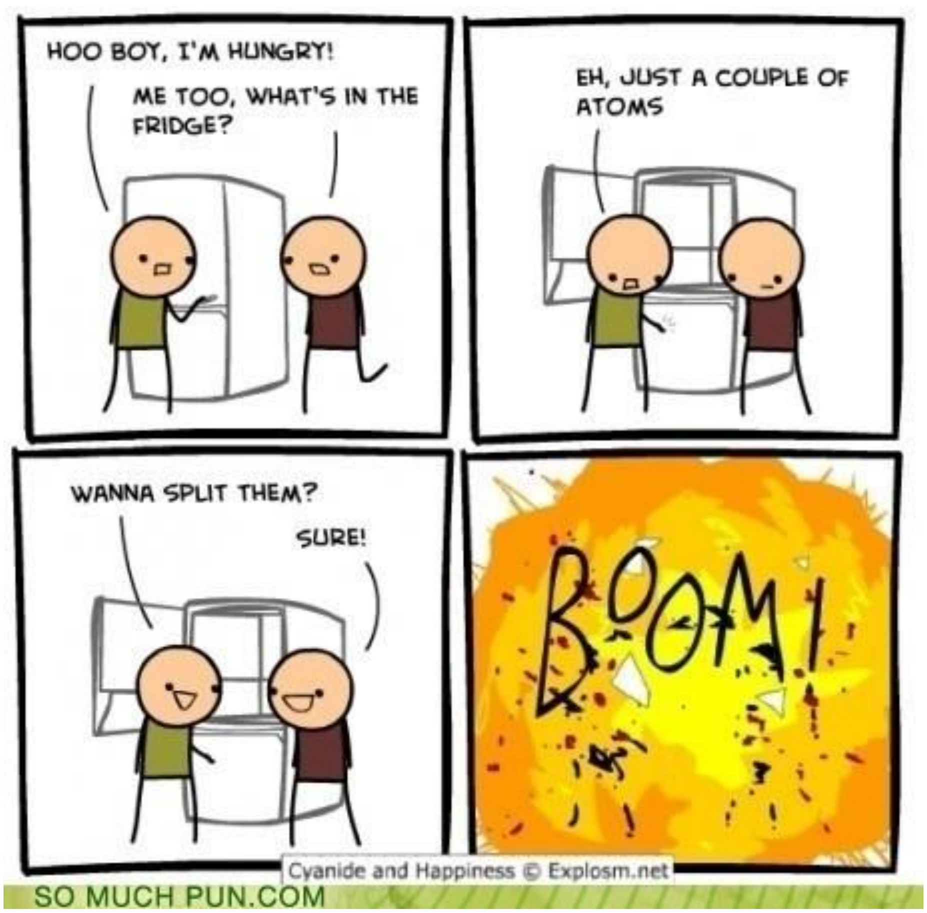
Atomic Structure and Nuclear Chemistry Notes Outline



These notes belong to:

____

Video 1: Atomic History and Theory

Democritus (460-370 BCE)

- He and his mentor, Leucippus, are credited with the first atomic model.
- His theory included:
  - Matter is made of **tiny, solid, indivisible particles** which they called atoms.
  - Different kinds of atoms have different **sizes** and **shapes**, therefore different **properties**.
  - Widely rejected by his peers, most notably Aristotle.

Aristotle (384-322 BCE)

- Believed that everything was composed of the fours elements: earth, air, fire, and water.
- Because Democritus' ideas could not be tested experimentally at the time Aristotle's ideas were accepted to be true**…**Until**…**

John Dalton (early 19th century)

- Revived Democritus' ideas and developed the first modern atomic theory.
- His theory is summarized by four main points:
  1. All matter is composed of extremely small particles called **atoms**.
  2. **All atoms of one element are identical** but different from those of any other element.
  3. Atoms of one element combine with atoms of another element to **form compounds during chemical reactions** and they combine **in simple whole number ratios**.
  4. **Atoms are indivisible**; they cannot be created or destroyed, only rearranged during reactions.

