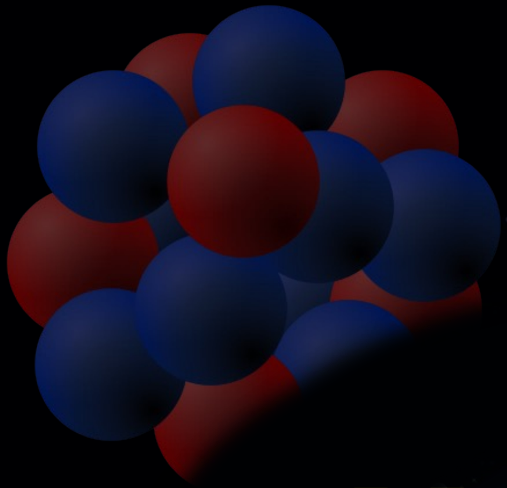

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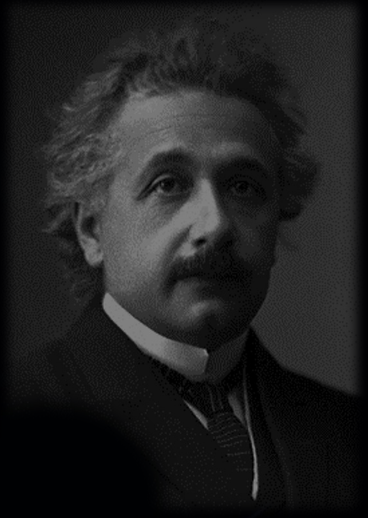
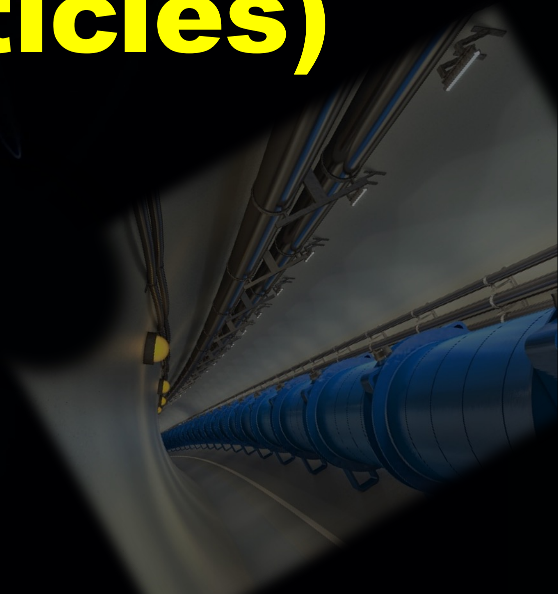


Albert Einstein

Nuclear Chemistry

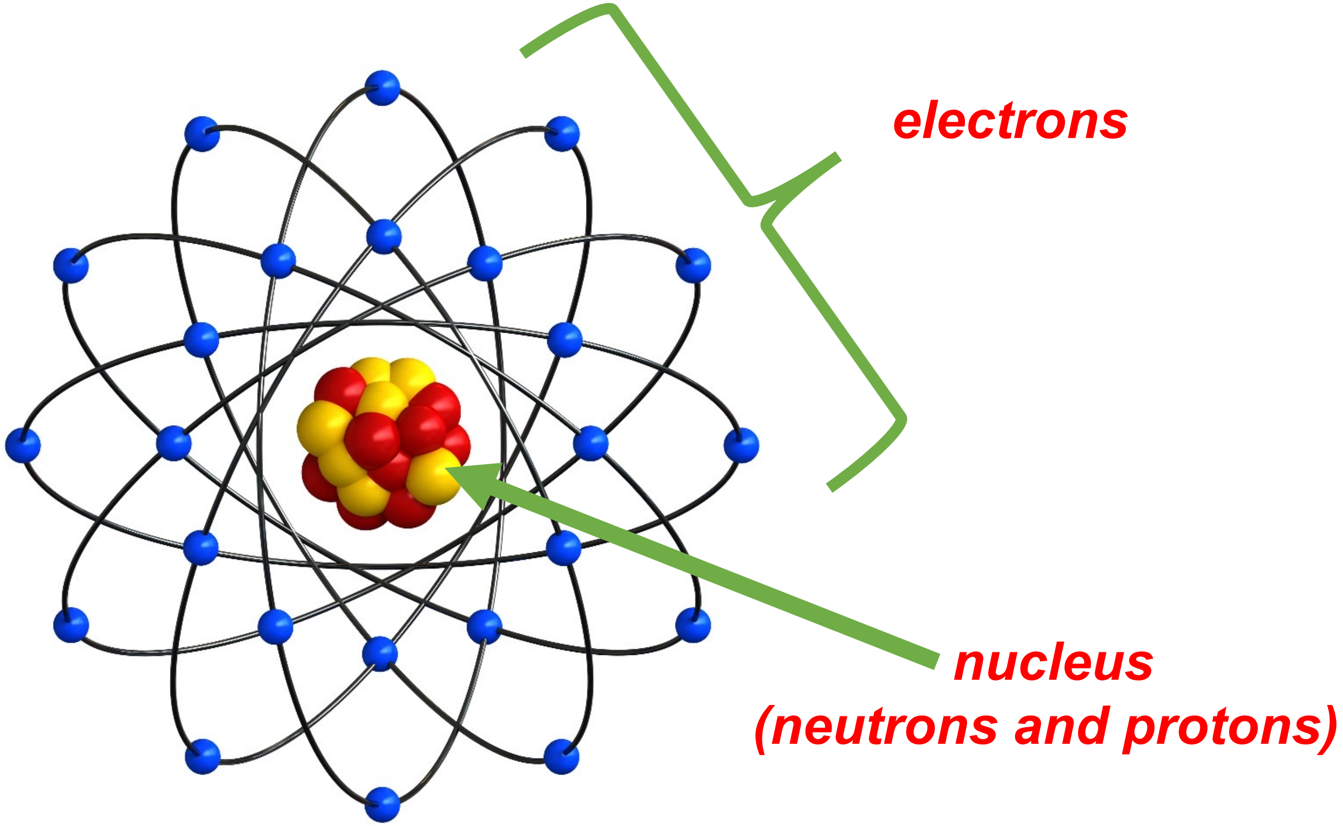
(The Nucleus and Nuclear Particles)



A Nuclear Intro

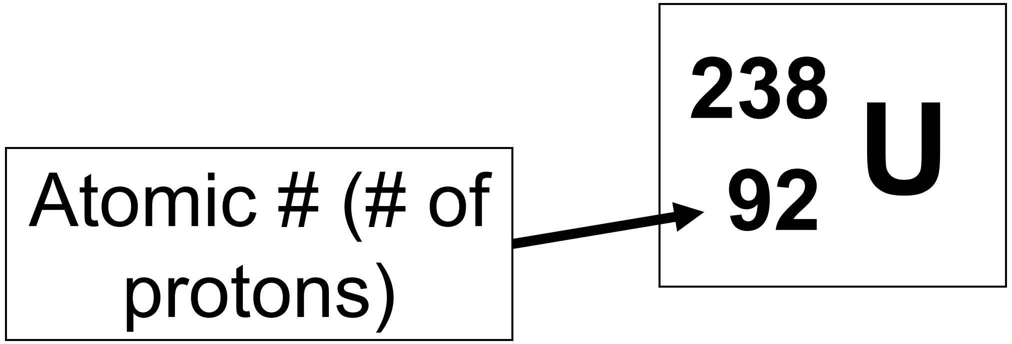
The vast majority of chemical reactions that we typically discuss involve atoms sharing or transferring their electrons, which are located in their outer orbitals. In contrast, ***nuclear reactions*** involve transfers or transformations of *neutrons or protons* from the atom's nucleus. In this video, we will discuss ***nuclear reactions***.

Before getting into that, however, we have to first review atomic structure.



Atomic (Chemical) Symbols

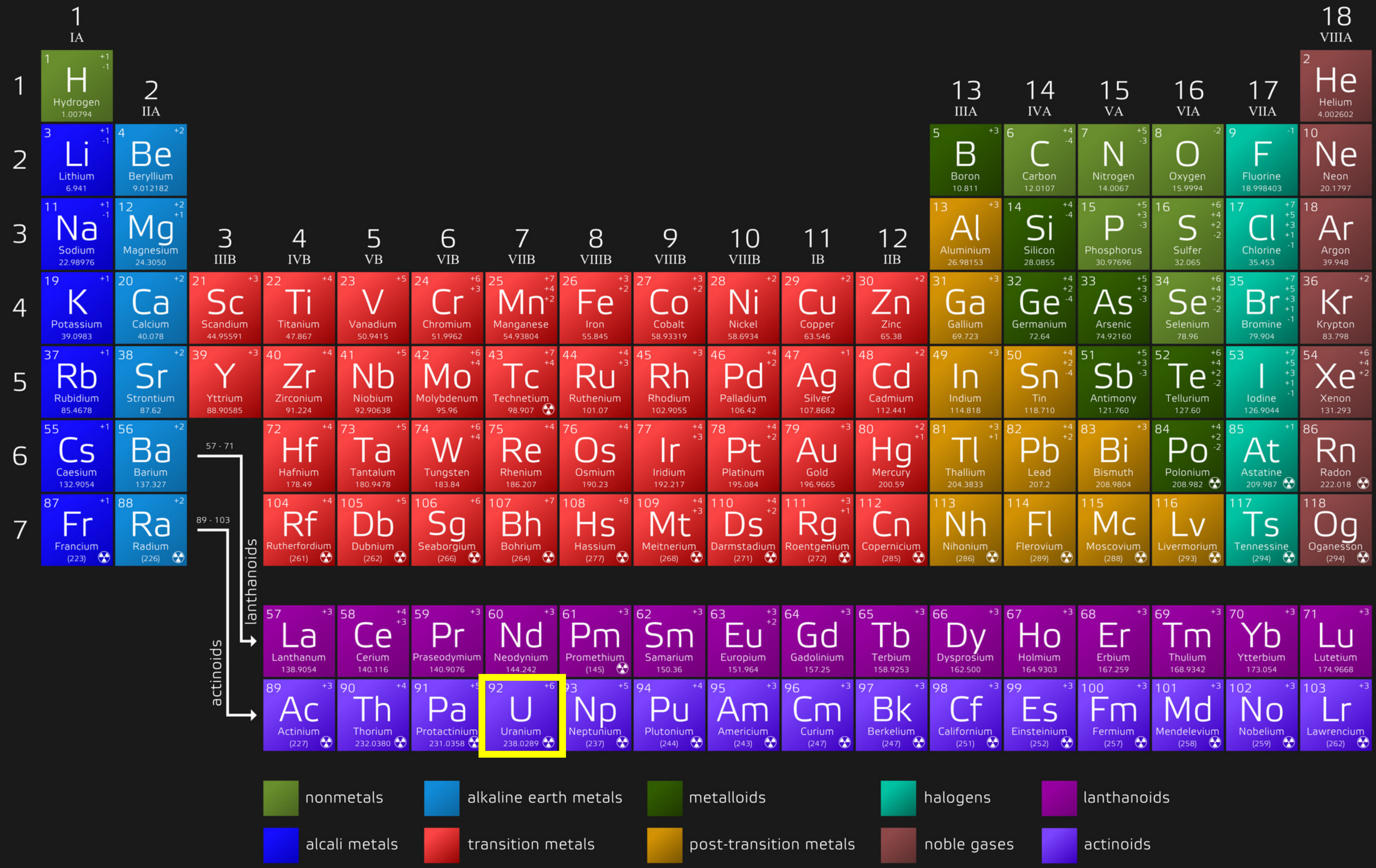
We use abbreviations called ***atomic symbols*** to identify elements. For example, here’s the symbol for Uranium (U):



Atomic # (# of protons)

$$^{238}_{92}\text{U}$$

The ***atomic number*** corresponds to the box in which the element appears on the periodic table. For example, U appears in box 92 of the periodic table, so its ***atomic number*** is 92. The ***atomic number*** is <u>always the same</u> for all elements with the same letter symbol. In other words, all Mg elements have an ***atomic number*** of 12. Also, the ***atomic number*** is equal to the number of protons, so all Mg atoms have 12 protons. <u>The number of protons in a given element never changes</u>.



