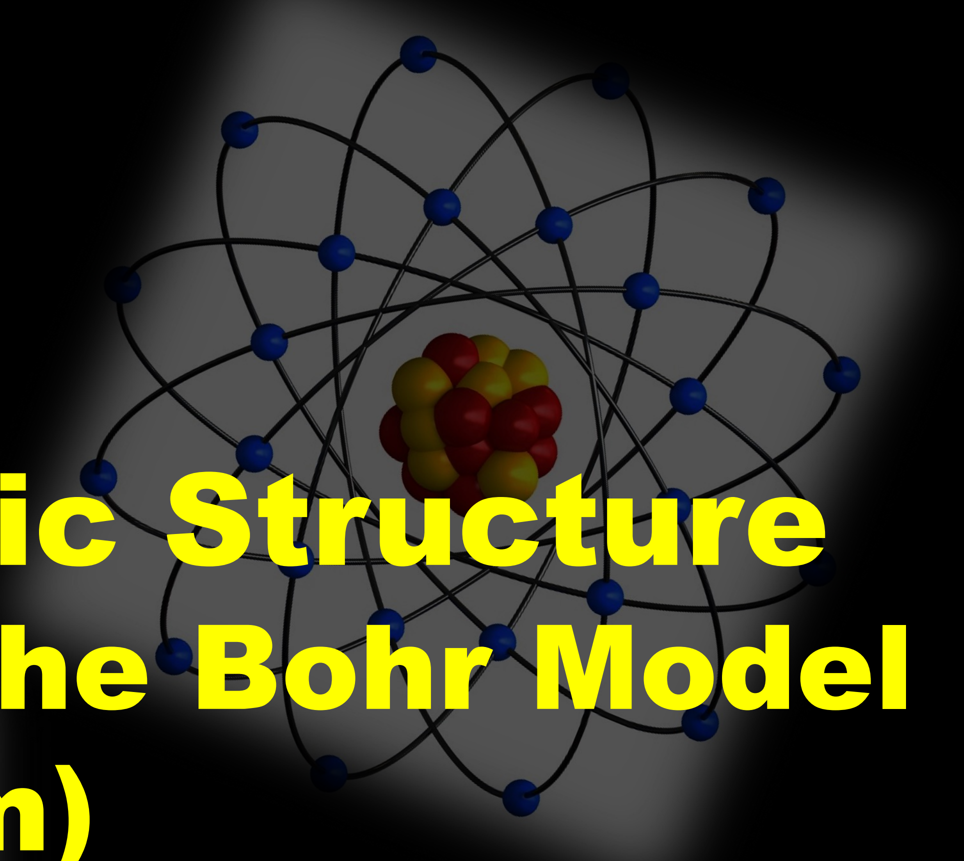
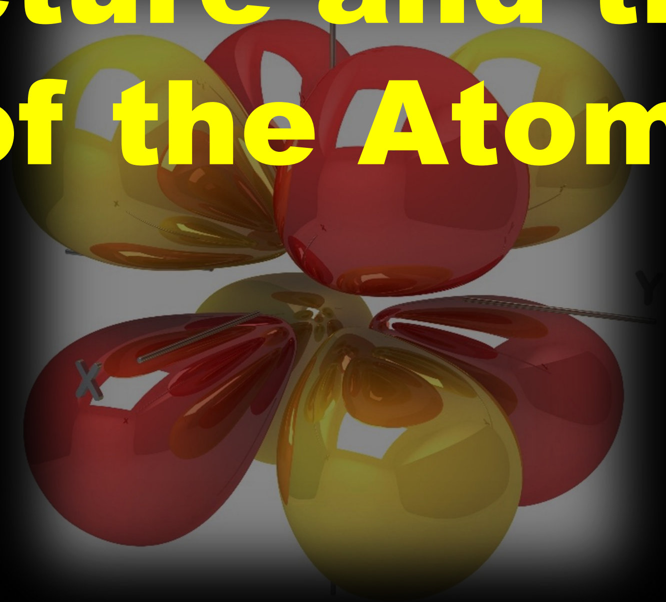


Atomic and Electronic Structure (Atomic Structure and the Bohr Model of the Atom)



Atomic (Chemical) Symbols

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

Atomic # (# of protons) $\rightarrow$ $^{24}_{12}\mathrm{Mg}$

The ***atomic number*** corresponds to the box in which the element appears on the periodic table. For example, Mg appears in box 12 on the periodic table, so its ***atomic number*** is 12. The ***atomic number*** is always the same 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. The number of protons in a given element never changes.

PERIODIC TABLE OF THE ELEMENTS

atomic number → 26
chemical symbol → Fe
name → Iron
atomic mass → 55.845
+3 +2 ← oxidation states (most common)
☢ radioactive elements

	1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIIIB	9 VIIIB	10 VIIIB	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
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 Neodynium 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 · alkaline earth metals · metalloids · halogens · lanthanoids

alcali metals · transition metals · post-transition metals · noble gases · actinoids

Mass Number

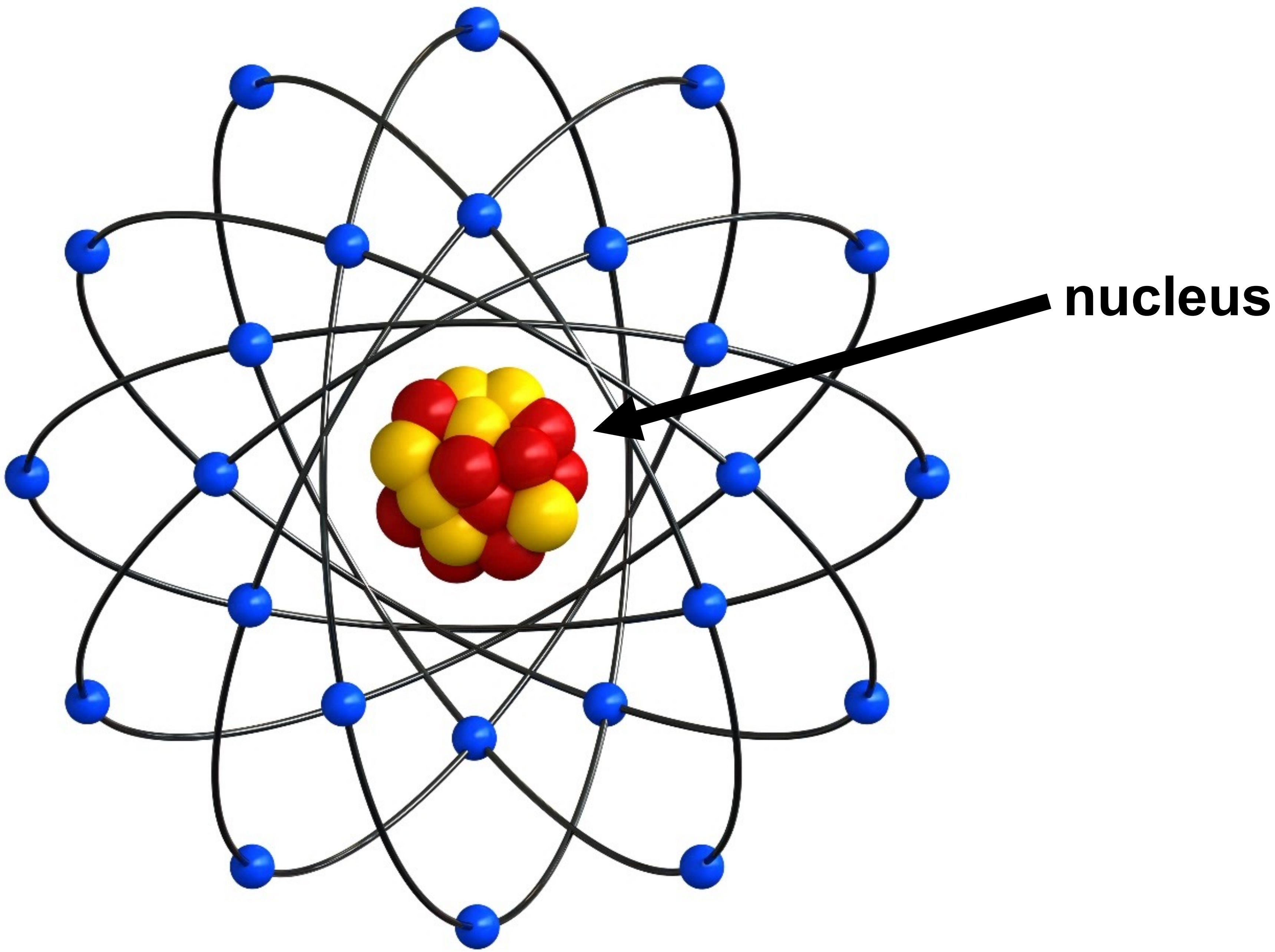
An atom's mass is equal to the combined masses of its ***protons*** and ***neutrons***, which are located in its nucleus.

Mass #

Atomic # (# of protons)

$$^{24}_{12}\mathrm{Mg}$$

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


nucleus

Mass Number

An atom's mass is equal to the combined masses of its ***protons*** and ***neutrons***, which are located in its nucleus.

Mass #

Atomic # (# of protons)

$$^{24}_{12}\mathbf{Mg}$$

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

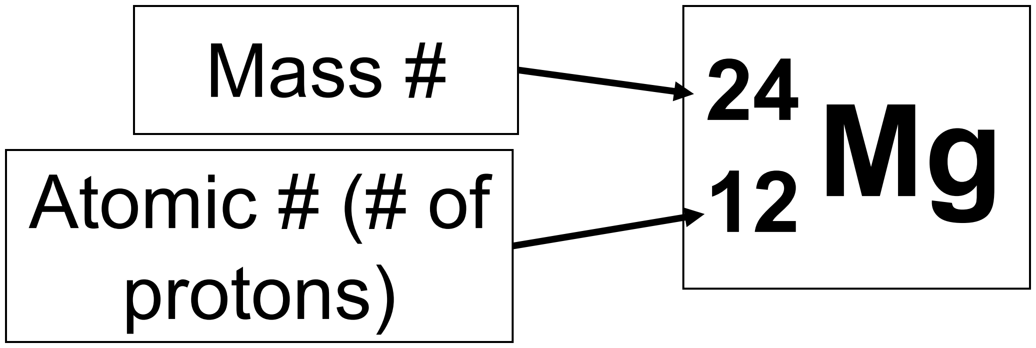
So how can you calculate an atom's number of ***neutrons?***

# of neutrons = Mass # – # of protons

= Mass # – atomic #

Mass Number

The ***mass number*** <u>CAN</u> change for different atoms with the same letter symbol.



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

As I mentioned before, all atoms of the same element have the <u>SAME number of protons and the SAME atomic number</u>. For example, all Mg atoms have 12 protons. You CANNOT change the number of protons in an element. If you do, you change the element.

Mass Number

The ***mass number*** CAN change for different atoms with the same letter symbol.

Mass # → 24

Atomic # (# of protons) → 12

$$^{24}_{12}\mathrm{Mg}$$

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

However, different atoms of the same element *can* have different numbers of neutrons. For example, a ^{24}Mg atom has 12 protons and 12 neutrons (12+12=24), but not all Mg atoms weigh 24. Some weigh 23, and some weigh 25. A ^{25}Mg atom has 12 protons and 13 neutrons (12+13=25), while a ^{23}Mg atom has 12 protons and 11 neutrons (12+11=23).

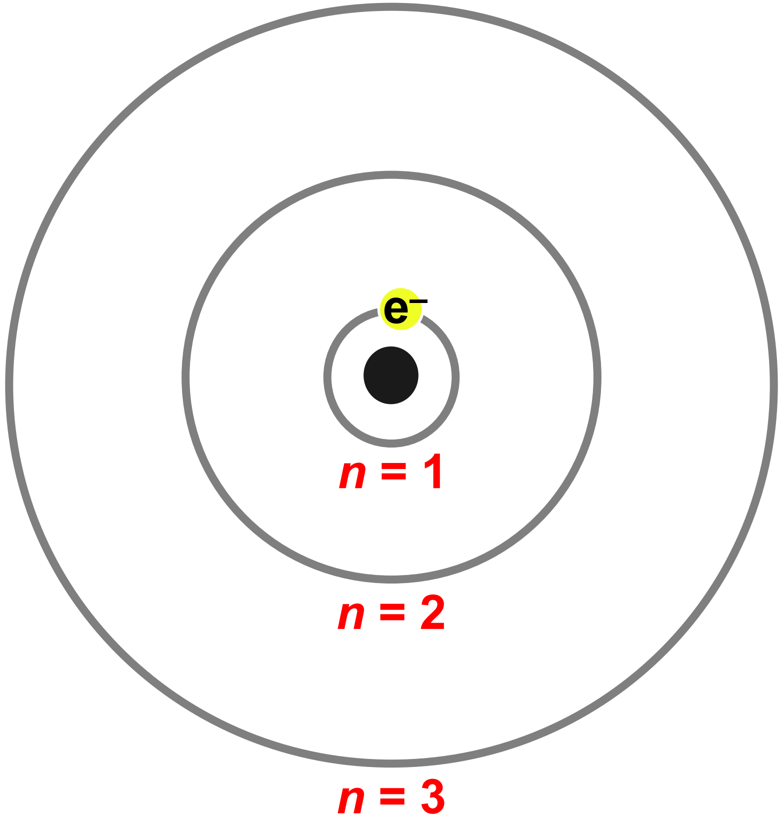
Atoms of the same element that have different ***mass numbers***, are called ***isotopes.***

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 <u>CAN</u> 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.

The Bohr Model of the Atom

In 1913, Niels Bohr and Ernest Rutherford created the first working model of a hydrogen atom.



What is Quantization?

When walking up a ramp, you can stand at any vertical position along the ramp. When walking up *stairs*, however, you cannot do that. You can only stand at certain vertical locations, where the steps are, not *in-between them*.

ramp

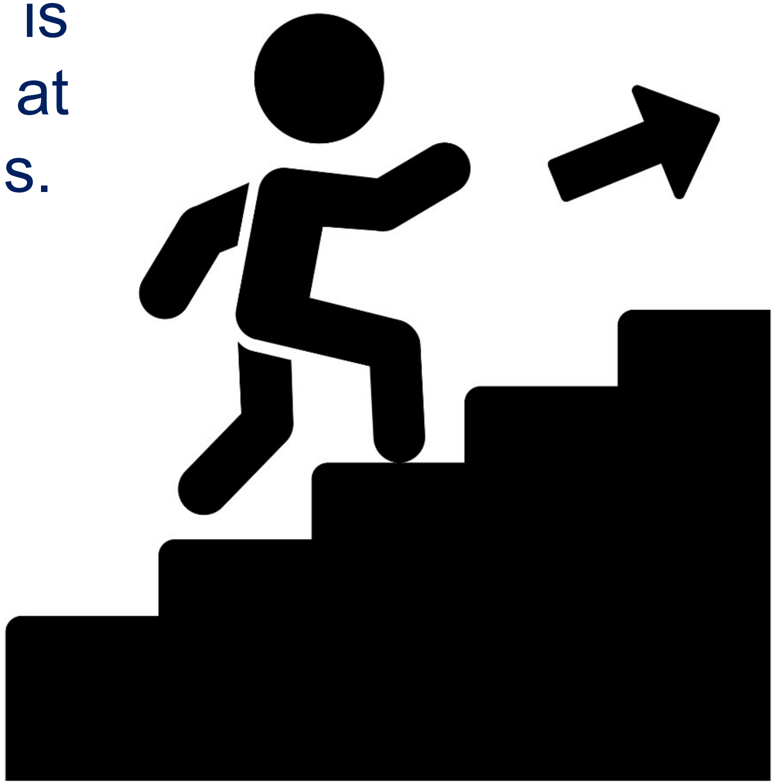
stairs

What is Quantization?

For instance, you cannot stand halfway between the first and second step on a set of stairs. Once again, you can only stand at *specific vertical positions* and not in-between those positions.

We say, then, that your position on a set of stairs is ***quantized***. That is, your position can only exist at certain distinct levels, but not in-between those levels.

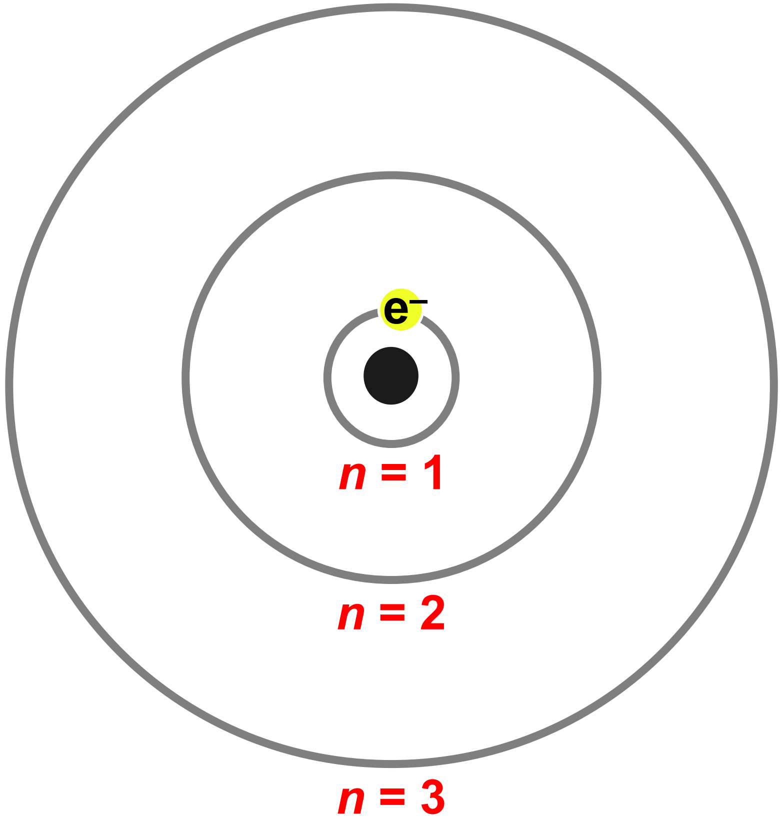
ramp



stairs

The Bohr Model of the Atom

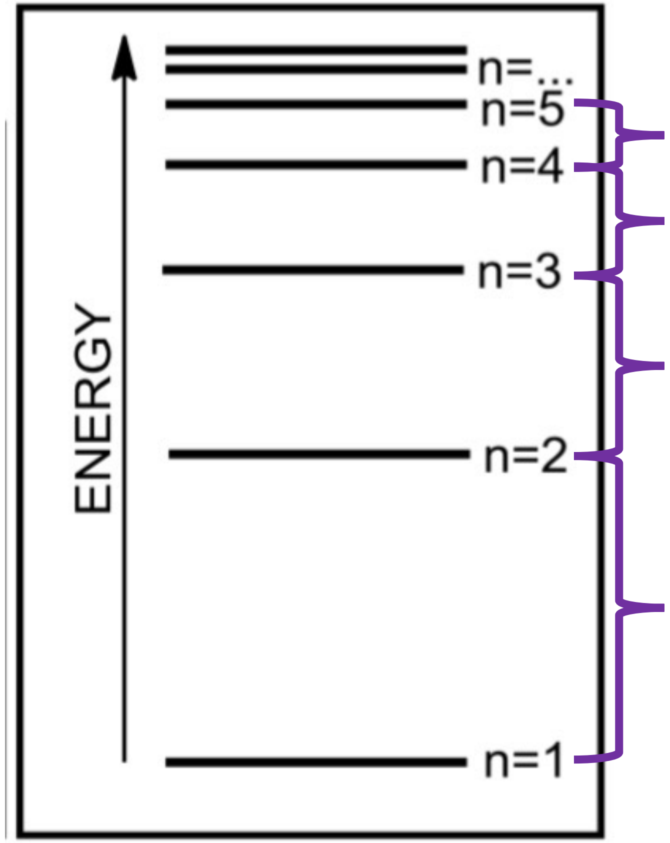
Similarly, the Bohr model for hydrogen (and by extension, all *other* elements' atoms) indicates that electrons can only exist at certain distances from the nucleus (***n* = 1**, **2**, **3**, **4**, etc.) and not anywhere in-between. Thus, electrons' distances from their nuclei are also ***quantized***. Hence, there is no ***n* = ½** shell.



The Bohr Model of the Atom

The distance between each energy shell and the *next* shell above it gets smaller as you get further away from the nucleus.

Thus, there is a larger energy difference between an electron that shifts from ***n*=2** down to ***n*=1** than (for example) an electron that shifts from ***n*=5** to ***n*=4**.



Questions

There are four isotopes of calcium. Their mass numbers are 40, 42, 43, and 44.

- Please write the complete chemical symbol (superscript and subscript) for each isotope.
- How many neutrons are in an atom of each isotope?
- How many neutrons, protons, and electrons are present in a $^{40}Ca^{2+}$ ion?

Give the chemical symbol, with its superscripted mass number, for:

- The element with 42 protons and 54 neutrons
- The ion with 15 protons, 16 neutrons, and 18 electrons
- The isotope of silver that has 61 neutrons

Questions

How many protons, neutrons and electrons does the $^{52}Cr^{3+}$ ion have?

A. 24, 52, 26
B. 52, 52, 55
C. 24 , 28, 26
D. 24, 28, 21
E. 24, 28, 25

Questions

The symbol of an ion is $^{238}_{92}X^{4+}$. Which of the following is true regarding the ion?

I. It has 92 protons.
II. It has 238 neutrons.
III. It has 88 electrons.
IV. The element X is uranium.

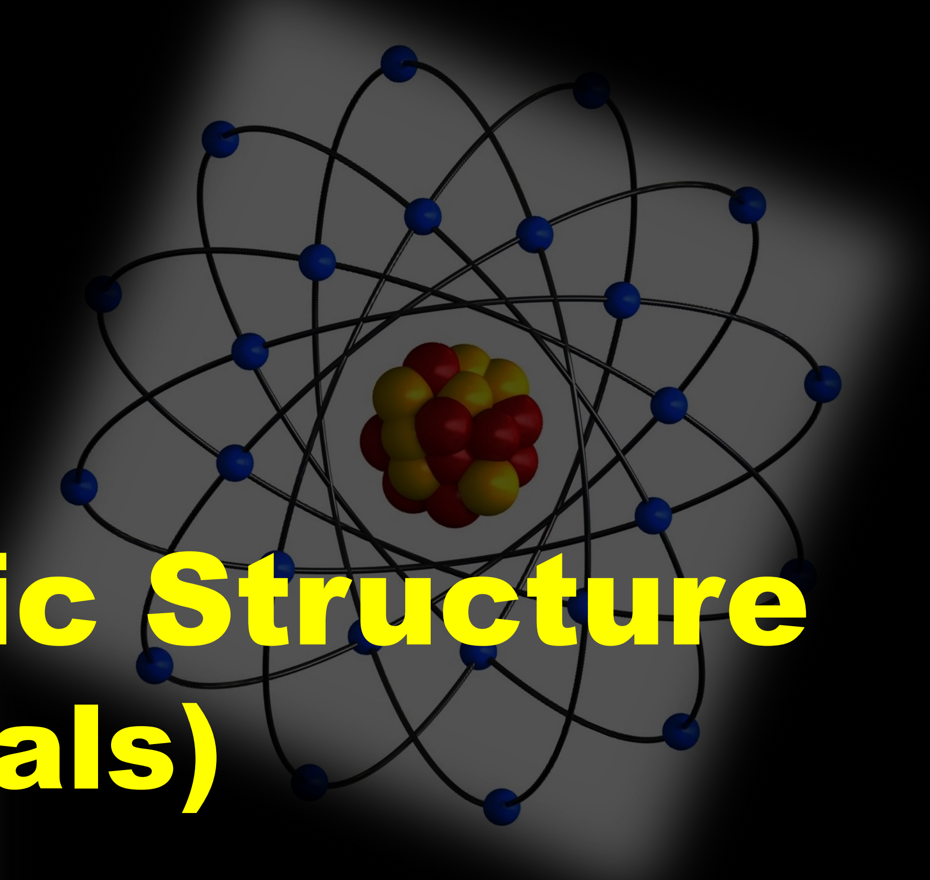
A. I and II only

B. I, III and IV only

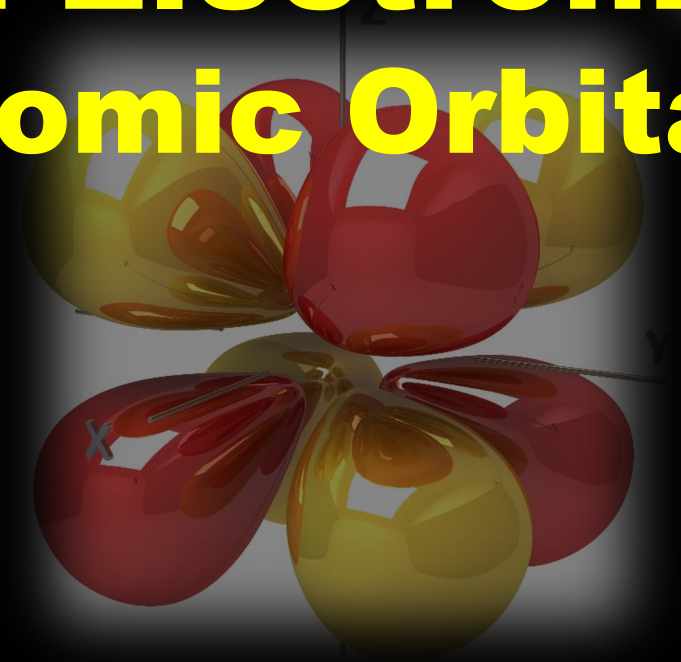
C. II and IV only

D. I and IV only

E. I, II, III, and IV

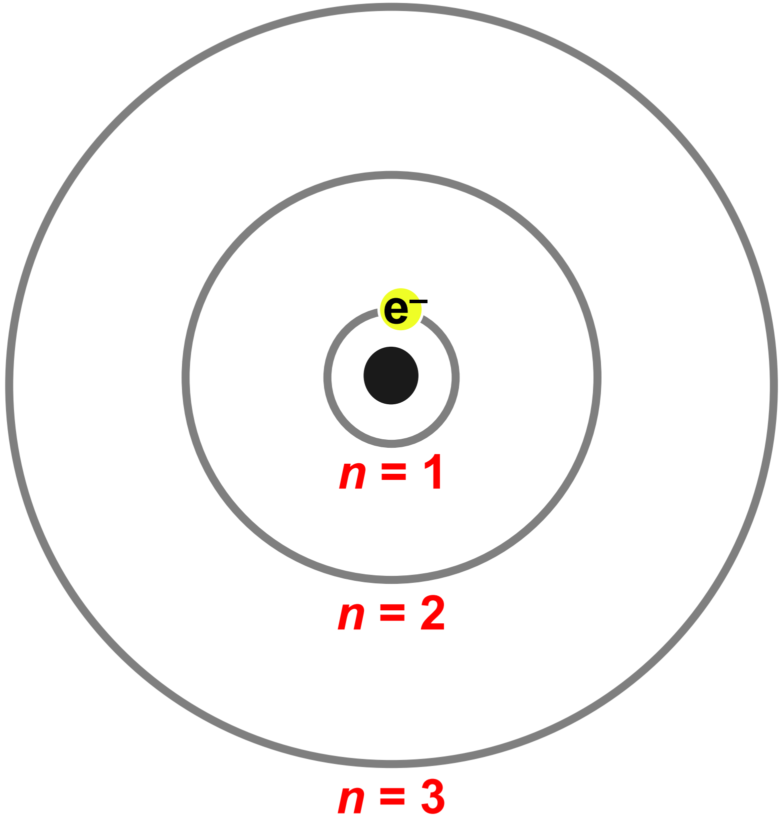


Atomic and Electronic Structure
(Atomic Orbitals)



Where Are an Atom's Electrons?