Draw a picture of Dalton's Model in the space provided:

Dalton's Theory had 2 Flaws

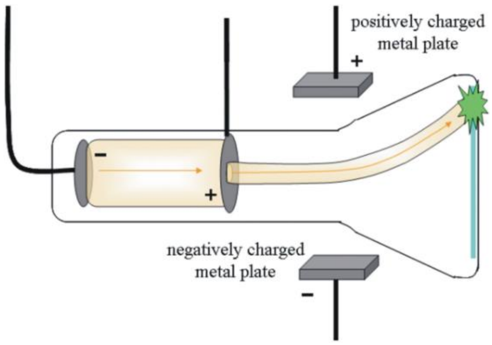
- He said all atoms of one element are identical – FALSE
  - ______________________ are atoms of the same element which have different numbers of ______________________, therefore they have different ____________________________.
- He said atoms are indivisible - FALSE
  - Atoms are divisible into **protons, neutrons**, and **electrons**.

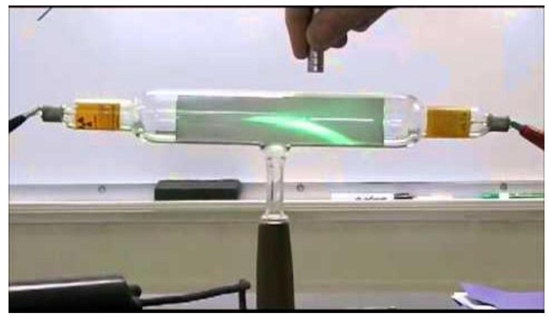
Scientists then quantitatively experimented with chemical reactions and developed **3 basic laws:**

1. **Law of Conservation of Mass –** matter cannot be created or destroyed during a chemical reaction, only rearranged.
2. **Law of Definite Proportions** – a chemical compounds always contains the same proportions of elements by mass**.**
3. **Law of Multiple Proportions** – elements combine with each other in ratios of small whole numbers.

J.J. Thomson (1856–1940)

- Discovered the ________________ and determined it had a ________________ charge by experimentation with **cathode ray tubes.**
    - Thomson placed oppositely charged electric plates around the cathode ray. The cathode ray was deflected towards the positive plate. This indicates that the cathode ray is composed of negatively-charged particles which became known as electrons.



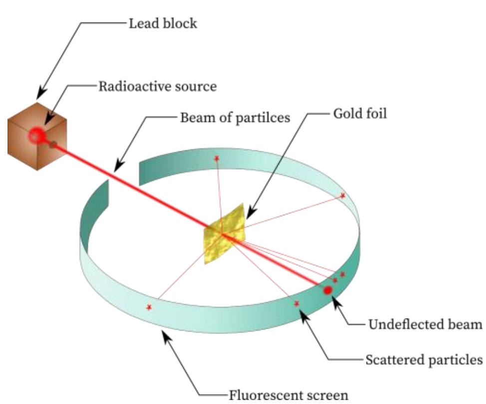


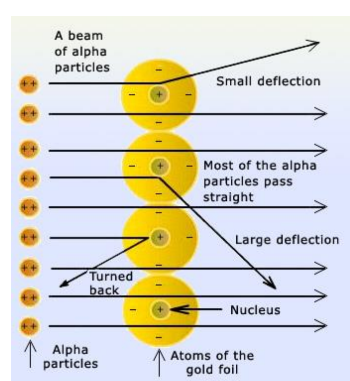
- Thomson repeated this experiment with numerous metals and gases and obtained the same results. He concluded that **electrons are present in atoms of ALL elements**.
- Thomson developed the ____ which shows evenly distributed negative electrons in a uniform positively charged matrix of material.

Draw a picture of Thomson's Model in the space provided:

Ernest Rutherford (early 20th century)

- Discovered the _______________ of the atom in his _____________________________________.
    - Alpha particles (positively charged nuclei) produced from the radioactive decay of radium salts were focused by a lead box and streamed toward a sheet of extremely thin gold foil (~1000 atoms thick). Rutherford expected all the alpha particles to pass through the atoms of gold in the foil, undeflected, but a small percentage of the alpha particles bounced off of the gold foil at great angles.