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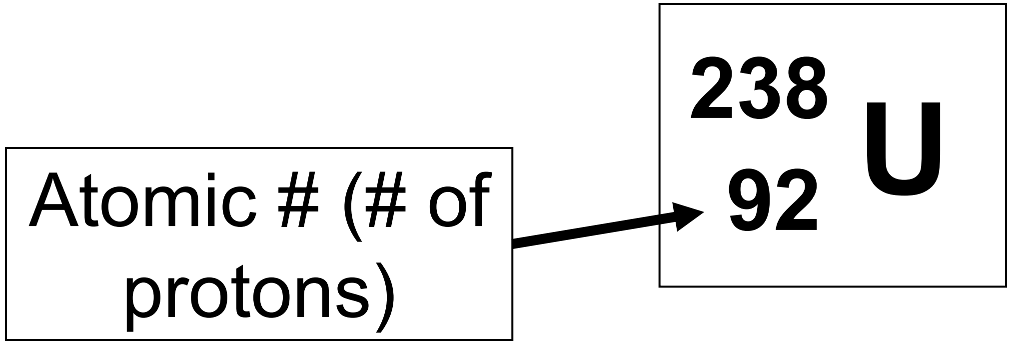
1	1 +1 -1 H Hydrogen 1.00794																	2 He Helium 4.002602
2	3 +1 -1 Li Lithium 6.941	4 +2 Be Beryllium 9.012182											5 +3 B Boron 10.811	6 +4 -4 C Carbon 12.0107	7 +5 -3 N Nitrogen 14.0067	8 -2 O Oxygen 15.9994	9 -1 F Fluorine 18.998403	10 Ne Neon 20.1797
3	11 +1 -1 Na Sodium 22.98976	12 +2 +1 Mg Magnesium 24.3050											13 +3 Al Aluminium 26.98153	14 +4 -4 Si Silicon 28.0855	15 +5 +3 -3 P Phosphorus 30.97696	16 +6 +4 +2 -2 S Sulfer 32.065	17 +7 +5 +3 +1 -1 Cl Chlorine 35.453	18 Ar Argon 39.948
4	19 +1 K Potassium 39.0983	20 +2 Ca Calcium 40.078	21 +3 Sc Scandium 44.95591	22 +4 Ti Titanium 47.867	23 +5 V Vanadium 50.9415	24 +6 +3 Cr Chromium 51.9962	25 +7 +4 +2 Mn Manganese 54.93804	26 +3 +2 Fe Iron 55.845	27 +3 +2 Co Cobalt 58.93319	28 +2 Ni Nickel 58.6934	29 +2 Cu Copper 63.546	30 +2 Zn Zinc 65.38	31 +3 Ga Gallium 69.723	32 +4 +2 -4 Ge Germanium 72.64	33 +5 +3 -3 As Arsenic 74.92160	34 +6 +4 +2 -2 Se Selenium 78.96	35 +7 +5 +3 +1 -1 Br Bromine 79.904	36 +2 Kr Krypton 83.798
5	37 +1 Rb Rubidium 85.4678	38 +2 Sr Strontium 87.62	39 +3 Y Yttrium 88.90585	40 +4 Zr Zirconium 91.224	41 +5 Nb Niobium 92.90638	42 +6 +4 Mo Molybdenum 95.96	43 +7 +4 Tc Technetium 98.907 ☢	44 +4 +3 Ru Ruthenium 101.07	45 +3 Rh Rhodium 102.9055	46 +4 +2 Pd Palladium 106.42	47 +1 Ag Silver 107.8682	48 +2 Cd Cadmium 112.441	49 +3 In Indium 114.818	50 +4 +2 -4 Sn Tin 118.710	51 +5 +3 -3 Sb Antimony 121.760	52 +6 +4 +2 -2 Te Tellurium 127.60	53 +7 +5 +3 +1 -1 I Iodine 126.9044	54 +6 +4 +2 Xe Xenon 131.293
6	55 +1 Cs Caesium 132.9054	56 +2 Ba Barium 137.327	57 - 71	72 +4 Hf Hafnium 178.49	73 +5 Ta Tantalum 180.9478	74 +6 +4 W Tungsten 183.84	75 +4 Re Rhenium 186.207	76 +4 Os Osmium 190.23	77 +4 +3 Ir Iridium 192.217	78 +4 +2 Pt Platinum 195.084	79 +3 Au Gold 196.9665	80 +2 +1 Hg Mercury 200.59	81 +3 +1 Tl Thallium 204.3833	82 +4 +2 Pb Lead 207.2	83 +3 Bi Bismuth 208.9804	84 +4 +2 -2 Po Polonium 208.982 ☢	85 +1 At Astatine 209.987 ☢	86 Rn Radon 222.018 ☢
7	87 +1 Fr Francium (223) ☢	88 +2 Ra Radium (226) ☢	89 - 103	104 +4 Rf Rutherfordium (261) ☢	105 +5 Db Dubnium (262) ☢	106 +6 Sg Seaborgium (266) ☢	107 +7 Bh Bohrium (264) ☢	108 +8 Hs Hassium (277) ☢	109 +4 +3 Mt Meitnerium (268) ☢	110 +4 +2 Ds Darmstadium (271) ☢	111 +3 +1 Rg Roentgenium (272) ☢	112 Cn Copernicium (285) ☢	113 Nh Nihonium (286) ☢	114 Fl Flerovium (289) ☢	115 Mc Moscovium (288) ☢	116 Lv Livermorium (293) ☢	117 Ts Tennessine (294) ☢	118 Og Oganesson (294) ☢

lanthanoids	57 +3 La Lanthanum 138.9054	58 +4 +3 Ce Cerium 140.116	59 +3 Pr Praseodymium 140.9076	60 +3 Nd Neodymium 144.242	61 +3 Pm Promethium (145) ☢	62 +3 Sm Samarium 150.36	63 +3 +2 Eu Europium 151.964	64 +3 Gd Gadolinium 157.25	65 +3 Tb Terbium 158.9253	66 +3 Dy Dysprosium 162.500	67 +3 Ho Holmium 164.9303	68 +3 Er Erbium 167.259	69 +3 Tm Thulium 168.9342	70 +3 Yb Ytterbium 173.054	71 +3 Lu Lutetium 174.9668
actinoids	89 +3 Ac Actinium (227) ☢	90 +4 Th Thorium 232.0380 ☢	91 +5 Pa Protactinium 231.0358 ☢	92 +6 U Uranium 238.0289 ☢	93 +5 Np Neptunium (237) ☢	94 +4 Pu Plutonium (244) ☢	95 +3 Am Americium (243) ☢	96 +3 Cm Curium (247) ☢	97 +3 Bk Berkelium (247) ☢	98 +3 Cf Californium (251) ☢	99 +3 Es Einsteinium (252) ☢	100 +3 Fm Fermium (257) ☢	101 +3 Md Mendelevium (258) ☢	102 +3 No Nobelium (259) ☢	103 +3 Lr Lawrencium (262) ☢

- 🟩 nonmetals
- 🟦 alcali metals
- 🟦 alkaline earth metals
- 🟥 transition metals
- 🟩 metalloids
- 🟧 post-transition metals
- 🟩 halogens
- 🟫 noble gases
- 🟪 lanthanoids
- 🟪 actinoids

Atomic (Chemical) Symbols

We use abbreviations called ***atomic symbols*** to identify elements. For example, here's the symbol for Uranium (U):



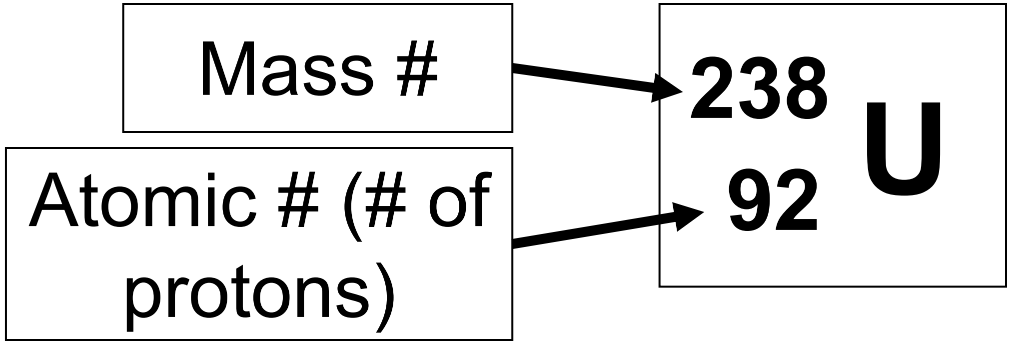
Atomic # (# of protons)

$$^{238}_{92}\mathrm{U}$$

The ***atomic number*** corresponds to the box in which the element appears on the periodic table. For example, U appears in box 92 of the periodic table, so its ***atomic number*** is 92. The ***atomic number*** is always the same for all elements with the same letter symbol. In other words, all U atoms have an ***atomic number*** of 92. Also, the ***atomic number*** is *equal to the number of protons*, so all U atoms have 92 protons. *The number of protons in a given element never changes!*

Atomic Weight

An atom’s mass is equal to the combined masses of its ***protons*** and ***neutrons***, which are (once again) located in its nucleus.

