Unit 2: Nuclear Chemistry and Chemical Periodicity

Chapter 28: Nuclear Chemistry

28.1: Nuclear Radiation

Nucleons: - the particles that make up a nucleus of an atom (**protons,($^{1}_{1}p^{+}$ or $^{1}_{1}$H) and neutrons, ($^{1}_{0}n$)**).

Isotopes: - atoms that have different mass number but the same atomic number or number of protons.

Nuclide: a particular atom or isotope containing specific numbers of protons and neutrons

Mass Number (# of p^+ and n) → Atomic Number (# of p^+) →	$^{A}_{Z}X$ ←	Element Symbol (based on Atomic #)

Radioisotopes: - unstable atoms that have different mass number but the same atomic number or number of protons.
- when decaying to more stable atoms, they tend to emit large amount of energy.

Radiation: - the penetrating rays and particles emitted by a radioisotopes that can also have ionizing power.

Radioactivity: - the particles and/or electromagnetic radiation that are emitted due to unstable nuclei.
- all elements having atomic number 84 (Polonium) and greater are radioactive.

History of Radioactivity

- In 1896, **Wilhelm Roentgen** discovered X-ray by examining ray emitting from the outside of the cathode ray glass tube. It has the capability of passing through solid materials, but can be blocked by denser matters. It can also be exposed to photographic plate, resulting from an image of "seeing through" an container made of less dense material.
- In the same year, **Antoine Becquerel** discovered that uranium emits a ray onto a photographic plate in the absence of sunlight or other forms of energy.
- A few years later, **Marie and Pierre Curie** demonstrated that radiation can be emitted by other elements. They discovered two new elements (polonium and radium) based on their tendency to emit radiation (**radioactivity**).

Radioactive Decay: - a process where a radioisotope decomposes into a different nucleus to achieve more stability.

Types of Radioactive Particle and Decay

1. **Alpha Particle (α particle)**: - basically a helium nucleus (**$^{4}_{2}$He**), commonly found during radioactive decay from heavier nuclide (the net result is to **increase the neutron to proton ratio** – more explanation in the next section).

Example: $^{218}_{84}\text{Po} \rightarrow \, ^{214}_{82}\text{Pb} + \, ^{4}_{2}\textbf{He}$

2. **Beta Particle** (**β particle**): - basically an electron (**$^{0}_{-1}e$ or $^{0}_{-1}\beta$**) that is emitted when the neutron to proton ratio is higher than the zone of stability **(a neutron is transformed to a proton and an electron as a result** – more explanation in the next section**)**.
 - electrons have a mass number of 0 and an atomic number assignment of –1, due to its charge.

Example: $^{214}_{82}\text{Pb} \rightarrow\ ^{214}_{83}\text{Bi} +\ ^{0}_{-1}e$ $\qquad$ ($^{1}_{0}n \rightarrow\ ^{1}_{1}p +\ ^{0}_{-1}e$)

3. **Gamma Ray** (**γ ray**): - also known as a high-energy photon (**$^{0}_{0}\gamma$**) that is usually a by-product of an alpha-particle decay.
 - photon has no mass and no atomic number.

Example: $^{238}_{92}\text{U} \rightarrow\ ^{234}_{90}\text{Th} +\ ^{4}_{2}\text{He} + 2\ ^{0}_{0}\gamma$

4. **Positron** (**e^+**): - an antimatter of electron (**$^{0}_{1}e$ or $^{0}_{1}\beta$**) that is emitted when the neutron to proton ratio is lower than the zone of stability (**a proton is transformed to a neutron** as a result - more explanation in the next section).
 - positrons have a mass number of 0 and an atomic number of 1, due to its charge.
 - when a positron and an electron collide, they **annihilate** themselves to produce energy **(matter-antimatter reaction)**.

Example: $^{15}_{8}\text{O} \rightarrow\ ^{15}_{7}\text{N} +\ ^{0}_{1}e$ $\qquad$ ($^{1}_{1}p \rightarrow\ ^{1}_{0}n +\ ^{0}_{1}e$)

5. **Electron Capture**: - an inner-orbital electron is "captured" by the nucleus to **increase neutron to proton ratio**. It is usually accompanied by an emission of gamma ray.

Example: $^{73}_{33}\text{As} +\ ^{0}_{-1}e \rightarrow\ ^{73}_{32}\text{Ge} +\ ^{0}_{0}\gamma$

Balancing Nuclear Equations: - the total atomic number (Z) and the total atomic mass (A) have to balance on both sides.

Example 1: Balance the following nuclear equations.

a. $^{222}_{86}\text{Rn}$ produces an α particle.

$^{222}_{86}\text{Rn} \rightarrow\ ^{218}_{84}\textbf{Po} +\ ^{4}_{2}\textbf{He}$ $\qquad$ A: 222 = **(218)** + **(4)**
Z: 86 = **(84 → Po)** + **(2)**

b. $^{14}_{6}\text{C}$ produces a β particle.

$^{14}_{6}\text{C} \rightarrow\ ^{14}_{7}\textbf{N} +\ ^{0}_{-1}e$ $\qquad$ A: 14 = **(14)** + **(0)**
Z: 6 = **(7 → N)** + **(–1)**

c. $^{49}_{21}\text{Sc}$ produces a β particle and a neutron.

$^{49}_{21}\text{Sc} \rightarrow\ ^{48}_{22}\textbf{Ti} +\ ^{0}_{-1}e +\ ^{1}_{0}\textbf{n}$ $\qquad$ A: 49 = **(48)** + **(0)** + **(1)**
Z: 21 = **(22 → Ti)** + **(–1)** + **(0)**

d. $^{11}_{6}\text{C}$ produces a positron.

$^{11}_{6}\text{C} \rightarrow\ ^{11}_{5}\textbf{B} +\ ^{0}_{1}e$ $\qquad$ A: 11 = **(11)** + **(0)**
Z: 6 = **(5 → B)** + **(1)**

e. $^{40}_{19}\text{K}$ captures an electron to produce γ ray

$^{40}_{19}\text{K} +\ ^{0}_{-1}e \rightarrow\ ^{40}_{18}\textbf{Ar} +\ ^{0}_{0}\gamma$ $\qquad$ A: 40 + **(0)** = **(40)** + **(0)**
Z: 19 + **(–1)** = **(18 → Ar)** + **(0)**

f. ${}^{1}_{1}\text{H}$ reacts with ${}^{15}_{7}\text{N}$ to produce an α particle with γ ray.

$${}^{1}_{1}\text{H} + {}^{15}_{7}\text{N} \rightarrow {}^{12}_{6}\textbf{C} + {}^{4}_{2}\textbf{He} + {}^{0}_{0}\gamma$$

A: 1 + 15 = **(12)** + **(4)** + (0)
Z: 1 + 7 = **(6 → C)** + **(2)** + (0)

g. ${}^{20}_{9}\text{F} \rightarrow {}^{20}_{10}\text{Ne} +$ ______

A: 20 = 20 + **(0)**
Z: 9 = 10 + **(–1)**

$$\left.\begin{matrix} A: 20 = 20 + (0) \\ Z: 9 = 10 + (-1) \end{matrix}\right\} {}^{0}_{-1}e$$

$${}^{20}_{9}\text{F} \rightarrow {}^{20}_{10}\text{Ne} + {}^{0}_{-1}e$$

h. ${}^{239}_{94}\text{Pu} +$ ____ $\rightarrow {}^{242}_{96}\text{Cm} + {}^{1}_{0}\text{n}$

$$\left.\begin{matrix} A: 239 + (4) \\ Z: 94 + (2) \end{matrix}\right\} {}^{4}_{2}\textbf{He} \quad \begin{matrix} = 242 + 1 \\ = 96 + 0 \end{matrix}$$

$${}^{239}_{94}\text{Pu} + {}^{4}_{2}\textbf{He} \rightarrow {}^{242}_{96}\text{Cm} + {}^{1}_{0}\text{n}$$

i. ${}^{54}_{27}\text{Co} \rightarrow$ ______ $+ {}^{0}_{1}e$

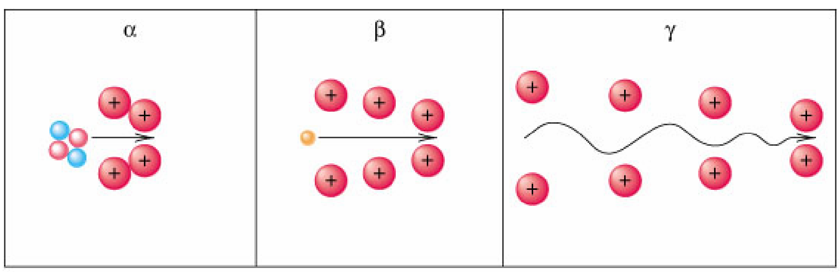
A: 54 = **(54)** + 0
Z: 27 = **(26 → Fe)** + 1

$${}^{54}_{27}\text{Co} \rightarrow {}^{54}_{26}\textbf{Fe} + {}^{0}_{1}e$$

Effects from Radiation

1. **Penetrating Ability**: - the lighter the particles, the more penetrating they can be. In terms of penetrating ability: ***γ* ray is the strongest, follows by *β* particles and α particle is the least penetrating.**

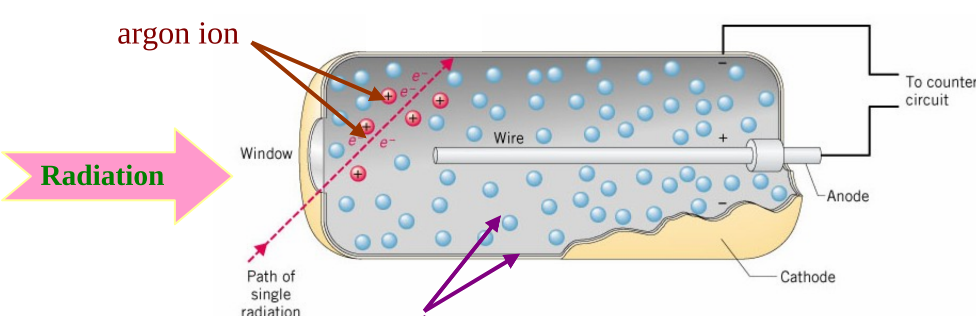
2. **Ionization Ability**: - as high-energy particles pass through tissue, it can cause ionization that is damaging to the organism. **α particles can ionize the most along its path whereas *γ* ray does not.** Therefore, α particle producers like plutonium and radon can cause severe radiation damage if ingested or inhaled.



High ionization along path but low penetration | Moderate ionization along path but low penetration | Low ionization along path but very high penetration

Units for Radiation: - ***rems,*** measured the ability to radiation to cause harm biologically.
- it takes 500 rems to be considered lethal radiation dosage if it was given in a short period of time.
- smaller radiation on a body can be measured in ***millirems***.

Geiger-Müller Counter: - sometimes refer to as the **Geiger Counter**.
- argon gas becomes ionized and when struck by high-energy particle from radioactive decay. The resulting electric potential is amplified and the current can show as the intensity of the radioactivity.



Effects of Radiation on Biological Matter

- depends on the intensity, dosage, and the period of time a body is exposed to the radiation.
- high intensity at large dosage in a short period of time can do irreversible damages such as **mutation** (an alteration of the DNA).
- reasonable dosage over a long period of time may not cause permanent damage as cells are allowed time to heal.

Natural Radiations: - small amounts of radiation happened naturally from the environment.

- **Cosmic radiation** from outer space and the sun (amounts vary by elevation from sea level).
- **Ground radiation** from Earth's interior that is responsible from heat of hot springs and geyser to the molten core of the planet. It is also present in organic food and water, building material such as bricks and wood products.
- **Air radiation** from randon-222 (inert gas from natural uranium deposits); easily inhale from basement cracks. Randon-222 is also found in tobacco smoke (man-made radiation).
- **Human tissues** contains about 20 mg of potassium-40 that emit pulses of radioactivity over time.