In an earlier video, I showed you the ***Bohr atomic model***, which proposes that electrons travel along two-dimensional orbits around their atom's nucleus, similar to the way that planets orbit the sun:



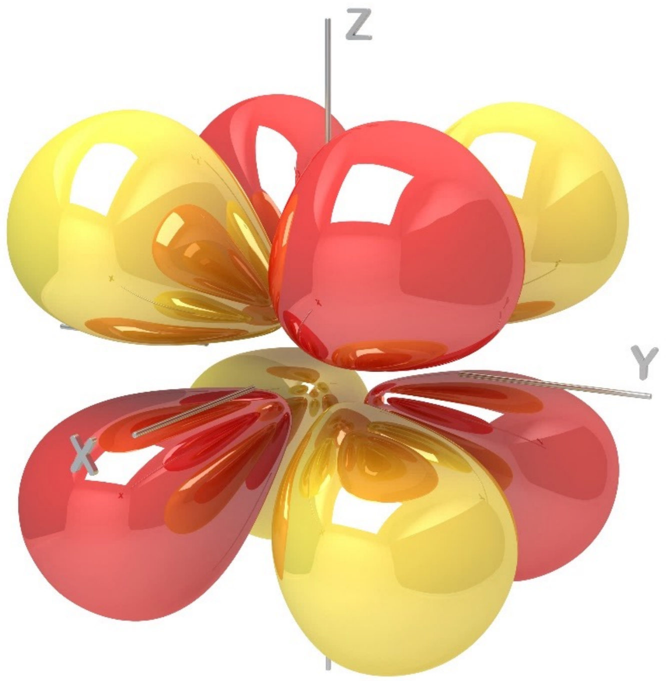
Where Are an Atom's Electrons?

In reality, electrons do _not_ travel around the nucleus in circular, two-dimensional orbits. Instead, they zoom throughout various three-dimensional regions of space called ***orbitals*** (sometimes called ***atomic orbitals*** for single atoms, or ***molecular orbitals*** for molecules).

These ***orbitals'*** shapes are not simple 2D circles. In fact, they have a variety of different shapes, which a physicist named Erwin Schrödinger identified through complex calculation. His work was accomplished by taking the ***Bohr model*** to the "next level," so to speak.

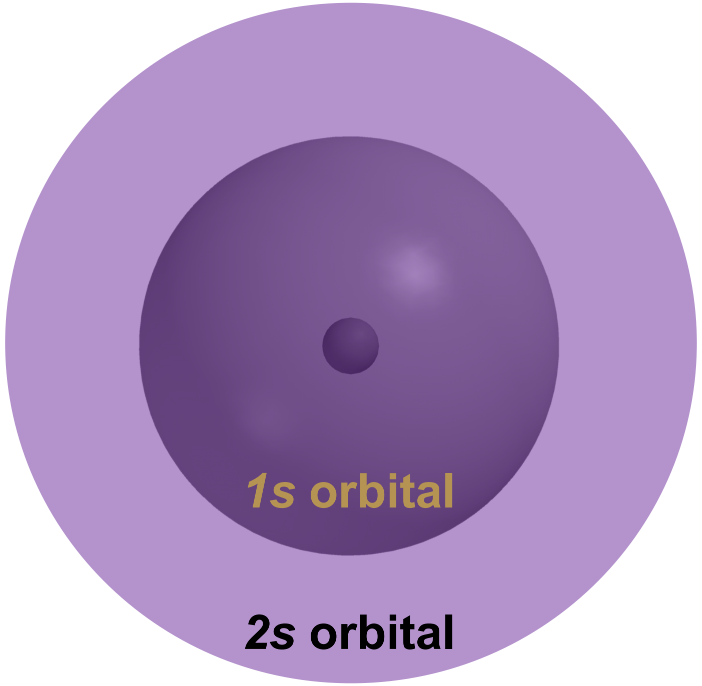
Four Different Kinds of Orbitals

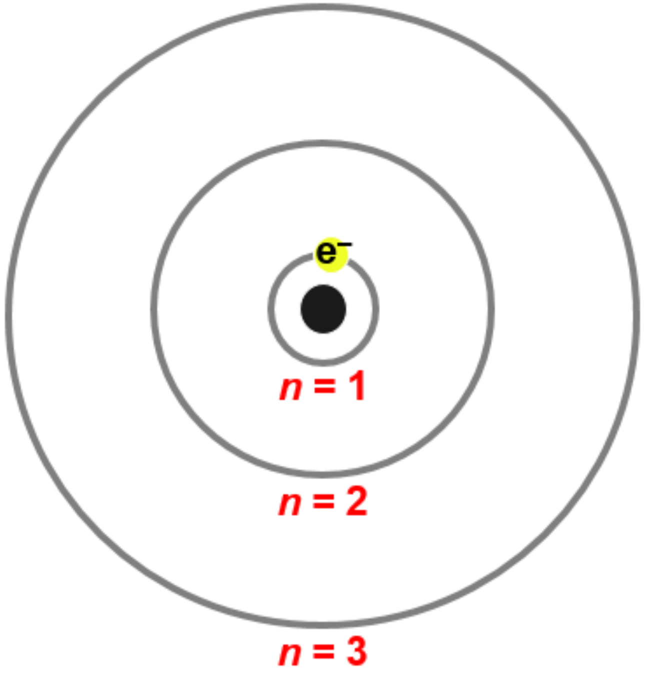
There are four different kinds of orbitals: ***s-orbitals***, ***p-orbitals***, ***d-orbitals***, and ***f-orbitals***. Each of these orbitals, or sets of orbitals, are simply regions of space around an atom's nucleus, where its electrons "live." ***Orbitals***, then, are kind of like empty parking places for electrons.



S-Orbitals

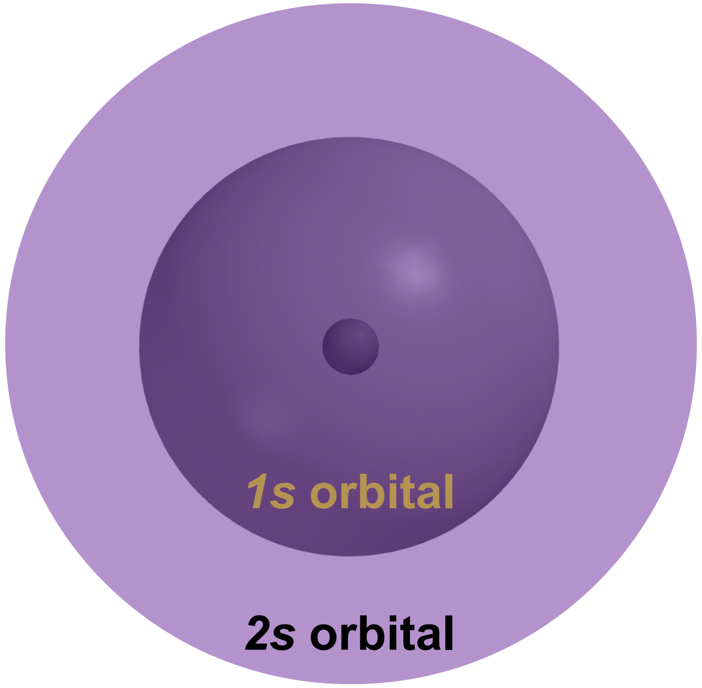
S-orbitals are shaped like spheres. They're very much like the 2D circles in the ***Bohr model***, except that they aren't 2D circles. They're 3D spheres. A ***1s*** orbital, for instance, surrounds the nucleus at an ***n*=1** distance. A ***2s*** orbital has the same shape. It's just further from the nucleus and surrounds the ***1s*** orbital.





S-Orbitals

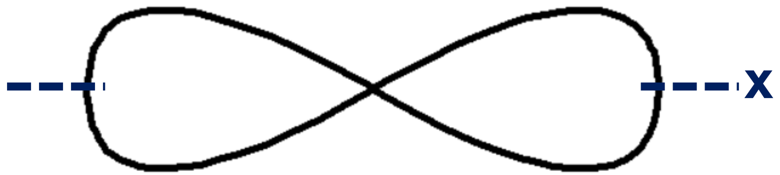
So, lower-energy (lower-***n***) orbitals and their electrons are buried inside higher-energy (lower-***n***) orbitals, kind of like nesting dolls. This same thing applies to ***p-orbitals***, ***d-orbitals***, and ***f-orbitals***.

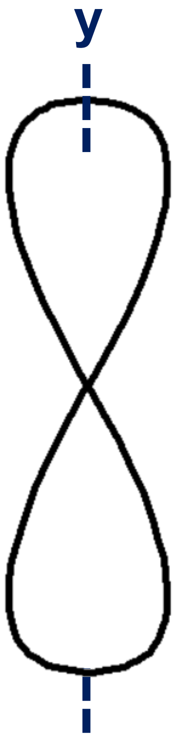


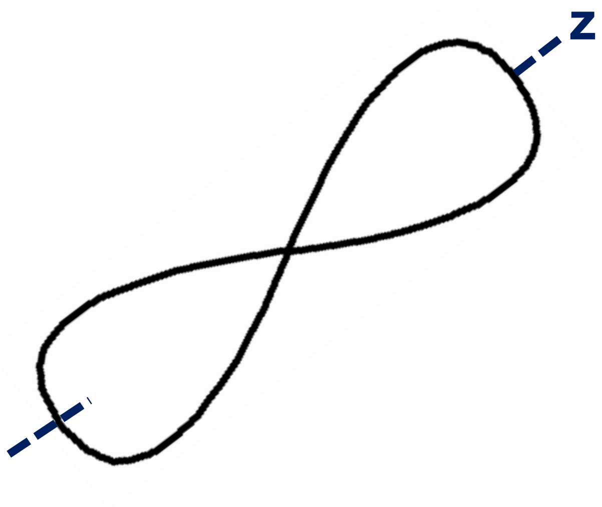
nesting dolls

P-Orbitals

P-orbitals are shaped like dumbbells. Each energy level, however, has *three* ***p-orbitals***: one straddling the x-axis, one straddling the y-axis, and one straddling the z-axis. So when we talk about a ***2p*** shell, we're talking about three separate ***2p*** orbitals: a $\boldsymbol{p_x}$, a $\boldsymbol{p_y}$, and a $\boldsymbol{p_z}$.

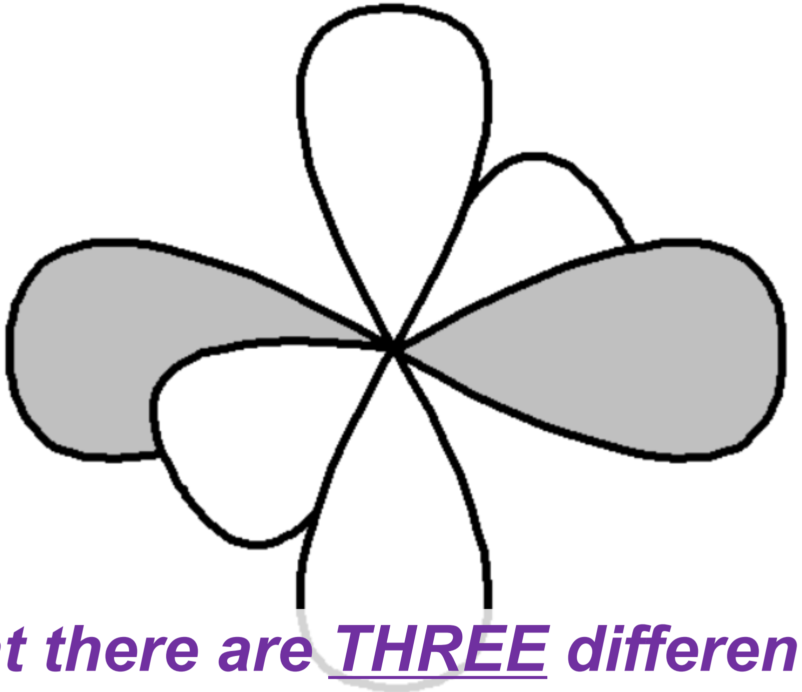






P-Orbitals

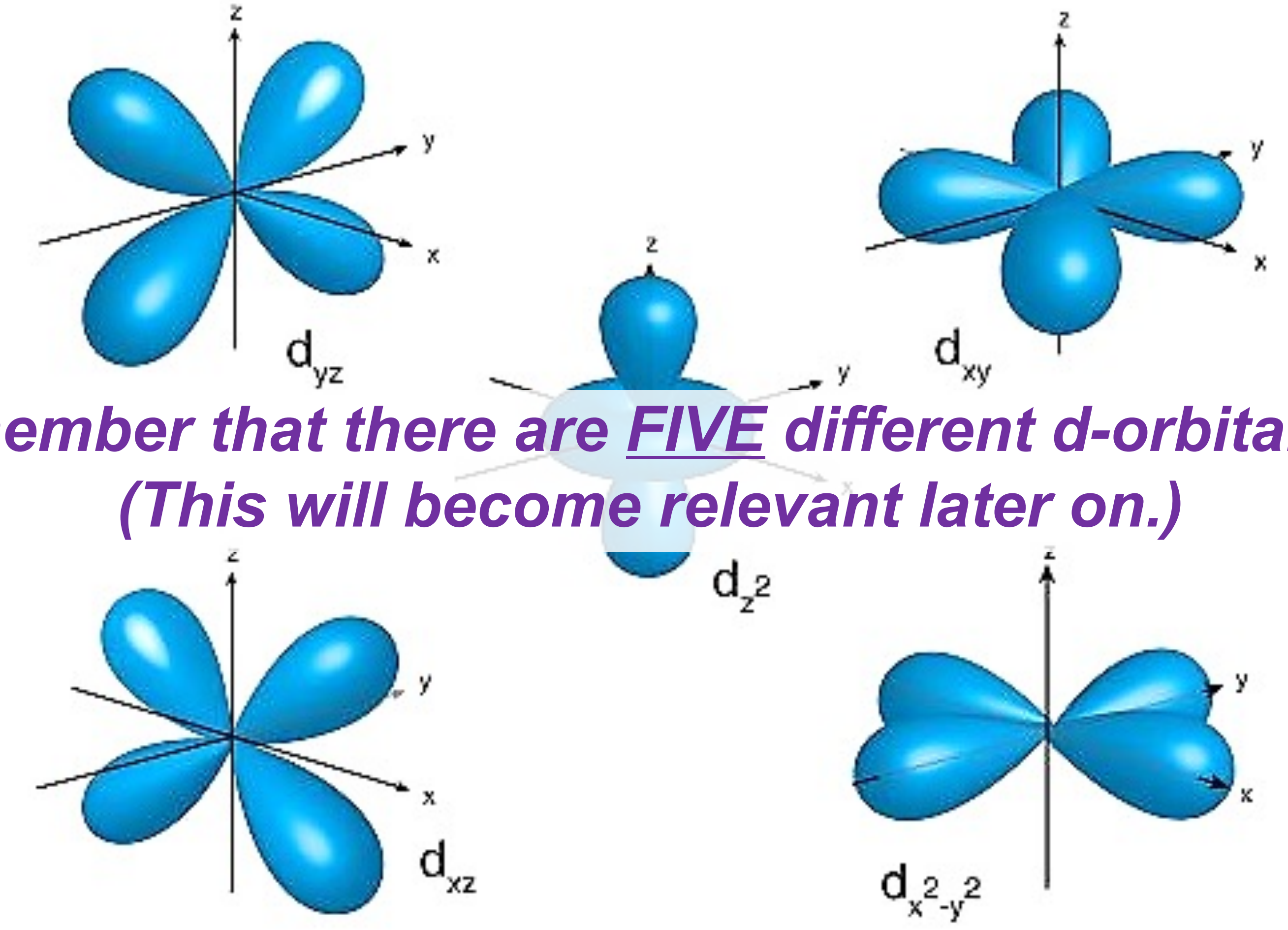
P-orbitals are shaped like dumbbells. Each energy level, however, has *three* ***p-orbitals***: one straddling the x-axis, one straddling the y-axis, and one straddling the z-axis. So when we talk about a ***2p*** shell, we're talking about three separate ***2p*** orbitals: a $\boldsymbol{p_x}$, a $\boldsymbol{p_y}$, and a $\boldsymbol{p_z}$.



Please remember that there are THREE different p-orbitals per shell.
(This will become relevant later on.)

D-Orbitals

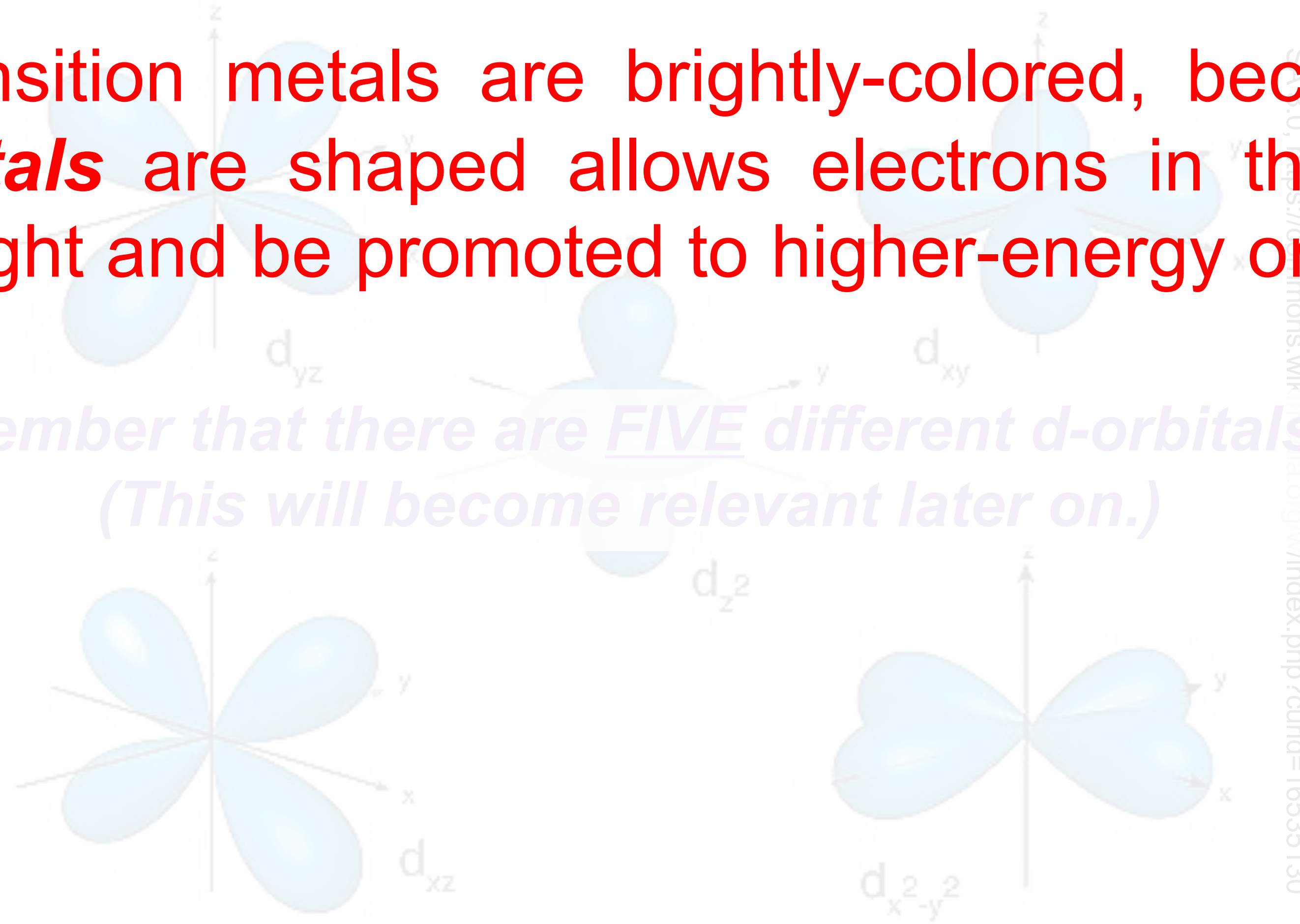
D-orbitals are shaped like four-leaf clovers, except for the fifth one, which looks like a donut/dumbbell combination.



Please remember that there are <u>FIVE</u> different d-orbitals per shell.
(This will become relevant later on.)

D-Orbitals

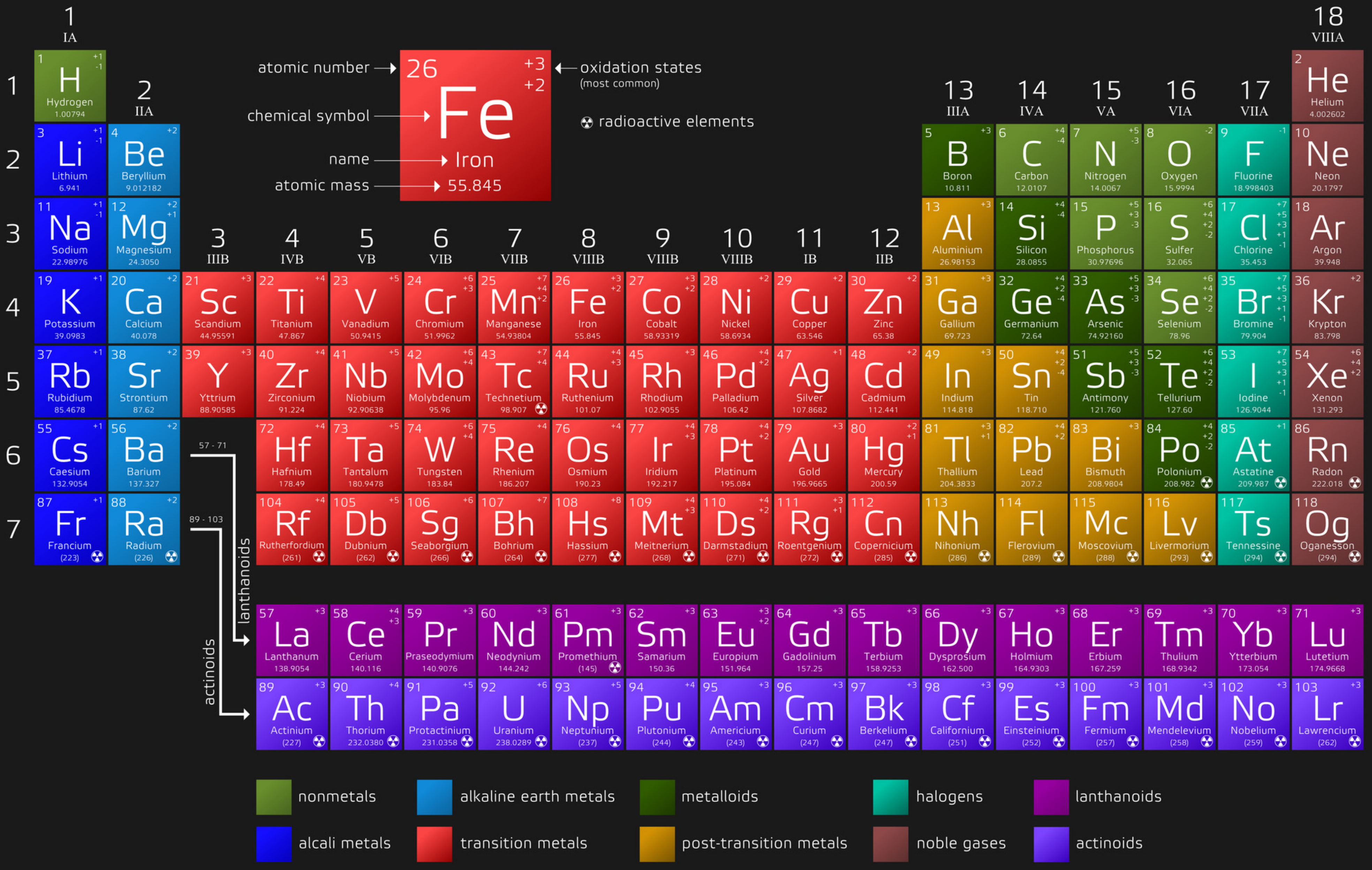
D-orbitals are shaped like four-leaf clovers, except for the fifth one which looks like a donut/dumbbell combination.



<u>**EXAM TIP**</u>**:** Transition metals are brightly-colored, because the funky way that ***d-orbitals*** are shaped allows electrons in the ***d-orbitals*** to absorb colored light and be promoted to higher-energy orbitals within the ***d-shell***.

