the tube to the other.

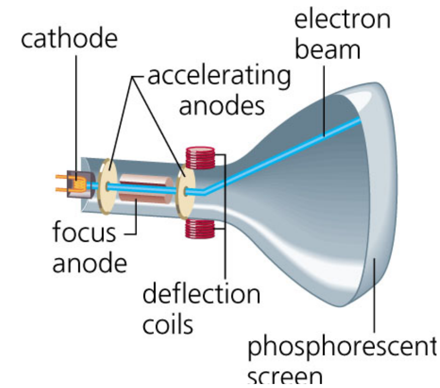
- the subatomic particle found when it is connected to a battery travels slower than the speed of light. Because of its negative charged, it is called an **electron**.
- **J.J. Thomson** found that when the beam of electrons deflects at an angle from its original path when it is placed in a magnetic field. He determined that the angle of the deflection can be expressed in a charge to mass ratio of the charged particle.

$$\text{Angle of Deflection} = \frac{\text{Charge of the Particle}}{\text{Mass of the Particle}}$$

A cathode ray tube (above)



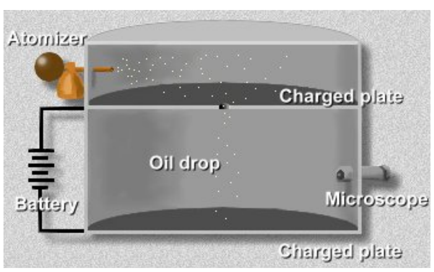
The beam is deflected in the presence of a magnet (above)



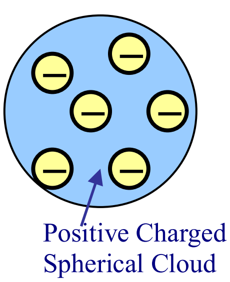
Television (commonly known as the tube) is an application of the cathode ray tube (above)

Millikan Oil Drop Experiment: - Millikan showed that charged oil drops can be suspended between two charged plates.

- using the fact that force of gravity and electric force of an oil drop are equalized in this arrangement, he was able to discover that the oil drops have charges with interval of **1.6×10^{-19} C** (Coulomb – a unit of charge). He deduced this must be the **elemental charge of an electron**.
- since the charge of an electron was known, he was able to compute **mass of an electron as 9.1×10^{-31} kg**.

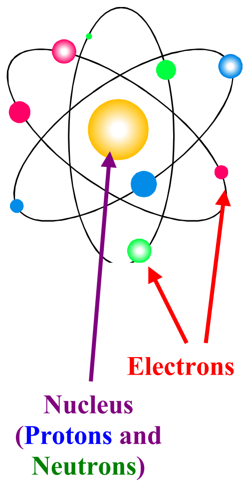


The Millikan Oil Drop Apparatus (above)



Plum Pudding Model:

In 1903, J.J. **Thomson** and Michael Faraday discovered **electrons** within an atom using a device called the **cathode ray tube**. Electrons are negatively charged subatomic particles with a **charge of −1**. The electrons were viewed as embedded in a positively charged spherical cloud. This is similar to the raisins distributed in a plum pudding. He also found the charge to mass ratio of an electron to be -1.76×10^{8} C/kg. In 1917, Robert **Millikan** used his **oil drop experiment** (by balancing the weight of an oil drop with electric force) to determine the elemental charge of the electron as -1.6×10^{-19} C and has a mass of 9.11×10^{-31} kg.



Nuclear Model:

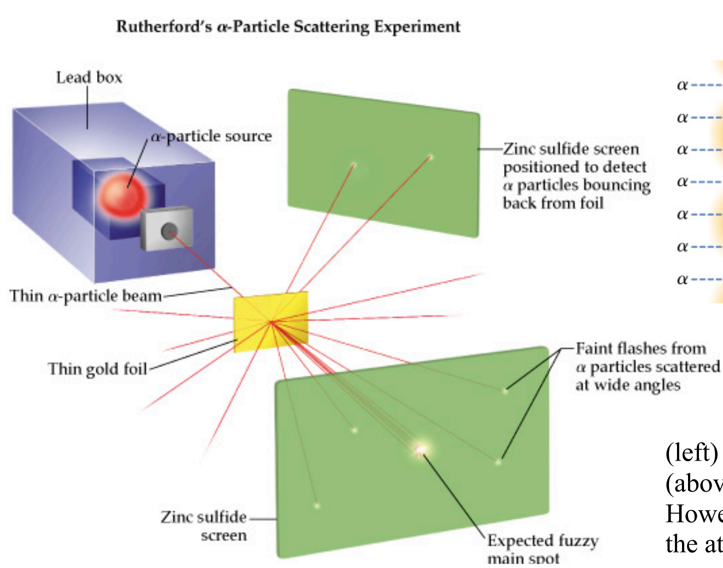
In 1895, Wilhelm **Röntgen** found that cathode rays caused glass and metals to emit unusual rays (he called them X-rays). Later, Antoine **Becquerel** discovered that a piece of radioactive metal (uranium) can expose a photographic plate. A student of Becquerel, **Marie Curie** termed this energy as radiation. Three types of rays were found as they deflected differently through a magnetic field. **Alpha (α) rays** – positively charged particles, **Beta (β) rays** – negatively charged particles, and **Gamma (γ) rays** – neutral particles with no deflection in a magnetic field.

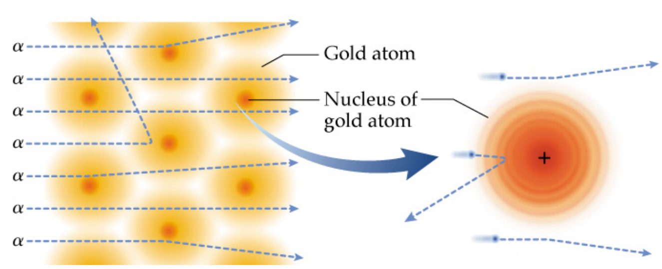
In 1912, Ernest **Rutherford** proposed the Nuclear Model for atoms after his famous **gold foil experiment** (he shot alpha particles into a piece of gold foil and found that the alpha particles passed through the gold foil – indicating the the atom is made up of mostly empty space). Earlier to this time, **E. Goldstein** discovered the **positively charged (+1)** subatomic particles called **protons**. Rutherford proposed that the protons are packed tightly together at the centre of the atom called the **nucleus**. In 1932, James Chadwick discovered **neutrons (no charged)**. Together, they suggested that the **nucleus was made up of both protons and neutrons** (the bulk of the atomic mass) since electrons are very light compared to the masses of the protons and neutrons.

On the other hand, negatively charged electrons move around the nucleus because of their attraction with the positively charged nucleus (contains protons). Since the nucleus is very small, the circling electrons make up almost all of the volume of the atom. If the atom has a size of a football field, the nucleus is about the size of a small nail at the centre of the field.

Rutherford's Gold Foil Experiment: - Rutherford fired alpha particles (positive particles that are more massive than electrons), to a sheet of gold foil. The result was that most alpha particles were not deflected and passed right through the gold foil. This leads to the conclusion that the "plum pudding" model was not correct (otherwise, the alpha particles would have deflected off the dense positive charges in an atom.

- instead, the ***atom must be organized such that most of its volume is really empty space***.
- a few alpha particles were deflected and even reflected back directly to the source. Hence, Rutherford concluded that the ***centre of an atom must consist of subatomic particles that are more massive than electrons and have positive charges***.

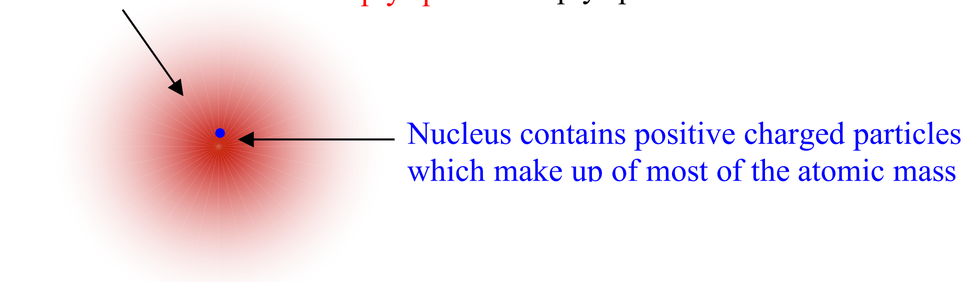




(left) An apparatus for Rutherford's Gold Foil Experiment.
(above) Most alpha particles pass through the gold atoms. However, some are deflected because they hit the centres of the atoms which are positively charged

Atomic Nucleus: - the central part of an atom that contains positive charged particles to balance the electrons of an atom.

Rutherford Nuclear Atomic Model: - the atom is consisted of a central nucleus where it consists of more massive subatomic particles that are positively charged. The electrons move around the nucleus where it is composed mainly of empty space.



Protons: - positively charged subatomic particles located at the nucleus (centre) of an atom.
- discovered by E. Goldstein around the time Rutherford performed his gold foil experiment.
- since it is known that electrons has a charge of -1.60×10^{-19} C, **protons have a charge of $+1.60 \times 10^{-19}$ C.** Goldstein also determine the **mass of a proton as 1.68×10^{-27} kg**.
- a group of two protons is called an **alpha (α) particle**.

Neutrons: - neutral subatomic particles also found at the nucleus of an atom.
- discovered by James Chadwick in 1932.
- it has the **same mass as an proton (1.68×10^{-27} kg)**. **Together with protons, they make up the total mass of an atom.**

Subatomic Particles	Charge	Relative Mass	Actual Mass	Location
Electrons (e^-)	-1	1	9.11×10^{-31} kg	Region around the center of the atom
Protons (p^+)	$+1$	1836.12	1.67×10^{-27} kg	Centre of the atom called Nucleus
Neutrons (n)	0	1836.65	1.67×10^{-27} kg	Inside the Nucleus with the protons

Atomic Number and Mass Number:

Recall from the Dalton's Atomic Theory, one of its points is that different elements have different atoms. For a long time, it was believed that the main difference between atoms of different elements is the **mass number** (the total mass of an atom). This is the mass characteristic of a given element**. The mass number of an element is relative to the mass of the carbon atom (6 protons and 6 neutrons with a mass number of 12)**. It is usually located at the right, top corner or directly below each element on the Table of Elements. For now, we will assume that **mass number has a unit of amu (Atomic Mass Unit).**

Because different elements have different mass number, the number of subatomic particles within an atom is also different for these elements. The **atomic number**, **a number assigned to each element based on its mass number**, is located at the top left corner of each element on the Table of Elements. **The atomic number is equated to the number of protons and electrons of that atom**. The **number of neutrons can be found by subtracting the mass number (rounded off whole number) with the atomic number**.

Atomic Number = Number of Protons and Electrons of an Atom

Number of Neutrons = Mass Number − Atomic Number

Example 1: State the Atomic Number, Mass Number, number of protons, neutrons, and electrons of the following elements.

17 35.45 **Cl** **Chlorine**	**Atomic Number = 17** (**17** p^+ and **17** e^-) **Mass Number = 35 (rounded)** # of Neutron = 35 − 17 = **18** n	**1** 1.01 **H** **Hydrogen**	**Atomic Number = 1** (**1** p^+ **and 1** e^-) **Mass Number = 1 (rounded)** # of Neutron = 1 − 1 = **0** n

Note: Because any given atom has the same number of protons and electrons (same atomic number), **all Atoms have a Net Charge of 0**.

Isotopes:

Isotopes are atoms of an element with the same atomic number but a different mass because of different number of neutrons. They are identified by the **element symbol** as well as the **mass number (total number of protons and neutrons)**. For a given mass of substance, there exist a certain percentage of isotopes. Some isotopes are stable. Others are unstable and they go through a decomposition process called **radioactive decay**.

Mass Number → A
Atomic Number → Z

$$^{A}_{Z}X$$

← **Element Symbol**

A common example is the isotope $^{14}_{6}C$ (Carbon-14: Carbon with an atomic mass of 14 amu, which has 8 n, 6 p^+ and 6 e^-). Naturally occur carbon contains 98.9 % of Carbon-12, 0.55% of Carbon-13 and 0.55% of Carbon-14. Chemists, physicists, archaeologists, geologists, and criminologists commonly use the carbon isotope. Because Carbon-14 is unstable and goes through radioactive decay at a definite rate, we can measure the amount of isotopes left in a substance to deduce its age. **Carbon-14 dating is a technique to date archaeological and geological findings by measuring the amount of Carbon-14 left in the artefacts**. Carbon-13 is used by chemists to assist in identifications of various chemical compounds.

Isotopes of other elements also have their uses in society. A tiny proportion of all water molecules (H_2O) compose of a hydrogen isotope called deuterium ($^{2}_{1}H$). Deuterium can be utilised as fuel in nuclear fusion reactors of the future. Other isotopes of various elements are used as **radiotracers**. These **are radioactive isotopes that can be introduced into organisms in food or drugs, and their pathways can be traced by monitoring their radioactivity**. These radiotracers have found their way into medical research. The list below shows some radiotracers and their medical applications.

Radiotracers	Area of the body examined	Radiotracers	Area of the body examined
$^{131}_{53}I$	Thyroid	$^{87}_{38}Sr$	Bones
$^{59}_{26}Fe$ and $^{51}_{24}Cr$	Red Blood Cells	$^{99}_{43}Tc$	Heart, Bones, Liver, and Lungs
$^{99}_{42}Mo$	Metabolism	$^{133}_{54}Xe$	Lungs
$^{32}_{15}P$	Eyes, Liver, Tumours	$^{24}_{11}Na$	Circulatory System

Recall that the superscript is the mass number of the isotope and the subscript is the atomic number. Since the mass number of the isotope is different than the mass number of the original element, the number of neutrons of the isotope is different than the number of neutrons of the element.

Meaning of Atomic Symbols:

With further studies on isotopes, scientists have found that **the atomic number, not the mass number, determines the characteristics of an atom**. Specifically, ***an element is defined by the number of protons in the nucleus***. When one changes the number of protons, he or she changes the identity of the element and a different element symbol is used.