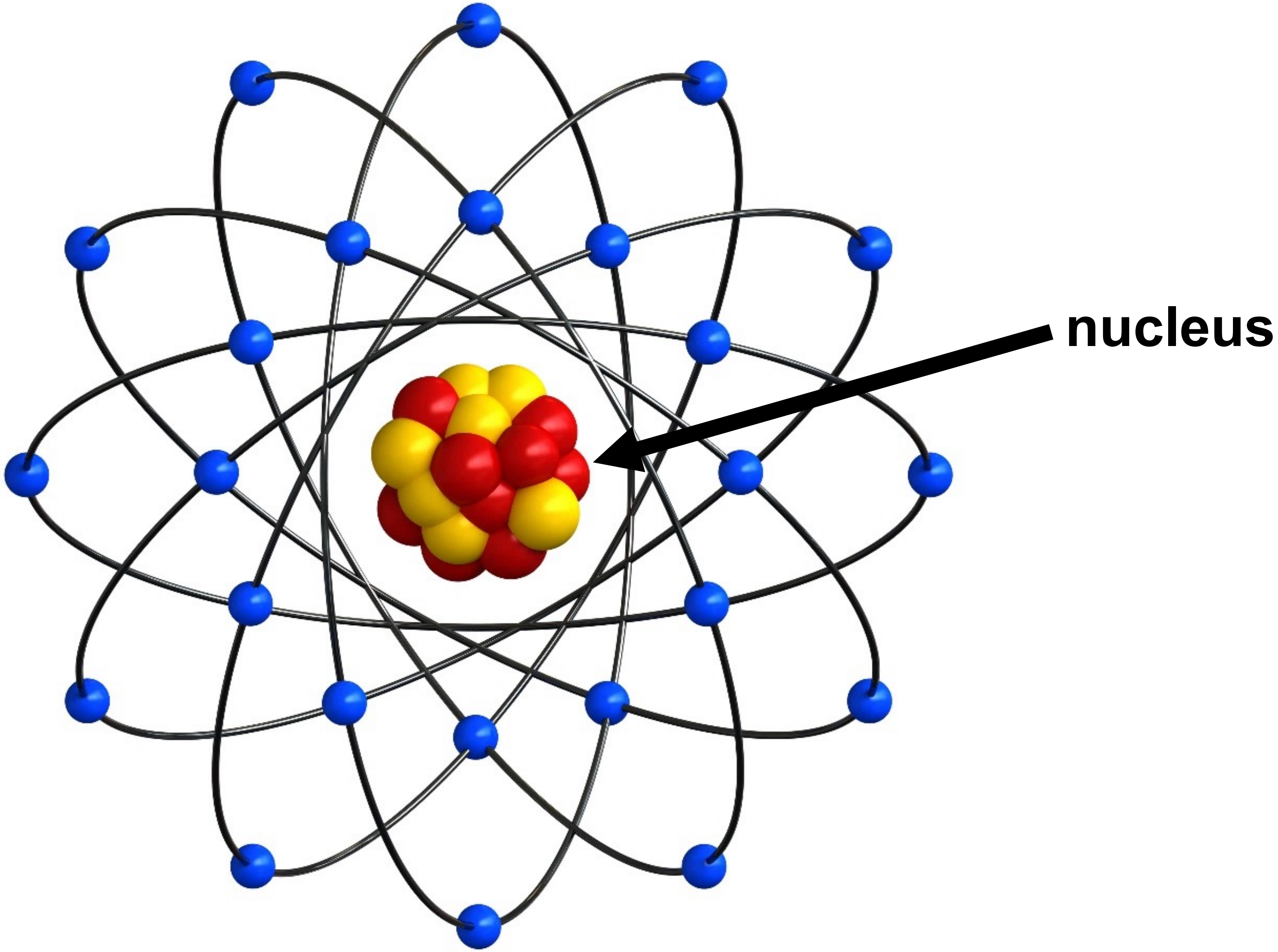


Mass #

Atomic # (# of protons)

$^{238}_{92}\mathrm{U}$

Mass # = (# of protons) + (# of neutrons)


nucleus

Atomic Weight

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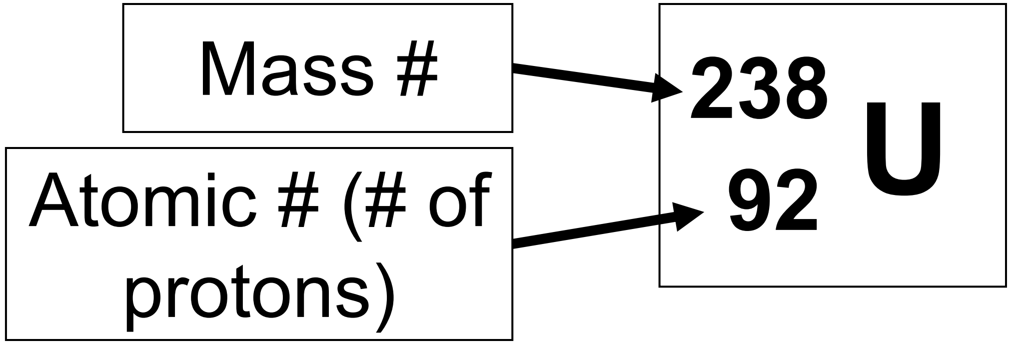


Mass #

Atomic # (# of protons)

$^{238}_{92}\mathrm{U}$

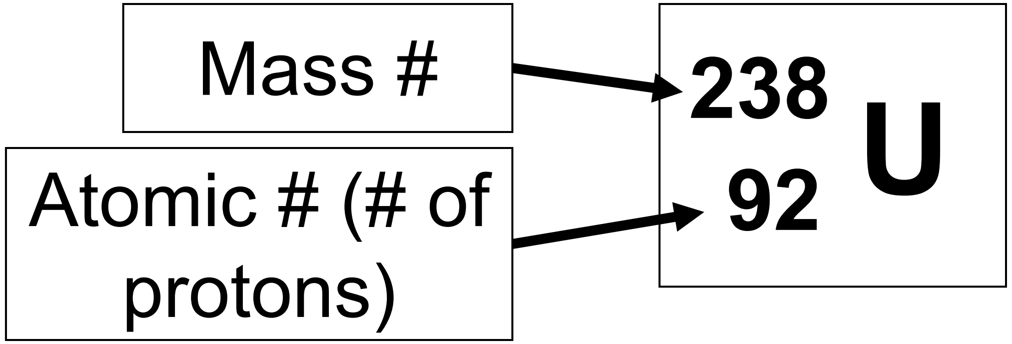
Mass # = (# of protons) + (# of neutrons)

So how can you calculate an atom’s number of ***neutrons?***

# of neutrons = Mass # – # of protons
= Mass # – atomic #

Atomic Weight

The ***mass number*** CAN change for different atoms of the same element.


Mass #

Atomic # (# of protons)

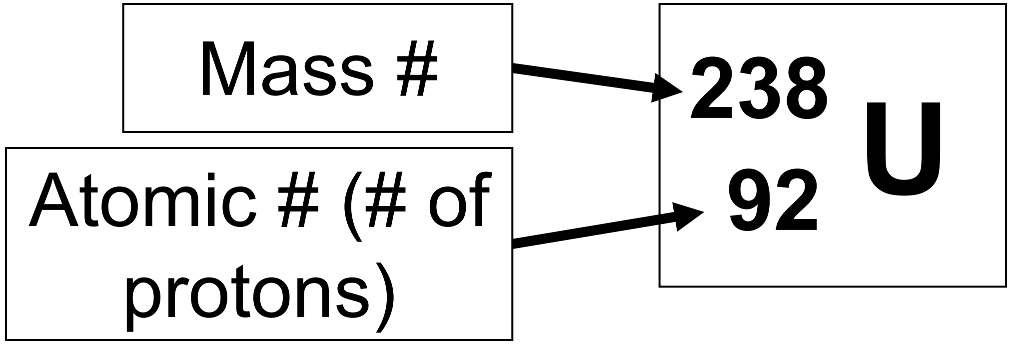
$^{238}_{92}\mathrm{U}$

Mass # = (# of protons) + (# of neutrons)

As I already mentioned, all atoms of the same element have the SAME number of *protons and the SAME atomic number*. For example, all U atoms have 92 protons. You CANNOT change the number of protons in an element. If you do, you change the element.

Atomic Weight

The ***mass number*** <u>CAN</u> change for different atoms of the same element.



Mass #

Atomic # (# of protons)

$$^{238}_{92}\mathrm{U}$$

Mass # = (# of protons) + (# of neutrons)

However, different atoms of the same element *<u>can</u>* have different numbers of neutrons. For example, a ^{238}U atom has 92 protons and 146 neutrons (92+146=238), but not all U atoms weigh 238. Some weigh 236, and some weigh 235. A ^{236}U atom, for example, has 92 protons and 144 neutrons (92+144=236), while a ^{235}U atom has 92 protons and 143 neutrons (92+143=235).

Atoms of the same element that have different ***mass numbers***, are called ***isotopes.***

For example, most carbon atoms have 6 neutrons. We call these “carbon 12” (or ^{12}C) atoms because they weigh 12 amu (6 neutrons + 6 protons = 12 amu). In contrast, other carbon atoms have 7 neutrons. These are called “carbon 13” (or ^{13}C) atoms because they weigh 13 amu. We would say that ^{12}C and ^{13}C atoms are different ***isotopes*** of carbon. All of these, however, have the same *atomic number*, 6, and the same number of protons (6).

Once again, you CAN change an element’s number of electrons or neutrons without changing which element it is, but you CANNOT change its number of protons. For example, if you change a carbon atom’s number of protons to 7, it is no longer carbon. It is now nitrogen.

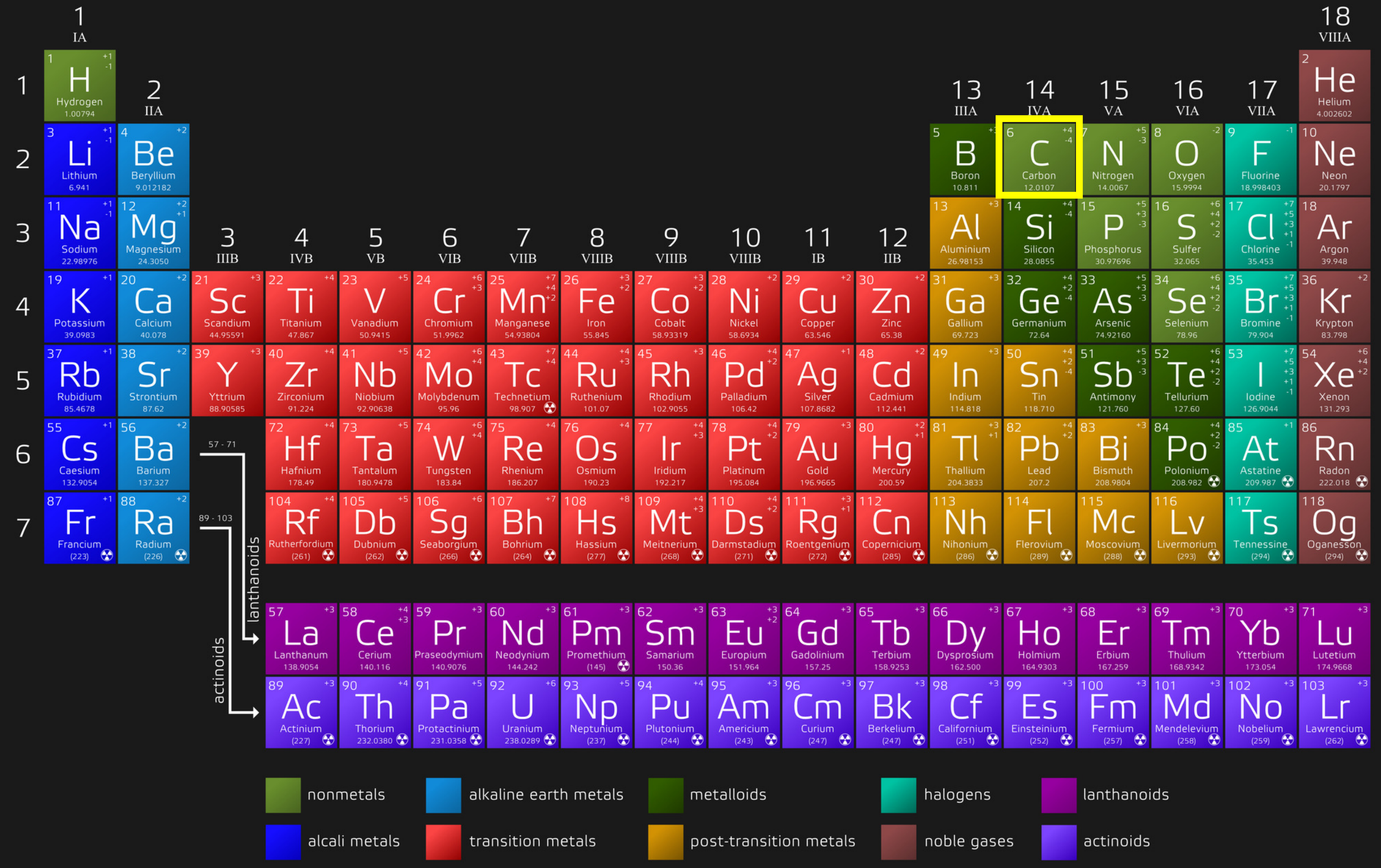
Atomic Weight vs. Mass Number

There is a slight, technical difference between an atom's ***atomic weight*** and an element's ***mass number*** (or ***atomic mass***). For example a ^{12}C atom has a ***mass number*** of 12 amu, which means that *it weighs 12 amu*, end of story.

