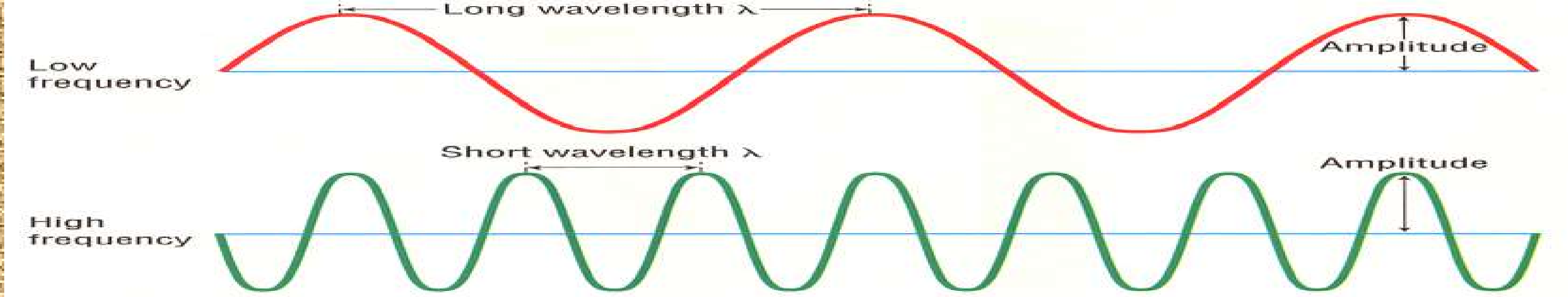
Atomic Structure and Periodicity

Chapter 7

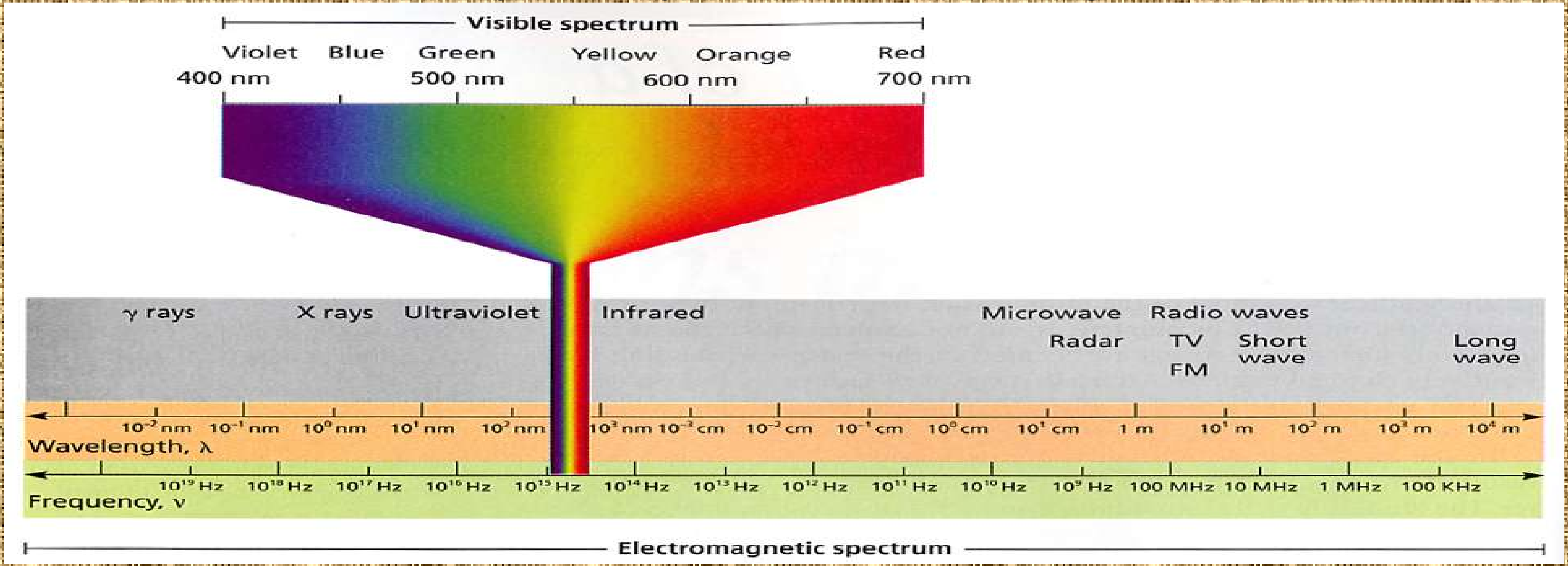
Electromagnetic Radiation

- Energy that travels through space in the form of waves. (simplified definition)
- Wavelength (λ)– distance between 2 peaks
- Frequency (ν)- # of waves that pass a point per second
- Speed – EM Radiation travels at speed of light ($c=3.0\times10^{8}$m/s)



EM Spectrum

- Which wavelength has the greatest energy?
- Which has the least energy?
- How are wavelength and energy related?
- How are frequency and wavelength related?



More About This Stuff

- Inverse relationship between frequency and wavelength.
- **$c = \lambda\nu$**

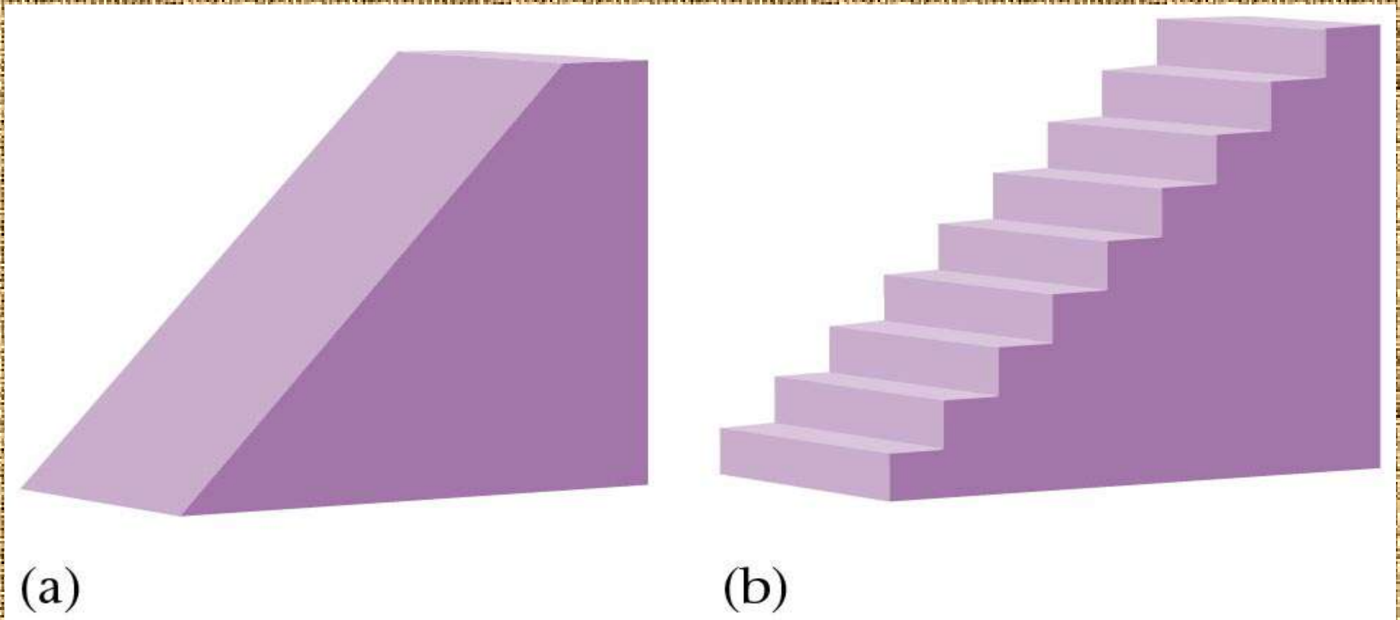
  λ = metersc = 3.0×10^8 m/s

  ν = cycles = 1/sec = sec^{-1} = hertz (Hz)
- 1. What is the wavelength of light with a frequency 5.89×10^5 Hz?
- 2. What is the frequency of blue light with a wavelength of 484 nm?

HW: 31,32

Planck and Quatized Energy

- Found that energy is quantized.
- Matter can only absorb or emit certain quantities of energy.
- Like stairs as opposed to a ramp.



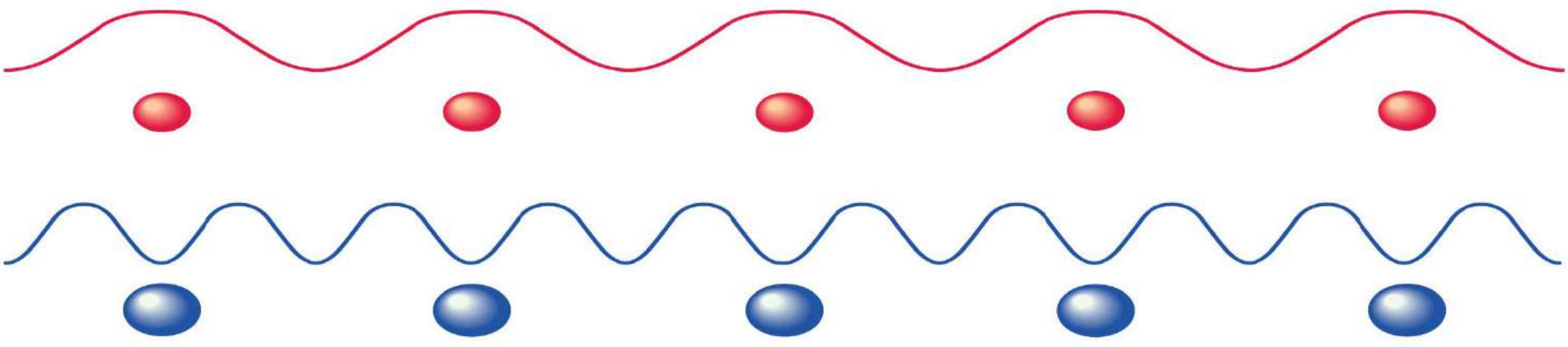
More Planck

- Determined a formula for the energy of a quantum (packet of quantized energy) and a constant that has his name (awesome).
- ΔE = hνor ΔE = hc/λ
  - ν = frequency of EM radiation
  - H = Planck's constant = 6.63 x 10-34 J·s

1. Calculate the quantum (increment of energy) emitted at a wavelength of 520nm.
2. What wavelength of light is capable of emitting a quantum of 4.41 x 10-19 J?

Einstein EM Radiation is Quantized

- Planck said energy is quantized.
- Einstein hypothesized EM radiation is quantized.
- EM Radiation is now seen as traveling as waves AND a stream of particles (called photons).
- Energy of each photon is given by the equation:
- $E_{photon} = h\nu = hc/\lambda$



Calculating Energy of Photons

34. What is the energy of a single UV photon and a mole of UV photons having a wavelength of 25nm?

40. It takes 208.4 kJ of energy to remove 1 mole of electrons from the surface of rubidium metal. How much E to remove a single electron? What is the maximum wavelength capable of doing this?

HW: 33-39 odd

De Broglie’s Equation

- Now able to calculate the λ of a particle.
- Knowing $E=hc/\lambda$and $E=mc^2$
- Then $mc^2 = hc/\lambda$
- Som $= h/\lambda c$
- But a particles speed is not the speed of light
- Som $= h/\lambda v$
- Therefore$\lambda = h/mv$

Note: Here v = speed of particle in m/s
m = mass of particle in kg

De Broglie Equation

1. Kyle Korver, an NBA player (and my future husband ☺), shoots a basketball (m=926 g) with a velocity of 6.21 mph (2.78 m/s). What is the de Broglie wavelength of the basketball?
2. What velocity would an electron (m=9.109x10^{-31} kg) need for its de Broglie wavelength to be that of red light (750 nm)?