Today, all elements and their symbols are listed in an orderly fashion with the help of the Periodic Table of Elements. The use of standardized symbols allows scientists from all over the world to share their knowledge despite the differences in language. Most elements are **monoatomic**. That means their atoms can exist individually ("*mono*" means one). Others are **diatomic**, atoms that exist in pairs ("*di*" means two). Some are **polyatomic**, atoms that exist in numbers more than one ("*poly*" means many). The table below shows all the diatomic and polyatomic elements.

Diatomic Elements (all the ~gens, including Halogens - second last column of the Periodic Table)	**Polyatomic Elements**	**Monoatomic Elements**
Hydrogen (H_2), Oxygen (O_2), Nitrogen (N_2), Fluorine (F_2), Chlorine (Cl_2), Bromine (Br_2), Iodine (I_2)	Phosphorus (P_4) Sulphur (S_8)	All other Elements. **Examples**: Helium (He), Iron (Fe), Calcium (Ca), Silver (Ag), Mercury (Hg)

Note: **Students should memorize all the diatomic and polyatomic elements. They are the only exceptions. All other elements are monoatomic.** Most symbols are recognizable from the name of the elements (Zinc: Zn; Carbon: C; Aluminium: Al). Others look somewhat different. This is because the symbols came from the elements' Latin names (Silver: Ag for "Argentum"; Gold: Au for "Aurum"). To save time, students should also familiarize themselves with the whereabouts of the elements on the Table.

Example 2: State the Atomic Number, Mass Number, number of protons, neutrons, and electrons of the following isotopes.

a. $^{87}_{38}Sr$

Atomic Number = 38
(**38 p^+** and **38 e^-**)
Mass Number = 87
# of Neutron = 87 – 38 = **49 *n***

b. $^{99}_{42}Mo$

Atomic Number = 42
(**42 p^+** and **42 e^-**)
Mass Number = 99
# of Neutron = 99 – 42 = **57 *n***

Example 3: Explain the differences between S_8 and 8 S.

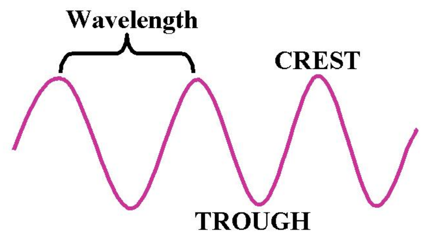
S_8 denotes a molecule that has 8 sulfur atoms ***chemically bonded together as one single molecular unit***.

8 S means that there are 8 ***individual sulfur atoms***.

Assignment

3.2 pg. 86 #1 to 4 (Practice); pg. 89 #1 & 2 (Practice); pg. 89 # 1 to 8

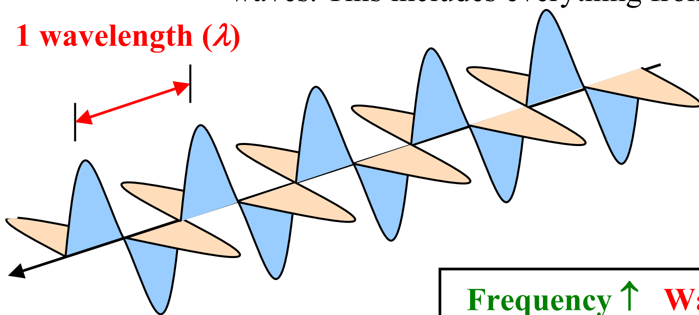
3.3: Electron Configurations



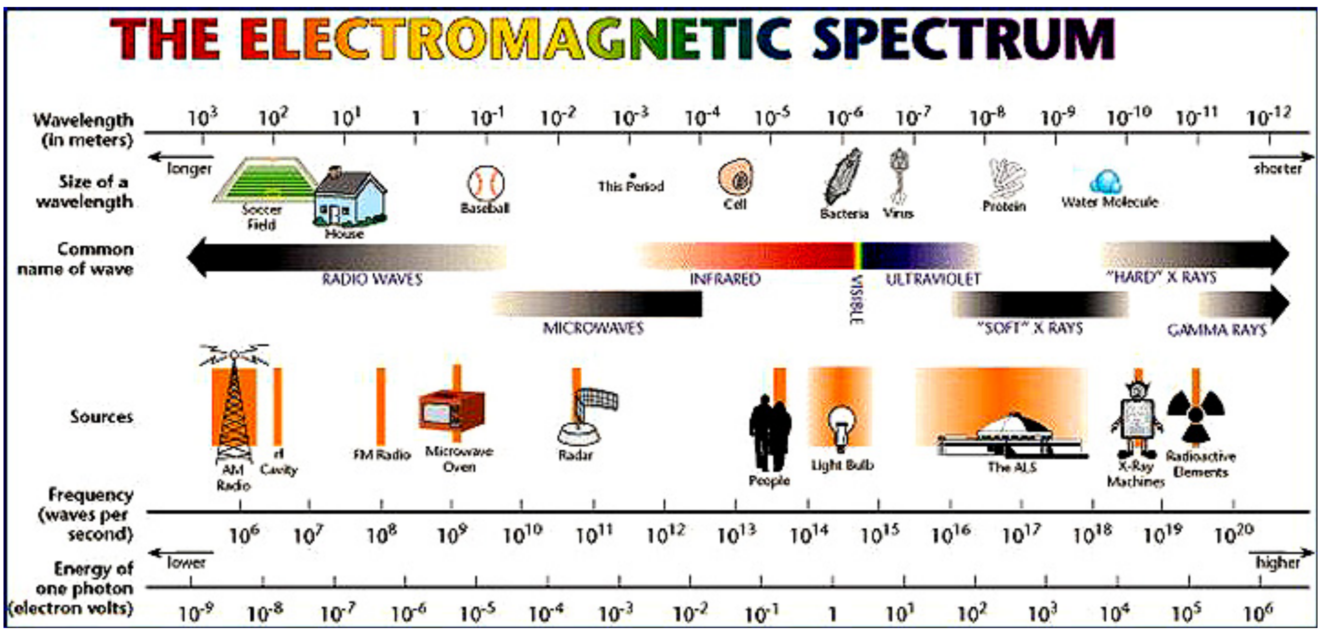
Wavelength: - the length of a wave (from crest to crest).
- measures in metres (m) or nanometres ($\times 10^{-9}$ m)

Frequency: - the number of wave in one second.
- measures in Hertz (Hz) or s^{-1}. (1 Hz = 1 s^{-1})

Electromagnetic (EM) Spectrum: - energy that travels at the speed of light in a form of perpendicular waves. This includes everything from cosmic rays to radio waves.

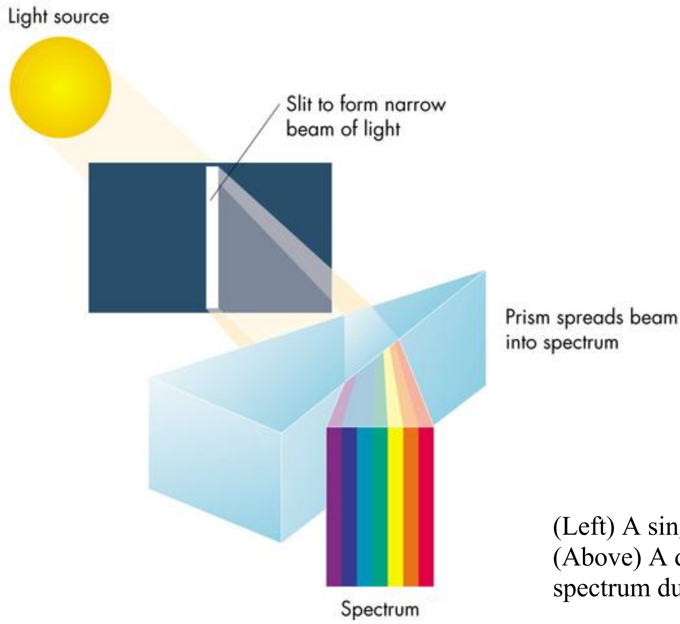


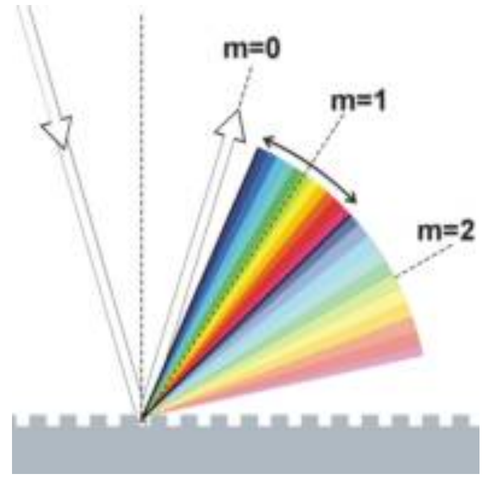
Frequency ↑ **Wavelength** ↓ **Energy** ↑



EM Wave	Frequency (Hz)	Wavelength (m)	EM Wave	Frequency (Hz)	Wavelength (m)
Cosmic Wave	10^{23}	10^{-15}	Infrared	10^{12}	10^{-4}
Gamma Wave	10^{20}	10^{-12}	Microwaves	10^{10}	10^{-2}
X-Ray	10^{18}	10^{-10}	FM Radio	10^8 or (100 MHz)	1 to 10
Ultraviolet	10^{16}	10^{-8}	Shortwave Radio	10^6 (1 MHz)	10^2
Visible	$(7.5 \text{ to } 4.3) \times 10^{14}$ (blue to red)	$(4 \text{ to } 7) \times 10^{-7}$ 400 nm to 700 nm (blue to red)	AM Radio	10^4 (10 kHz)	10^4

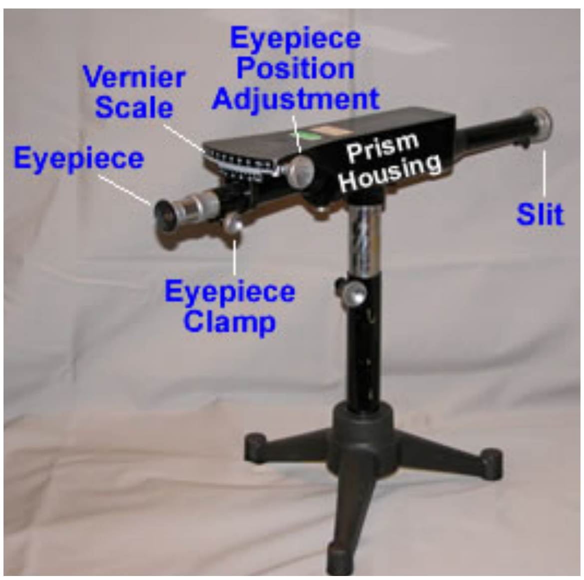
Diffraction Grating: - a surface that contains a series of prisms that diffracts incoming light into their individual wavelengths.

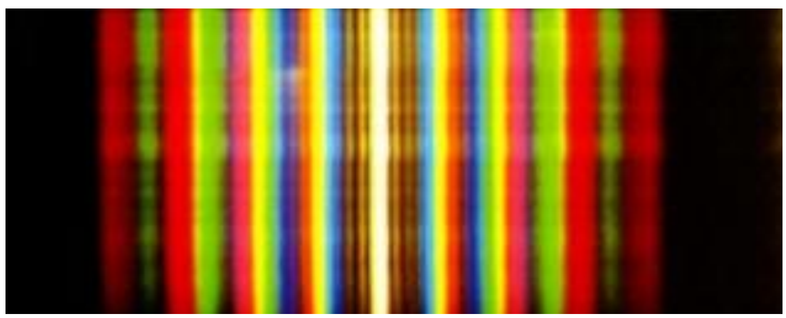




(Left) A single prism breaks up white light into the visible spectrum.
(Above) A diffraction grating is able to generate series of visible spectrum due to the uneven surface.

Spectroscope: - a device to break down light into its component colors using a diffraction grating





(Left) A Spectroscope. Light enters the slit and goes through the prism housing (may be replaced prism with a diffraction grating). The observer uses the eyepiece to see the image.

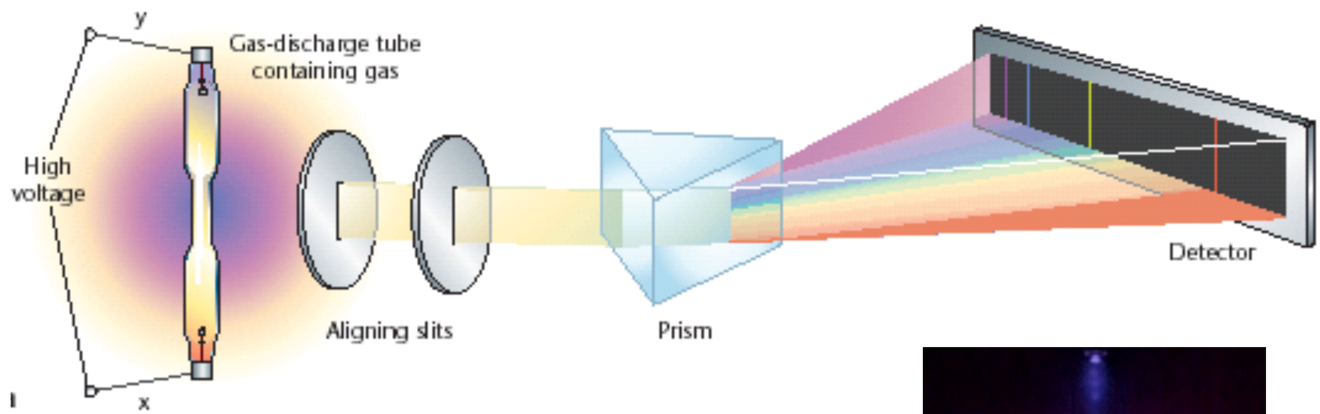
(Above) An image of white light diffracted through a spectroscope. Note the series of visible spectrum on either end of the central bright band.

