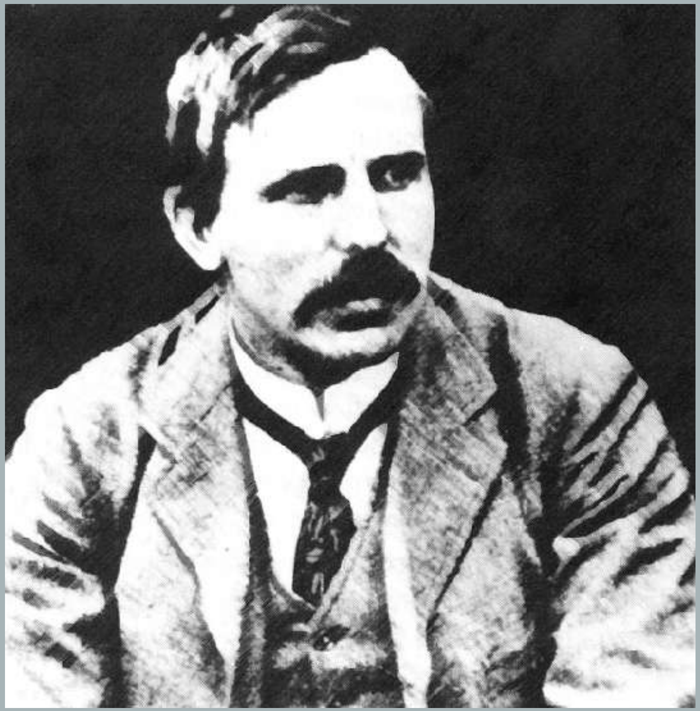
Chapter 11
Modern Atomic Theory

This chapter helps us understand why chemicals do the things they do!

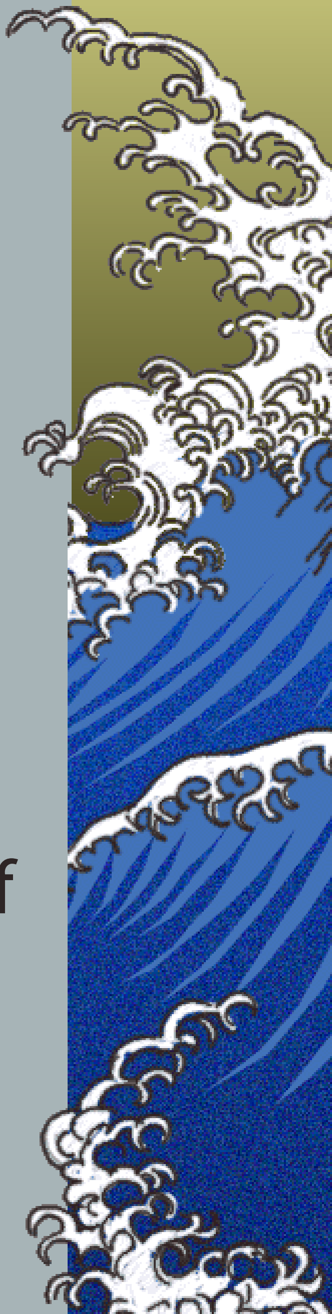
Chapter 11
Modern Atomic Theory

- *What is the structure of an atom?*
- how are elements grouped (columns) on the Periodic Table?
- It has everything to do with ***electrons*** and how they are arranged

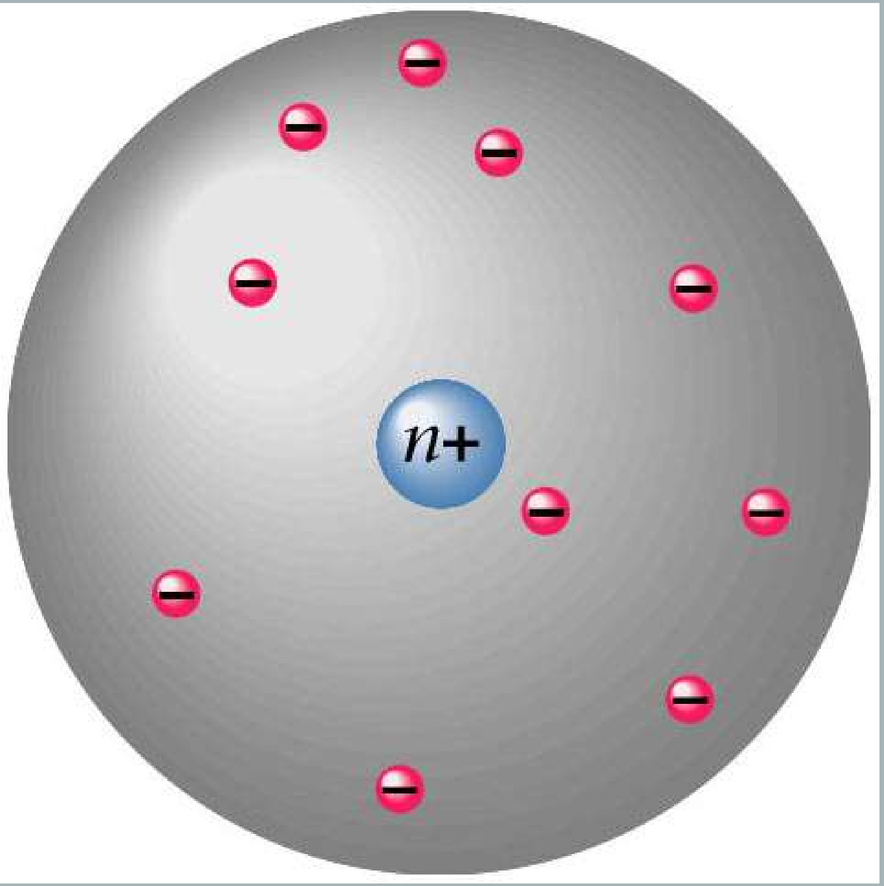
11.1 rutherford's atom



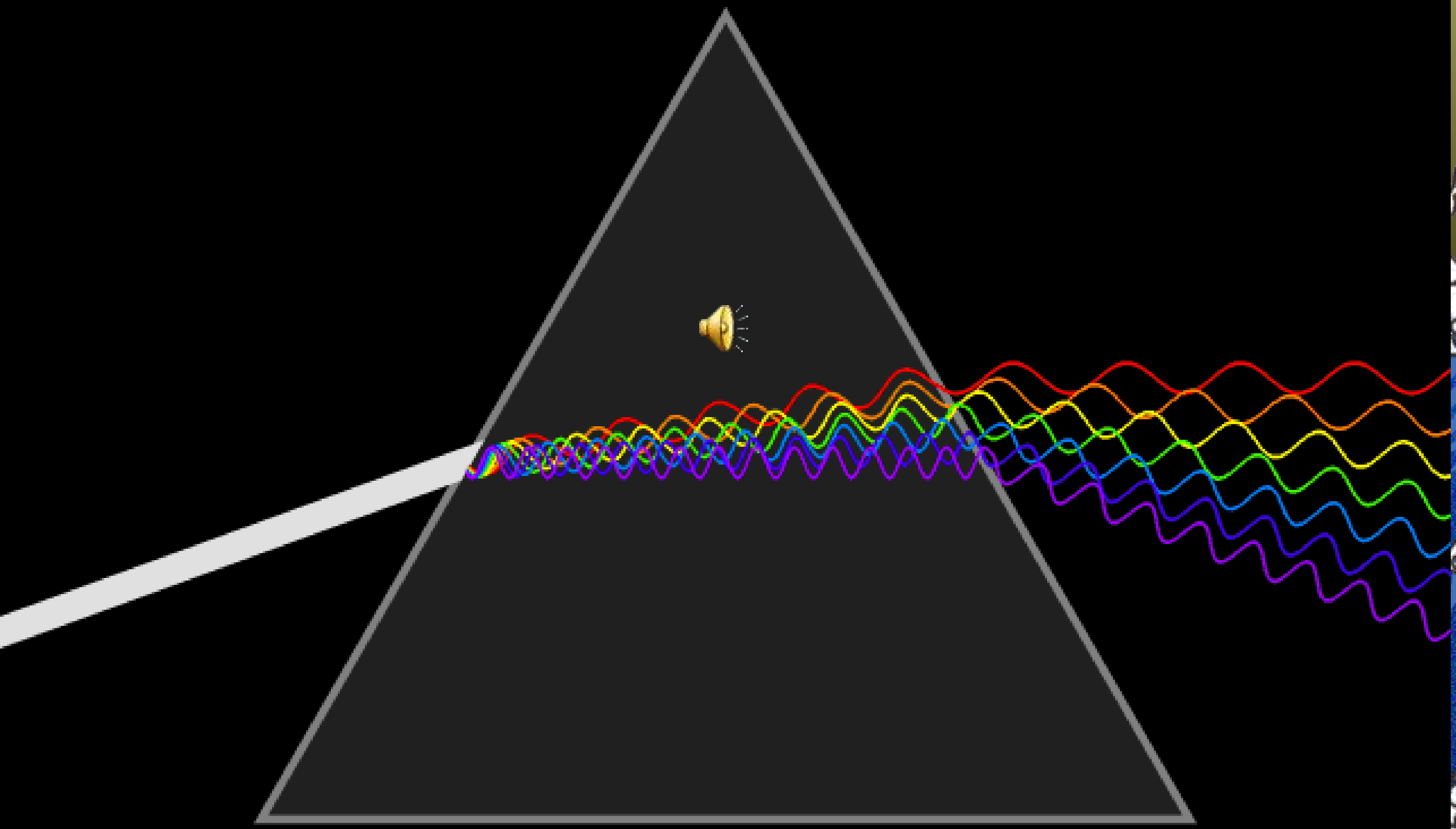
- Remember: **Rutherford** and his buddies (in Ch 3) found that the atom had a nucleus with electrons on the outside
- The nucleus was very small and was made of protons and neutrons
- The e- made up the rest of the atom…



- *But what were the e- doing? How were they arranged? Why didn't they just crash into the nucleus?*
- Something more was needed!!!
- We need enl**ight**enment!

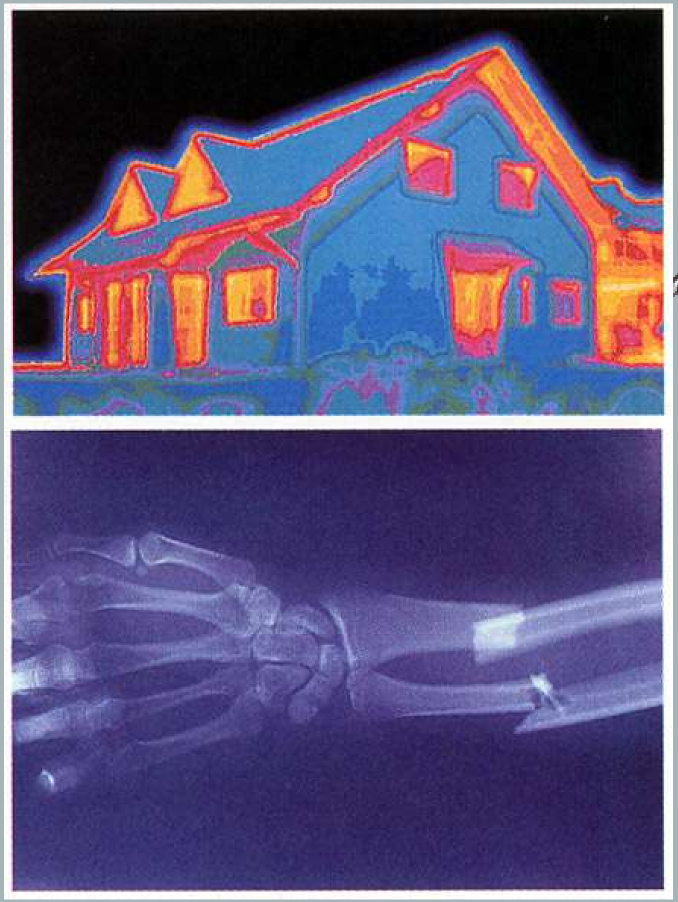


11.2 energy and light

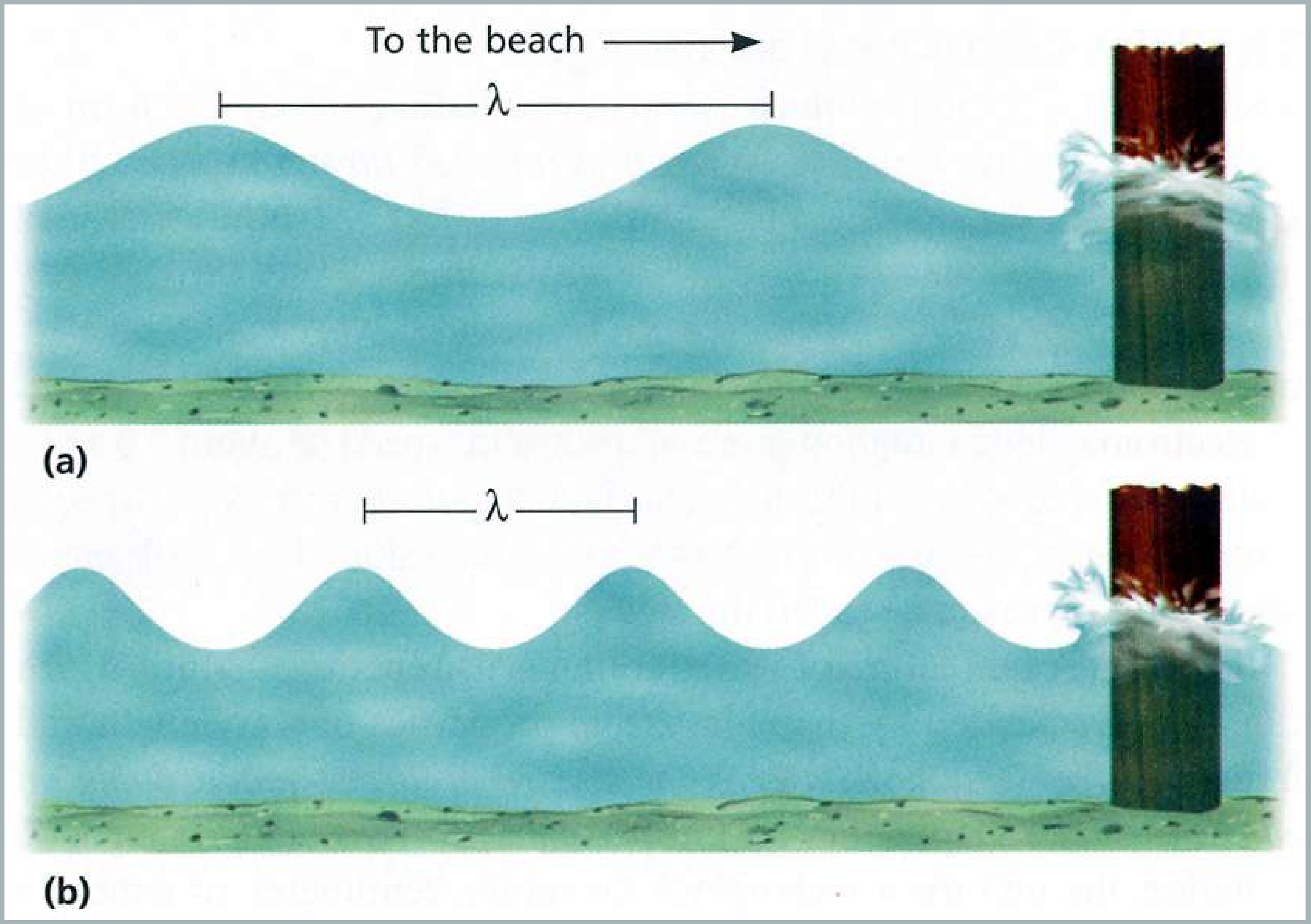


11.2 energy and light

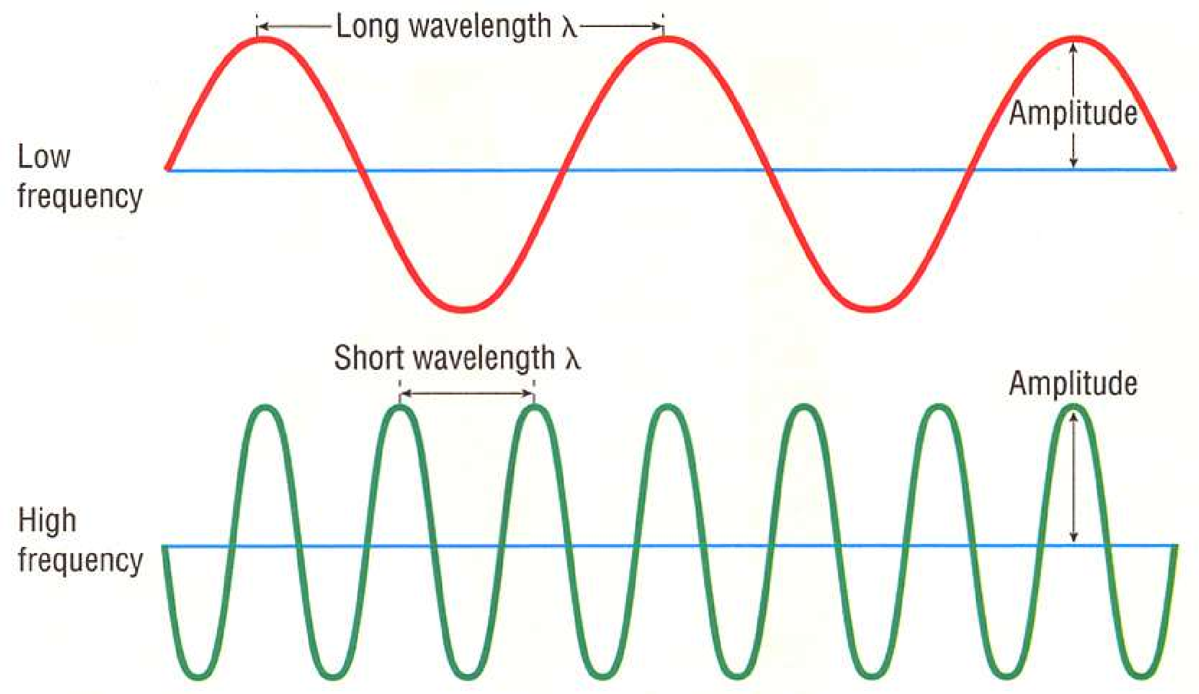
- When Energy is transmitted from one place to another *by light* we call it **Electromagnetic radiation**
- There are several types of electromagnetic radiation: light bulbs, fire places, the sun, X-rays, microwave ovens, you!
- what’s the difference between them?



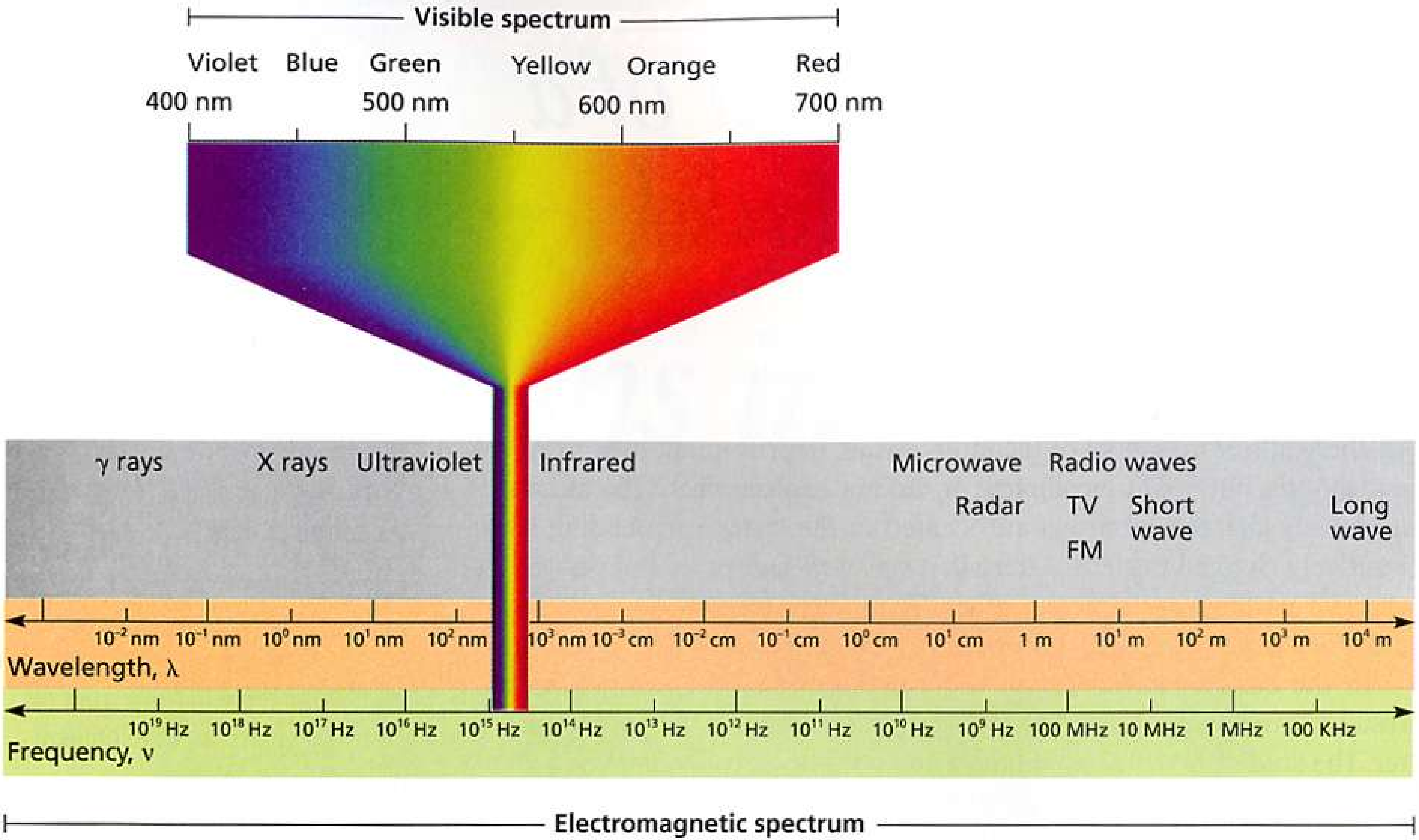
- It has to do with **waves** and their properties
- Waves have three main characteristics: Wavelength, Frequency, & Speed…



- **Wavelength:** (λ) is the length of one full wave cycle
- **Frequency:** (ν) is how many waves pass a given point per given time
- **Speed:** how fast a wave travels
- Light does a similar thing to water waves; it travels at light speed with λ and ν…



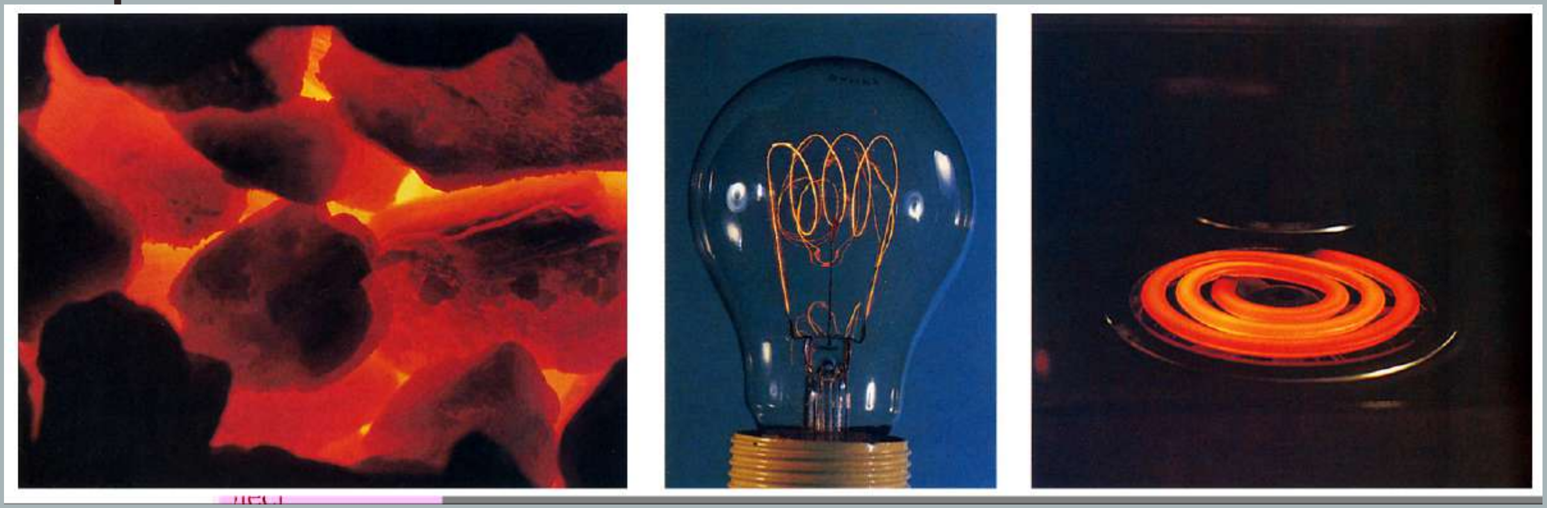
- What is a wave?