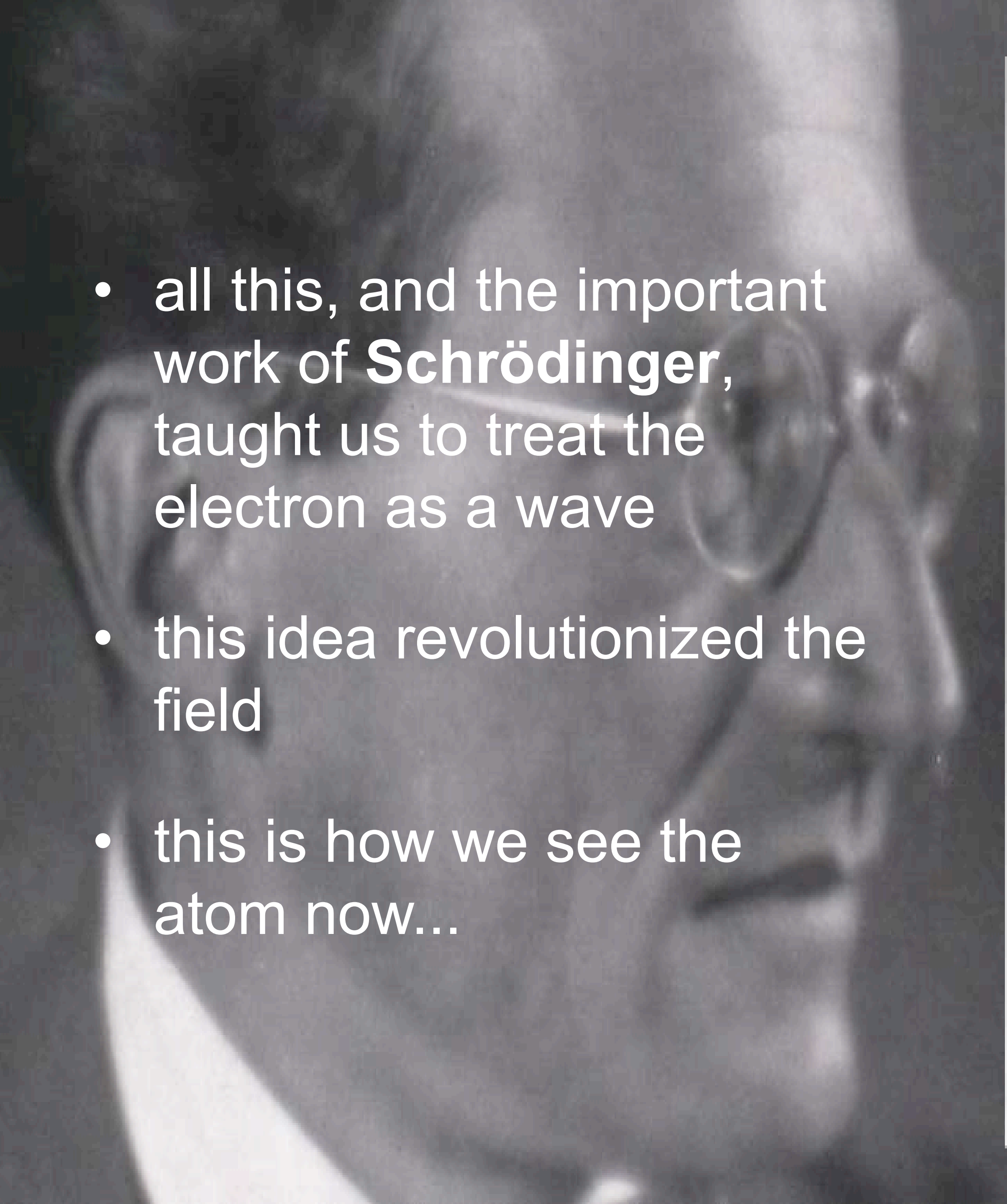
HW: 41,43

Diamagnetism & Paramagnetism

- **diamagnetic** just means all e- are paired up
- **paramagnetic** means there is a loner in there somewhere
- paramagnetic atom (and molecules) are strongly affected by magnetic fields; not too much for diamagnetic

The Heisenberg Uncertainty Principle

- Werner H came along and said we can’t know both the position and the momentum of an electron
- meaning: we can’t know that the electrons go around like little planets; we can just know probably where it is



- all this, and the important work of **Schrödinger**, taught us to treat the electron as a wave
- this idea revolutionized the field
- this is how we see the atom now...

Quantum Numbers

- the positions of electrons with respect to the nucleus are described by **quantum numbers**
- each electron has four QN to narrow down where it ***might*** be
- like four bits of an address to help specify
- ***shell, subshell, orbital, spin...***

7.5-7.8

- **shells: n = 1, 2, 3...**
- “**n**” is called the ***principal quantum number***
- almost always tells us ***average distance*** from nucleus (higher numbers indicate greater distance)
- also tell us ***relative energy*** (higher numbers indicate higher energy, less stability)

- **subshells: ℓ = 0, 1, 2...**
- aka *angular momentum number*
- this second QN tells us the ***shape*** of the electron's orbital..

- first shell ($n=1$) has just one subshell:
  s ($\ell=0$)

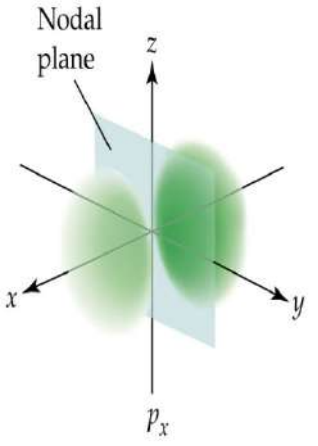
- second shell ($n=2$) has two subshells:
  s ($\ell=0$) and **p** ($\ell=1$)

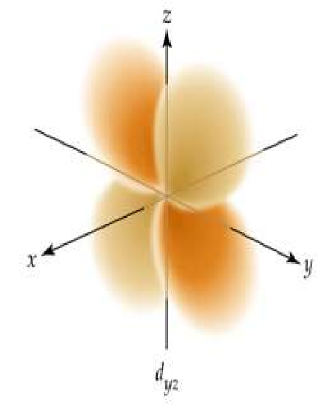
- third shell ($n=3$) has three subshells:
  s ($\ell=0$), **p** ($\ell=1$), and **d** ($\ell=2$)

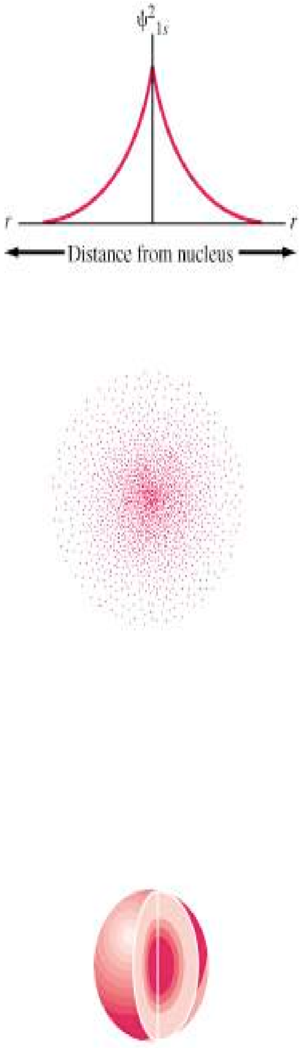
- see a number pattern?

- **ℓ will always be from 0 to ($n-1$)**

- the shapes are:
- *s = spherical*
- *p = dumbbell shaped*
- *d = cloverleaf*

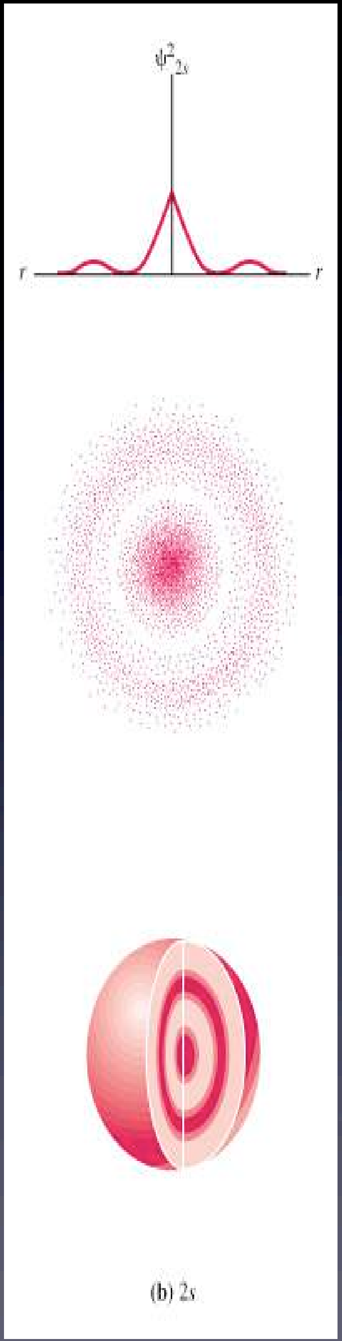


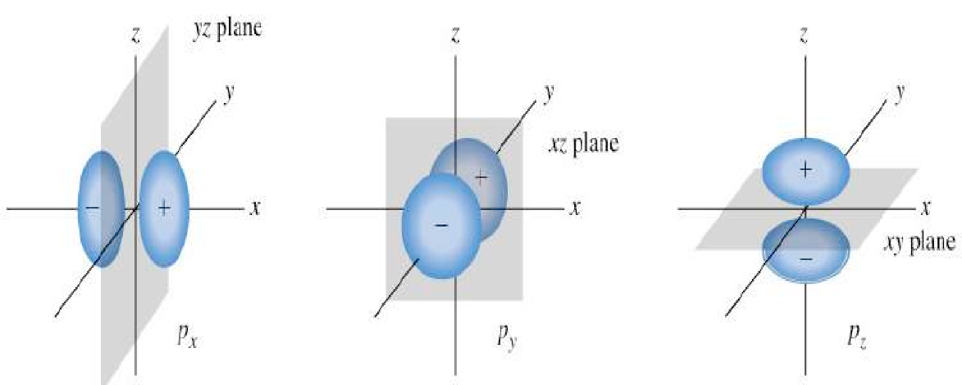




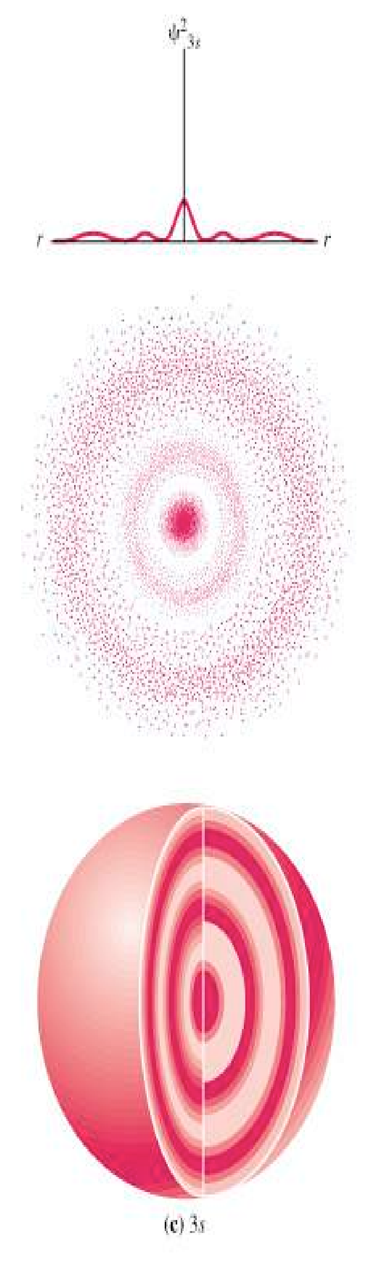
(a) 1s

- at the first level there is just this one orbital (1s)

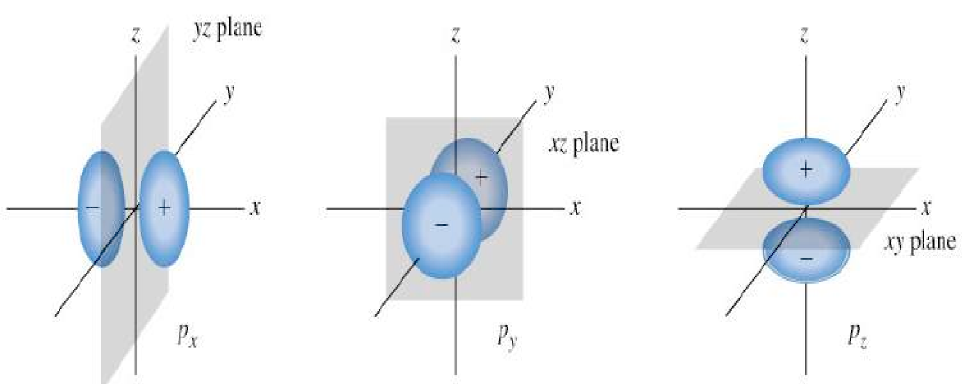




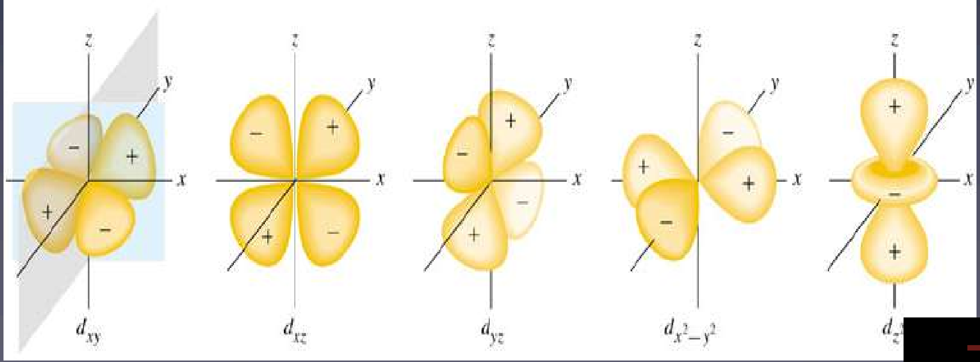
- at level two there are the one 2s and three 2p’s



(c) 3s



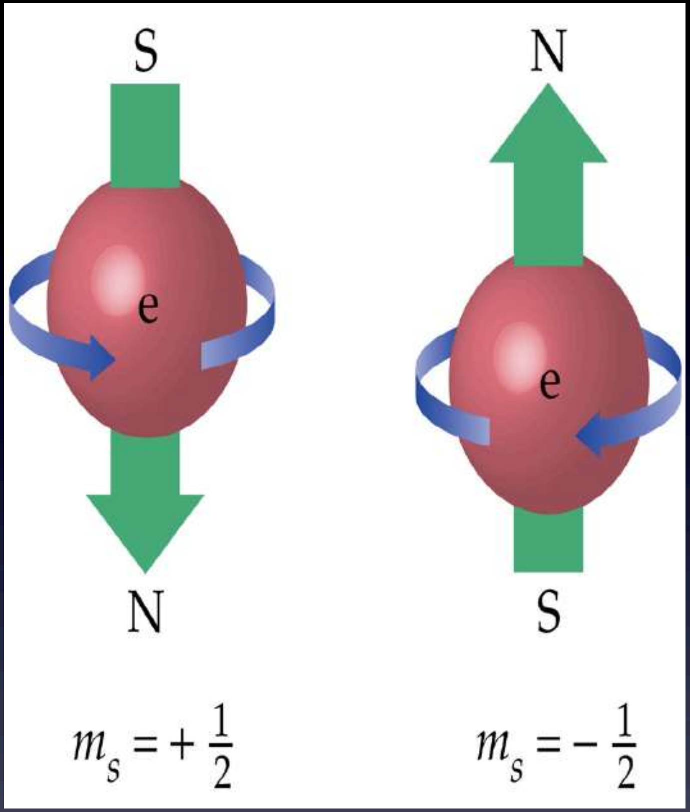
- at the level 3 there are one 3s, three 3p’s and five 3d’s
- see another pattern yet?