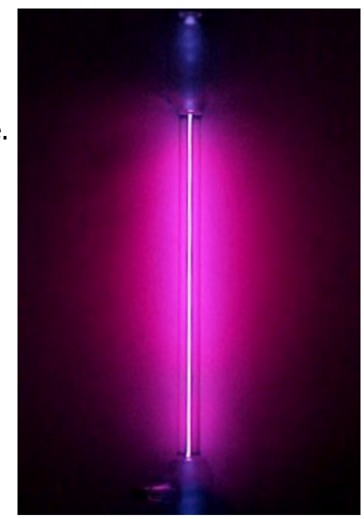
Quantum Hypothesis: - in 1900, Maxwell Planck reasoned that light energy moves in "*packets*" of energy called **quantum**.
- in 1905, Albert Einstein proved that the quantum hypothesis was correct as light can exist as a particle as well as a wave. Each light particle is called a **photon**.

Wave-Particle Duality of Light: - EM Radiation has characteristics of wave (reflection, refraction, and diffraction) and particles (collision and kinetic energy as demonstrated by Einstein).

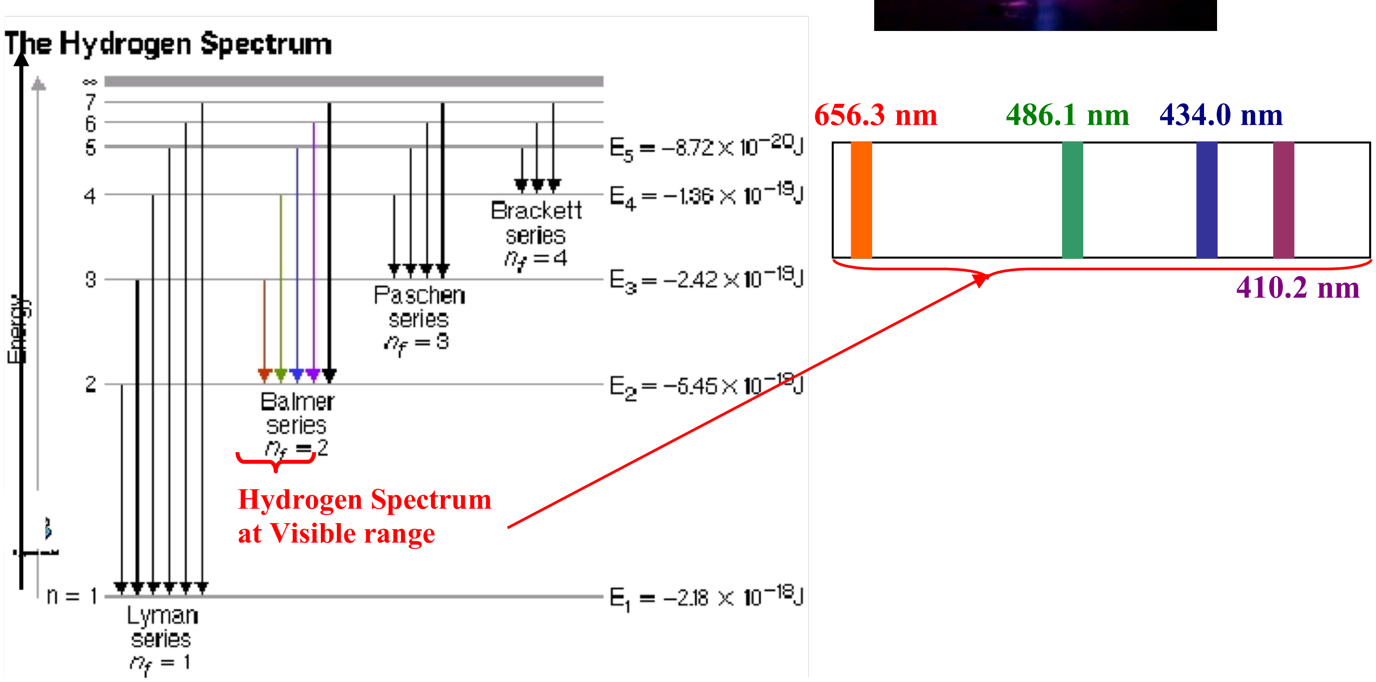
Neil Bohr and The Hydrogen Spectrum

In 1913, a Danish scientist named Neil Bohr performed the **Hydrogen Spectrum Experiment** by passing electricity through a hydrogen filled tube and put the light through a **spectroscope** (a device with diffraction grating to measure wavelength emitted).

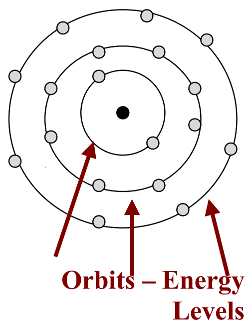




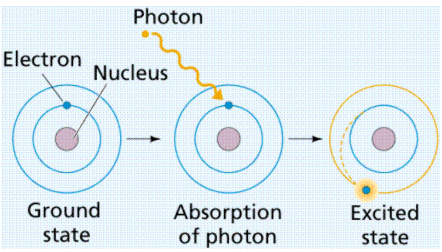
(Above) The inner working of a spectroscope and a gas discharge tube.
(Below) A common spectroscope (Right) A Hydrogen discharge tube



Bohr Atomic Model: - in 1913, Neil Bohr refined the Nuclear Model by suggesting that electrons move around the nucleus in specified **orbits**. These orbits are called **energy shells**.



1. **Electrons cannot exist between the orbits,** much like one cannot stand between steps on a set of stairs.
2. **The further the orbit is from the nucleus, the higher its energy level for the electrons in that orbit.** This is very similar to the planetary model of our Solar system. The **quantum number, *n*** is a natural number than indicates the energy level of an electron.
3. **An electron can "jump" from lower to higher shells when given sufficient amount of energy**.
4. **When an electron "relaxes" from a higher shell to a lower shell, a specific amount of energy is released in a form of photons.**



(Left) As a photon of light is applied to an electron at the lowest shell (ground state), it gains energy and move into a higher shell (excited state).

(Below) The hydrogen spectrum results a electrons move from higher shells to various lower shells. Each series is defined by the final state of the electrons. The Balmer Series, where the electrons rest at $n = 2$, generates specific lines in the visible range of the EM spectrum. **Since different atoms from different elements have specific energy levels, their spectrum is unique and serve as a fingerprint for identification.**

Electrons as Particles and Wave

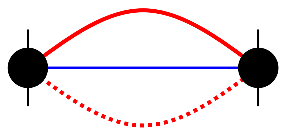
de Broglie Wavelength: - since light can behave like particles; particles can have wave properties.
- in 1924, Louis de Broglie proved that electrons can behave like waves when they are moving very fast. This led in the development of electron microscope when electrons are accelerated to 0.25% the speed of light (750,000 m/s), it can view image with a diameter of 10^{-9} m.



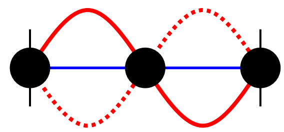
Some images from a scanning electron microscope. (Left) the head of a hypodermic needle, (top) a staple through a piece of paper, and (above) red blood cells

Schrödinger Wave Equation: - recognized that since electron can behave like waves (de Broglie). In 1926, Erwin Schrödinger proposed that electron could have **quantized energy** when it achieved a **standing wave** (wave that appears standing since it contains nodes that signifies exactly half a wavelength).
- when electrons move at a speed that achieved a standing wave, it is considered **self-reinforcing**. This is the reason why electrons can sustain themselves in an orbital and not fall into the nucleus.
- the specific wavelengths of these electron standing waves are what give rise to the unique energy levels of these orbitals. Hence, wavelengths emitted in a spectrum are specific for a certain type of atoms. This

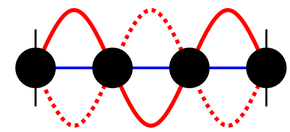
enables scientists to use various spectra as a tool to identify different material. (For example, astronomers pass lights from distant stars through a spectroscope to identify their compositions.)



Standing Wave
(2 nodes – 1 half-wavelength)

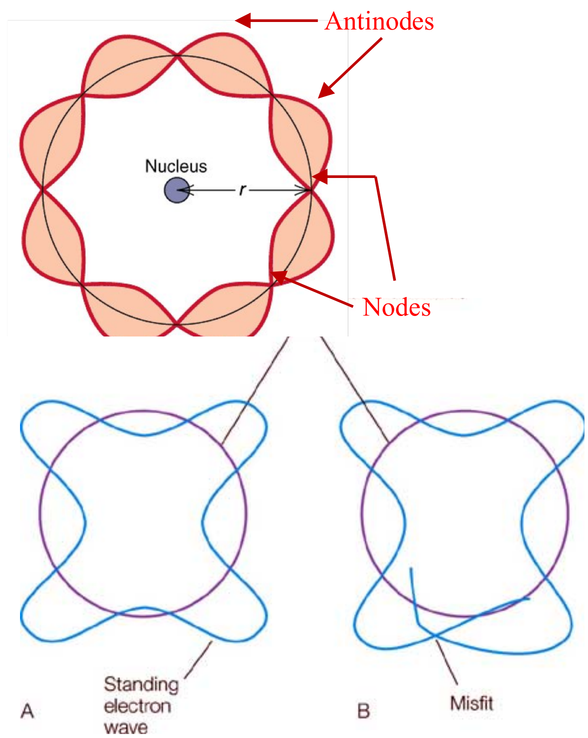


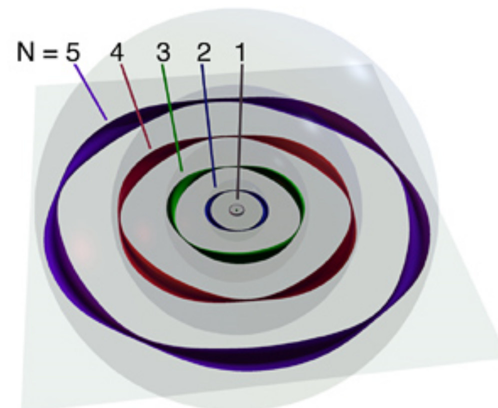
Standing Wave
(3 nodes – 2 half-wavelength)



Standing Wave
(4 nodes – 3 half-λ)

(Standing Wave Video at https://www.youtube.com/watch?v=wYoxOJDrZzw)





Bohr-de Broglie electron matterwave orbits shells 1-5

(Left) (A) An allowable orbit as the standing wave forms an integral number of wavelengths (In this case, $n = 4$.). (B) A standing wave that does not have integral number of wavelengths is considered a misfit. Hence, this orbit is not allowed. (Above) Electron waves for the 5 orbits (or electron shells) of a hydrogen atom.

1-Dimensional Animations of Electron Waves in a Column
http://www.glafreniere.com/matter.htm
2-Dimensional Animations of Electron Waves in a Box
http://www.youtube.com/watch?v=ufdS-fvoK8U

Heisenberg Uncertainty Principle:

It states that locating a particle in a small region of space makes the momentum of the particle (mass × velocity) uncertain; and conversely, measuring the momentum of a particle precisely makes the position uncertain. When we "observe" an electron using light (a photon), we would invariably change its position or speed (or both) due to their interactions. Hence, treating electrons as particles would create uncertainties and we could not pinpoint the speed and position of an electron in the atom. Therefore, we must employ the wave nature of particles to find their energy levels. This is why electrons are in a ***probable region of space*** where they are ***"likely" to be located***. ***In the quantum universe, we should really view electrons as waves.***

Videos and Information for Quantum Mechanics:

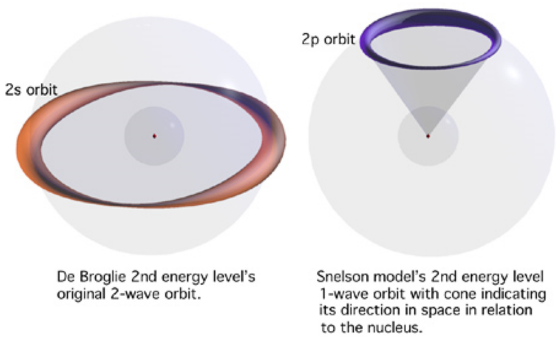
1. Video on Light through Single Slit: http://www.youtube.com/watch?v=KT7xJ0tjB4A
2. Video on Electron through Double Slits: https://www.youtube.com/watch?v=ZqS8Jjkk1HI
3. History of Quantum Physics (BBC): (3 Parts)
  a. https://www.youtube.com/watch?v=GOJFznzSZhM
  b. https://www.youtube.com/watch?v=CYQwrAhT7HA
  c. https://www.youtube.com/watch?v=KFS4oiVDeBI
4. Video on Properties of a Quantum Electron: http://www.youtube.com/watch?v=uq1h6jg61yI
5. Important People of Quantum Mechanics:
http://doctortang.com/AP%20Chemistry/Historical%20Development%20of%20Quantum%20Mechanics.pdf

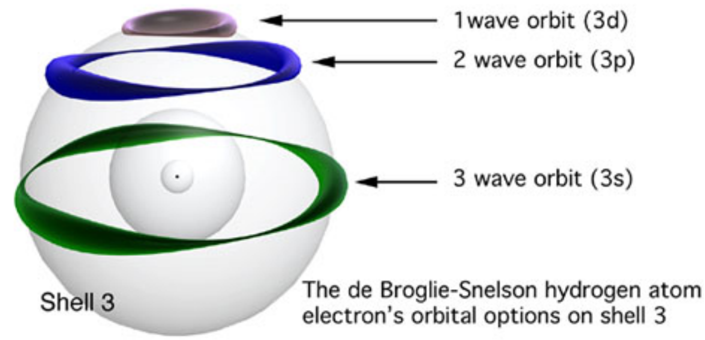
Quantum Numbers: - a series of number that ***describe the distribution of electrons in hydrogen and other atoms***. They are derived from the *mathematical solutions of the Schrödinger Wave Equation* for the hydrogen atom. There are ***four sets of quantum numbers used to describe any single electron***.