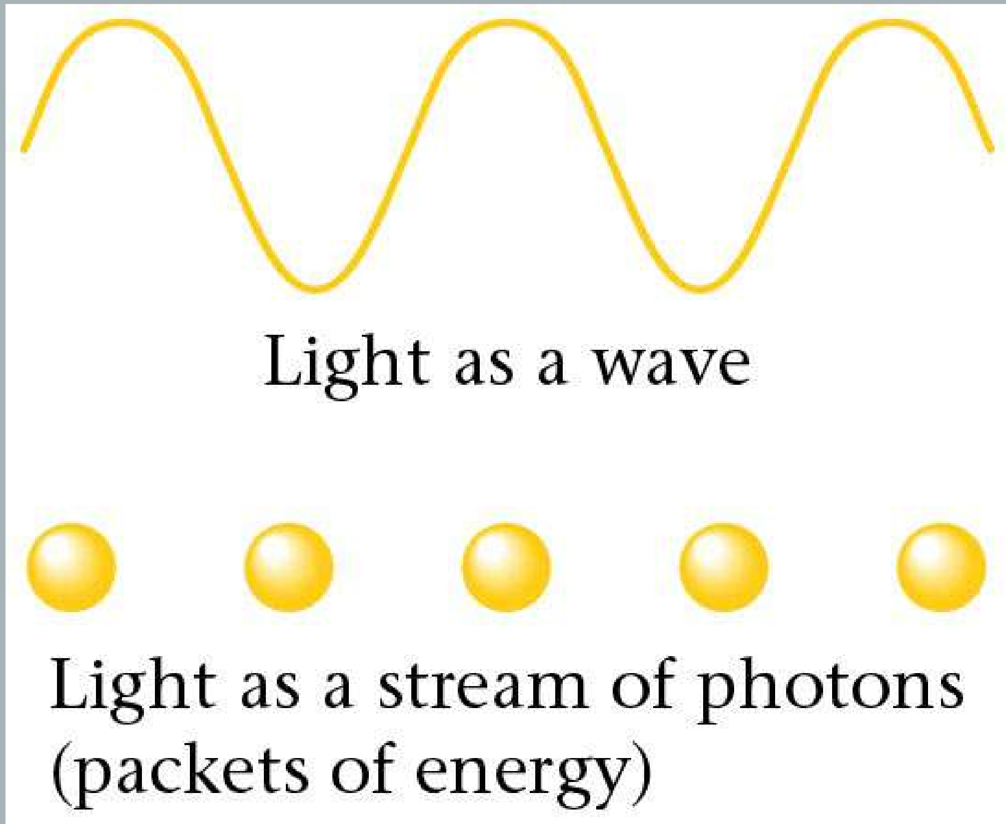
The ELECTROMAGNETIC SPECTRUM!!

Which of the above transfer energy?

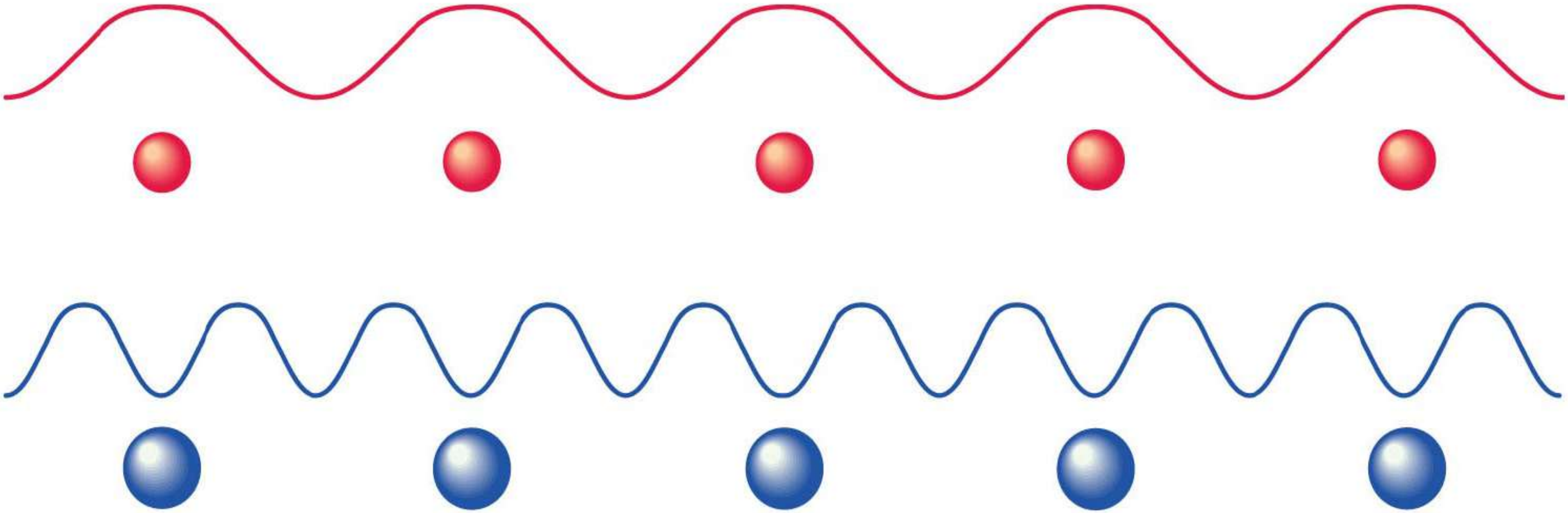
- The EM radiation mentioned earlier all have their λ and ν, and all transfer Energy from one place to another - even through space!



- a wave consists of discrete packets of energy, or quanta



- The light "particles" are called **photons**
- So is it a ***wave*** or a ***particle***?
- Both: **wave-particle nature of light**



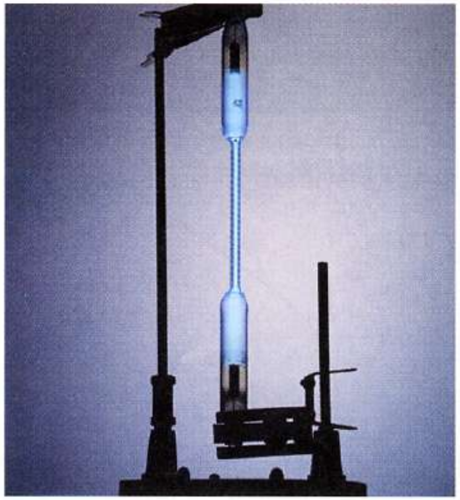
- Different wavelengths carry different amounts of energy, blue more than yellow more than red…
- X-rays more than uv, visible more than IR, IR more than radio, etc.

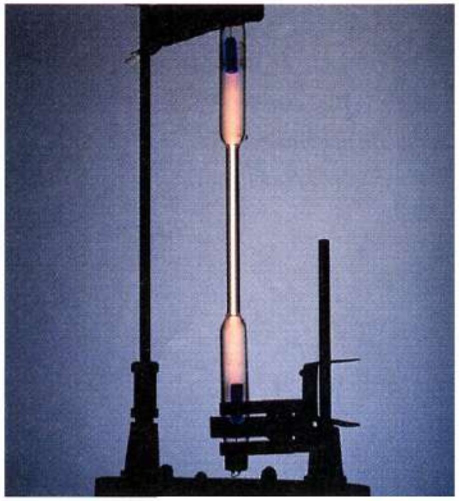
11.3 Emission of energy by atoms

- Why the different colors here?

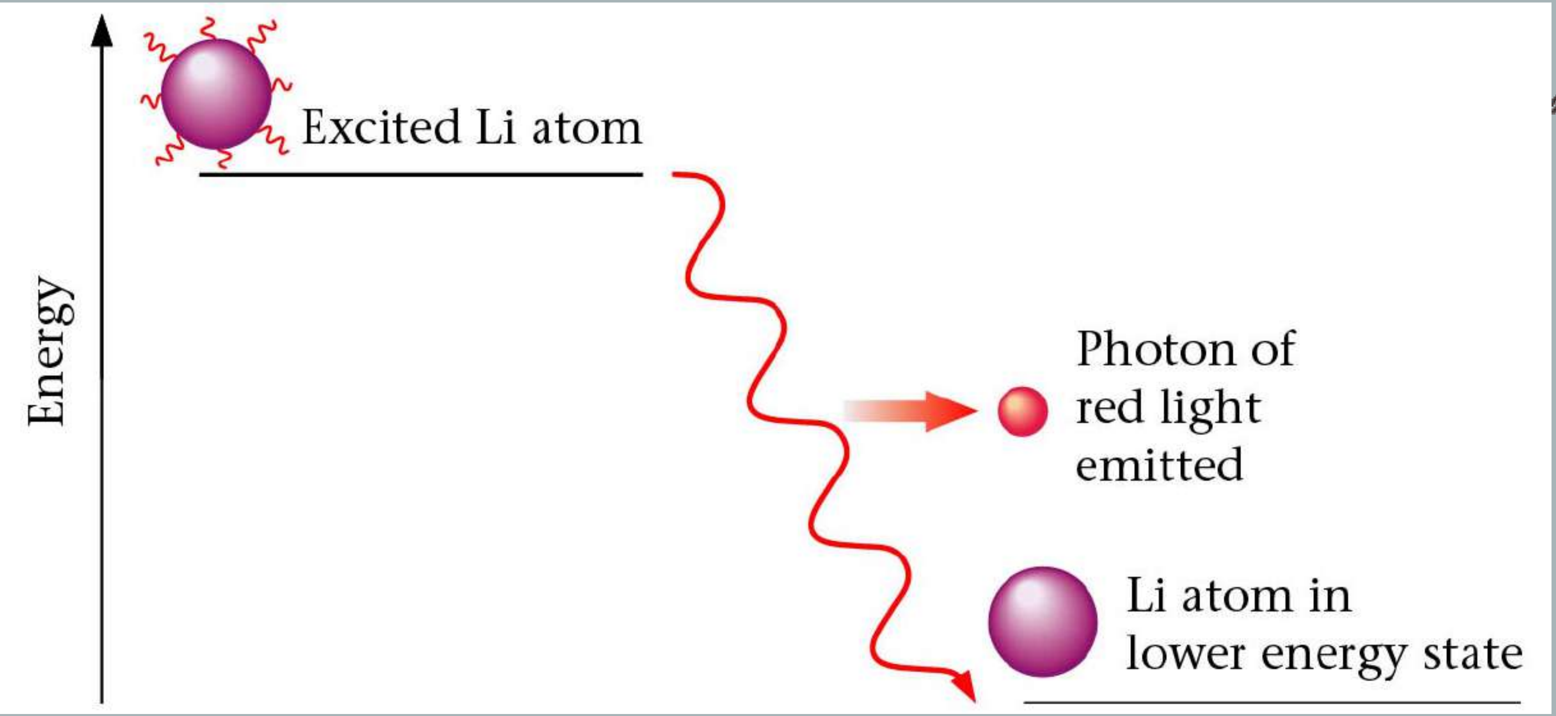
Flame Tests


Sodium light


Mercury light


Nitrogen light

- *they only light up when they receive energy* (e.g. flame or electricity); but why?
- ***Absorb*** the energy
- when they release that same amount of energy, they ***give off in the form of light***



- Li is gives off reddish light

Cu would give off green

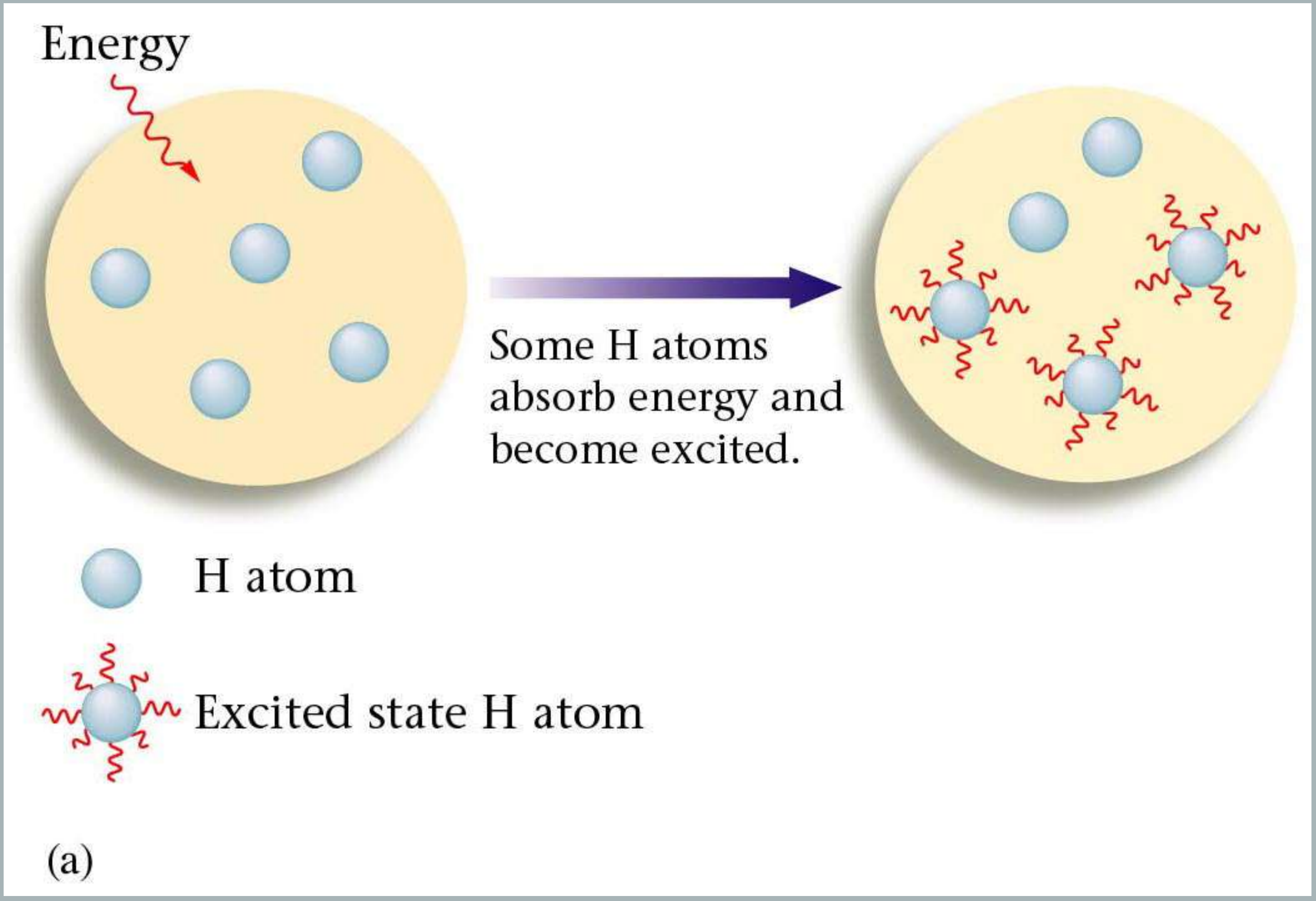
sodium yellow

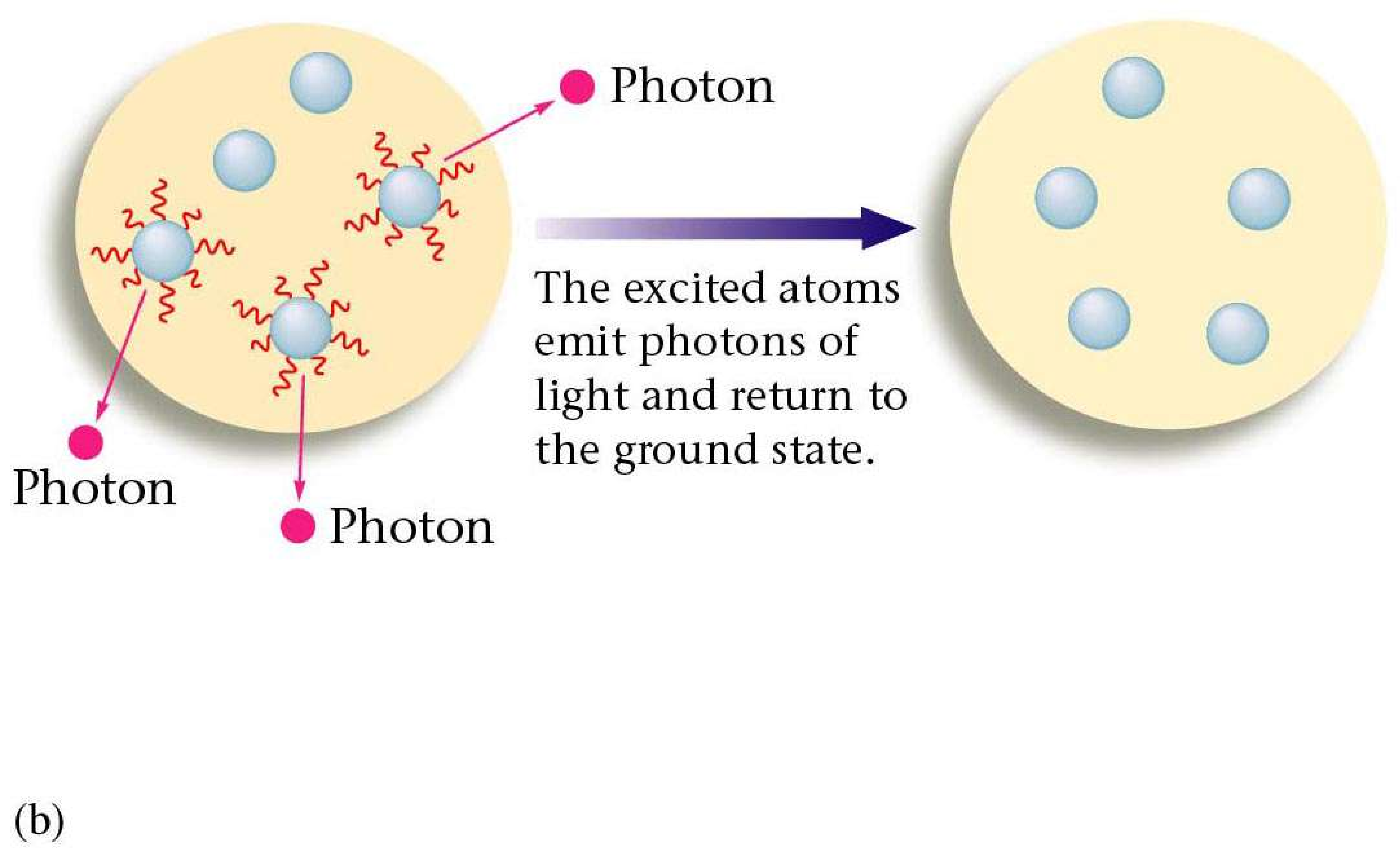
- why?
- *It has to do with how much energy an atom can absorb and then release*

11.4 the energy levels of hydrogen

- an atom with excess energy is said to be ***excited***
- When it emits the photon it goes back to unexcited state called ***ground state***

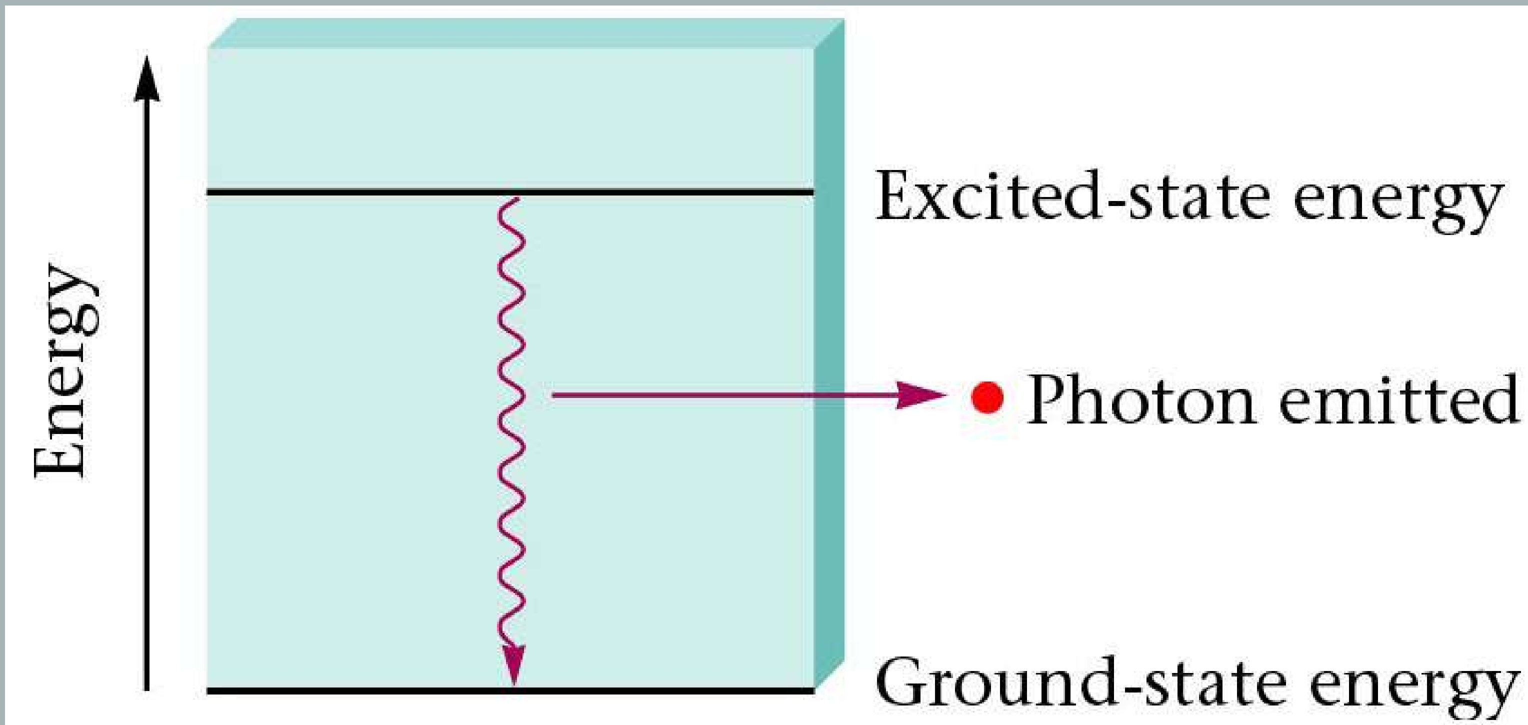
- Now we look specifically at **hydrogen**, but remember: *different wavelengths carry different amounts of Energy per photon*



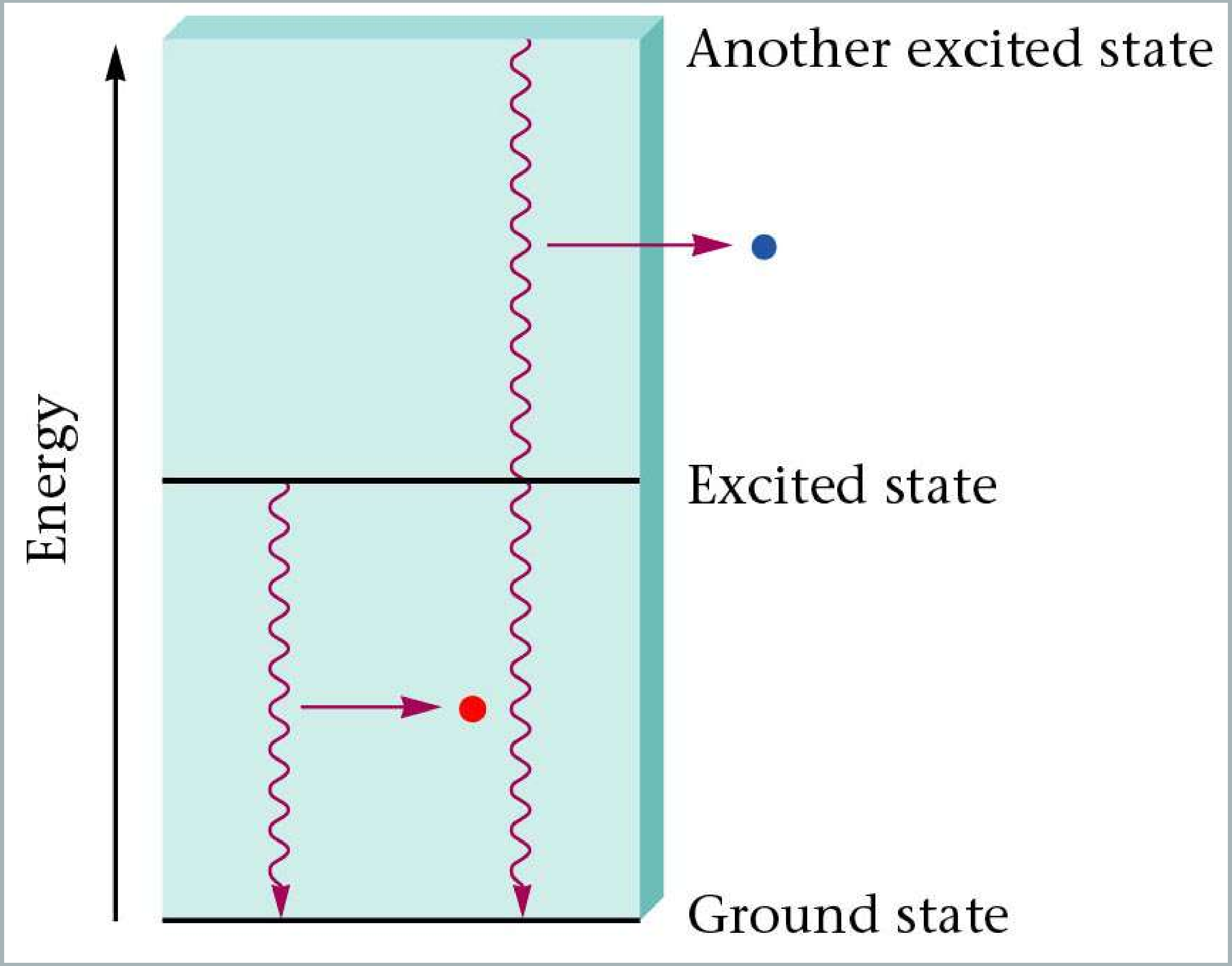


Photon
The excited atoms emit photons of light and return to the ground state.
Photon
Photon
(b)

- another way to look at it…



- Important point! T*he Energy contained in the photon corresponds exactly to the* ΔE *that the atom experiences*

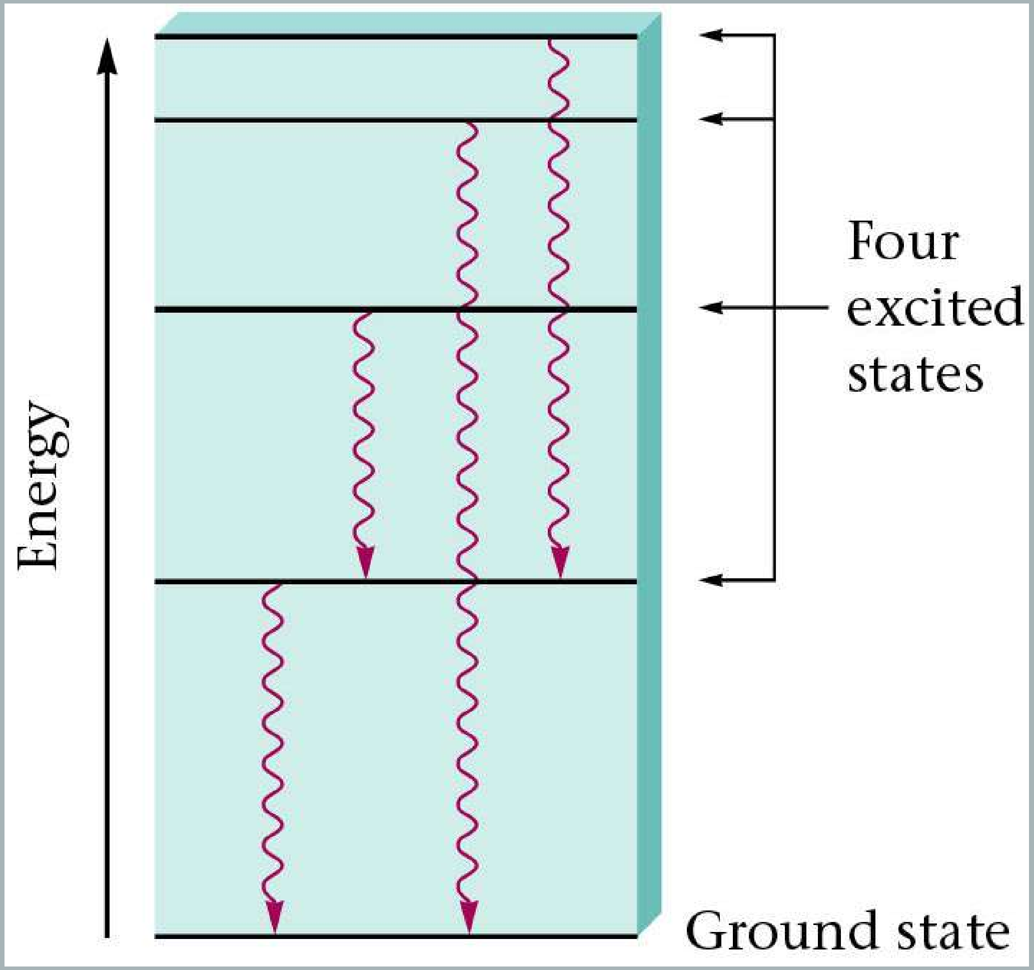


- notice that the **color** of the photon is related to **how much nrg** was given off

- when we put Energy into a hydrogen sample, and look at the light produced, specific wavelengths shine through

- there must be more than one energy level!