7.5-7.8

- **orbitals: m_ℓ = ... -1, 0, +1...**
- this is called the *magnetic quantum number*, or **orbital**, and describes ***orientation in space***
- *where is orbital on the x, y, z axes?* m_ℓ can tell you
- for now..

- the s subshell (ℓ=0) being spherical, can only be placed in **one** orientation: **$m_\ell=0$**
- the p subshell (ℓ=1) has **three** different ways it can be placed: **$m_\ell=-1$, $m_\ell=0$, $m_\ell=+1$**
- the d subshell (ℓ=2) has **5** different ways to be placed in space: **m_ℓ = -2, -1, 0, +1, +2**
- **m at any level will have the value $-\ell$ to $+\ell$**



- **spin: m_s = +1/2 or -1/2**
- each orbital we've seen so far can house two electrons, one spinning one way, one the other way

TABLE 5.2 Allowed Combinations of Quantum Numbers *n*, *l*, and m_l for the First Four Shells

n	l	m_l	Orbital Notation	Number of Orbitals in Subshell	Number of Orbitals in Shell
1	0	0	1*s*	1	1
2	0	0	2*s*	1	4
	1	−1, 0, +1	2*p*	3	
3	0	0	3*s*	1	9
	1	−1, 0, +1	3*p*	3	
	2	−2, −1, 0, +1, +2	3*d*	5	
4	0	0	4*s*	1	16
	1	−1, 0, +1	4*p*	3	
	2	−2, −1, 0, +1, +2	4*d*	5	
	3	−3, −2, −1, 0, +1, +2, +3	4*f*	7	

7.5-7.8

- now we describe the electrons *in a given element* using all this QN junk
- be patient; it is really much easier than it first appears
- but first: gotta know a couple rules...

The Aufbau Principle

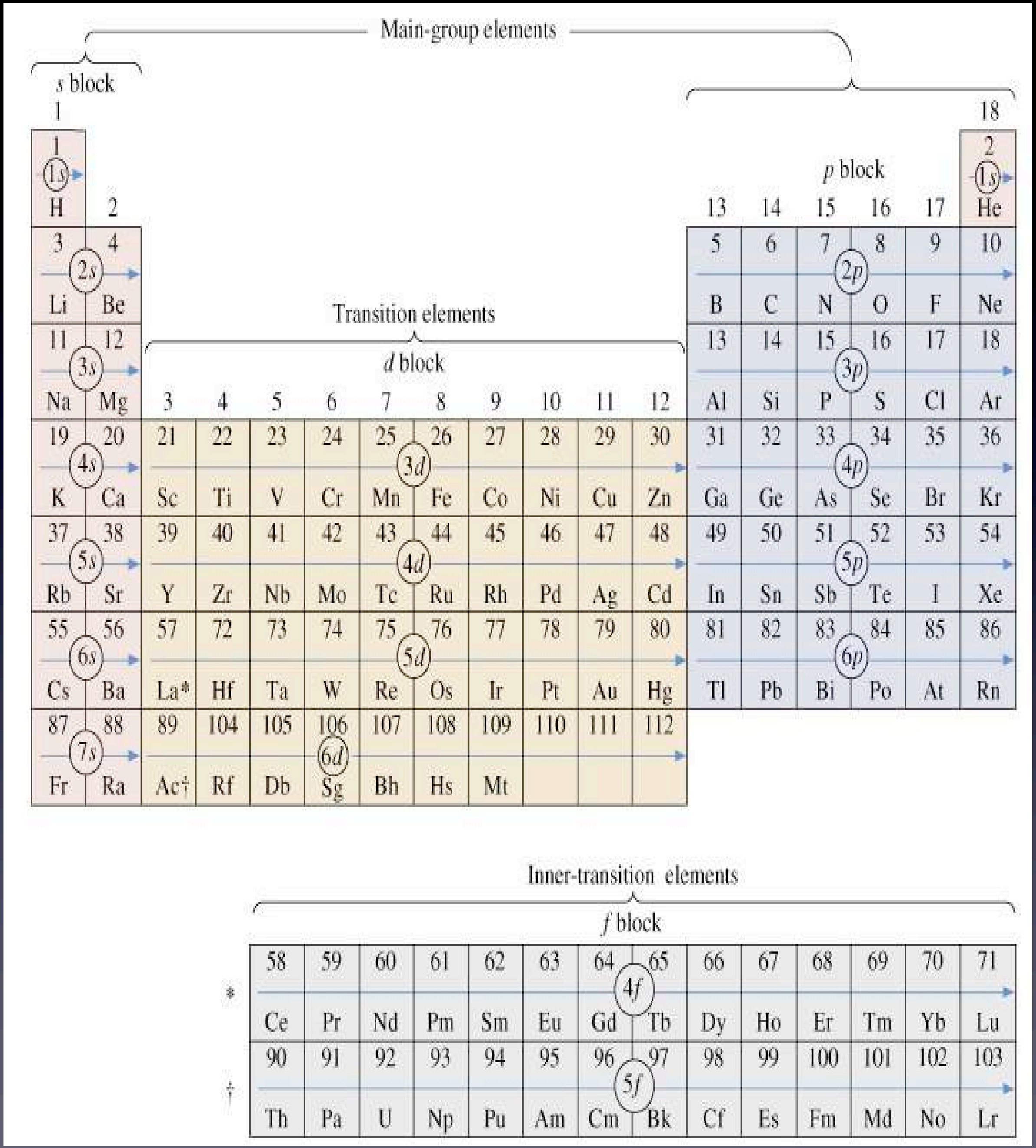
- **aufbau** from the German “to build up”
- when seeing how electrons are “configured” in an atom we have to start from the bottom and work the way up, low energy to hi energy

The Pauli Exclusion Principle

- no two electrons can occupy the exact same orbital, so...
- *every electron has it own unique set of 4 QNs*

QN and the PT

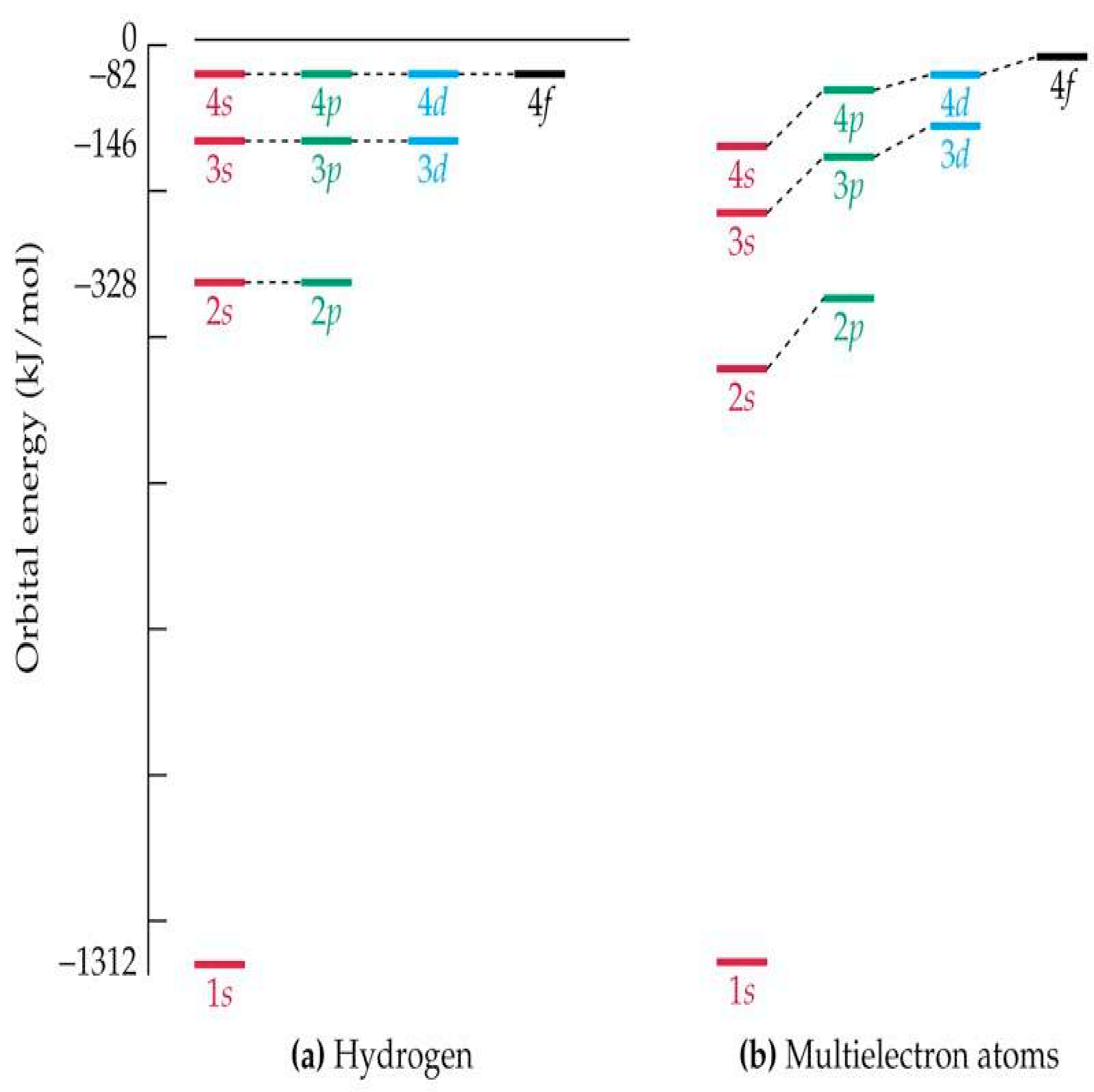
- using the PT to tell QN is profoundly important and easy!
- the first two QNs (e.g. 3s) can be seen here....

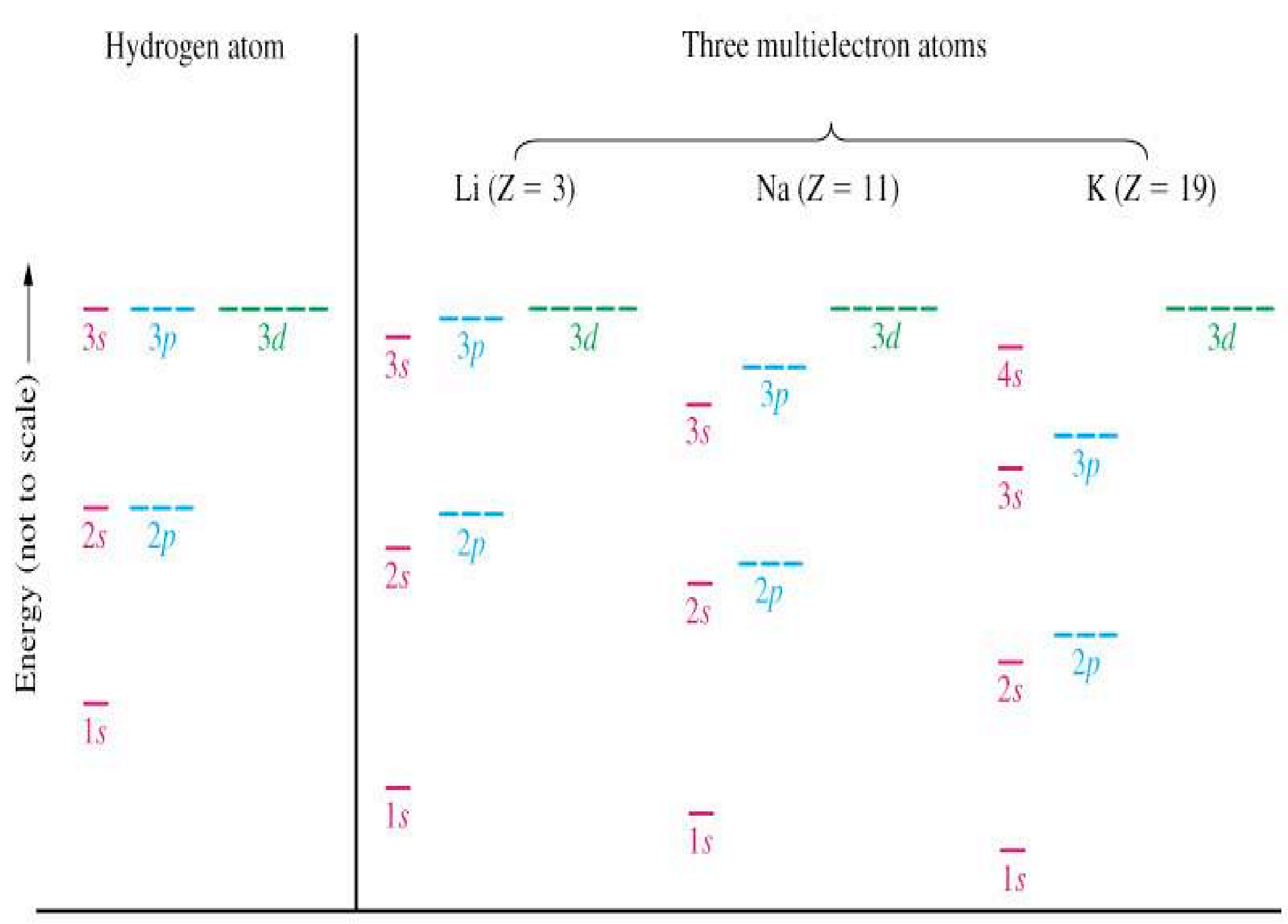


1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr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

*	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
†	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

- you can also see the Aufbau Principle when reading it left to right
- notice what happens to the filling order once you get past 3p... does it go to 3d?