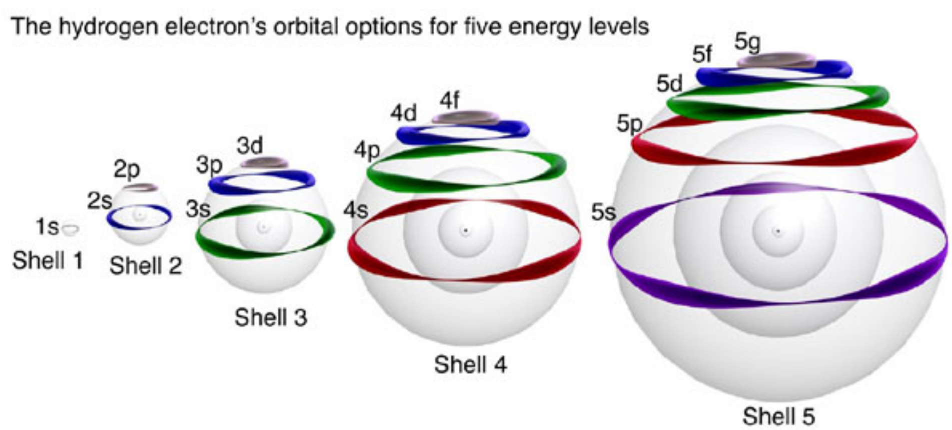
a. **Principal Quantum Number (*n*)**: - natural number values {1, 2, 3, …} are used to describe the energy and size of the orbital.

b. **Angular Momentum Quantum Number (ℓ)**: - whole numbers $\{0 \leq \ell \leq (n-1)\}$ are used to indicate the shape of the atomic orbitals.

$\ell = 0$ (*s* orbital)	$\ell = 1$ (*p* orbital)	$\ell = 2$ (*d* orbital)	$\ell = 3$ (*f* orbital)	$\ell = 4$ (*g* orbital)
"sharp"	"principal"	"diffuse"	"fundamental"	(follows the alphabet after *f*)

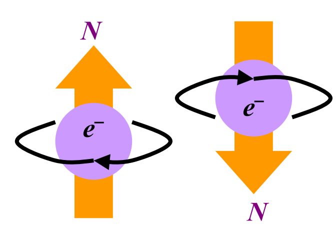






c. **Magnetic Quantum Number (m_ℓ)**: - integral numbers $\{-\ell \leq m_\ell \leq \ell\}$ are used to show the orientation of the orbital in space relative to the other orbitals in the atom.
- each m_ℓ value represents an atomic orbital.

Electron Spin: - when electron spins clockwise, it creates a magnetic north pole in the upward direction. Conversely, when electron spins counter-clockwise, it creates a magnetic north pole in the downward direction.



d. **Electron Spin Quantum Number (m_s)**: - values of ±½ to denote the electron spin direction.

Check out these websites at:

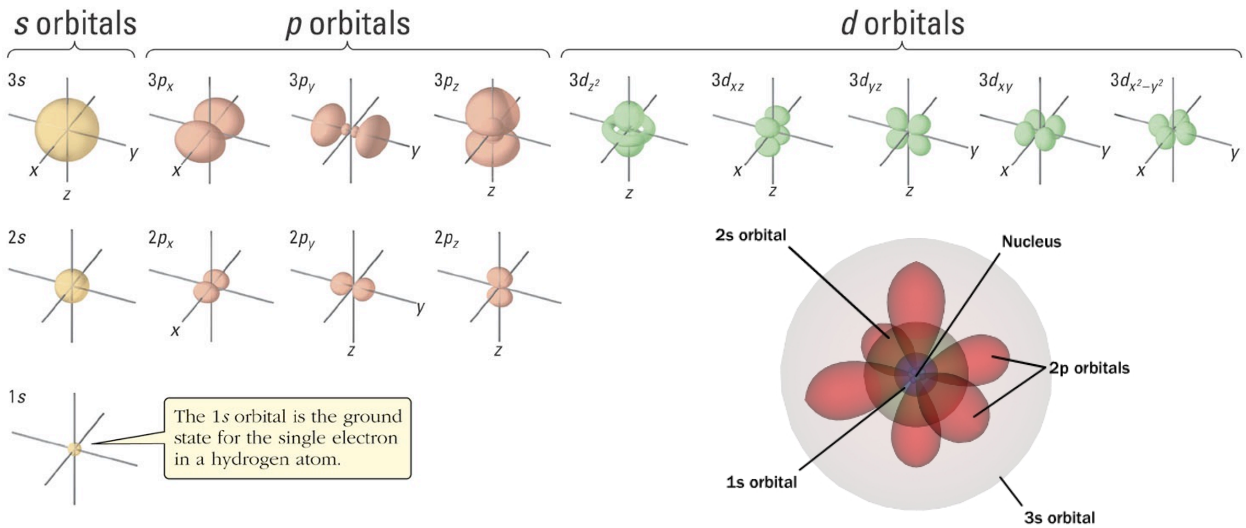
1. **Schrödinger Equation in Three Dimensions:** **http://hyperphysics.phy-astr.gsu.edu/hbase/quantum/sch3d.html#c1**
2. **Quantum Numbers of Hydrogen Atom:** **http://hyperphysics.phy-astr.gsu.edu/hbase/quantum/hydsch.html#c1**

Subshells: - are electron designations as indicated by the first two quantum numbers (Principal Quantum Number (n) and the letter used for Angular Momentum Quantum Number (ℓ)).

Atomic Orbitals: - are electron designations as indicated by the first three quantum numbers (Principal Quantum Number (n), the letter used for Angular Momentum Quantum Number (ℓ), and the Magnetic Quantum Number (m_ℓ)).

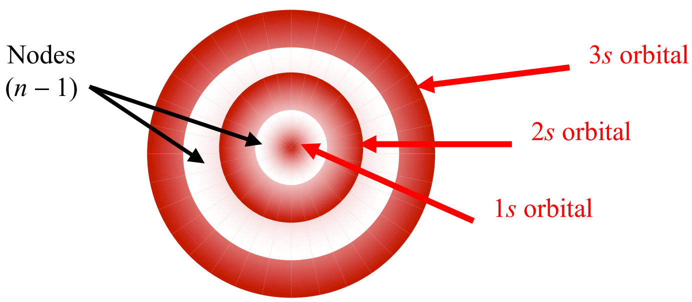
n	ℓ	Subshell	m_ℓ (Orientation of Orbitals)	Number of Orbitals in each Subshell	Total Number of Orbitals in each Energy Level
1	0	1s	0	1	1
2	0	2s	0	1	4
	1	2p	−1, 0, 1 (p_y, p_z, p_x)	3	
3	0	3s	0	1	9
	1	3p	−1, 0, 1 (p_y, p_z, p_x)	3	
	2	3d	−2, −1, 0, 1, 2 ($d_{x^2-y^2}$, d_{yz}, d_{z^2}, d_{xz}, d_{xy})	5	
4	0	4s	0	1	16
	1	4p	−1, 0, 1 (p_y, p_z, p_x)	3	
	2	4d	−2, −1, 0, 1, 2 ($d_{x^2-y^2}$, d_{yz}, d_{z^2}, d_{xz}, d_{xy})	5	
	3	4f	−3, −2, −1, 0, 1, 2, 3 ($f_{z(x^2-y^2)}$, $f_{y(z^2-x^2)}$, f_{y^3}, f_{z^3}, f_{x^3}, $f_{x(z^2-y^2)}$, f_{xyz})	7	
5	0	5s	0	1	25
	1	5p	−1, 0, 1 (p_y, p_z, p_x)	3	
	2	5d	−2, −1, 0, 1, 2 ($d_{x^2-y^2}$, d_{yz}, d_{z^2}, d_{xz}, d_{xy})	5	
	3	5f	−3, −2, −1, 0, 1, 2, 3 ($f_{z(x^2-y^2)}$, $f_{y(z^2-x^2)}$, f_{y^3}, f_{z^3}, f_{x^3}, $f_{x(z^2-y^2)}$, f_{xyz})	7	
	4	5g	−4, −3, −2, −1, 0, 1, 2, 3, 4 ($g_{xy(x^2-y^2)}$, g_{zy^3}, $g_{z^2(x^2-y^2)}$ g_{z^3y}, g_{z^4}, g_{z^3x}, g_{z^2xy}, g_{zx^3}, $g_{(x^4+y^4)}$)	9	

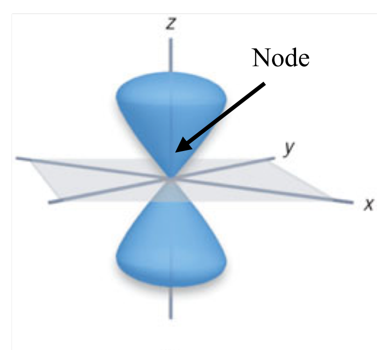
Atomic Orbital of the first three energy levels:



Check out the Hydrogen Atom Orbital Viewer at http://www.falstad.com/qmatom/

Nodes: - the areas between orbitals where there is zero probability of electron distribution.
- however, electrons can "slide" through the node when it takes on the properties of a wave.





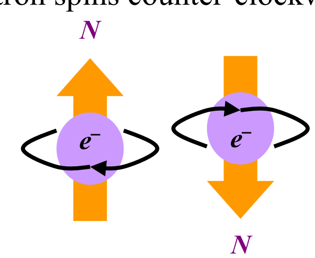
Advantages of using the Bohr Atomic Model

1. It provides a physical and conceptual model that can easily explain the quantized nature of atomic spectra.
2. It is easy for beginning chemistry student to understand the structure of an atom, periodic trends, and chemical bonding (will be discuss in later chapters).

Advantages of using the Atomic Orbital Model (Cloud Probability Model)

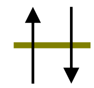
1. It is the most accurate atomic model that combines the wave-duality nature of both energy and matter.
2. In higher level chemistry, it can explain the actual 3 dimensional geometry of molecules and the complex ways in which these molecules can form.

Electron Spin: - when electron spins clockwise, it creates a magnetic north pole in the upward direction. Conversely, when electron spins counter-clockwise, it creates a magnetic north pole in the downward direction.



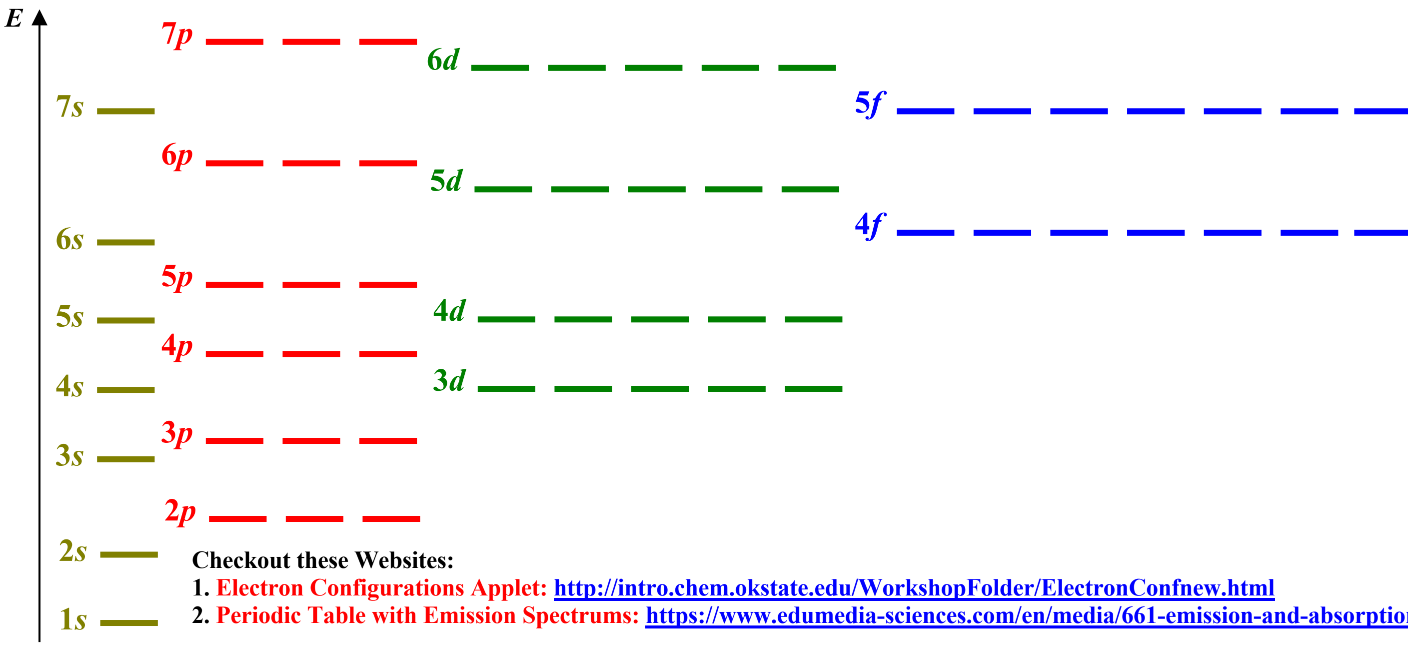
<u>Pauli Exclusion Principle</u>: - in a given atom, each atomic orbital can only have a maximum of two electrons with opposing spins.

Example: In the 1*s* orbital, the two electrons can be represented by



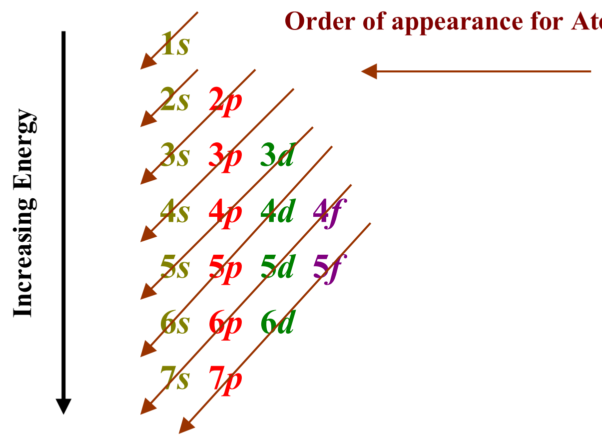
1*s* orbital

<u>Energy Level Diagram</u>: - a diagram that shows the arrangements of electrons in atomic orbitals.



Checkout these Websites:
1. **Electron Configurations Applet: http://intro.chem.okstate.edu/WorkshopFolder/ElectronConfnew.html**
2. **Periodic Table with Emission Spectrums: https://www.edumedia-sciences.com/en/media/661-emission-and-absorption-spectra**

Order of appearance for Atomic Orbitals



<u>Aufbau Principle</u>: - "*Aufbau*" German for "building up"
- for each element, electrons are added up into the atomic orbitals as protons are being added to the nucleus.