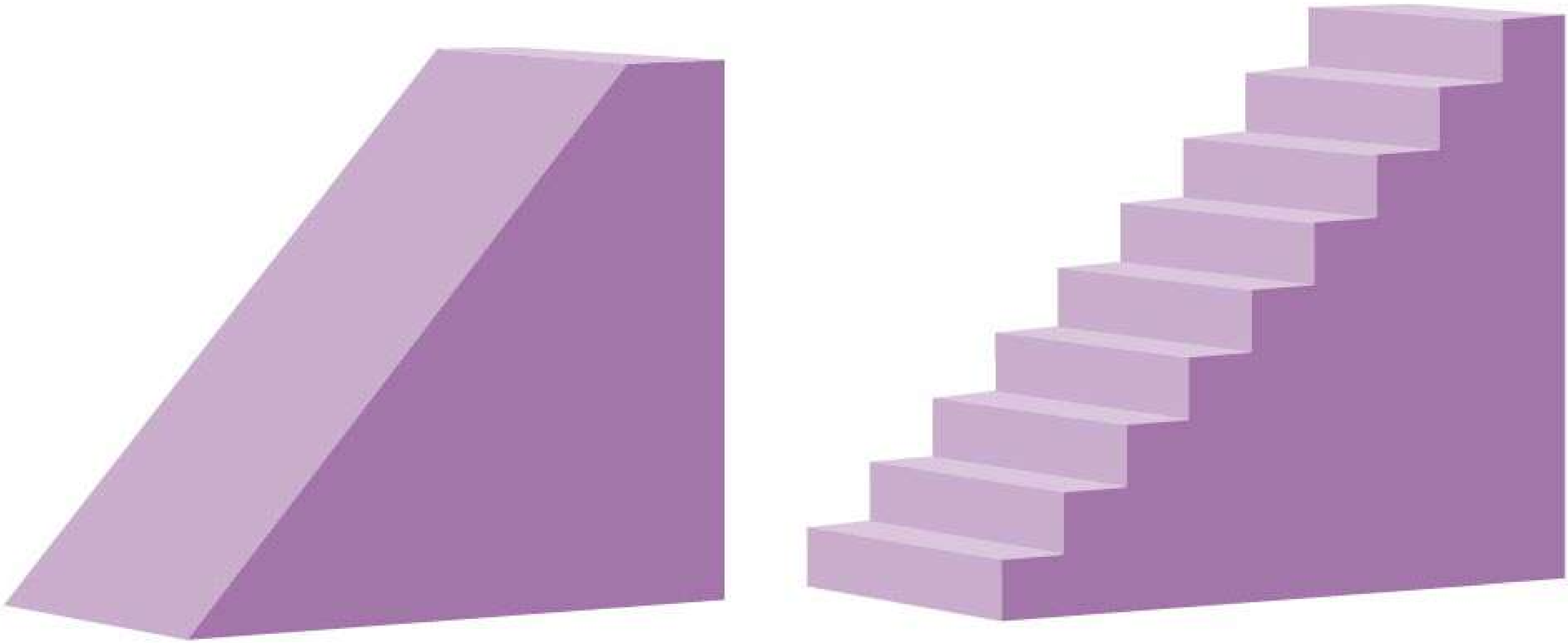
- ***Big picture for H?*** The different wavelengths mean there must be several ways for the e- to get back to ground state!



- That hydrogen has only *certain allowed* ways for the electron to return to ground means the energy levels **are not continuous (a)**
- **The energy levels are distinct steps, they are: quantized** (b)

quantum (plural: **quanta**)

- It is the fundamental notion that a physical property may be "quantized", or "quantization".
- This means that the magnitude can take on only certain discrete numerical values, rather than any value, at least within a range.



(a) (b)

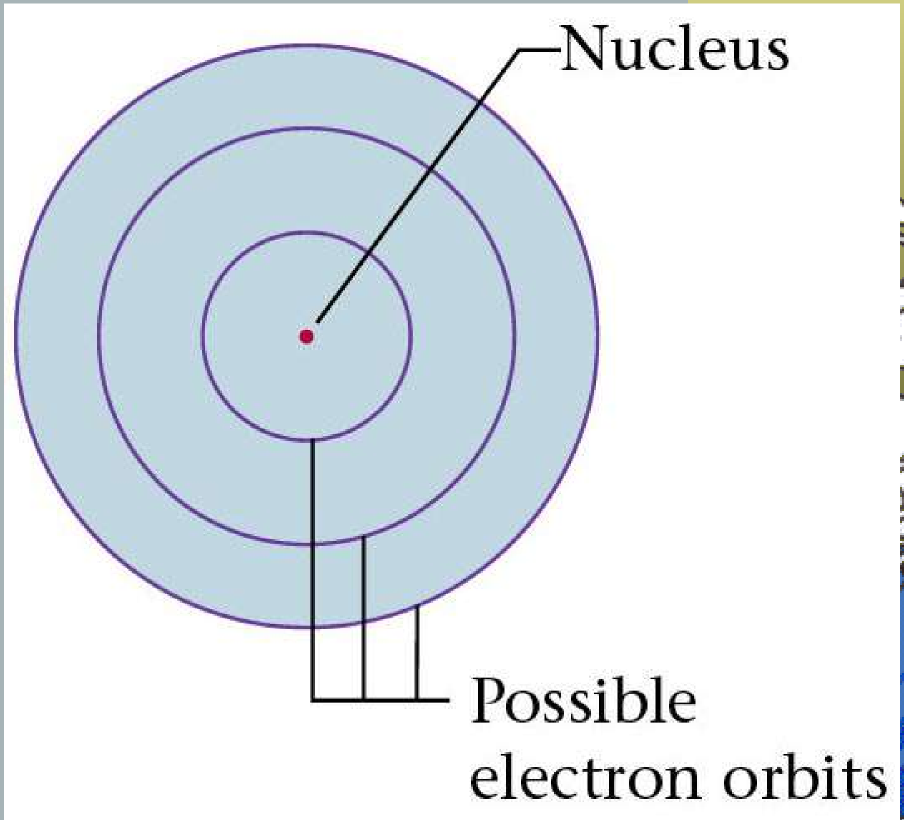
- One can “exist” anywhere on the ramp
- There are discrete locations (steps) that can be occupied on the stairs, one cannot stand between steps

- http://ioannis.virtualcomposer2000.com/spectroscope/amici.html#colorphotos

11.5 the bohr model of the atom

- depicts the atom as a small, positively charged nucleus surrounded by electrons that travel in circular orbits around the nucleus—similar in structure to the solar system, but with electrostatic forces providing attraction, rather than gravity.

- The e only orbit the nucleus in certain orbits
- When the e- gained energy it went up to a higher level
- When it gave off its energy it fell back down, giving off a photon in the process
- this actually worked well for hydrogen, but…



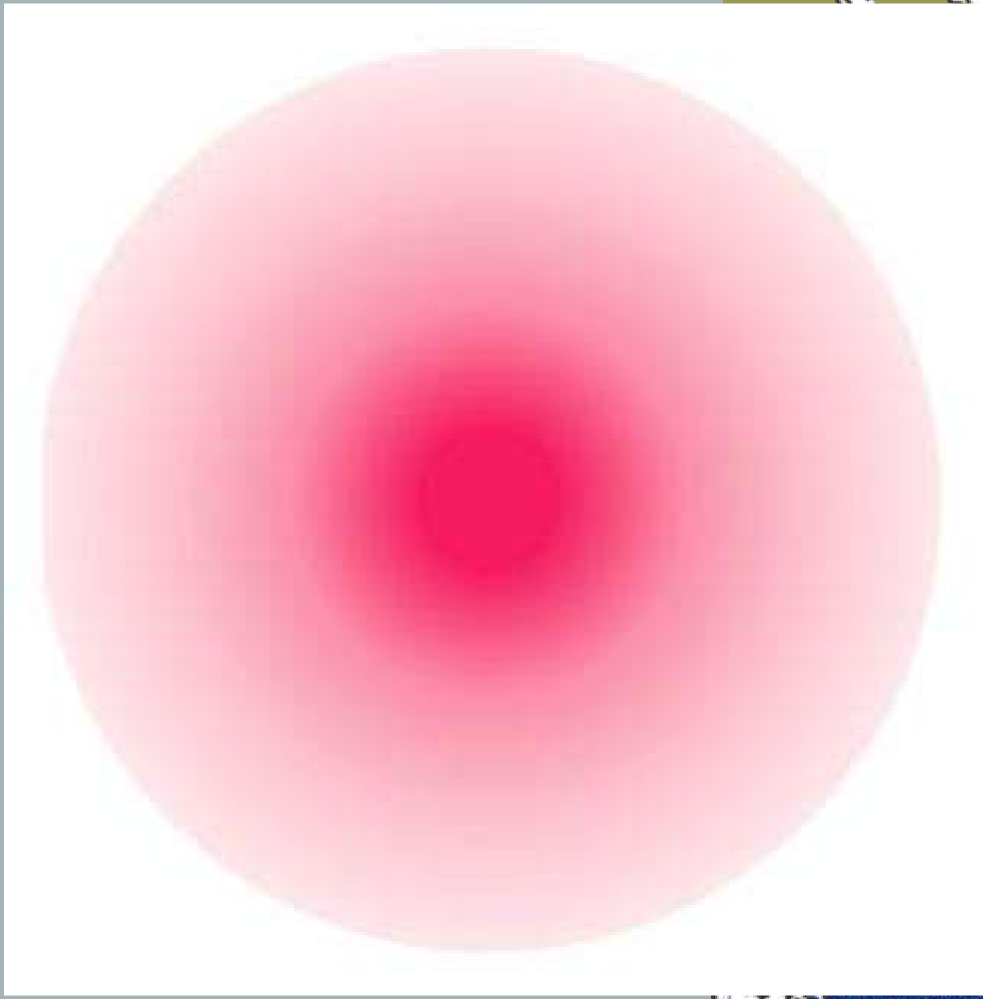
- **It didn’t work for *any other atom***
- In fact, it turns out e- don’t move around the atom like planets, either
- Despite the fact that it ***seemed*** nice and elegant, a better model for the atom was needed…

11.6 the wave mechanical model of the atom

- We needed a radical new approach to looking at the atom
- By the mid-1920’s that would happen…

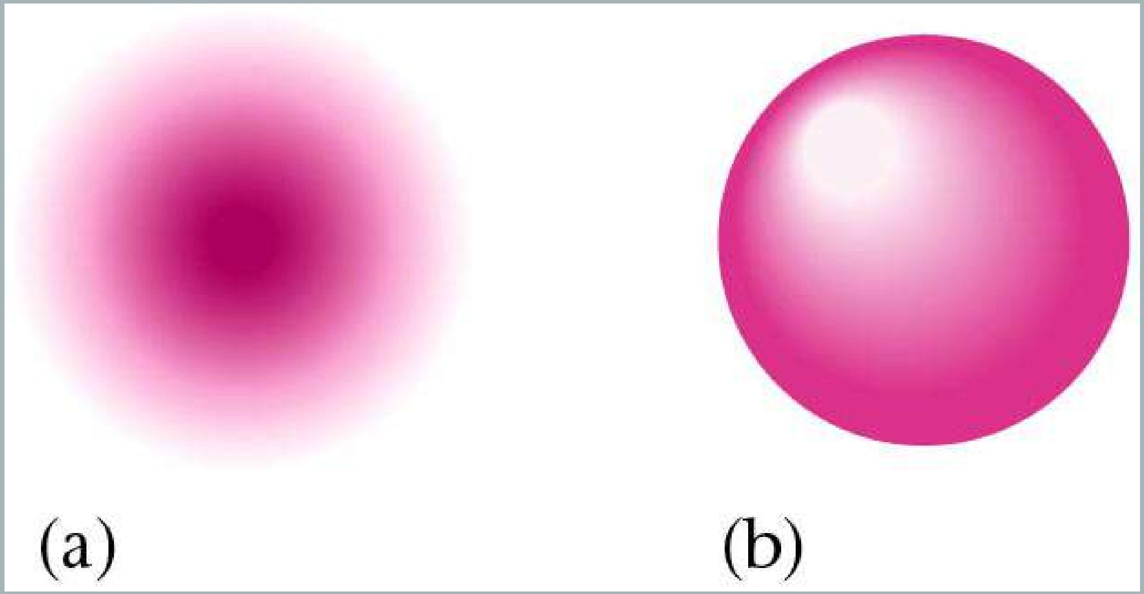
- Schrödinger said maybe we should look at the electron not as a particle ***but as a wave!***
- Schrödinger even developed an equation to describe what the wave-electron was doing
- called **wave mechanical model** of the atom…

- According to the wave-mechanical model, the hydrogen electron's lowest energy can be pictured like this →
- Schrödinger found that he could not tell exactly where the e- was or where it was going, only where it probably is
- Those areas of probability are called: **orbitals**



11.7 the hydrogen orbitals

- The fuzzy drawing we see here just represents where the electron probably is
- It's easier to draw it as a sphere, which represents the volume in space where there is a 90% chance of finding it



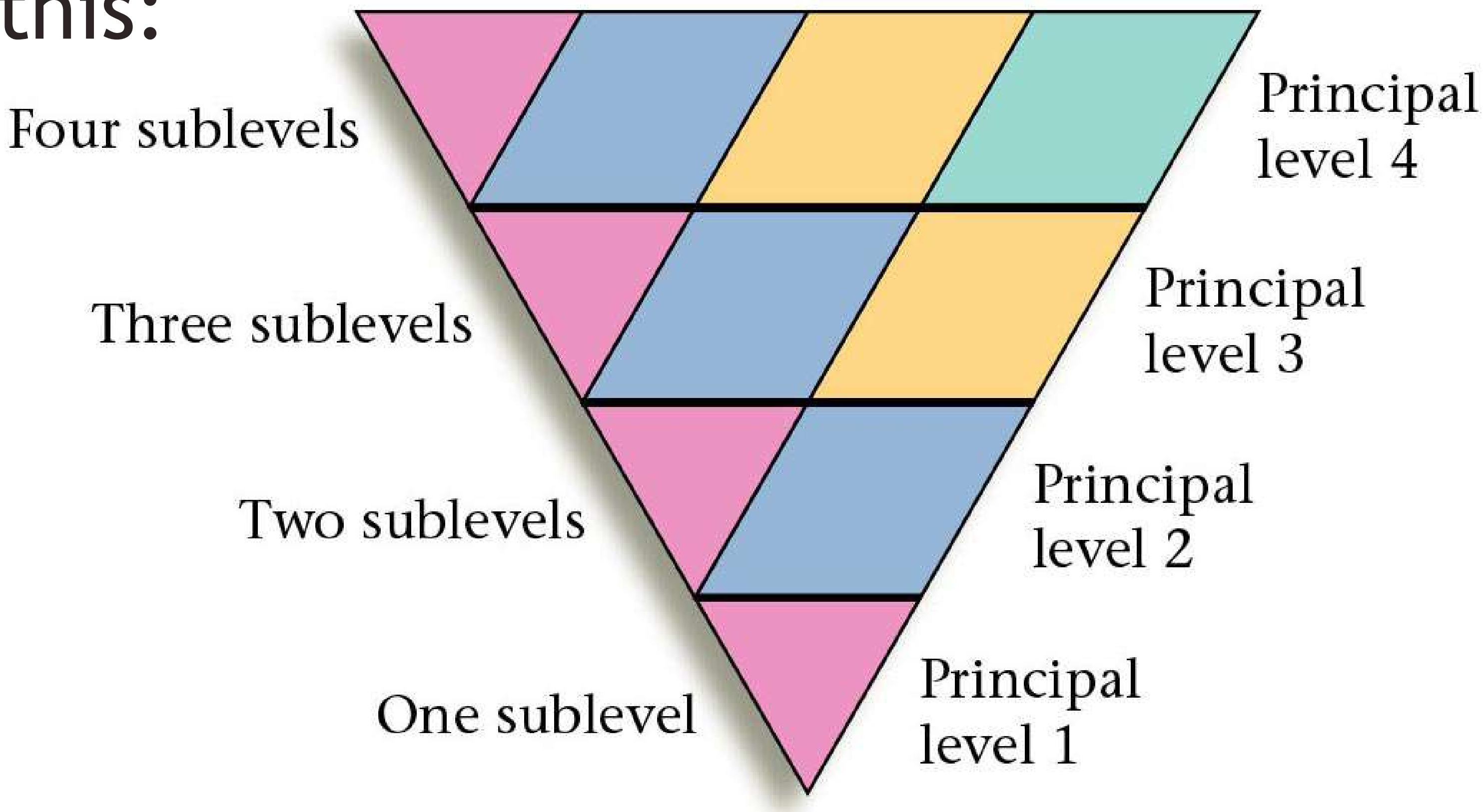
- This orbital has a name: **1s**
- It is hydrogen's lowest energy state
- What happens when the e- goes into an excited state?
- first…

The S Orbital

Hydrogen energy levels

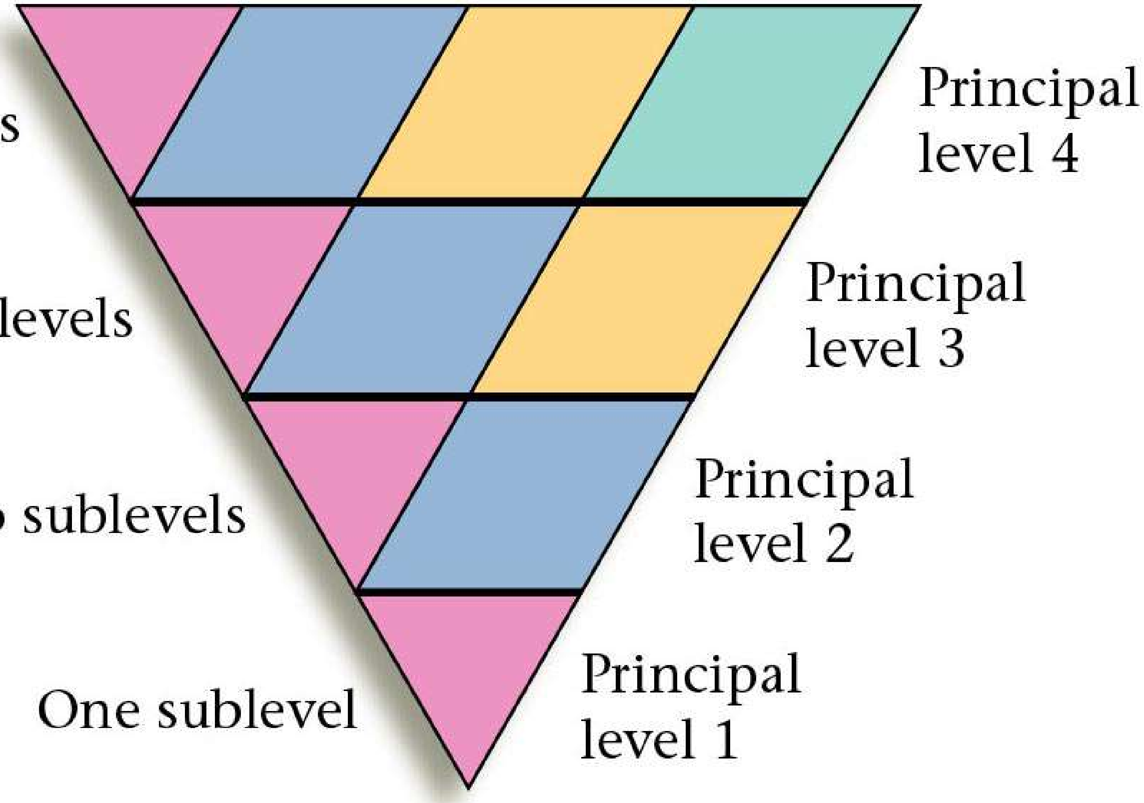
- Remember that the H atom has discrete E levels? They have a name…
- = **principal energy levels**
- Labeled by integers from n = 1 $\rightarrow \infty$
- Each level has **sublevels**, like rooms on a floor

- It can be pictured like this:



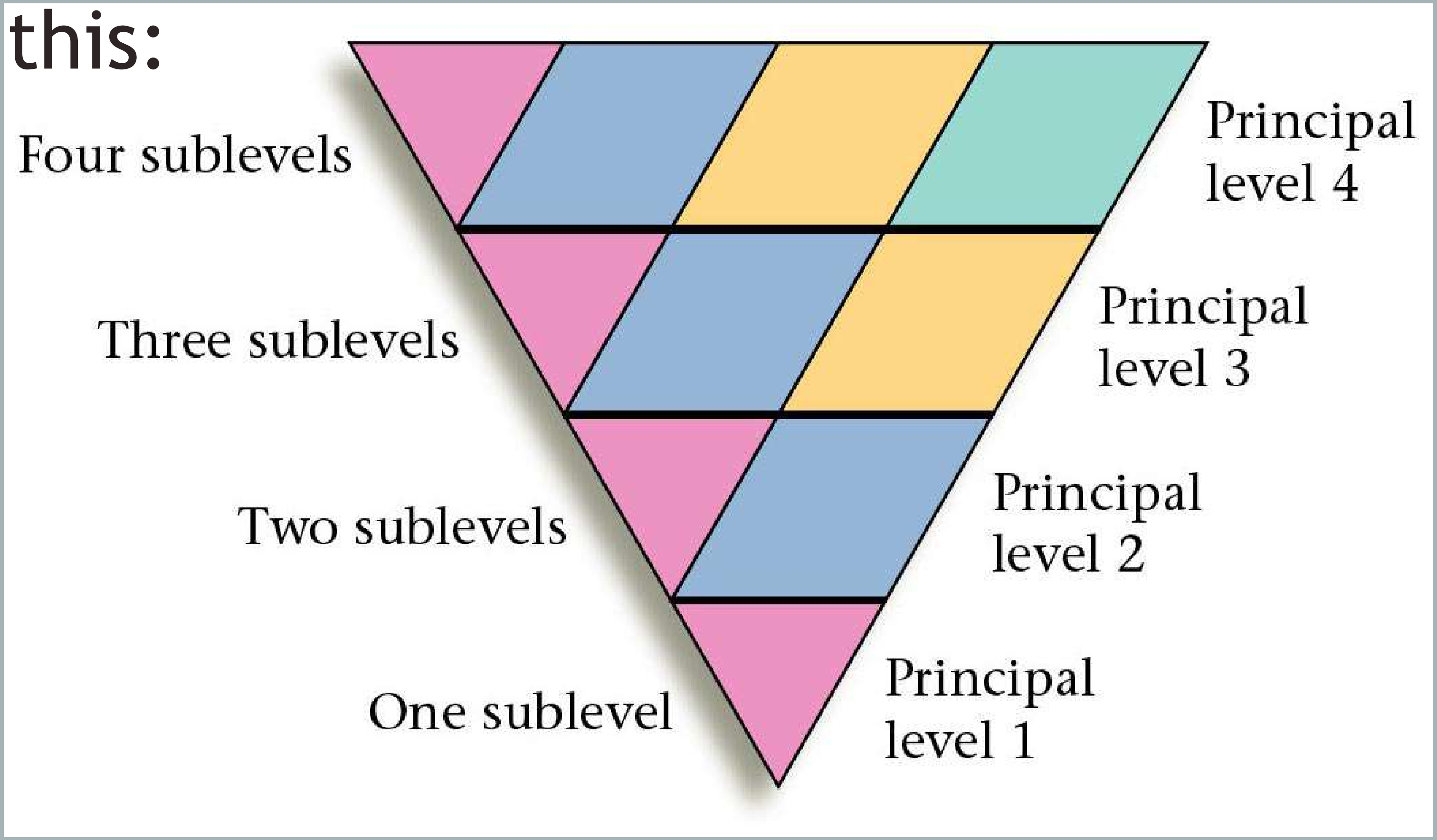
see a pattern?