However, if you look on the periodic table, you'll see that as an element, carbon's ***atomic weight*** is listed as 12.0107, not 12.

Why? Because in real life, there are *four* different isotopes of carbon (^{11}C, ^{12}C, ^{13}C, and ^{14}C), each present in a different percentage in nature. ^{12}C is by far the most abundant.

The ***atomic weight*** on the periodic table, then, is a *weighted average* of all of these four isotopes' individual ***atomic masses***, mathematically combined and averaged into one number. The reason it's so close to 12 is because >98% of all C atoms are ^{12}C.



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4	19 K Potassium 39.0983 +1	20 Ca Calcium 40.078 +2	21 Sc Scandium 44.95591 +3	22 Ti Titanium 47.867 +4	23 V Vanadium 50.9415 +5	24 Cr Chromium 51.9962 +6 +3	25 Mn Manganese 54.93804 +7 +4 +2	26 Fe Iron 55.845 +3 +2	27 Co Cobalt 58.93319 +3 +2	28 Ni Nickel 58.6934 +2	29 Cu Copper 63.546 +2	30 Zn Zinc 65.38 +2	31 Ga Gallium 69.723 +3	32 Ge Germanium 72.64 +4 +2 -4	33 As Arsenic 74.92160 +5 +3 -3	34 Se Selenium 78.96 +6 +4 +2 -2	35 Br Bromine 79.904 +7 +5 +3 +1 -1	36 Kr Krypton 83.798 +2
5	37 Rb Rubidium 85.4678 +1	38 Sr Strontium 87.62 +2	39 Y Yttrium 88.90585 +3	40 Zr Zirconium 91.224 +4	41 Nb Niobium 92.90638 +5	42 Mo Molybdenum 95.96 +6 +4	43 Tc Technetium 98.907 ☢ +7 +4	44 Ru Ruthenium 101.07 +4 +3	45 Rh Rhodium 102.9055 +3	46 Pd Palladium 106.42 +4 +2	47 Ag Silver 107.8682 +1	48 Cd Cadmium 112.441 +2	49 In Indium 114.818 +3	50 Sn Tin 118.710 +4 +2 -4	51 Sb Antimony 121.760 +5 +3 -3	52 Te Tellurium 127.60 +6 +4 +2 -2	53 I Iodine 126.9044 +7 +5 +3 +1 -1	54 Xe Xenon 131.293 +6 +4 +2
6	55 Cs Caesium 132.9054 +1	56 Ba Barium 137.327 +2	57 - 71	72 Hf Hafnium 178.49 +4	73 Ta Tantalum 180.9478 +5	74 W Tungsten 183.84 +6 +4	75 Re Rhenium 186.207 +4	76 Os Osmium 190.23 +4	77 Ir Iridium 192.217 +4 +3	78 Pt Platinum 195.084 +4 +2	79 Au Gold 196.9665 +3	80 Hg Mercury 200.59 +2 +1	81 Tl Thallium 204.3833 +3 +1	82 Pb Lead 207.2 +4 +2	83 Bi Bismuth 208.9804 +3	84 Po Polonium 208.982 ☢ +4 +2 -2	85 At Astatine 209.987 ☢ +1	86 Rn Radon 222.018 ☢
7	87 Fr Francium (223) ☢ +1	88 Ra Radium (226) ☢ +2	89 - 103	104 Rf Rutherfordium (261) ☢ +4	105 Db Dubnium (262) ☢ +5	106 Sg Seaborgium (266) ☢ +6	107 Bh Bohrium (264) ☢ +7	108 Hs Hassium (277) ☢ +8	109 Mt Meitnerium (268) ☢ +4 +3	110 Ds Darmstadium (271) ☢ +4 +2	111 Rg Roentgenium (272) ☢ +3 +1	112 Cn Copernicium (285) ☢	113 Nh Nihonium (286) ☢	114 Fl Flerovium (289) ☢	115 Mc Moscovium (288) ☢	116 Lv Livermorium (293) ☢	117 Ts Tennessine (294) ☢	118 Og Oganesson (294) ☢

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- nonmetals
- alcali metals
- alkaline earth metals
- transition metals
- metalloids
- post-transition metals
- halogens
- noble gases
- lanthanoids
- actinoids

Atomic Weight vs. Mass Number

Thus, an atom's ***atomic mass*** is its *exact mass*. For example, a ^{12}C atom weighs *exactly* 12 amu. A ^{13}C weighs *exactly* 13 amu, and so forth.

In contrast, an element's ***atomic weight*** is the number that appears on the periodic table, which is the sum of all of its isotopes' individual ***atomic masses***, each multiplied by its own abundance percentage.

As I already explained, an individual atom's ***mass number*** (***atomic mass***) is equal to its number of protons plus its number of neutrons. Thus:

# of neutrons = mass # – # of protons

Nuclear Particles

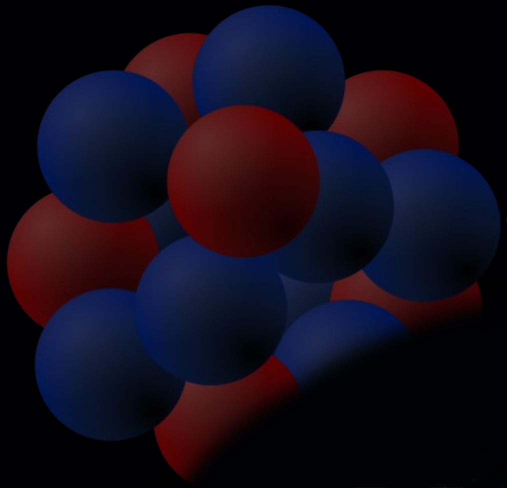
In the realm of nuclear chemistry and nuclear physics, for the EXAM, you should memorize the six nuclear particles' names and formulas shown below. These are listed here in order of increasing mass and penetration, which (you'll notice) have inverse relationships. We'll discuss and use these nuclear particles more later on.

not true He atoms because they have no electrons

increasing mass

Particle	Symbol
Alpha Particle	${}^{4}_{2}\mathbf{He}$ or ${}^{4}_{2}\alpha$
Neutron	${}^{1}_{0}\mathbf{n}$
Proton	${}^{1}_{1}\mathbf{H}$ or ${}^{1}_{1}\mathbf{p}$
Beta Particle (high-speed electron)	${}^{0}_{-1}\mathbf{e}$ or ${}^{0}_{-1}\beta$
Positron	${}^{0}_{+1}\mathbf{e}$
Gamma particle	${}^{0}_{0}\gamma$

increasing penetration



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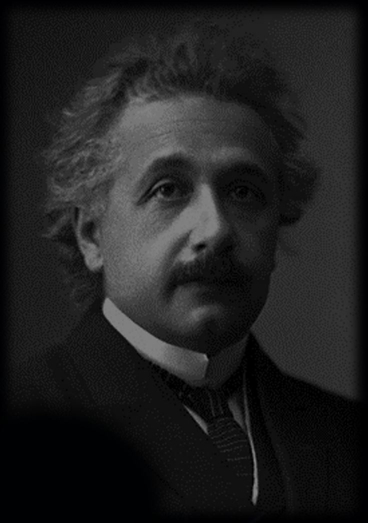
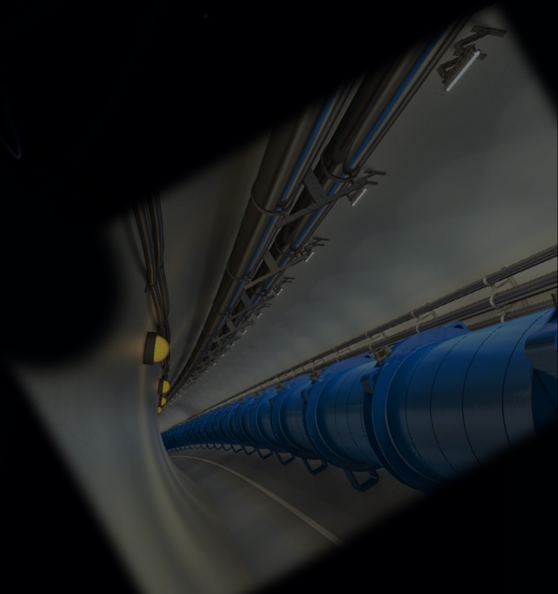
Nuclear Chemistry

(Nuclear Reactions)



Albert Einstein



To Review . . .

In a prior video, I taught you that for the EXAM, you need to memorize the six nuclear particles’ names and formulas shown below, listed here in order of increasing mass and penetration, which are inversely related.

increasing mass

Particle	Symbol
Alpha Particle	${}^{4}_{2}\mathbf{He}$ or ${}^{4}_{2}\alpha$
Neutron	${}^{1}_{0}\mathbf{n}$
Proton	${}^{1}_{1}\mathbf{H}$ or ${}^{1}_{1}\mathbf{p}$
Beta Particle (high-speed electron)	${}^{0}_{-1}\mathbf{e}$ or ${}^{0}_{-1}\beta$
Positron	${}^{0}_{+1}\mathbf{e}$
Gamma particle	${}^{0}_{0}\gamma$

increasing penetration

Nuclear Reactions

Unlike traditional chemical reactions, which involve the giving or sharing of *electrons*, ***nuclear reactions*** involve ***nuclear particles***: specifically, the ones listed in this table.