- Rutherford concluded that:
    - There is a **massive, dense, positively-charged structure** that was very small – this became known as the ______________________.
    - Most alpha particles saw no influence to their path, so atoms must be mostly ______________ ________________.
- He then discovered the **proton** in further studies using alpha particles, and his student, James Chadwick, discovered the **neutron.** Both of these are located in the **nucleus.**

Draw a picture of Rutherford's Model in the space provided:

Niels Bohr (early 20th century)

- Developed the **Bohr Model**.
- His theory stated that **electrons are restricted to specific energies and follow paths called orbits at a fixed distance from the nucleus.** This is similar to the way the planets orbit the sun. *We now know electrons **do not** have neat orbits like the planets.*

Draw a picture of Bohr's Model in the space provided:

Quantum Mechanical Model (current model)

- Mathematical model developed in the 1920s by many scientists, including de Broglie, Schrodinger, Pauli, and Heisenberg. Also known as the **"Electron Cloud"** model.
- Electrons **exist in regions of space around the nucleus which are called orbitals**. The paths of electrons are random and cannot be predicted. We can only talk about the probability of an electron being in a certain region (likely to be near the nucleus).

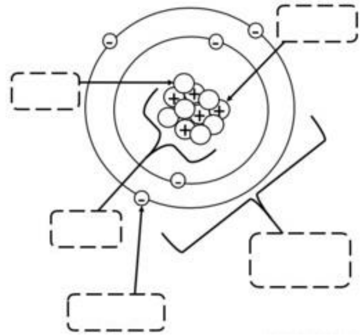
Draw a picture of the Quantum Mechanical Model in the space provided:

Scientist Whiteboard Activity

Use the space below to complete each card’s activity.

Video 2: Reviewing Atoms, Isotopes, and Ions

Review: Parts of the Atom



An **atom** is the smallest unit of an element that maintains the chemical properties of that element. *During chemical reactions, the atom* *<u>does not</u>* *split!*

There are 2 regions of an atom that contain particles of matter: the **nucleus** – contains protons and neutron, and the **electron cloud** – contains electrons. The rest is empty space.

Subatomic Particles Summary Chart				
Particles	**Symbol**	**Charge**	**Mass (g)**	**Location**
Electron	e^-		$9.11\text{x}10^{-28}$ g	
Proton	p^+		$1.673\text{x}10^{-24}$ g	
Neutron	n^0		$1.675\text{x}10^{-24}$ g	

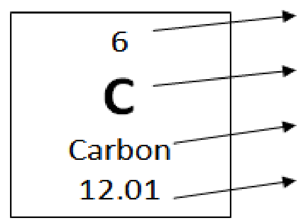
Protons and electrons have equal but opposite charges. Because the **magnitude** (size) of the charge is identical, when we assign charges to these particles we use whole numbers:

1 electron = -1 charge — 1 proton = +1 charge
2 electrons = -2 charge — 2 protons = +2 charge

The nucleus contains almost all of the atom's **mass,** but the electron cloud is responsible for almost all of the **volume.**

Review: Calculating Protons, Electrons, and Neutrons

The periodic table is organized in order of increasing atomic number. The **atomic number** is the whole number that is unique for each element on the periodic table which **defines the element.** The atomic number represents the number of protons in one atom of an element.