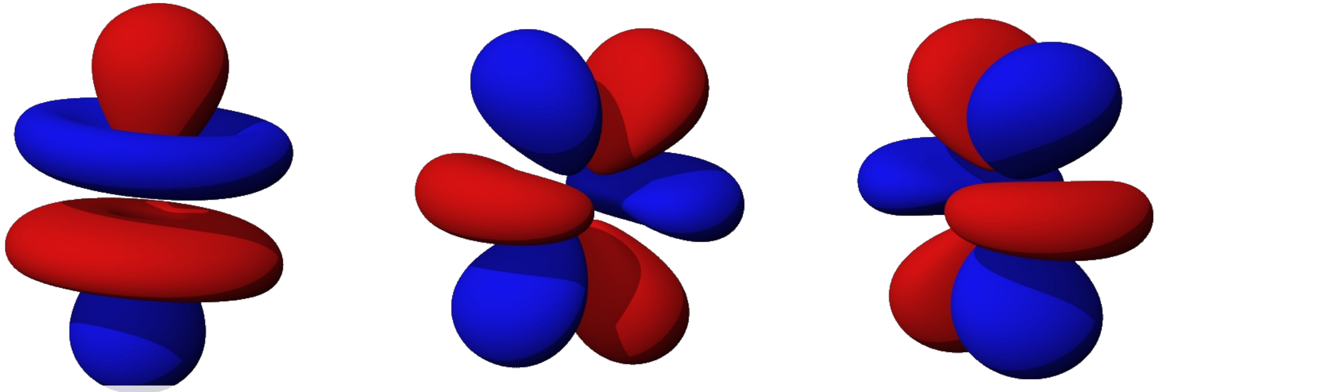
Please remember that there are <u>FIVE</u> different d-orbitals per shell.
(This will become relevant later on.)

PERIODIC TABLE OF THE ELEMENTS

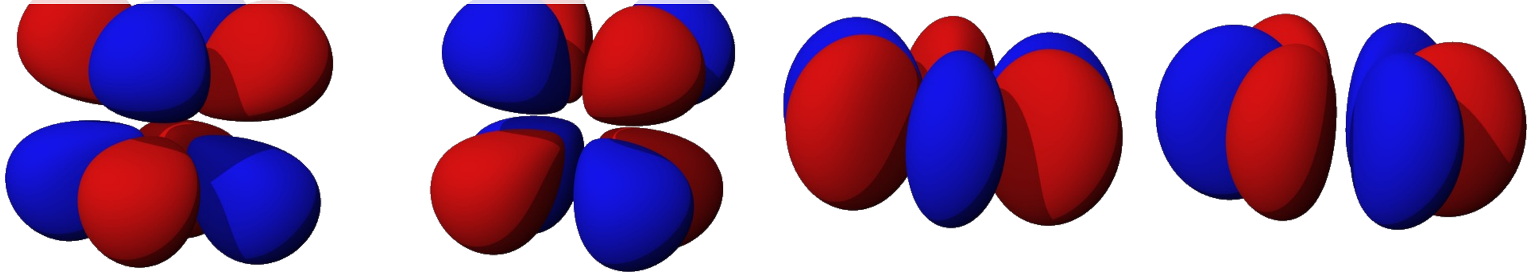


F-Orbitals

F-orbitals have seven different shapes, as seen here:

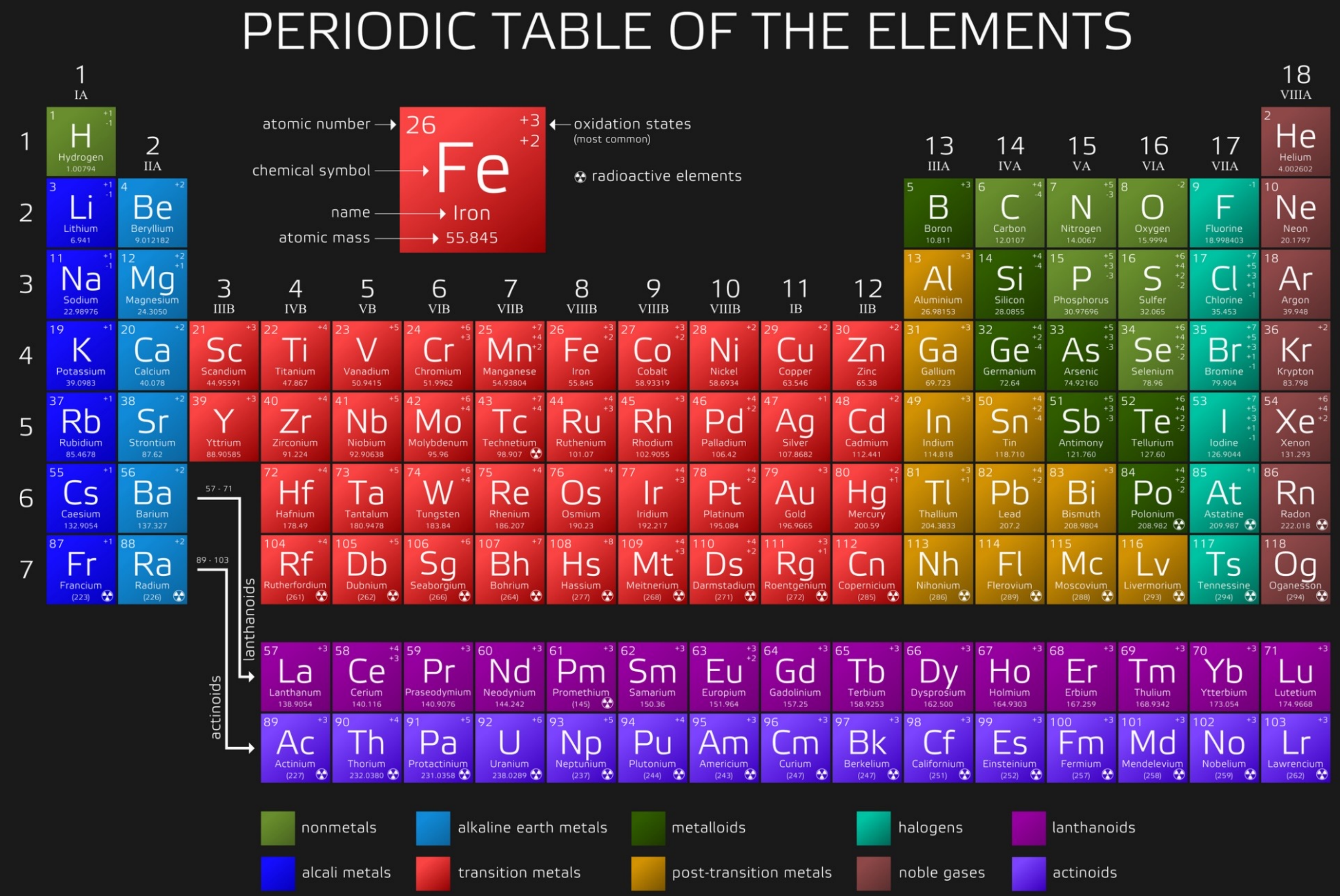


Please remember that there are <u>SEVEN</u> different f-orbitals per shell.
(Again, this will become relevant later on.)



m = +2 m = -2 m = +3 m = -3

We’re used to seeing the periodic table shown like this:



Each orbital can hold up to two electrons.

***s*-block**
(2 wide)

***f*-block**
(14 columns wide)

***d*-block**
(10 columns wide)

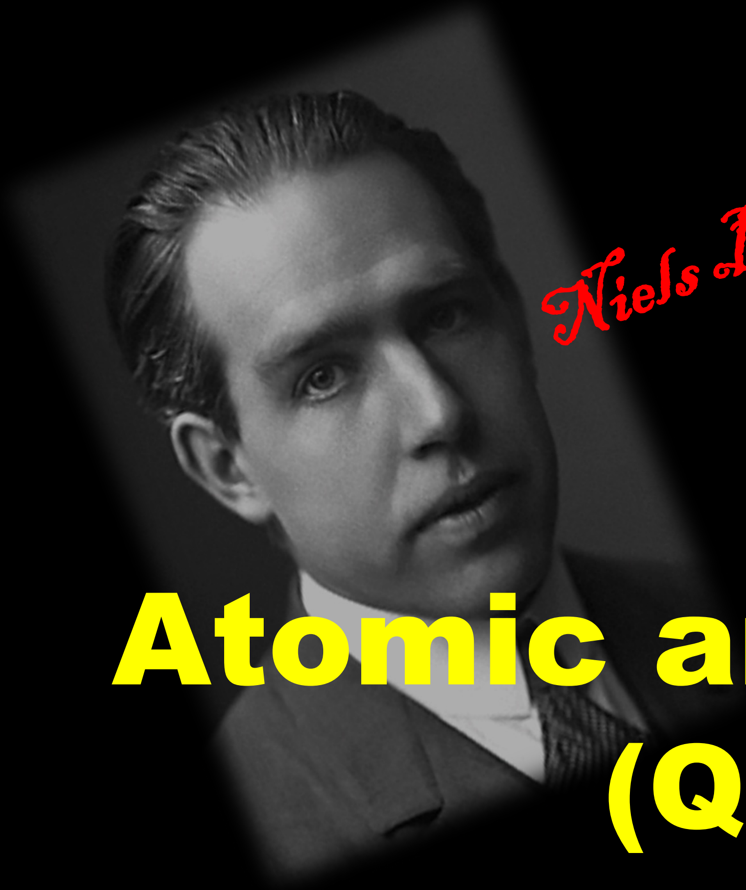
***p*-block**
(6 columns wide)

ONE ***s*-orbital per shell**

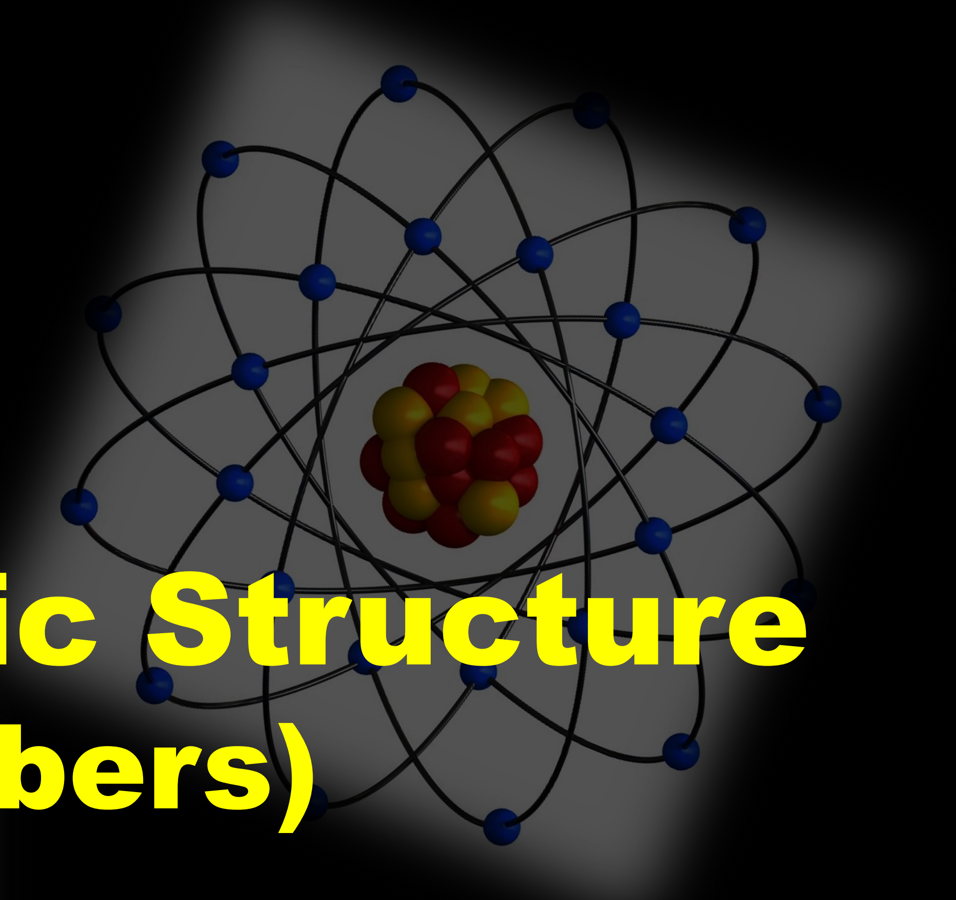
SEVEN ***f*-orbitals per shell**

FIVE ***d*-orbitals per shell**

THREE ***p*-orbitals per shell**



Atomic and Electronic Structure (Quantum Numbers)



Quantum Numbers

In an earlier video, we learned that electrons “live” (reside) in regions of space around their atom’s nucleus. These regions of space are called ***orbitals***. We also learned that there are four different kinds of orbitals: ***s-orbitals***, ***p-orbitals***, ***d-orbitals***, and ***f-orbitals***.

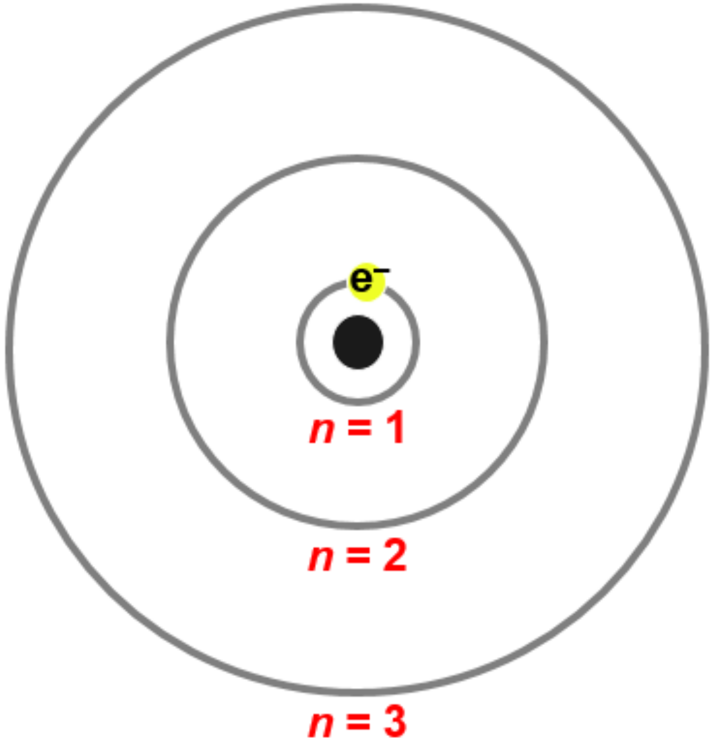
Scientists have a numerical system for describing an electron’s location: that is, how far it is from its nucleus and in what kind of orbital it resides. This system involves four different numbers, called ***quantum numbers***. These numbers are very much like an electron’s address.

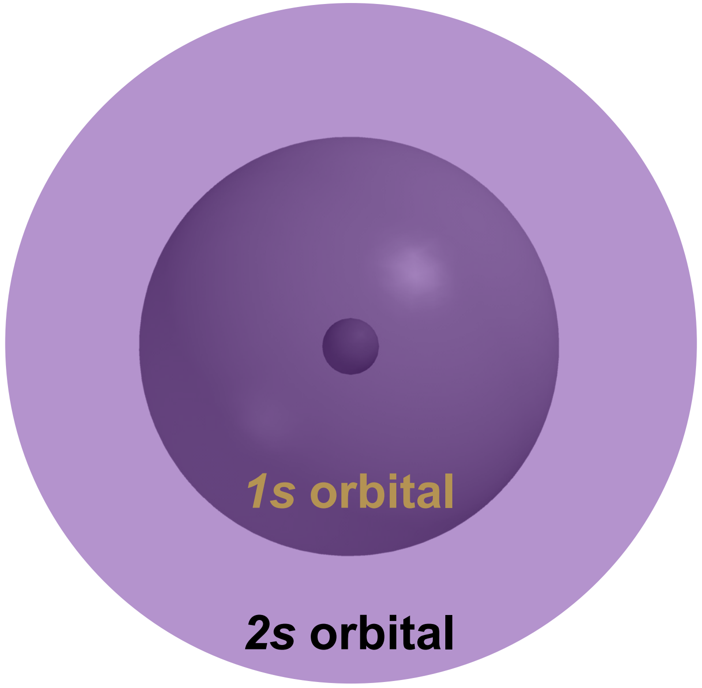
For instance, the address **1234 North Broadway, Scooby Doo City, New Jersey** tells us a building or a house’s location: its state, city, street, and the house number <u>on</u> that street, which allows us to zero in on its exact position. ***Quantum numbers*** are similar, just for electrons.

The Principle Quantum Number

There are four ***quantum numbers***. I'll now describe each one.

n **(the principle quantum #):** This number can be any integer **1**, **2**, **3**, **4**, etc. This number tells us which shell the electron is in, or (in other words) how far the electron is from the nucleus.





The Azimuthal Quantum Number

l* (the azimuthal quantum #):** This tells us what kind of orbital (s***, ***p***, ***d***, or ***f***) the electron is in.

l	**0**	**1**	**2**	**3**
orbital	***s***	***p***	***d***	***f***

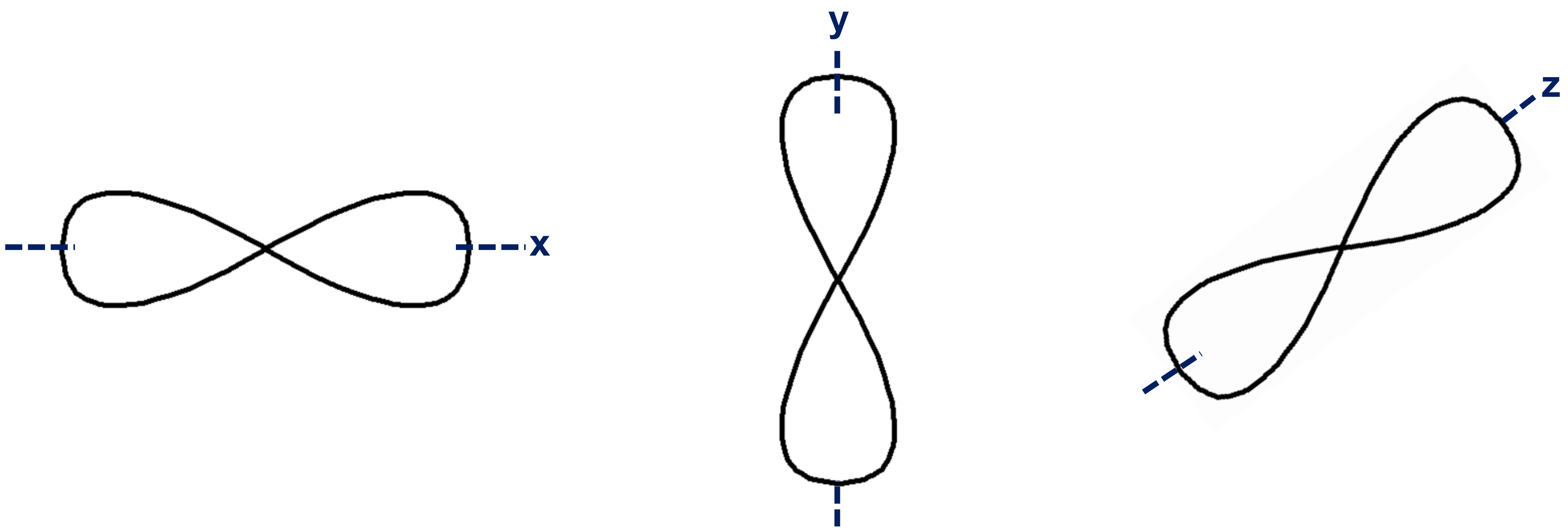
The Magnetic Quantum Number

***m_l* (the magnetic quantum #):** This number can be any integer from $-l$ to $+l$, including zero.

This number tells us which one of the ***p-orbitals***, ***d-orbitals***, or ***f-orbitals***, the electron is in.

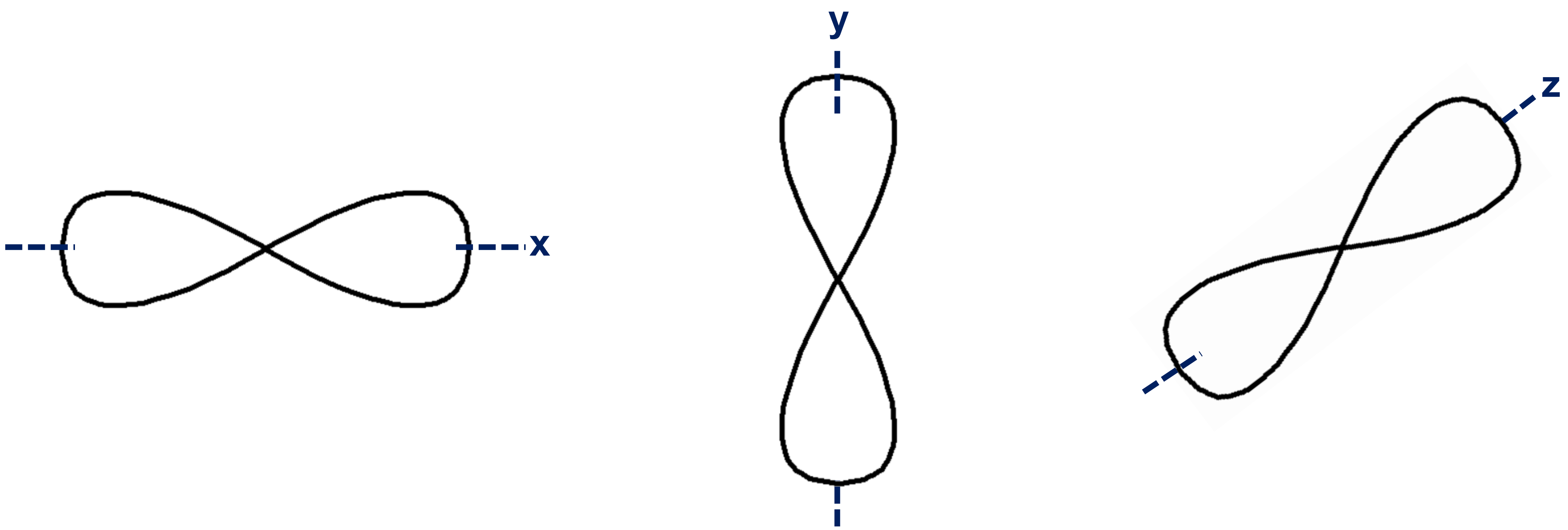
The Magnetic Quantum Number

For example, each ***p-orbital*** really represents three different ***p-orbitals***: one on the x-axis, one on the y-axis, and one on the z-axis. For a ***p-orbital*** electron, l = 1. Thus, m_l = -1, 0, or +1.



The Magnetic Quantum Number

What do you think the numbers -1, 0, and +1 mean? One of these numbers refers to the ***p**$_x$* orbital, another to the ***p**$_y$* orbital, and the other to the ***p**$_z$* orbital.



Each orbital can hold up to two electrons. A *d*-shell, for instance, has five different shapes, or kinds, or *d*-orbitals. Each one can hold up to two electrons, which amounts to 10 total electrons per *d*-shell. That's why the *d*-block is 10 columns wide.

***s*-block**

$l = 0$

$m_l = 0$

***f*-block**
(14 columns wide)

$l = 3$

$m_l = -3, -2, -1, 0, +1, +2, +3$

***d*-block**
(10 columns wide)

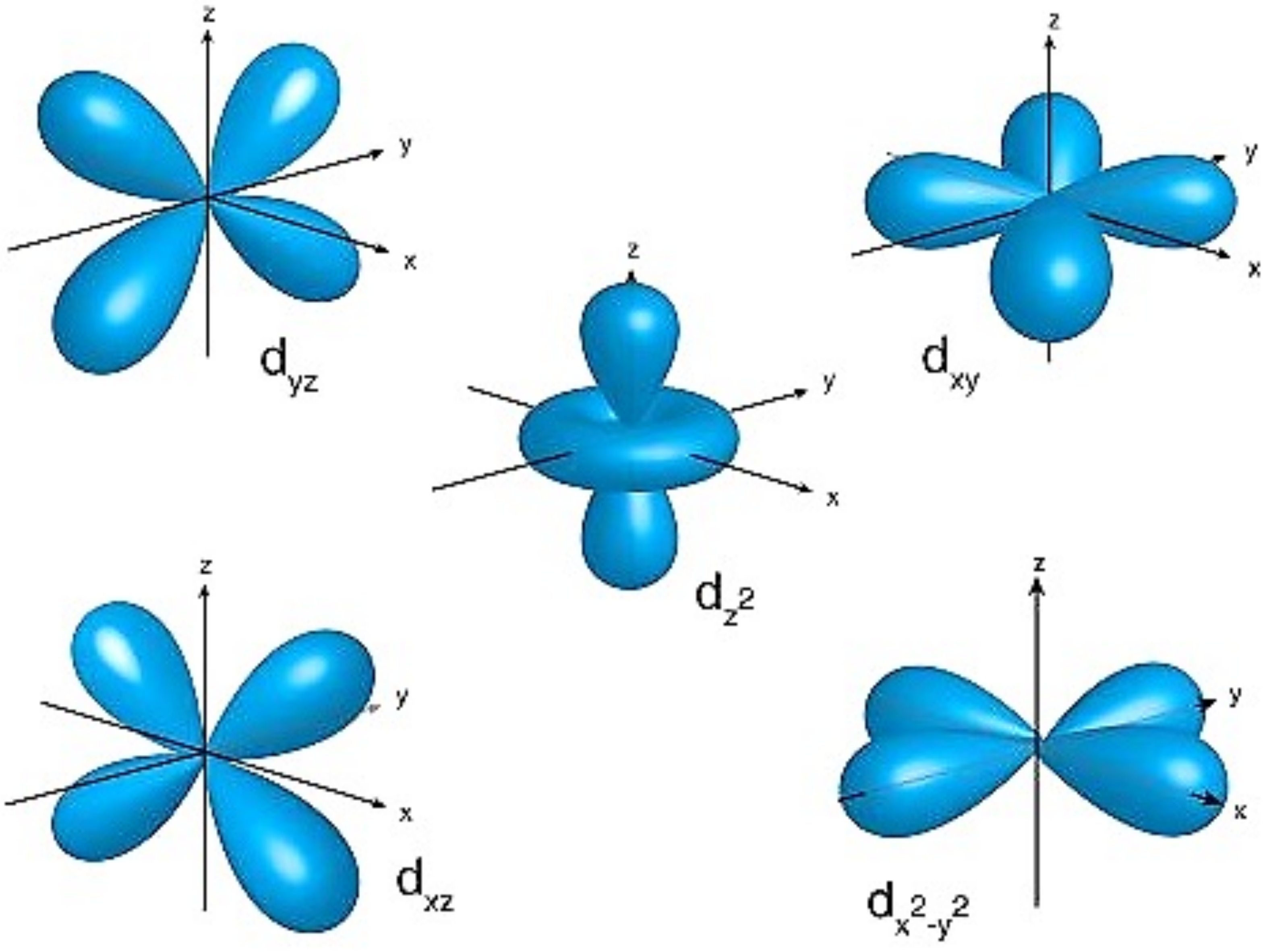
$l = 2$

$m_l = -2, -1, 0, +1, +2$

***p*-block**
(6 columns wide)

$l = 1$

$m_l = -1, 0, +1$

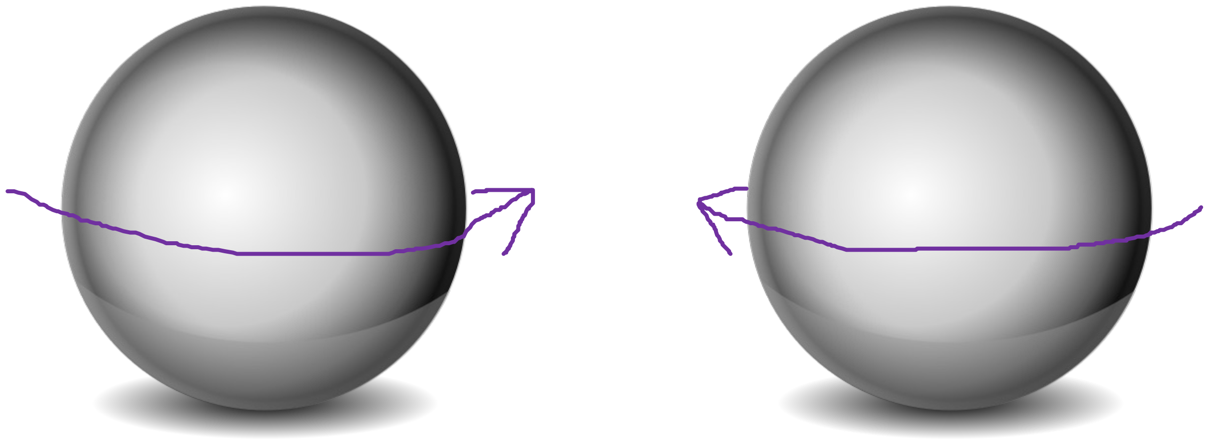


z
y
x
d_{yz}
z
y
x
d_{z^2}
z
y
x
d_{xy}
z
y
x
d_{xz}
z
y
x
$d_{x^2-y^2}$

The Spin Number

m_s* (the spin quantum #):** The spin number is the last ***quantum number and is equal to ±½. That is, for any two electrons occupying the same orbital, one is assigned a +½ spin and the other a –½ spin.