When an atom *gives off* (**emits**) these particles, it's called ***nuclear emission***, ***decay***, or ***fission***.

increasing mass

Particle	Symbol
Alpha Particle	${}^{4}_{2}\mathbf{He}$ or ${}^{4}_{2}\alpha$
Neutron	${}^{1}_{0}\mathbf{n}$
Proton	${}^{1}_{1}\mathbf{H}$ or ${}^{1}_{1}\mathbf{p}$
Beta Particle (high-speed electron)	${}^{0}_{-1}\mathbf{e}$ or ${}^{0}_{-1}\beta$
Positron	${}^{0}_{+1}\mathbf{e}$
Gamma particle	${}^{0}_{0}\gamma$

increasing penetration

Decay (Fission) Reactions

In general, ***nuclear decay (fission*** or ***emission) reactions*** look like this:

$$\underset{\text{reactant}}{{}^{A}_{B}X} \rightarrow \underbrace{{}^{C}_{D}Y \quad + \quad {}^{E}_{F}Z}_{\text{products}}$$

Where **A = C + E** and **B = D + F**. The identities of **X, Y,** and **Z** are determined completely by their ***atomic numbers***, **B**, **D**, and **F**.

Capture (Fusion) Reactions

In contrast, ***nuclear capture (fusion) reactions*** look like this:

$$\underbrace{{}^{A}_{B}X + {}^{C}_{D}Y}_{\text{reactants}} \rightarrow \underset{\text{product}}{\boxed{{}^{E}_{F}Z}}$$

Where **A + C = E** and **B + D = F**. The identities of **X, Y,** and **Z** are determined completely by their ***atomic numbers***, **B**, **D**, and **F**.

Nuclear Reactions Involve Mass Change

Just so you know, both classes of nuclear reactions (***fission*** and ***fusion***) also have an accompanying release of energy. All ***nuclear reactions***, then, look generically like this:

reactants → products + energy

This release of energy always involves a change in mass, according to Einstein's famous equation $E = mc^2$, or (in this context) $\Delta E = \Delta mc^2$.

So yes, in nuclear reactions, mass and energy are interchanged. We'll discuss this more in a later video.

How To Do Nuclear Reaction Questions

For now, please understand that for most ***nuclear reaction*** questions, you'll be given all the reactants and products except for one, and will then be asked to identify the missing piece.

To do this, you simply have to make sure that your subscript and superscript numbers all add up to equal each other on both sides of the equation.

$$^{A}_{B}X + ^{C}_{D}Y \rightarrow ^{E}_{F}Z$$

Please note: Even though, in these problems, we balance these nuclear reactions completely and usually leave ***energy*** out of the equation, *there is still some amount of* ***energy*** *given off in every nuclear reaction*, which means that *in every nuclear reaction, some mass is lost*. For these problems, we just use rounded numbers that are so much larger than the lost mass that we don't really have to account for it.

Questions

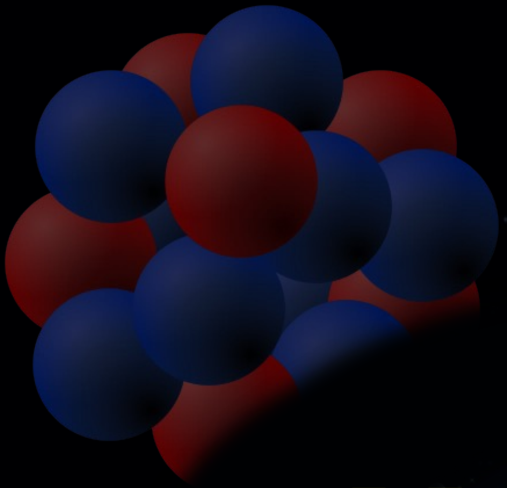
Please identify the missing atom or particle (**???**) in the following ***nuclear reaction:***

$$^{131}_{53}\mathrm{I} \rightarrow {}^{0}_{-1}\beta + ???$$

Questions

Please identify the missing atom or particle (**???**) in the following ***nuclear reaction:***

$$^{214}_{84}\mathrm{Po} \; + \; 2\,^{4}_{2}\alpha \; + \; 2\,^{0}_{-1}\beta \;\rightarrow\; ???$$



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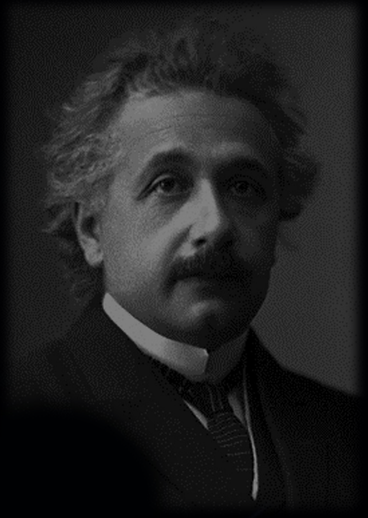
Nuclear Chemistry



Albert Einstein

(Nuclear Stability and Radioactive Decay)

The Band of Stability

Because nuclei contain a bunch of protons that are squished tightly together in a very small space, you might wonder why their repulsion for each other (caused by their all having the same positive charge) wouldn't cause the nucleus to simply blow apart.

As it turns out, at close distance, nuclear particles (protons and neutrons) are actually *attracted* to each other by a force called the ***strong force*** (also called the ***nuclear force***), notwithstanding the repulsion caused by protons all having the same charge as each other.

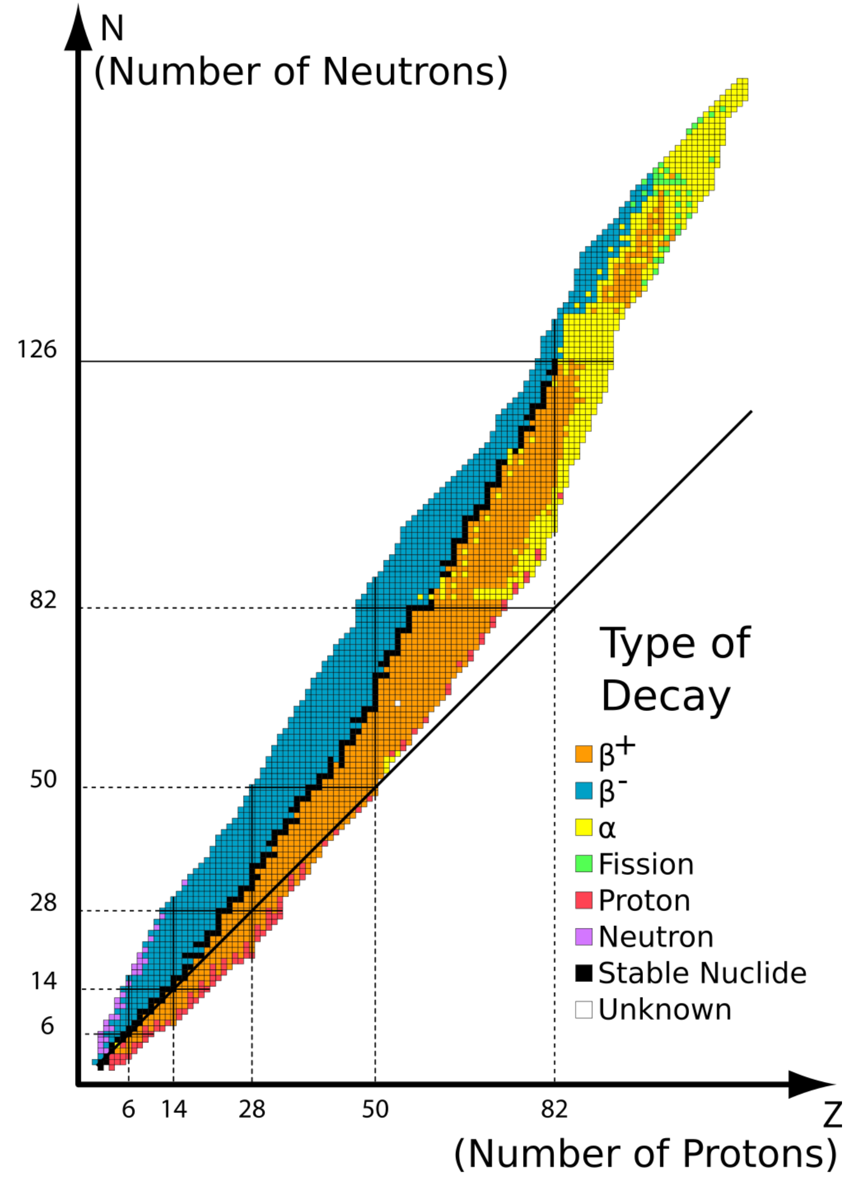
In addition to that, neutrons help offset the repulsive forces between protons in an atom's nucleus. Thus, an atom's ratio of neutrons-to-protons is intimately connected to its stability.

The Belt of Stability

For reasons that aren’t extremely clear to me, for the first 20 elements on the periodic table, atoms must have close to a 1-to-1 ratio of protons-to-neutrons in order to be stable.

For atoms beyond that, their *neutron number can exceed their proton number*, **to a point**, while still maintaining stability.

This **chart** shows a blueish/yellowish-colored band (called the ***belt of stability***), which represents all the neutron-to-proton ratios that are stable. All isotopes that fall *outside* of this band are unstable (***radioactive***) and will spontaneously undergo ***nuclear reactions***.

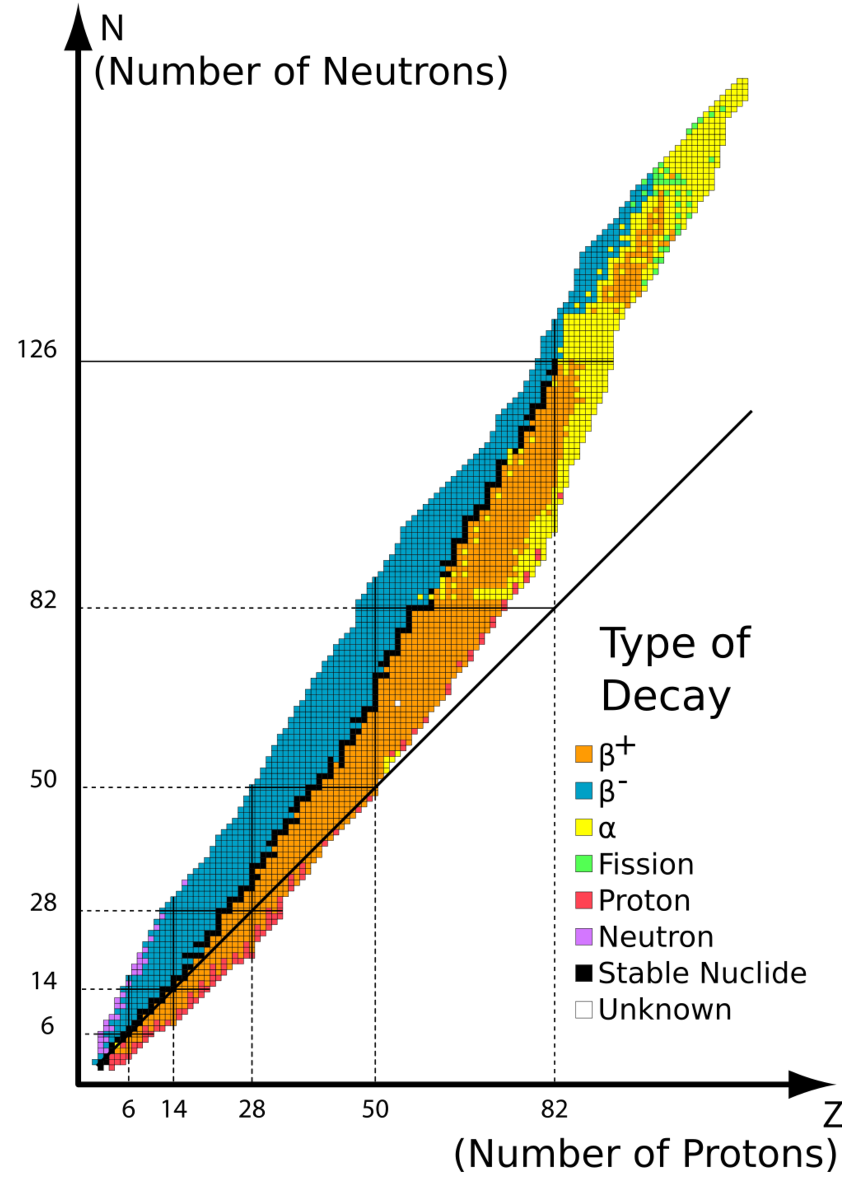


The Belt of Stability

For reasons that aren't extremely clear to me, for the first 20 elements on the periodic table, atoms must have close to a 1-to-1 ratio of protons-to-neutrons in order to be stable.