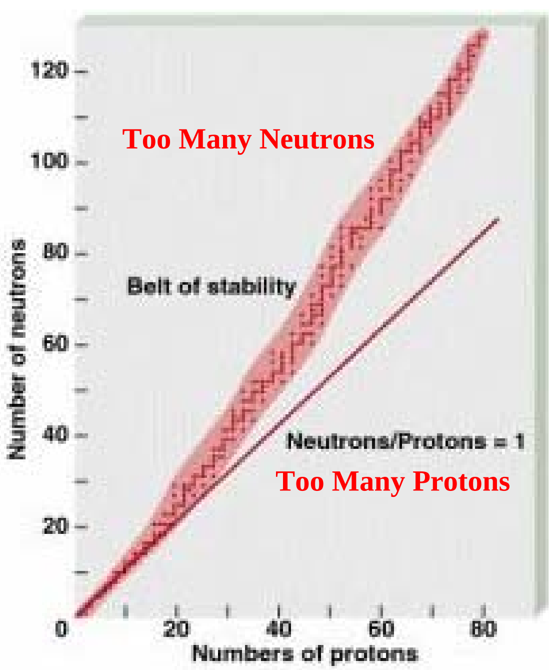
Man-Made Radiations: - radiations that are artificially created (intentional and unintentional).

- **Medical procedures** such as X-ray and radiotherapy and radiodiagnostic procedures.
- **Consumer products** like television tubes.
- **Proximity to Power Generators** like nuclear and coal power plants.
- **Aviation** from airline travels.

28.2: Nuclear Transformation

Radioactive Decay: - a process where a radioisotope decomposes into different nuclei to gain stability.

Common Observations of Radioactive Decay



1. When a nuclide has 84 or more protons ($Z \geq 84$), it tends to be **unstable** and likely undergo **radioactive decay**.

2. Lighter nuclides are stable when $Z = n$ (or $n : p^+$ ratio = 1). However, heavier nuclides are stable only when $Z < n$ (or $n : p^+$ ratio > 1).

3. Nuclides with **even # of p^+ with even # of n are more stable** than nuclides with odd # of p^+ and odd # of n.

   Example: Most Stable to Least Stable Nuclides
   ($^{12}_{6}C$, $^{13}_{6}C$, $^{19}_{9}F$, $^{6}_{3}Li$)

Nuclide	$^{12}_{6}C$	$^{13}_{6}C$	$^{19}_{9}F$	$^{6}_{3}Li$
# of p^+	6 (even)	6 (even)	9 (odd)	3 (odd)
# of n	6 (even)	7 (odd)	10 (even)	3 (odd)

Stability: **Most** ⟶ **Least**

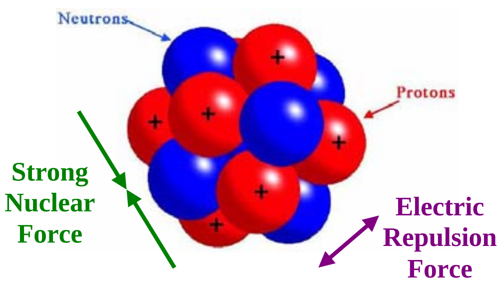
4. **Magic Numbers of protons or neutrons (2, 8, 20, 28, 50, 82 and 126) results in very stable nuclides**.

Zone of Stability: - a graph that depicts the relationship between the number of neutrons versus the number of protons, and the area where there are stable nuclides.

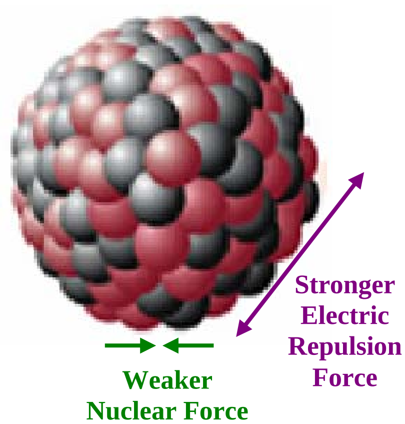
Strong Nuclear Force: - a force of attraction that is present over extremely short distance (1×10^{-15} m) between all nucleons (protons and neutrons).
- it is much stronger than electromagnetic force in short distances. However, electromagnetic force is more significant over longer distances.

Properties of Neutrons:

1. **Neutrons** serve as ***"nuclear cement"***, gluing neighbouring protons together despite the electric repulsion of positive charges, but only **over short distances**.
2. At large distances, strong nuclear force becomes less significant. Hence, **the MORE Protons in the nucleus (Heavier Atoms), the MORE Neutrons are needed to hold them together**. Sometimes this means for every 1 proton, there are 1.5 times to twice as many neutrons.
3. **SMALLER Atomic Nuclei usually have the SAME Number of Protons as Neutrons.**
4. **A Single Neutron is rather UNSTABLE. It will CONVERT itself to a Proton and an Electron.**



For Small Nucleus, Strong Nuclear Force is MORE significant compared to electric repulsive force between protons due to small distances between nucleons.

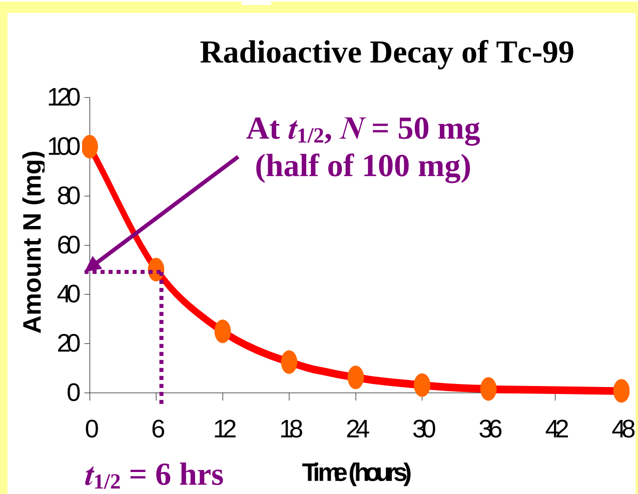


For Large Nucleus, Electric Repulsive Force between Protons is MORE significant compared to strong nuclear force because of the large distances between nucleons. Hence, they are inherently unstable and likely undergo alpha decay (see below).

Half-Life ($t_{1/2}$): - the amount of time it takes to half the amount of radioactive nuclides.

Example 1: Technetium-99, the first synthetic element in the Table, is used as a radiotracer for many organs such as heart, liver and lungs. It has a half-life of 6.00 hours. Draw a graph showing how 100 mg of $^{99}_{43}$Tc decays over time. What is the radioactive amount of $^{99}_{43}$Tc after 2.00 days?

Time	Amount of $^{99}_{43}$Tc
0 hour	100 mg
6.00 hours	50 mg
12.0 hours	25 mg
18.0 hours	12.5 mg
24.0 hours	6.25 mg
30.0 hours	3.125 mg
36.0 hours	1.5625 mg
42.0 hours	0.78125 mg
48.0 hours = 2 days	**0.39625 mg = 0.40 mg**



Radiocarbon Dating: - sometimes called **carbon-14 dating**. ${}^{14}_{6}C$ can be found naturally in organic material and the atmosphere. It decays as soon as the organism dies.

$$ {}^{14}_{6}C \rightarrow {}^{0}_{-1}e + {}^{14}_{7}N $$

- uses the known ratio of ${}^{14}_{6}C/{}^{12}_{6}C$ of similar organic sample of the day with the ratio in the artefact and the half-life of ${}^{14}_{6}C$ being 5730 years to determine the age of the artefact.
- it is accurate within 50,000 years in the past. Beyond that time, the amount of carbon-14 is too little to make an accurate analysis.

Uranium-238 Dating: - due to its lengthy half-life (4.5×10^9 years), it is used to date rocks and other ancient inorganic material. ${}^{238}_{92}U/{}^{206}_{82}Pb$ ratio is used as ${}^{238}_{92}U$ eventually decays to stable ${}^{206}_{82}Pb$.

Example 2: An archaeologist unearthed a femur that contained one-sixteenth the amount of carbon-14 compared to a similar size femur of a person today. Knowing that carbon-14 has a half-life of 5730 years, what is the approximate age of this femur?

$\frac{1}{16}$ of the original amount is **4 half-lives**. $\left[\left(\frac{1}{2}\right)^4 = \frac{1}{16}\right]$ **Hence, it has been 4 × 5730 years = 22,920 years**

Isotopic Dating Using Logarithm

Logarithm can be used to solve for an exponent. This is helpful when we need to find the time elapsed or the half-life in isotopic dating. In general, we can solve for an exponent by doing the conversion to the right.

$$ b^x = a \qquad x = \frac{\log a}{\log b} $$

Example 3: Solve for x

a. $3^x = 21$

b. $(½)^x = 0.457$

c. $(0.5)^{\frac{x}{12.6}} = 0.761$

a.
$$ 3^x = 21 $$
$$ x = \frac{\log 21}{\log 3} $$
$$ x = 2.77 $$

```
log(21)/log(3)
     2.771243749
3^Ans
              21
```
Verify Answer

b.
$$ (½)^x = 0.457 $$
$$ x = \frac{\log 0.457}{\log(1/2)} $$
$$ x = 1.13 $$

```
log(0.457)/log(1
/2)
      1.12973393
(1/2)^Ans
            .457
```
Verify Answer

c.
$$ (0.5)^{\frac{x}{12.6}} = 0.761 $$
$$ \frac{x}{12.6} = \frac{\log 0.761}{\log 0.5} $$
$$ x = \frac{12.6 \times \log 0.761}{\log 0.5} $$
$$ x = 4.96 $$

```
12.6*log(0.761)/
log(0.5)
     4.964798679
0.5^(Ans/12.6)
            .761
```
Verify Answer

Radioactive Decay Formula

$$ N = N_o \left(\frac{1}{2}\right)^{\left(\frac{t}{t_{1/2}}\right)} $$

N = Final Amount **t = time elapsed**
N_o = Initial Amount **$t_{½}$ = half-life**

Example 4: Cobalt-60 is commonly used in chemotherapy and it has a half-life of 5 years. If a patient has received 6.00 mg of cobalt-60 during her treatment, how many years will it take until 0.25 mg of the original isotope remains?

N = 0.25 mg
N_o = 6.00 mg
t = ?
$t_{½}$ = 5 years

$$N = N_o\left(\frac{1}{2}\right)^{\left(\frac{t}{t_{1/2}}\right)}$$

$$0.25 = 6.00\,(0.5)^{(t/5)}$$

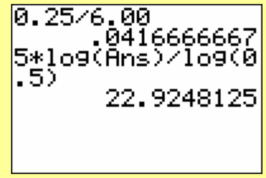
$$\frac{0.25}{6.00} = (0.5)^{(t/5)}$$

$$0.04166666 = (0.5)^{(t/5)}$$

$$\frac{\log 0.04166666}{\log 0.5} = \frac{t}{5}$$

$$\frac{5 \times \log 0.04166666}{\log 0.5} = t$$

t = 22.9 years t = **23. years**



Example 5: A radioactive isotope decayed to 35.4% of its original value in 4.26 millisecond. What is its half-life?