If, for example, two electrons had identical quantum numbers ***n***, ***l***, and ***m_l***, then we would assign one electron a +½ spin and the other a –½ spin. This would mean that these two electrons, both shooting around inside the same orbital, are revolving around their individual axes in opposite directions from each other, like this:



To Summarize . . .

#	name	what it is	range
n	principal	Energy level (distance from nucleus)	[1…infinity]
l	azimuthal	what type of orbital it is l = 0 (***s***) l = 1 (***p***) l = 2 (***d***) l = 3 (***f***)	0, 1, 2, or 3
m_l	magnetic	which ***p***, ***d***, or ***f*** orbital you have (orientation in space)	[$-l$ …$+l$]
m_s	spin	The electron's spin (up or down)	+1/2 or -1/2

Questions

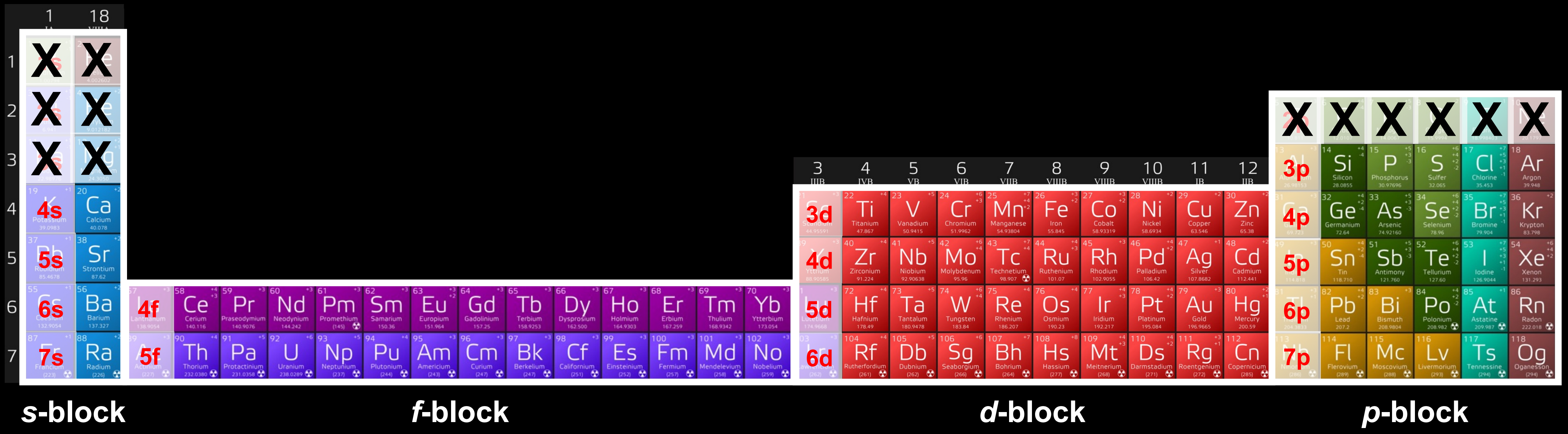
Which of the following sets of quantum numbers represents a valence electron in phosphorous?

A. $n = 4$, $l = 2$, $m_l = -2$, $m_s = -½$

B. $n = 2$, $l = 0$, $m_l = 0$, $m_s = +½$

C. $n = 3$, $l = 3$, $m_l = -2$, $m_s = -½$

D. $n = 3$, $l = 1$, $m_l = -1$, $m_s = +½$



4s
5s
6s
7s
4f
5f
3d
4d
5d
6d
3p
4p
5p
6p
7p
s-block
f-block
d-block
p-block

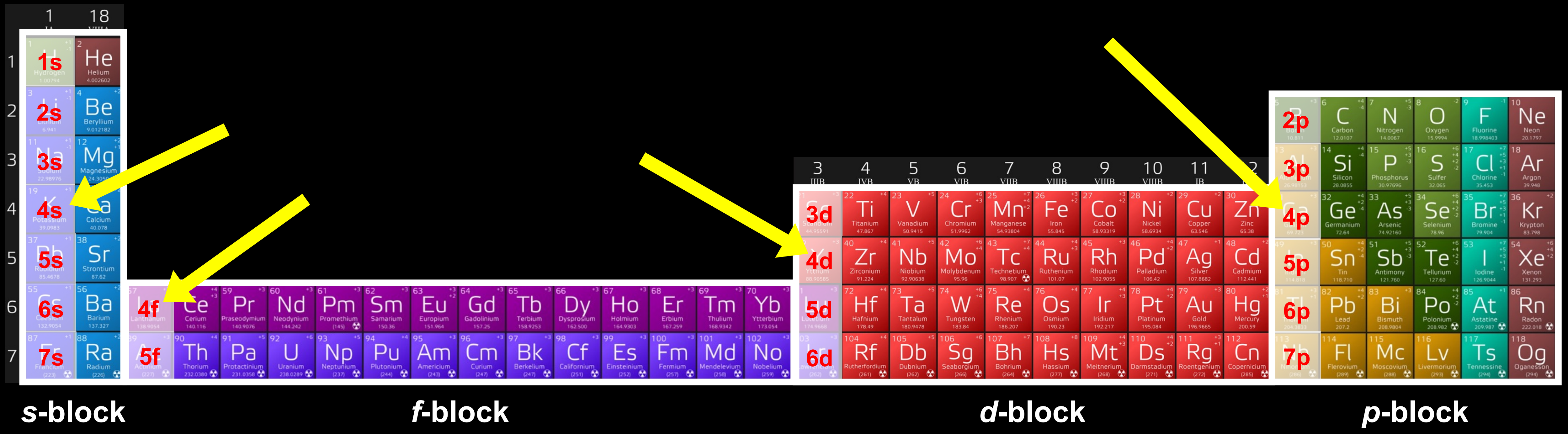
Questions

Please answer each of the following:

a. For $n = 4$, what are the possible values of l?

b. For $l = 2$, what are the possible values of m_l?

c. If m_l is 2, what are the possible values for l?


1s
2s
3s
4s
5s
6s
7s
4f
5f
3d
4d
5d
6d
2p
3p
4p
5p
6p
7p
s-block
f-block
d-block
p-block

Questions

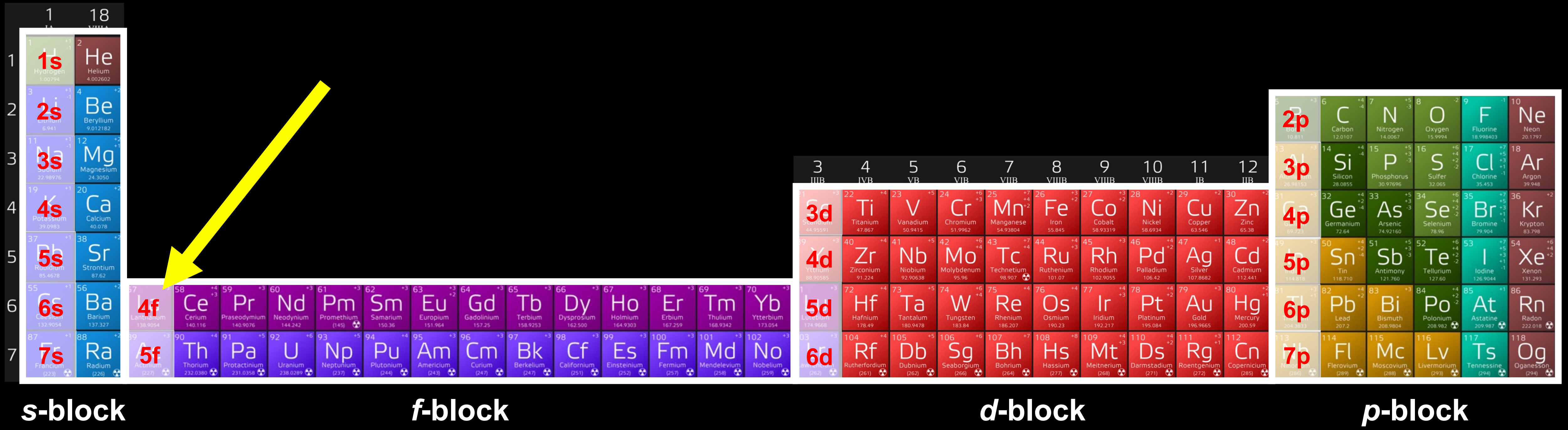
Which of the following is a possible set of quantum numbers?

A. $n = 5, l = 3, m_l = -2, m_s = -½$

B. $n = 3, l = 3, m_l = 0, m_s = +½$

C. $n = 6, l = 5, m_l = -2, m_s = -½$

D. $n = 0, l = 0, m_l = -1, m_s = +½$


1s
2s
3s
4s
5s
6s
7s
4f
5f
3d
4d
5d
6d
2p
3p
4p
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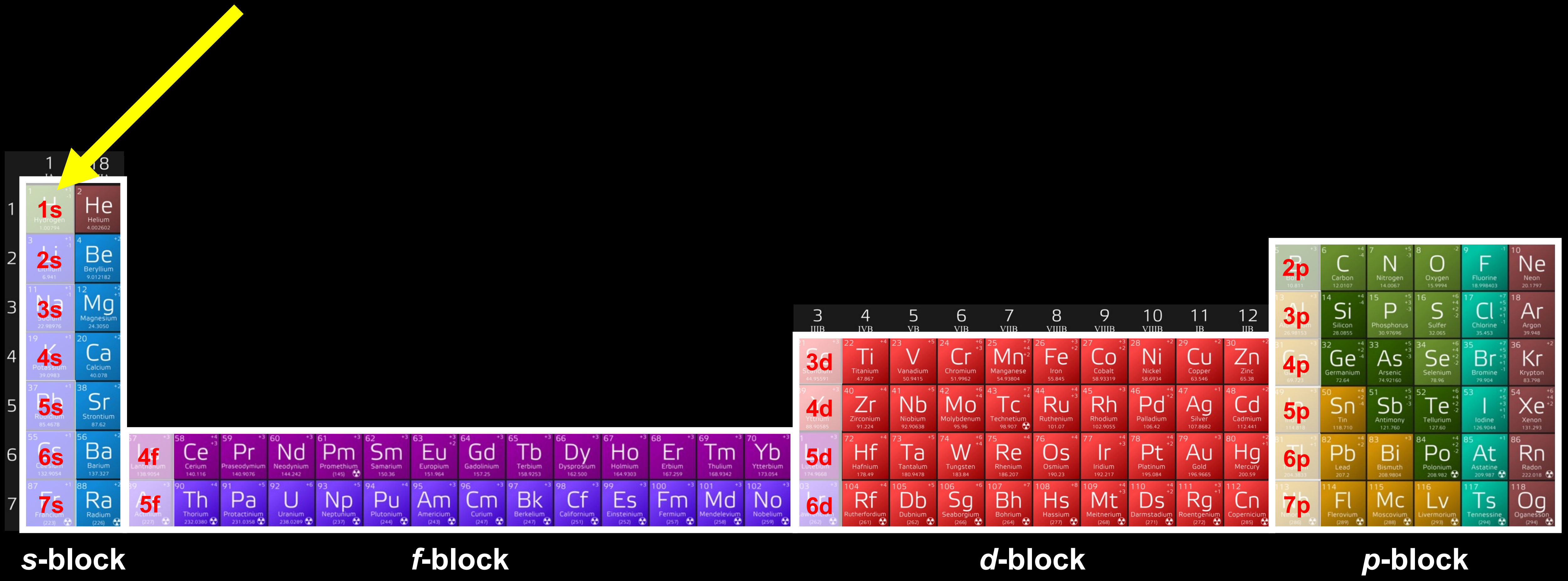
7p
s-block
f-block
d-block
p-block


1s
2s
3s
4s
5s
6s
7s
4f
5f
3d
4d
5d
6d
2p
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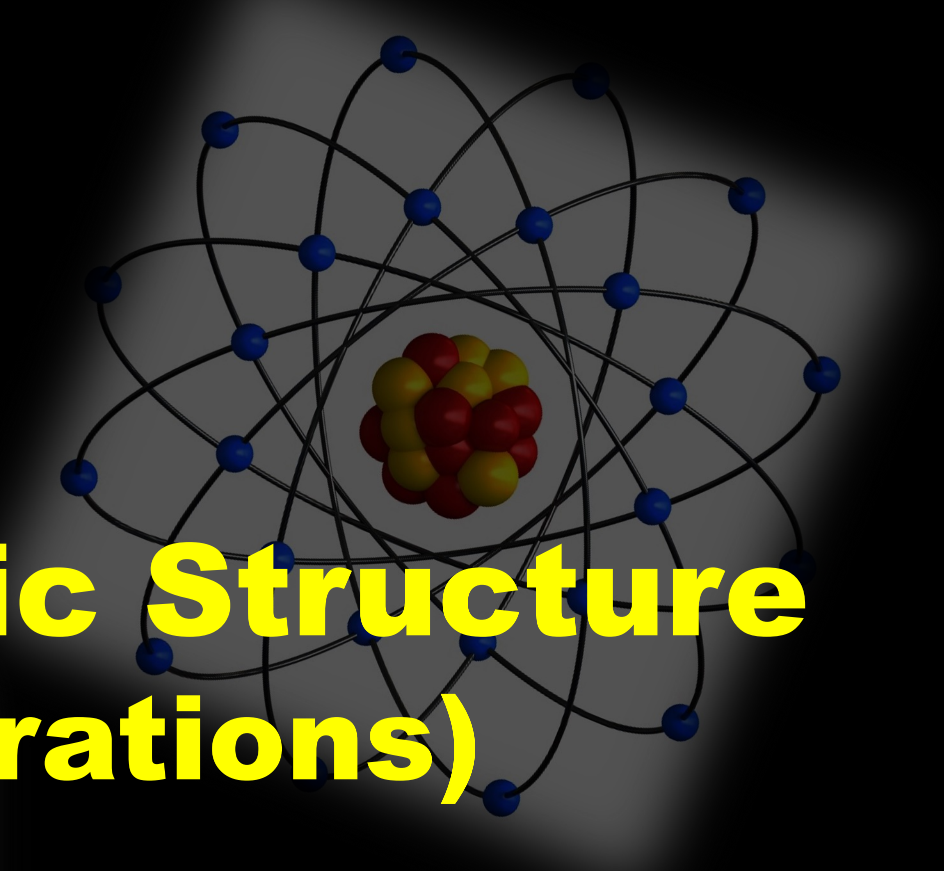
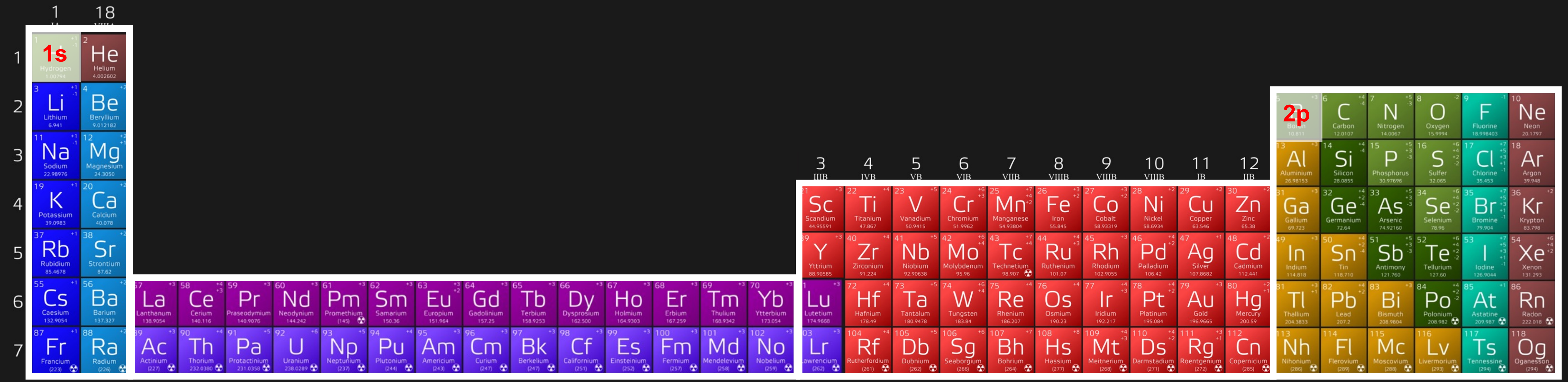
f-block
d-block
p-block



Atomic and Electronic Structure (Electron Configurations)

Electron Configurations

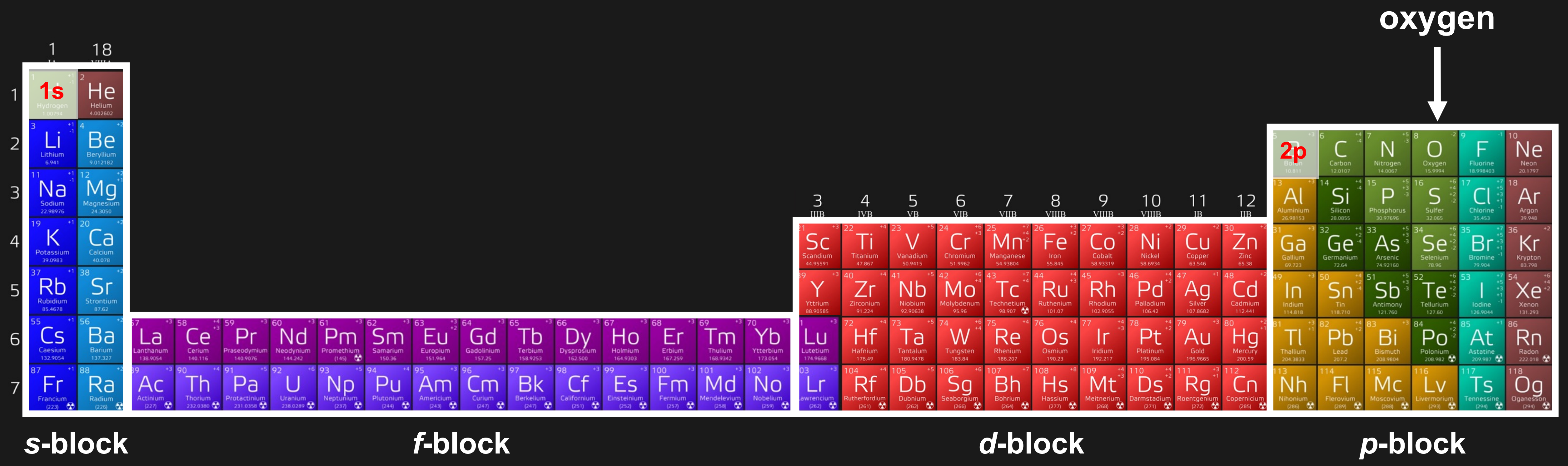
An ***electron configuration*** is a list of all the electrons in an atom and the ***orbitals*** that they occupy. To determine an element's electron configuration, we have to begin by breaking up the periodic table like this. The first box in the ***s***-block is **1s**, and the first box in the ***p***-block is **2p**.



***s*-block** ***f*-block** ***d*-block** ***p*-block**

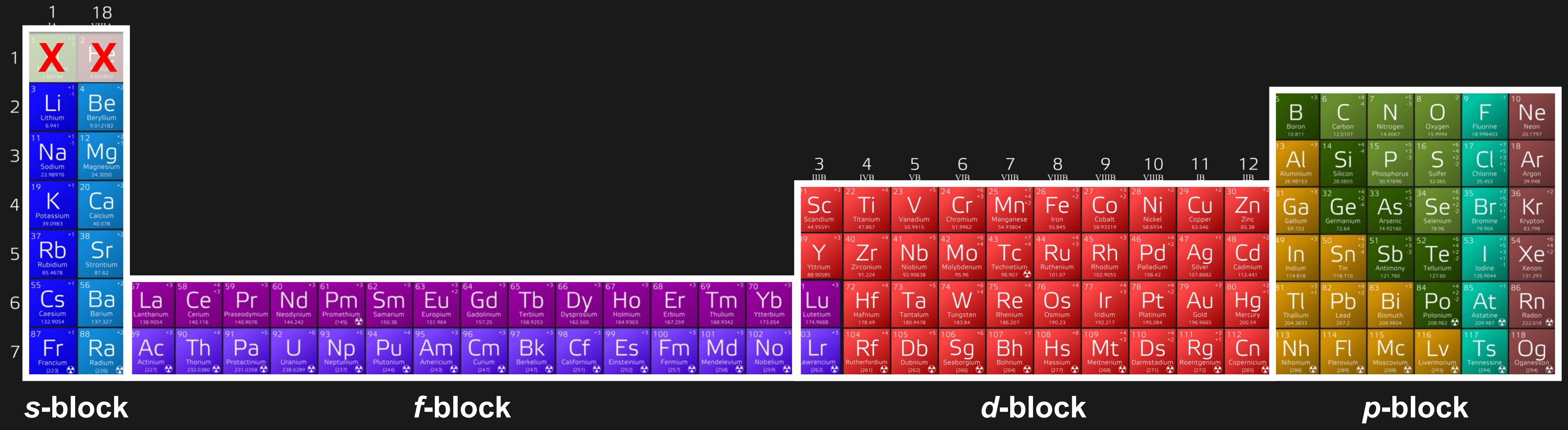
Electron Configurations

What is the ***electron configuration*** of oxygen? First, find oxygen. Next, count from left-to-right, top-to-bottom on the periodic table, remembering that each box represents one electron.



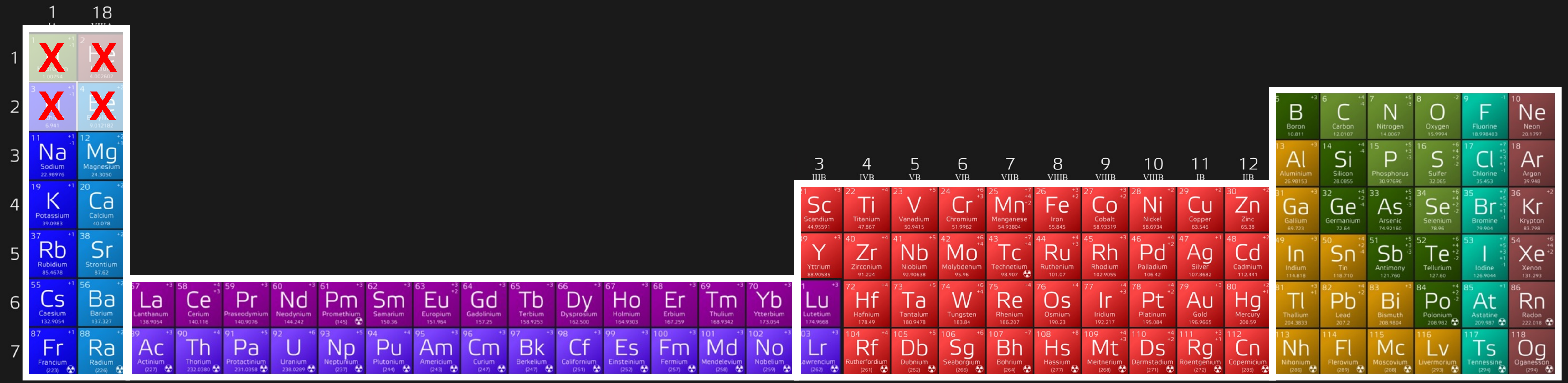
Electron Configurations

So we start with row **1s** (the hydrogen box) and count left-to-right. Thus far, then, we've counted **two** electrons on row 1 (the **1s** row). So we would say that thus far, our electron configuration is **1s²**.