In a **NEUTRAL atom,** # of **protons** (P) **= # of electrons**
Therefore, **atomic number** (A) = # of **electrons** (E) in a **neutral atom**

Remember: "Ape" or **A = P = E**
****In an <u>ion</u>, electrons* ***do not*** *= protons****

The **atomic mass number** is the sum of the **protons** and **neutrons** for a particular atom or isotope. It is the **mass** of the atom *(electron mass is negligible).*

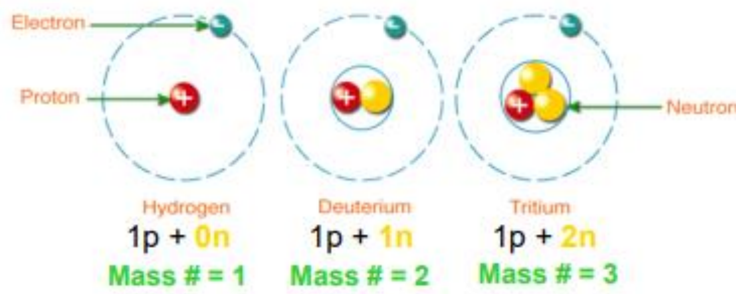
Mass number (M) = # of **protons** + # of **neutrons** (N)
# of protons = atomic number (A)

Remember: "Man" or **M = A + N**

NEW: Isotopes

<u>Isotopes</u> are two or more forms of the same element that differ only in the number of neutrons. Chemically they react identically.

Example: Hydrogen has 3 isotopes.
Because isotopes of the same element have **different numbers of neutrons**, they will also have a **different mass.**

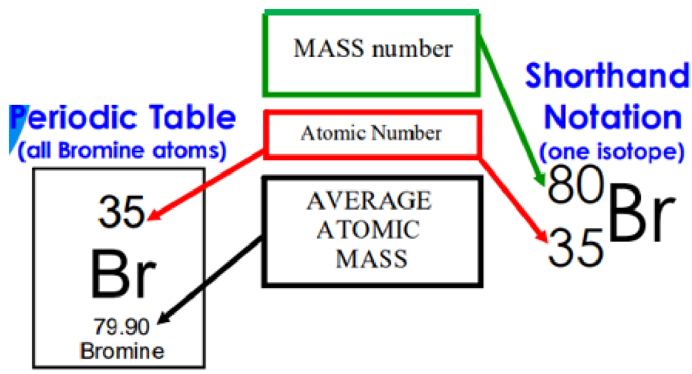


Types of Notation

<u>Isotope notation</u> allows us to express the identity or isotope quickly and easily. When shorthand notation is used, it will appear one of four ways.

Example: The Bromine atom with a mass number of 80 can be written using any of these formats –

__________________ __________________ __________________ __________________



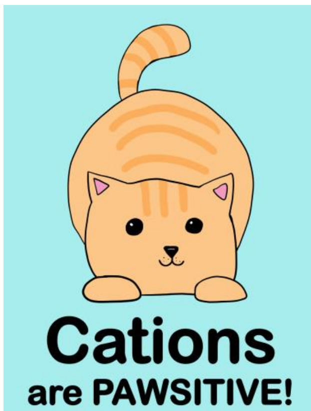
Notation Practice

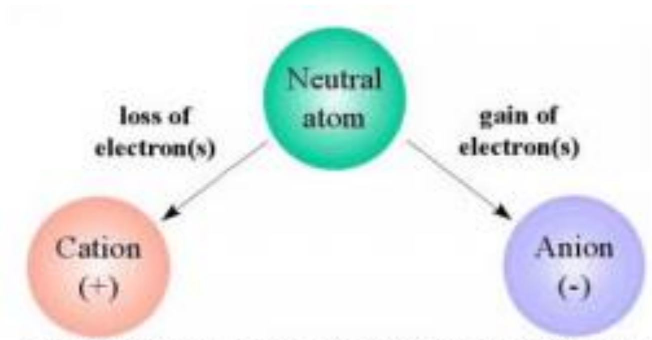
How many protons, electrons, and neutrons are in each isotope of carbon below?