N = 35.4 %
N_o = 100 %
t = 4.26 ms
$t_{½}$ = ?

$$N = N_o\left(\frac{1}{2}\right)^{\left(\frac{t}{t_{1/2}}\right)}$$

$$35.4 = 100\,(0.5)^{(4.26/t_{1/2})}$$

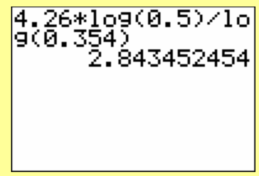
$$\frac{35.4}{100} = (0.5)^{(4.26/t_{1/2})}$$

$$0.354 = (0.5)^{(4.26/t_{1/2})}$$

$$\frac{\log 0.354}{\log 0.5} = \frac{4.26}{t_{1/2}}$$

$$t_{½} = \frac{4.26 \times \log 0.5}{\log 0.354}$$

$t_{½}$ = **2.84 ms**



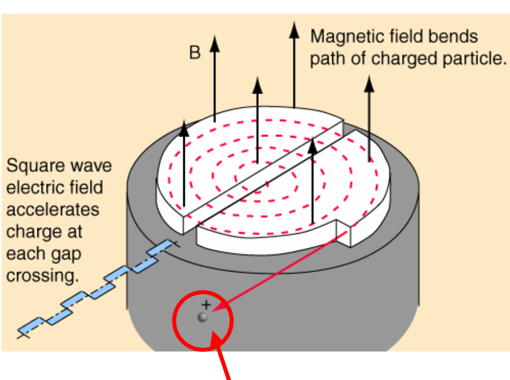
Transmutation: - the reaction where one element is converted to another element by changing the number of protons. This includes alpha and beta decays.

Transuranium Elements: - elements that have been synthesized by nuclear transformation after the last natural element, uranium.

Example: $^{244}_{94}\text{Pu} + {}^{48}_{20}\text{Ca} \rightarrow {}^{289}_{114}\text{Uuq} + 3\,{}^{1}_{0}\text{n}$ (Discovered in 1998 and $t_{1/2}$ = 30 seconds)

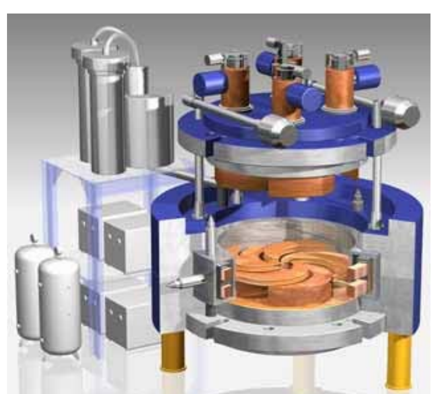
Particle Accelerator: - a device that alternates electric field to speed up a particle to add into a target nuclide.

a. **Cyclotron**: - a type of particle accelerator that utilizes a changing electric field along with a magnetic field to increase the speed of an ion around a disc before hitting a target nuclide.



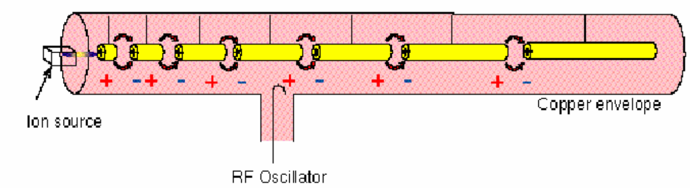
Schematic of a Cyclotron

Accelerated charged particle to collide with target nuclide



COMET: A medical superconducting cyclotron. It is used to generate thallium-201 (coronary arteries) and gallium-67 (soft-tissue tumors). It can also produce radio-pharmaceutical needed for PET and SPECT scans

b. Linear Accelerator: - a particle accelerator that speeds up a particle by using an alternating electric field at different segment of a linear tube to add an ion into a target nuclide.



Left: Schematic of a Linear Accelerator

Right: Stanford Linear Accelerator

Assignment

28.1 pg. 844 #1 to 3

28.2 pg. 851 #6, 8 and 10 and Worksheet: Isotopic Dating and Half Life

Worksheet: Isotopic Dating and Half-Life

1. Strontium-90 has a half-life of 28 years. How long will it take for all of the strontium-90 presently on earth to be reduced to 1/32 of its present amount?
2. The half-life of plutonium-239 is 24,110 years. Of an original mass of 100. g, how much remains after 96,440 years?
3. The half-life of thorium-227 is 18.72 days. How many days are required for only 1/8 of the original amount to be left?
4. The half-life of cesium-137 is 30 years. Of an initial 100. g sample, how much cesium-137 will remain after 300 years?
5. A 24.0 g sample of Ra-226 is supplied to a lab. It is predicted 9540 years later, only 0.375 g of Ra-226 will remain. What is the half-life of Ra-226?
6. Mercury-197 is used for kidney scans and has a half-life of 3 days. If the amount of mercury-197 needed for a study is 4.00 gram and the time allowed for shipment is 14 days, how much mercury-197 will need to be ordered?
7. If one million atoms of element-106 were prepared and after 12.33 seconds, only 75 atoms remain, what is the half-life of element-106?
8. Iodine-131 is to treat hyperthyroidism and it has a half-life of 8 days. If the amount of iodine-131 used was 50.0 mg on a patient, how much time must elapse before his iodine-131 level reaches 3.00 mg?
9. Uranium-235 is used in geological dating and it has a half-life of 4.5 billion years. If a piece of rock contains 2.00 mg of uranium-235 out of an estimated 3.20 mg as initial mass. How old is this rock?
10. A new element is found and it decayed to 15.3% of its original amount in 2.4 nanoseconds. What is its half-life?

Answers

1. 140 years	**2.** 6.25 g	**3.** 56.16 days	**4.** 0.0977 g	**5.** 1.59×10^3 years
6. 102. g	**7.** 0.8998 seconds	**8.** 32.5 days	**9.** 3.2 billion years	**10.** 0.89 ns

28.3: Fission and Fusion of Atomic Nuclei

Nuclear Fission: - the breaking up of a heavier nucleus into two nuclei with small mass number.

Example: $^{235}_{92}\text{U} + ^{1}_{0}\text{n} \rightarrow ^{141}_{56}\text{Ba} + ^{92}_{36}\text{Kr} + 3\,^{1}_{0}\text{n}$ $\Delta H = 1.692 \times 10^{10}$ kJ

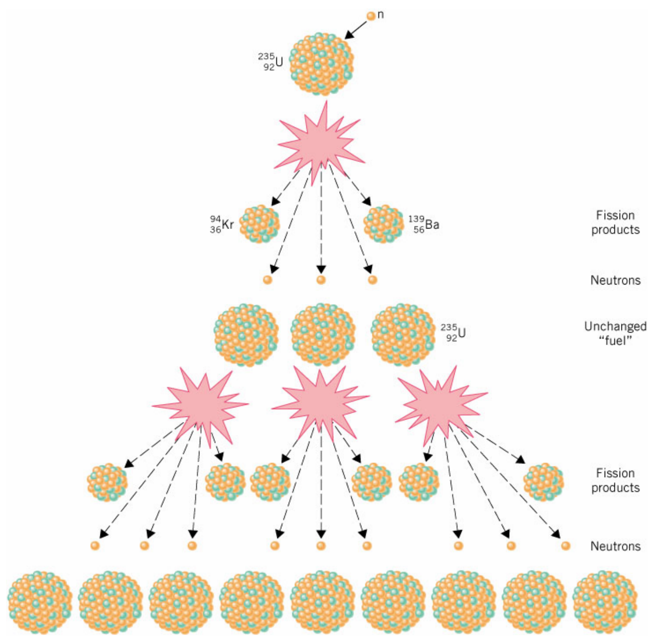
(See Animation at http://reactor.engr.wisc.edu/fission.htm)

Chain Reaction: - when the nuclear fission is self-sustaining.

a. **Subcritical**: - when there is on average, less than one neutron produced per $^{235}_{92}$U consumed. The fission will eventually stop.

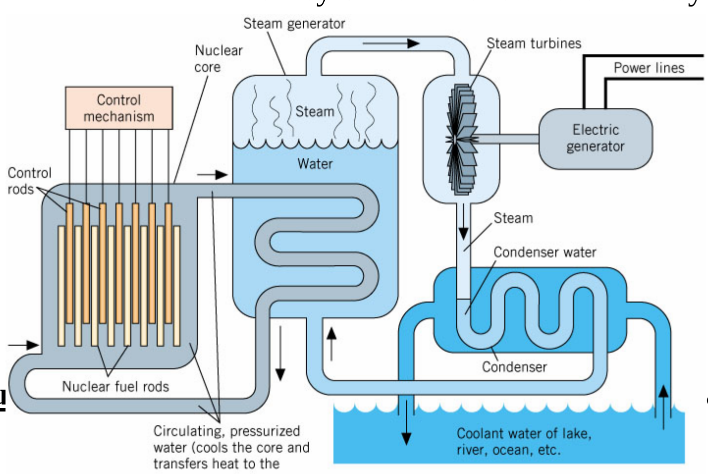
b. **Critical**: - when there is on average, exactly one neutron produced per $^{235}_{92}$U consumed. The fission can then be self-sustaining at the same level.

c. **Supercritical**: - when there is one average, more than one neutron produced per $^{235}_{92}$U consumed. The fission can increase its rate rapidly and a violent explosion can result.



Nuclear Reactors: - fission reactors where enriched $^{235}_{92}$U is placed in the **reactor core (nuclear core)**, **control rods (usually made of carbon to absorb extra neutrons)** can be lifted or lowered to control the rate of the fission process. As the water of the surrounding is heated, it transferred the heat to a steam generator to generate electricity via a steam turbine. The water is then cooled and recycled.

- the by-products of $^{235}_{92}$U fission have very long half-lives and can remain radioactive for a long time. Great efforts are needed to dispose of the wastes properly. The danger of a nuclear meltdown is also a constant danger as in the cases of Three Mile Island, Pennsylvania in 1979 and Chernobyl, Ukraine in 1986.

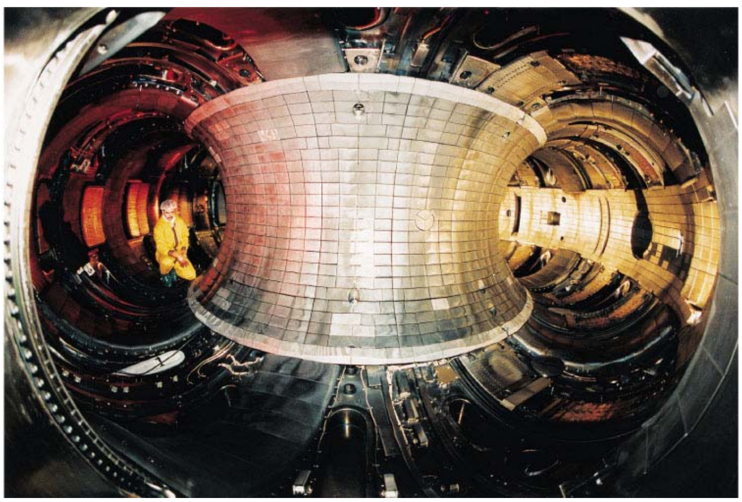


Nu a h

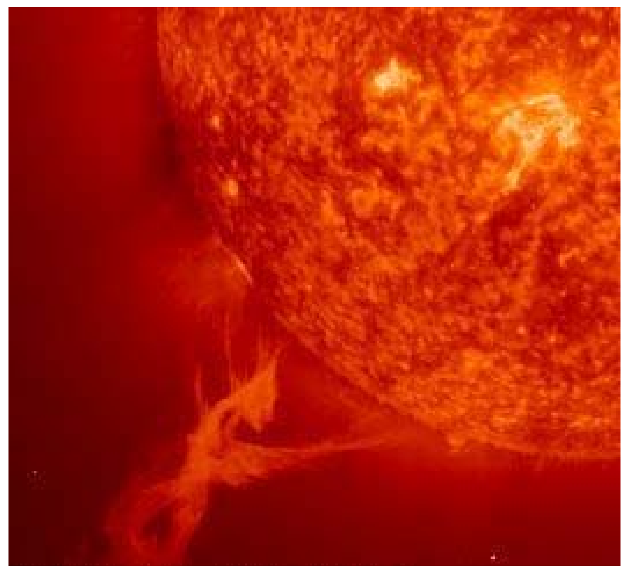
Example: ${}^{2}_{1}H + {}^{3}_{1}H \rightarrow {}^{4}_{2}He + {}^{1}_{0}n$ $\Delta H = 1.698 \times 10^{9}$ kJ

(See Animation at http://reactor.engr.wisc.edu/fission.htm)

- the availability of hydrogen isotopes, deuterium (${}^{2}_{1}H$) and tritium (${}^{3}_{1}H$), in sea water and the harmless product, ${}^{4}_{2}He$, makes nuclear fusion an environmental friendly alternative to generate power.
- however, fusion reactions such as the one above usually require initial temperature above 4×10^{7} K to overcome the strong electrostatic repulsion between the two protons (the release of significant binding energy can only achieve when the distance between the two protons is approximately 10^{-15} m). High-powered laser and heating by electric currents are being studied as methods to attain this high temperature to initial a control fusion reaction.