For atoms beyond that, their *neutron number can exceed their proton number*, **to a point**, while still maintaining stability.

You might notice by examining this chart that as the number of protons increases, the number of neutrons must *also* increase –but at a faster, disproportion rate, relative to protons– to maintain a stable nucleus.

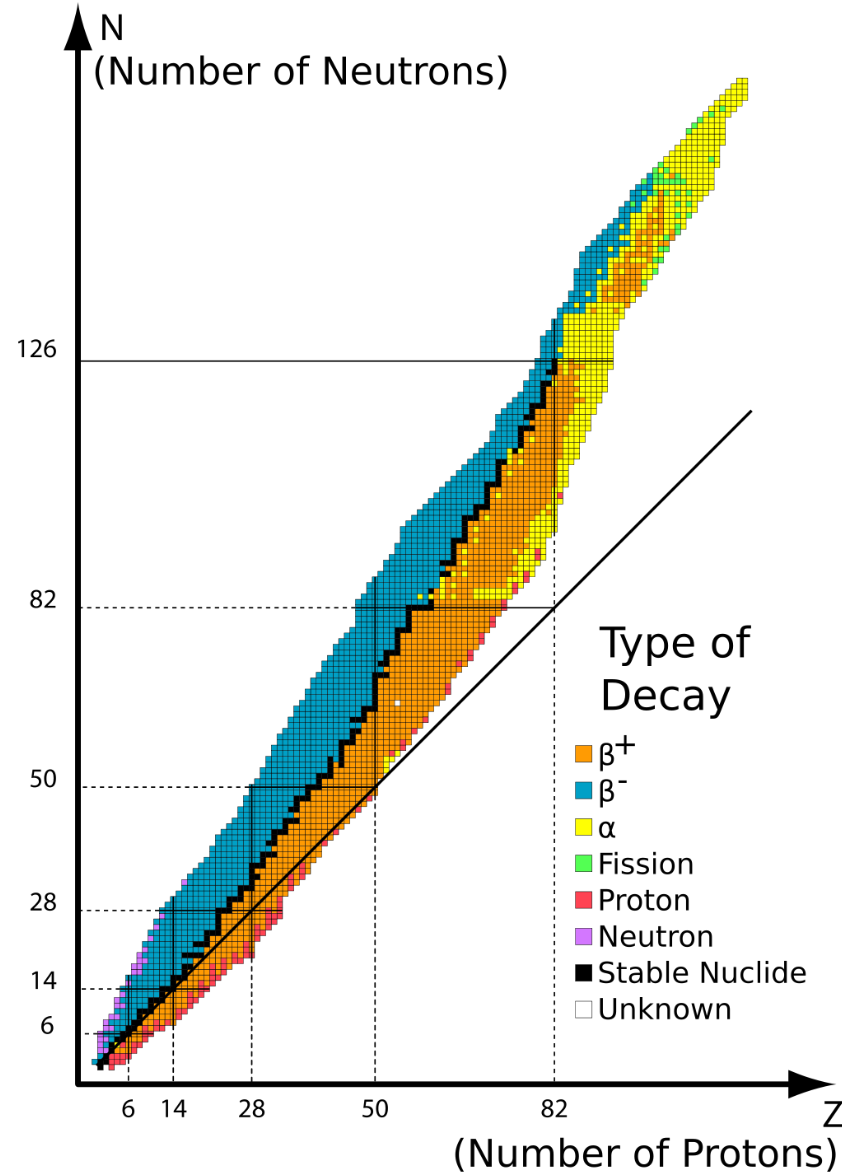


The Belt of Stability

You can predict what type of ***nuclear reaction*** a radioactive isotope will undergo by looking at its neutron-to-proton ratio, relative to this ***belt of stability***. For example:

- **Atoms above the belt** – These isotopes have too many neutrons, relative to protons. They decrease this neutron-to-proton ratio by emitting beta particles through ***beta decay***. For example:

$$^{135}_{54}\mathrm{Xe} \rightarrow {}^{0}_{-1}\beta + {}^{135}_{55}\mathrm{Cs}$$



The Belt of Stability

You can predict what type of ***nuclear reaction*** a radioactive isotope will undergo by looking at its neutron-to-proton ratio, relative to this ***belt of stability***. For example:

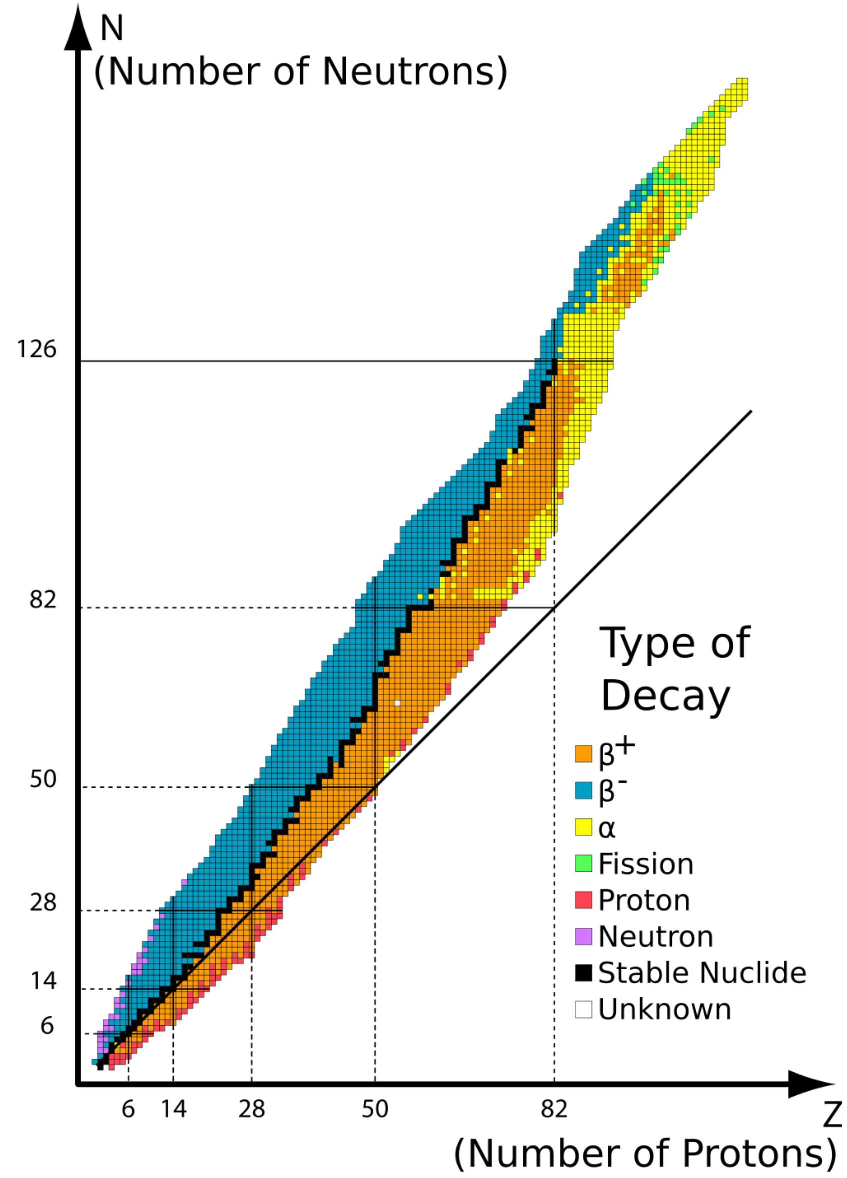
- **Atoms below the belt** – These isotopes have too many protons, relative to neutrons. They fix this problem by either emitting a ***positron*** (***positron emission***) or capturing an electron (***electron capture***). For instance:

Positron emission

$$ {}^{118}_{54}\text{Xe} \rightarrow {}^{0}_{+1}\text{e} + {}^{118}_{53}\text{I} $$

Electron capture

$$ {}^{118}_{54}\text{Xe} + {}^{0}_{-1}\beta \rightarrow {}^{118}_{53}\text{I} $$



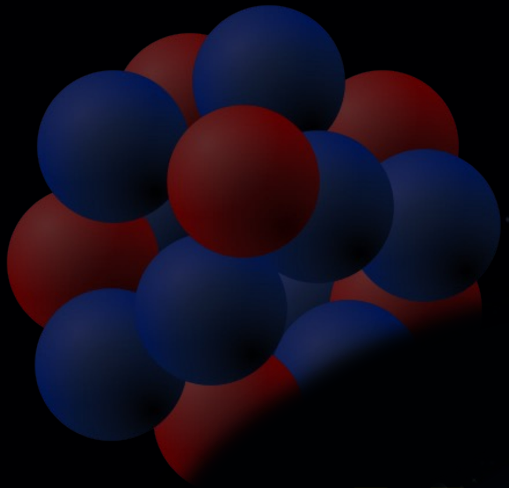
Emission Summary

This table summarizes the types of radioactive decays or particle captures that unstable isotopes undergo in their move toward stability. Please note for this table that ***N*** = # of neutrons, and ***Z*** = # of protons.

Route of decay	Nuclear particle involved	Result	Likely For?
Alpha (α) decay	${}^{4}_{2}\mathrm{He}$ or ${}^{4}_{2}\alpha$	reduces mass #	Large nuclei (Z>83)
β decay (β emission)	${}^{0}_{-1}e$ or ${}^{0}_{-1}\beta$	neutron → proton	***N***/***Z*** ratio too high (too many neutrons)
β^{+} decay (positron emission)	${}^{0}_{+1}e$ or ${}^{0}_{+1}\beta$	proton → neutron	***N***/***Z*** ratio too low (too many protons)
Electron capture	${}^{0}_{-1}e$ or ${}^{0}_{-1}\beta$	proton → neutron	***N***/***Z*** ratio is too low (too many protons)
γ decay	${}^{0}_{0}e$	no ***N***/***Z*** change	unpredictable

Questions

Oxygen-17 and a proton combine and then undergo radioactive decay to emit a nitrogen-14 atom and another particle. What type of radioactive decay is oxygen-17 undergoing in this scenario?