Electron Configurations

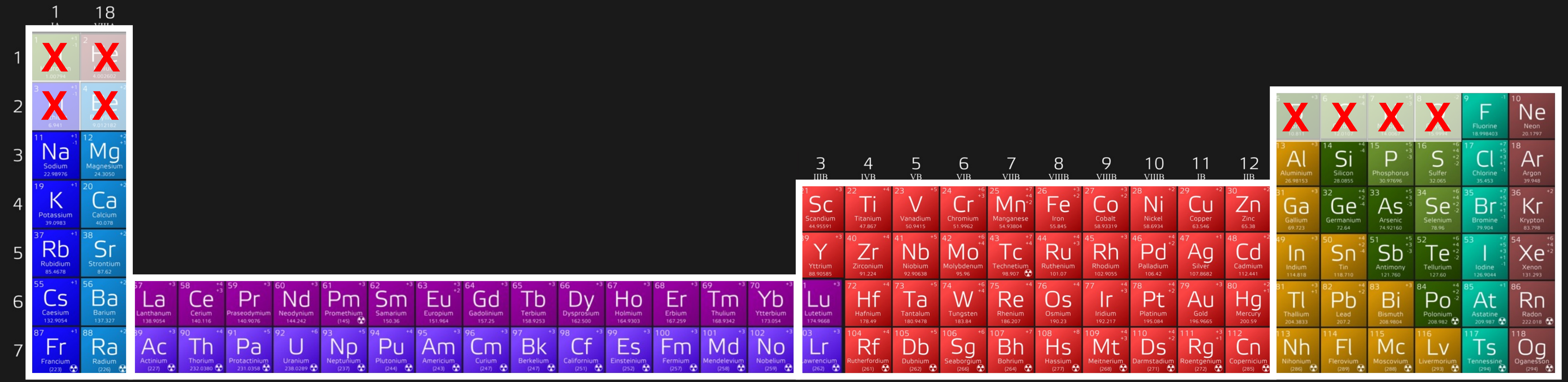
But we're not to oxygen yet. We have to keep going, beginning on row two, Box 3 (the lithium box). We count that up . . . until we fill up the **2s** row. Because we've filled it up with 2 electrons, we can say that the electron configuration we've made, so far, is now **$1s^22s^2$**.



***s*-block** ***f*-block** ***d*-block** ***p*-block**

Electron Configurations

But we're still not to oxygen. So we keep counting, starting now at Box 5 (the boron box), which is the first box of the ***p***-block. We count that up until we get to oxygen . . . So we just filled up four boxes in the ***2p***-block to get to oxygen. Thus, oxygen's full electron configuration is **$1s^2 2s^2 2p^4$**.

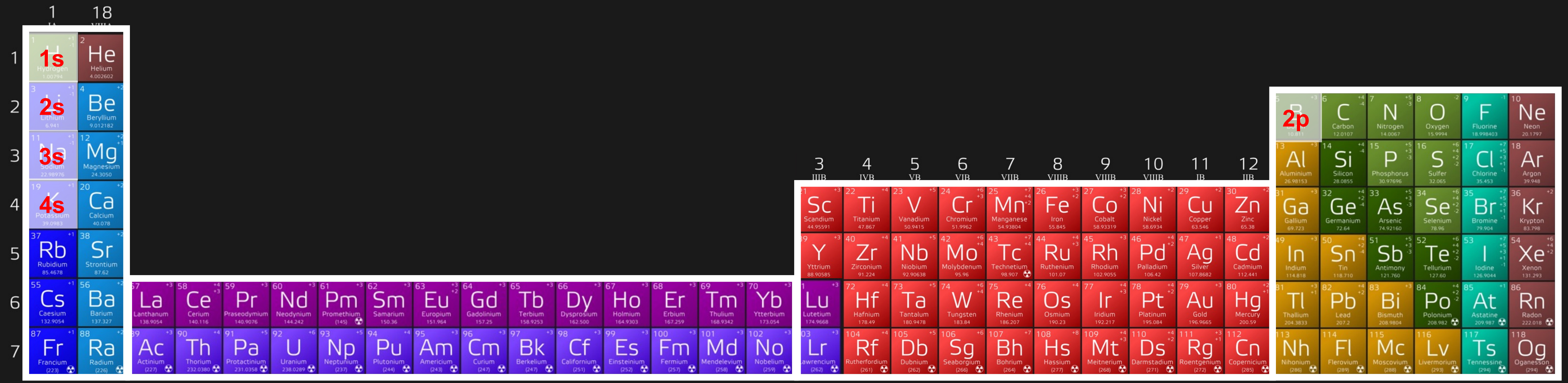


***s*-block** ***f*-block** ***d*-block** ***p*-block**

Notice that if you add up the superscript numbers in $1s^2 2s^2 2p^4$ (2+2+4), you get 8, which is the atomic # and total # of electrons in oxygen.

Electron Configurations

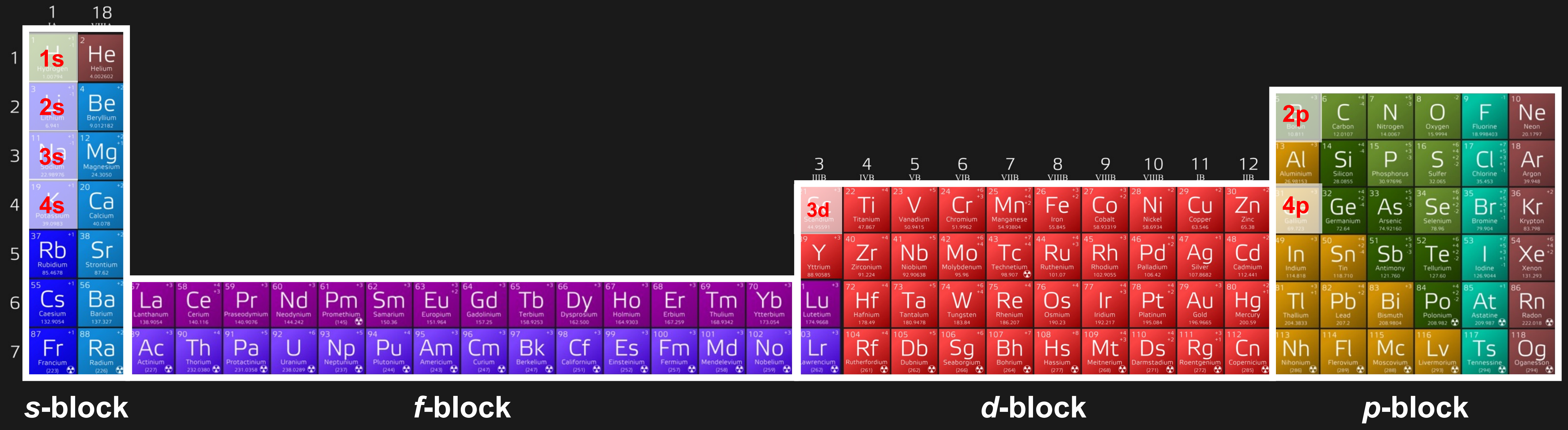
Now to something weird. Going left-to-right, from ***s***- to the ***p***-block, we eventually get to Row 4, which starts with **4s** electrons at element #19 (potassium). You'll notice, as we move from potassium to calcium (#20) to scandium (#21), that we now enter the ***d***-block.



***s*-block** ***f*-block** ***d*-block** ***p*-block**

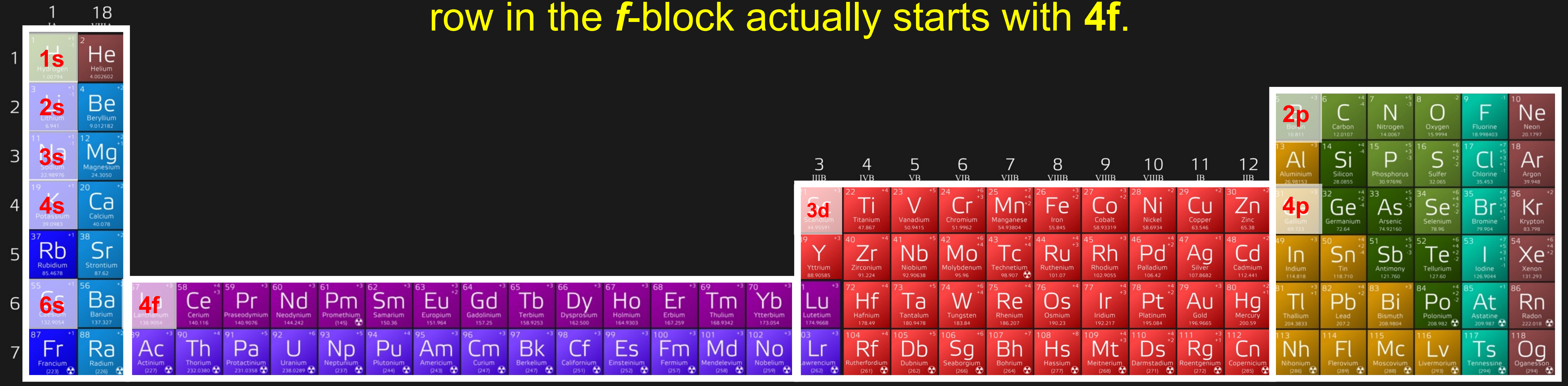
Electron Configurations

Here's the weird. Even though the first ***d***-block element (Sc, #21) is on Row 4, the ***d***-block actually reverts backwards by one number. So starting in Row 4, #21, the first row in the ***d***-block actually starts with **3d**.



Electron Configurations

Something similar happens in the ***f***-block. As we go across row 6, from Cs (#55) to Ba (#56), the next element on the period table is actually lanthanum (#57), the first element in the ***f***-block. Please note that even though the ***f***-block is on Row 6, the ***f***-block actually reverts backwards by two numbers. So starting in Row 6, #57, the first row in the ***f***-block actually starts with **4f**.



***s*-block** ***f*-block** ***d*-block** ***p*-block**

What do Electron Configurations Mean?

At this stage you may be asking yourself: "So now that I know how to assign an element's electron configuration, what in the world does that tell me?"

Do you remember all the stuff I taught you in earlier videos about ***quantum numbers*** and ***atomic orbitals?*** Well, this stuff all goes hand-in-hand! You see, every part of an element's electron configuration tells you something about its quantum numbers and orbitals:

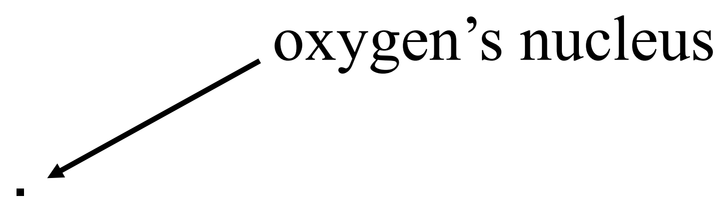
This represents the principle quantum number, or energy level, ***n***. (**n = 1**) → **$1s^2$** ← This represents how many electrons (in this case, 2) are <u>*in*</u> the orbital in question.

↑

This represents which kind of the orbital these particular electrons are in. In this case, it's an ***s***-orbital, so ***l*** **= 0**.

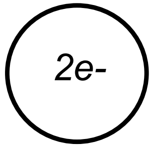
What do Electron Configurations Mean?

So let’s return to oxygen, which has an electron configuration of **$1s^2 2s^2 2p^4$**. Here’s oxygen’s tiny nucleus, where its neutrons and protons are kept:



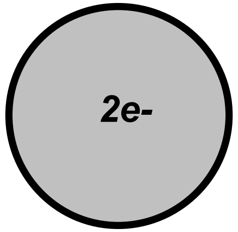
What do Electron Configurations Mean?

Around the nucleus is a spherically-shaped **1s** orbital. It contains two electrons.



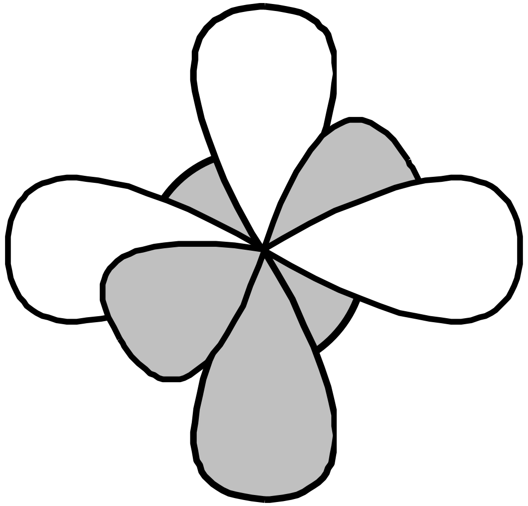
What do Electron Configurations Mean?

Outside of the **1s** orbital is a larger (also spherically-shaped) **2s** orbital, which also contains two electrons:



What do Electron Configurations Mean?

Outside of the **2s** orbital are three dumbbell-shaped **2p** orbitals, oriented along the x, y, and z axes.



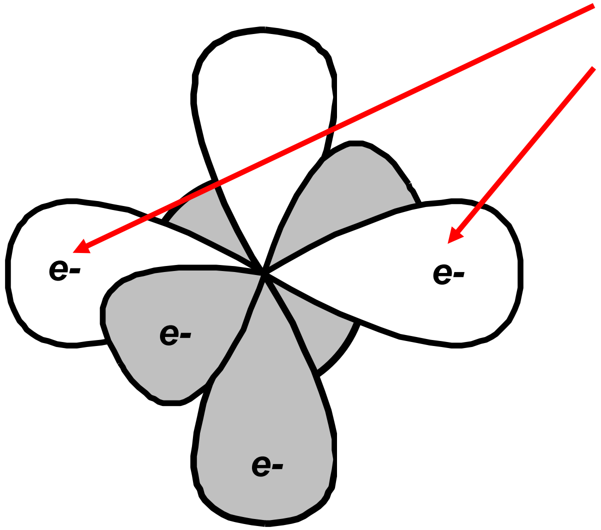
What do Electron Configurations Mean?

You'll notice that the last part of oxygen's electron configuration (**$1s^2 2s^2 2p^4$**) says **$2p^4$**. This means that there are four electrons occupying oxygen's three **2p** orbitals. So where do these electrons go?

Well, they fill up these orbitals one at a time. In other words, electrons won't pair up to occupy the same orbital until there's no other option. This is called ***Hund's rule***, which I'll discuss in greater detail later on.

What do Electron Configurations Mean?

So this means that the four electrons occupying each **2p** orbital fill up like this:

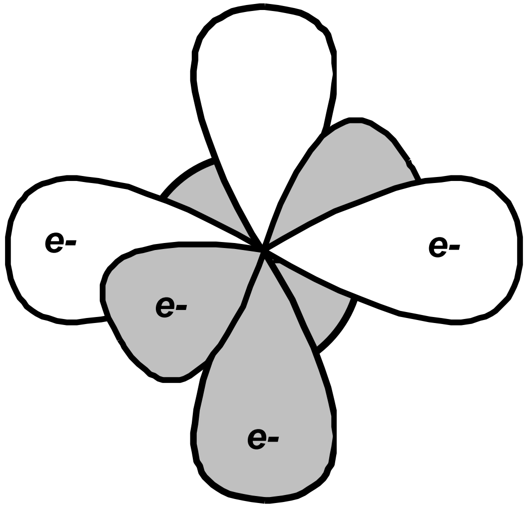


These two are paired, occupying the same **2p** orbital.

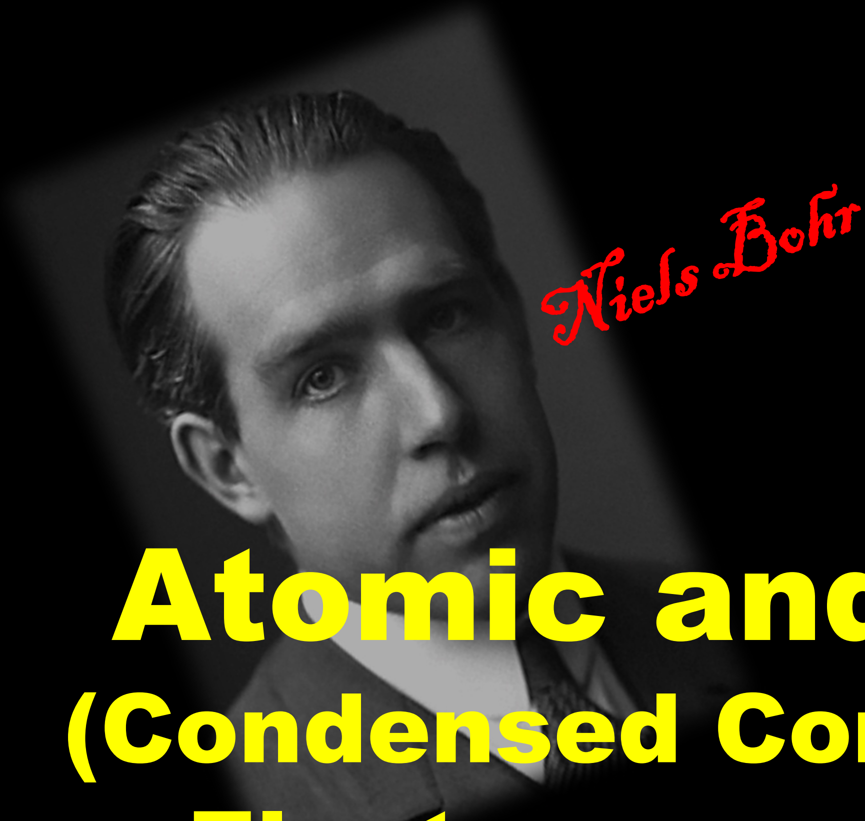
The two paired electrons in the **2p** orbital along the x-axis are different from each other because one has a **+½** spin, and the other has a **–½** spin.

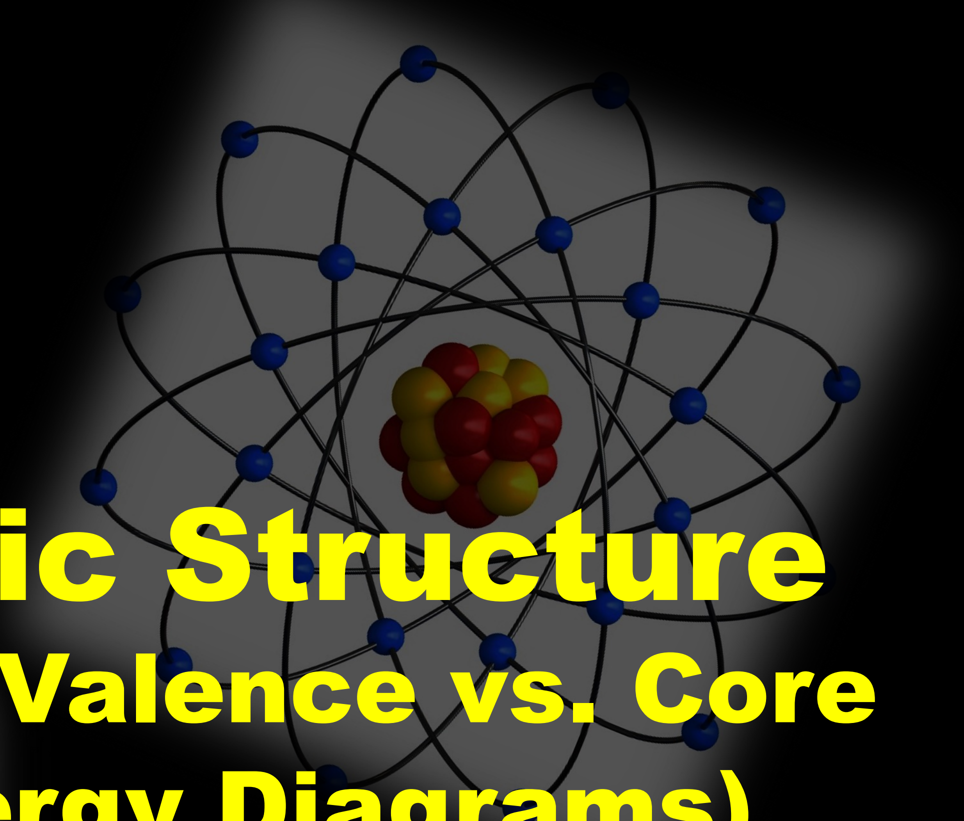
What do Electron Configurations Mean?

So when we talk about oxygen having an electron configuration of **$1s^22s^22p^4$**, this is what we're talking about. We're saying that in oxygen's **1s** orbital, there are two paired electrons. In its **2s** orbital, there are two more paired electrons. And in its three **2p** orbitals, there are four electrons, two of which are unpaired and two of which are paired.



For a cool video on this, check out this html: http://www.youtube.com/watch?v=sMt5Dcex0kg





Atomic and Electronic Structure

(Condensed Configurations, Valence vs. Core Electrons, and Orbital Energy Diagrams)

Condensed Electron Configurations

As you may imagine, writing out full ***electron configurations*** can get tedious for elements that are further down on the periodic table. For example, bromine's full ***electron configuration*** is $\underbrace{1s^22s^22p^63s^23p^6}_{[Ar]}4s^23d^{10}4p^5$, which is really long.

Fortunately, there is a shorthand –or "condensed" way– of writing ***electron configurations***. Using this condensed approach, we just write, in brackets, the elemental symbol of the noble gas that comes *just before* the element in question. We then write the rest of the electron configuration after it.

In bromine's case, the noble gas the precedes Br is Ar, whose electron configuration is $1s^22s^22p^63s^23p^6$. We can therefore abbreviate Br's configuration as $[Ar]4s^23d^{10}4p^5$. This is Br's ***condensed electron configuration***.

Valence vs. Core Electrons

The vast majority of chemical reactions involve the exchange or sharing of electrons: specifically, of <u>*outer*</u> electrons. These outer electrons are called ***valence electrons***.

Valence electrons are different from ***core electrons***. ***Valence electrons*** are the ones in an atom's outermost shells, while ***core electrons*** are the ones that are buried inside the <u>*inner*</u> shells.

You can quickly identify ***valence*** versus ***core electrons*** by looking at an element's ***electron configuration***. For example, Br's full electron configuration is **$1s^22s^22p^63s^23p^64s^23d^{10}4p^5$**. Its ***valence electrons*** are in its outermost shells. Its ***core electrons*** are all the inner shells:

$$\underbrace{1s^22s^22p^63s^23p^6}_{\text{core}}\,4s^2\,\underbrace{3d^{10}}_{\text{core}}\,4p^5$$

valence

So Br has 7 ***valence electrons***.

Valence vs. Core Electrons

In general, then, if you look at an atom's ***electron configuration***, its electrons whose orbitals are the furthest away from the nucleus (the ones with the largest ***principle quantum number***, **n**) are that atom's ***valence electrons***. These are its highest-energy electrons: the electrons that are at play in a chemical reaction with that element. All the other electrons in the ***electron configuration*** are ***core electrons***.

There is one type of exception. For elements that are *in* the ***d*-block**, the ***d*-shell** electrons are also ***valence electrons***. For example, Ti's electron configuration is:

$$\underbrace{1s^2 2s^2 2p^6 3s^2 3p^6}_{\text{core}}\underbrace{4s^2 3d^2}_{\text{valence}}$$

Ti, then, has 4 ***valence electrons***. Again, you only count ***d*-shell** electrons as ***valence electrons*** for elements that are *in* the ***d*-block**.

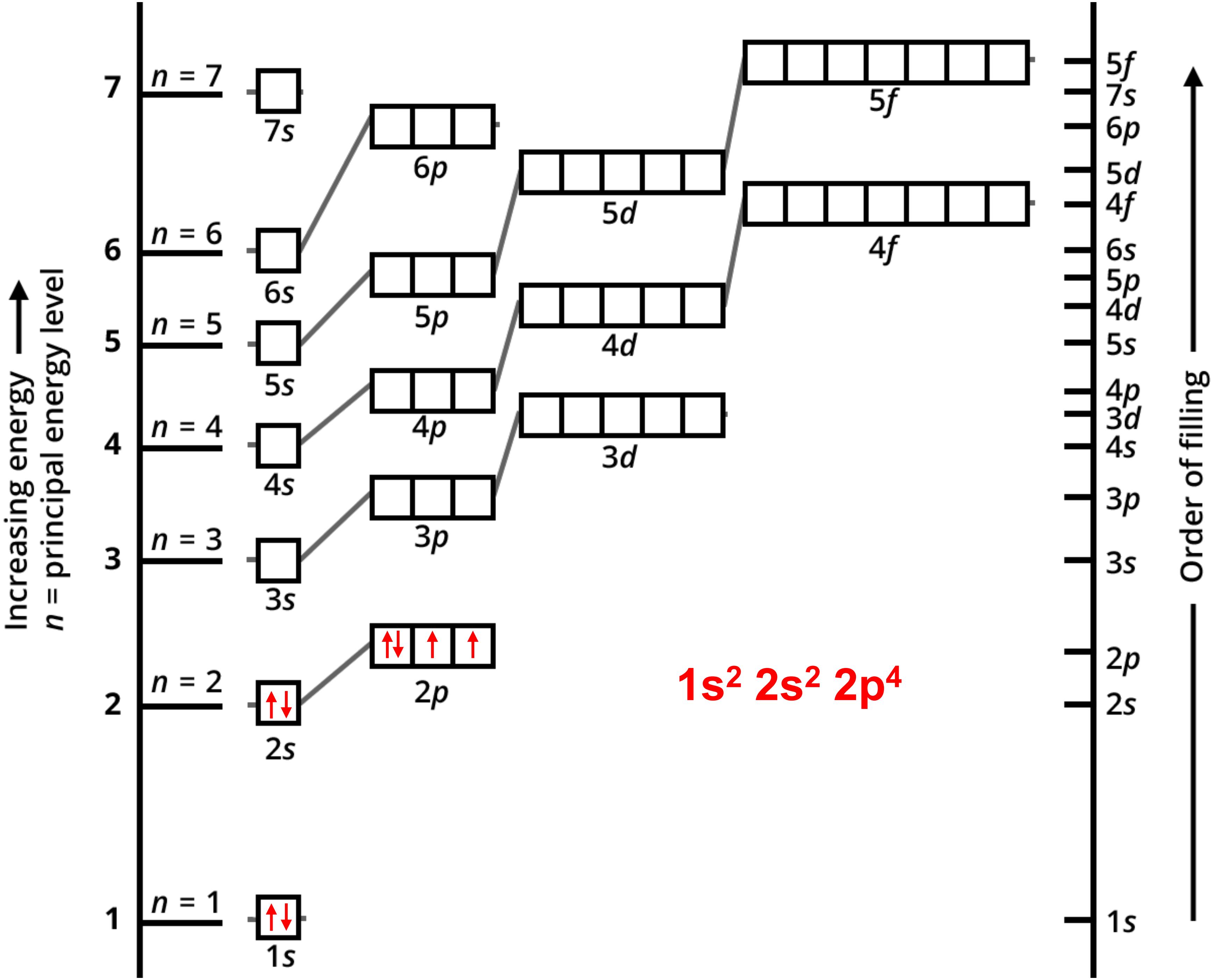
Atomic Orbital Energy Diagrams

As I mentioned in an earlier video, the larger an electron's ***principle quantum number*** (**n**), the further away from the nucleus, higher in energy, and more reactive it is:

↑ ***principle quantum number*** (**n**) **=** ↑ **distance from nucleus, energy, reactivity**

We can draw diagrams that show the relative energies of the various atomic orbitals and shells. In these diagrams, called ***atomic orbital energy diagrams***, each box represents an orbital. So, the higher up the box is, the further the orbital is from the atom's nucleus, the higher in energy the orbital is, and the more reactive the orbital is.