0
–82
–146
–328
–1312
Orbital energy (kJ/mol)
4s
4p
4d
4f
3s
3p
3d
2s
2p
1s
(a) Hydrogen
4f
4d
4p
3d
4s
3p
3s
2p
2s
1s
(b) Multielectron atoms


Hydrogen atom
Three multielectron atoms
Li (Z = 3)
Na (Z = 11)
K (Z = 19)
Energy (not to scale)
3s
3p
3d
2s
2p
1s
4s

eocs

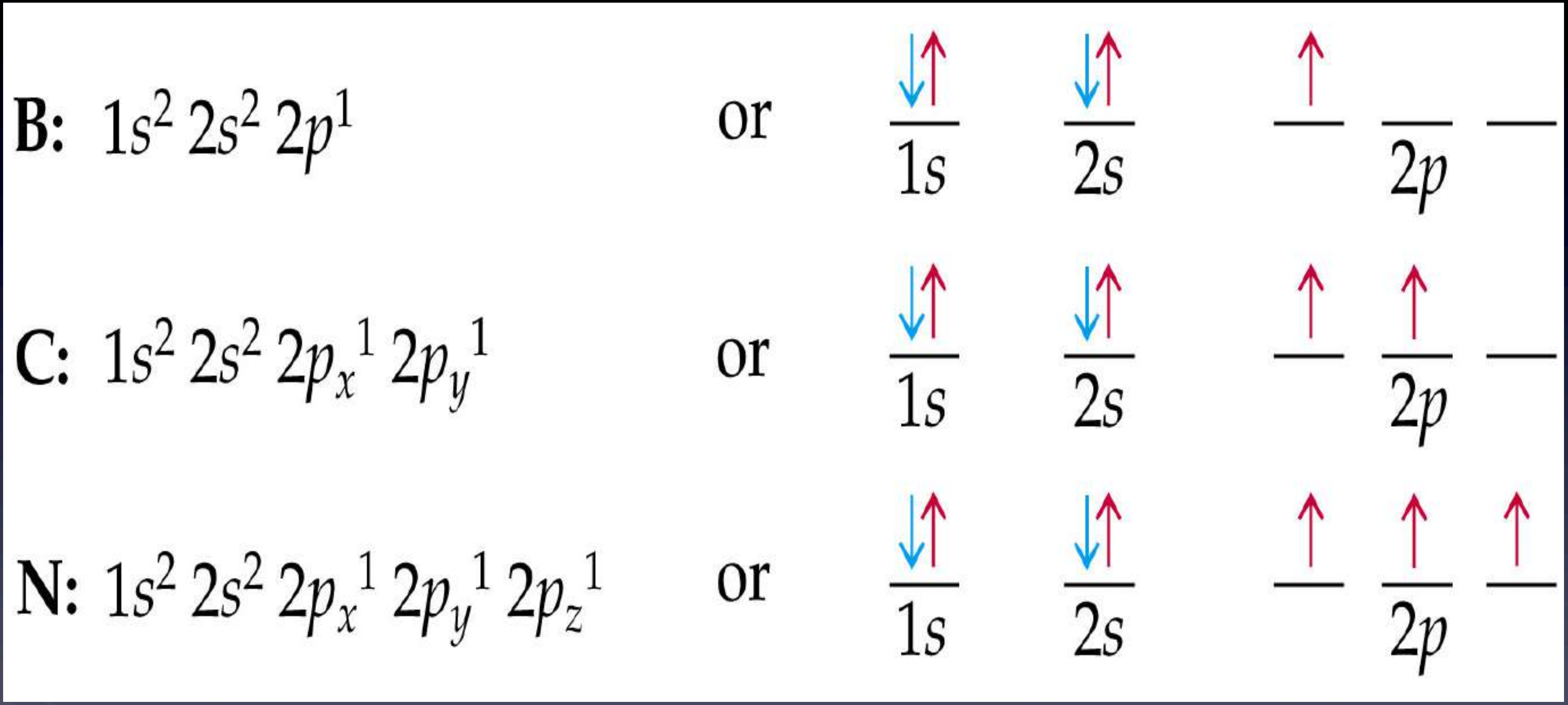
- 7.57-60

Hund’s Rule

- the rule says that if you add an electron to a subshell it will go into one that is empty and only “pair up” if it has to

Electron Configurations

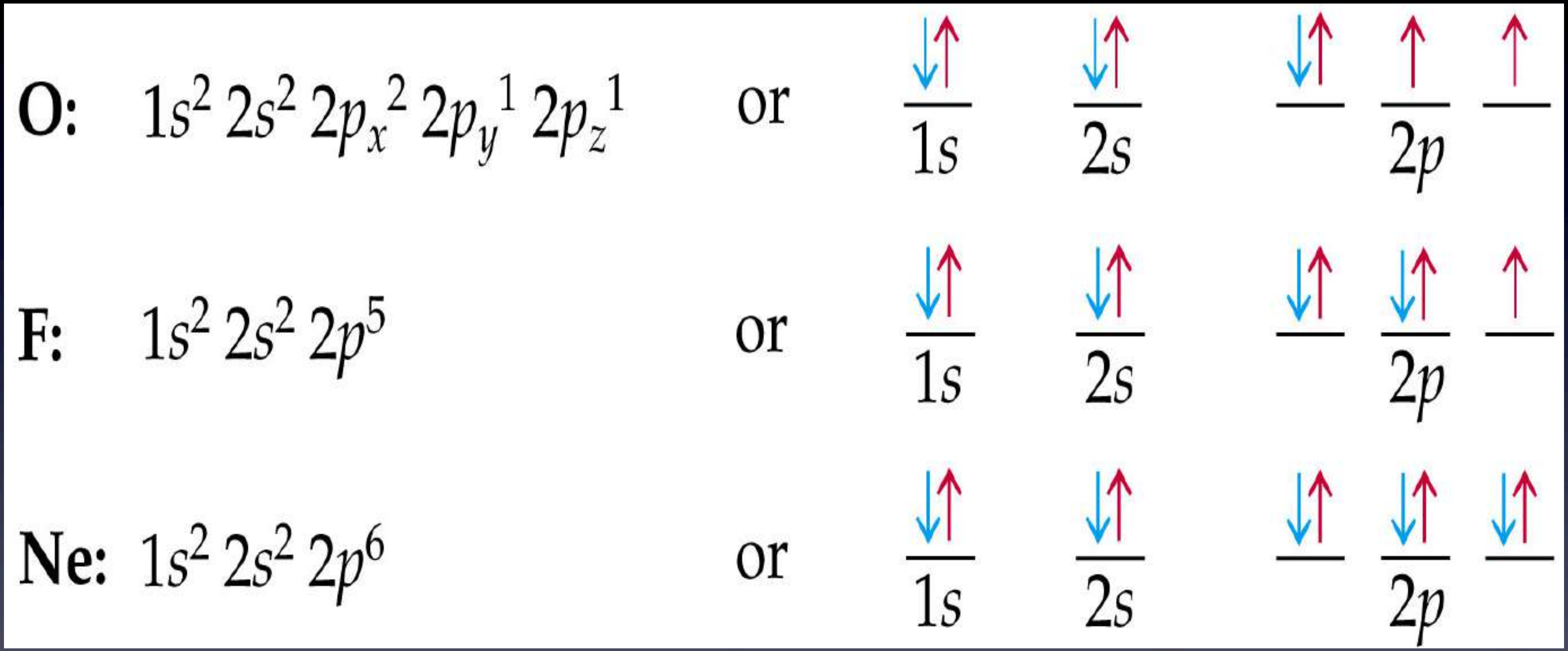
- putting all these rules together we can see the way the electrons are configured around any element on the PT
- watch how I do it...
- now use the PT to do any electron configuration



B: $1s^2\,2s^2\,2p^1$ or

C: $1s^2\,2s^2\,2p_x^{\,1}\,2p_y^{\,1}$ or

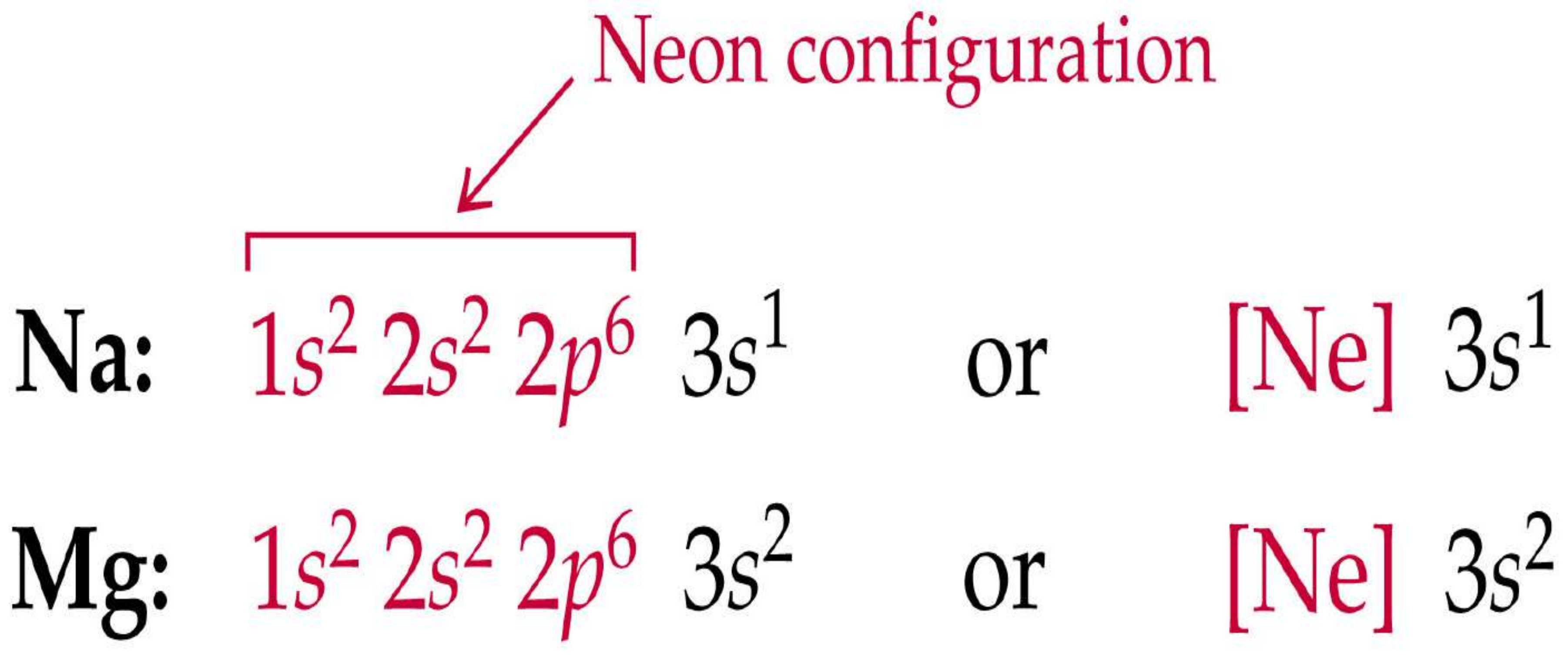
N: $1s^2\,2s^2\,2p_x^{\,1}\,2p_y^{\,1}\,2p_z^{\,1}$ or