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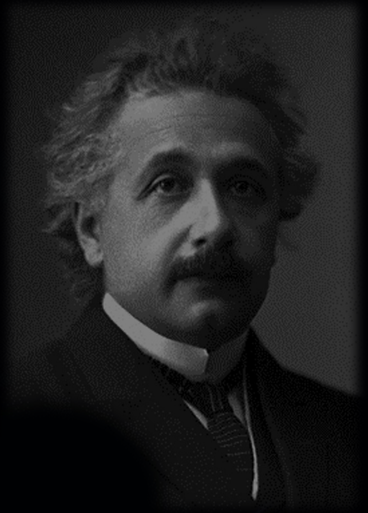


Albert Einstein

Nuclear Chemistry

(The Kinetics of Radioactive Decay)

To Review . . .

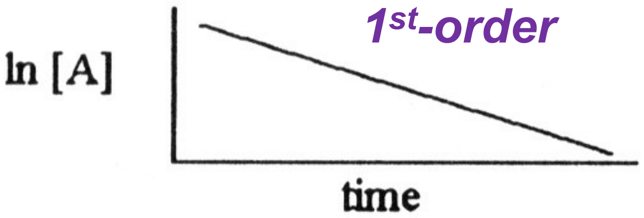
Back in our chapter on ***chemical kinetics***, I taught you about ***zero-order***, ***first-order***, and ***second-order*** reactions. I also taught you that when given a *straight-line graph* of an ***integrated rate law***, you can quickly determine the reaction's order (zero, one, or two) by looking at its y-axis units:

- ***zero-order*** = **[concentration]**
- ***first-order*** = **ln [concentration]**
- ***second order*** = **1 / [concentration]**

To Review . . .

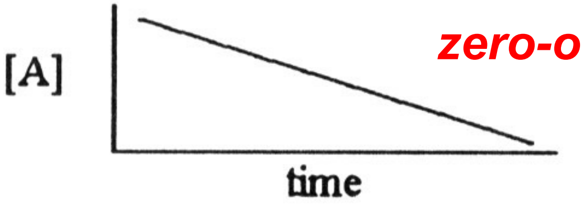
For example:

a.



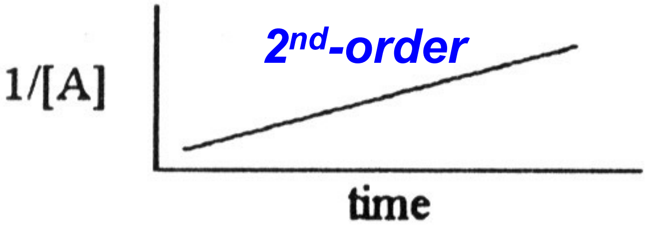
1st-order

b.



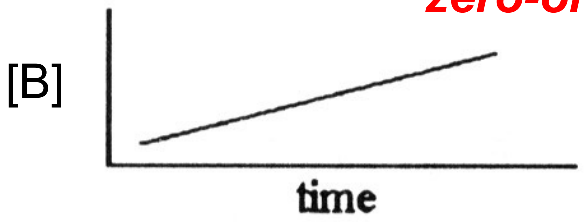
zero-order

c.



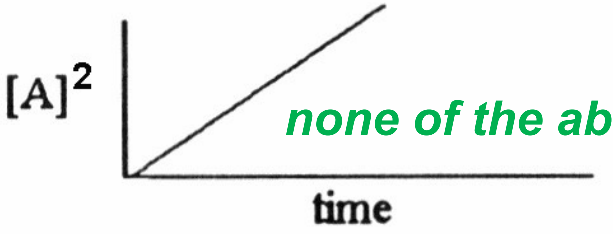
2nd-order

d.



zero-order

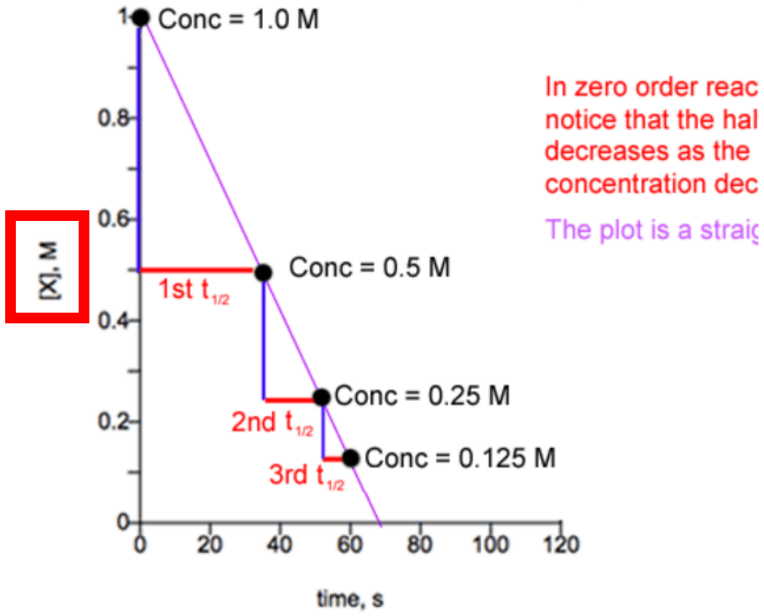
e.

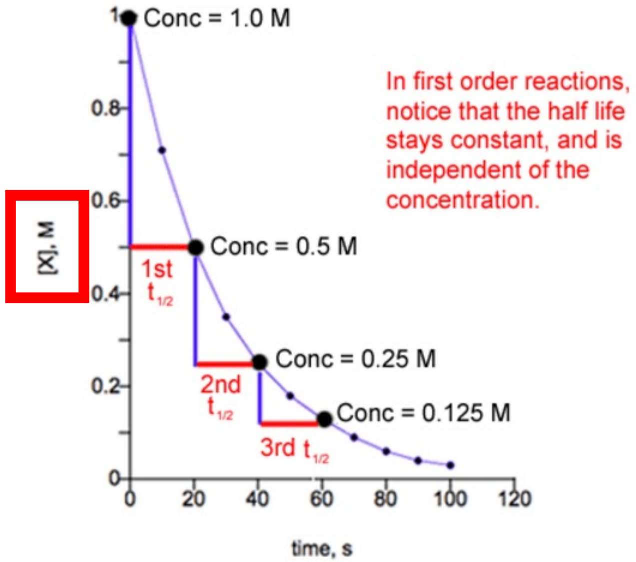


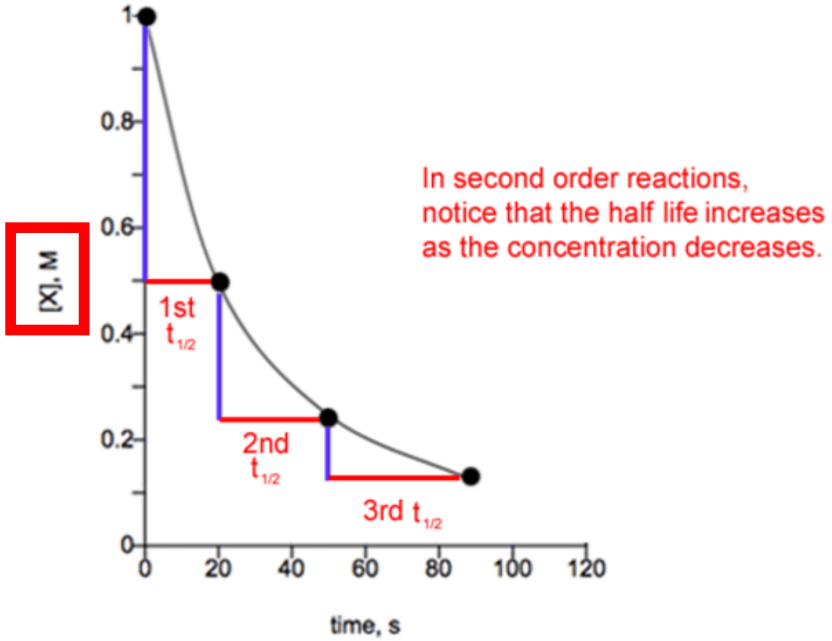
none of the above

To Review . . .

I also explained that these y-axis units only apply to *straight-line graphs* of your ***integrated rate law***. For curved-line graphs:







zero-order reactions:

- The plot is a straight line
- ↓Half-life ≈ ↓concentration

first-order reactions:

- NOT a straight line
- Half-life constant is independent of concentration

second-order reactions:

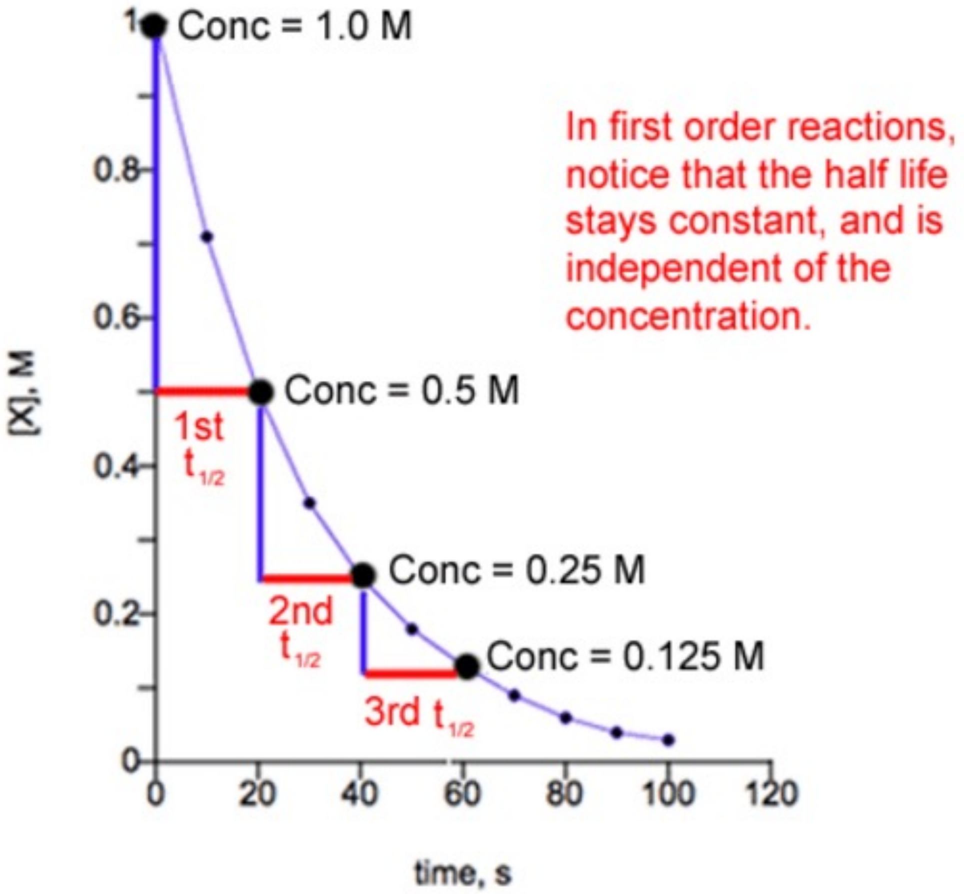
- NOT a straight line
- ↑Half-life ≈ ↓concentration

The Kinetics of Radioactive Decay

As it turns out, all ***radioactive decay*** reactions are ***first-order***. Thus, their curved-line ***integrated rate law*** graphs look like this.