Isotopes	Protons	Electrons	Neutrons
Carbon-12			
Carbon-13			
Carbon-14			

Ions

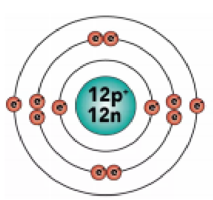
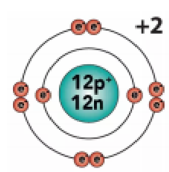
- Atoms form ions when ____________________ **are lost or gained**.
- The number of **protons is NOT EQUAL to the number of electrons.**
- Ions carry an electrostatic charge (positive or negative). An atom that has **gained** electrons is an ________________. An atom that has **lost** electrons is a _______________.
- In a **neutral** atom the number of **protons =** number of **electrons.**
- In an **ion** the number of **protons ≠** number of **electrons**.

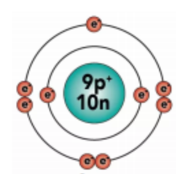
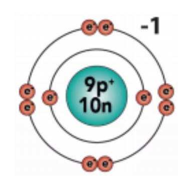




A neutral atom becomes an ion either by losing an electron (cation) or by gaining an electron (anion).

Ions Practice

Neutral Mg	Cation Mg^{2+}
	
# e^- =	# e^- =

Neutral F	Anion F^-
	
# e^- =	# e^- =

Questions

1. Write all the isotope notations for an isotope of the element sodium with a mass number of 23.

2. How many protons, neutrons, and electrons are in Al^{3+} and N^{3-}?

3. Why is the atomic mass number on the periodic table generally not a whole number?

Video 3: Average Atomic Mass

Average Atomic Mass – represents the average mass of ALL isotopes of an element.

- Found using a weighted average technique considering abundance and mass of each isotope.
- Measured in **atomic mass units** (amu).
- 1 amu = 1.660538×10^{-24} g (it is 1/12 the mass of Carbon-12)

Calculating Relative Abundance

Round all average atomic masses to the one-hundredths place (2 places after the decimal)

Example: Lithium exists as 2 isotopes with the following abundances and masses –

Isotope	Mass (amu)	% Abundance	Relative Abundance
Lithium-6	6.015	7.50%	
Lithium-7	7.016	92.50%	

Rules for Calculating Average Atomic Mass

1. Calculate the percent abundance of each isotope *(recall that percent = part / whole)*. This value may be given to you in the question.
2. Convert percent abundance to relative abundance (% → decimal)
3. Multiply each isotope's mass by its relative abundance
4. Add these values together **(Round to 2 places after the decimal!)**

Example:

Neon has 3 isotopes: Neon-20 has a mass of 19.992 amu and an abundance of 90.51%. Neon-21 has a mass of 20.994 amu and an abundance of 0.27%. Neon-22 has a mass of 21.991 amu and an abundance of 9.22%. What is the average atomic mass of neon?

Isotope	Mass (amu)	% Abundance	Relative Abundance
Neon-20	19.992	90.51	
Neon-21	20.994	0.27	
Neon-22	21.991	9.22	

Questions

1. Fictitious element X has 2 isotopes. Isotope 1 has an atomic mass of 21.022 amu and isotope 2 has an atomic mass of 23.034 amu. The average atomic mass of element X is 21.95 amu. What is the percentage of each isotope in element X?

Video 4: Nuclear History and Nuclide Stability

Nuclear chemists study the structure of atomic nuclei and the changes they undergo.

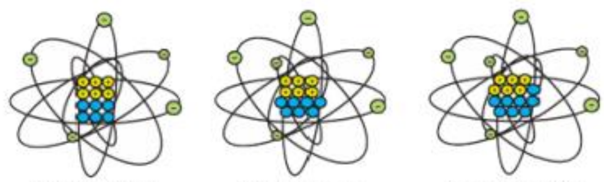
Wilhelm Röntgen - found that invisible rays were emitted when electrons bombarded the surface of certain matierals.

Henri Becquerel - accidentally discovered that phosphorescent uranium salts produced spontaneous emissions that darkened photographic plates.

Marie Curie - isolated the components emitting the invisible rays (uranium); defined radioactivity and radiation; discovered Polonium, Radium, and Curium; died of aplastic anaemia resulting from large amounts of radiation exposure.