O: $1s^2\,2s^2\,2p_x{}^2\,2p_y{}^1\,2p_z{}^1$ or

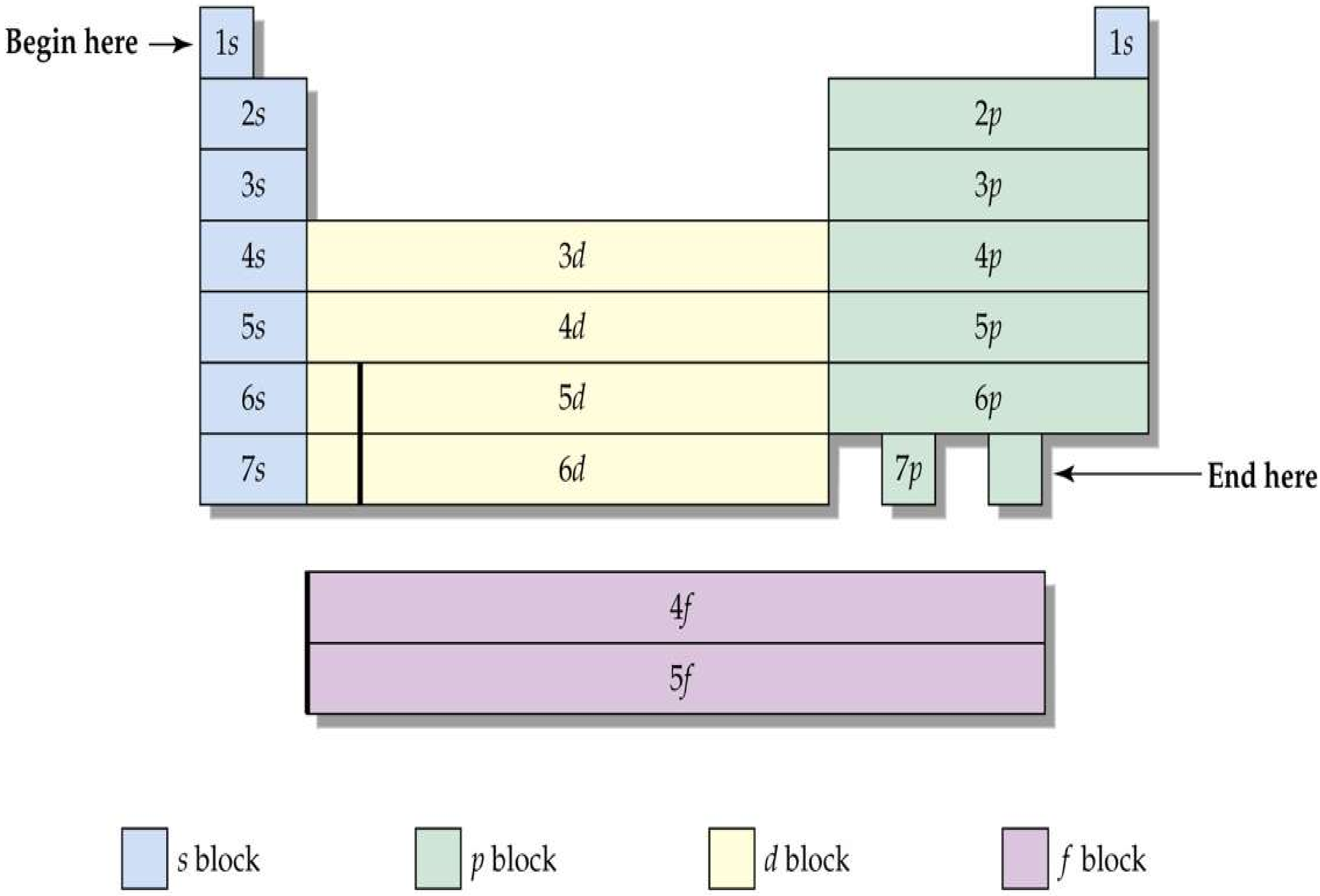
F: $1s^2\,2s^2\,2p^5$ or

Ne: $1s^2\,2s^2\,2p^6$ or



Na: $\overbrace{1s^2\,2s^2\,2p^6}^{\text{Neon configuration}}\;3s^1$ or $[\mathrm{Ne}]\,3s^1$

Mg: $1s^2\,2s^2\,2p^6\;3s^2$ or $[\mathrm{Ne}]\,3s^2$


Begin here
1s
1s
2s
3s
4s
5s
6s
7s
3d
4d
5d
6d
2p
3p
4p
5p
6p
7p
End here
4f
5f
s block
p block
d block
f block

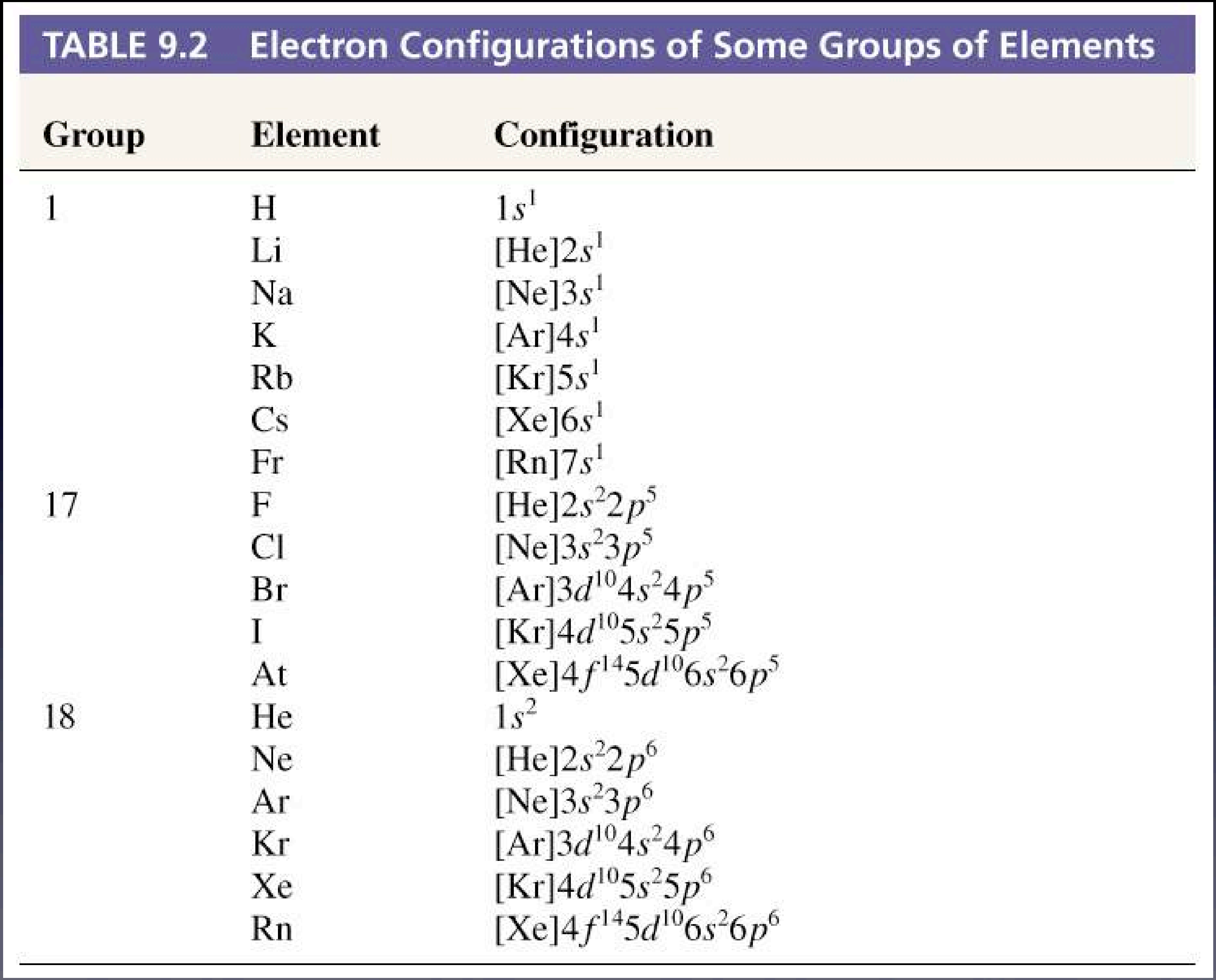
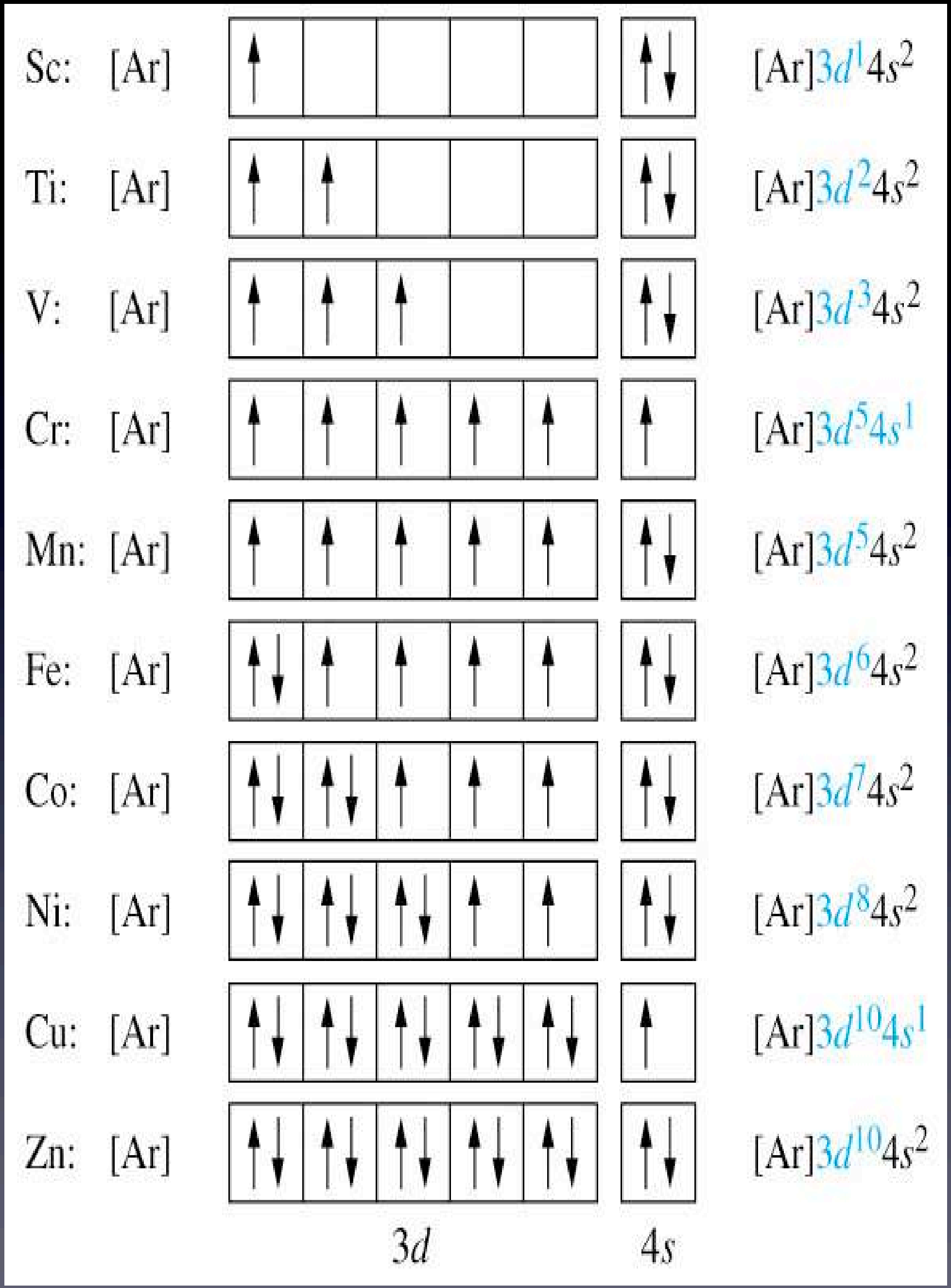
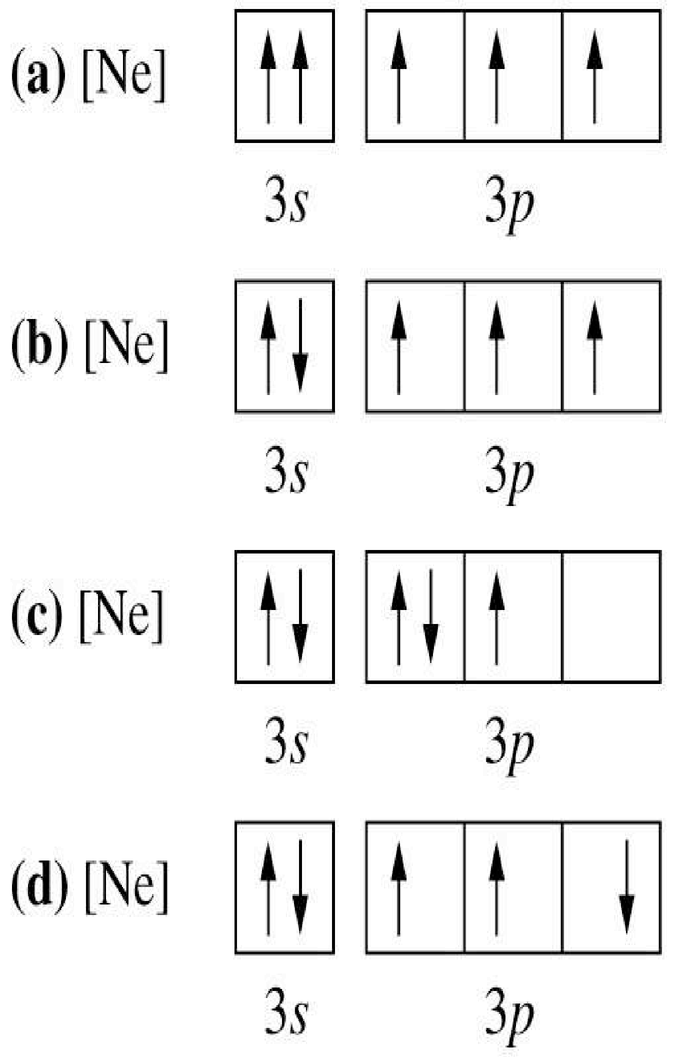


TABLE 9.2 Electron Configurations of Some Groups of Elements

Group	Element	Configuration
1	H	$1s^1$
	Li	$[\mathrm{He}]2s^1$
	Na	$[\mathrm{Ne}]3s^1$
	K	$[\mathrm{Ar}]4s^1$
	Rb	$[\mathrm{Kr}]5s^1$
	Cs	$[\mathrm{Xe}]6s^1$
	Fr	$[\mathrm{Rn}]7s^1$
17	F	$[\mathrm{He}]2s^2 2p^5$
	Cl	$[\mathrm{Ne}]3s^2 3p^5$
	Br	$[\mathrm{Ar}]3d^{10} 4s^2 4p^5$
	I	$[\mathrm{Kr}]4d^{10} 5s^2 5p^5$
	At	$[\mathrm{Xe}]4f^{14} 5d^{10} 6s^2 6p^5$
18	He	$1s^2$
	Ne	$[\mathrm{He}]2s^2 2p^6$
	Ar	$[\mathrm{Ne}]3s^2 3p^6$
	Kr	$[\mathrm{Ar}]3d^{10} 4s^2 4p^6$
	Xe	$[\mathrm{Kr}]4d^{10} 5s^2 5p^6$
	Rn	$[\mathrm{Xe}]4f^{14} 5d^{10} 6s^2 6p^6$



		3d					4s	
Sc:	[Ar]	↑					↑↓	[Ar]$3d^1 4s^2$
Ti:	[Ar]	↑	↑				↑↓	[Ar]$3d^2 4s^2$
V:	[Ar]	↑	↑	↑			↑↓	[Ar]$3d^3 4s^2$
Cr:	[Ar]	↑	↑	↑	↑	↑	↑	[Ar]$3d^5 4s^1$
Mn:	[Ar]	↑	↑	↑	↑	↑	↑↓	[Ar]$3d^5 4s^2$
Fe:	[Ar]	↑↓	↑	↑	↑	↑	↑↓	[Ar]$3d^6 4s^2$
Co:	[Ar]	↑↓	↑↓	↑	↑	↑	↑↓	[Ar]$3d^7 4s^2$
Ni:	[Ar]	↑↓	↑↓	↑↓	↑	↑	↑↓	[Ar]$3d^8 4s^2$
Cu:	[Ar]	↑↓	↑↓	↑↓	↑↓	↑↓	↑	[Ar]$3d^{10} 4s^1$
Zn:	[Ar]	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	[Ar]$3d^{10} 4s^2$



- which of these is correct?

example

- sample exercise 7.7 (326)

eocs

- 7.69-72
- 7.75, 76, 80

Periodic Trends

- using the PT can show you how atoms might act
- understand these basic rules and you’ll go far...