European Tokamak Fusion Test Reactor Vacuum Vessel employs the design of a toroid with a super strength magnetic field to contain plasma without having it touch the wall of the reactor. A similar experimental fusion reactor can also be found at Princeton, USA.



Fusion reaction is the driving force of our sun's energy.

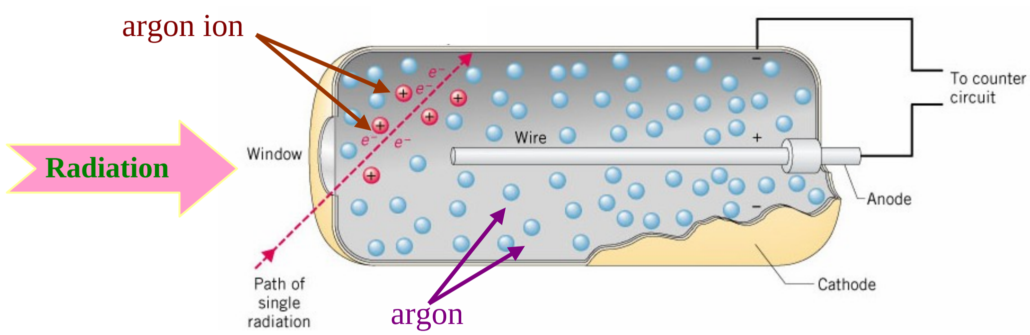
Relative Amounts of Energy Involved

Nuclear >> Chemical Reactions >> Physical Potential (Phase Changes) >> Kinetic (Temperature Change)

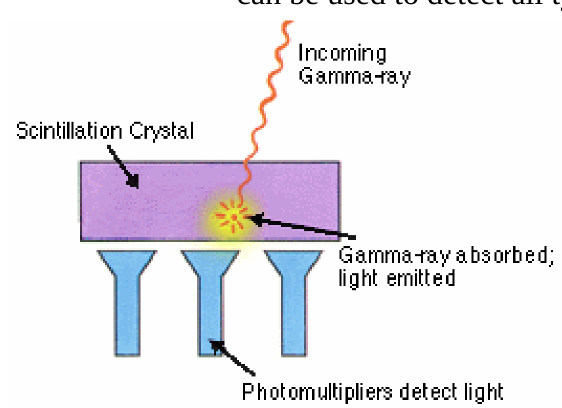
28.4: Radiation in Your Life

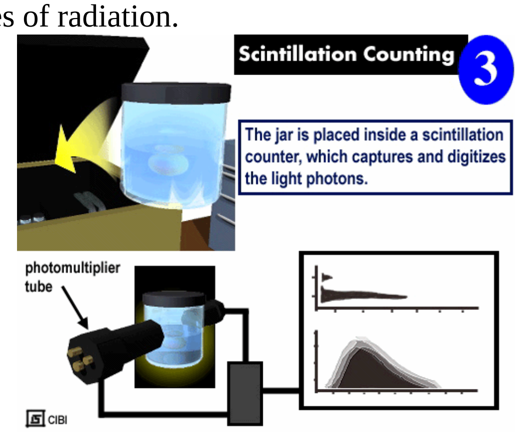
Detecting Radiations

1. **Geiger-Müller Counter**: - sometimes refer to as the **Geiger Counter**.
   - argon gas becomes ionized and when struck by high-energy particle from beta radioactive decay. The resulting electric potential is amplified and the current can show as the intensity of the radioactivity.
   - only detects beta radiation but not alpha radiation.



2. **Scintillation Counter**: - zinc sulfide and other substances give off light when struck by high-energy particle from any radioactive decay. A photocell measure the intensity of the light produced and gives the measure as the number of decay events per unit of time.
   - can be used to detect all types of radiation.





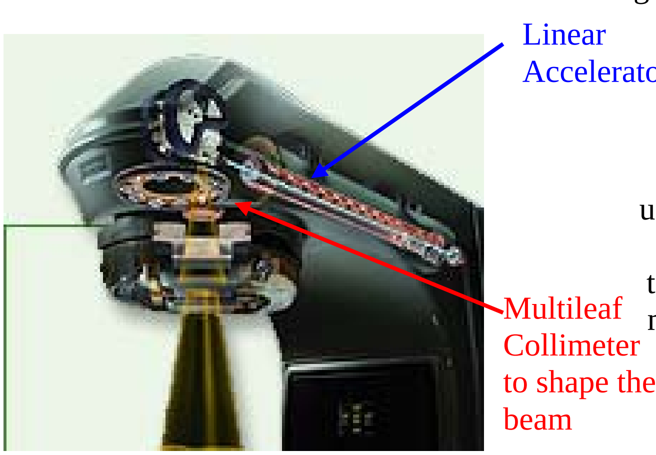
Medical Uses of Radiations

1. **Radioactive Tracers**: - radioactive elements that leave a path of radiation that can be imaged to determine how an organism takes up the material.
   - radioactive isotope can be included in fertilizer to study how plants absorb nutrients.
   - compound that contain a radioactive tracer (**carrier compound**) can be introduced to a patient, and an device that can pick up radiation produces an image for the purpose of diagnosis (**medical imaging**).

Some Common Tracers and their Usages

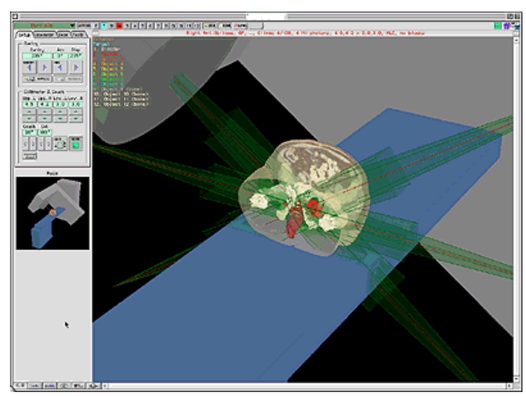
Radiotracers	Area of the body examined	Other Usages
${}^{131}_{53}I$	Thyroid	
${}^{59}_{26}Fe$ and ${}^{51}_{24}Cr$	Red Blood Cells	
${}^{252}_{98}Cf$		Explosive Detections
${}^{99}_{42}Mo$	Metabolism	
${}^{32}_{15}P$	Eyes, Liver, Tumours	
${}^{192}_{77}Ir$		Metal Integrity Tests
${}^{87}_{38}Sr$ and ${}^{47}_{20}Ca$	Bones	
${}^{99}_{43}Tc$	Heart, Bones, Liver, and Lungs	
${}^{241}_{95}Am$		Smoke Detection
${}^{133}_{54}Xe$	Lungs and Blood flow	
${}^{24}_{11}Na$	Circulatory System	

2. **Radiation Therapy**: - in treating cancer, when traditionally surgeries are impossible to perform or fail to eliminate the cancer tissues, radiation is directed onto the cancer area in the hopes it will kill of the tumour.
   - in most cases, some healthy tissue surrounding the tumour is taken out as well so there is no recurrence. However, the amount of healthy tissues sacrifice is minimized using previous 3-D imaging and pre-programming computer to target the radiation into the right area.



Schematic of a Medical Linear Accelerator

Computer Program is used to control radiotherapy treatment by a medical linear accelerator



Assignment

28.3 pg. 856 #11 to 15
28.4 pg. 861 #16 to 21
Chapter 28 Review: pg. 864–865 #22 to 29, 34 to 36, 39, 41, 45 to 47; pg. 867 #1 to 6, 10 to 13

Chapter 13: Electrons in Atoms

13.1: Models of Atom

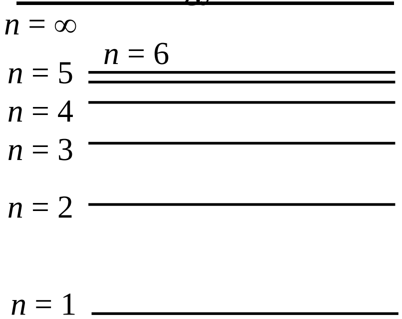
As mentioned in Section 5.2: Structure of the Nuclear Atom, the evolution on the Atomic Theories is as followed:

1. **Dalton's Atomic Model**: - John Dalton theorized that atoms are the smallest fundamental particle unique to different elements. They can never be created or destroyed.

2. **Thomson's Plum Pudding Model**: - negative subatomic particles called electrons are spread out around a "positive charged spherical cloud" much like plum pudding.

3. **Rutherford's Nuclear Model**: - the innermost part of the atom is called a nucleus, which consists of neutrons and protons. An equal number of electrons fly around the nucleus so an atom is neutral and mainly composed of empty space.

4. **Bohr's Energy Level Model**: - the electrons around the nucleus move about in a preset **orbits (energy level)** like planets around the sun. Each orbit can house different number of electrons. Electrons cannot exist between energy level. ***(Electrons are quantized* – can exist at one level at a time.*)* Quantum refers to the amount of energy from one orbit to the next.** The number of electrons in each energy level is the same as the number of elements in each row of the Table of Elements. The higher the energy level, the unstable it is for the electrons because it has higher potential energy.

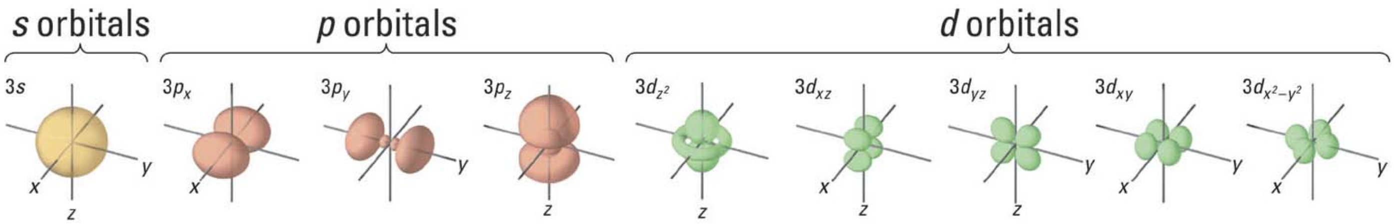


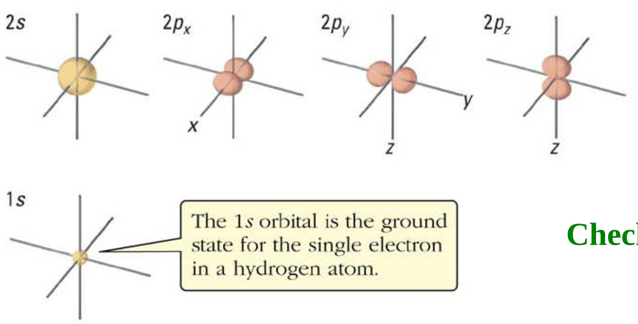
5. **Quantum Mechanical Model**: - proposed by **Erwin Schrödinger (electrons have wave properties)** and **Werner Heisenberg (due to the speed of an electron, we can never locate its position and determine its speed at the same time)**, electrons in an atom moves around the nucleus in specific **orbitals (an region defined by the probability of where the electrons might be found).** *Note*: Orbitals do not have to be circular but it is three-dimensional in form and can come in different shape.

Quantum Atomic Orbitals: - are consisted of two parts: **the Principle Energy Level (*n*) and Type of Sublevel (*s*, *p*, *d* or *f*).**
- each atomic orbitals can house two electrons only.