- It can be pictured like this:



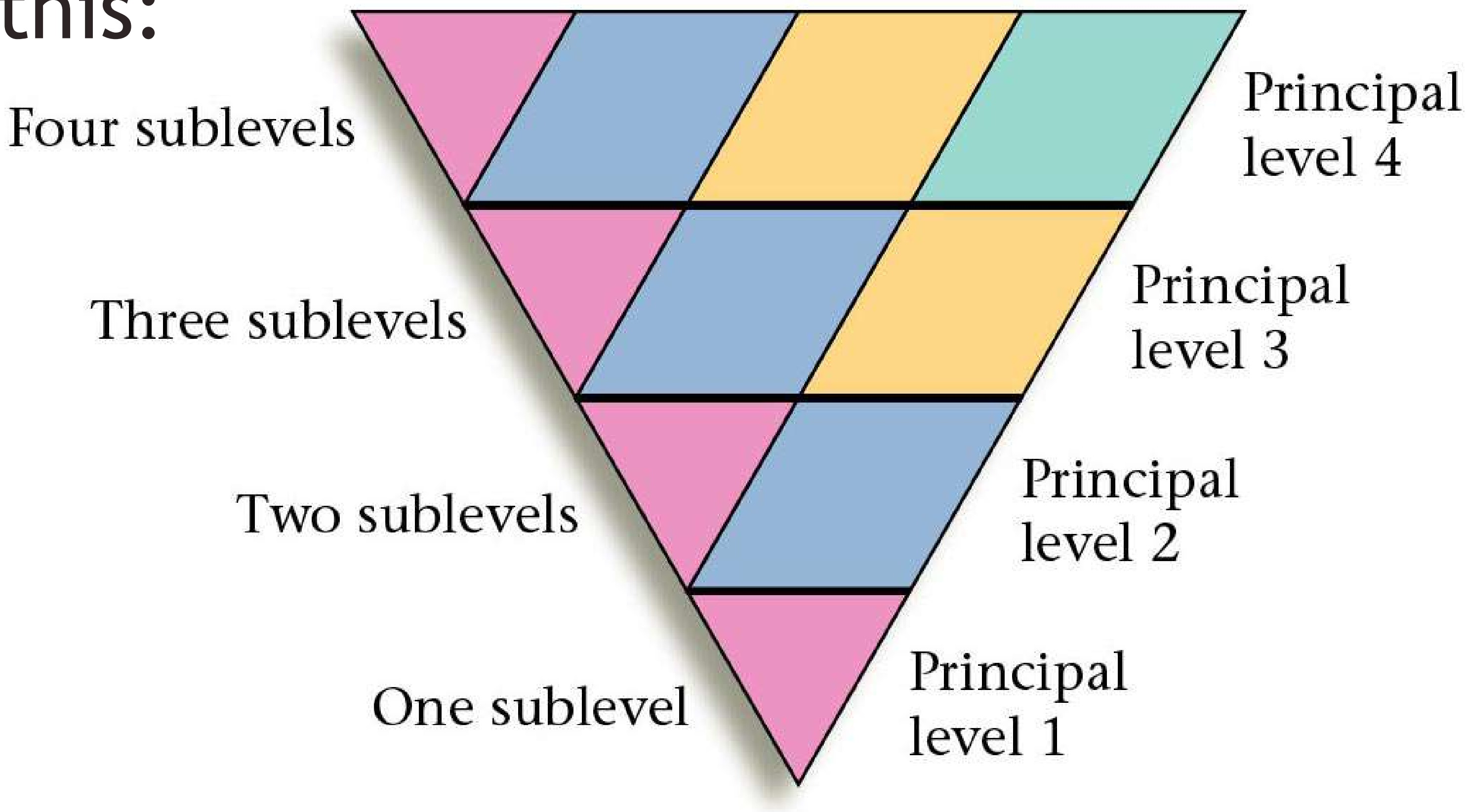
The lowest level (ground state) contains just one orbital; the 1s

- It can be pictured like this:



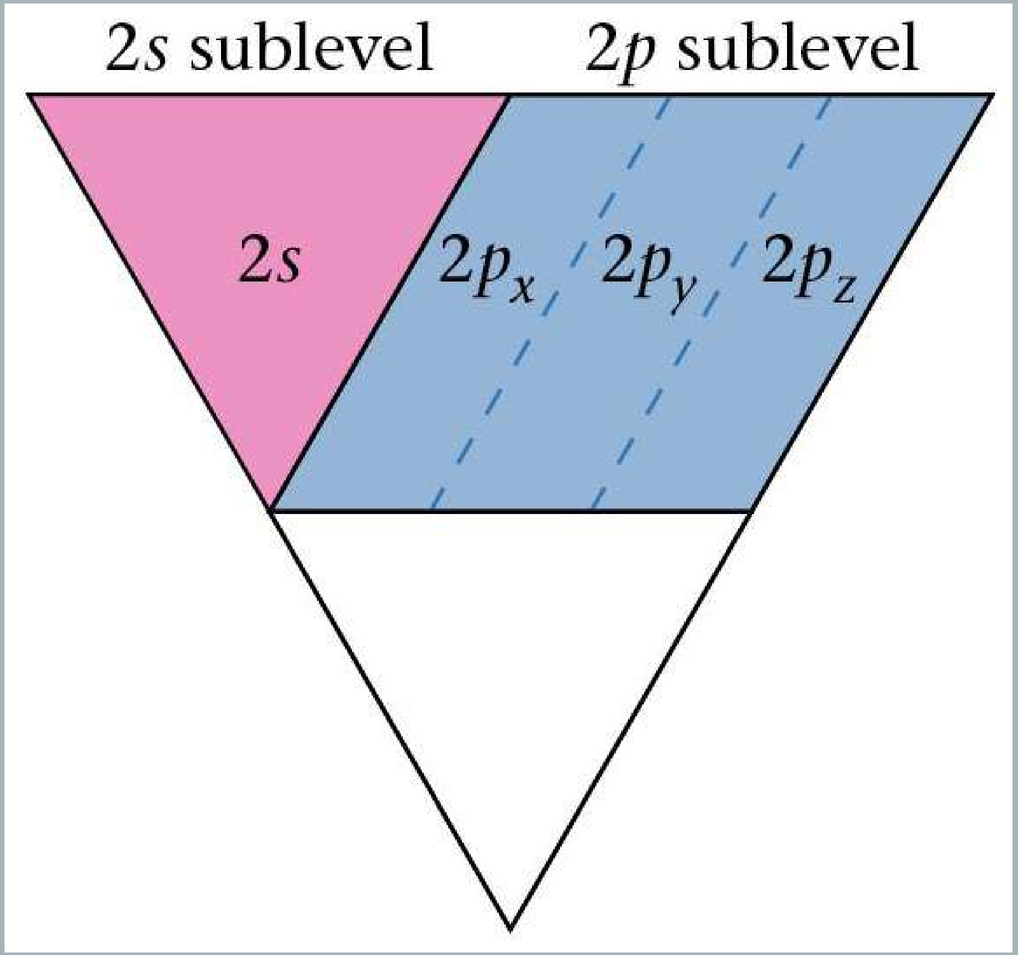
1 stands for the ***1st principal quantum level***

- It can be pictured like this:

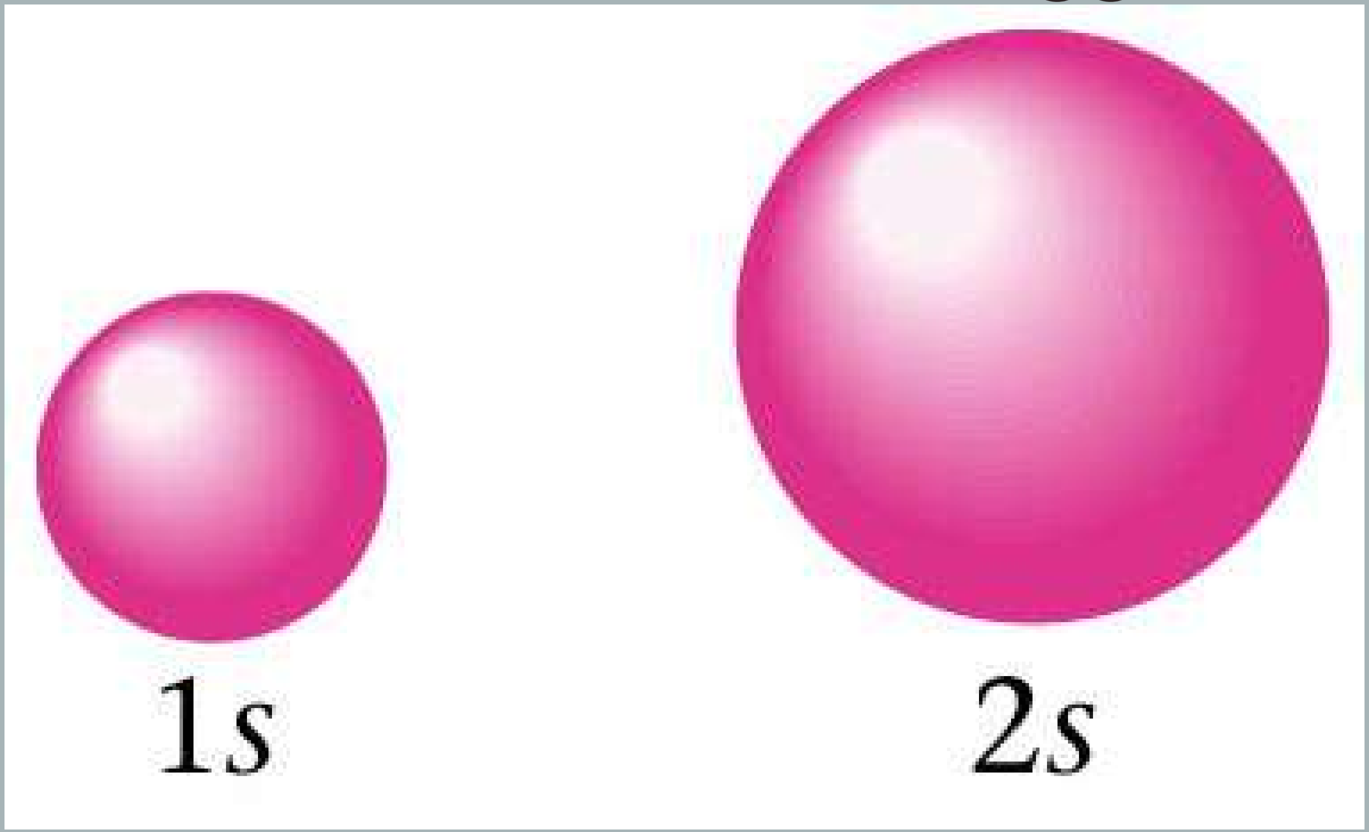


s is the abbreviation for the ***sublevel*** that is there and tells us its ***shape*** (spherical)

The ***second*** level has 2s ***and*** 2p sublevels

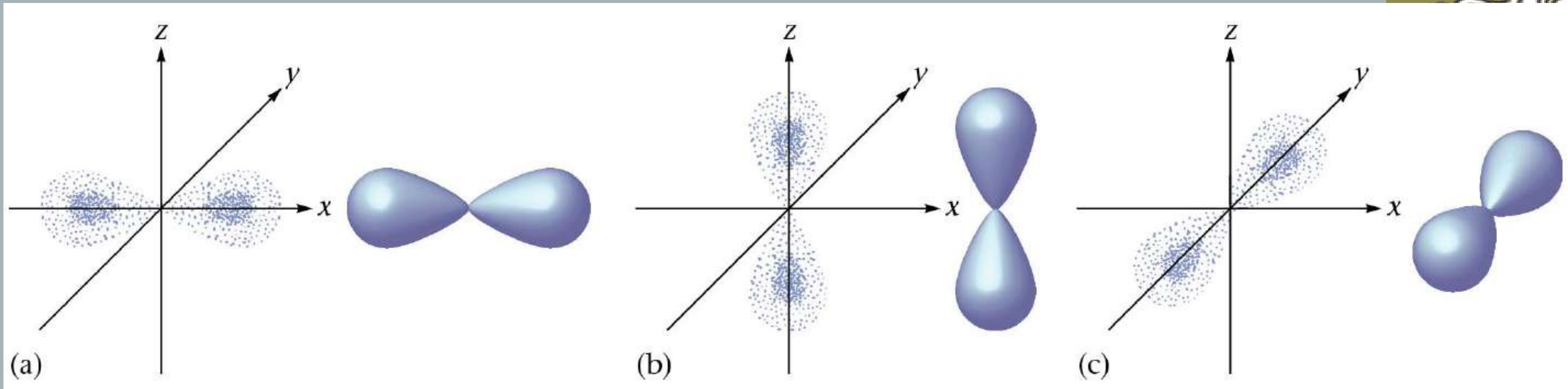


the 2s is just like the 1s, but **bigger**



but the 2p’s are entirely different…

- There are three of them and they are dumbbell shaped (lobed)



- The x, y, and z tells us which axis they are lined up on
- (note: these are **single** orbitals with **double** lobes)

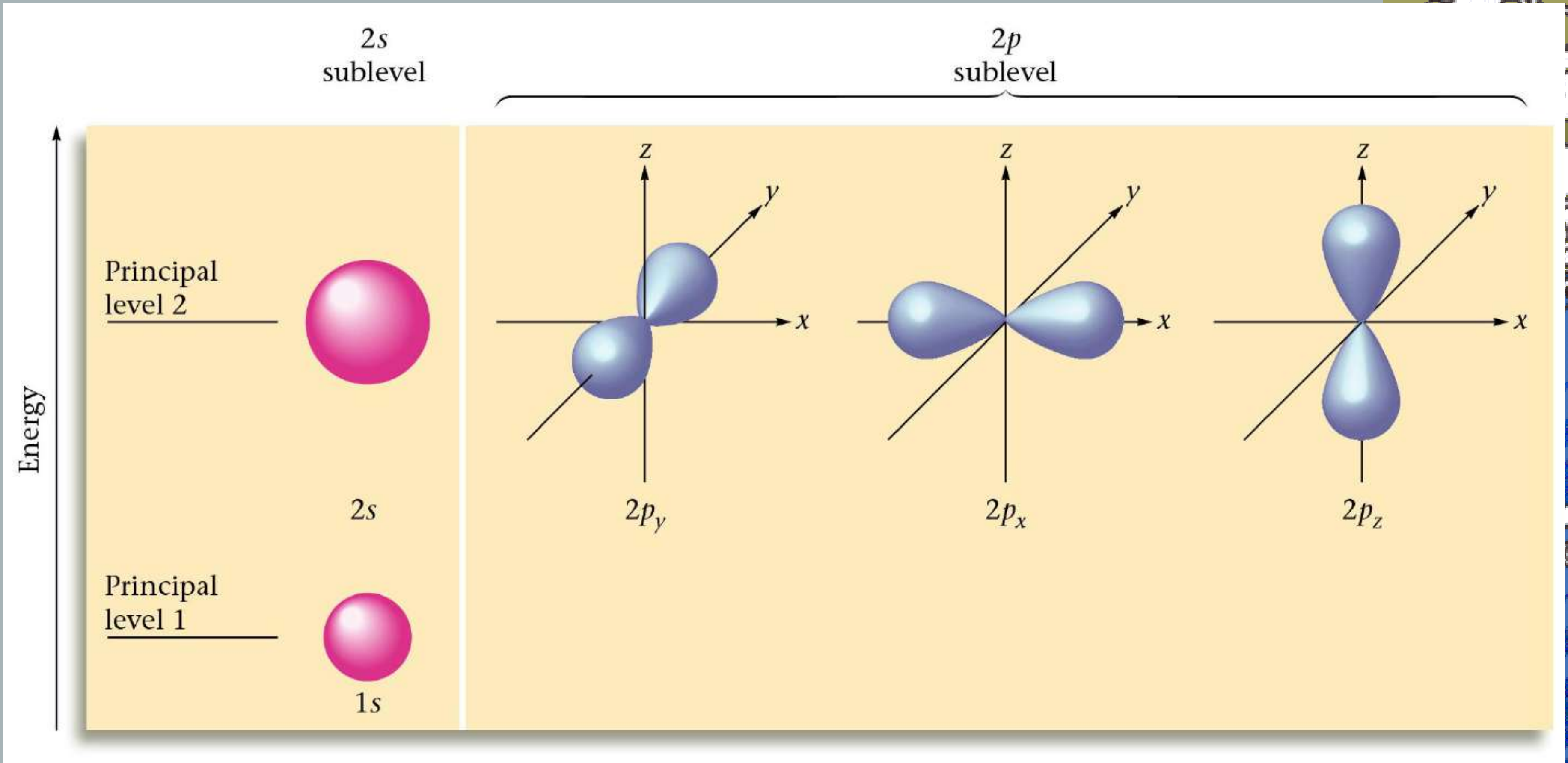
The three 2p Orbitals

The $2p_x$ orbital

The $2p_y$ orbital

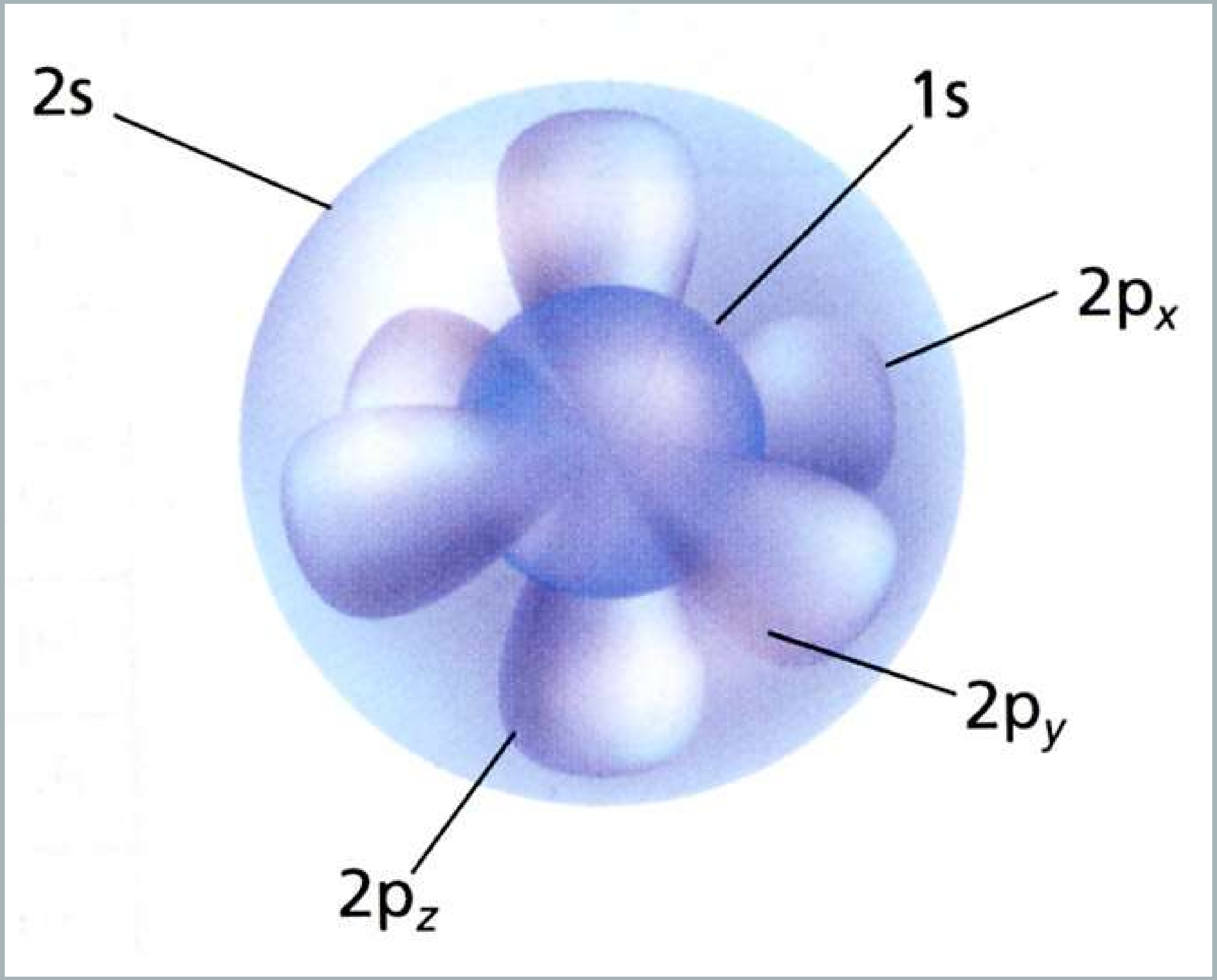
The $2p_z$ orbital

- the story so far…



- Each can hold two electrons

- If overlapped they look like this
  (this view will play a big role beyond H)



- 1S, 2S and sP Orbital animation

- the summary so far:

Orbital Labels

1. The number tells the principal energy level.
2. The letter tells the shape. The letter *s* means a spherical orbital; the letter *p* means a two-lobed orbital. The *x*, *y*, or *z* subscript on a *p* orbital label tells along which of the coordinate axes the two lobes lie.

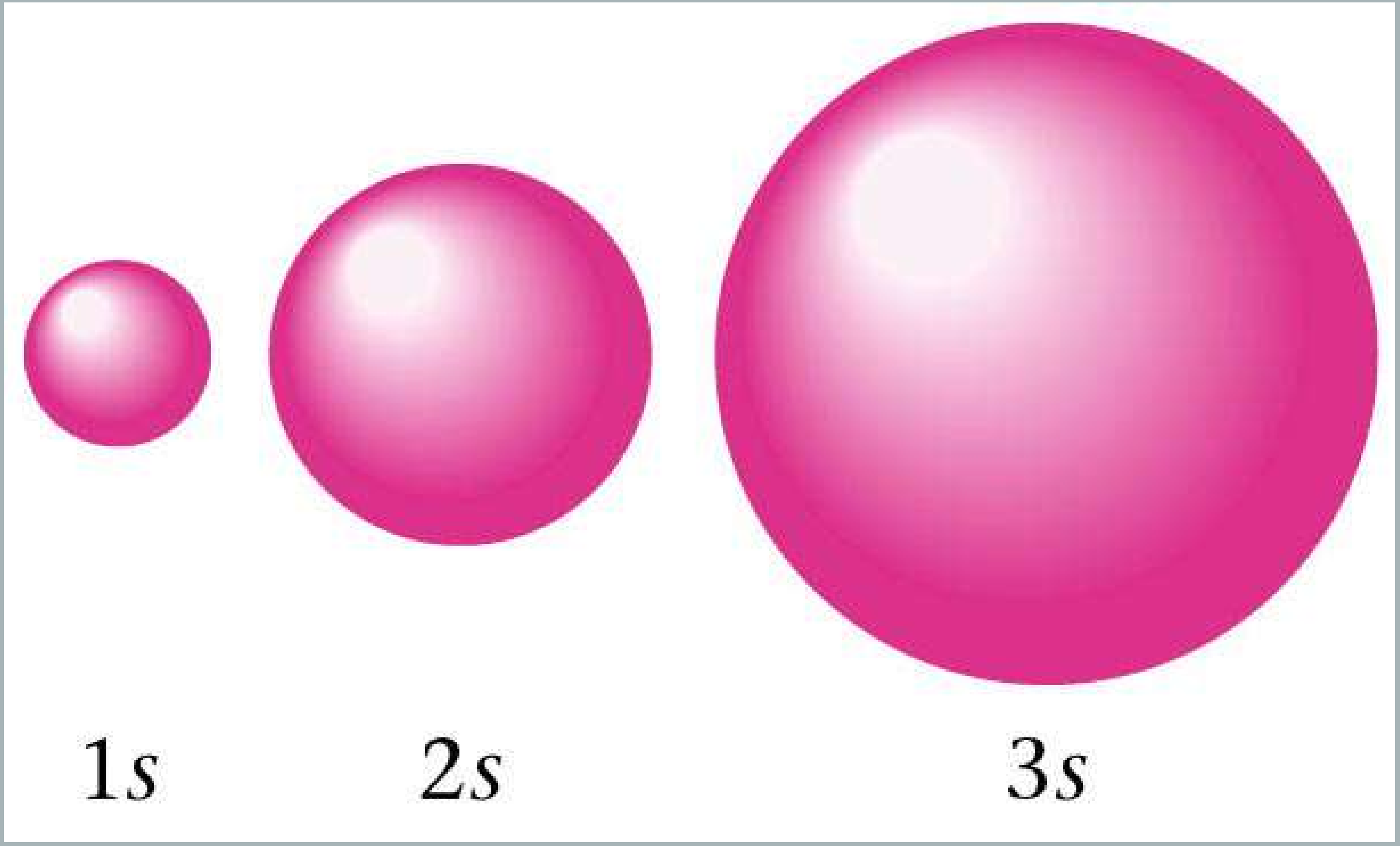
- Important to note that as ***level number goes up so does average distance from nucleus***

- So if H has only one e- why does it have so many orbitals?…

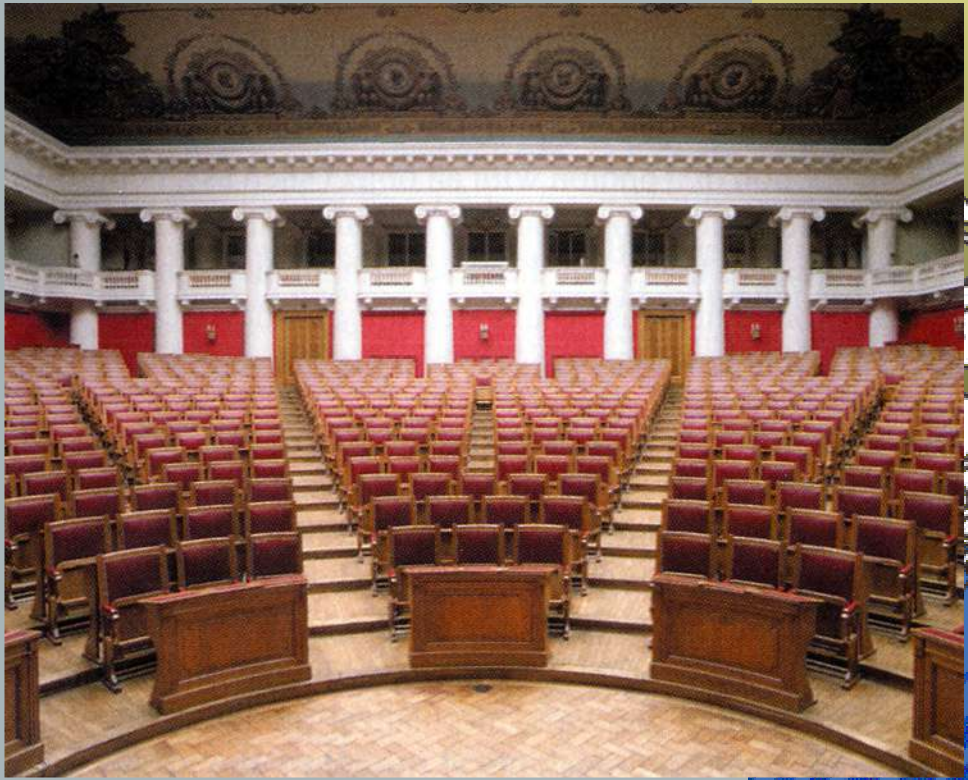
Hydrogen orbitals

- The extra orbitals are just **potential** orbitals for an excited e-
- The e- can only occupy one space at a time!