<u>Hund's Rule</u>: - for atomic orbitals other than the *s* orbital (as in *p*, *d*, *f* …), the lowest energy can be achieved when the electrons are arranged so that there are a maximum number of unpaired electrons. These unpaired electrons are drawn "spinning up" (↑) in the orbital diagram.

<u>Electron Configuration</u>: - the arrangement of electrons in atomic orbitals.

Example 1: Draw the energy level diagrams and state the electron configurations for the following atoms.

Atomic #	Atom	Electron Configuration	Orbital Diagram 1*s*	2*s*	2*p*
1	H	$1s^1$	↑	—	— — —
2	He	$1s^2$	↑↓	—	— — —
3	Li	$1s^2\ 2s^1$	↑↓	↑	— — —
4	Be	$1s^2\ 2s^2$	↑↓	↑↓	— — —
5	B	$1s^2\ 2s^2\ 2p^1$	↑↓	↑↓	↑ — —
6	C	$1s^2\ 2s^2\ 2p^2$	↑↓	↑↓	↑ ↑ —
7	N	$1s^2\ 2s^2\ 2p^3$	↑↓	↑↓	↑ ↑ ↑
8	O	$1s^2\ 2s^2\ 2p^4$	↑↓	↑↓	↑↓ ↑ ↑
9	F	$1s^2\ 2s^2\ 2p^5$	↑↓	↑↓	↑↓ ↑↓ ↑
10	Ne	$1s^2\ 2s^2\ 2p^6$	↑↓	↑↓	↑↓ ↑↓ ↑↓

<u>Core Electrons</u>: - inner electrons that have completed a row in the Periodic Table of Elements.
- instead of writing the full electron configuration from the very beginning, we can ***<u>abbreviate</u>*** this process by stating the previous noble, then writing out the rest of the element's electron configuration.

Example 2: Draw the quantum orbital diagrams and state the electron configurations for the following atoms.

Atomic #	Atom	Electron Configuration	Orbital Diagram	4*s*	3*d*	4*p*
31	Ga	$[Ar]\ 4s^2\ 3d^{10}\ 4p^1$	[Ar]	↑↓	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑ — —
32	Ge	$[Ar]\ 4s^2\ 3d^{10}\ 4p^2$	[Ar]	↑↓	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑ ↑ —
33	As	$[Ar]\ 4s^2\ 3d^{10}\ 4p^3$	[Ar]	↑↓	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑ ↑ ↑
34	Se	$[Ar]\ 4s^2\ 3d^{10}\ 4p^4$	[Ar]	↑↓	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑↓ ↑ ↑
35	Br	$[Ar]\ 4s^2\ 3d^{10}\ 4p^5$	[Ar]	↑↓	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑↓ ↑↓ ↑
36	Kr	$[Ar]\ 4s^2\ 3d^{10}\ 4p^6$	[Ar]	↑↓	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑↓ ↑↓ ↑↓

Example 3: Draw the quantum orbital diagram and state the electron configurations of Antimony (Sb).

Sb (Atomic # = 51 = 51e^-). Sb = $1s^2\ 2s^2 2p^6\ 3s^2 3p^6\ 4s^2 3d^{10} 4p^6\ 5s^2 4d^{10}\ 5p^3$ **or** $[Kr]\ 5s^2\ 4d^{10}\ 5p^3$

[Kr] 5*s* ↑↓   4*d* ↑↓ ↑↓ ↑↓ ↑↓ ↑↓   5*p* ↑ ↑ ↑

Example 4: Identify the element with the following ground state electron configurations.

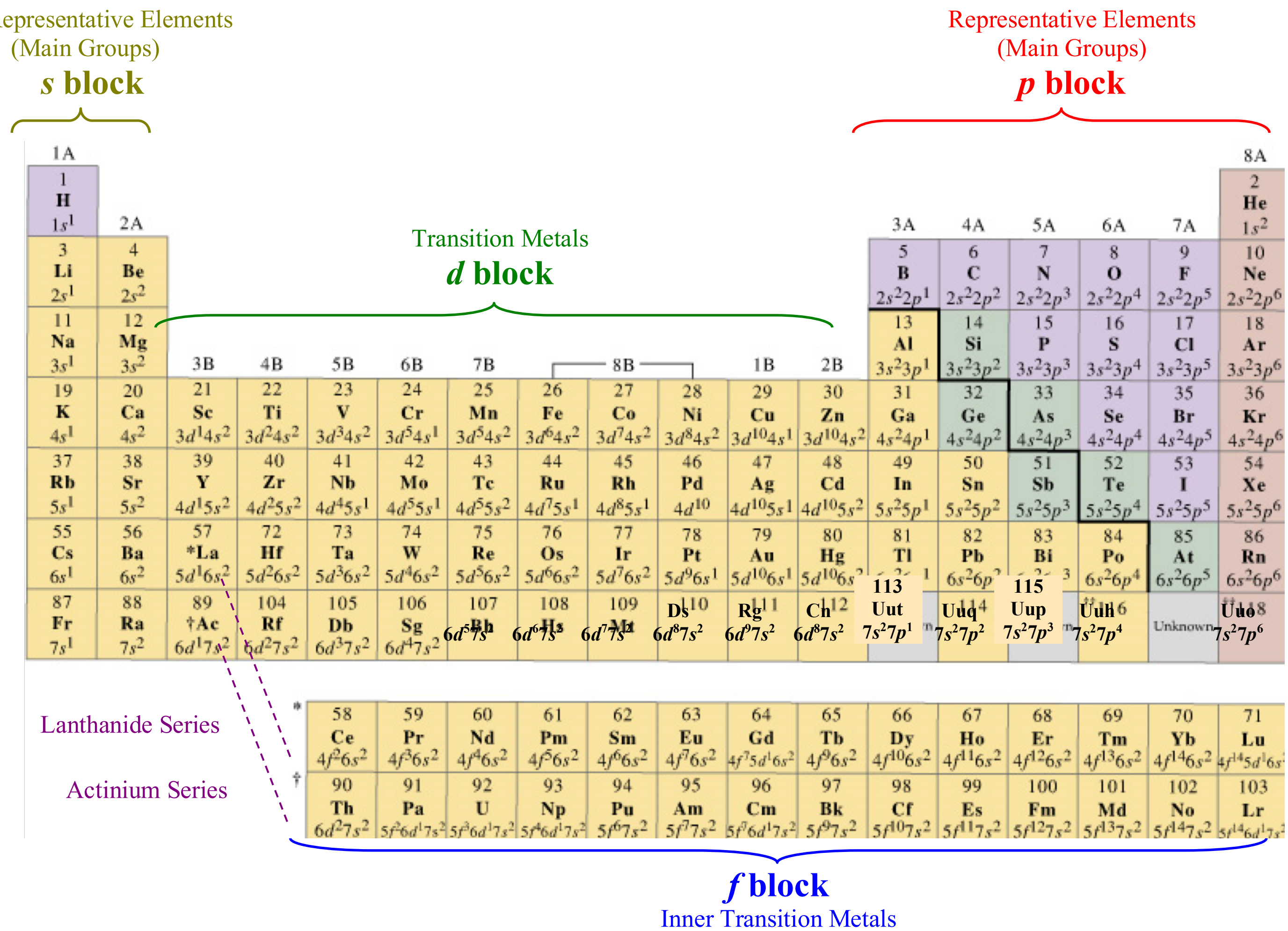
a. $[Kr]\ 5s^2\ 4d^{10}\ 5p^5$

There are 17 electrons (2 + 10 + 5) after the element Kr (which has 36 electrons). **Hence, there are a <u>total 53 electrons</u> in this element. The element would be <u>Iodine (I)</u> because its atomic number is 53.**

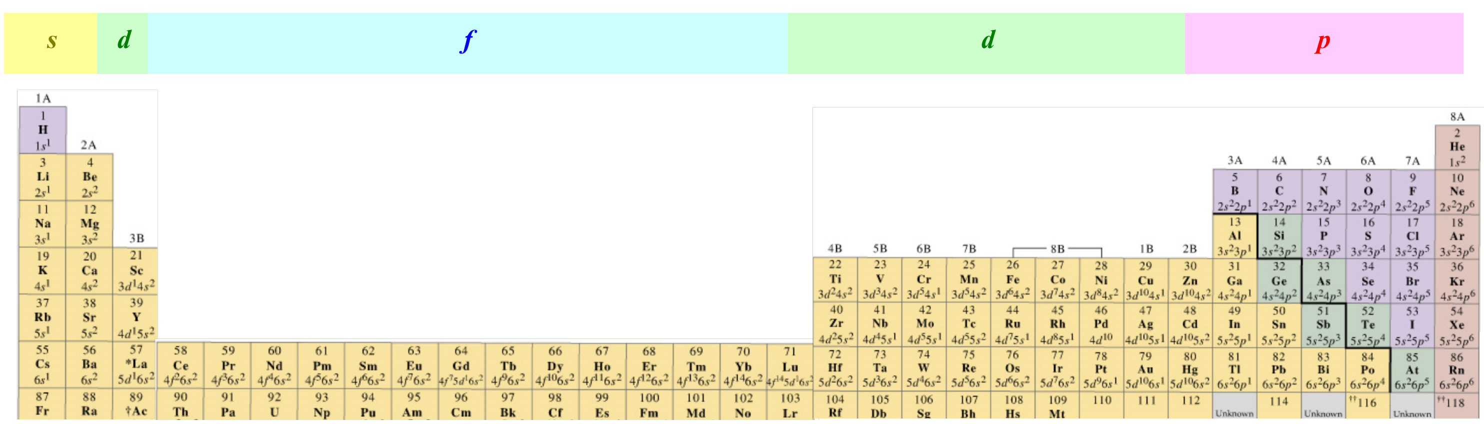
b. $[Xe]\ 6s^2\ 4f^{14} 5d^{10}$

There are 26 electrons (2 + 14 + 10) after the element Xe (which has 54 electrons). **Hence, there are a <u>total 80 electrons</u> in this element. The element would be <u>Mercury (Hg)</u> because its atomic number is 80.**

Electron Configurations in the Periodic Table of Elements



Extended Periodic Table of Elements



Example 5: Identify the element with the following excited state electron configurations.

a. [Ne] $3s^1\ 3p^5$

Excited state element has the same number of electrons as its ground state.

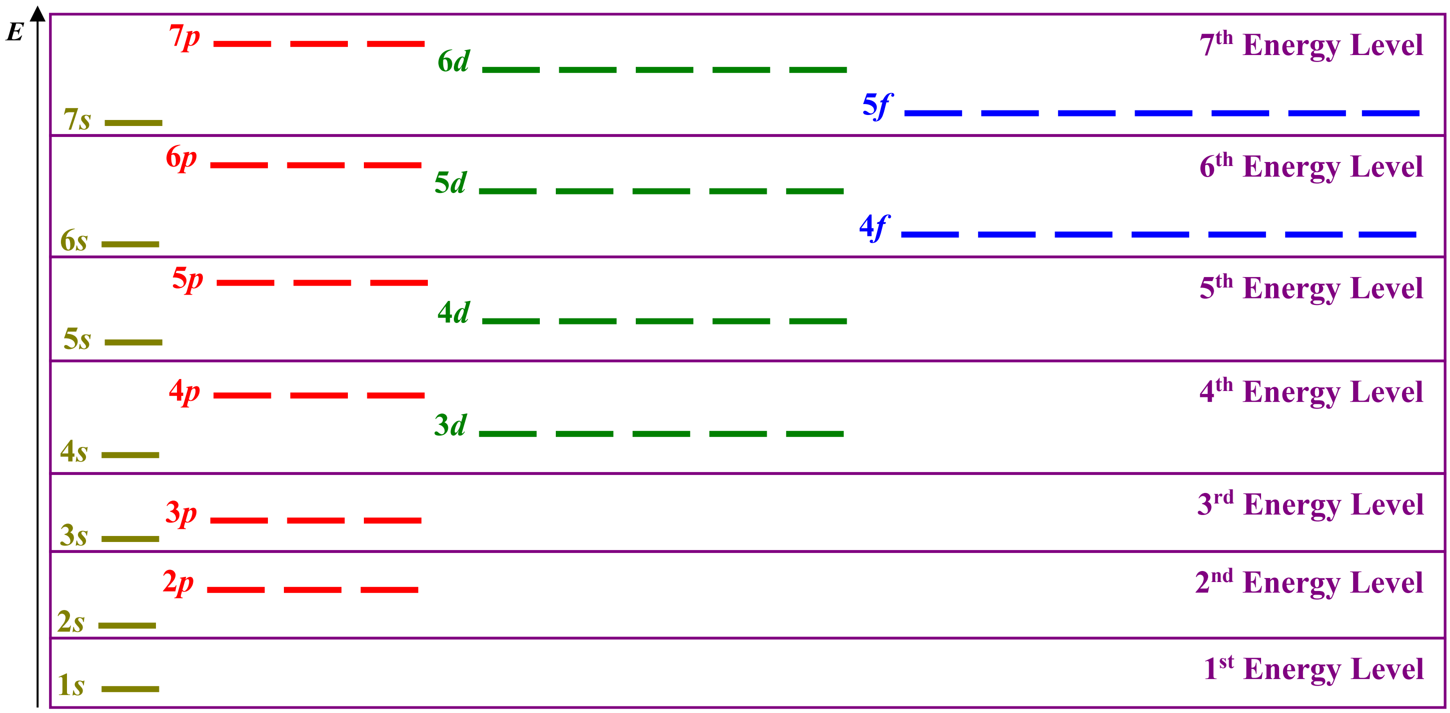
There are 6 electrons (1 + 5) after the element Ne (which has 10 electrons). **Hence, there are a <u>total 16 electrons</u> in this element. The element would be <u>Sulfur (S)</u> because its atomic number is 16.**

b. [Ar] $4s^2\ 3d^7\ 4p^3$

There are 12 electrons (2 + 7 + 3) after the element Ar (which has 18 electrons). **Hence, there are a <u>total 30 electrons</u>. The element would be <u>Zinc (Zn)</u> because its atomic number is 30.**

Electron Shells and Valence Electrons:

Recall from the Bohr Model studied earlier. It states that **electrons travel around the nucleus in specified orbits (electrons are quantum)**. These orbits are called **energy shells**. They can also be called **electron shells**. These orbits are very similar to the planets orbiting our sun. The only difference is that each orbit can accommodate more than one electron at a time. This can be explained as the simplified energy level diagram based on the probability cloud model. The following diagram and table shows the maximum number of electrons each successive "shell" allows.