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

b. **Type of Sublevel**: - indicate the shape of the atomic orbitals (***s***, ***p***, ***d*** **or** ***f***).





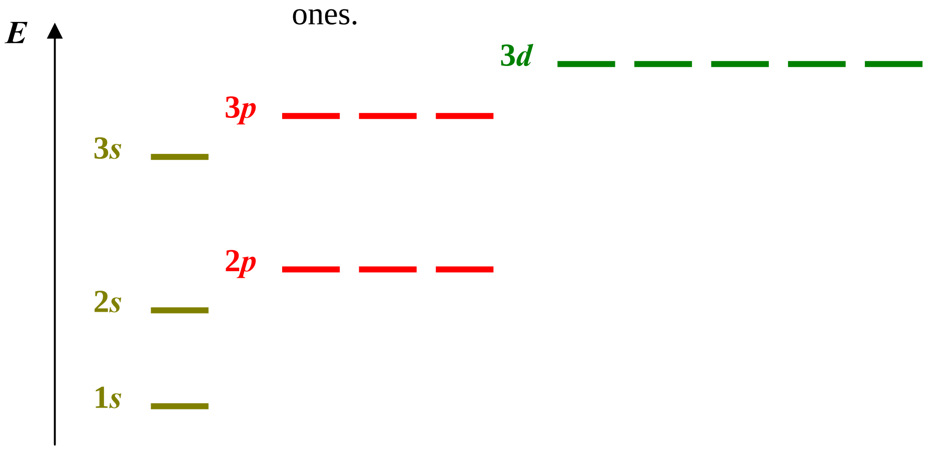
s **orbital – "sharp" – 1 orbital**
p **orbitals – "principal" – 3 orbitals**
d **orbitals – "diffuse" – 5 orbitals**
f **orbitals – "fundamental" – 7 orbitals**
g **orbitals – follows the alphabet – 9 orbitals**

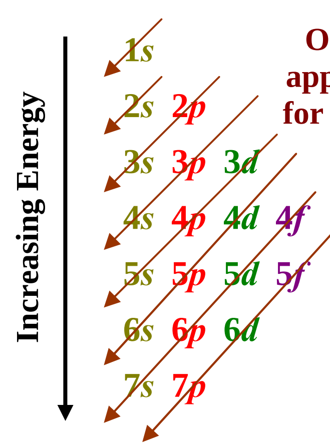
Check out the Hydrogen Atom Orbital Viewer at
http://www.falstad.com/qmatom/

n	Sublevel	# of Orbitals	Total # of Orbitals	Total # of Electrons
1	1*s*	1	1	2
2	2*s*	1	4	8
	2*p*	3		
3	3*s*	1	9	18
	3*p*	3		
	3*d*	5		
4	4*s*	1	16	32
	4*p*	3		
	4*d*	5		
	4*f*	7		

13.2: Electron Arrangement in Atoms

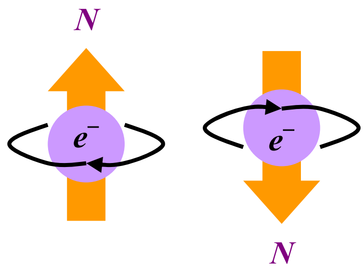
Aufbau Principle: - "*Aufbau*" German for "building up"
- for each element, electrons are built up from the lowest energy orbital to the higher ones.



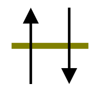


Order of appearance for Sublevel

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.



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

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

1*s* orbital

Hund's Rule: - for sublevels that can have more than 1 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.

Electron Configuration: - the arrangement of electrons in atomic sublevels.

Orbital Diagram: - a diagram that shows the arrangements of electrons in quantum sublevels.

Use the Electron Configuration Animation/Applet at
http://intro.chem.okstate.edu/WorkshopFolder/Electronconfnew.html to do the next two examples.
Note the exceptions for Cr and Cu.

Example 2: Draw the quantum orbital 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	$1s^2\ 2s^2\ 2p^6$	↑↓	↑↓	↑↓ ↑↓ ↑↓

Example 3: Draw the quantum orbital diagrams and state the electron configurations for the following atoms.

Atomic #	Atom	Electron Configuration	Orbital Diagram	4s	3d	3d	3d	3d	3d	4p	4p	4p
19	K	[Ar] $4s^1$	[Ar]	↑								
20	Ca	[Ar] $4s^2$	[Ar]	↑↓								
21	Sc	[Ar] $4s^2\,3d^1$	[Ar]	↑↓	↑							
22	Ti	[Ar] $4s^2\,3d^2$	[Ar]	↑↓	↑	↑						
23	V	[Ar] $4s^2\,3d^3$	[Ar]	↑↓	↑	↑	↑					
24	*Cr	[Ar] $4s^1\,3d^5$	[Ar]	↑	↑	↑	↑	↑	↑			
25	Mn	[Ar] $4s^2\,3d^5$	[Ar]	↑↓	↑	↑	↑	↑	↑			
26	Fe	[Ar] $4s^2\,3d^6$	[Ar]	↑↓	↑↓	↑	↑	↑	↑			
27	Co	[Ar] $4s^2\,3d^7$	[Ar]	↑↓	↑↓	↑↓	↑	↑	↑			
28	Ni	[Ar] $4s^2\,3d^8$	[Ar]	↑↓	↑↓	↑↓	↑↓	↑	↑			
29	*Cu	[Ar] $4s^1\,3d^{10}$	[Ar]	↑	↑↓	↑↓	↑↓	↑↓	↑↓			
30	Zn	[Ar] $4s^2\,3d^{10}$	[Ar]	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓			
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]	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓

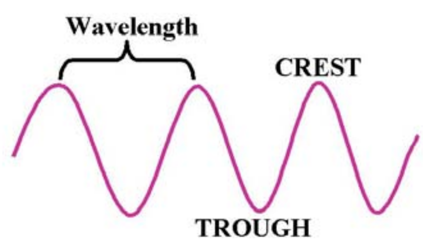
***From Hund's Rule, Cr and Cu can achieve lowest energy if the $4s^2$ e^- was moved to the $3d^5$ or $3d^{10}$.**

Some Exceptions in Electron Configurations

1. Chromium ($_{24}$Cr): [Ar] $4s^1\,3d^5$
2. Copper ($_{29}$Cu): [Ar] $4s^1\,3d^{10}$
3. Molybdenum ($_{42}$Mo): [Kr] $5s^1\,4d^5$
4. Technetium ($_{43}$Tc): [Kr] $5s^1\,4d^6$
5. Palladium ($_{46}$Pd): [Kr] $4d^{10}$
6. Silver ($_{47}$Ag): [Kr] $5s^1\,4d^{10}$

13.3: Physics and the Quantum Mechanical Model

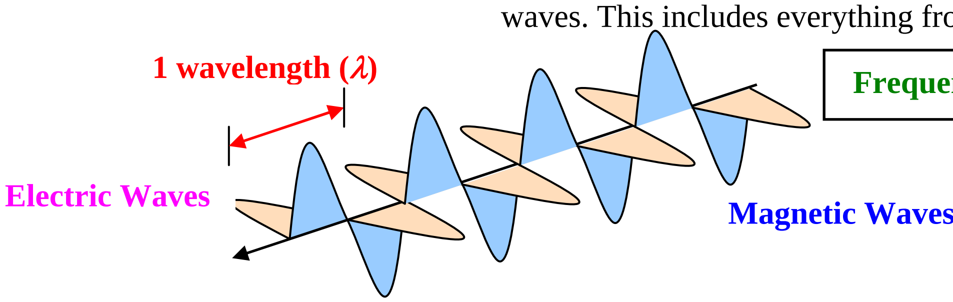
Wavelength: - the length of a wave (from crest to crest).
- measures in metres (m) or nanometres ($\times 10^{-9}$ m)

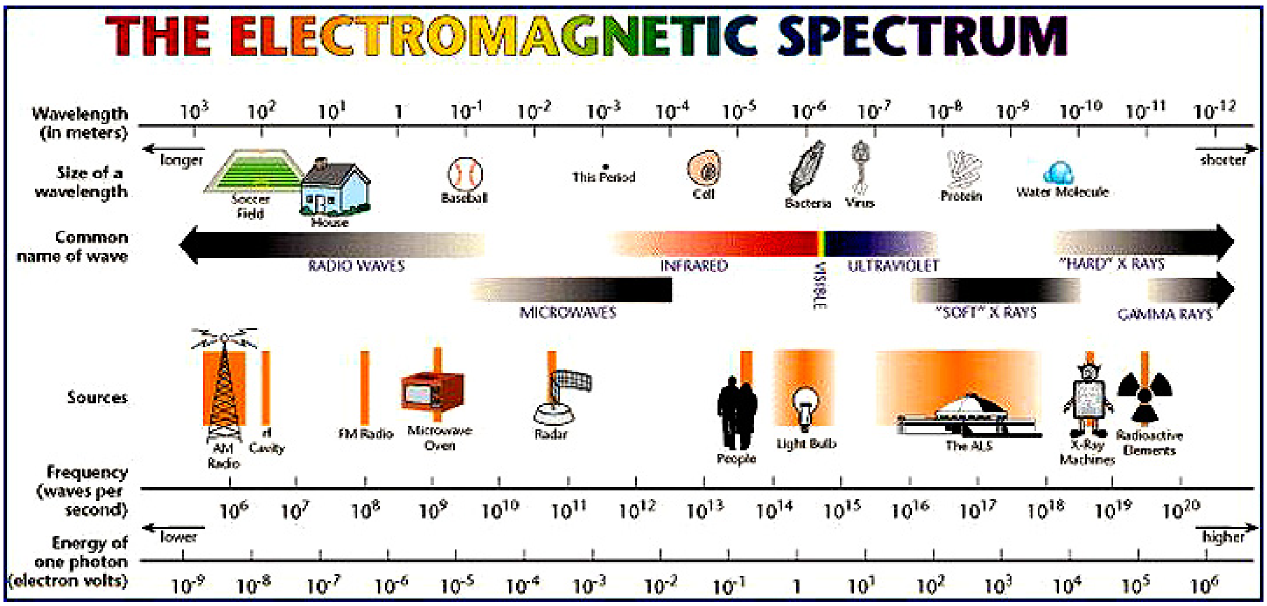


Frequency: - the number of wave in one second.
- measures in Hertz (Hz) or s^{-1}. (1 Hz = 1 s^{-1})

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.

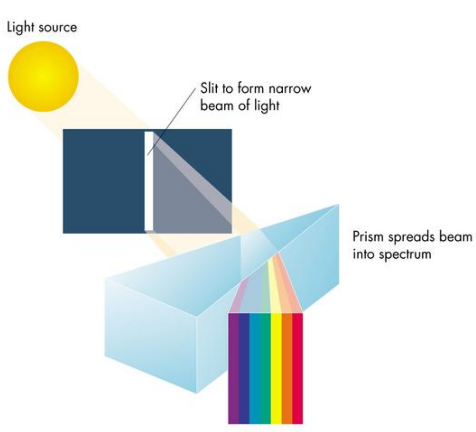
Frequency ↑ Wavelength ↓ Energy ↑



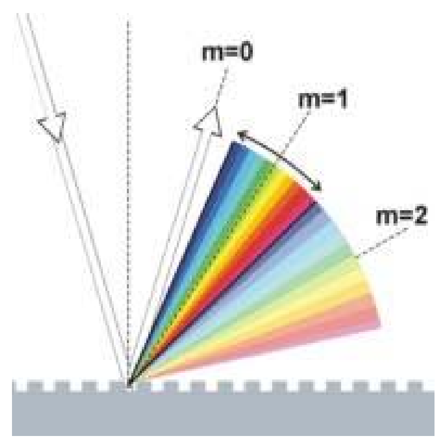


EM Wave	Frequency (Hz)	Wavelength (m)	EM Wave	Frequency (Hz)	Wavelength (m)
Cosmic Wave	10^{23}	10^{-15}	Infrared	10^{12}	10^{-4}
Gamma Wave	10^{20}	10^{-12}	Microwaves	10^{10}	10^{-2}
X-Ray	10^{18}	10^{-10}	FM Radio	10^{8} or (100 MHz)	1 to 10
Ultraviolet	10^{16}	10^{-8}	Shortwave Radio	10^{6} (1 MHz)	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}

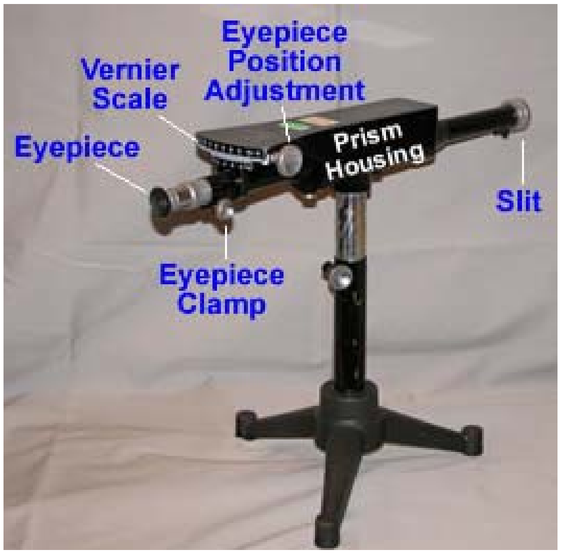
Diffraction Grating: - a surface contains a series of prisms that diffracts incoming light into their individual wavelengths.