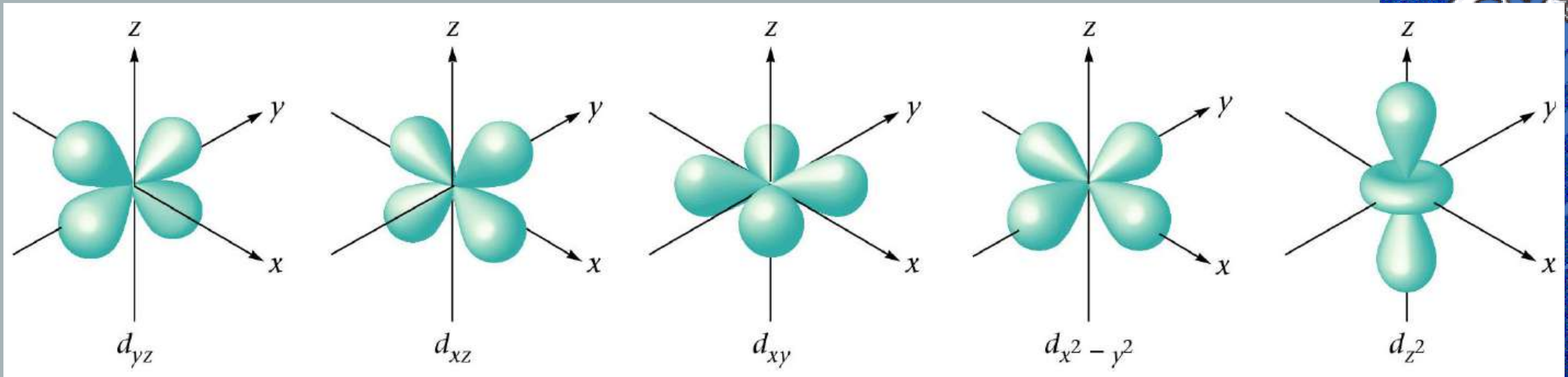
- At **level 3** there are **s and p orbitals** just like the previous levels - ***only bigger***



- There is room for even more orbitals out there



- there are d orbitals!



The shapes and labels of the five 3d orbitals

The $3d_{xz}$ orbital

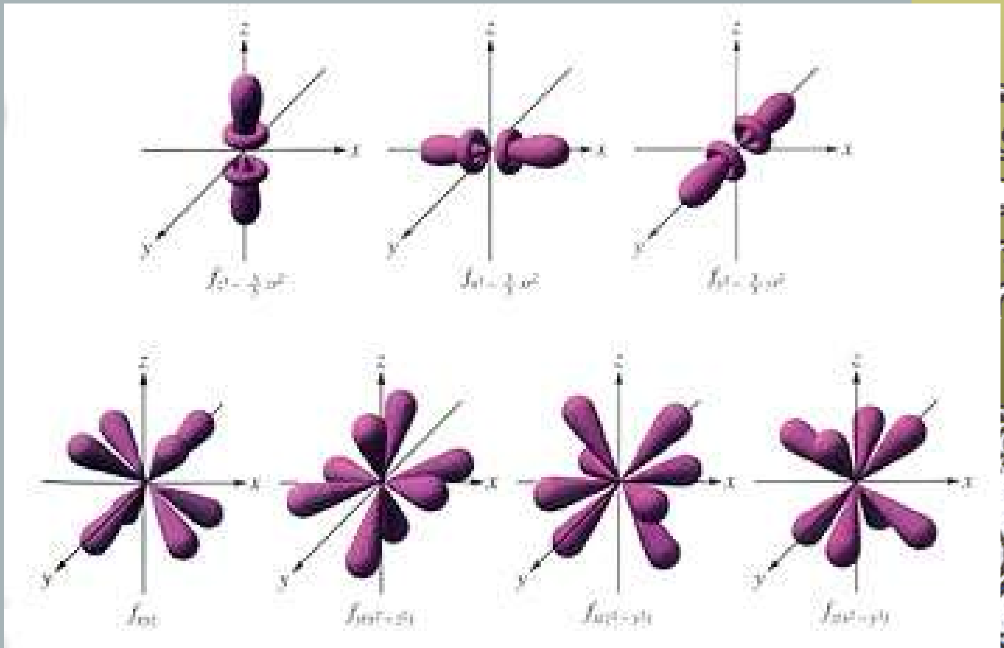
The $3d_{yz}$ orbital

The $3d_{xy}$ orbital

The $3d_{z^2}$ orbital

The $3d_{x^2 z^2}$ orbital

- At the **4th level** there are ***f orbitals*** and so on

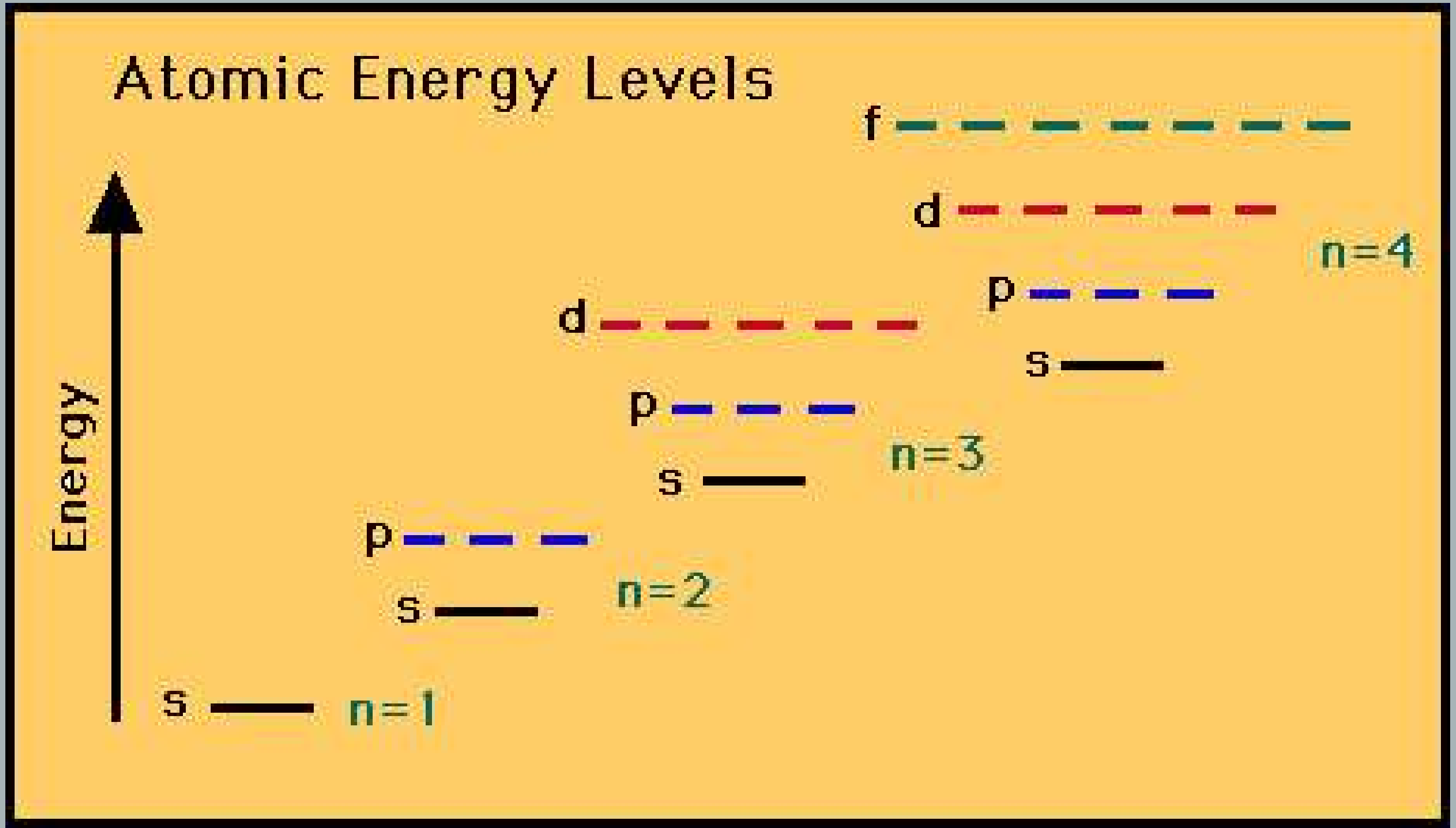


- You are responsible for knowing up to the **4th level** and only the **shapes** of s, p, and d orbitals *(s and p in detail)*

11.9 Electron Arrangements

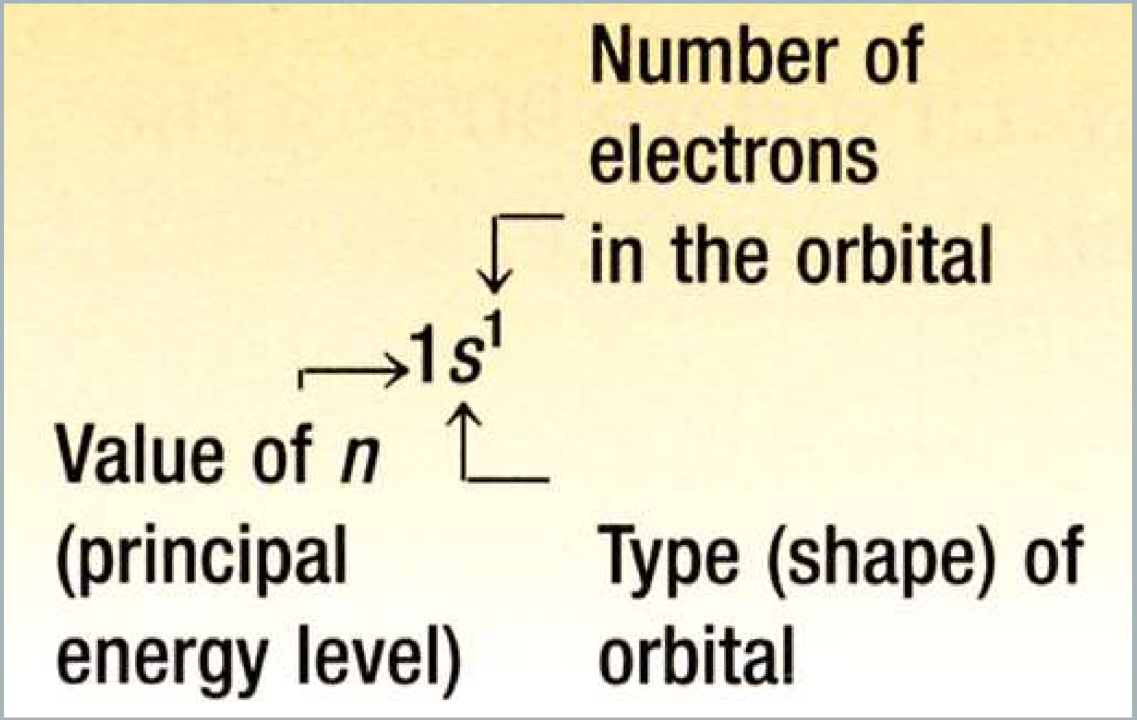
- **Electron configurations** and **electron diagrams**
- Using the Aufbau Principle

  Electrons fill energy levels starting with the lowest



Writing electron configurations

- *Hydrogen*



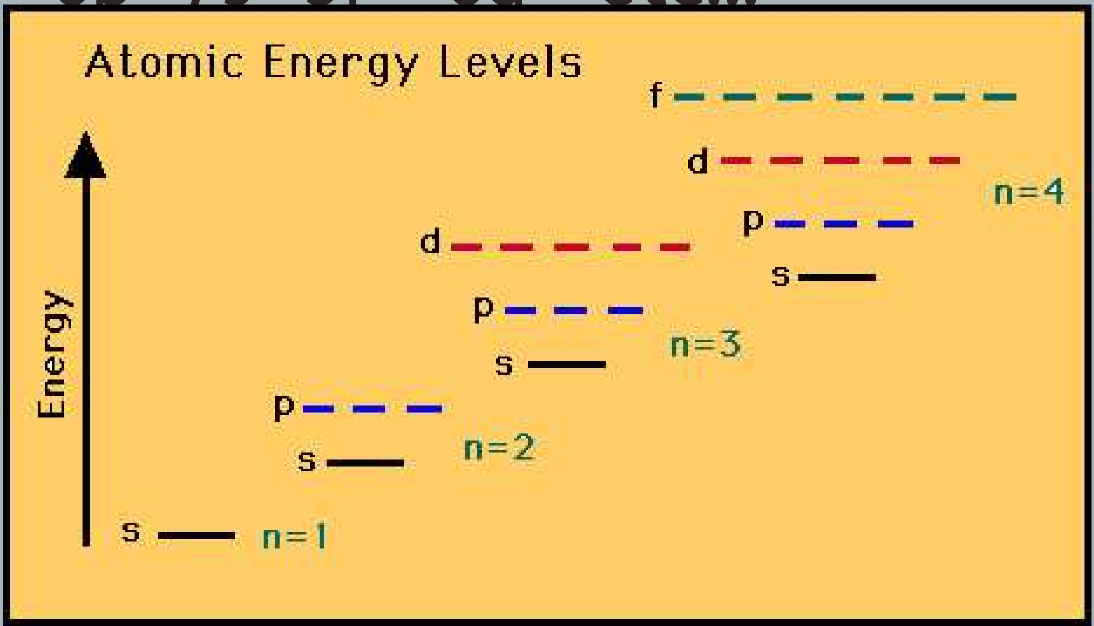
- What would the electron configuration of He be?
- After that, the electrons go into principal energy level 2

Electron configuration

- As larger atoms are configured, the electrons should be assigned to increasing energy levels, according to the Aufbau Principle

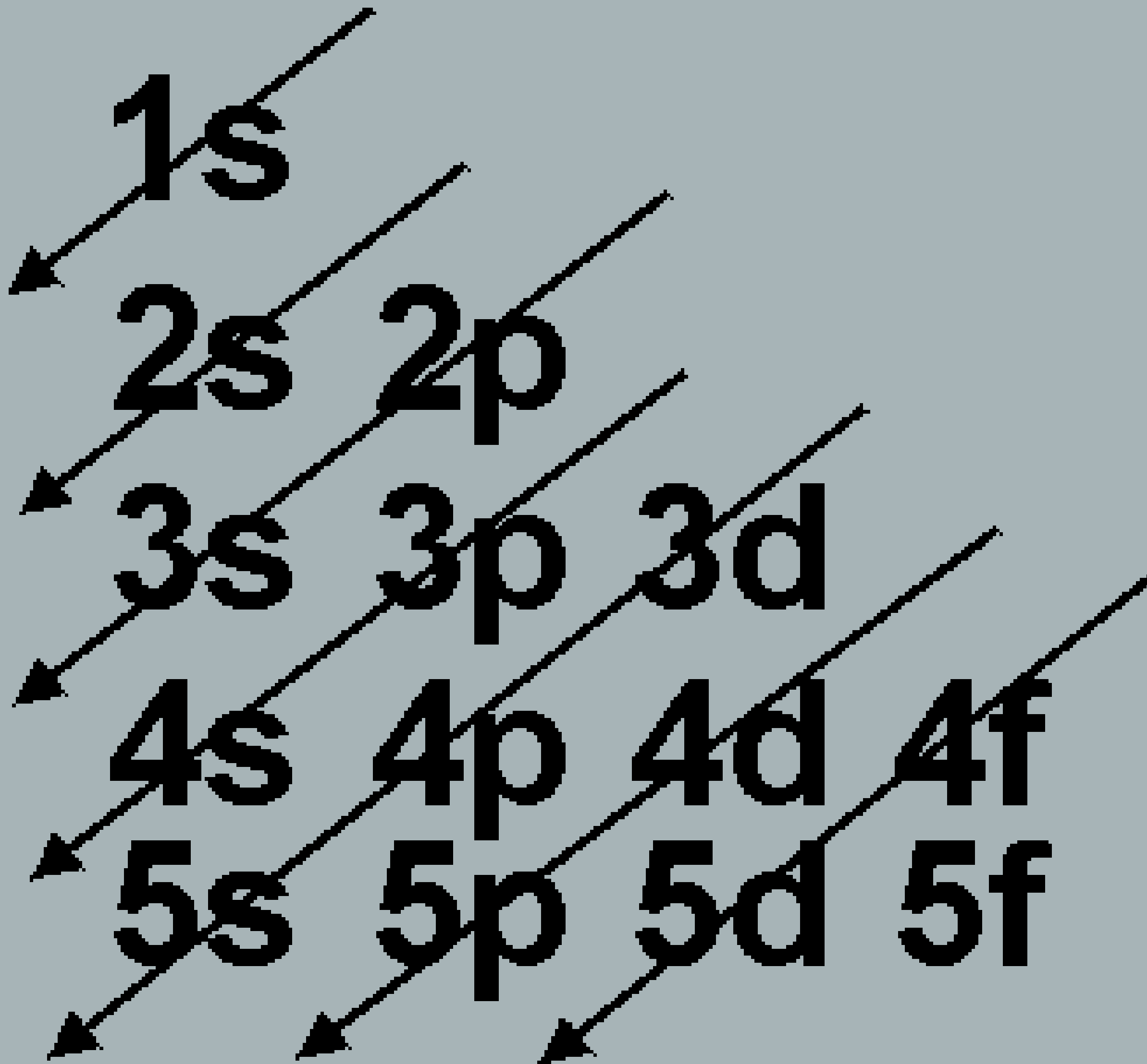
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2$

$4f^{14}\ 5d^{10}\ 6p^6\ 7s^2\ 5f^{14}\ 6d^{10}$ etc...



Aufbau Principle

Graphic



Questions?

Write Electron Configurations for the following

- Boron
- Nitrogen
- Neon
- Sodium
- Phosphorus
- Argon

Write Electron Configurations for the following

- Boron

$1s^2\ 2s^2\ 2p^1$

- Nitrogen

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

- Neon

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

- Sodium

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

- Phosphorus

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

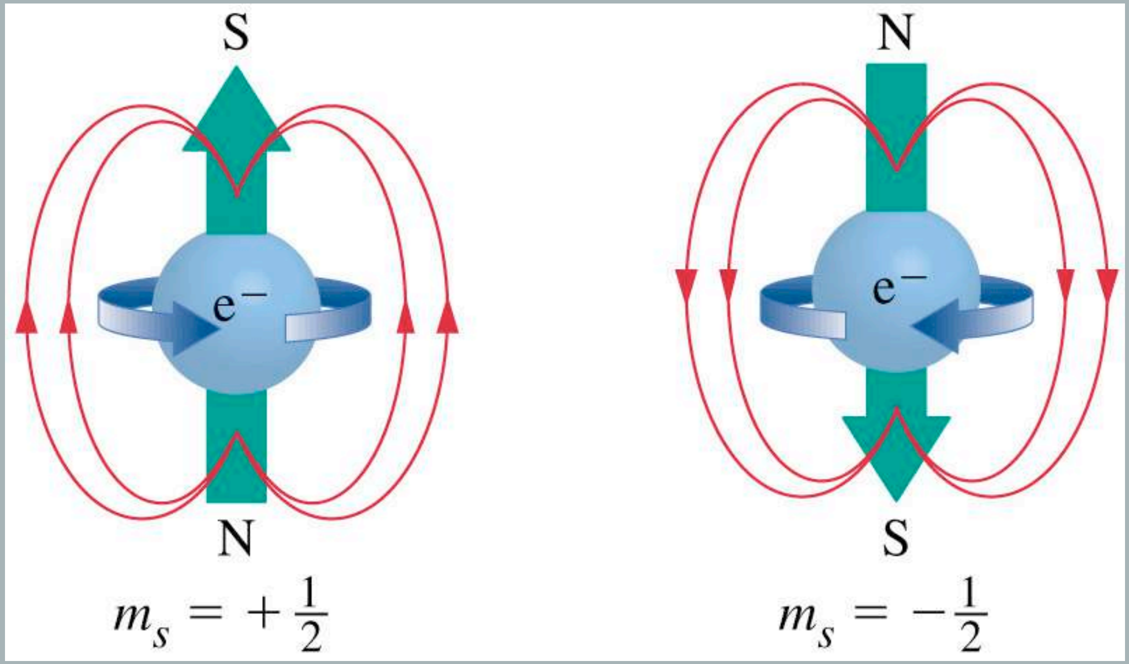
- Argon

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

11.9 Electron Arrangements

- **Electron diagrams**
- Using the Aufbau Principle

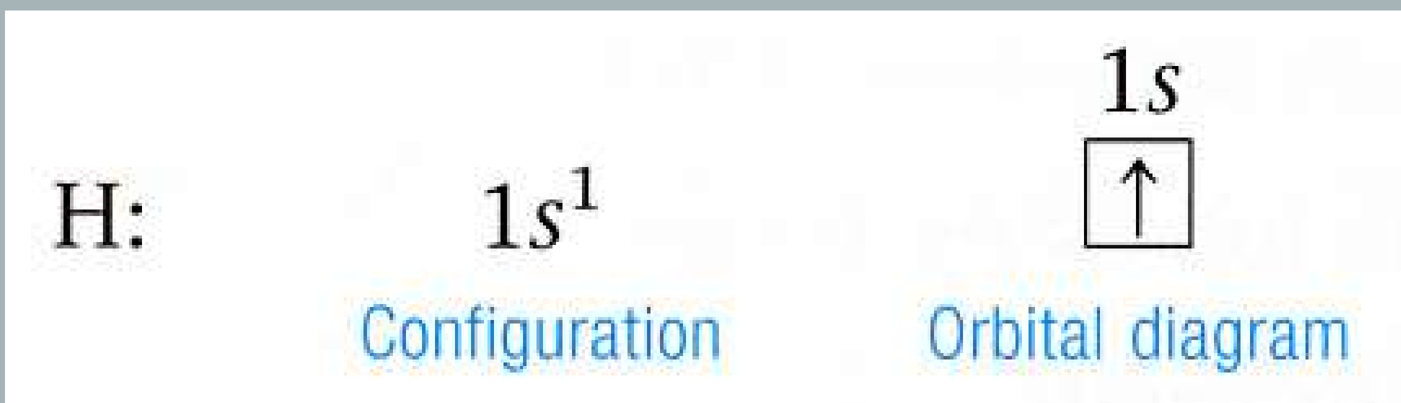
- We have to know one more wave property - **spin**



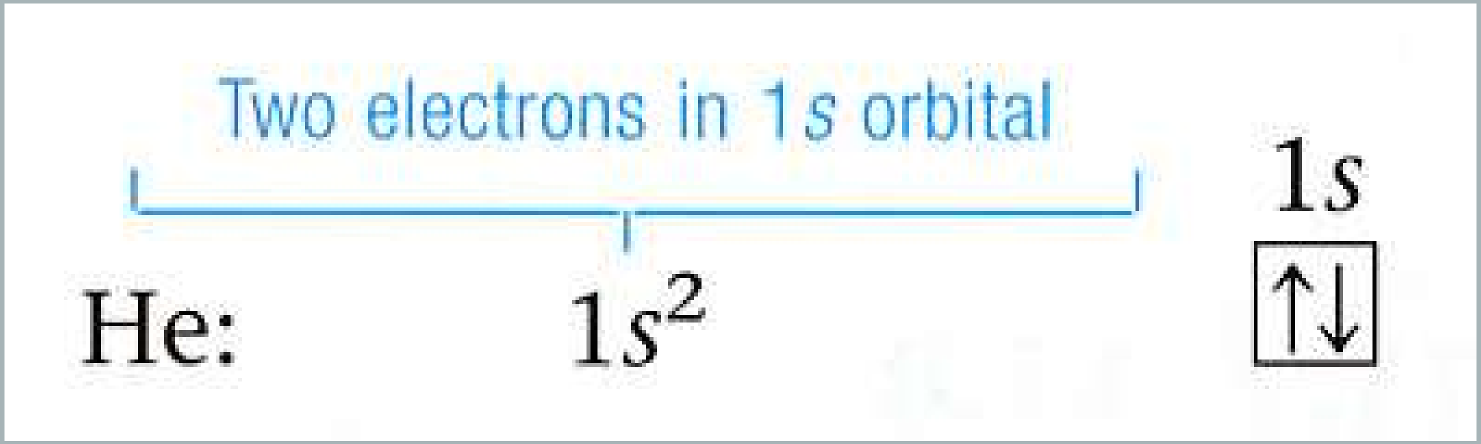
- Little spinning charged things (like e-) have a magnetic field
- Their orientation and proximity must be paired up properly

- for two e- to occupy the same orbital they must spin ***in opposite directions***
- this is the **Pauli exclusion principle**

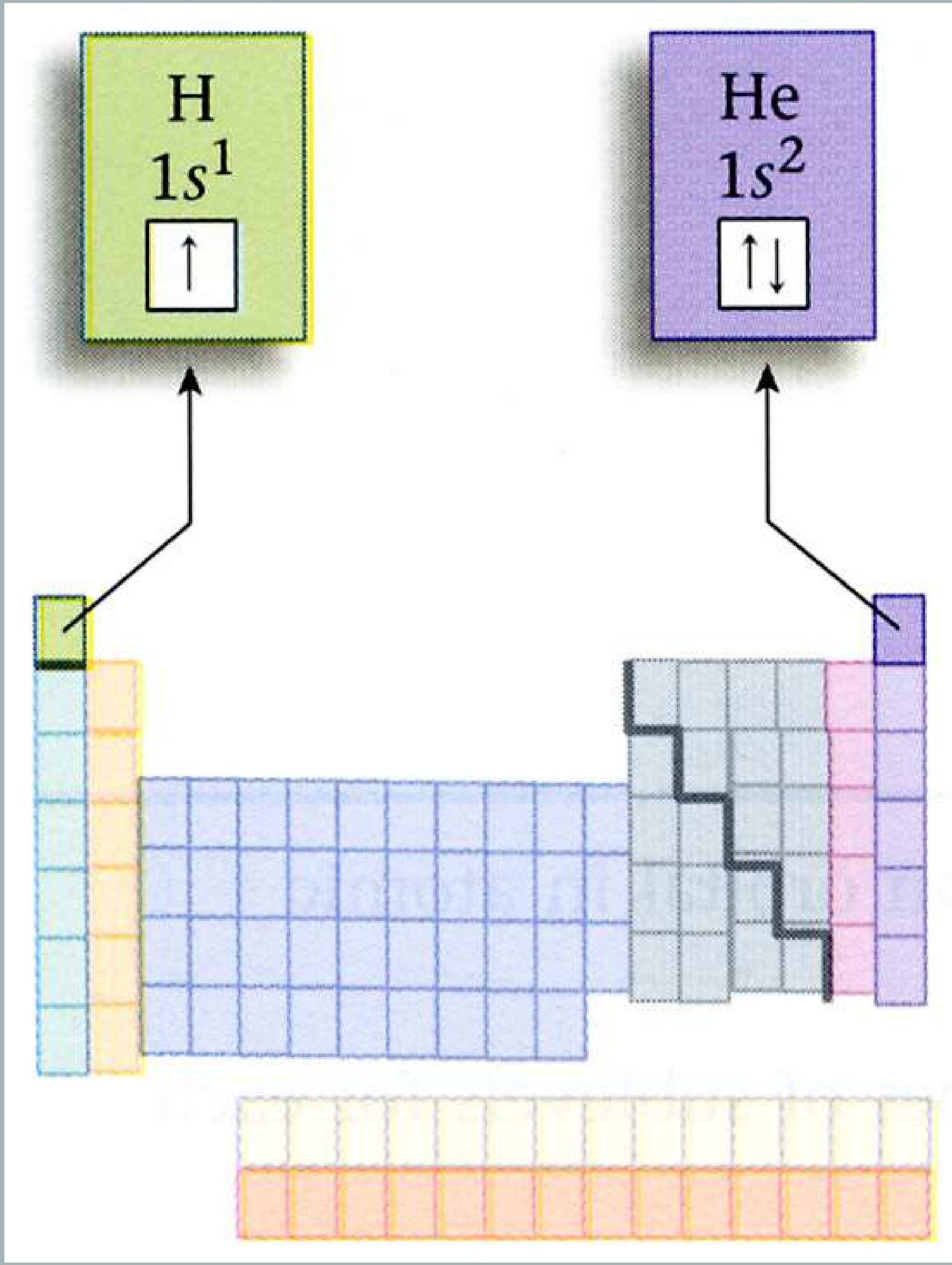
- Atoms can also be represented in an **orbital diagram** (aka **box diagram**) where the arrow represents the electron