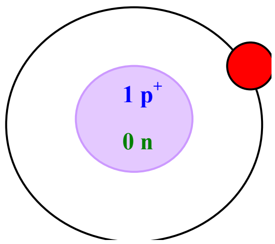
Energy Level	Maximum Number of Electrons Allowed
1st	2
2nd	8
3rd	8
4th	18
5th	18
6th	32
7th	32

To put electrons in the shells, we have to fill the first energy level until it is full before we can start filling the next energy level. If the second energy level is filled, then we can put electrons in the third energy level and so on. This process is repeated until all the electrons are used up. One way to remember the maximum number of electrons for each energy level is to look at the Periodic Table of Elements. There are 2 elements in the first row, hence 2 electrons are allowed in the first energy level. There are 8 elements each in the second and third rows, hence 8 electrons are allowed in each of the second and third energy level. This pattern repeats itself for higher energy levels.

Valence Electrons: - the electrons in the **outermost** shell.

Shell Model Diagram: - a diagram that emphasize the number of electrons in their respective shells along with the number of protons and neutrons of an atom.

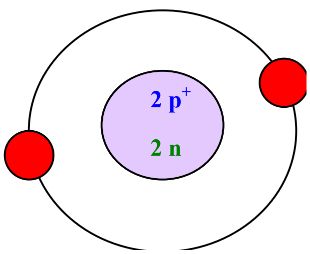
Example 6: Draw the shell model diagram for the first 10 elements.



Element: Hydrogen (H)

Atomic Number: 1 **Atomic Mass: 1.01** **Nucleus: 1 p^+ and 0 n**

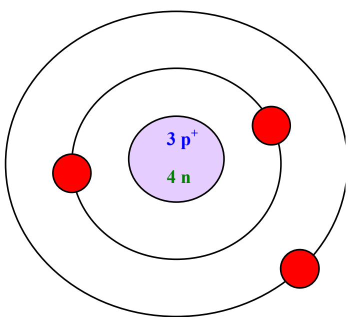
1st Energy Level: 1 e^- **(1 valence e^-)**
Total: 1 e^-



Element: Helium (He)

Atomic Number: 2 **Atomic Mass: 4.00** **Nucleus: 2 p^+ and 2 n**

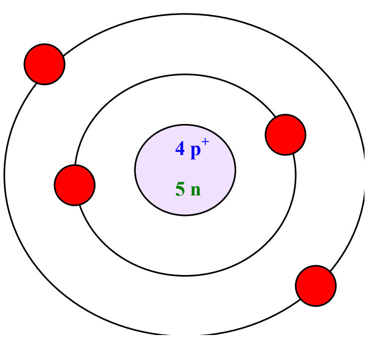
1st Energy Level: 2 e^- **(2 valence e^- - Filled)**
Total: 2 e^-



Element: Lithium (Li)

Atomic Number: 3 **Atomic Mass: 6.94** **Nucleus: 3 p^+ and 4 n**

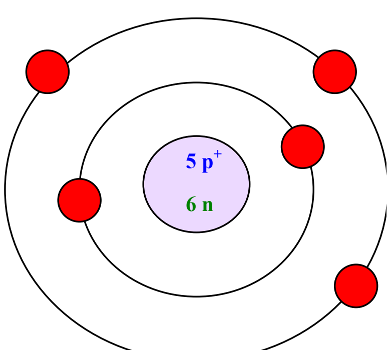
2nd Energy Level: 1 e^- **(1 valence e^-)**
1st Energy Level: 2 e^-
Total: 3 e^-



Element: Beryllium (Be)

Atomic Number: 4 **Atomic Mass: 9.01** **Nucleus: 4 p^+ and 5 n**

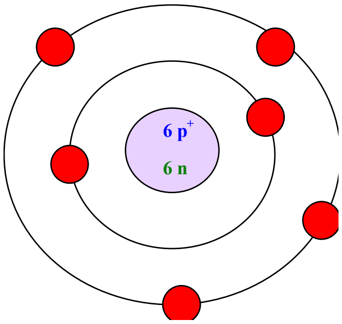
2nd Energy Level: 2 e^- **(2 valence e^-)**
1st Energy Level: 2 e^-
Total: 4 e^-



Element: Boron (B)

Atomic Number: 5 **Atomic Mass: 10.81** **Nucleus: 5 p^+ and 6 n**

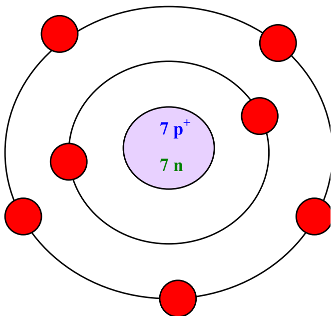
2nd Energy Level: 3 e^- **(3 valence e^-)**
1st Energy Level: 2 e^-
Total: 5 e^-



Element: Carbon (C)

Atomic Number: 6 **Atomic Mass: 12.01** **Nucleus: 6 p^+ and 6 n**

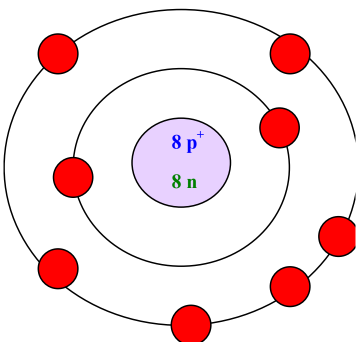
2^{nd} Energy Level: 4 e^- **(4 valence e^-)**
1^{st} Energy Level: 2 e^-
Total: 6 e^-



Element: Nitrogen (N)

Atomic Number: 7 **Atomic Mass: 14.01** **Nucleus: 7 p^+ and 7 n**

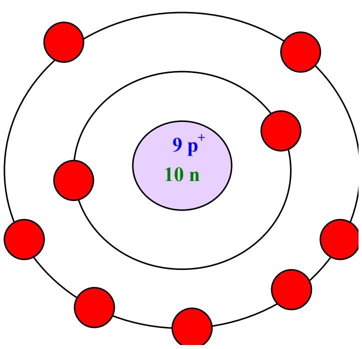
2^{nd} Energy Level: 5 e^- **(5 valence e^-)**
1^{st} Energy Level: 2 e^-
Total: 7 e^-



Element: Oxygen (O)

Atomic Number: 8 **Atomic Mass: 16.00** **Nucleus: 8 p^+ and 8 n**

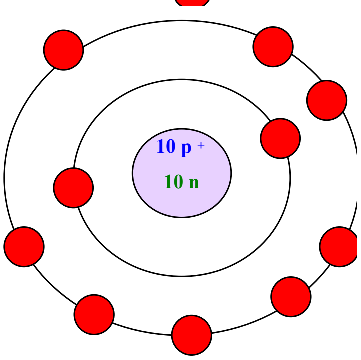
2^{nd} Energy Level: 6 e^- **(6 valence e^-)**
1^{st} Energy Level: 2 e^-
Total: 8 e^-



Element: Fluorine (F)

Atomic Number: 9 **Atomic Mass: 19.00** **Nucleus: 9 p^+ and 10 n**

2^{nd} Energy Level: 7 e^- **(7 valence e^-)**
1^{st} Energy Level: 2 e^-
Total: 9 e^-



Element: Neon (Ne)

Atomic Number: 10 **Atomic Mass: 20.18** **Nucleus: 10 p^+ and 10 n**

2^{nd} Energy Level: 8 e^- **(8 valence e^- - Filled)**
1^{st} Energy Level: 2 e^-
Total: 10 e^-

Assignment

3.3 pg. 99 # 1 & 2 (Practice); pg. 99 #1 to 11

3.4: Counting Atoms

Atomic Mass: - sometimes called **atomic weight**.
- the mass of the atom in atomic mass unit (amu).
- 1 amu = exactly one-twelfth the mass of one carbon-12 atom ≈ 1.67×10^{-27} kg.

Average Atomic Mass: - Average Mass of an atom and its isotopes after accounting their proportions of abundance (as stated on the Periodic Table of Elements).

Relative Abundance: - the relative proportion of various isotopes of an element.

Example 1: State the Average Atomic Mass Unit for hydrogen if it is made of 99.48% of ${}^{1}_{1}\text{H}$, 0.2400% of ${}^{2}_{1}\text{H}$, and 0.2800% of ${}^{3}_{1}\text{H}$.

$$\text{Average amu of Hydrogen} = \underbrace{(0.9948)(1)}_{\text{99.48\% of Atomic Mass 1}} + \underbrace{(0.002400)(2)}_{\text{0.2400\% of Atomic Mass 2}} + \underbrace{(0.002800)(3)}_{\text{0.2800\% of Atomic Mass 3}}$$

Average amu of Hydrogen = 1.008

Mole (mol): - a group of atoms or molecules numbered **6.022×10^{23} (*Avogadro's Number*, N_A)**

Examples: 1 mol of carbon (C) = 6.022×10^{23} carbon atoms = 12.01 g (same as the amu)
1 mol of oxygen (O_2) = 6.022×10^{23} oxygen molecules = 32.00 g (include subscripts with amu)

Example 2: Calculate the mass of 250 atoms of gold.

1 mol Au = 196.97 g Au = 6.022×10^{23} Au atoms

$$250\ \cancel{\text{atoms}} \times \frac{196.97\text{ g}}{6.022 \times 10^{23}\ \cancel{\text{atoms}}}$$

8.177×10^{-20} g of Au

Example 3: Determine the number of molecules for 50.0 mg of oxygen. How many atoms of oxygen are there?

1 mol O_2 = 32.00 g O_2 = 6.022×10^{23} O_2 molecules
(16.00 × 2)

50.0 mg = 0.0500 g

$$0.0500\ \cancel{\text{g}} \times \frac{6.022 \times 10^{23}\text{ molecules}}{32.00\ \cancel{\text{g}}}$$

9.41×10^{20} molecules of O_2

There are two oxygen atoms in one molecular unit of O_2.

$$9.41 \times 10^{20}\text{ molecules of } O_2 \times \frac{2\text{ atoms of oxygen}}{1\text{ molecule of } O_2} = \mathbf{1.88 \times 10^{21}\text{ oxygen atoms}}$$

Molar Mass (g/mol): - sometimes refer to as **molecular mass** or **molecular weight**, is the mass per one mole of atoms or molecules.
- molar mass of a mono-atomic element is the same as the atomic mass.
- molar mass of a compound, diatomic element, or polyatomic element is the same as the combine atomic masses of all atoms in the molecule.

Example 4: Find the molar mass of the following.

a. aluminium

Al = 26.98 g/mol

b. nitrogen

$N_2 = 2 \times 14.01$ g/mol

$\mathbf{N_2 = 28.02}$ **g/mol**

c. water (H_2O)

$H_2O = (2 \times 1.01) + 16.00$

$\mathbf{H_2O = 18.02}$ **g/mol**

d. lead (IV) nitrate – $Pb(NO_3)_4$

$Pb(NO_3)_4 = 207.2 + (4 \times 14.01) + (12 \times 16.00)$

$\mathbf{Pb(NO_3)_4 = 455.2}$ **g/mol**

e. sucrose – $C_{12}H_{22}O_{11}$

$C_{12}H_{22}O_{11} = (12 \times 12.01) + (22 \times 1.01) + (11 \times 16.00)$

$\mathbf{C_{12}H_{22}O_{11} = 342.30}$ **g/mol**

Converting between Mass and Moles:

$$\text{Moles (mol)} = \frac{\text{Mass (g)}}{\text{Molar Mass (g/mol)}} \qquad n = \frac{m}{M}$$

***n* = moles** ***m* = mass** ***M* = Molar mass**

Example 5: Calculate the number of moles for

a. 20.0 g of magnesium chloride – $MgCl_2$

$MgCl_2 = 24.31 + 2(35.45)$ $M = 95.21$ g/mol

$$n = \frac{m}{M} = \frac{20.0\text{ g}}{95.21\text{ g/mol}}$$

$\mathbf{n = 0.210}$ **mol**

b. 4.52 mg of glucose – $C_6H_{12}O_6$

$C_6H_{12}O_6 = 6(12.01) + 12(1.01) + 6(16.00)$

$M = 180.18$ g/mol

$$n = \frac{m}{M} = \frac{4.52\text{ mg}}{180.18\text{ g/mol}}$$

$\mathbf{n = 0.0251}$ **mmol**

Example 6: Determine the mass of the following amount.

a. 8.52 mol of ozone – O_3

$O_3 = 3(16.00)$ $M = 48.00$ g/mol

$$n = \frac{m}{M}$$

$m = nM = (8.52 \cancel{\text{mol}})(48.00\text{ g}/\cancel{\text{mol}})$

$\mathbf{m = 409}$ **g**

b. 24.7 mmol of phosphoric acid – H_3PO_4

$H_3PO_4 = 3(1.01) + 30.97 + 4(16.00)$ $M = 97.99$ g/mol

$$n = \frac{m}{M}$$

$m = nM = (24.7 \text{ m}\cancel{\text{mol}})(97.99\text{ g}/\cancel{\text{mol}}) = 2420.353$ mg

$\mathbf{m = 2.42 \times 10^3}$ **mg = 2.42 g**

Assignment

3.4 pg. 102 # 1 to 4 (Practice); pg. 103 # 1 to 3 (Practice); pg. 104 #1 to 13

~~**Chapter 3 Review pg. 107–108 #12, 13, 16, 17, 18, 20 to 26, 28 to 34, 36, 38 to 46**~~

Chapter 4: The Periodic Table

4.1 & 4.2: Development and Tour of the Periodic Table

Johann Dobereiner: - first to discover a pattern of a group of elements like Cl, Br, and I (called triads).

John Newland: - suggested elements should be arranged in "octaves" because they repeat their properties for every eighth elements.

Demitri Mendeleev: - conceived the first modern periodic table of elements.
- insisted certain spots of the table be left blank until the actual element is found that matched the predicted properties. This was done to preserve the elements with similar properties called groups or families.