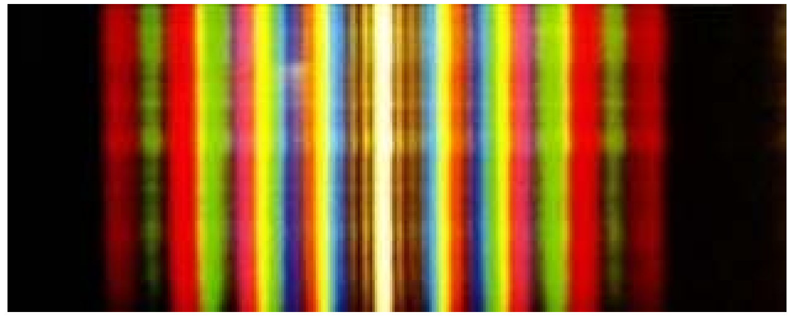


(Left) A single prism breaks up white light into the visible spectrum. (Right) A diffraction grating is able to generate series of visible spectrum due to the uneven surface.



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



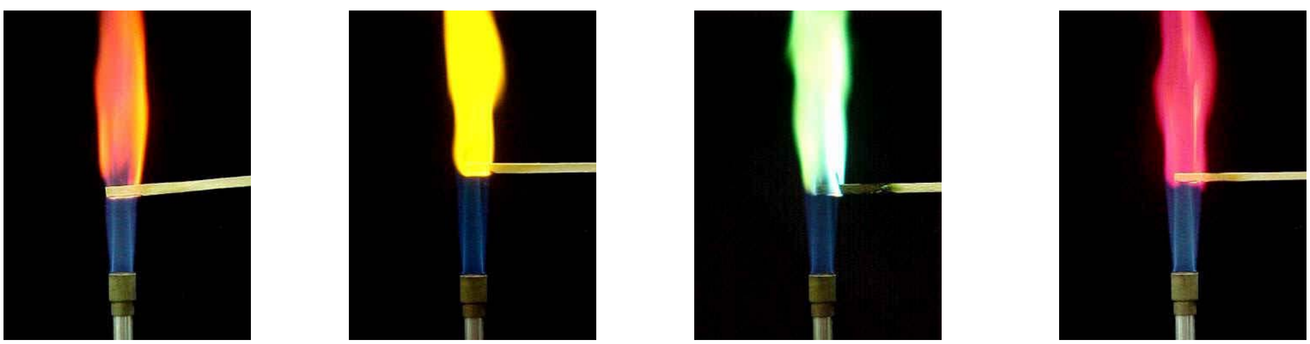


(Left) A Spectroscope. Light enters the slit and goes through the prism housing (may be replaced prism with a diffraction grating). The observer uses the eyepiece to see the image.

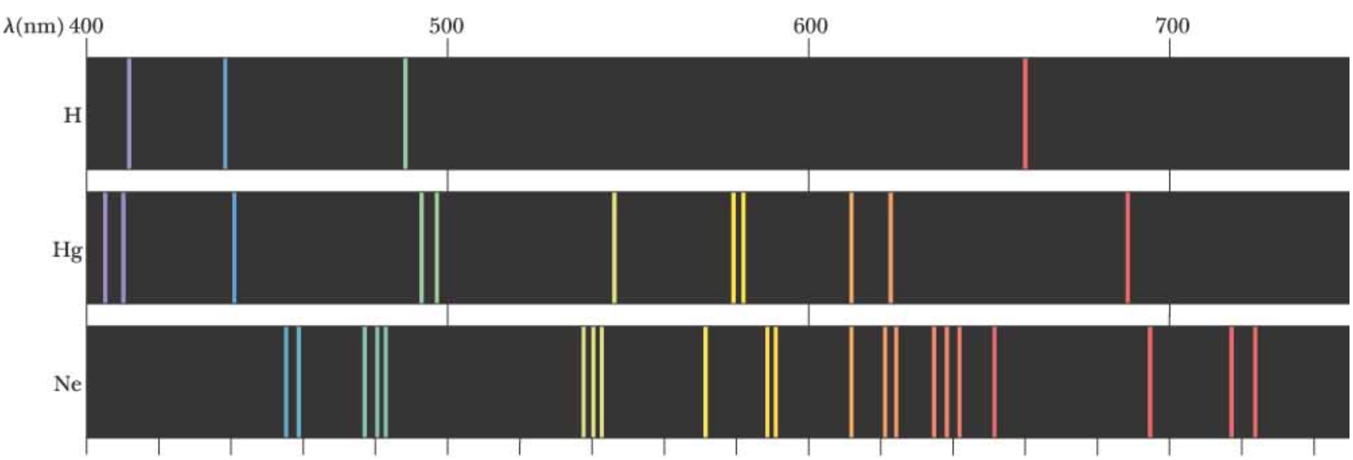
(Above) An image of white light diffracted through a spectroscope. Note the series of visible spectrum on either end of the central bright band.

Atomic Spectrum: - when atoms of an element are given energy and become excited, they emit a distinctive glow by colours. If these colours pass through a spectroscope (with diffraction grating), a series of lines will appear. These lines are unique for each element.

- Johann Balmer and Johannes Rydberg discovered separate series of the hydrogen atomic spectrum.
- sometimes the **frequencies of these spectral lines can be added** to form other spectral lines that might not be in the visible spectrum.



From left to right: Lithium, Sodium, Copper, and Strontium compounds produce different colours during the flame tests

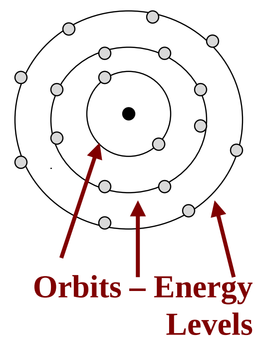


(Above) The atomic spectrums of hydrogen, mercury and neon. The scale indicates the wavelength of these lines in nanometre.

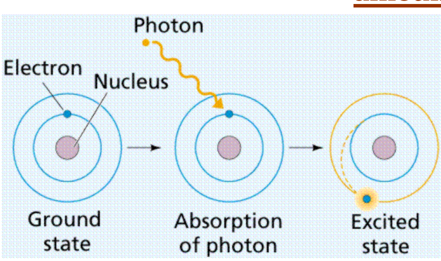
Quantum Hypothesis: - in 1900, Maxwell Planck reasoned that light energy moves in "*packets*" of energy called **quantum**.
- in 1905, Albert Einstein proved that the quantum hypothesis was correct as light can exist as a particle as well as a wave. Each light particle is called a **photon**.

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).

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**.

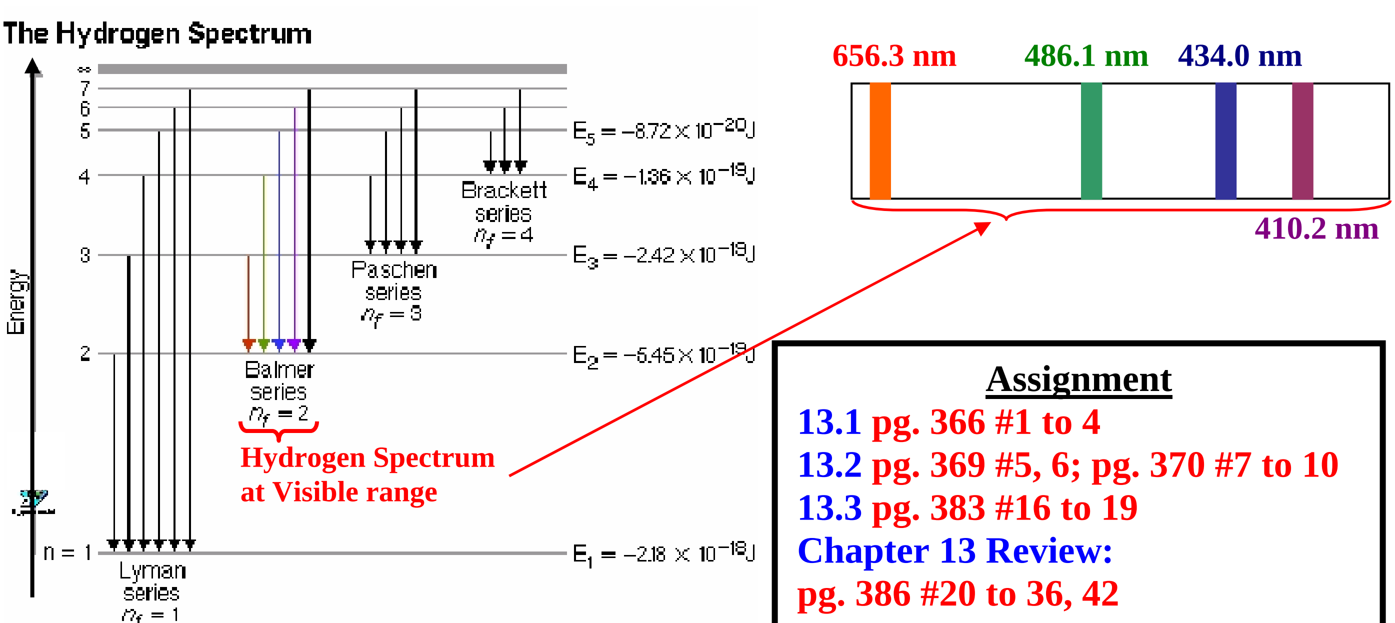


1. **Electrons cannot exist between the orbits,** much like one cannot stand between steps on a set of stairs.
2. **The further the orbit is from the nucleus, the higher its energy level for the electrons in that orbit.** This is very similar to the planetary model of our Solar system. The **quantum 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.**



Assignment

13.1 pg. 366 #1 to 4
13.2 pg. 369 #5, 6; pg. 370 #7 to 10
13.3 pg. 383 #16 to 19
Chapter 13 Review:
pg. 386 #20 to 36, 42
pg. 387 #45 to 48, 52