We can fill in each box with electrons, drawing an "up" arrow for electrons with a **+1/2** spin, and a "down" arrow for electrons with a **-1/2** spin.



Increasing energy
n = principal energy level
7
n = 7
6
n = 6
5
n = 5
4
n = 4
3
n = 3
2
n = 2
1
n = 1
7s
6s
5s
4s
3s
2s
1s
6p
5p
4p
3p
2p
5d
4d
3d
5f
4f
1s2 2s2 2p4
5f
7s
6p
5d
4f
6s
5p
4d
5s
4p
3d
4s
3p
3s
2p
2s
1s
Order of filling

Energy Diagrams Rules

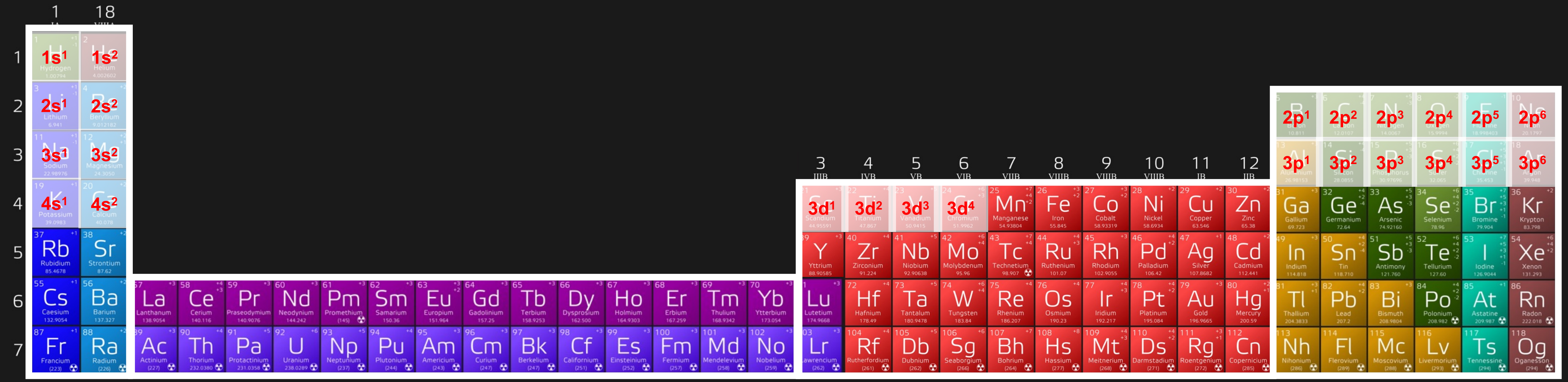
Aufbau Principle: electrons fill the lowest energy orbitals first

Hund's Rule: Don't pair up electrons until you have to.

Pauli Exclusion Principle: no two electrons in the same atom can have the same four ***quantum numbers***. In other words, no two electrons in the same atom can have the exact same address.

Electron Configuration Exceptions

When determining an element's ***electron configuration***, nearly all of the elements follow the pattern you would normally predict. However, there are five exceptions. You would assume, then, that Cr's configuration would be **[Ar]4s²3d⁴**. Unfortunately, it's not. In reality, Cr's configuration is **[Ar]4s¹3d⁵**. I'll now explain why.

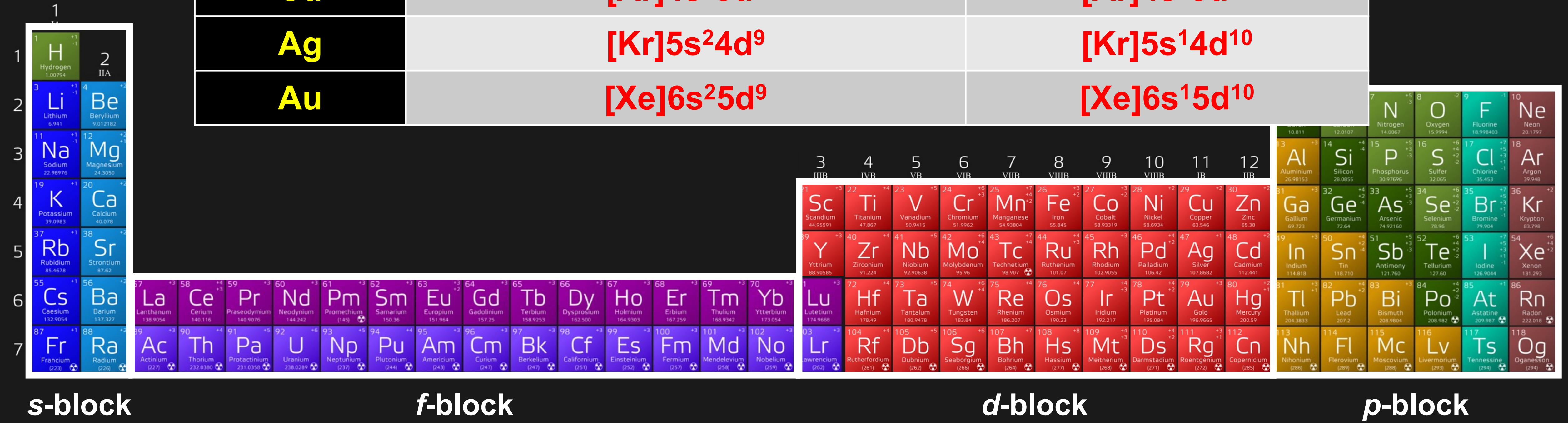


***s*-block** ***f*-block** ***d*-block** ***p*-block**

Electron Configuration Exceptions

Here are the five exceptions you should memorize for the EXAM:

Element	What you think it should be	What it actually is
Cr	$[Ar]4s^2 3d^4$	$[Ar]4s^1 3d^5$
Mo	$[Kr]5s^2 4d^4$	$[Kr]5s^1 4d^5$
Cu	$[Ar]4s^2 3d^9$	$[Ar]4s^1 3d^{10}$
Ag	$[Kr]5s^2 4d^9$	$[Kr]5s^1 4d^{10}$
Au	$[Xe]6s^2 5d^9$	$[Xe]6s^1 5d^{10}$



Electron Configuration Exceptions

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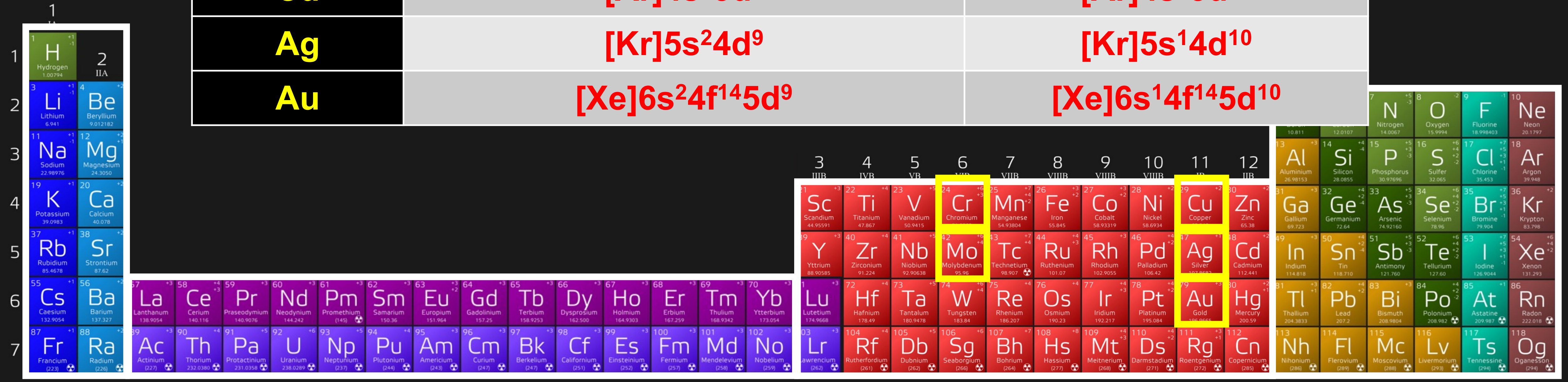
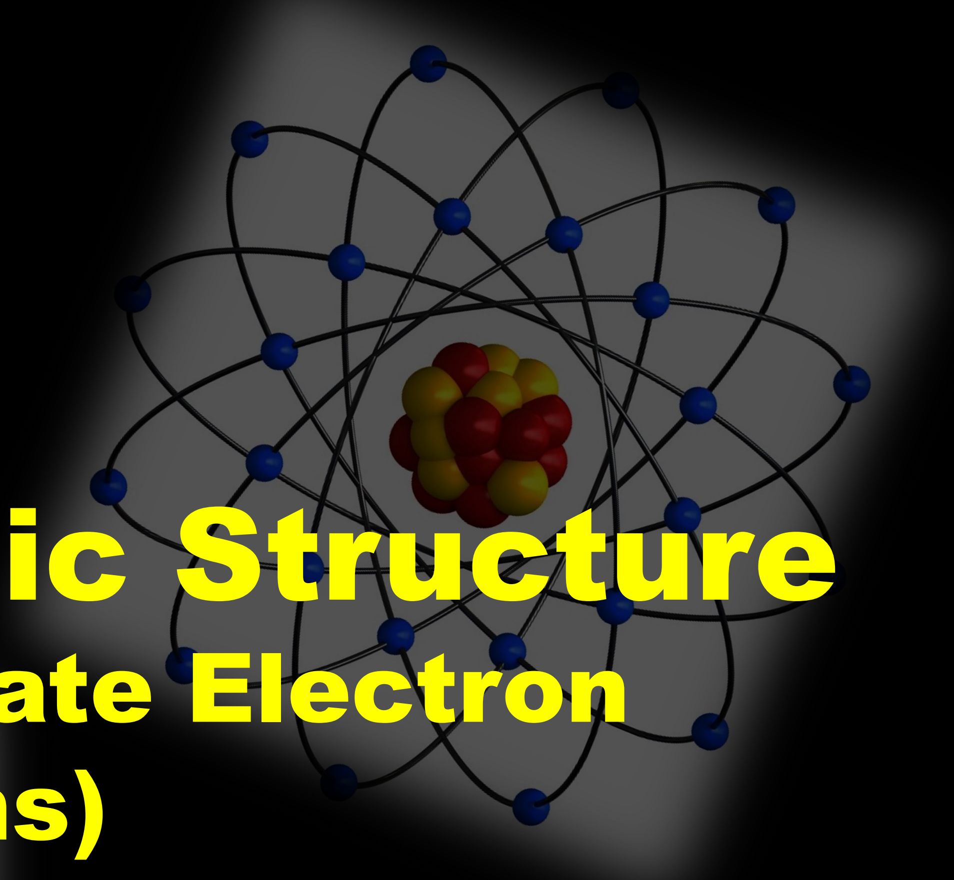
Mo	$[Kr]5s^24d^4$	$[Kr]5s^14d^5$
Cu	$[Ar]4s^23d^9$	$[Ar]4s^13d^{10}$
Ag	$[Kr]5s^24d^9$	$[Kr]5s^14d^{10}$
Au	$[Xe]6s^24f^{14}5d^9$	$[Xe]6s^14f^{14}5d^{10}$



s-block *f*-block *d*-block *p*-block

Atomic and Electronic Structure

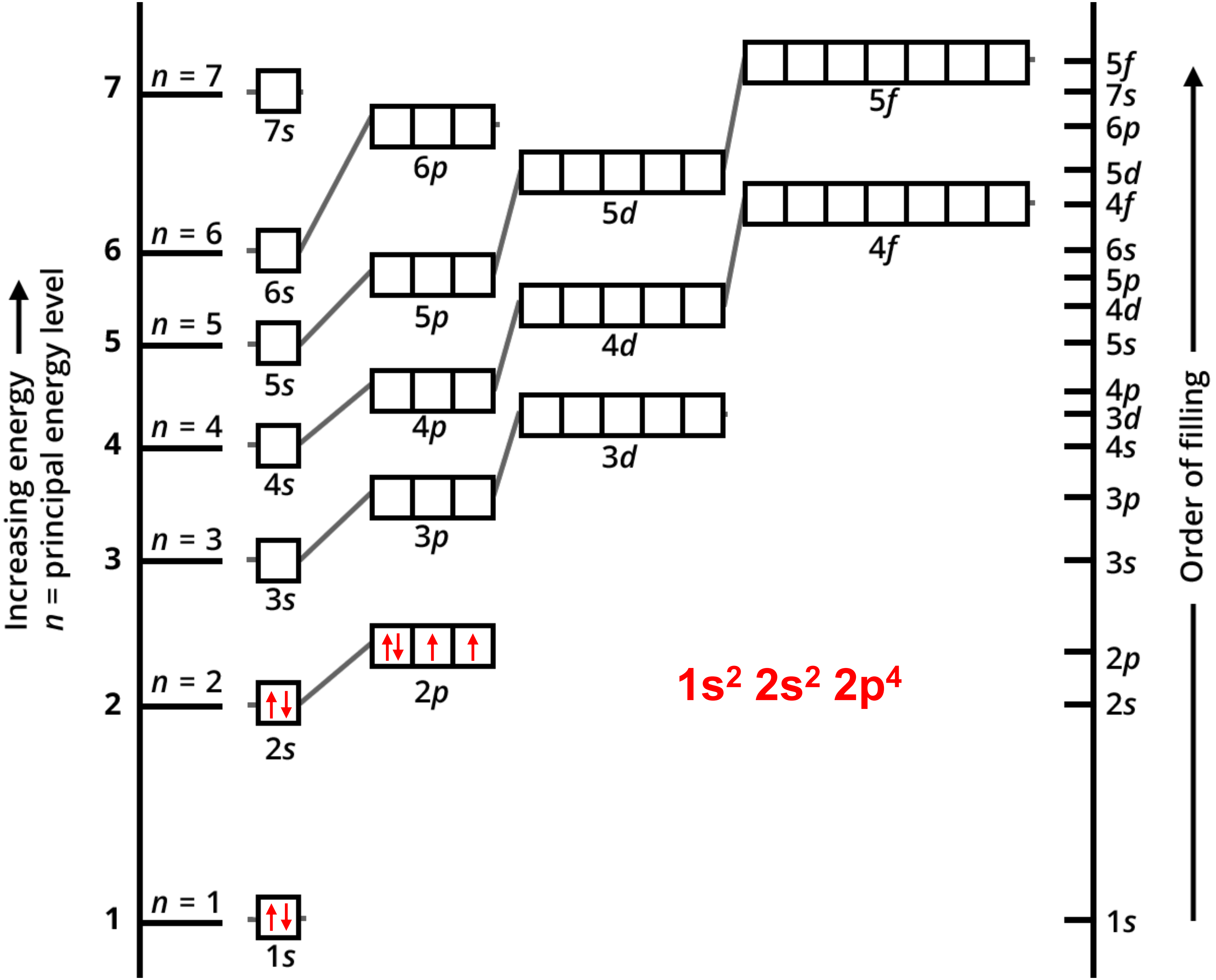
(Excited- vs. Ground-State Electron Configurations)



Excited vs. Ground States

All the ***electron configurations*** I've taught you thus far are element's ***ground-state electron configurations***. That is, the ways in which their electrons are typically arranged when nothing "exciting" is going on.

As it turns out, when an atom absorbs a photon of light (which you can think of as being a particle of light) or other form of energy, its electrons can get "promoted" up to a higher-energy shell or orbital.



Increasing energy
n = principal energy level
7
6
5
4
3
2
1
n = 7
n = 6
n = 5
n = 4
n = 3
n = 2
n = 1
7s
6s
5s
4s
3s
2s
1s
6p
5p
4p
3p
2p
5d
4d
3d
5f
4f
1s2 2s2 2p4
5f
7s
6p
5d
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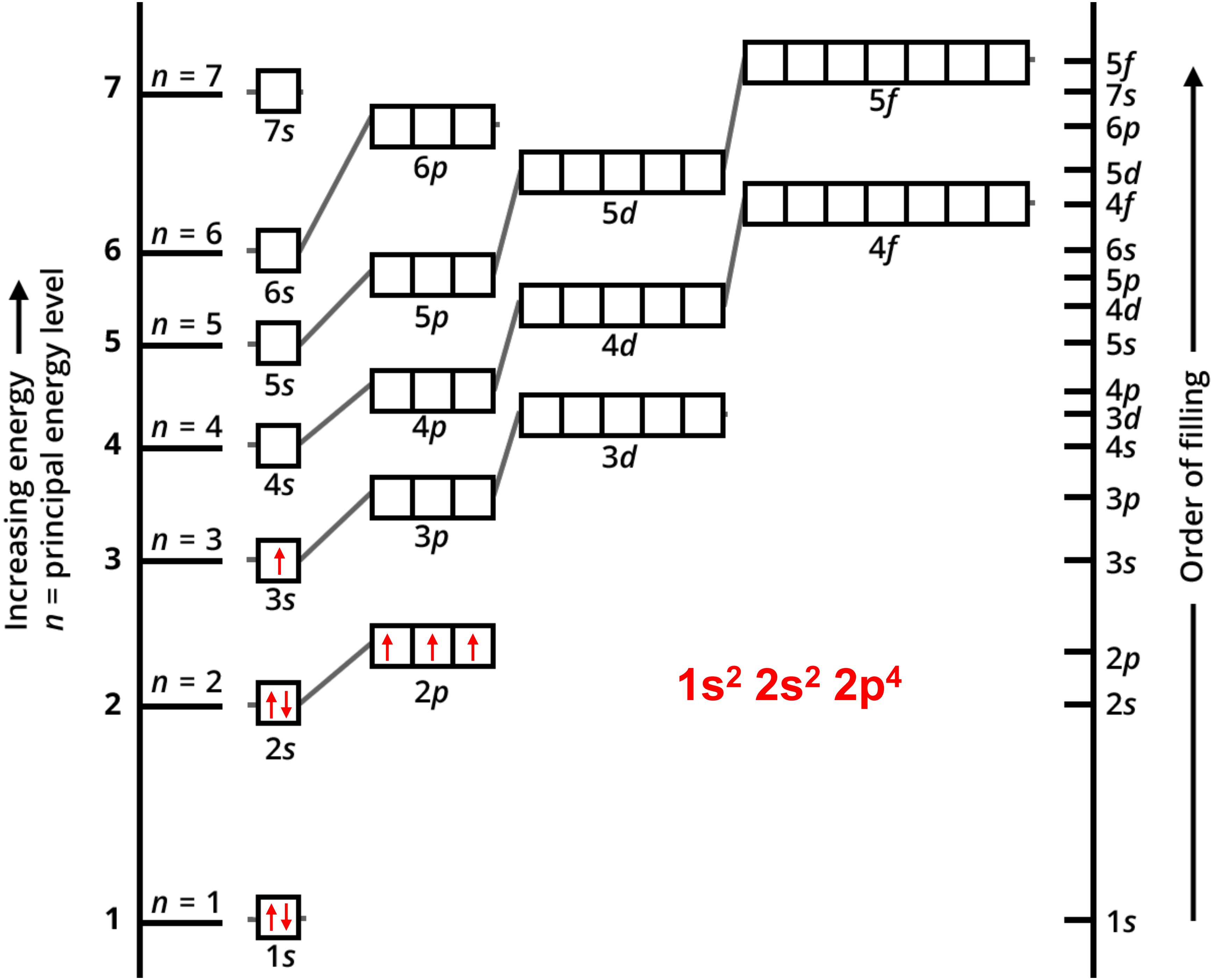
4p
3d
4s
3p
3s
2p
2s
1s
Order of filling



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6p
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3p
2p
5d
4d
3d
5f
4f
1s2 2s2 2p4
5f
7s
6p
5d
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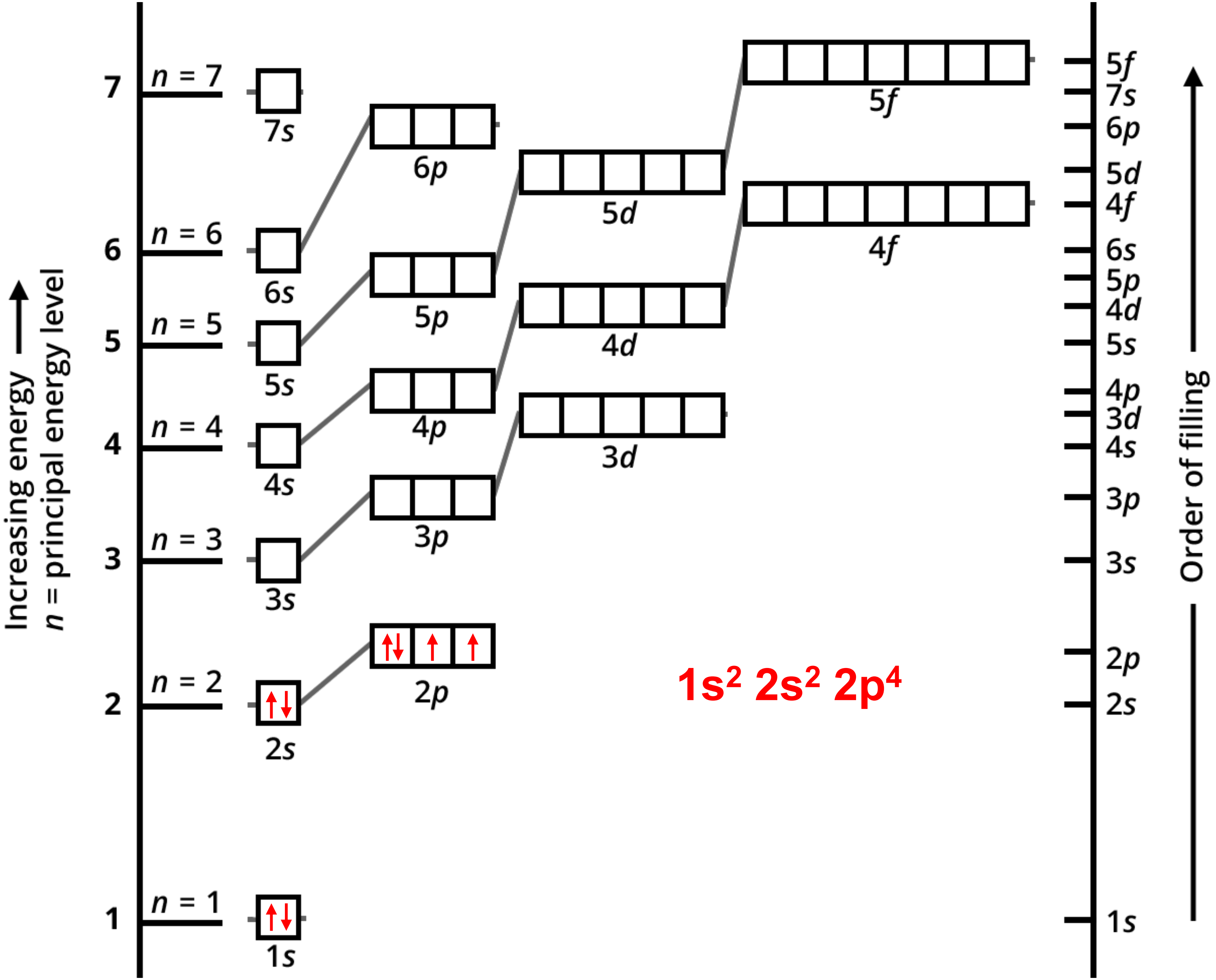
4s
3p
3s
2p
2s
1s
Order of filling



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5d
4d
3d
5f
4f
1s2 2s2 2p4
5f
7s
6p
5d
4f
6s
5p
4d
5s
4p
3d
4s
3p
3s
2p
2s
1s
Order of filling

Question

Which of the following ***electron configurations*** represents an element in an excited state?

A. $1s^2 2s^2 2p^5$

B. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$

C. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$

D. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$

E. $1s^2 2s^2 2p^6 3s^2 3p^2 4s^1$

Strategy: find the element that doesn't have its regular ground-state configuration.

Questions

How many unpaired electrons are present in the +2 cobalt ion, Co^{2+}, in its ground state?

a. 2
b. 3
c. 4
d. 5
e. 6

Questions

Please indicate which element corresponds to each of the following ***electron configurations***.

A. $1s^2 2s^2 2p^6$

B. $1s^2 2s^2 2p^5$

C. $[Kr] 5s^1 4d^5$

D. $[Ar] 4s^2 3d^{10} 4p^4$

Questions

Why is each of the following ***electron configurations*** incorrect for atoms in their ground states? How could you correct each of these?