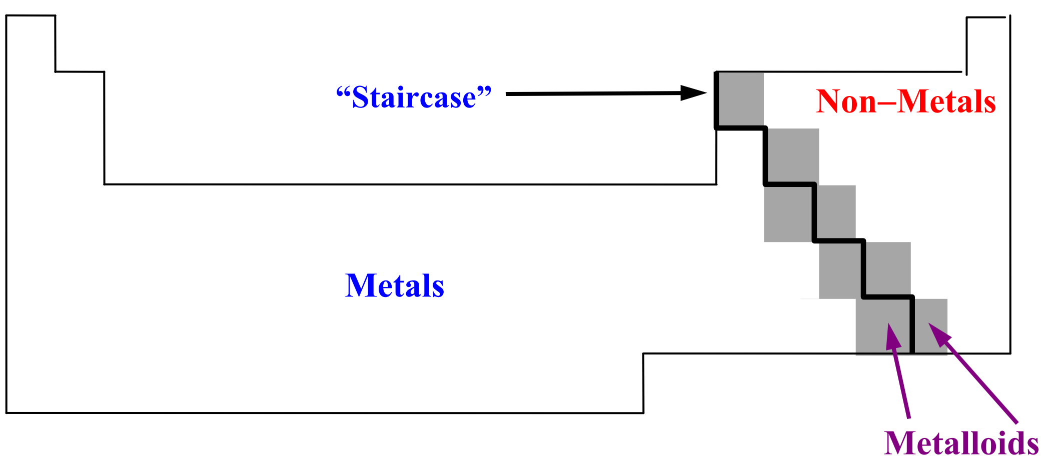
Periodic Law: - when elements arranged according to their atomic numbers, elements with similar properties occur at regular interval. (Chemical and physical properties of elements change periodically with their atomic numbers.)

Valence Electrons: - the outer electrons of an atom that are involved in chemical bonding.
- for representative elements, it takes 8 valence electrons (ns^2np^6) to achieve stability.
- in most cases, this means having the same electron arrangement of the **nearest** noble gas, except helium (only 2 electrons to fill the first energy level), and the transition metals.

Octet Rule: - the tendency for electrons to fill the second and third energy levels (8 valence electrons – for main groups – IA to VIIIA columns) to achieve stability.
- in most cases, this means having the same electron arrangement of the **nearest** noble gas.
- exceptions to the rule include helium (only 2 electrons to fill the first energy level), and the transition metals.

Metals and Non-Metals

The 2 main categories of the Periodic Table of Elements are the **metals** and **non-metals**. They are divided by the "staircase" on the table. This "staircase" can be found at the element Boron extending down to the element Astatine. **Metals are the elements at the left side of the "staircase"**, and **non-metals are the elements at the right side of the "staircase"**.



Note: **Hydrogen is placed at the metal side but is considered a non-metal.**

Physical Properties: - are the properties or characteristics of a substance that can be change without involving the chemical change in its composition.

Physical Properties of Metals (with the exception of hydrogen):

1. Metals are mostly solids at room temperature (with the exception of mercury).
2. Metals are malleable (they can be hammered into thin sheets).
3. Metals are ductile (they can be pulled into wires).
4. Metals are good conductors of heat and electricity.
5. Metals are lustrous (shiny).

Physical Properties of Non-Metals:

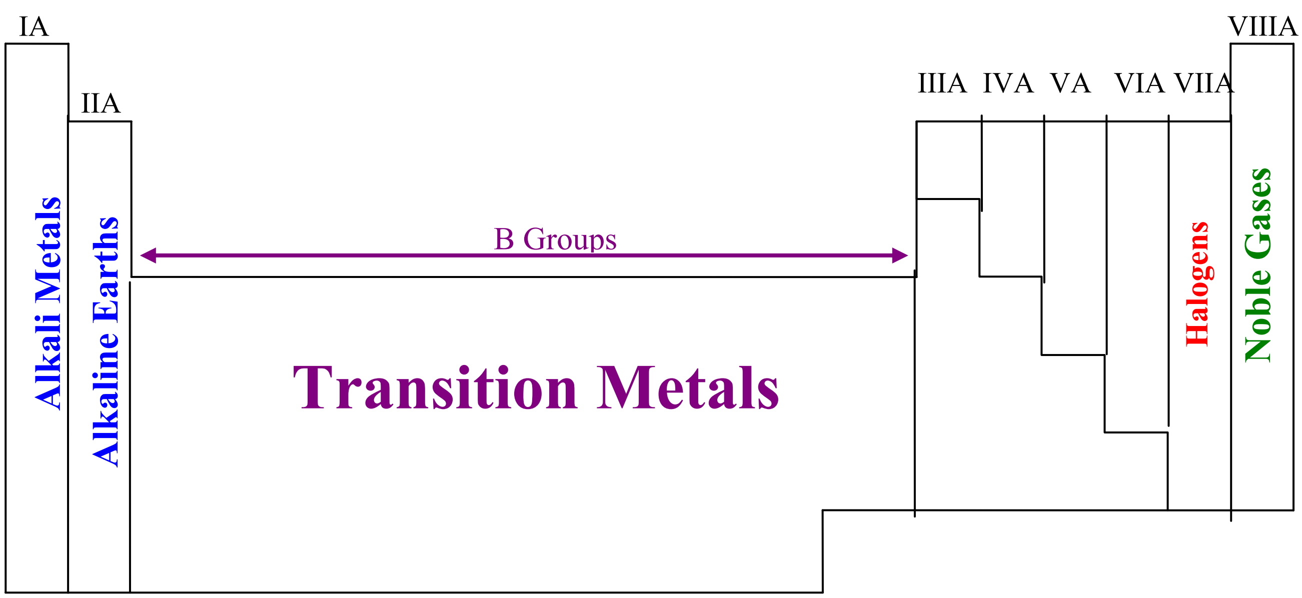
1. Non-metals are mostly gases or solids at room temperature (with the exception of bromide).
2. Non-metals usually do not have the physical properties of metals.

Some **elements near the "staircase" exhibit both the properties of metals and non-metals**. These elements are referred to as **metalloids**. An example is silicon. It is often used as a semiconductor material (an electrical conductor that can conduct and an insulate electricity). Other metalloids are boron, germanium, arsenic, antimony, tellurium, polonium, and astatine.

Periods and Groups: Chemical Properties of Elements

Chemical Properties: - the properties of a substance that involves a change in the organisation of atoms (mainly the sharing or transfer of electrons).

The shape of the Periodic Table of Elements is a structural way to organize elements. The vertical columns of the Table are called **groups or families**. As we have seen before, the column number is the same as the number of valence electrons of the elements. Since chemical properties depend greatly on the number of valence electrons, all elements within the same group or family must have similar chemical properties. We have already seen one such family, the noble gases. All elements of this group are non-reactive and very stable (recall the valence electron shell of these elements is full). The names of other families and their general chemical properties are listed below.



Alkali Metals (Group IA or Group 1): - derived from the Arabic word "*al-qali*" (meaning ashes), commonly use as a slippery solution when mixed with water (alkaline solutions) to remove grease. They are very reactive when initially mixed with water. All have one valence electron.

Alkaline Earth Metals (Group IIA or Group 2): - similar to alkali metal, they form alkaline (basic) solutions when mixed with water. However, they do not melt or change when put in a fire. Hence, we use the word "earth. All have two valence electrons.

Halogens (Group VIIA or Group 17): - derived from the Greek meaning of "salt-forming", as this group forms different type of salts. It is the most reactive type of non-metals. All of them require one more valence electron to complete the energy level.

Noble Gases (Group VIIIA or Group 18): - non-reactive gases of the last column of the table.

Groups or Families	Chemical Properties
Alkali Metals (IA)	very reactive metals
Alkaline Earth Metals (IIA)	less reactive than alkali metals
Halogens (VIIA)	very reactive non-metals
Noble Gases (VIIIA)	very stable; all are gaseous state at room temperature

Periods: - "rows" of elements that are identify by their highest energy level.
- the pattern of chemical properties "repeats" for every row.

Representative Elements: - also called ***Main Group Elements***.
- elements in Group IA to VIIIA (Groups 1, 2 and 13 to 18) that consists of partial or filled *s* and *p* orbitals of the same principal quantum number, *n*.
- other than helium, $1s^2$, all noble gases have ns^2np^6 as their electron configurations.

Transition Metals: - elements in Group 1B and 3B through 8B (Groups 3 to 11).
- all transition metals consists of *d* orbital electrons in their outer electron configurations. The group designation acknowledges the number of outer electrons. For examples, Mn has 7 valence electrons ($4s^23d^5$) and it is in Group 7B.
- Group 2B (Zn, Cd and Hg) are neither transition metals nor representative elements due to the fact they have filled $ns^2\ (n-1)d^{10}$ as their electron configurations.

Lanthanides and Actinides: - are elements with incompletely filled *f* subshells.
- are sometimes called **Inner Transition Elements**.

Alloys: - physical mixture of elements (mostly metals) in different composition. The resulting metal mixture either have enhance properties or physical and chemical characteristics of component elements.

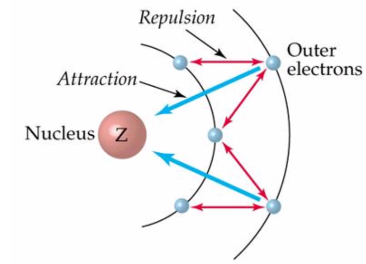
Examples: Brass (37% Zinc and 63% Copper); Bronze (10% Zinc and 90% Copper);
Stainless Steel (97% Iron and 3% Carbon)

Assignment
4.1 pg. 122 #1 to 14
4.2 pg. 131 #1 to 13

4.3: Trends in the Periodic Table

There are many different trends regarding the physical and chemical properties of the elements in the Periodic Table. However, we will limit to three atomic properties. They are atomic size, ionic radii, ionization energy, and electron affinity.

Electron Shielding Effect: - the outer electrons are pushed away because of the repulsion between them and the core electrons. The net result is that the protons in the nucleus cannot hold on to these outer electrons as tightly as they would for the core electrons.



Effective Nuclear Charge (Z_{eff}): - the net nuclear charge actually experienced by an electron (the difference between the number of protons, Z, and the number of "shielded" core electrons).
- **the higher it is for Z_{eff}, the less shielding effect the outer electrons will experience**.

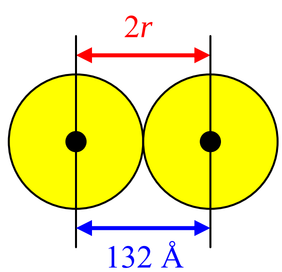
$$Z_{eff} = Z - \text{“Shield” Core Electrons}$$

Example 1: Calculate the effective nuclear charge of Na and Ar (first and last elements of period 3).

Sodium (Na): $Z = 11$ protons "Shield" Core $e^- = 10$ (e^- in the first two shells)
$Z_{eff} = 11 - 10$ **Z_{eff} = 1 for Na**

Argon (Ar): $Z = 18$ protons "Shield" Core $e^- = 10$ (e^- in the first two shells)
$Z_{eff} = 18 - 10$ **Z_{eff} = 8 for Ar**
(Ar has experience LESS shielding effect than Na)

Atomic Radius: - the size of an atom as measured by the distances between atoms in chemical compound.
- sometimes refer to as **bond radius** (half the distance between the chemical bond).



Cl_2 Molecule
$2r = 132$ Å

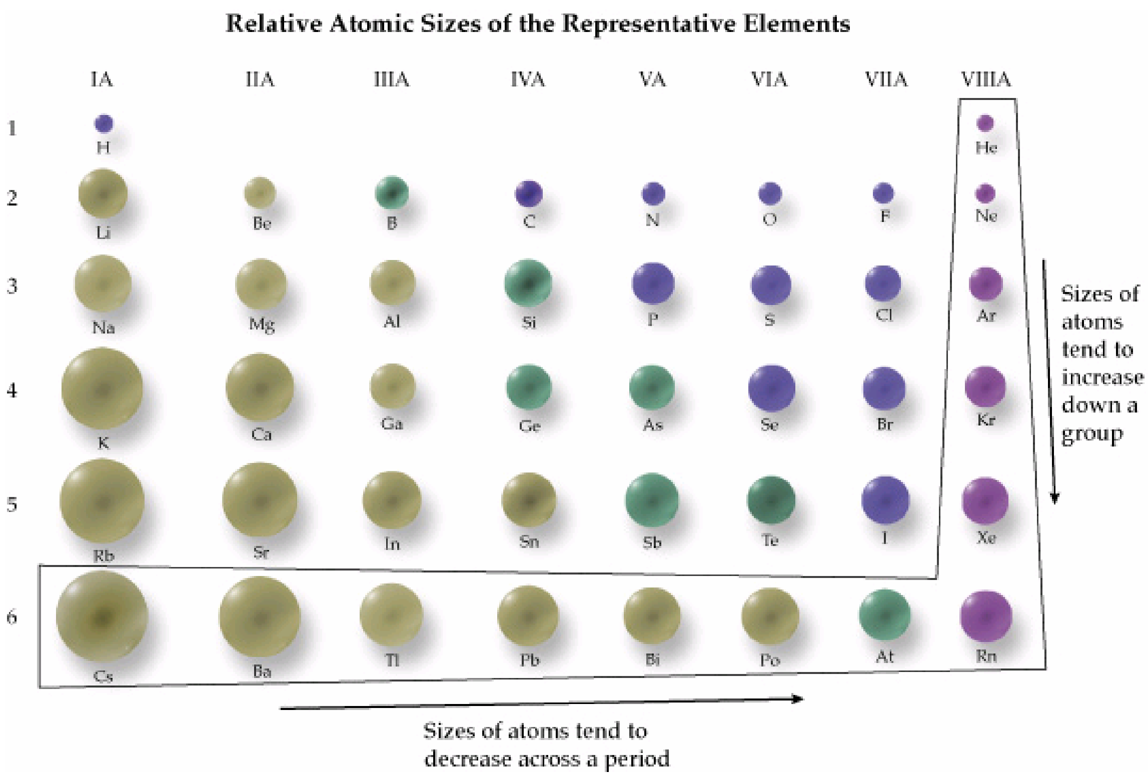
Cl Atom
$r = 66$ Å

(1 Å = 0.1 nm = 1×10^{-10} m)

Several Notes on Trends in Atomic Radii

1. In general, **Atomic Radii decrease as one move to the right of a period**. This is because the increases in protons in the same row increase the effective nuclear charge (the protons in the nucleus have more pull on the outer electrons, decreasing shielding), thus drawing these outer electrons closer to the nucleus, decreasing in sizes as the result.