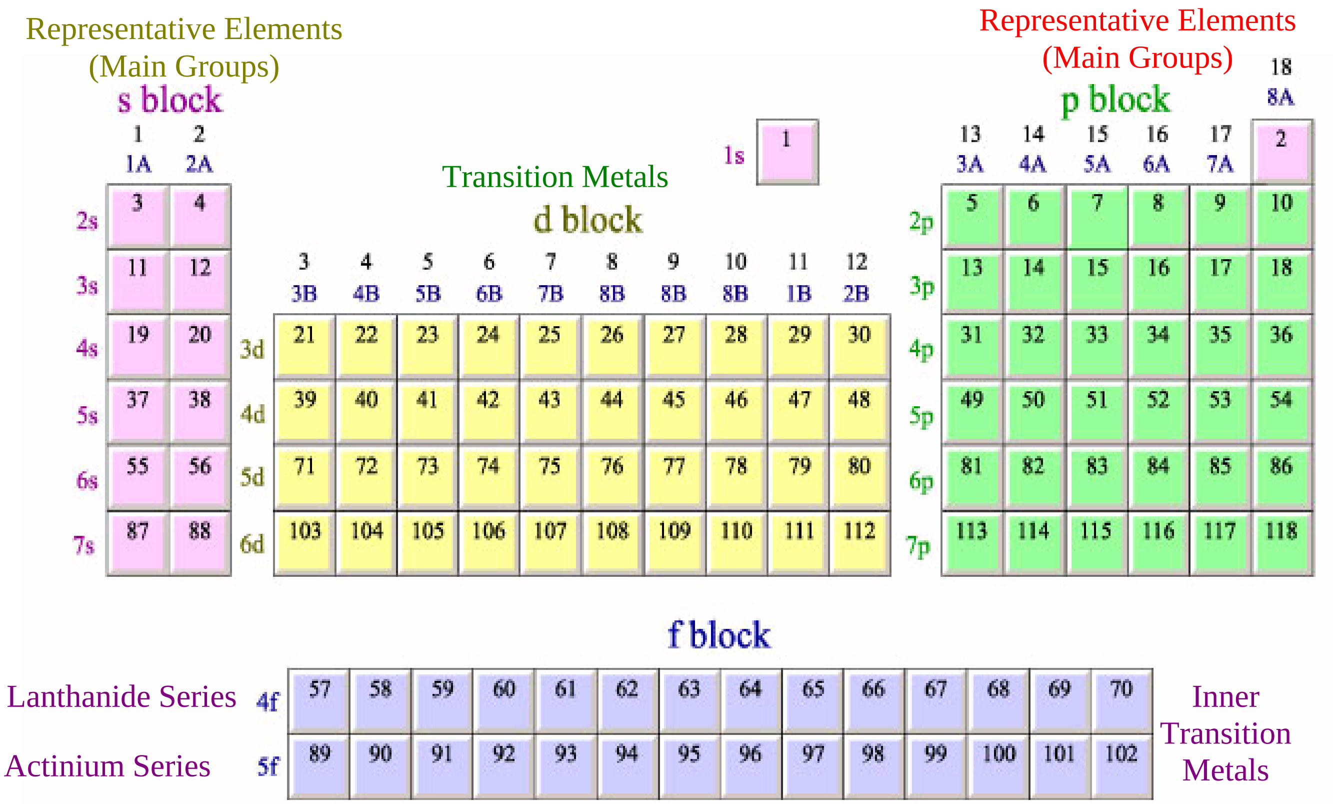
Chapter 14: Chemical Periodicity

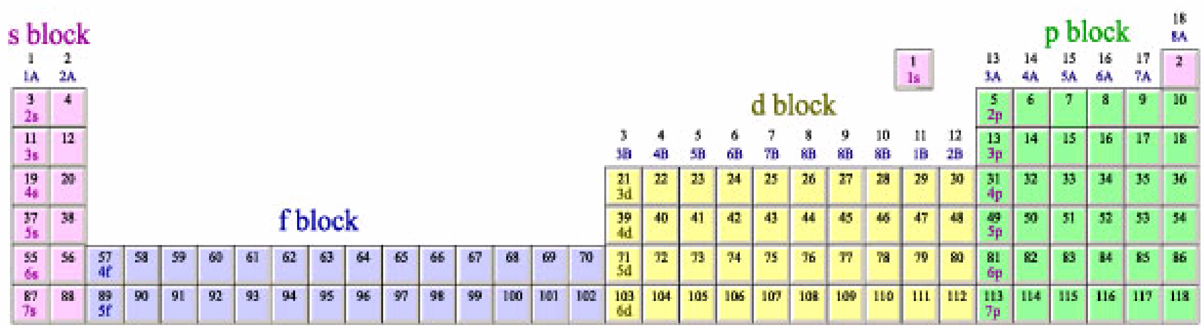
14.1: Classification of Elements

Different Groups and Periods in the Periodic Table of Elements

1. **Noble Gases**: - elements in the last column of the Table.
   - they are stable and generally non-reactive because their **outermost shell is completely filled** in the 1*s* (in the case of He) or the *p* orbitals (Ne, Ar, Kr, Xe and Rn).
2. **Representative Elements**: - elements in the *A* groups where the ***s* or *p* orbitals are partially filled**.
3. **Transition Metals**: - elements in the *B* groups characterized by the ***d* orbitals**.
4. **Inner Transition Metals**: - elements in the ***f* orbitals** in the Lanthanide and Actinide Series.

Electron Configurations in the Periodic Table of Elements





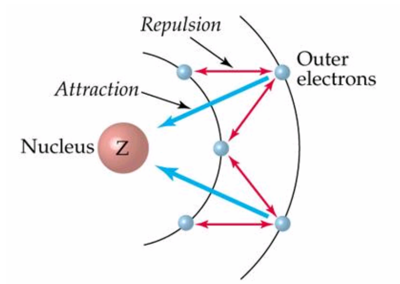
Core Electrons: - inner electrons that have completed a row in the Periodic Table of Elements.

Valance Electrons: - electrons in the outermost principal energy level of an atom.
- elements in the same group or family contains the same valence electron configuration.

14.2: Periodic Trends

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

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 (The nucleus will have more pull on the outermost electrons).

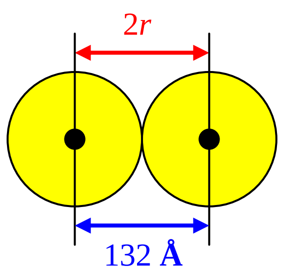
$$Z_{eff} = \text{Atomic Number } (Z) - \text{"Shield" Core Electrons}$$

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

For Na: ($1s^2\ 2s^2\ 2p^6\ 3s^1$) = 11 electrons Z_{eff} = 11 – 10 ($1s^2\ 2s^2\ 2p^6$) **Z_{eff} = 1 for Na**	For Ar: ($1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$) = 18 electrons Z_{eff} = 18 – 10 ($1s^2\ 2s^2\ 2p^6$) **Z_{eff} = 8 for Ar**

Therefore, there is less shielding for Ar and the nucleus for Ar can hold on to its valence electrons a lot better than Na.

Atomic Radius: - the size of an atom as measured by the distances between atoms in chemical compound.



Cl_2 Molecule
$2r$ = 132 Å

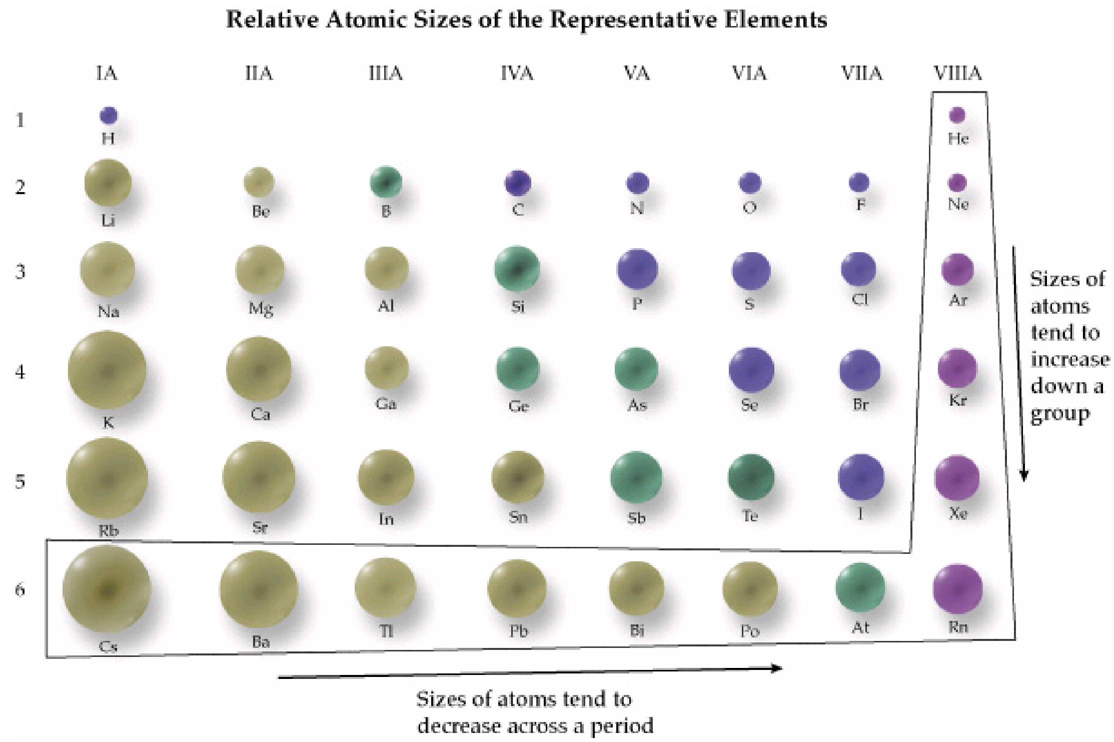
Cl Atom
r = 66 Å

1 Å = 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.

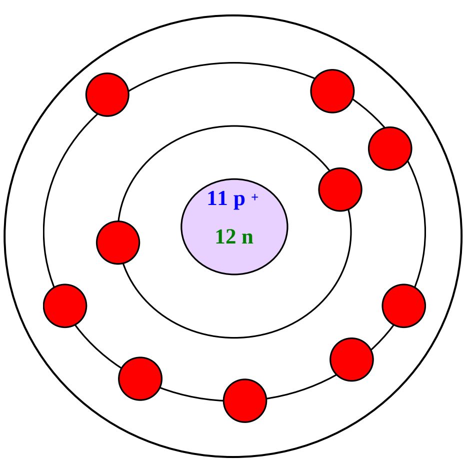


Ions: - when atoms lose or gain electrons, they attain a positive or negative charge.

1. **Cations**: - **positive** charged ions (atoms that lose electrons).
   - naming cation (element name follow by "ion")

Example 2: Draw the energy level diagrams for the following cations.

a. Sodium ion = Na^+ (11 p^+ and 10 e^-)



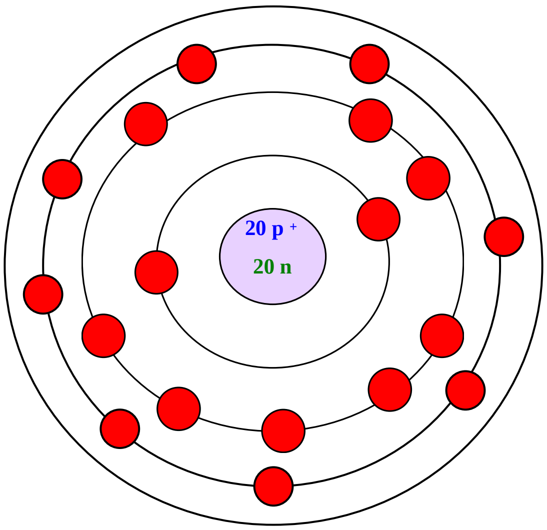
Sodium Ion (Na^+)

Atomic Number: 11 **Atomic Mass: 22.99** **Nucleus: 11 p^+ & 12 n**

2nd Energy Level: 8 e^- (8 valence e^- - Filled)
1st Energy Level: 2 e^-
Total: 10 e^- **Net Charge = 1+**

Location on the Period Table of Elements: Third Row; Column IA

b. Calcium ion = Ca^{2+} (20 p^+ and 18 e^-)



Calcium Ion (Ca^{2+})