Please notice that for ***first-order reactions***, *the half-life ($t_{\frac{1}{2}}$) width stays constant and is independent of concentration*.

$$half-life\ of\ first\ order\ reaction = \frac{\ln 2}{k} = \frac{0.693}{k}$$



In first order reactions, notice that the half life stays constant, and is independent of the concentration.

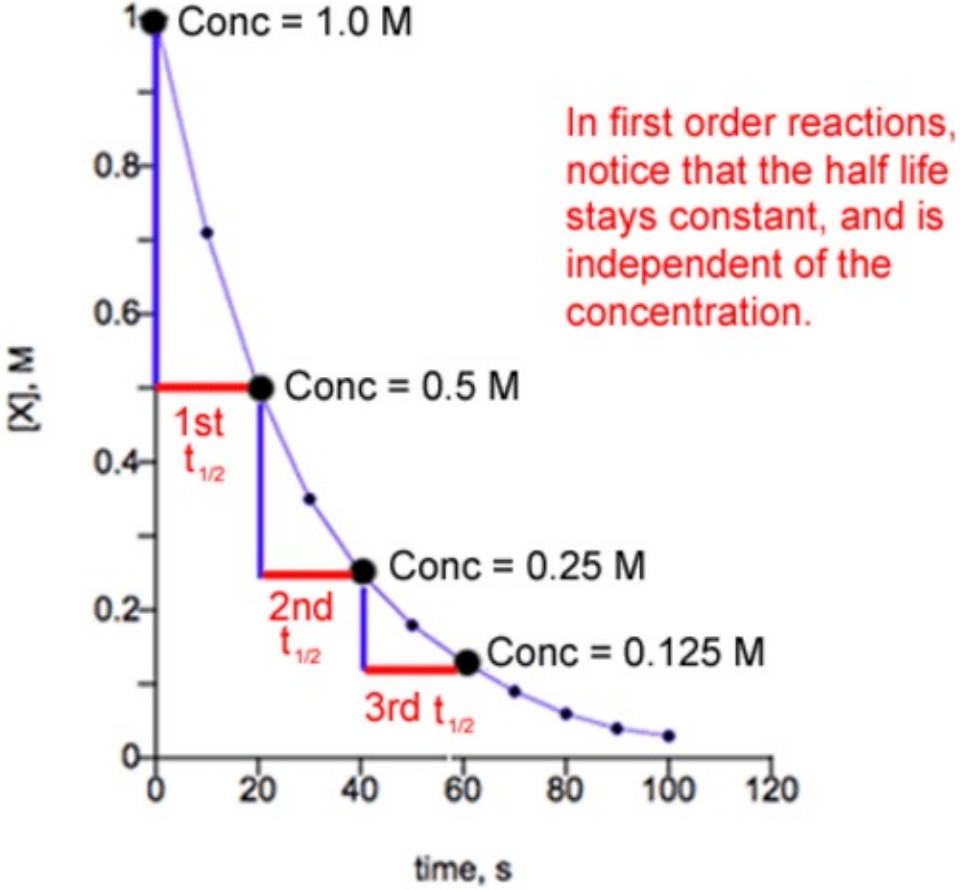
The Kinetics of Radioactive Decay

This means that every time one ***half-life*** goes by, half of your sample decays.

For example, if you start off with 100 g of a radioactive sample, then after the first ***half-life***, you'll have 50 g left. After the second ***half-life*** (which takes the same amount of time as the first ***half-life***), you'll have 25 g left.

Make sense?

$$half-life\ of\ first\ order\ reaction = \frac{\ln 2}{k} = \frac{0.693}{k}$$



In first order reactions, notice that the half life stays constant, and is independent of the concentration.

The Kinetics of Radioactive Decay

With that said, the following ***half-life*** equation would be really helpful to remember:

$$t_{½} = \frac{0.693}{k}$$

. . . Where:

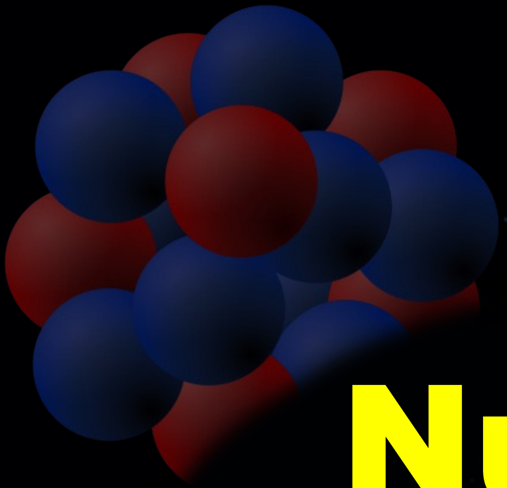
- $t_{½}$ is the substance's ***half-life***.
- k is the radioactive decay's ***rate constant.***

Questions

Given a ***half-life*** of 30 minutes and a starting sample of 100 grams, how many grams would you have left after 2 hours?

Questions

Given a ***half-life*** of 15 minutes, how long would it take for 87.5 grams of your 100-gram sample to decay?



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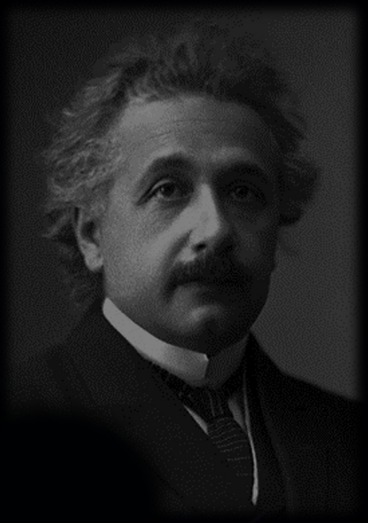
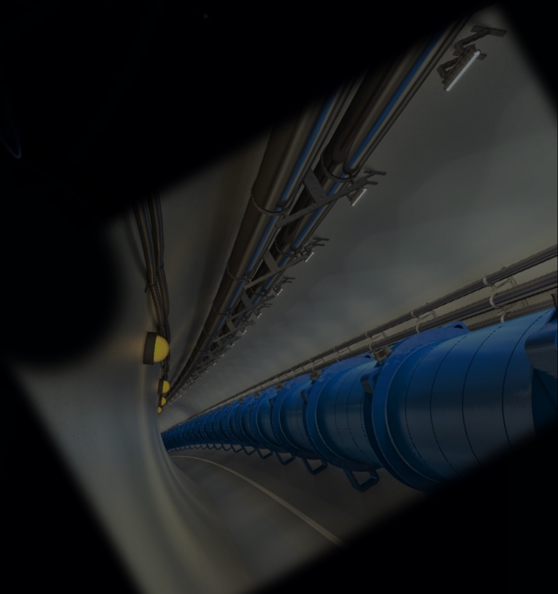
Nuclear Chemistry

(Fission, Fusion, and Nuclear Binding Energy)



Albert Einstein



To Review . . .

As we discussed in an earlier video in this chapter:

- **Fission** – This happens when larger nuclei decay (break apart) into smaller nuclei. This is how the first atomic bombs worked and how nuclear reactors work.

- **Fusion** – This happens when two or more small nuclei combine to form larger nuclei. This is what happens in stars and is how all the heavier elements in the universe were formed.

To Review . . .

I also taught you that ***fission*** and ***fusion*** reactions both have an accompanying release of energy:

reactants → products + energy

This release of energy always involves a change in mass, according to Einstein's famous equation $E = mc^2$, or (in this context) $\Delta E = \Delta mc^2$.

So yes, in nuclear reactions, a small amount of mass is changed into energy, which means that *in every nuclear reaction, some mass is lost*.

Why?

Nuclear Binding Energies

In the 1930s, scientists discovered that the masses of nuclei are always less than the masses of their individual ***nucleons*** (protons and neutrons) added up.

For example, Helium-4’s nucleus has a mass of **4.00150 amu**, but its *calculated* mass (the mass of two protons + two neutrons) should be **4.03188 amu**.

So, why is Helium-4’s calculated mass **0.03038 amu** higher than its *actual* mass?

The answer (as weird as it is) is that protons and neutrons actually weigh a little bit less when they’re bound together in a nucleus than they do when they’re individually separate.

Nuclear Binding Energies

Why? Well, when neutrons and protons combine into a nucleus, a certain amount of energy is required to keep them together. This energy, called the ***nuclear binding energy***, is obtained by converting some of those protons' and neutrons' masses into energy, according to Einstein's equation ($E = mc^2$). The difference between a nucleus' calculated mass and its *actual* mass, then, is called its ***mass defect***.

In the case of Helium-4, we can calculate its ***nuclear binding energy*** from its ***mass defect***, as follows:

$$\Delta E \quad = \quad \Delta mc^2 \quad = \quad c^2 \Delta m$$

$$= \ (3.0 \times 10^8 \text{ m/s})^2 (0.03038 \text{ amu}) (1 \text{ g}/6.022 \times 10^{23} \text{ amu}) (1 \text{ kg}/1000 \text{ g})$$

$$= \ \underline{4.534 \times 10^{-12} \text{ kg}\bullet\text{m}^2/\text{s}^2}$$

Nuclear Binding Energies

Why? Well, when neutrons and protons combine into a nucleus, a certain amount of energy is required to keep them together. This energy, called the ***nuclear binding energy***, is obtained by converting some of those protons' and neutrons' masses into energy, according to Einstein's equation (**$E = mc^2$**). The difference between a nucleus' calculated mass and its *actual* mass, then, is called its ***mass defect***.

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$$= \underline{4.534 \times 10^{-12} \text{ J}}$$

Nuclear Binding Energies

This means that it takes **4.534 × 10^{-12} J** of energy to keep the neutrons and protons together in a single Helium-4 atom.

As a result, if we blew up a Helium-4 atom, we would get **4.534 × 10^{-12} J** out of it. This is how ***nuclear-energy*** reactors get their power, except that they use heavier elements than **^{4}He**.

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$$= \ \underline{-\,4.534 \times 10^{-12} \text{ J}}$$

To Summarize

- The difference between what a nucleus *should* weigh (based on the total mass its neutrons and protons) and what it *actually* weighs is called its ***mass defect*** (***Δm***).

- This ***mass defect*** arises from the fact that a very small amount of the protons and neutrons' combined masses in a nucleus gets converted into energy (***ΔE***), to hold that nucleus together. This energy is called the atom's ***nuclear binding energy***.

- Thus, if you can figure out ***Δm***, then you can use the equation $\Delta E = \Delta m c^2$ to calculate the ***nuclear binding energy***. ***Δm***, by the way, must be in **kg** (the SI unit of mass).

To Summarize

- When comparing elements' ***nuclear binding energies*** to their total numbers of protons and neutrons, **^{56}Fe** happens to be the most stable nucleus out of all the elements.

Questions

A proton's mass is 1.00728 amu. A neutron's mass is 1.00867 amu.

- What is the ***mass defect*** of a ^{178}Au nucleus, if a ^{178}Au nucleus weighs 177.9760 amu?

- What is the binding energy (in J) of a ^{178}Au nucleus?