Nuclear Vocabulary

Radioactivity – the process by which particles give off rays.
Radiation – the penetrating rays and particles emitted by a radioactive source.
Radioisotopes – isotopes of atoms with unstable nuclei (too many or too few neutrons). Many of them are radioactive and undergo radioactive decay.
Radioactive Decay – a spontaneous process where unstable nuclei naturally break down into smaller elements and emit particles and/or energy (radiation) to attain more stable atomic configurations.
Nuclide – an isotope; identified by the number of protons and neutrons in the atom's nucleus.



C-12 Stable C-13 Stable C-14 Unstable

Transuranium Elements – elements with atomic number 92 or higher.

- Very few are found in trace amounts in nature. Most were synthesized in nuclear reactors and particle accelerators.
- They are very unstable and most exist for infinitely small moments and decay into smaller, more stable atoms.

The Periodic Table of the Elements

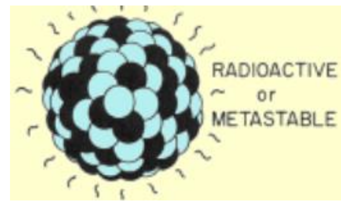
	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H 1.01																	2 He 4.00
2	3 Li 6.94	4 Be 9.01											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
3	11 Na 22.99	12 Mg 24.30											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
4	19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.69	33 As 74.92	34 Se 78.97	35 Br 79.90	36 Kr 83.80
5	37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.95	43 Tc (98)	44 Ru 101.10	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
6	55 Cs 132.91	56 Ba 137.33	71 Lu 174.97	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.20	77 Ir 192.20	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.20	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
7	87 Fr (223)	88 Ra (226)	103 Lr (262)	104 Rf (267)	105 Db (268)	106 Sg (271)	107 Bh (272)	108 Hs (270)	109 Mt (276)	110 Ds (281)	111 Rg (280)	112 Cn (285)	113 Nh (284)	114 Fl (289)	115 Mc (288)	116 Lv (293)	117 Ts (294)	118 Og (294)

57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.40	63 Eu 151.97	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.05
89 Ac (227)	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)

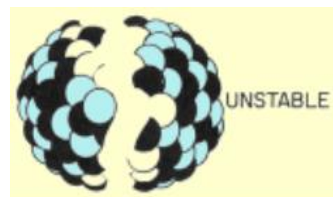
Nuclear Stability of Elements

An isotope of an elements is considered stable if **the nucleus will NOT spontaneously decay.**

- Elements with atomic numbers 1-20 have at least 1 nuclide that is very stable.
  - Nucleus has **proton : neutron ratio of 1:1**
  - Example:



- Elements with atomic numbers 21-83 have at least 1 nuclide that is somewhat stable (metastable).
  - Nucleus has **proton : neutron ratio of 2:3**
  - Example:



- Elements with atomic numbers greater than or equal to 84 do not have a stable isotope and are **unstable AND radioactive**.
  - Nucleus has **proton : neutron ratio of 1: >2**
  - Example:

Transmutation

Transmutation is the conversion of an atom of one element into an atom of another element.

- Occurs in spontaneous radioactive decay
- Can be induced when **high energy particles are ________________________ into the nucleus of an atom.**
- An unstable nucleus decays when **its nuclear configuration is _____________________ - has unfavorable proton : neutron ratio or contains too much energy.**
- The decaying nucleus **emits energy as particles and rays** and transmutates or decays into **a more stable isotope or a different element.**

Decay Chains (Series) -the decay of a radioisotope into other radioisotopes eventually resulting in a stable isotope using a series of alpha, beta, and gamma emissions.

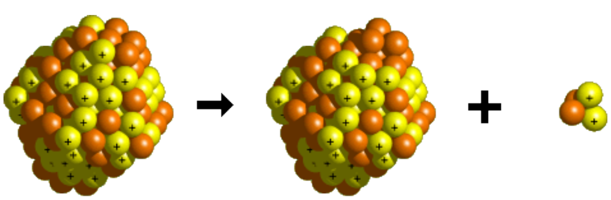
Parent nuclide – the heaviest initial nuclide in the series.