Atomic Number: 20 **Atomic Mass: 40.08** **Nucleus: 20 p^+ & 20 n**

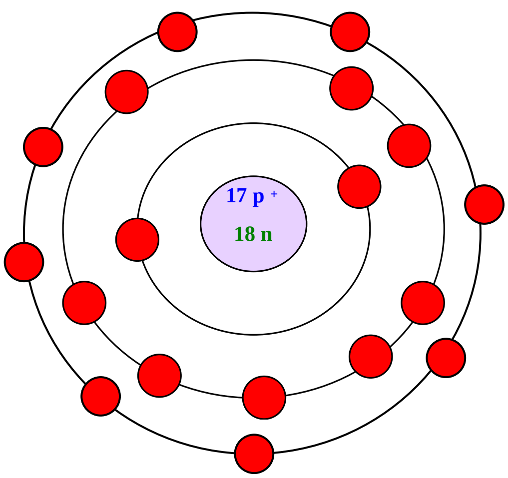
3rd Energy Level: 8 e^- (8 valence e^- - Filled)
2nd Energy Level: 8 e^-
1st Energy Level: 2 e^-
Total: 18 e^- **Net Charge = 2+**

Location on the Period Table of Elements: Fourth Row; Column IIA

2. **Anions**: - **negative** charged ions (atoms that gain electrons).
   - naming anion (first part of element name follow by suffix ~*ide*)

Example 3: Draw the energy level diagrams for the following anions.

a. Chloride = Cl^- (17 p^+ and 18 e^-)



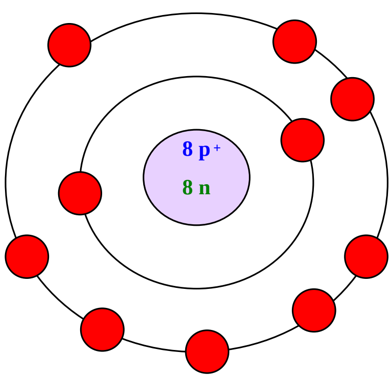
Chloride (Cl^-)

Atomic Number: 17 **Atomic Mass: 35.45** **Nucleus: 17 p^+ & 18 n**

3rd Energy Level: 8 e^- (8 valence e^- - Filled)
2nd Energy Level: 8 e^-
1st Energy Level: 2 e^-
Total: 18 e^- **Net Charge = 1–**

Location on the Period Table of Elements: Third Row; Column VIIA

b. Oxide = O^{2-} (8 p^+ and 10 e^-)



Oxide (O^{2-})

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

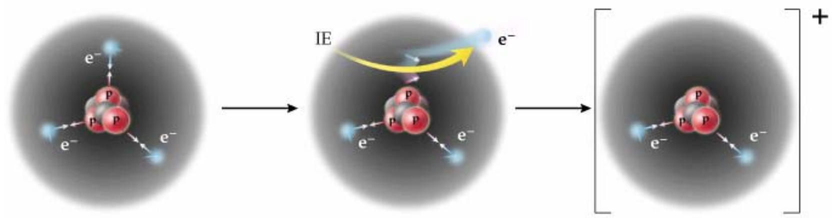
2nd Energy Level: 8 e^- (8 valence e^- - Filled)
1st Energy Level: 2 e^-
Total: 10 e^- **Net Charge = 2–**

Location on the Period Table of Elements: Second Row; Column VIA

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 the 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.

$X^+_{(g)} \rightarrow X^{2+}_{(g)} + e^-$ $Na^+_{(g)} \rightarrow Na^{2+}_{(g)} + e^-$ $I_2 = 4560$ 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

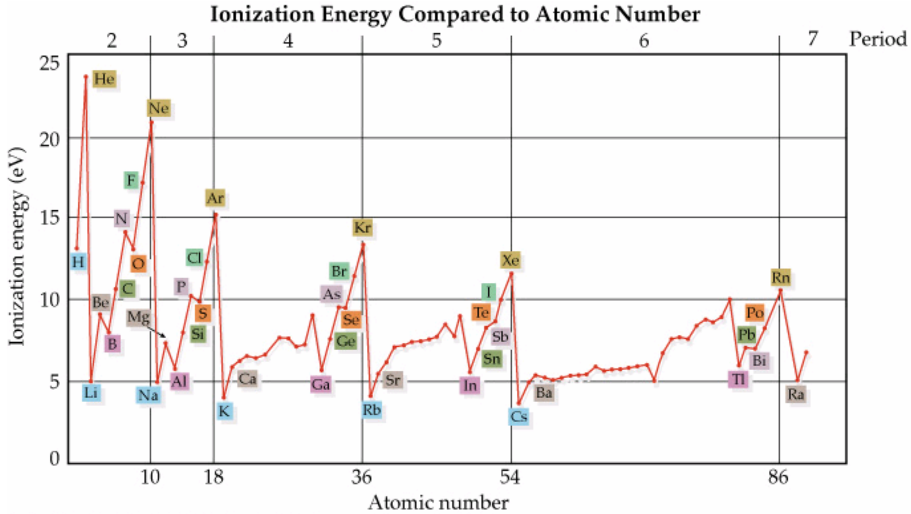
Successive Ionization Energies INCREASES within each element. ($I_1 < I_2 < I_3 < \ldots$)

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.



Ions Sizes: - the size of cations and anions.

Several Notes on Trends in Ion Size

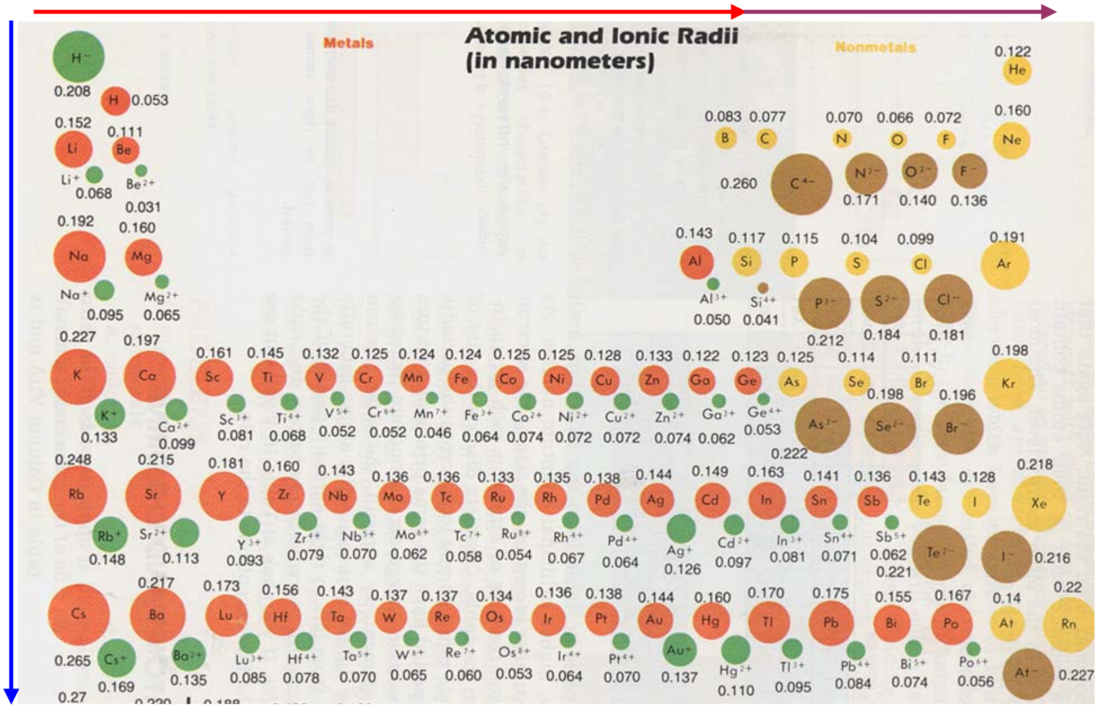
1. **Metal Cations are generally Smaller than Non-Metal Anions WITHIN the Same Period**. Metal cations tend to lose electrons to achieve the same number of valence electrons as the previous noble gas. Thereby, losing an energy level in the process. **Metal Cations are always Smaller than the Parent Neutral Atoms. Non-Metal Anions are always Larger than the Parent Neutral Atoms.**

2. In general, **Ion Sizes Decrease as one move from LEFT to RIGHT of a period WITHIN the METAL GROUPS and WITHIN the NON-METAL GROUPS**. 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.

3. **Ionic 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

Cations Decreases from Left to Right **Anions Decreases from Left to Right**

Ions Increases Down the column



Example 4: Order the following ions from the smallest to the largest.

a. Te^{2-}, S^{2-}, Se^{2-}, O^{2-}

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

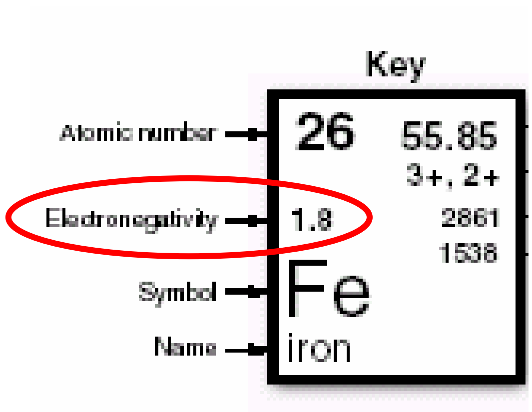
$O^{2-} < S^{2-} < Se^{2-} < Te^{2-}$

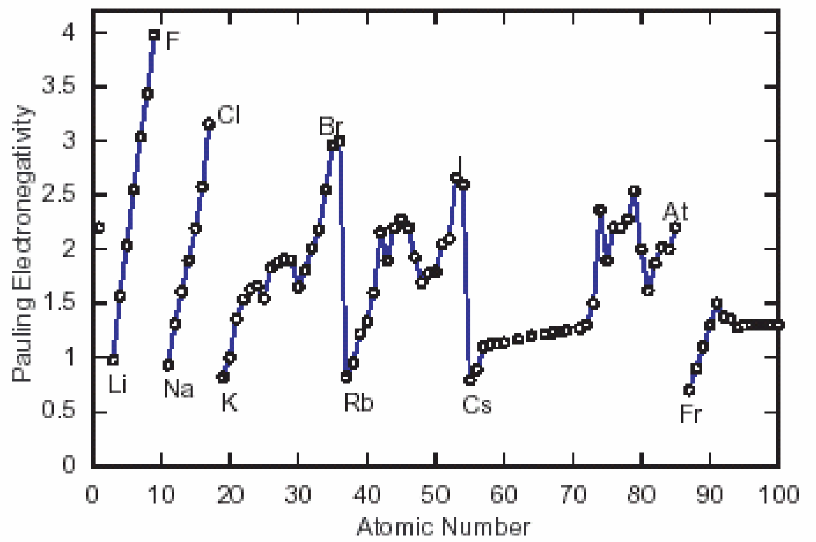
b. Na^{+}, Mg^{2+}, P^{3-}, S^{2-}, Cl^{-}

These ions are within the same Period (row). As we move to the right, ion size <u>decreases within each of the metal and non-metal groups</u>. Therefore,

$Mg^{2+} < Na^{+} < Cl^{-} < S^{2-} < P^{3-}$

<u>Electronegativity</u>: - first determined by **Linus Pauling**, it is a measure of the **capability of an atom within a molecule to attract shared electrons around it**.
- 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.





<u>Several Notes on Trends in Electronegativity</u>

1. In general, **<u>Electronegativities INCREASE as one move to the right of a period (up to and including halogens)</u>**. This is because non-metals tend to form anions to fill the valence orbitals, and therefore, they can better attract electrons. Metals, which have high ionization energies, like to give away electrons to form cations. Hence, they do not like to attract electrons and a lower electronegativity is the result.

2. **<u>Electronegativities DECRESE Down a Group</u>**. 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.

<u>Assignment</u>

14.1 pg. 396 #1 to 5

14.2 pg. 406 #6 to 9

Chapter 14 Review: pg. 409–410 #10 to 32