A. $1s^2 2s^2 3s^2$

B. $[Ar]3s^2 3p^3$

C. $[Ar]4s^2 4d^3$

Questions

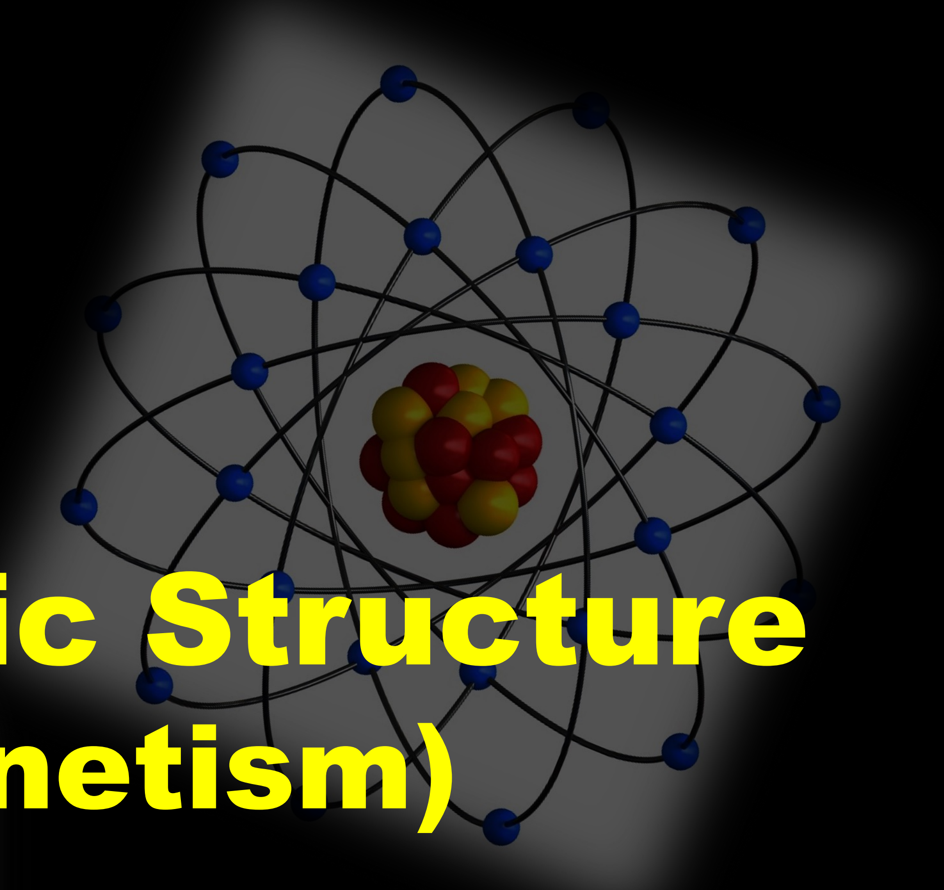
Please write the electron configuration of each of the following elements:

A. Oxygen
B. Potassium
C. Arsenic
D. Manganese
E. Copper

Questions

Please write the ***condensed electron configuration*** of each of the following elements:

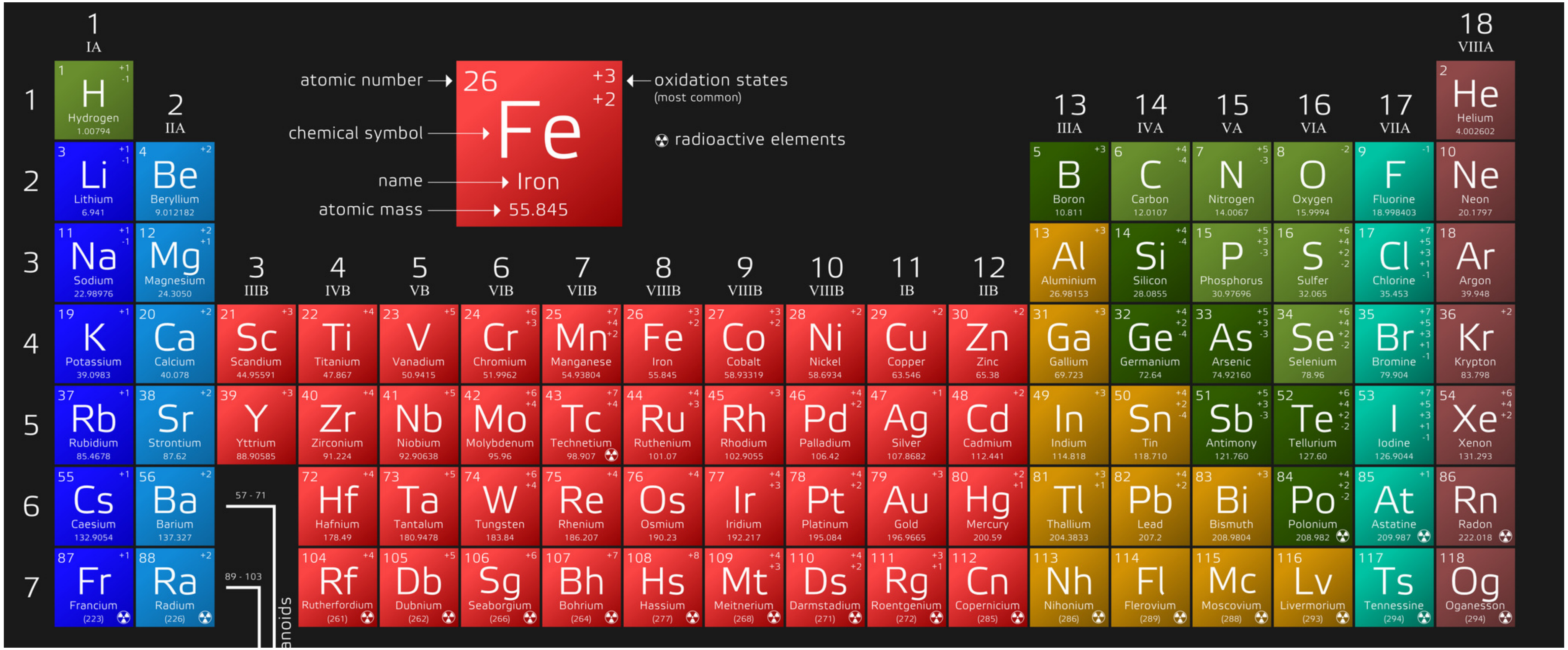
A. Fr
B. Br
C. Ca
D. Zn

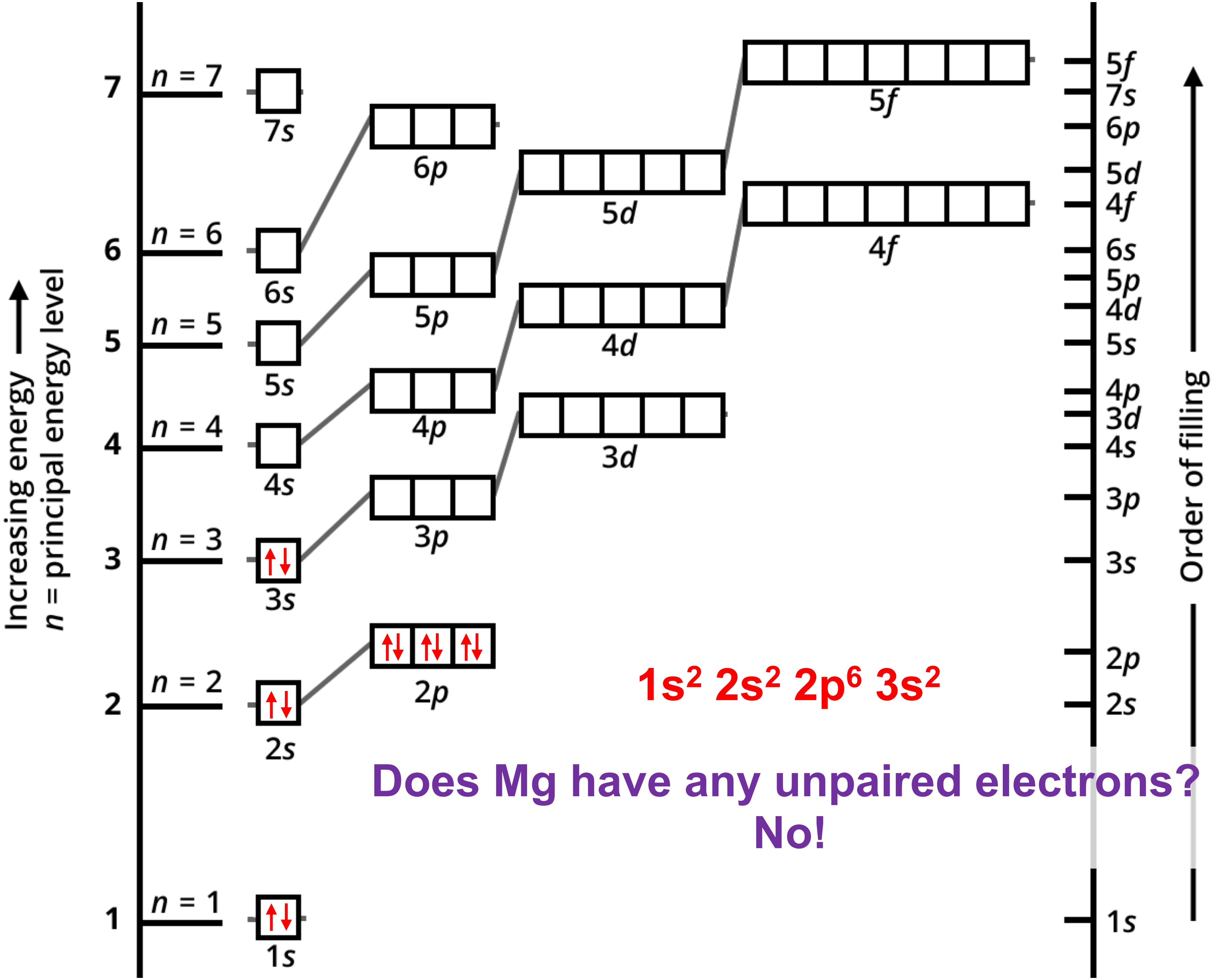


Atomic and Electronic Structure (Para vs. Diamagnetism)

Paramagnetic vs. Diamagnetic

Building on principles we've discussed in prior videos, you can hopefully see that magnesium (**Mg**) has an electron configuration of **$1s^2\ 2s^2\ 2p^6\ 3s^2$**.



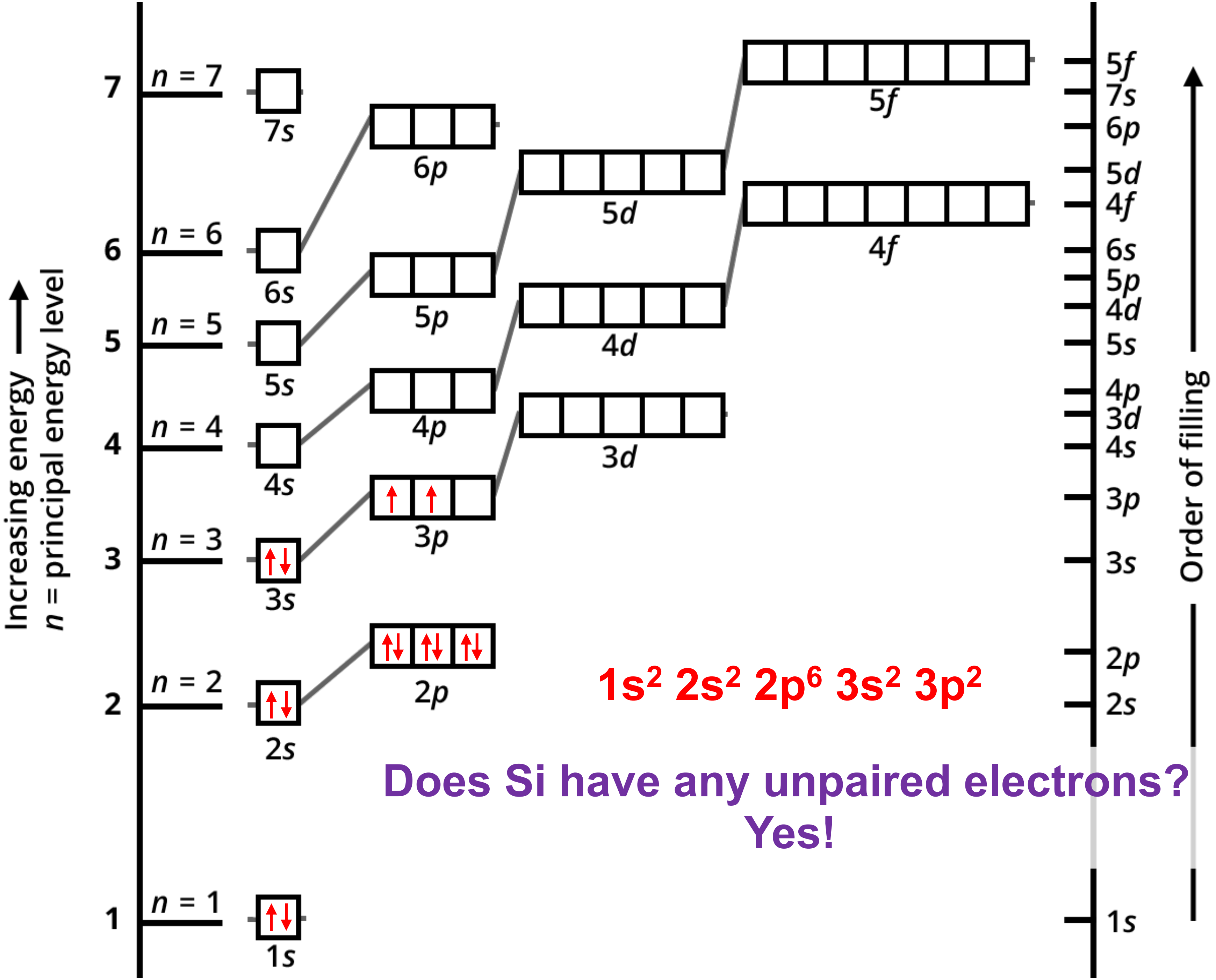


$1s^2\ 2s^2\ 2p^6\ 3s^2$

Does Mg have any unpaired electrons?

No!

14 +4 -4
Si
Silicon
28.0855

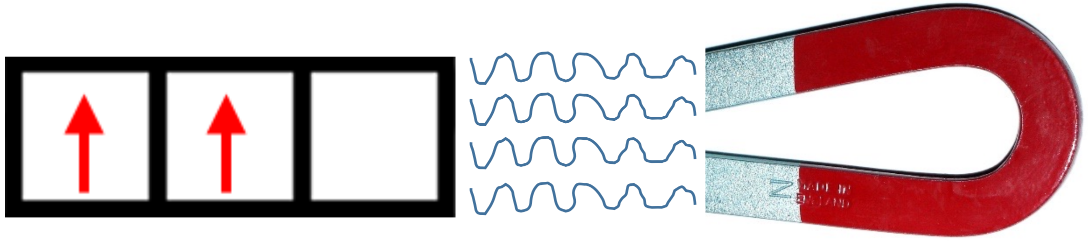


$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$

Does Si have any unpaired electrons?
Yes!

Paramagnetic vs. Diamagnetic

Paramagnetic: has unpaired electrons. Paramagnetic elements are attracted to magnets:



Paramagnetic vs. Diamagnetic

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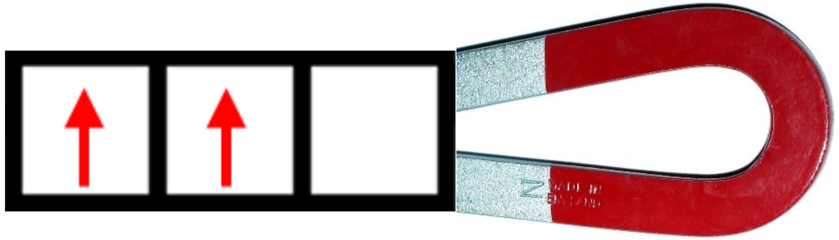
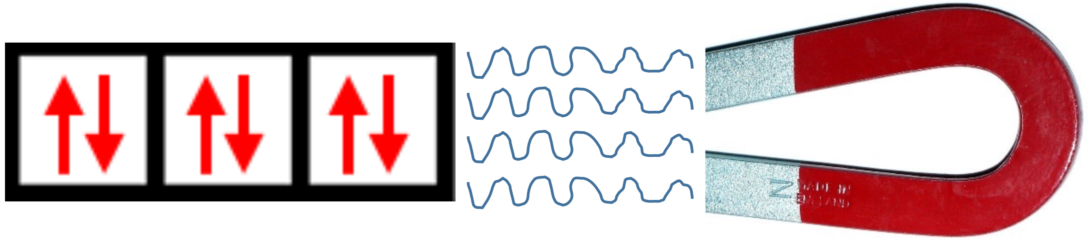


Diamagnetic: all electrons are paired. Diamagnetic elements are slightly repelled by magnetic fields:



Paramagnetic vs. Diamagnetic

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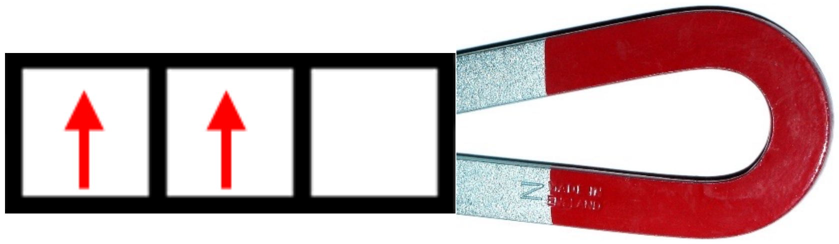

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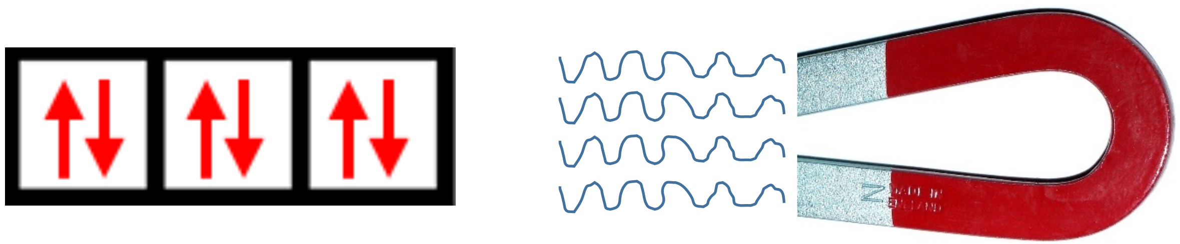


Paramagnetic vs. Diamagnetic

When I hear . . .	I think . . .	I then think . . .
Paramagnetic	"Unpaired" -a-magnetic	Attracted to magnets because it has unpaired electrons.
Diamagnetic	The other one (all paired)	Slightly repelled by magnets because it has all paired electrons.

Quick tip: If an element has an ODD number of electrons, then it's ***paramagnetic***.
If it has an EVEN number of electrons, then it can be either ***paramagnetic*** or ***diamagnetic***; you have to fill out the electron configuration energy diagram to find out.

Questions

Which of the following statements is true regarding the Mn^+ ion?

I. It is paramagnetic.
II. It has five d-electrons.
III. It has 2 valence electrons.

A. I only
B. II and III only
C. II only
D. I and III only
E. I and II only

Questions

Which of the following are paramagnetic? (Circle all that apply.)

A. Cl
B. Pt
C. Cr^{3+}
D. Br^{-}
E. O

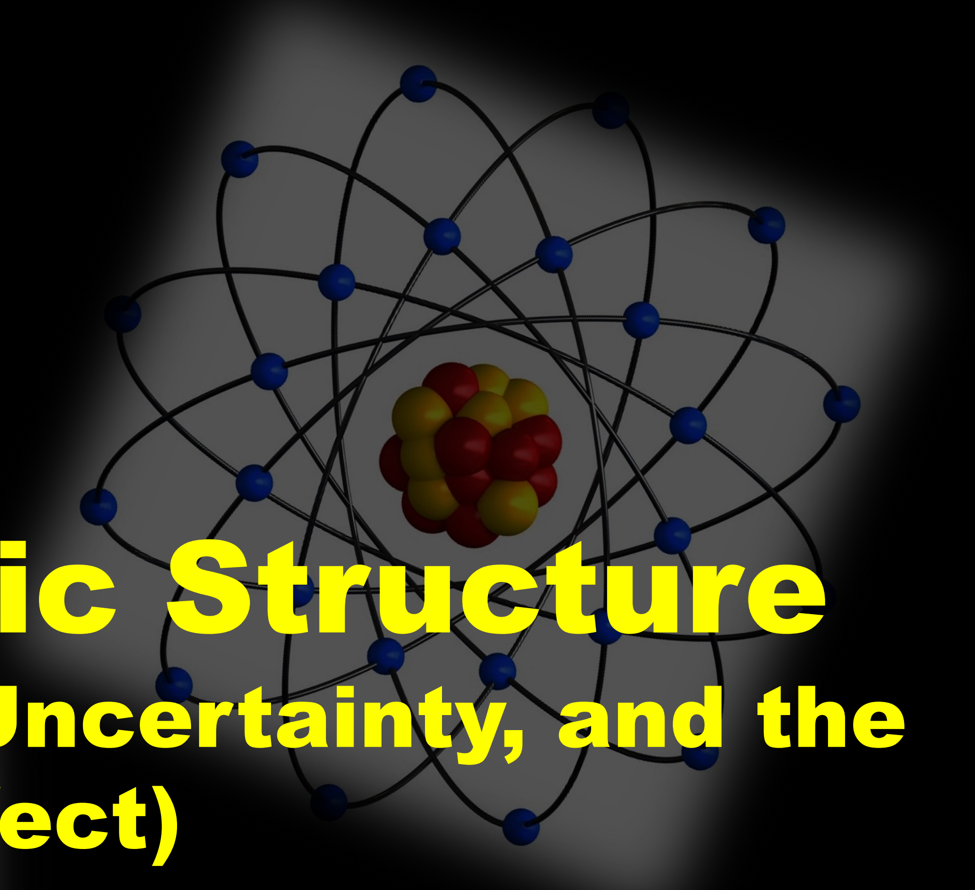
Questions

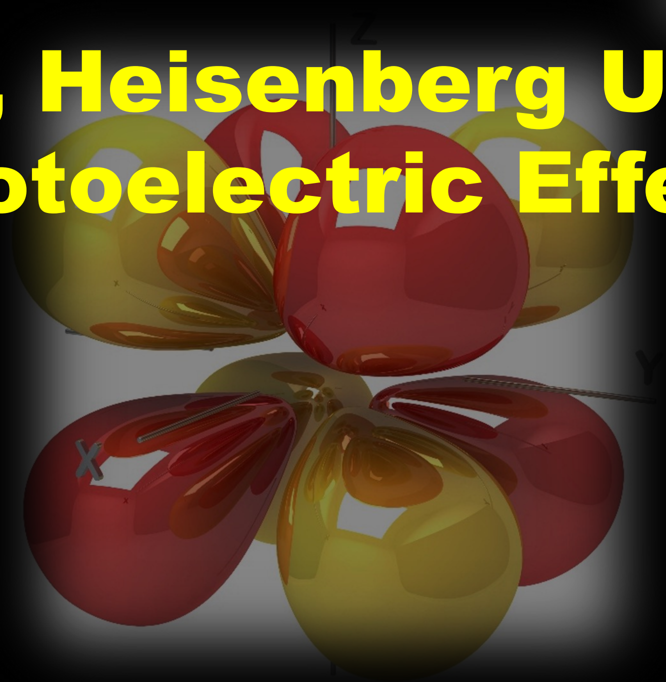
Which of the following would interact with a magnetic field? (Circle all that apply.)