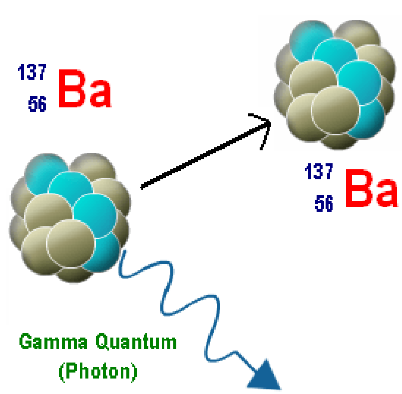
Daughter nuclide – the nuclides produced throughout the decay.

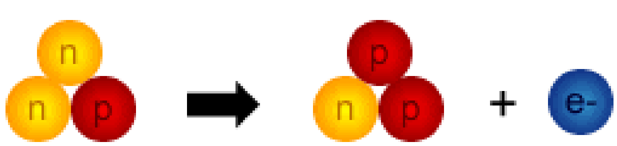
Types of Radioactive Decay

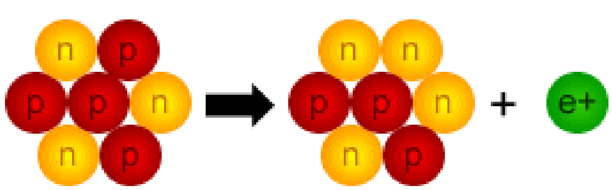
Name	Symbol	Particle	Description	Example
Alpha		Helium Nucleus ______	**The nucleus ejects an alpha particle as a product. Only occurs in large nuclides.**	$^{238}_{92}U \rightarrow$ _____ + _____ Mass drops by 4, atomic number drops by 2.
Beta-minus		High Speed Electron ______	**A neutron in the original atom decays to form a proton and high-speed electron.** $^{1}_{0}n \rightarrow ^{1}_{1}p + ^{0}_{-1}e$ **Occurs in neutron rich nuclides (#n^0 > #p^+)**	$^{210}_{82}Pb \rightarrow$ _____ + _____ The neutron is represented in the original element; the proton is represented in the new element. Atomic number increases by 1, mass is unchanged.
Beta-plus		High Speed Positron ______	**A proton in the original atom decays to form a neutron and a high-speed positron.** $^{1}_{1}p \rightarrow ^{1}_{0}n + ^{0}_{+1}e$ **Occurs in neutron poor nuclides (#n^0 < #p^+)**	$^{12}_{7}N \rightarrow$ _____ + _____ The proton is represented in the original element; the neutron is represented in the new element. The atomic number drops by 1, mass is unchanged.
Gamma		Gamma Radiation (Energy) ______	**High energy radiation in which no particles are emitted, just energy. The resulting nuclide is stable.**	$^{137}_{56}Ba \rightarrow$ _____ + _____ Gamma radiation does NOT have mass. Only excess energy held in the nucleus is shed.

Label the images to indicate what type of radiation they are modeling.