- **Electrons are attracted to the protons in the nucleus of an atom.**
  - the closer they are the more they are attracted
  - the more protons, the more attraction

- **Electrons are repelled by each other**
    - important b/c if there are outside electrons, they are prevented from being attracted by the inside electrons
    - called **shielding**

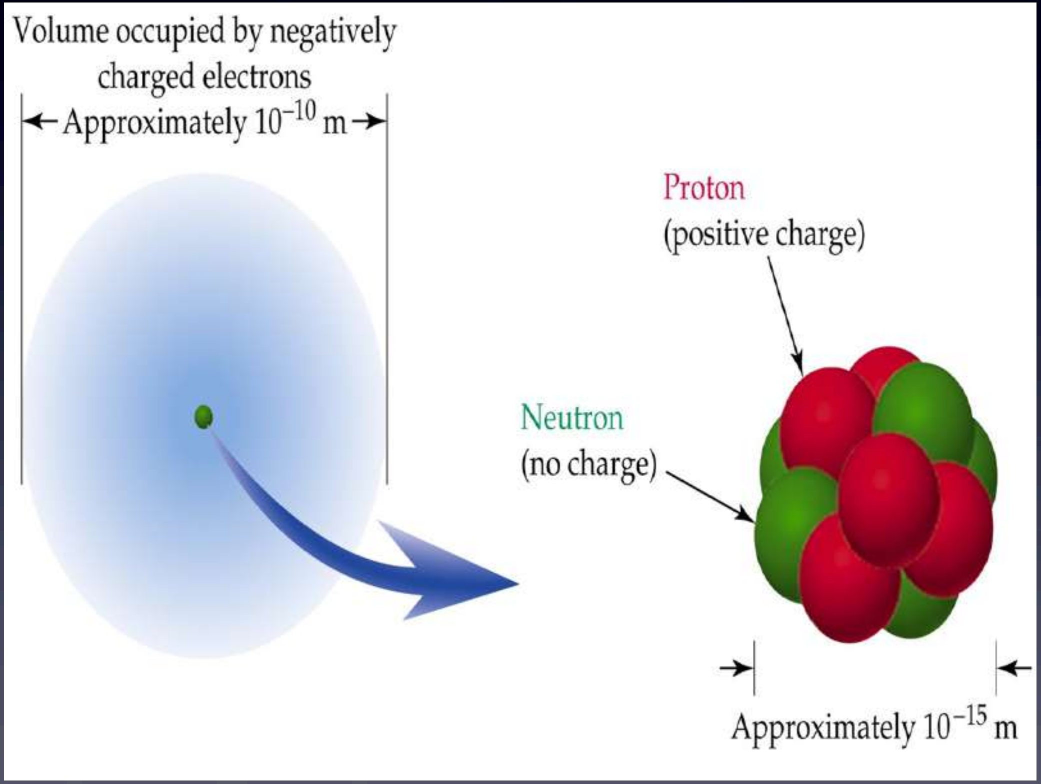
- **Completed shells are a good thing.**
- so if shells are incomplete they will lose or steal or share to get complete

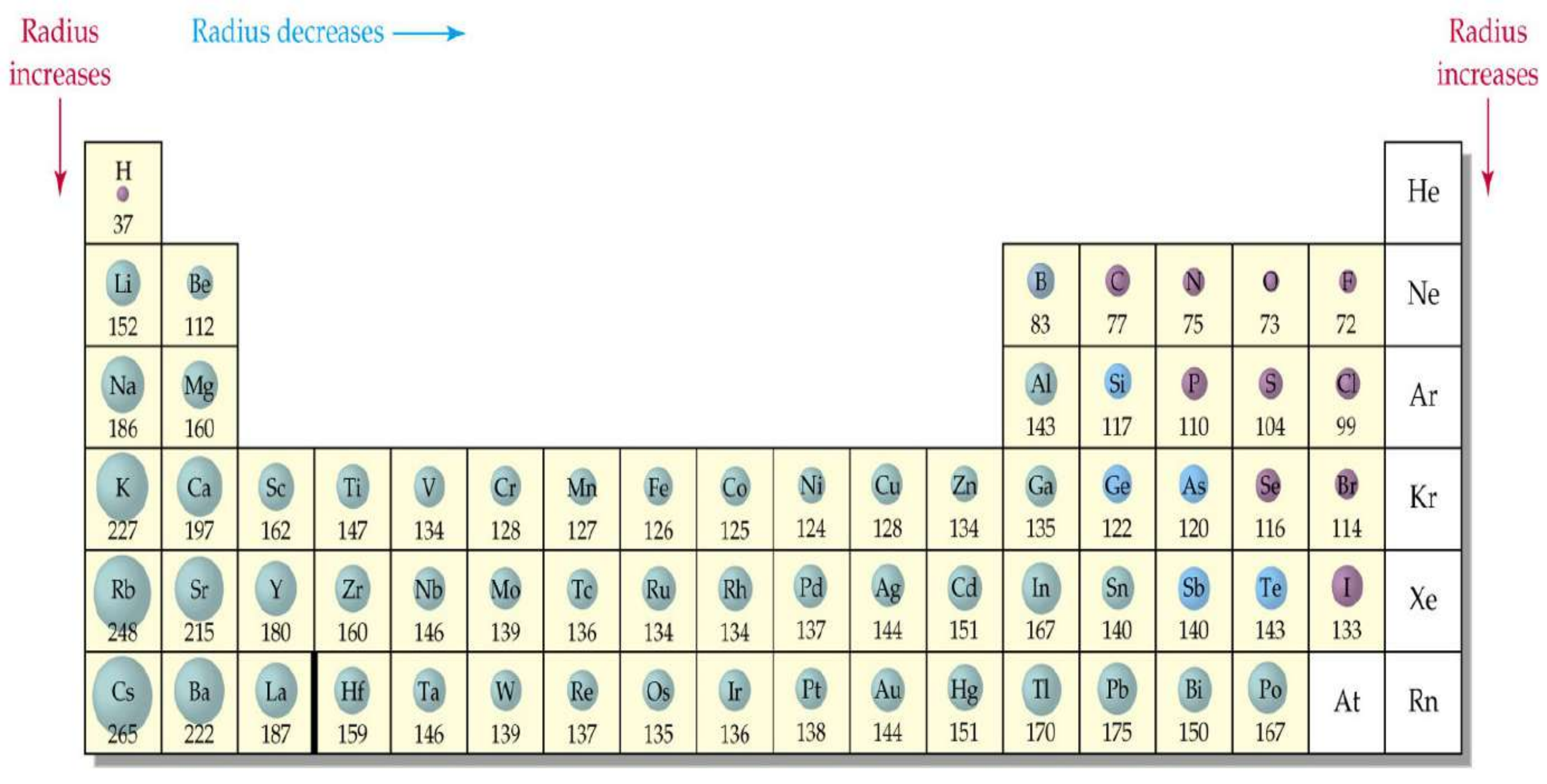
- **metals** are on the left side of the PT
  - they like to lose e-
  - they make up the majority of the PT

- **nonmetals** are on the right side
    - they tend to gain e- to form bonds
- **metalloids** are on the border b/t the two
    - they have properties of both metals and nonmetals

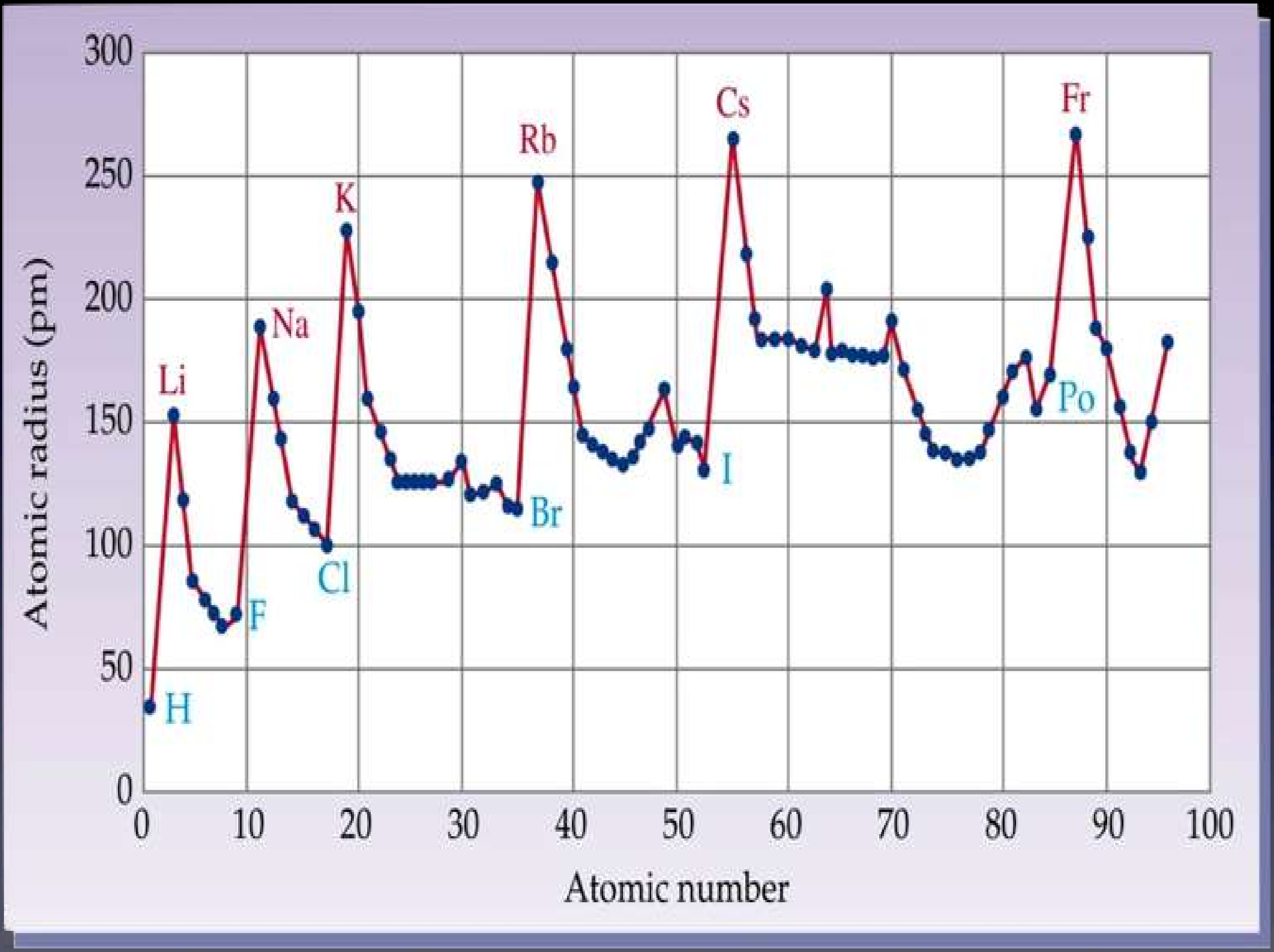
Atomic Radius

- the atomic radius is roughly the distance b/t nucleus and edge of cloud
- here are some things you should know...