A. Cl
B. Pt
C. Cr^{3+}
D. Br^{-}
E. O

Atomic and Electronic Structure

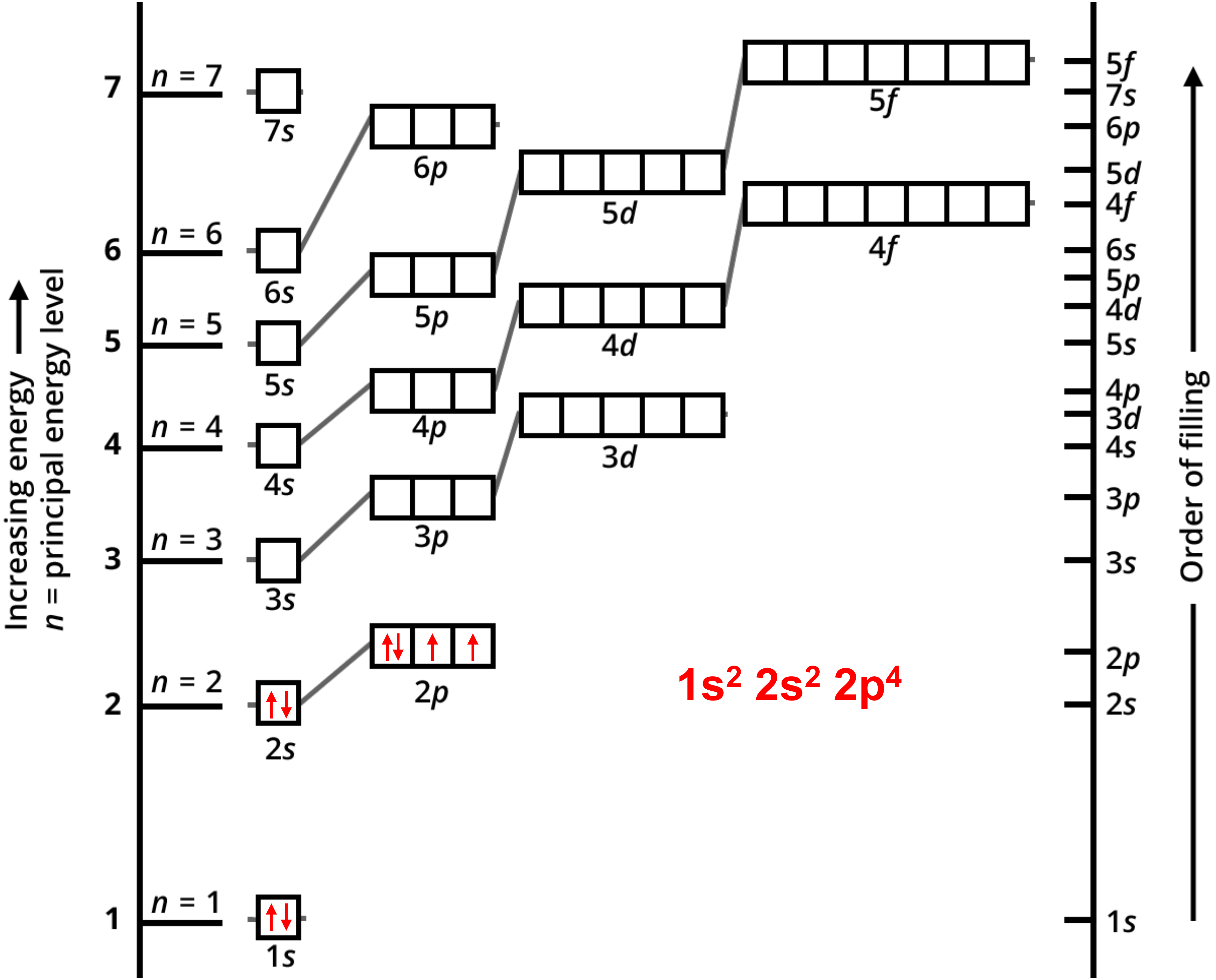
(Emission Spectra, Heisenberg Uncertainty, and the Photoelectric Effect)





Excited vs. Ground States

As I taught you an earlier video, when photons or other forms of energy pass through an element or molecule, it can “excite” electrons to a higher-energy state.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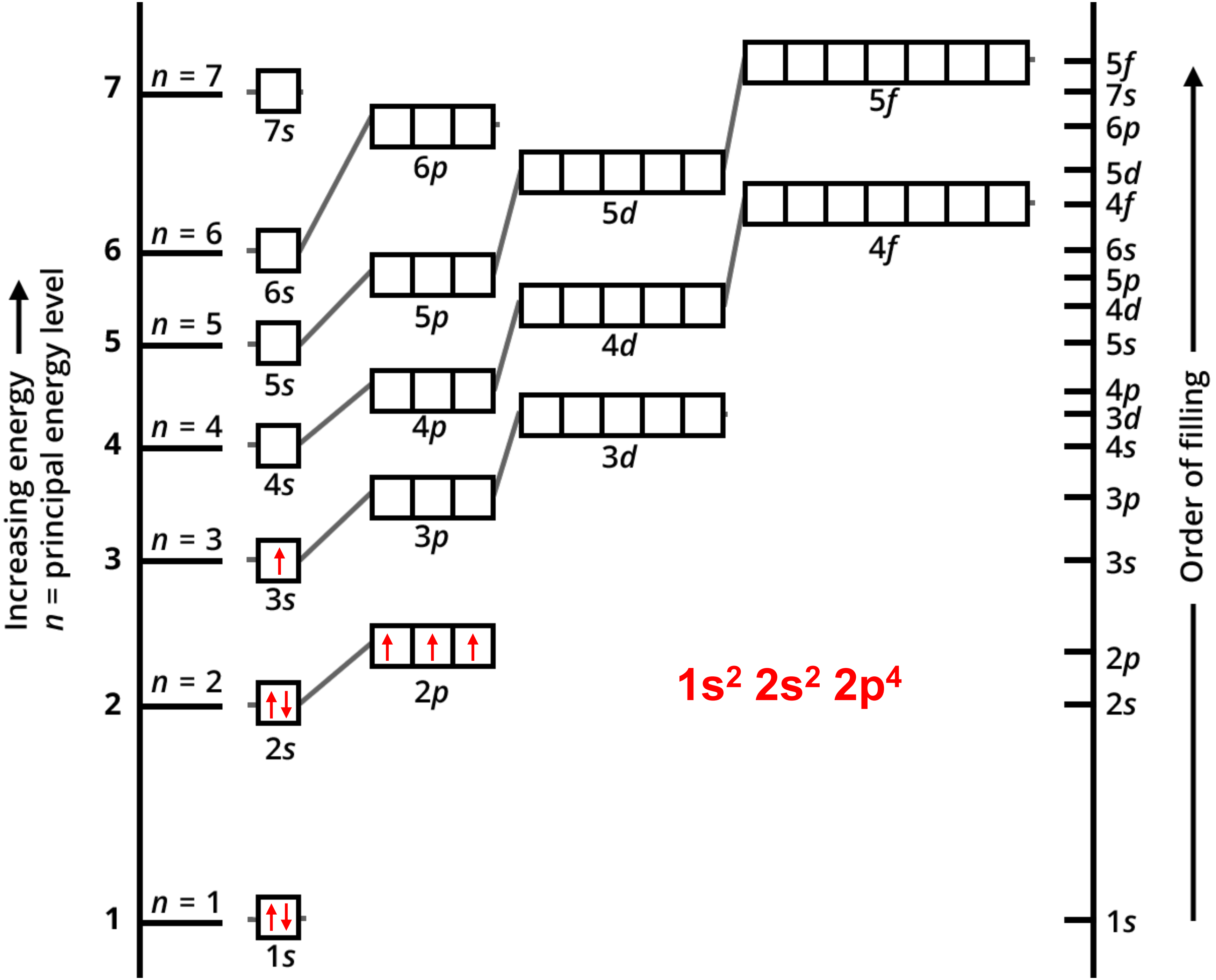
Order of filling



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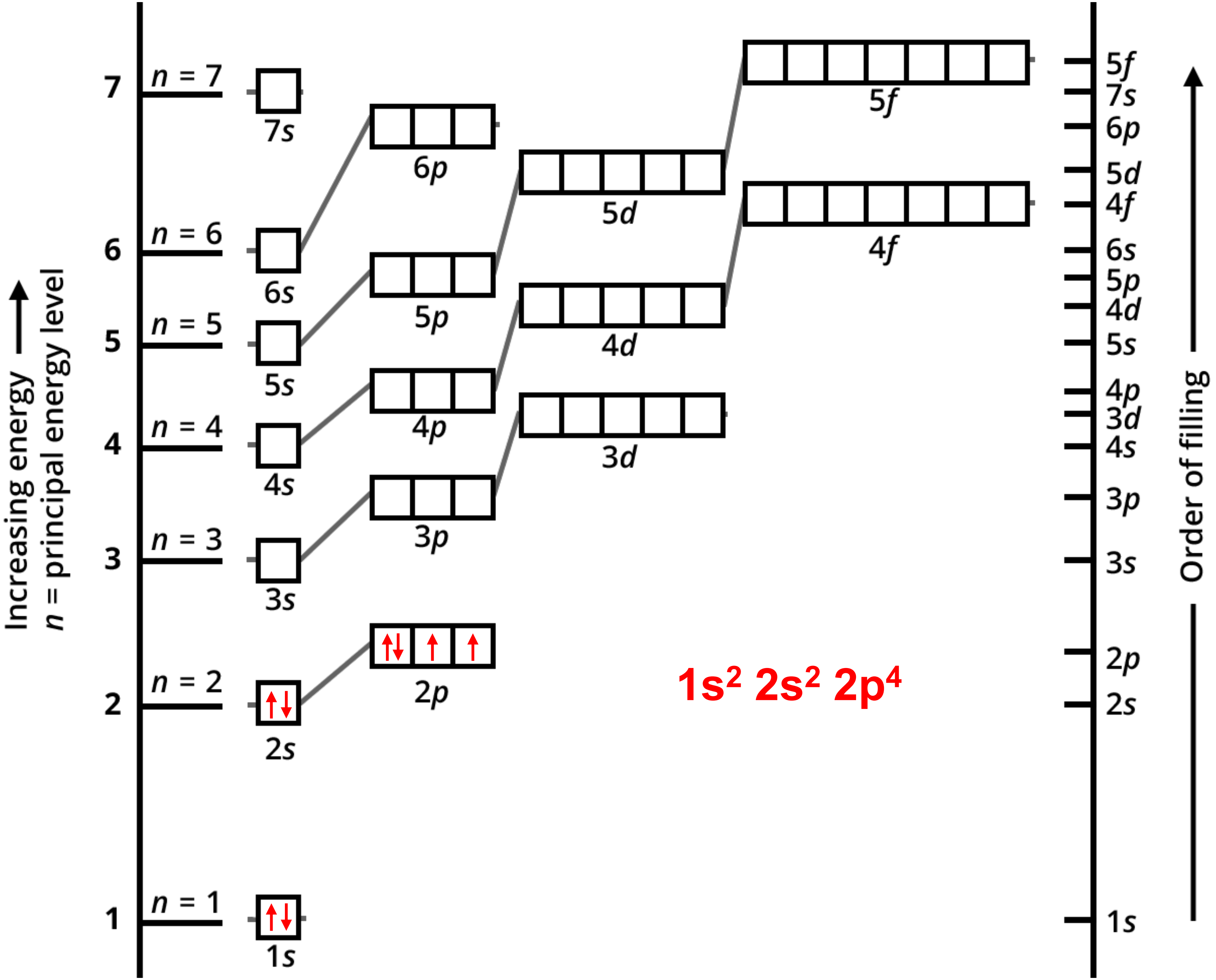

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4f
1s2 2s2 2p4
5f
7s
6p
5d
4f
6s
5p
4d
5s
4p
3d
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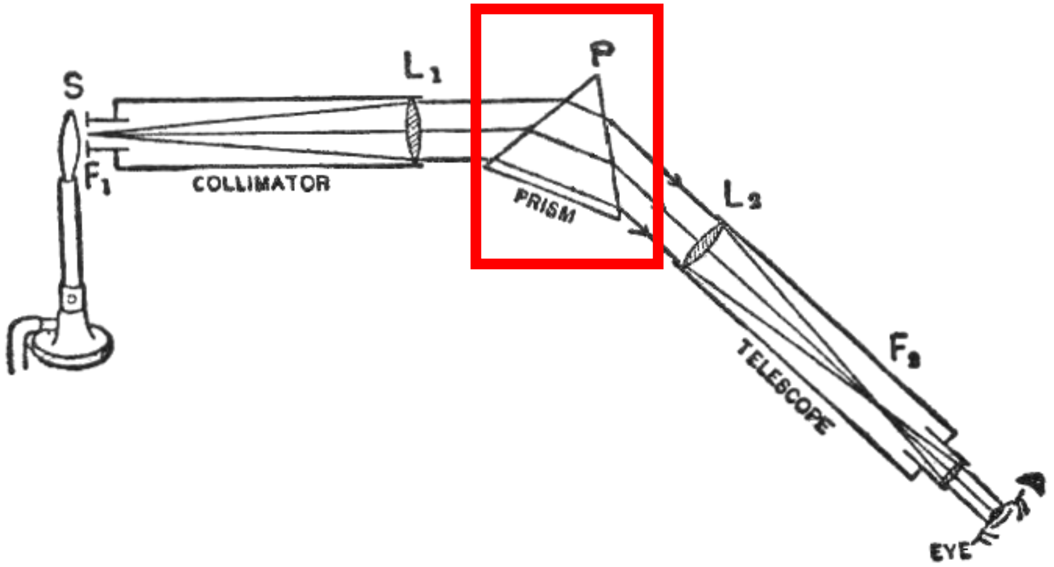
3s
2p
2s
1s
Order of filling

Excited vs. Ground States

When an electron “falls” back down to its original ground-state energy level, it emits colored light.

Why do we care?

Well, each element absorbs and emits different colors of light. We can look at those colors with a ***spectrophotometer***, which uses a prism to split the colors apart into their individual pieces. This gives us something called a ***line emission spectrum***, which is unique for each individual element. This ***line spectrum*** is kind of like an element's "fingerprint," which we can use to identify the element we're analyzing.

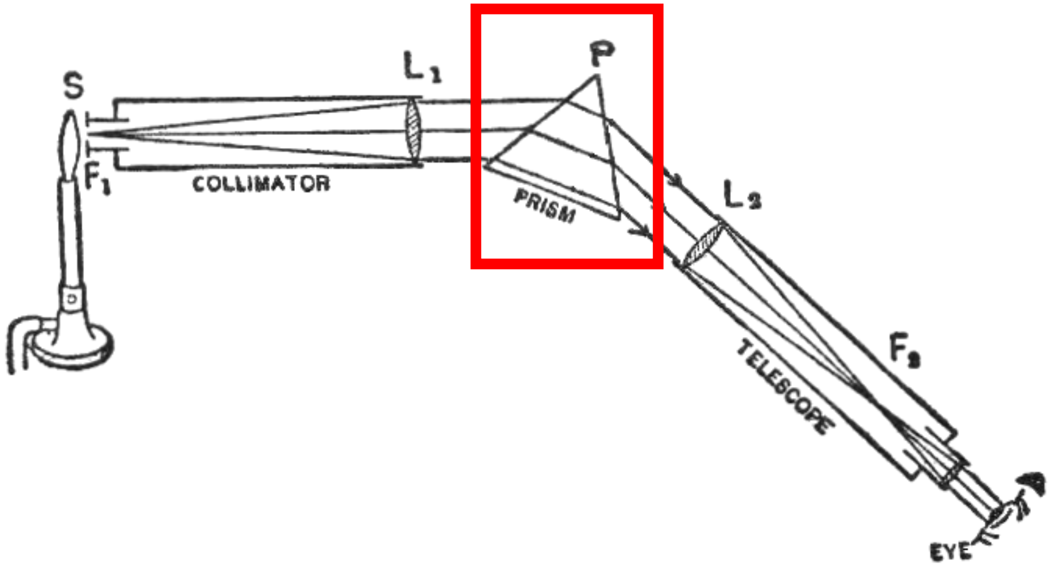


line emission spectrum for hydrogen

(by Jkasd - Own work with information taken from NIST., CC BY 3.0, https://commons.wikimedia.org/w/index.php?curid=8102837)

Why do we care?

For example, when the sun's white light is passed through a prism and analyzed in this way, we find that the resulting ***emission spectrum*** includes lines that match the ***emission spectrum*** lines of hydrogen gas. Thus, the sun must contain hydrogen gas. By extension, we can identify the elements in other stars and planets by looking at their ***emission line spectra***.



line emission spectrum for hydrogen

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A Photon's Energy

The energy of a photon (which you can think of as being a particle of light) can be calculated using the following equation:

$$E_{photon} = hf = \frac{hc}{\lambda}$$

Where:

f is the photon's ***frequency*** (this can be different for different photons)
c is ***speed of light***, which is **3.0×10^{8} *m/s***
h is ***Planck's constant***, which is **6.63×10^{-34} *J · s***
λ is the photon's ***wavelength***

You're not likely to get any calculation questions involving this equation on the EXAM.

A Photon's Energy

However, please remember the following:

Energy ≈ frequency ≈ 1/λ

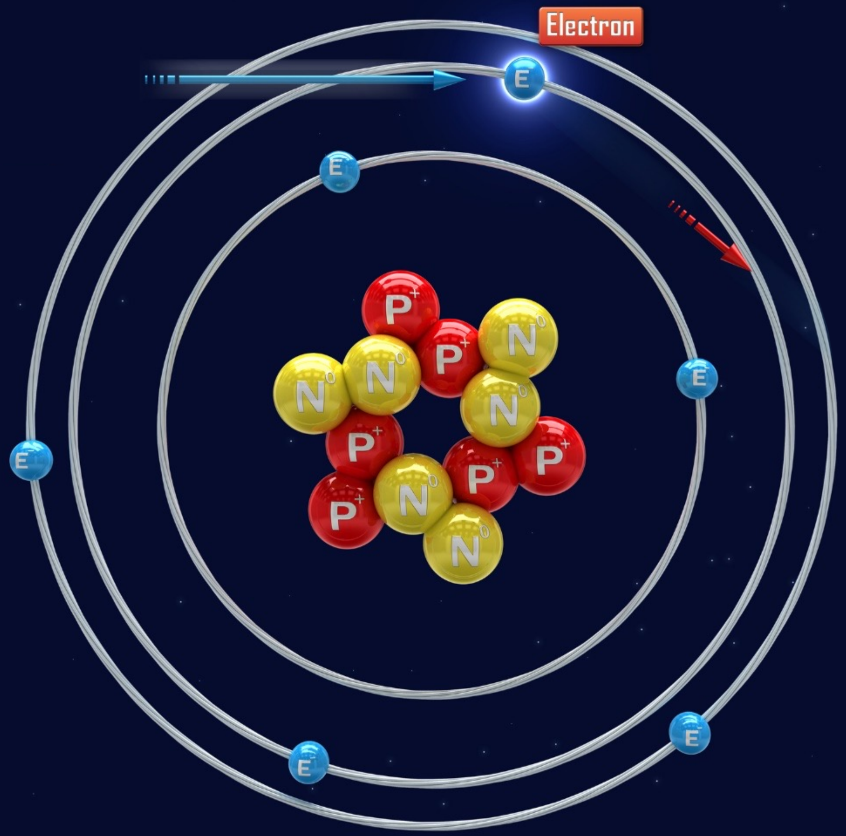
This means that:

↑Energy ≈ ↑frequency ≈ ↓λ

The Photoelectric Effect

If you shine a photon with enough energy at a metal plate, it causes the metal plate to eject electrons.

This is how solar panels work. Photons from the sun have a high enough energy to cause metal plates to eject electrons and produce electricity.



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The Photoelectric Effect

For the ***photoelectric effect*** to work, the metal has to be hit with a photon with enough energy.

$$\uparrow \text{Energy}_{\text{photon}} \approx \uparrow \text{frequency} \approx \downarrow \lambda$$

In other words, the photon must have a sufficiently ***high frequency*** and ***low wavelength***.

The Photoelectric Effect

For the ***photoelectric effect*** to work, we also have to care about this equation:

$$\textbf{Kinetic Energy}_{\text{electron}} = E_{\text{photon}} - \Phi$$

Where: Φ = work function (the minimum amount of energy required to ionize the electron). In order to expel an energized electron, the **Kinetic Energy**$_{\text{electron}}$ must be <u>greater than zero</u>.

The Heisenberg Uncertainty Principle

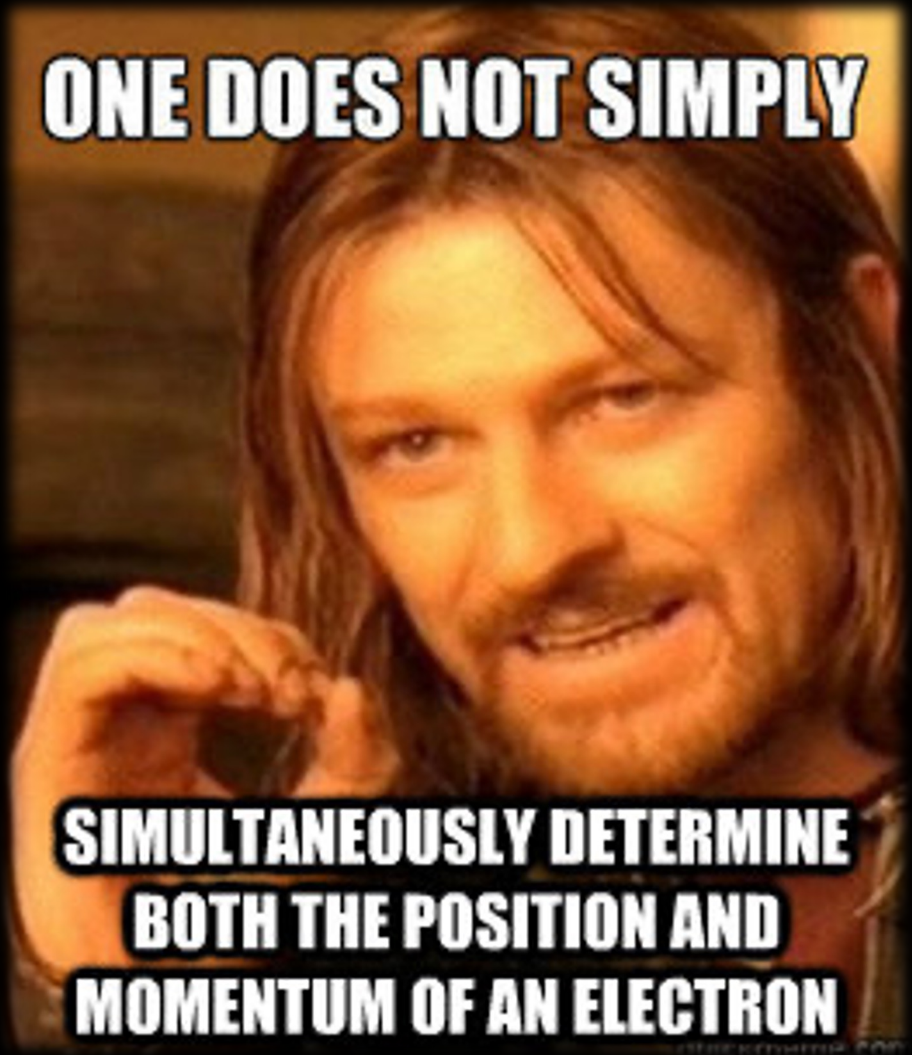
Subatomic (really tiny) particles, like neutrons, protons, and electrons, have some wavelike properties and particle-like properties.

The German physicist Werner Heisenberg proposed that as a result of this, it's impossible to determine a subatomic particle's location and its momentum with perfect accuracy. In other words, the more you know about its location, the less you know about its momentum, and vice-versa.

This is called the ***Heisenberg uncertainty principle***.

Werner Heisenberg

The Heisenberg Uncertainty Principle



http://www.quickmeme.com/Uncertainty-Principle/?upcoming

Question

Electronic transitions are accompanied by absorption or emission of light. Which electronic transition corresponds to emission of light with the shortest wavelength?

A. $n = 1 \rightarrow n = 5$

B. $n = 3 \rightarrow n = 1$

C. $n = 5 \rightarrow n = 4$

D. $n = 2 \rightarrow n = 3$

E. $n = 4 \rightarrow n = 3$

Question

Electronic transitions are accompanied by absorption or emission of light. Which electronic transition corresponds to emission of light with $\boxed{\text{the highest energy?}}$

A. n = 1 → n = 5

B. n = 3 → n = 1

C. n = 5 → n = 4

D. n = 2 → n = 3

E. n = 4 → n = 3

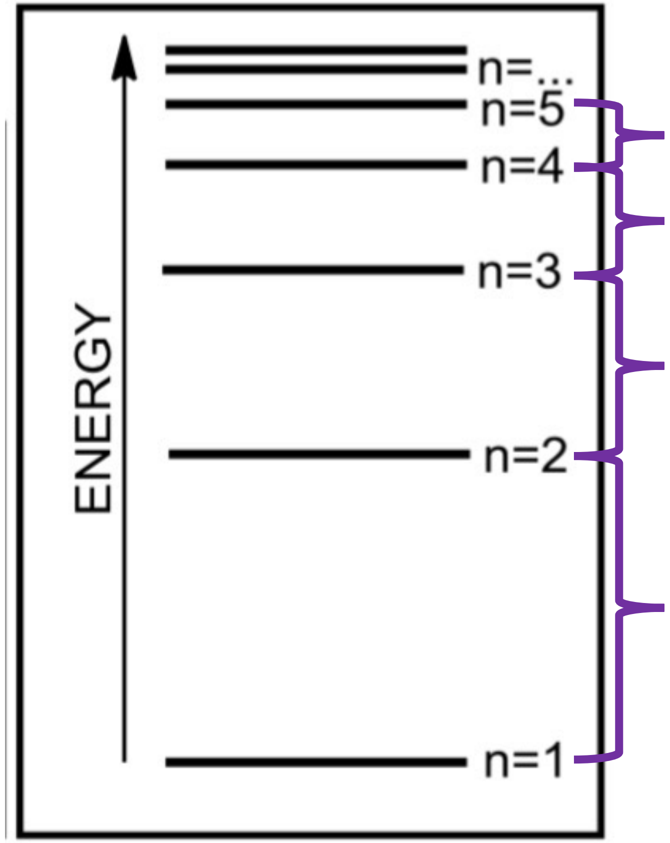
$$E_{photon} = hf = \frac{hc}{\lambda}$$

$$\uparrow \text{Energy} \approx \uparrow \text{frequency} \approx \downarrow \lambda$$

Question

The distance between each energy shell and the *next* shell above it gets smaller as you get further away from the nucleus.

Thus, there is a larger energy difference between an electron that shifts from ***n*=2** down to ***n*=1** than (for example) an electron that shifts from ***n*=5** to ***n*=4**.



Question

- You have a photocell whose work function value equals 3 eV. If this photocell is struck with a photon that has an energy of 2 eV, will it produce a current?

- In this circumstance, if your photocell is struck by a photon with an energy of 5 eV, what will the resulting energy of the expelled electron be?

The Photoelectric Effect

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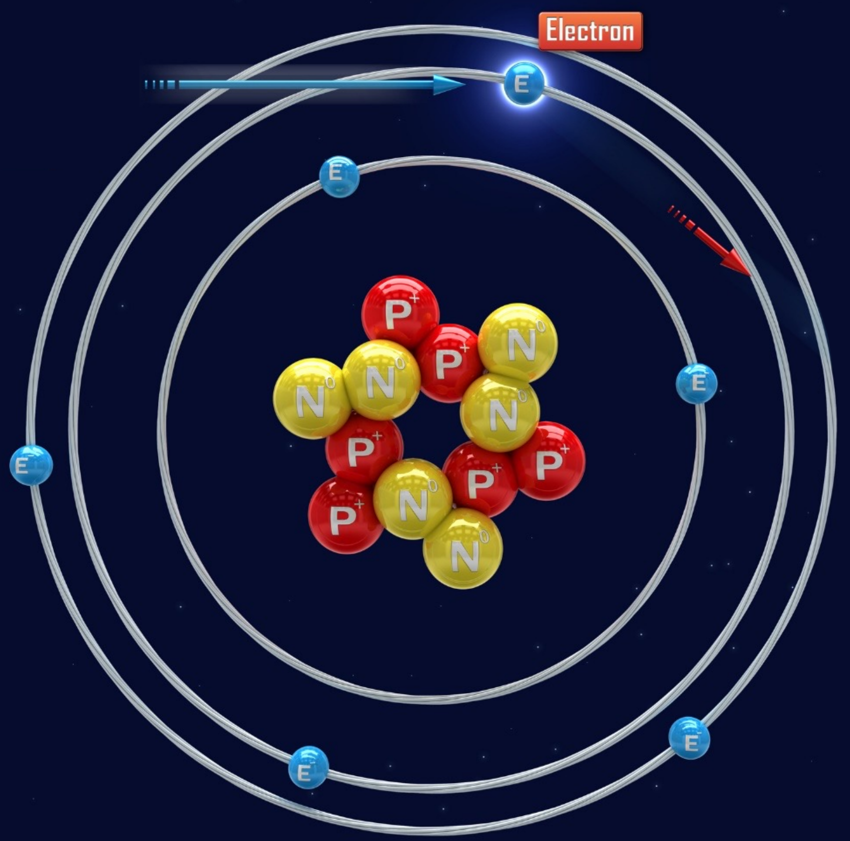

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