- Helium, with Z=2, has a configuration $1s^2$



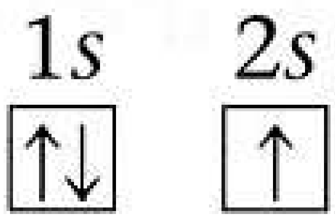
- notice the arrows are opposing, representing opposite spins



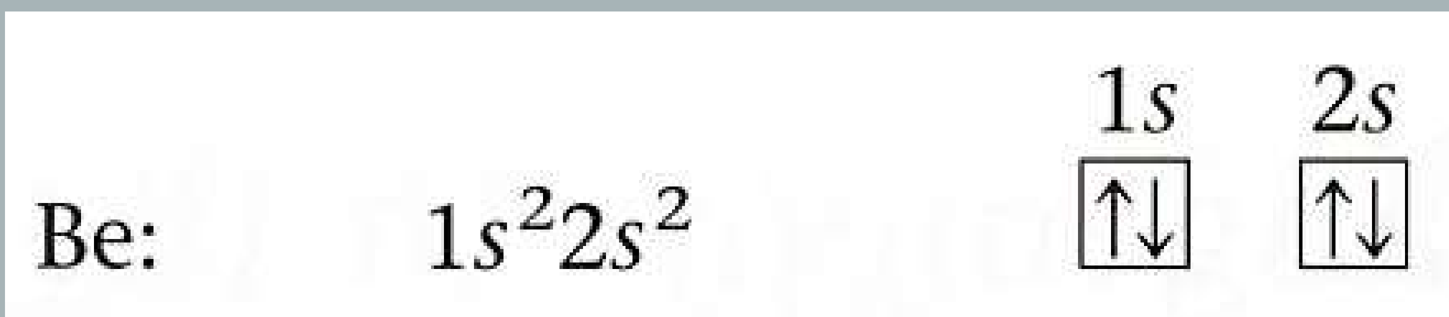
H
1s1
He
1s2

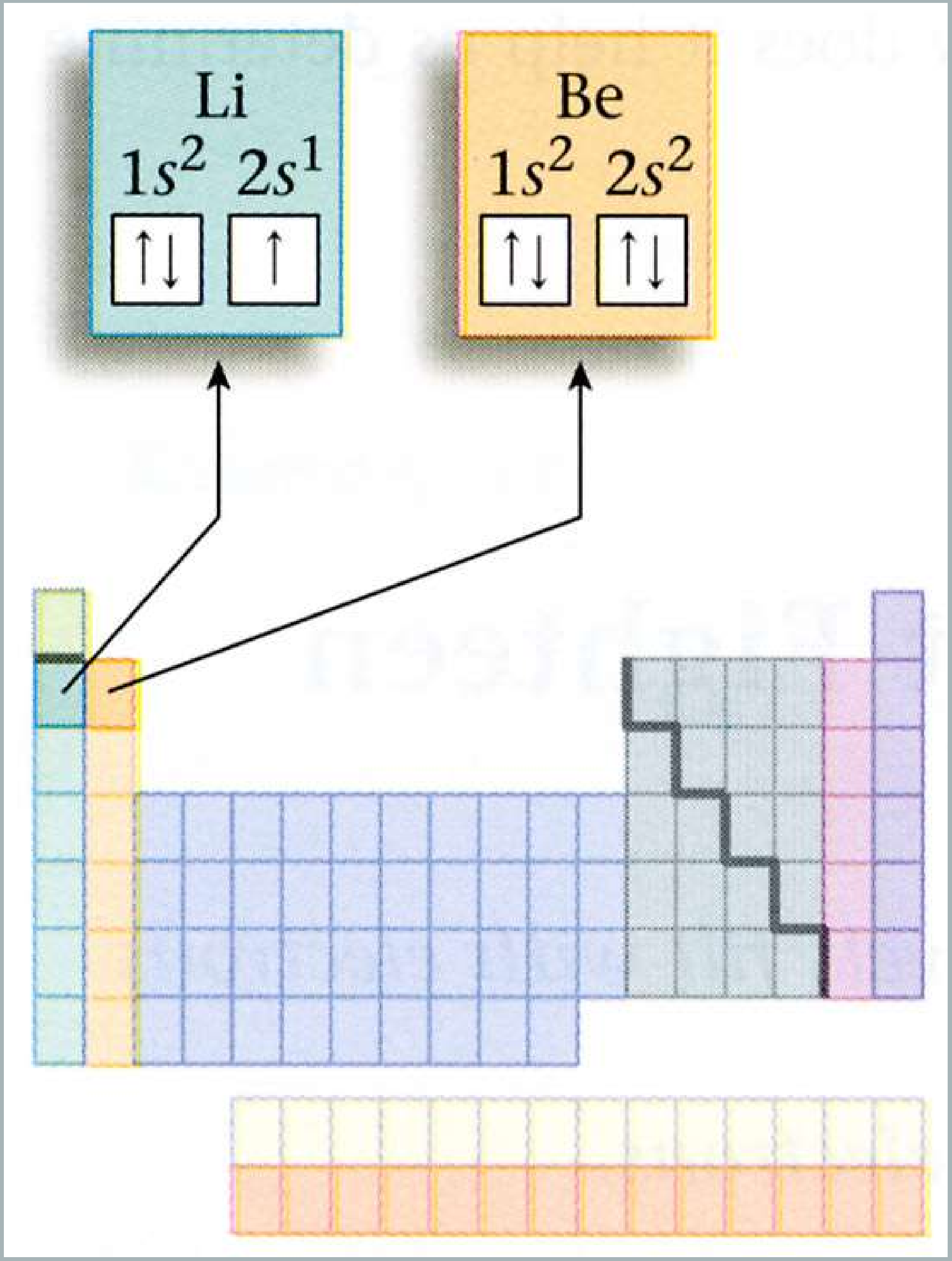
- lithium (Z=3) finds the 1s filled and so throws its third into the next lowest energy orbital
- that is the **2s**!
- when we regard the “polyelectronic” atoms the ***2s orbital is lower in E than the 2p***

Li: $1s^2 2s^1$



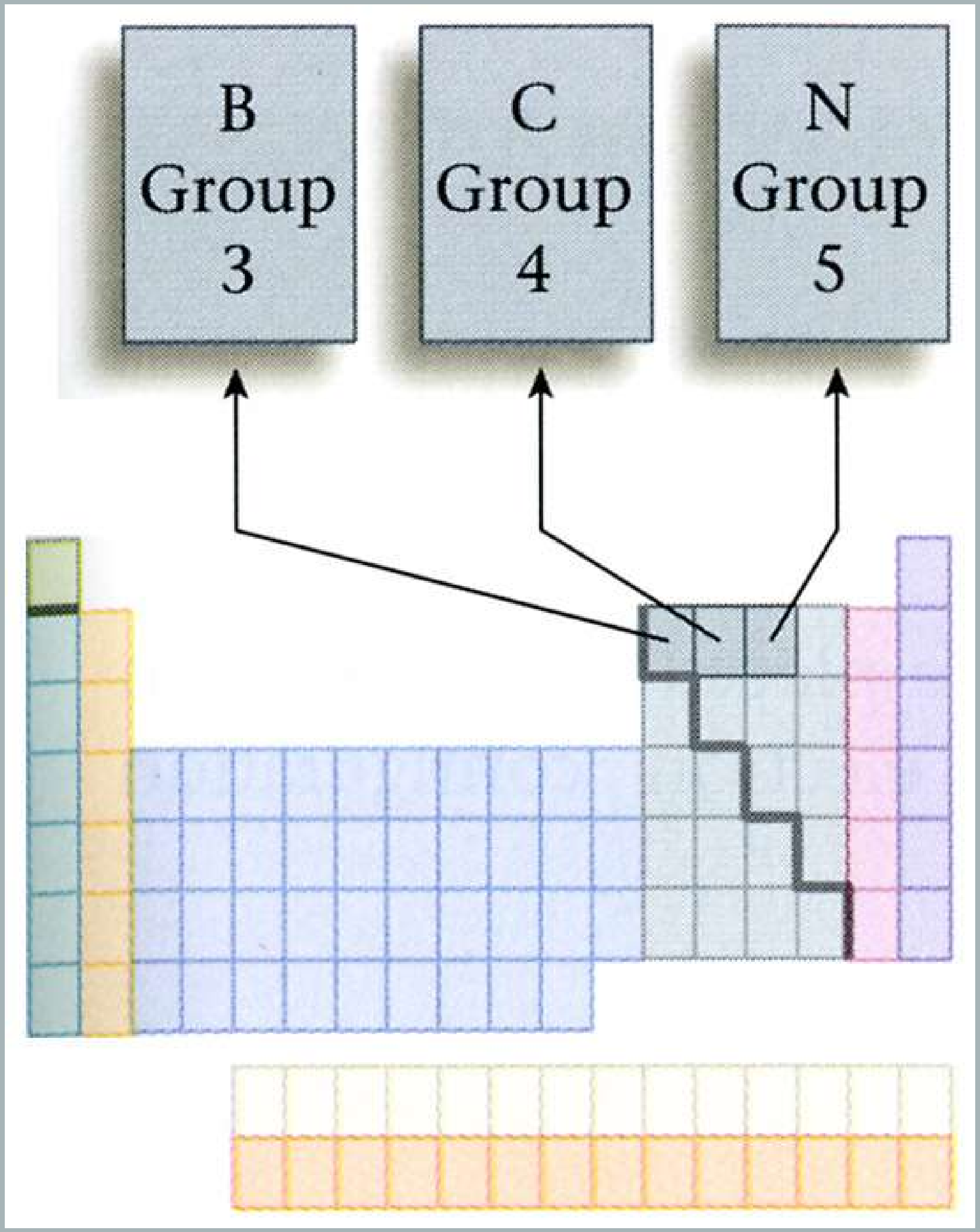
- beryllium’s 4th e- goes into the empty space of the 2s, *but opposite*
- *the 2s must be filled before going to next level*





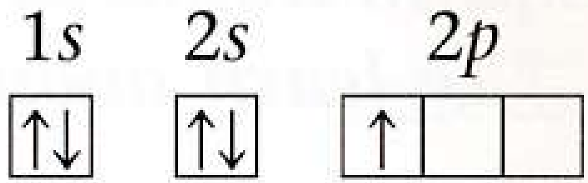
Li
$1s^2$ $2s^1$
Be
$1s^2$ $2s^2$

- now B, C, and N...



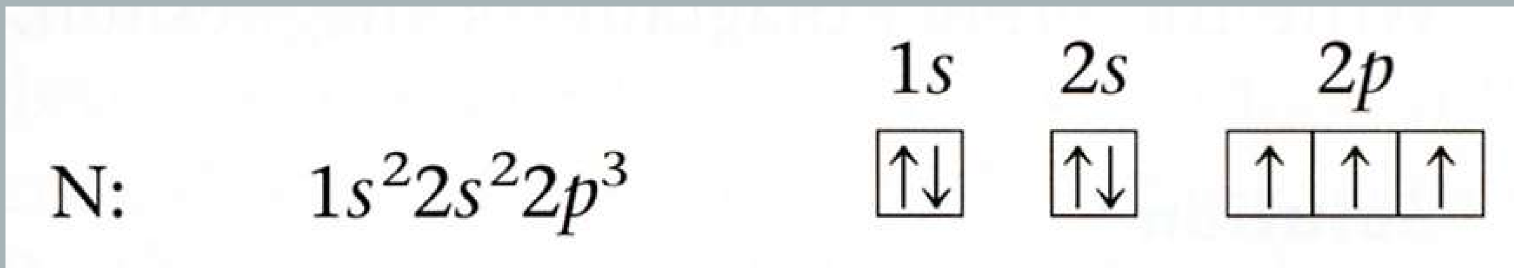
- B's 5th electron goes into any of the 2p's (since they are all the same E, it doesn't matter which)

B: $1s^2 2s^2 2p^1$

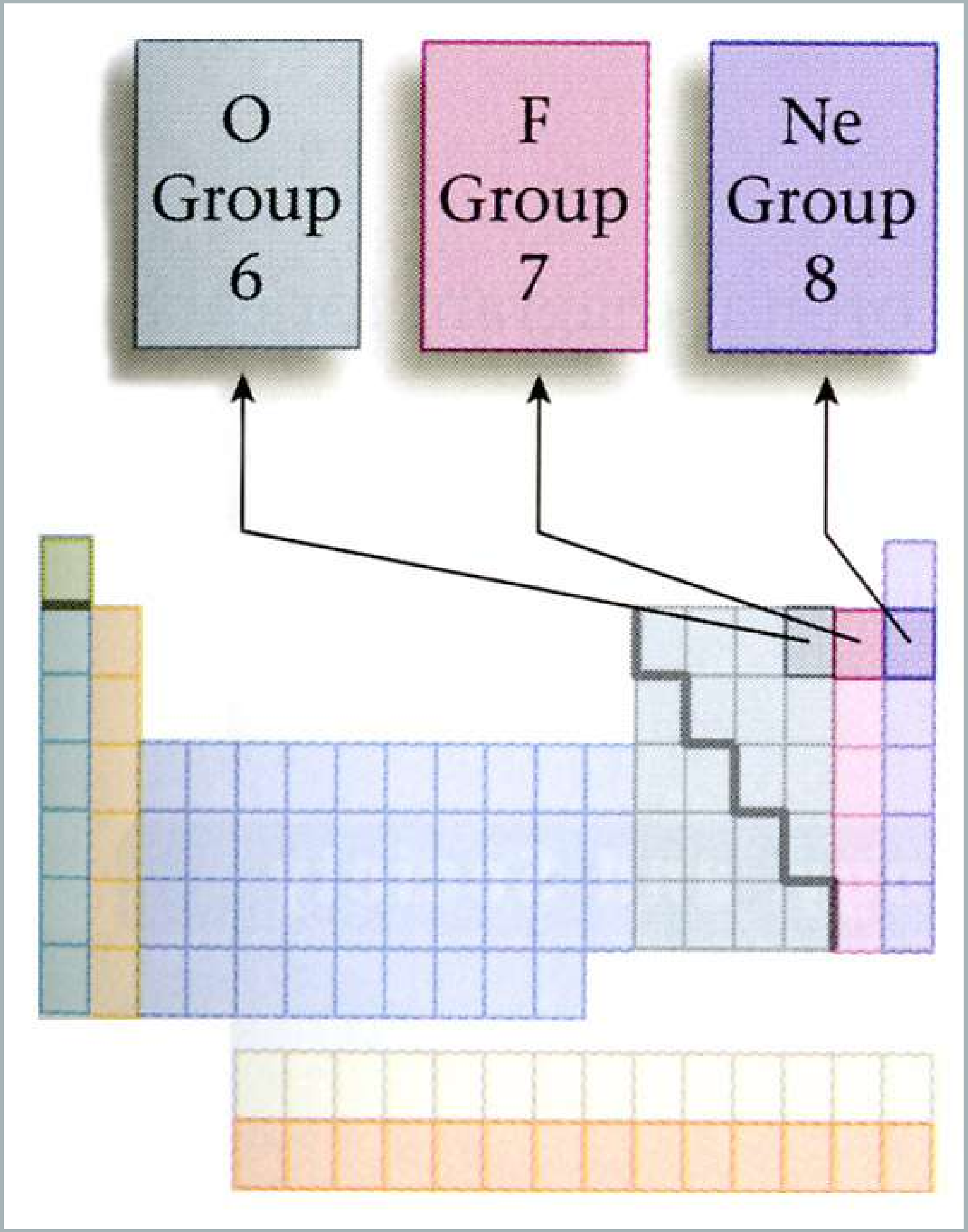


- when C adds its 6th, it has to occupy an empty one (remember, they aren't attracted each other)
- it doesn't matter which of the two, but they must be different from the first!

- (minor point: make sure they spin the same way **until** they are in same orbital)



- nitrogen’s 7th will go in the last unoccupied, keeping them all separate but with same spin


O
Group
6
F
Group
7
Ne
Group
8

- oxygen's 8th can buddy up with any of the others, ***but it must spin opposite!***

		1s	2s	2p		
O:	$1s^2 2s^2 2p^4$	↑↓	↑↓	↑↓	↑	↑

- same with F and Ne

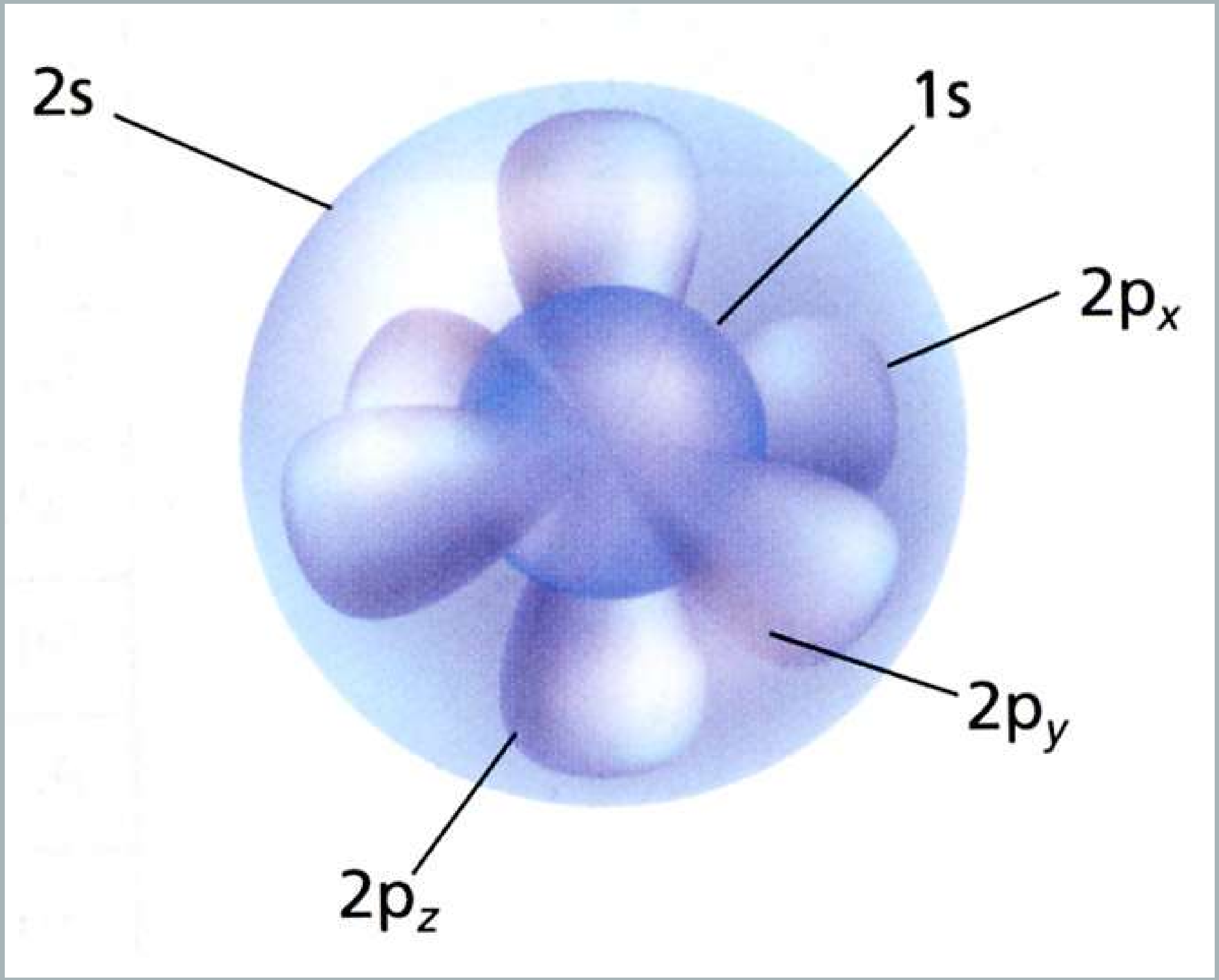
		1s	2s	2p		
F:	$1s^2 2s^2 2p^5$	↑↓	↑↓	↑↓	↑↓	↑
Ne:	$1s^2 2s^2 2p^6$	↑↓	↑↓	↑↓	↑↓	↑↓

- the story so far:

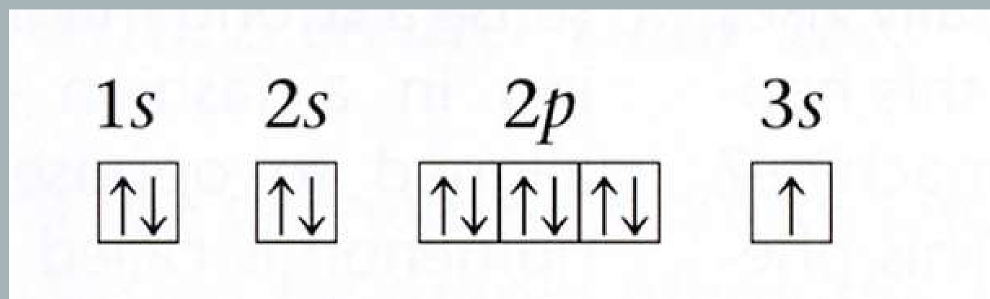
Electron Configurations and Orbital Diagrams for Elements in the First Two Periods

Element	Atomic number	Orbital diagram 1s	2s	$2p_x$	$2p_y$	$2p_z$	Electron configuration notation
Hydrogen	1	↑					$1s^1$
Helium	2	↑↓					$1s^2$
Lithium	3	↑↓	↑				$1s^2 2s^1$
Beryllium	4	↑↓	↑↓				$1s^2 2s^2$
Boron	5	↑↓	↑↓	↑			$1s^2 2s^2 2p^1$
Carbon	6	↑↓	↑↓	↑	↑		$1s^2 2s^2 2p^2$
Nitrogen	7	↑↓	↑↓	↑	↑	↑	$1s^2 2s^2 2p^3$
Oxygen	8	↑↓	↑↓	↑↓	↑	↑	$1s^2 2s^2 2p^4$
Fluorine	9	↑↓	↑↓	↑↓	↑↓	↑	$1s^2 2s^2 2p^5$
Neon	10	↑↓	↑↓	↑↓	↑↓	↑↓	$1s^2 2s^2 2p^6$

- these are the orbitals, all put together as before - just now they are all filled



- next comes Na!
- we've filled up the first and second level; now what?!
- go to the next level! the 3s



- but now a shortcut for the e-configuration
- instead of writing $1s^2 2s^2 2p^6 3s^1$ we abbreviate the inner-level electrons with [Ne], so Na becomes **[Ne]3s**$^{\mathbf{1}}$

For the following elements, write the electron configuration, shortcut and the electron diagram.

- oxygen
- boron
- chlorine
- iron
- aluminum
- nitrogen
- sulfur
- silicon
- krypton
- titanium
- magnesium
- phosphorus
- neon
- copper
- barium

- Mg (Z=12) is $1s^2 2s^2 2p^6 3s^2$ or $[Ne]3s^2$
- Al (Z=13) is $1s^2 2s^2 2p^6 3s^2 3p^1$ or $[Ne]3s^2 3p^1$
- and so on…

Electron Configurations for Elements in Period Three

Element	Atomic number	Complete electron configuration	Electron configuration using noble-gas notation
Sodium	11	$1s^2 2s^2 2p^6 3s^1$	$[Ne]3s^1$
Magnesium	12	$1s^2 2s^2 2p^6 3s^2$	$[Ne]3s^2$
Aluminum	13	$1s^2 2s^2 2p^6 3s^2 3p^1$	$[Ne]3s^2 3p^1$
Silicon	14	$1s^2 2s^2 2p^6 3s^2 3p^2$	$[Ne]3s^2 3p^2$
Phosphorus	15	$1s^2 2s^2 2p^6 3s^2 3p^3$	$[Ne]3s^2 3p^3$
Sulfur	16	$1s^2 2s^2 2p^6 3s^2 3p^4$	$[Ne]3s^2 3p^4$
Chlorine	17	$1s^2 2s^2 2p^6 3s^2 3p^5$	$[Ne]3s^2 3p^5$
Argon	18	$1s^2 2s^2 2p^6 3s^2 3p^6$	$[Ne]3s^2 3p^6$ or [Ar]

- **valence electrons** are those on the outermost principal energy level; the **core electrons** are the ones inside the outermost (yellow)

Electron Configurations for Elements in Period Three

Element	Atomic number	Complete electron configuration	Electron configuration using noble-gas notation
Sodium	11	$1s^2 2s^2 2p^6 3s^1$	$[Ne]3s^1$
Magnesium	12	$1s^2 2s^2 2p^6 3s^2$	$[Ne]3s^2$
Aluminum	13	$1s^2 2s^2 2p^6 3s^2 3p^1$	$[Ne]3s^2 3p^1$
Silicon	14	$1s^2 2s^2 2p^6 3s^2 3p^2$	$[Ne]3s^2 3p^2$
Phosphorus	15	$1s^2 2s^2 2p^6 3s^2 3p^3$	$[Ne]3s^2 3p^3$
Sulfur	16	$1s^2 2s^2 2p^6 3s^2 3p^4$	$[Ne]3s^2 3p^4$
Chlorine	17	$1s^2 2s^2 2p^6 3s^2 3p^5$	$[Ne]3s^2 3p^5$
Argon	18	$1s^2 2s^2 2p^6 3s^2 3p^6$	$[Ne]3s^2 3p^6$ or $[Ar]$

- here’s a summary of the outer orbitals; see a pattern?