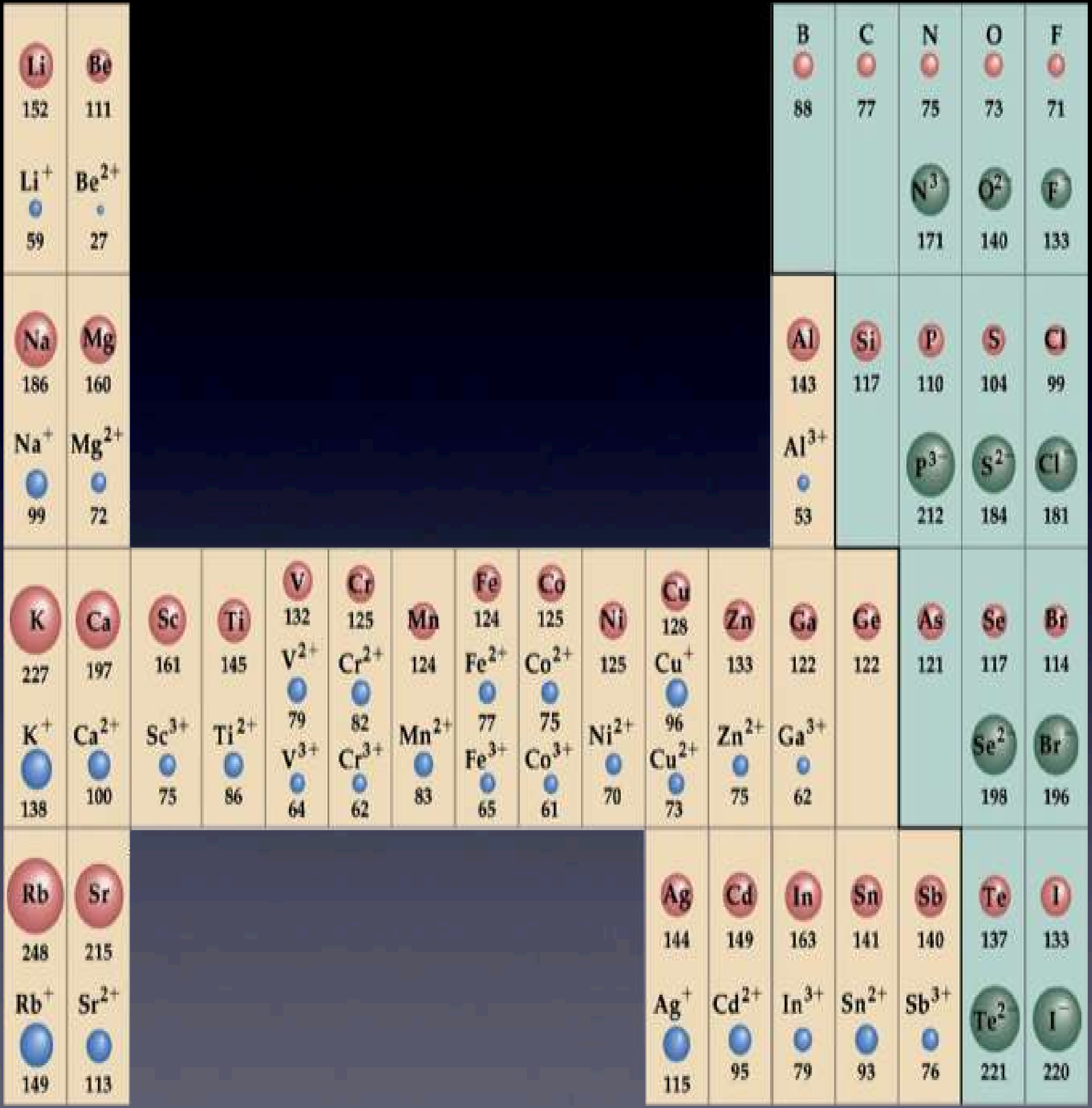


H 37																	He
Li 152	Be 112											B 83	C 77	N 75	O 73	F 72	Ne
Na 186	Mg 160											Al 143	Si 117	P 110	S 104	Cl 99	Ar
K 227	Ca 197	Sc 162	Ti 147	V 134	Cr 128	Mn 127	Fe 126	Co 125	Ni 124	Cu 128	Zn 134	Ga 135	Ge 122	As 120	Se 116	Br 114	Kr
Rb 248	Sr 215	Y 180	Zr 160	Nb 146	Mo 139	Tc 136	Ru 134	Rh 134	Pd 137	Ag 144	Cd 151	In 167	Sn 140	Sb 140	Te 143	I 133	Xe
Cs 265	Ba 222	La 187	Hf 159	Ta 146	W 139	Re 137	Os 135	Ir 136	Pt 138	Au 144	Hg 151	Tl 170	Pb 175	Bi 150	Po 167	At	Rn


Atomic radius (pm)
300
250
200
150
100
50
0
Atomic number
0
10
20
30
40
50
60
70
80
90
100
H
Li
F
Na
Cl
K
Br
Rb
I
Cs
Po
Fr

- **moving left to right we see the atoms getting smaller**
- more protons mean a greater pull
- and although more e-, they are being added to same shell, so not much shielding from nucleus

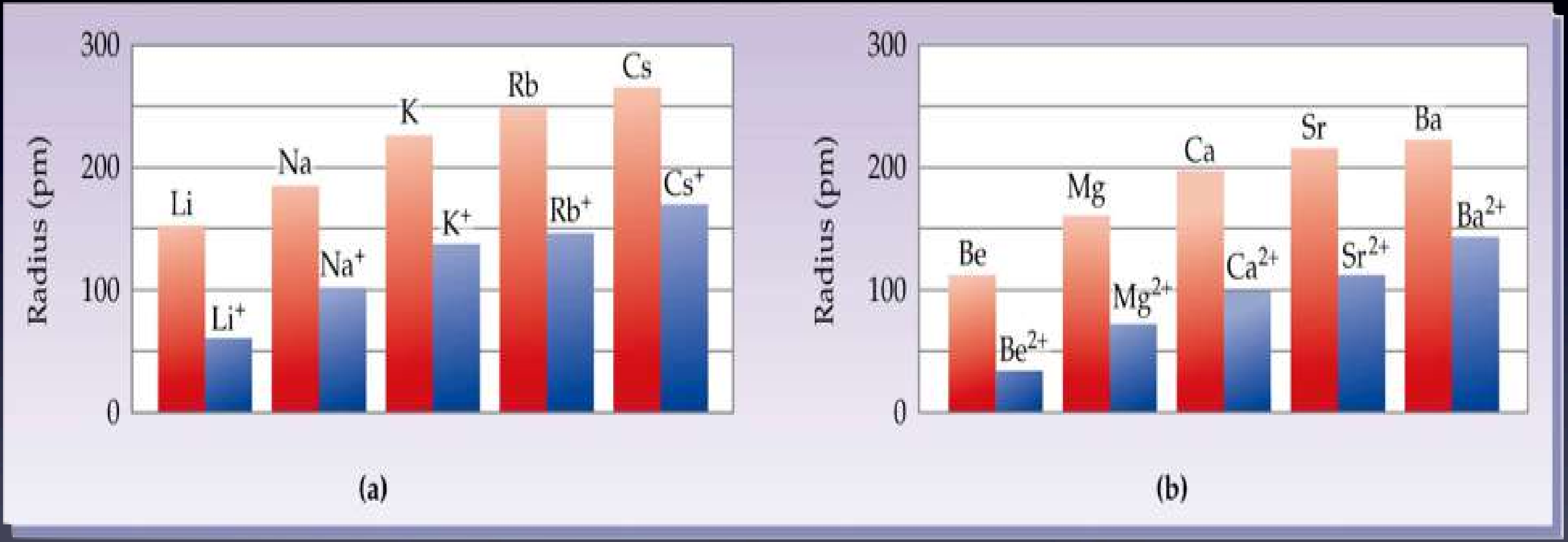
- **moving down a group, atoms get bigger**
- a whole new shell is being added here so shielding really kicks in
- the greater p number is not able to whoop shielding
- now look at the ion size relationship (after this film)...



Li 152 Li^+ 59	Be 111 Be^{2+} 27											B 88	C 77	N 75 N^{3-} 171	O 73 O^{2-} 140	F 71 F^- 133
Na 186 Na^+ 99	Mg 160 Mg^{2+} 72											Al 143 Al^{3+} 53	Si 117	P 110 P^{3-} 212	S 104 S^{2-} 184	Cl 99 Cl^- 181
K 227 K^+ 138	Ca 197 Ca^{2+} 100	Sc 161 Sc^{3+} 75	Ti 145 Ti^{2+} 86	V 132 V^{2+} 79 V^{3+} 64	Cr 125 Cr^{2+} 82 Cr^{3+} 62	Mn 124 Mn^{2+} 83	Fe 124 Fe^{2+} 77 Fe^{3+} 65	Co 125 Co^{2+} 75 Co^{3+} 61	Ni 125 Ni^{2+} 70	Cu 128 Cu^+ 96 Cu^{2+} 73	Zn 133 Zn^{2+} 75	Ga 122 Ga^{3+} 62	Ge 122	As 121	Se 117 Se^{2-} 198	Br 114 Br^- 196
Rb 248 Rb^+ 149	Sr 215 Sr^{2+} 113									Ag 144 Ag^+ 115	Cd 149 Cd^{2+} 95	In 163 In^{3+} 79	Sn 141 Sn^{2+} 93	Sb 140 Sb^{3+} 76	Te 137 Te^{2-} 221	I 133 I^- 220

- **cations are smaller than their atoms**
- when making positive ions, usually the whole valence shell falls off
- result is smaller

- **anions are larger than their atoms**
- adding e- increases the e-/e- repulsion so it gets bigger



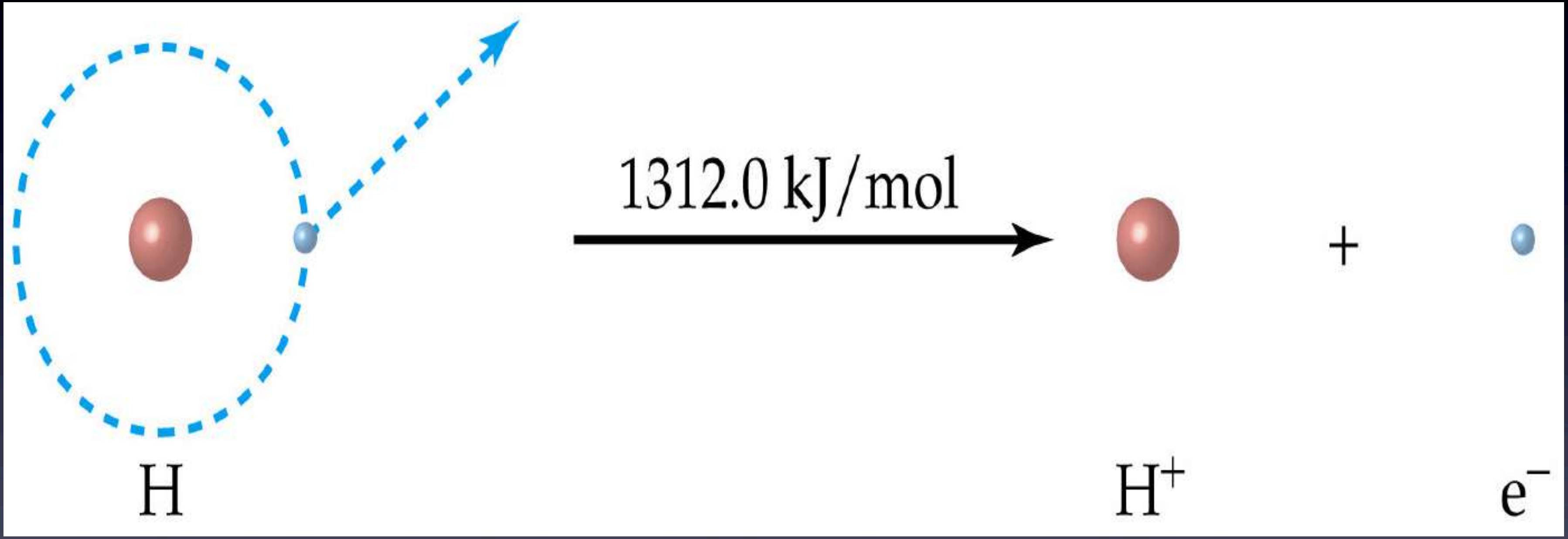
Radius (pm)
300
200
100
0
Li
Li⁺
Na
Na⁺
K
K⁺
Rb
Rb⁺
Cs
Cs⁺
(a)
Radius (pm)
300
200
100
0
Be
Be²⁺
Mg
Mg²⁺
Ca
Ca²⁺
Sr
Sr²⁺
Ba
Ba²⁺
(b)