2. **Atomic Radii INCRESES Down a Group**. This is due to the fact there are more orbitals as the number of row increases. The outer electrons are, of course, further away from the nucleus



Example 2: Order the following atoms from the smallest to the largest.

a. Te, S, Se, O

These atoms are within the same Group (column). As we move down the column, atomic size increases. Therefore,

O < S < Se < Te

b. Na, S, Mg, Cl, P

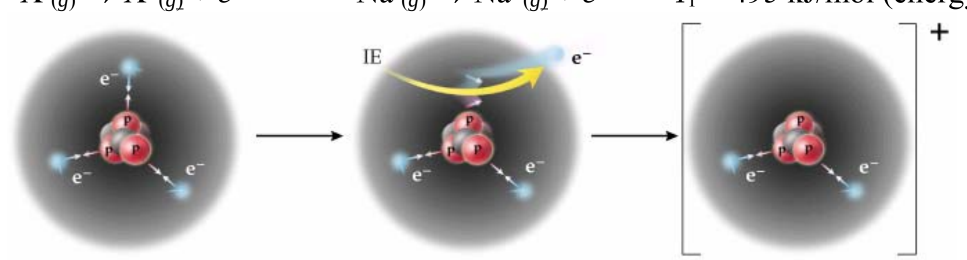
These atoms are within the same Period (row). As we move to the right, atomic size decreases. Therefore,

Cl < S < P < Mg < Na

Ionization Energy: - the energy needed to completely remove an electron from a gaseous atom or gaseous ion (plasma).

First Ionization Energy: - the ionization energy required to remove the highest-energy electron from an atom.

$X_{(g)} \rightarrow X^+_{(g)} + e^-$ $Na_{(g)} \rightarrow Na^+_{(g)} + e^-$ I_1 = 495 kJ/mol (energy is required)



Negative electrons cannot leave the atom unless energy is supplied to overcome their attraction to the nucleus.

Ionization energy (IE) kicks an electron out of the atom.

The resulting surplus of one proton gives the atom an overall +1 charge. The atom has been ionized and is now a cation.

Second Ionization Energy: - the ionization energy required to remove the second highest-energy electron from the ion.

$$X^{+}_{(g)} \rightarrow X^{2+}_{(g)} + e^{-} \qquad Na^{+}_{(g)} \rightarrow Na^{2+}_{(g)} + e^{-} \qquad I_2 = 4560 \text{ kJ/mol}$$

Successive Ionization Energies (kJ/mol) for Elements in Row 3 of the Periodic Table

First Ionization Energies generally INCREASE within a Period.

Elements	I_1	I_2	I_3	I_4	I_5	I_6	I_7
Na	495	4560					
Mg	735	1445	7730				
Al	580	1815	2740	11600			
Si	780	1575	3220	4350	16100		
P	1060	1890	2905	4950	6270	21200	
S	1005	2260	3375	4565	6950	8490	27000
Cl	1255	2295	3850	5160	6560	9360	11000
Ar	1527	2665	3945	5770	7230	8780	12000

Core Electrons

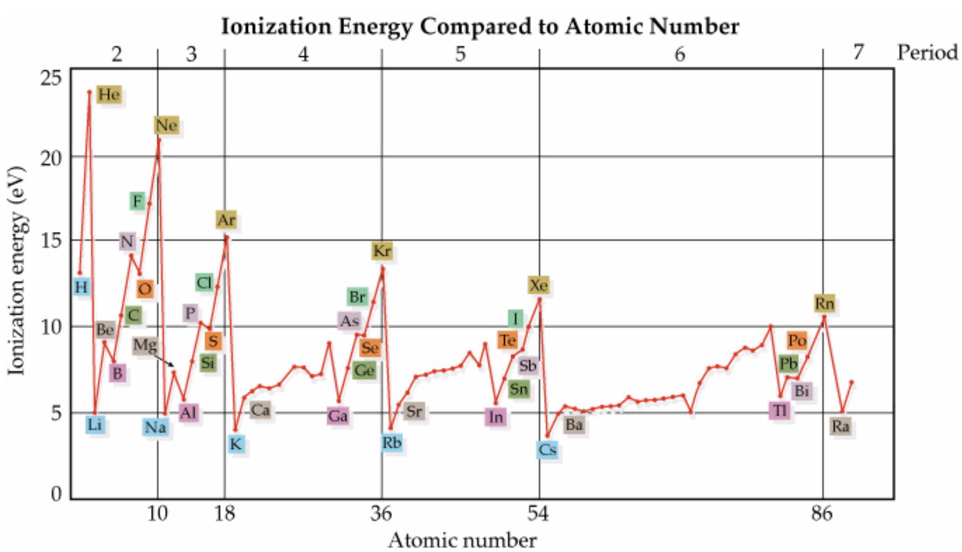
Successive Ionization Energies INCREASES within each element. ($I_1 < I_2 < I_3 < \ldots$)

Several Notes on Trends in Ionization Energies

1. There is an **Increase in Successive Ionization Energies** because each successive electron has to jump from a lower level. Besides, these successive electrons are bind more tightly with the nucleus because they are closer to the protons.

2. **Ionization Energies Decrease Down a Group**. This is due to the fact as the atom has more orbitals, it is increasing in size. It is easier (takes less energy) to take away a valence electron because the protons are having a more difficult time to "hold on" to the electron.

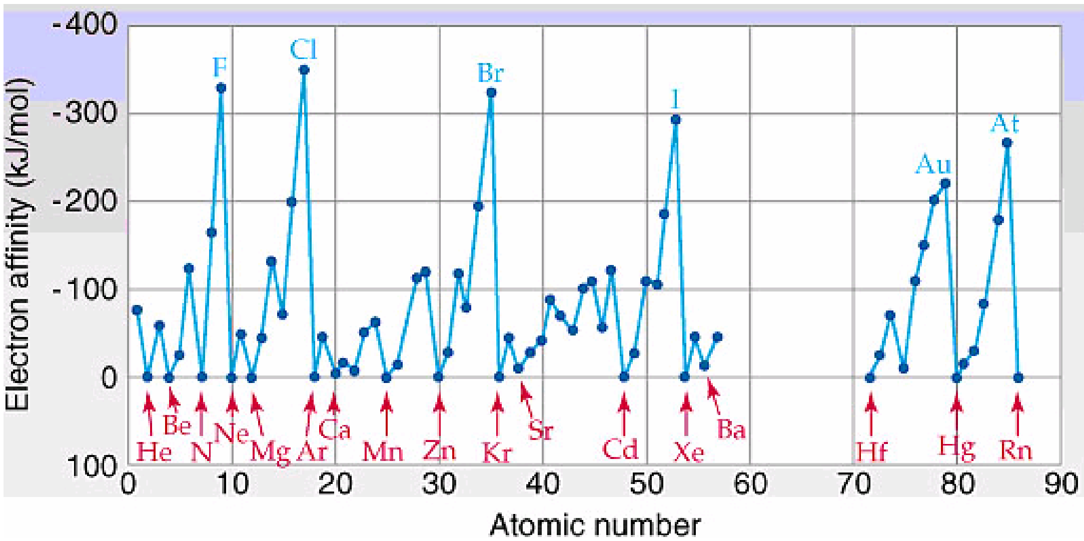
3. In general, **Ionization Energies Increase as one move from Left to the Right of a Period**. This is because the increases in protons in the same row increase the effective nuclear charge (the protons in the nucleus have more pull on the outer electrons, decreasing shielding), thus requiring more energy to ionize them.



Electron Affinity: - the change in energy associates with an addition of an electron to a gaseous atom.
- the larger negative electron affinity, the more stable the anion formed.

$$X_{(g)} + e^- \rightarrow X^-_{(g)}$$

$$Cl_{(g)} + e^- \rightarrow Cl^-_{(g)} \qquad \Delta EA = -348.7 \text{ kJ/mol} \qquad (348.7 \text{ kJ/mol is released})$$

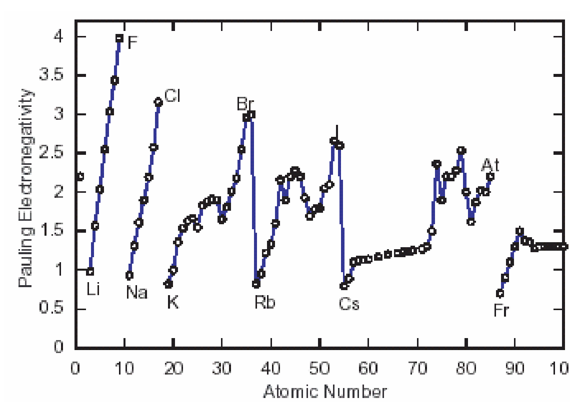


Several Notes on Trends in Electron Affinities

1. In general, **Electron Affinity DECRESES Down a Group (less energy is released)**. This is due to the fact there are more orbitals as the number of row increases. Since the protons in the nucleus cannot attract another electron as effectively due to the increase distance involved, less energy is released. **The trend holds except for row 2. The reason is because of the small size of the 2*p* orbitals. Electron repulsion cause smaller values of electron affinities than expected for row 2.**

2. In most cases, **Electron affinity INCRESASES (becomes more negative) across the Period from Left to Right up to the Halogen group (more energy is released)**. Metals have $EA > 0$ because they like to form cations (low ionization energy). Non-metals have $EA < 0$ because they like to form anions in order to form a stable octet.

Electronegativity: - first determined by **Linus Pauling**, it is a measure of the **capability of an atom within a molecule to attract shared electrons around itself**.
- the better the atom is able to attract electrons, the higher the electronegativity value.
- electronegativity of noble gases is 0 as their outer orbitals are filled and do not attract electrons.



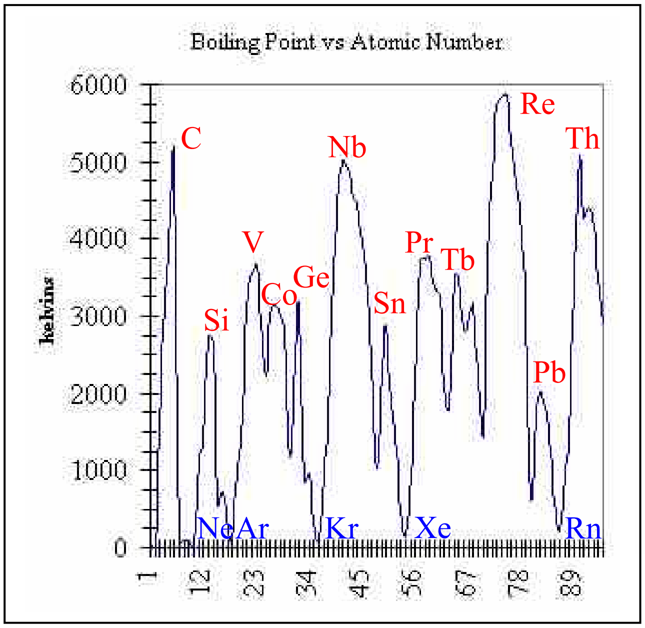
Several Notes on Trends in Electronegativity

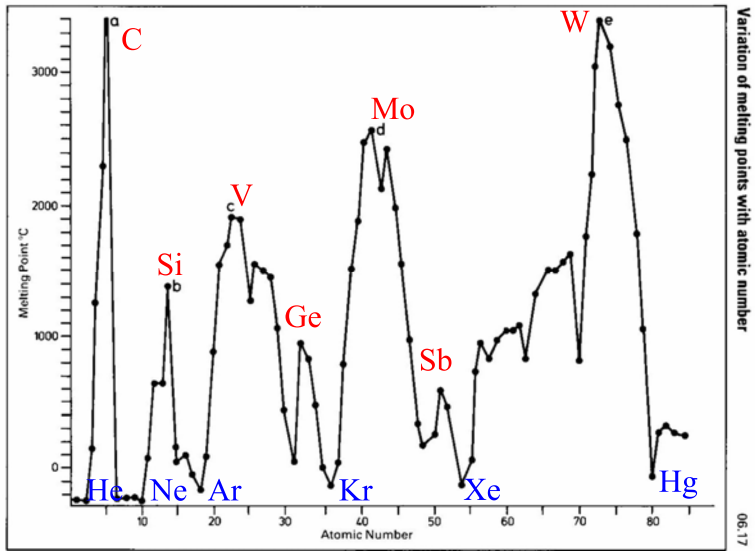
1. In general, **Electronegativities INCREASE as one move to the right of a period (up to and including halogens)**. This is because of the increase in electron affinity of the non-metals. These non-metals like to form anions to fill the valence orbitals. Metals tend to the high ionization energy because they like to give away electrons to form cations.

2. **Electronegativities DECRESE Down a Group**. This is due to the fact there are more orbitals as the number of row increases. The outer electrons are, of course, further away from the nucleus. Hence, it is more difficult for the protons of the nucleus to attract electrons into the valence orbitals.

Boiling and Melting Points: - in elements, the boiling and melting points corresponds to the strength of connection atoms have with each other.

- ***the more unpaired electrons an atom has, the higher the boiling and melting points***. This is because the unpaired electrons of an atom can form bonds with adjacent atoms (two unpaired electrons from different atoms can result in a bond).
- on the other hand, the less unpaired electrons there are, the lower the boiling and melting points for the element. This is due to the lack of unpaired electrons to form bonds with adjacent atoms.
- Hence, we find ***elements in the middle of the p-block and d-block have the highest boiling and melting points***. Conversely, elements at the beginning or the end of the *p*-block and *d*-block have the lowest boiling and melting points.





Assignment

4.3 pg. 141 # 1 to 7, 9 to 12, 14, 16, 17

Chapter 4 Review pg. 150−153 #14 to 16, 18 to 27, 29 to 33, 48, 51, 52, 59 to 64