H $1s^1$							He $1s^2$
Li $2s^1$	Be $2s^2$	B $2p^1$	C $2p^2$	N $2p^3$	O $2p^4$	F $2p^5$	Ne $2p^6$
Na $3s^1$	Mg $3s^2$	Al $3p^1$	Si $3p^2$	P $3p^3$	S $3p^4$	Cl $3p^5$	Ar $3p^6$

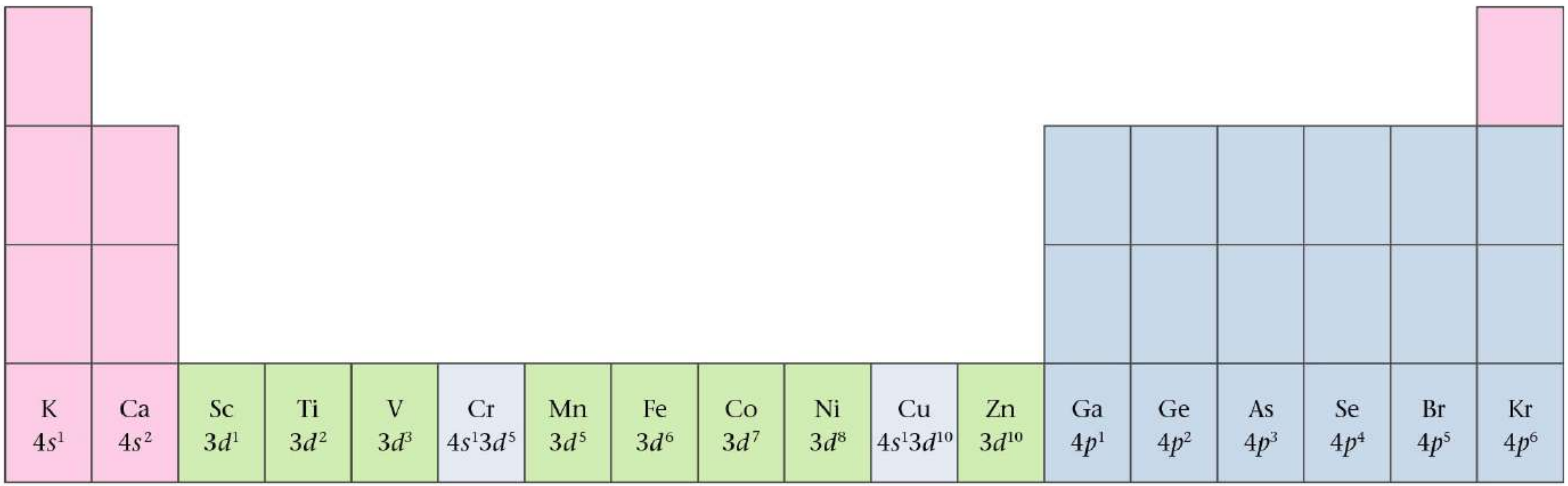
- next we tackle the other elements…

11.10 Electron Configurations & the Periodic Table

- the first 18 seem OK; they fill normally
- but when we get to K, things start getting a little weirder
- K's 19th electron should go to **3d**, right?
- but K behaves a lot like **Li and Na**, both of which have an s^1 as their last e-
- **what to do???!**

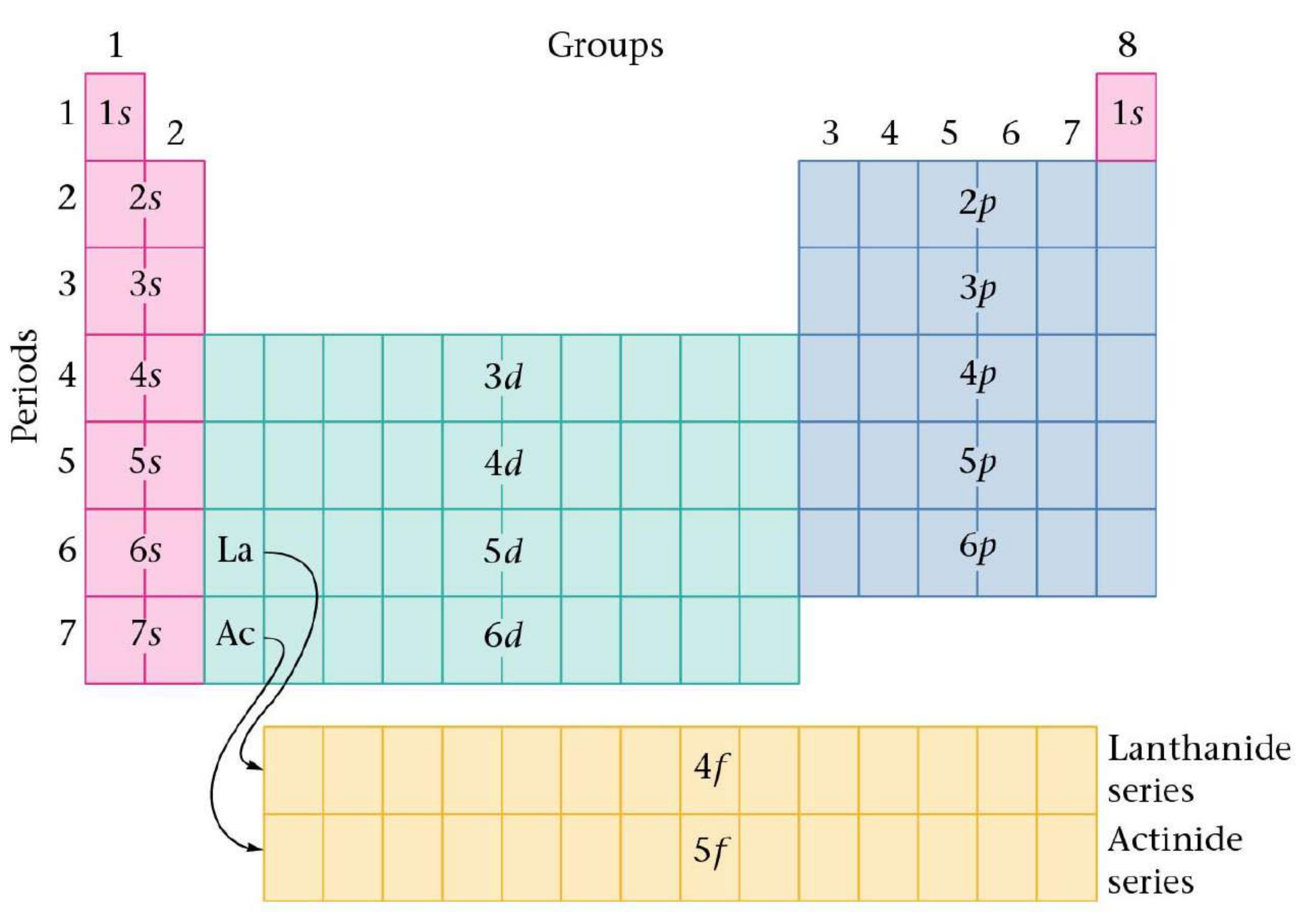
- we follow the lead **nature** gives us and say: K’s 19th will go in the **4s**, ***not the 3d***, so…
- K is $\mathbf{1s^2 2s^2 2p^6 3s^2 3p^6 4s^1}$ or $\mathbf{[Ar]4s^1}$
- *experiments show this is right*
- and Ca will be: $\mathbf{1s^2 2s^2 2p^6 3s^2 3p^6 4s^2}$ or $\mathbf{[Ar]4s^2}$
- ***now*** the 4s is full!!!
- so, b/c *where the electrons go is where there is lower E*, we conclude 4s **must** be lower in E than 3d
- then where next?

- **now** it appears the **3d's get filled**, **then** the 4p's like this

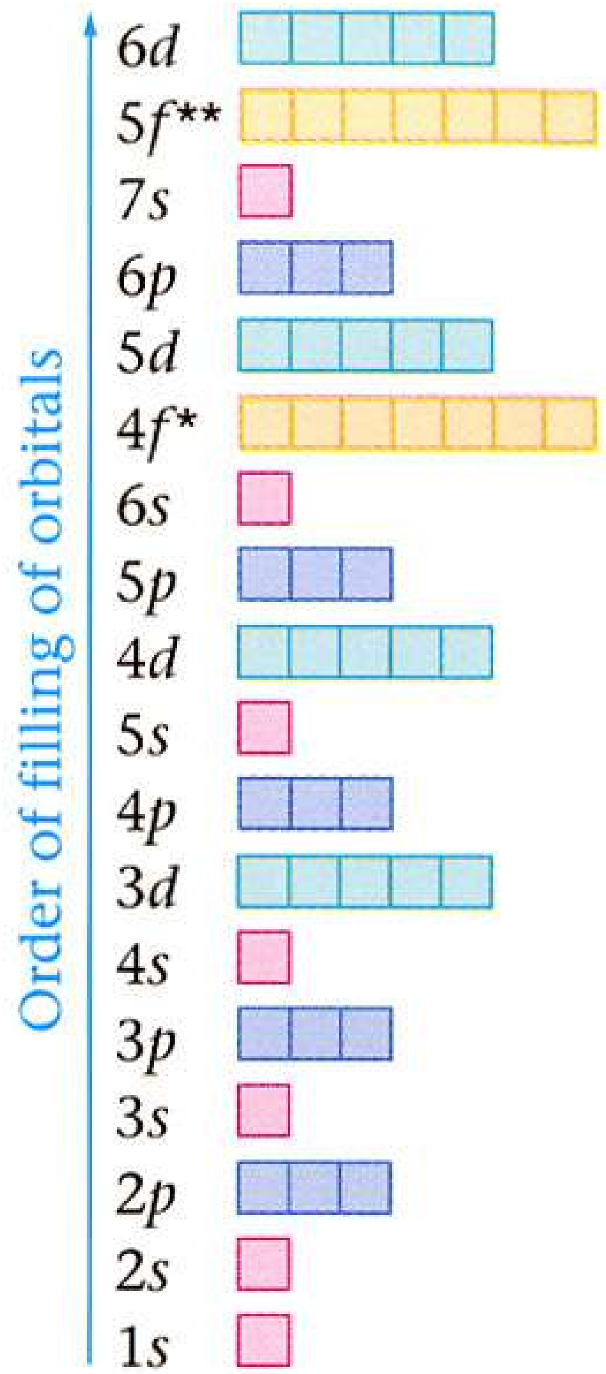


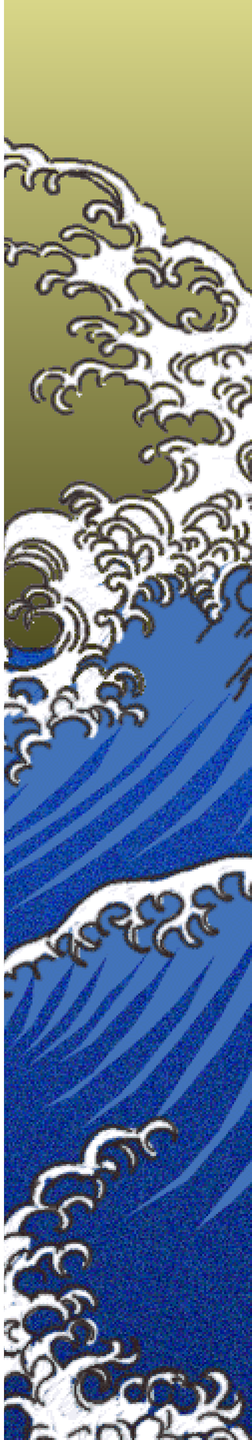
- [remember those 3d critters are called **transition metals**]
- [and, don't be too concerned right now about why Cr and Cu break the rules]

the big picture so far…



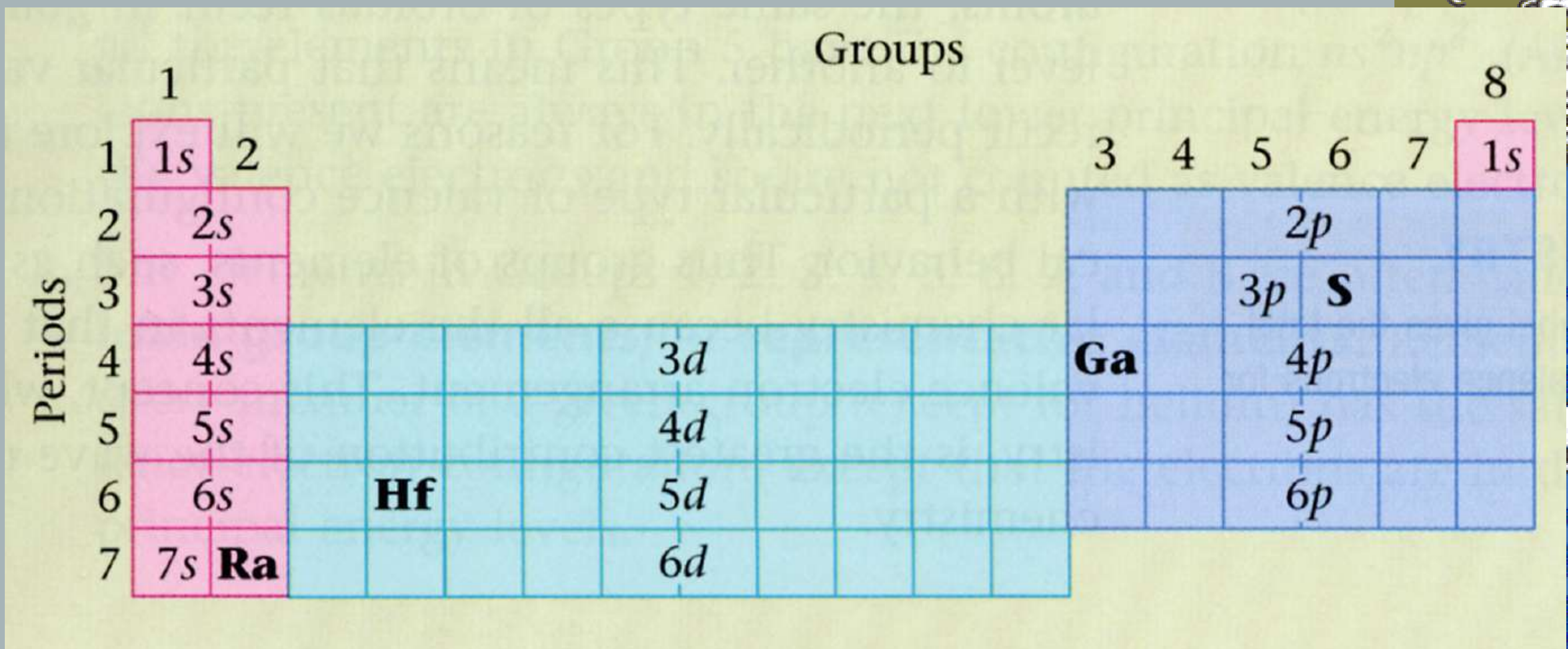
- here is how they fill using a **box diagram**
- you can use this *or* the PT to figure out just about any e-conf on the whole Table
- [sometime you'll **travel through** the **lanthanide (4f)** and **actinide (5f)** series, but I won't ask for any **in there**





example 11.3

- what are their e- configurations?



- and now, the Really Big Picture…

Representative Elements — *d*-Transition Elements — Representative Elements — Noble Gases

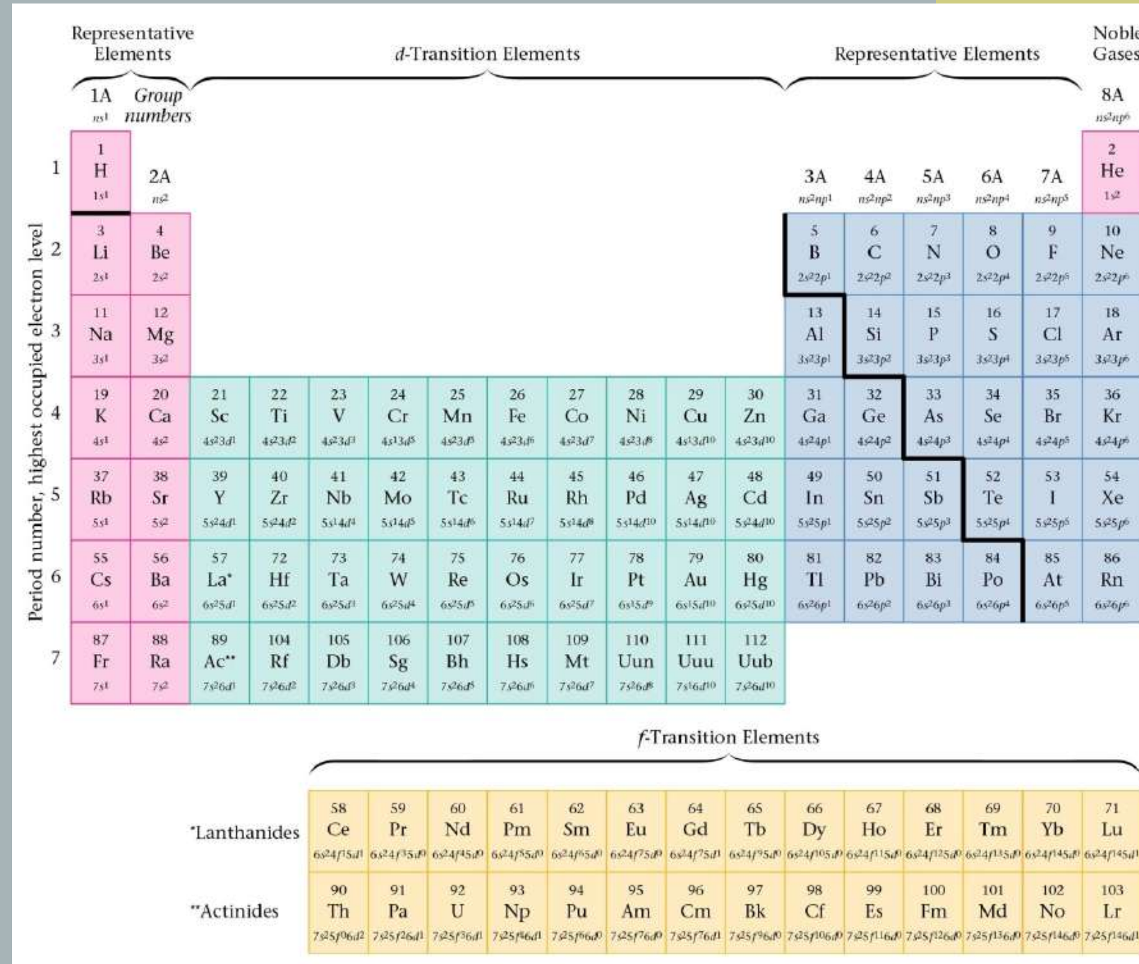
Group numbers; Period number, highest occupied electron level

Period	1A ns^1	2A ns^2											3A ns^2np^1	4A ns^2np^2	5A ns^2np^3	6A ns^2np^4	7A ns^2np^5	8A ns^2np^6
1	1 H $1s^1$																	2 He $1s^2$
2	3 Li $2s^1$	4 Be $2s^2$											5 B $2s^22p^1$	6 C $2s^22p^2$	7 N $2s^22p^3$	8 O $2s^22p^4$	9 F $2s^22p^5$	10 Ne $2s^22p^6$
3	11 Na $3s^1$	12 Mg $3s^2$											13 Al $3s^23p^1$	14 Si $3s^23p^2$	15 P $3s^23p^3$	16 S $3s^23p^4$	17 Cl $3s^23p^5$	18 Ar $3s^23p^6$
4	19 K $4s^1$	20 Ca $4s^2$	21 Sc $4s^23d^1$	22 Ti $4s^23d^2$	23 V $4s^23d^3$	24 Cr $4s^13d^5$	25 Mn $4s^23d^5$	26 Fe $4s^23d^6$	27 Co $4s^23d^7$	28 Ni $4s^23d^8$	29 Cu $4s^13d^{10}$	30 Zn $4s^23d^{10}$	31 Ga $4s^24p^1$	32 Ge $4s^24p^2$	33 As $4s^24p^3$	34 Se $4s^24p^4$	35 Br $4s^24p^5$	36 Kr $4s^24p^6$
5	37 Rb $5s^1$	38 Sr $5s^2$	39 Y $5s^24d^1$	40 Zr $5s^24d^2$	41 Nb $5s^14d^4$	42 Mo $5s^14d^5$	43 Tc $5s^14d^6$	44 Ru $5s^14d^7$	45 Rh $5s^14d^8$	46 Pd $5s^14d^{10}$	47 Ag $5s^14d^{10}$	48 Cd $5s^24d^{10}$	49 In $5s^25p^1$	50 Sn $5s^25p^2$	51 Sb $5s^25p^3$	52 Te $5s^25p^4$	53 I $5s^25p^5$	54 Xe $5s^25p^6$
6	55 Cs $6s^1$	56 Ba $6s^2$	57 La* $6s^25d^1$	72 Hf $6s^25d^2$	73 Ta $6s^25d^3$	74 W $6s^25d^4$	75 Re $6s^25d^5$	76 Os $6s^25d^6$	77 Ir $6s^25d^7$	78 Pt $6s^15d^9$	79 Au $6s^15d^{10}$	80 Hg $6s^25d^{10}$	81 Tl $6s^26p^1$	82 Pb $6s^26p^2$	83 Bi $6s^26p^3$	84 Po $6s^26p^4$	85 At $6s^26p^5$	86 Rn $6s^26p^6$
7	87 Fr $7s^1$	88 Ra $7s^2$	89 Ac** $7s^26d^1$	104 Rf $7s^26d^2$	105 Db $7s^26d^3$	106 Sg $7s^26d^4$	107 Bh $7s^26d^5$	108 Hs $7s^26d^6$	109 Mt $7s^26d^7$	110 Uun $7s^26d^8$	111 Uuu $7s^16d^{10}$	112 Uub $7s^26d^{10}$						

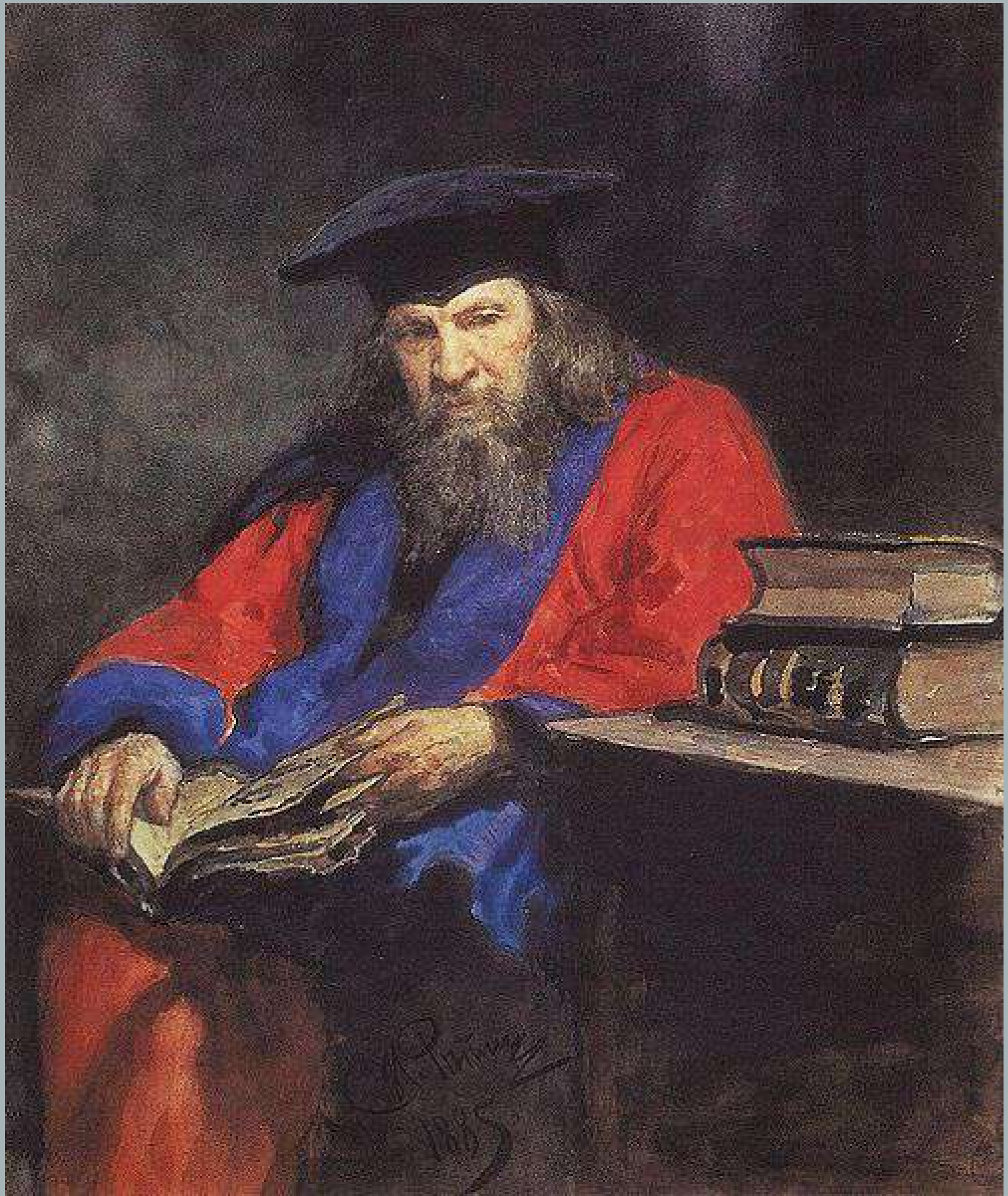
f-Transition Elements

*Lanthanides	58 Ce $6s^24f^15d^1$	59 Pr $6s^24f^35d^0$	60 Nd $6s^24f^45d^0$	61 Pm $6s^24f^55d^0$	62 Sm $6s^24f^65d^0$	63 Eu $6s^24f^75d^0$	64 Gd $6s^24f^75d^1$	65 Tb $6s^24f^95d^0$	66 Dy $6s^24f^{10}5d^0$	67 Ho $6s^24f^{11}5d^0$	68 Er $6s^24f^{12}5d^0$	69 Tm $6s^24f^{13}5d^0$	70 Yb $6s^24f^{14}5d^0$	71 Lu $6s^24f^{14}5d^1$
**Actinides	90 Th $7s^25f^06d^2$	91 Pa $7s^25f^26d^1$	92 U $7s^25f^36d^1$	93 Np $7s^25f^46d^1$	94 Pu $7s^25f^66d^0$	95 Am $7s^25f^76d^0$	96 Cm $7s^25f^76d^1$	97 Bk $7s^25f^96d^0$	98 Cf $7s^25f^{10}6d^0$	99 Es $7s^25f^{11}6d^0$	100 Fm $7s^25f^{12}6d^0$	101 Md $7s^25f^{13}6d^0$	102 No $7s^25f^{14}6d^0$	103 Lr $7s^25f^{14}6d^1$