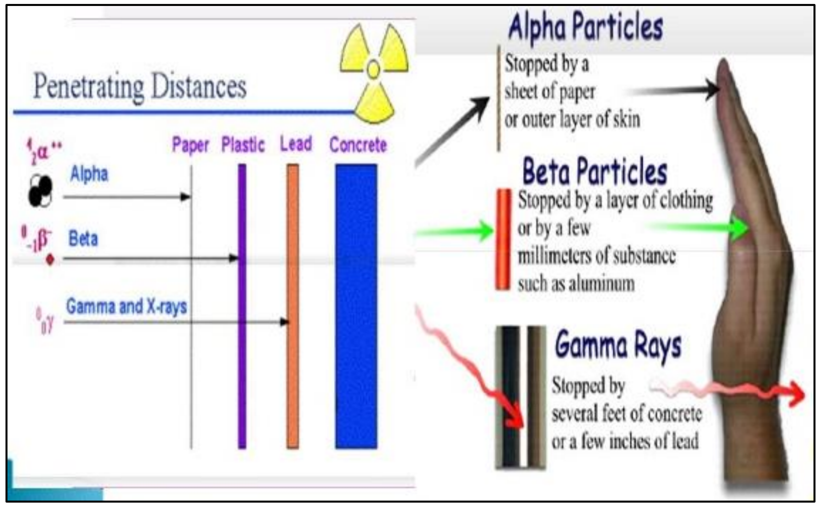






Comparing Radiation

Alpha



- More massive (4 amu)
- Cause the **most damage over a short range.**
- Least able to penetrate surfaces: **stopped by a piece of paper.**
- Charge = **+2**

Beta + or –

- Less massive (1/1837 amu)
- Causes **less damage over a short range.**
- More able to penetrate surfaces: **stopped by thin aluminum foil, wood, or glass.**
- Charge = **-1 (β^-) or +1 (β^+)**

Gamma

- No mass; all energy (not a particle!)
- Causes **least damage over short range.**
- Most able to penetrate surfaces: **stopped by 1in thick lead or foot of concrete.**
- Charge = **0**

Questions

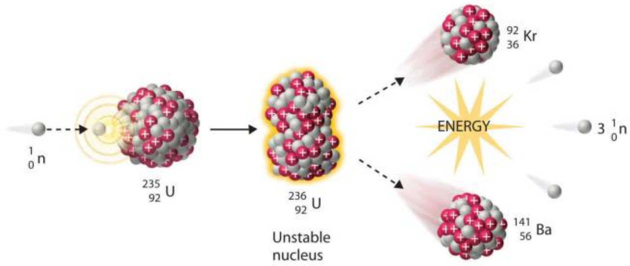
1. Write the equation and explain what happens to the nuclei in terms of mass, protons and neutrons for each decay:
   a. Alpha decay of Plutonium-248

   b. Beta-minus decay of Vanadium-52

   c. Beta-plus decay of Copper-59

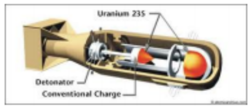
Video 5: Fusion and Fission

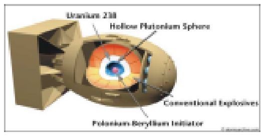
Nuclear Fission is when a large nuclide such as Uranium or Plutonium, is bombarded with neutrons resulting in the large nuclide splitting into several smaller daughter nuclide. It also releases more neutrons and a large quantity of energy.
Example:



Nuclear reactors use controlled fission to produce energy. Nuclear weapons use uncontrolled nuclear fission reactions with 90-95% enriched U-235 or Pu-239.
Examples of fission in nuclear weapons:

- "Little Boy" Bomb used U-235; dropped on Hiroshima on Aug. 6, 1945 in WWII
- "Fat Man" Bomb used Pu-239; dropped on Nagasaki on Aug. 9, 1945 in WWII





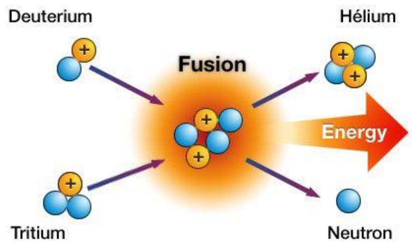
Examples of nuclear reactor disasters:

- Three Mile Island, Pennsylvania – 1979
- Chernobyl, Ukraine – 1986
- Fukushima Daiichi, Japan – 2011

Nuclear Fusion forcing 2 relatively small nuclei to combine (fuse) into 1 nucleus.

- This can only happen at extremely high temperatures.

Example:



Nuclear fusion powers the stars, including our sun. Scientists have not found a way to sustain a beneficial nuclear fusion reaction on earth.

Modern thermonuclear weapons, such as the Hydrogen Bomb, use a combination of fusion and fission reactions.

Questions

1. What types of nuclides undergo fission?

2. What types of nuclides undergo fusion?

3. List 2 possible products of the spontaneous fission of curium (Cm-250).

Use the space below for any questions or additional notes you may have.