- you'll notice the **group A** numbers indicate the ***total valence electrons***
- remember from CHMA that these groups (1-8) are called **representative elements** (aka **main-group elements**)



11.11 atomic properties and the periodic table



11.11 atomic properties and the periodic table

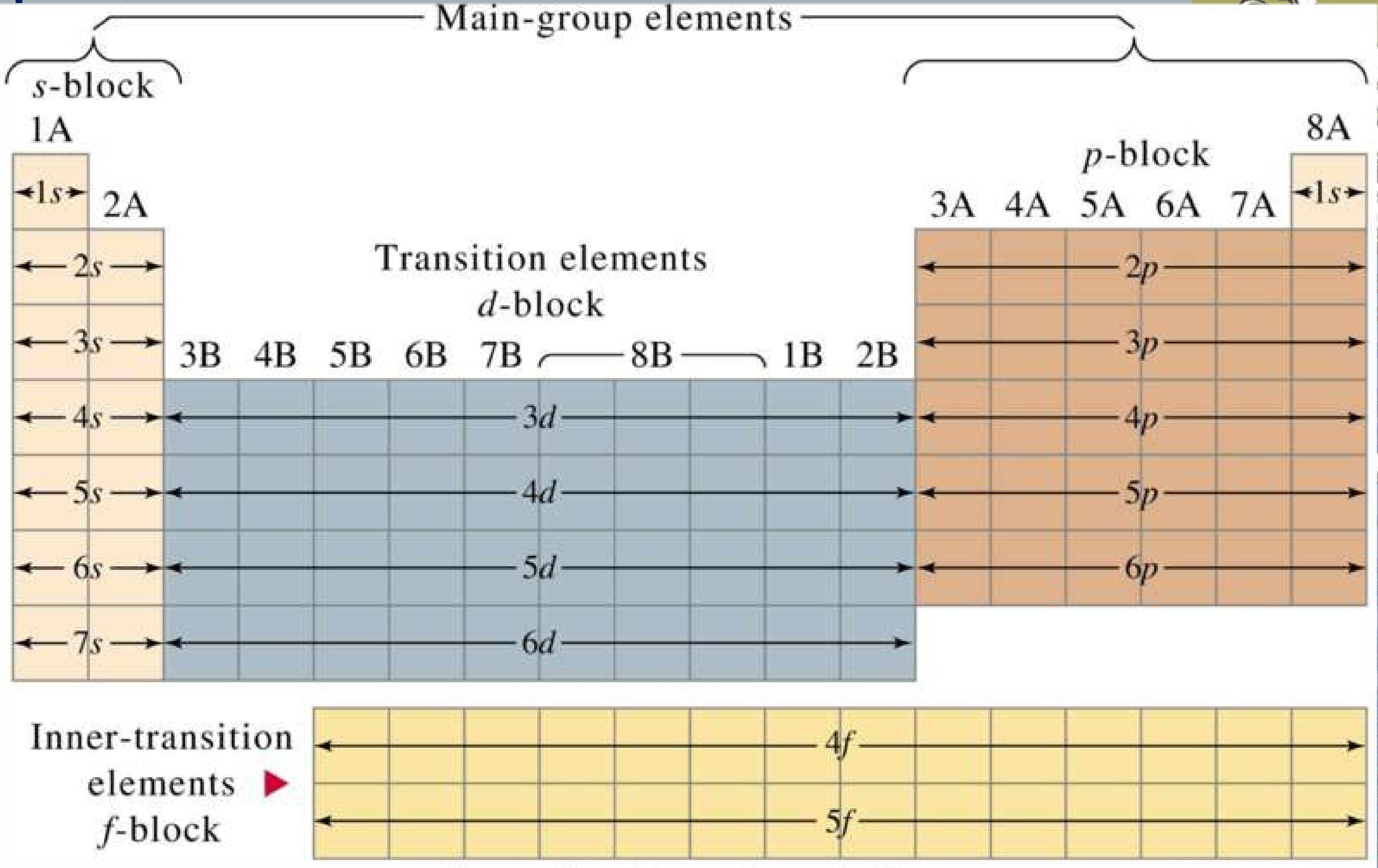
ОПЫТЪ СИСТЕМЫ ЭЛЕМЕНТОВЪ.

ОСНОВАННОЙ НА ИХЪ АТОМНОМЪ ВѢСѢ И ХИМИЧЕСКОМЪ СХОДСТВѢ.

			Ti=50	Zr=90	?=180.
			V=51	Nb=94	Ta=182.
			Cr=52	Mo=96	W=186.
			Mn=55	Rh=104,4	Pt=197,4.
			Fe=56	Rn=104,4	Ir=198.
			Ni=Co=59	Pl=106,6	Os=199.
H=1			Cu=63,4	Ag=108	Hg=200.
	Be=9,4	Mg=24	Zn=65,2	Cd=112	
	B=11	Al=27,4	?=68	Ur=116	Au=197?
	C=12	Si=28	?=70	Sn=118	
	N=14	P=31	As=75	Sb=122	Bi=210?
	O=16	S=32	Se=79,4	Te=128?	
	F=19	Cl=35,5	Br=80	I=127	
Li=7	Na=23	K=39	Rb=85,4	Cs=133	Tl=204.
		Ca=40	Sr=87,6	Ba=137	Pb=207.
		?=45	Ce=92		
		?Er=56	La=94		
		?Yt=60	Di=95		
		?In=75,6	Th=118?		

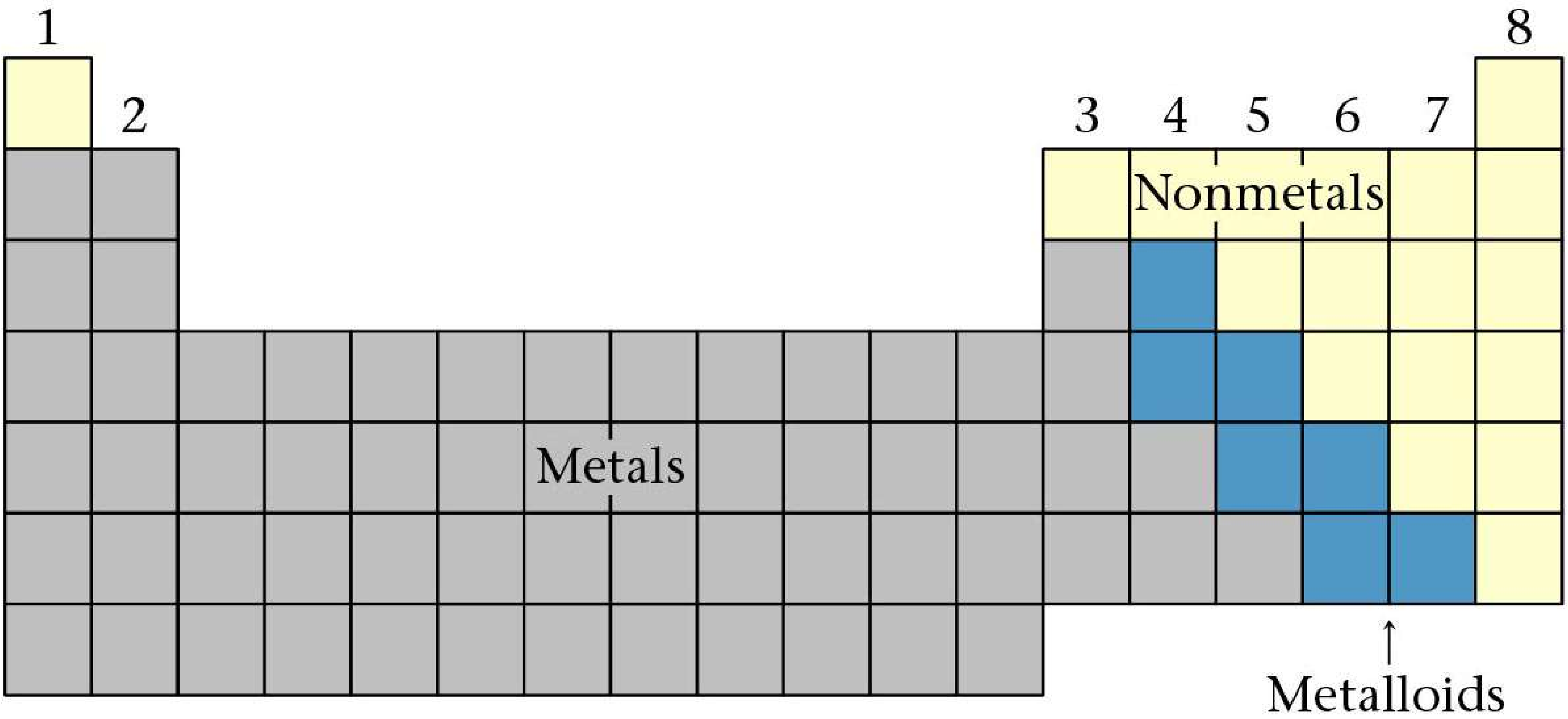
Д. Менделѣевъ

11.11 atomic properties and the periodic table



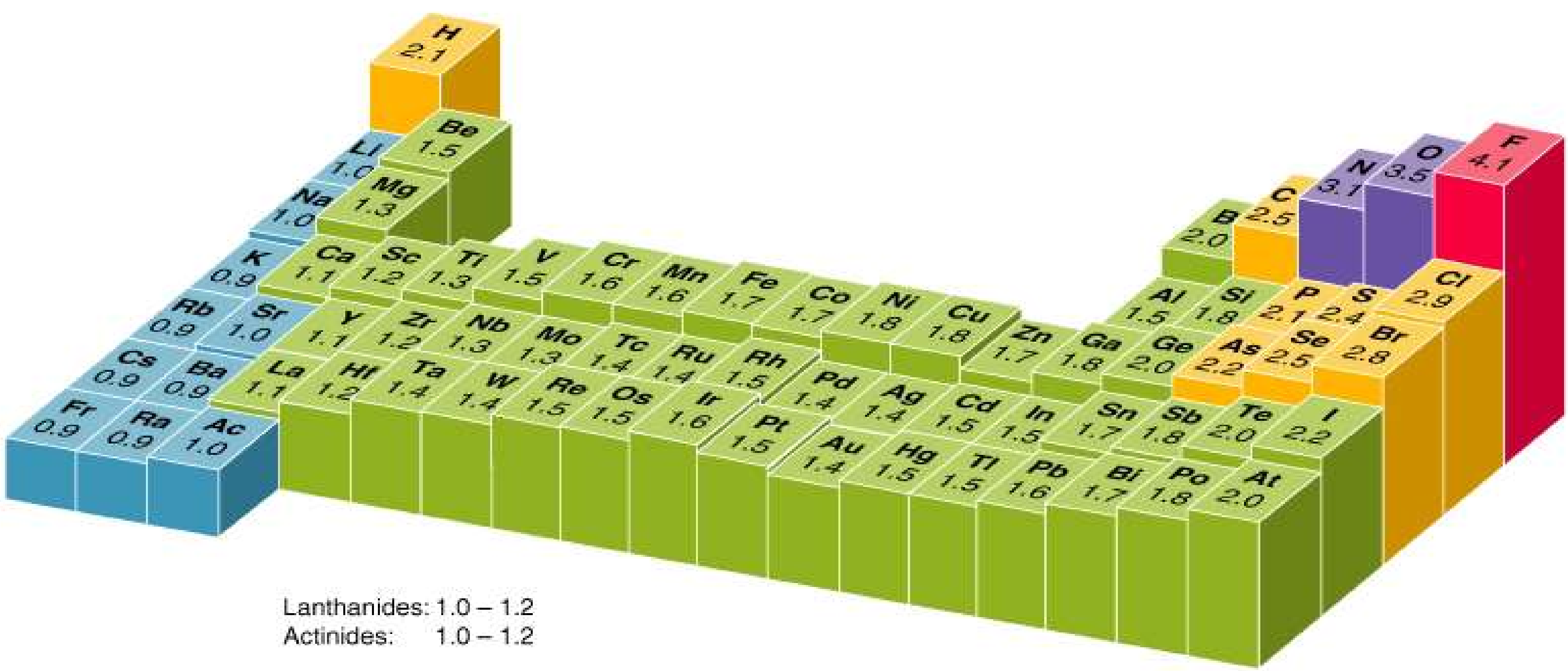
metals and nonmetals

- metals tend to lose e- to make positive ions, nonmetals will gain to form negative ions



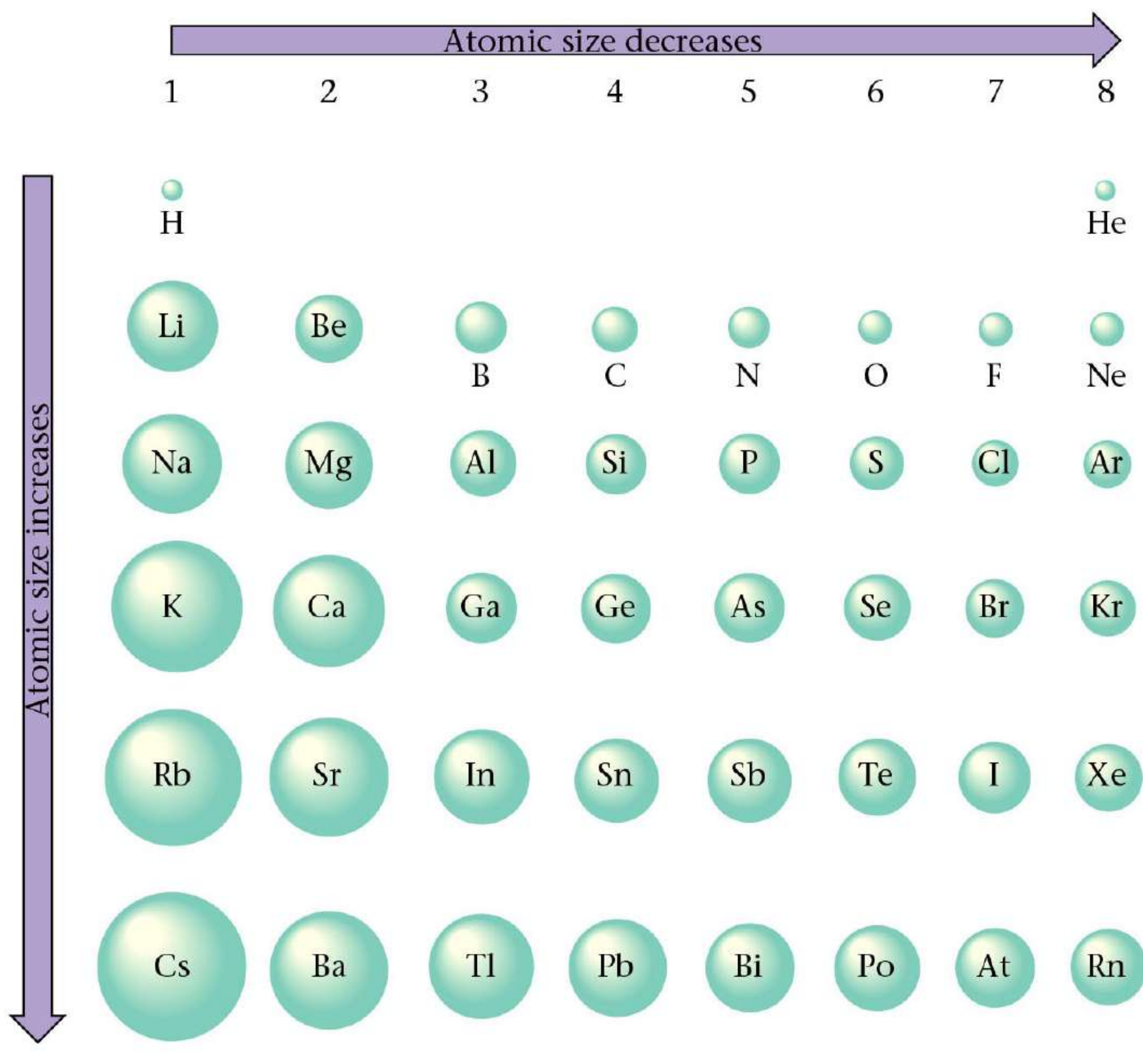
Electronegativity

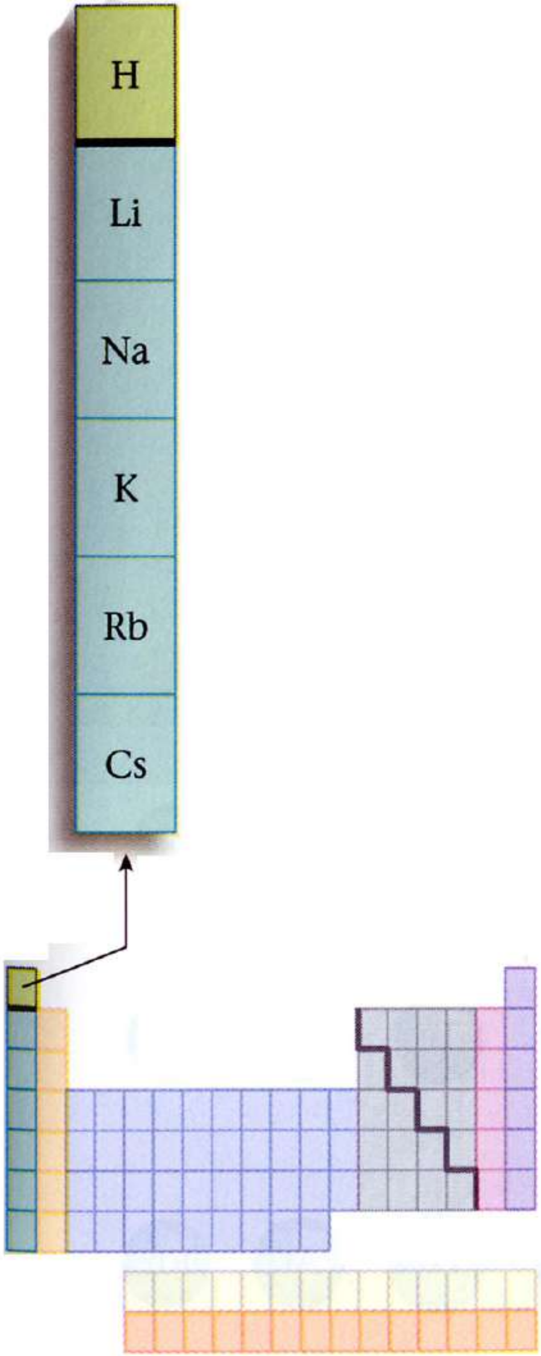
- a chemical property that describes the ability of an atom to attract electrons towards itself. An atom's electronegativity is affected by both its atomic weight and the distance that its valence electrons reside from the charged nucleus.



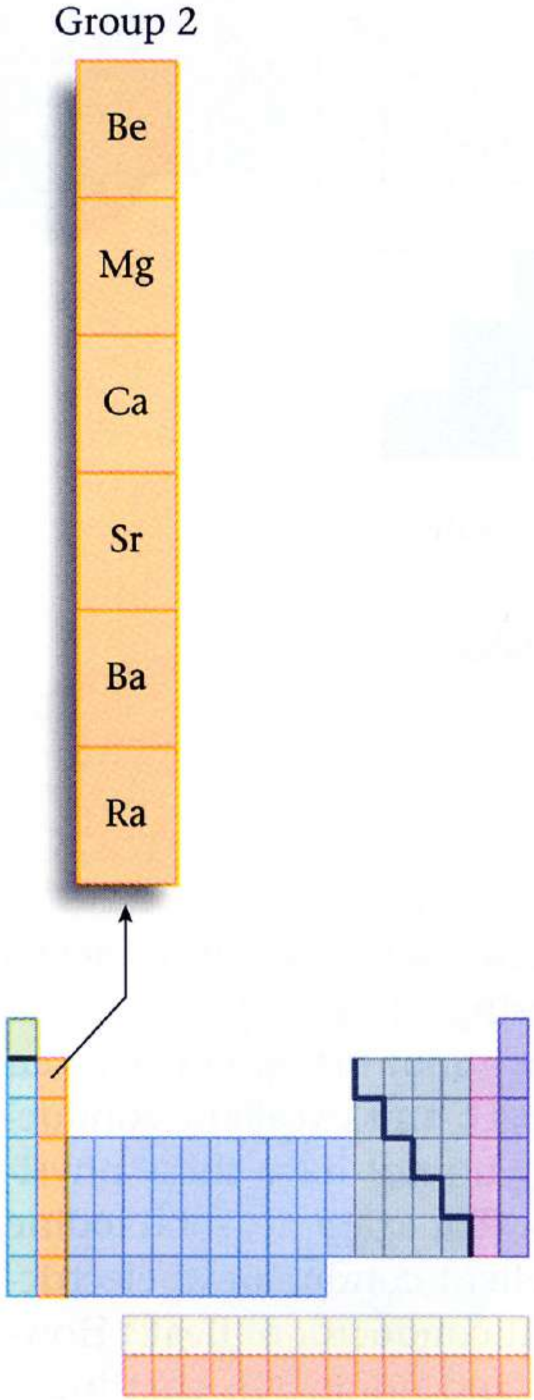
atomic size

- see a trend?
- can you explain the trend?
- going down a group is just adding shells; that’s easy
- but why the horizontal trend?





- **exp** shows for the Gr1A, Cs loses electrons ***very*** easily, then Rb, then K, Na; Li doesn’t lose them easily



- Consider group 2
- Ra is easiest to steal from; Be least likely to give up electron
- on the other side (nonmetals) the *opposite* occurs; the **upper** guys are most likely to ***steal*** **electrons**
- **summary: the most active metals are lower left; the most active nonmetals are upper right**

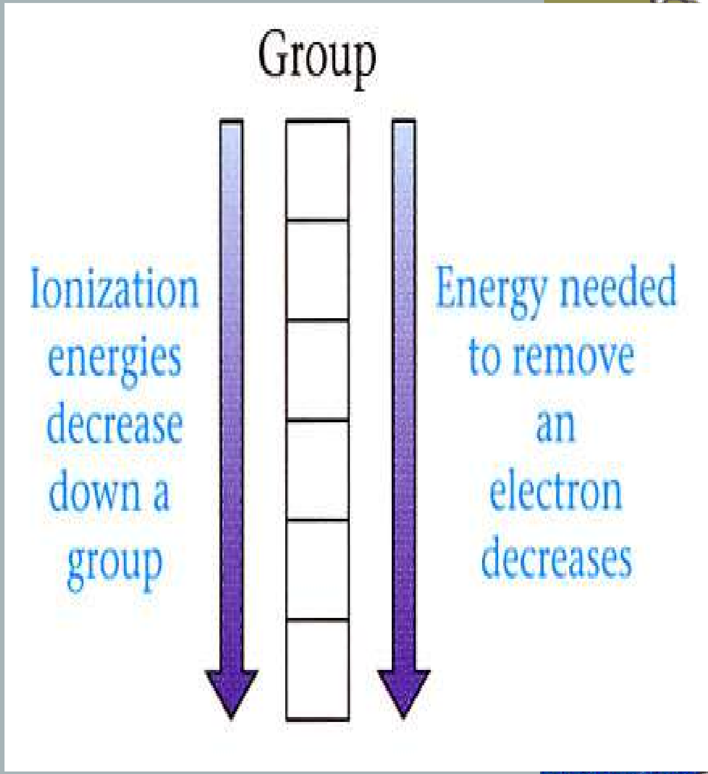
- *as we go across* **the atomic number (# protons) is increasing!**
- and the protons pull in the electron cloud a little tighter each time
- ∴ it gets smaller going across

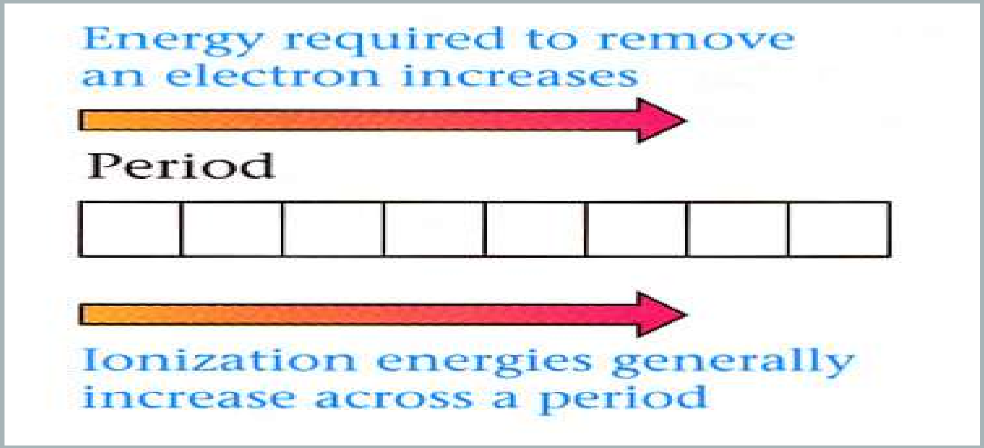
ionization energies

- you put E into an atom, and eventually the most loosely held electron comes flying off

$$M_{(g)} \rightarrow M^{+}_{(g)} + e^{-}$$

- the energy required to do this is less and less as we go down a group (the e- is farther away and easier to remove)





- but as we go ***across***, as we go more towards the area of the Nonmetals it becomes more difficult to strip one off, and the ionization energy gets higher and higher

mini-summary

- the elements at the **lower left** have lowest ionization energy and very easily react w/ things *(they are the most chemically active metals!)*
- the elements at the **upper right** have the highest ionization energy; they won't lose electrons, *(they are most chemically active nonmetals)*

Questions?