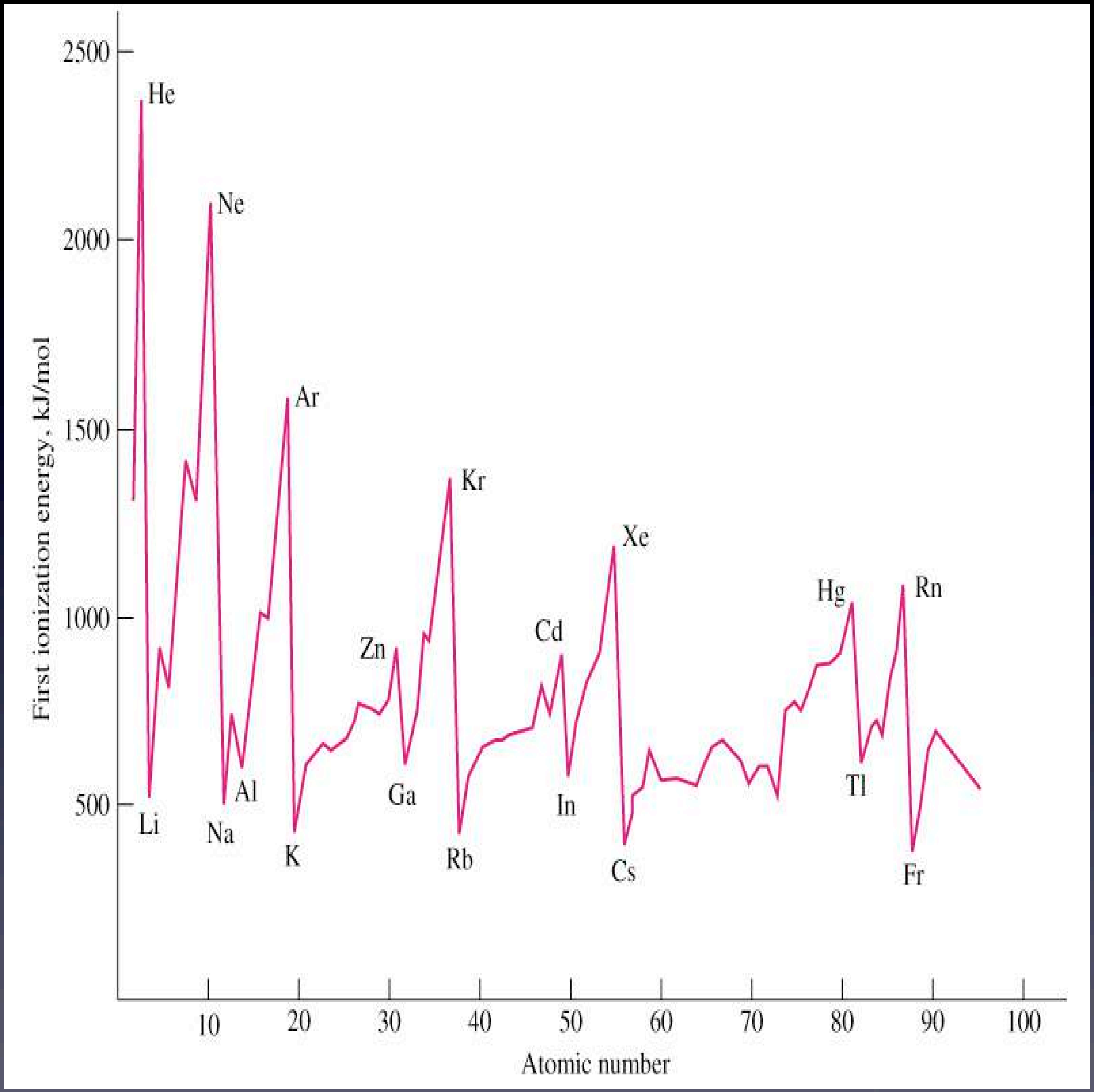
eocs

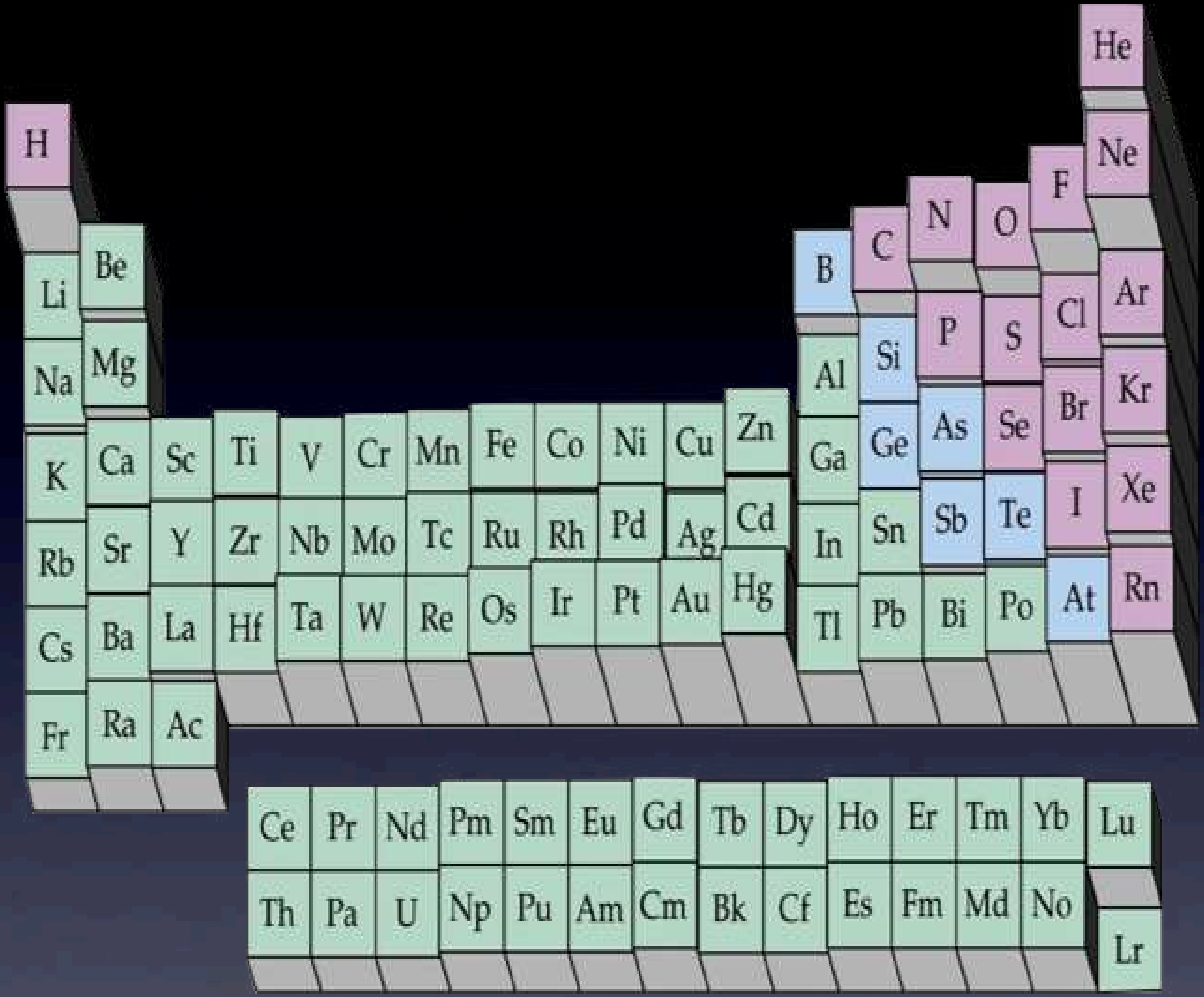
- 7.85, 86, 89, 90

Ionization Energy

- knocking an e- off will cost you
- the amt of E needed to knock off the most loosely held e- is called **first ionization energy**
- knocking another one off that new cation is called second ionization energy...

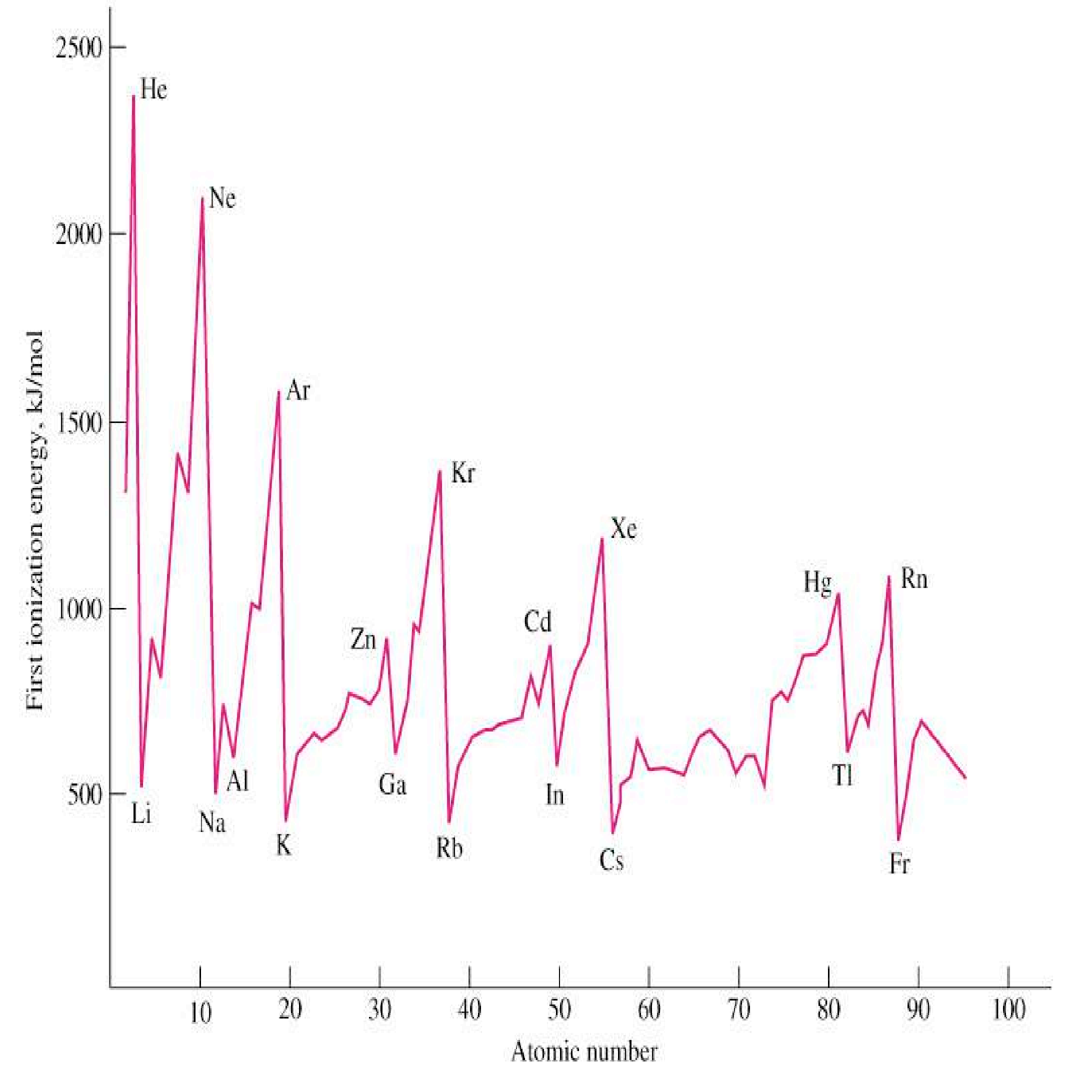

1312.0 kJ/mol
H
H⁺
+
e⁻


2500
2000
1500
1000
500
First ionization energy, kJ/mol
10
20
30
40
50
60
70
80
90
100
Atomic number
He
Ne
Ar
Kr
Xe
Rn
Hg
Cd
Zn
Li
Na
Al
K
Ga
Rb
In
Cs
Tl
Fr

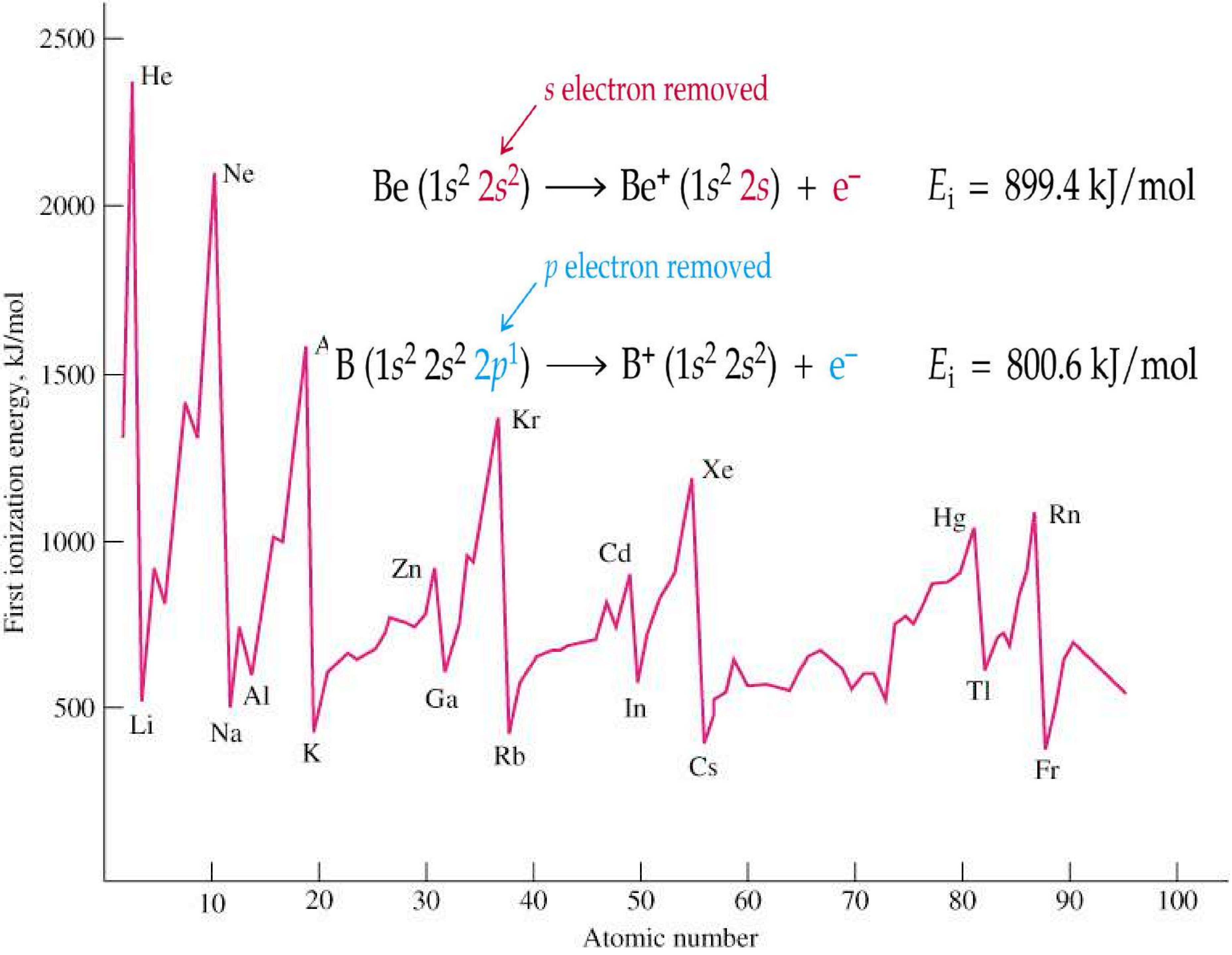


- this shows relative ionization energy
- which side doesn’t like losing electrons?

- **IE increases going left to right**
- as you go to the right, we get more protons
- more p’s hold on more tightly to the e-’s
- so it’s tougher to take them off
- so the IE increases



- why there are valleys along the way?



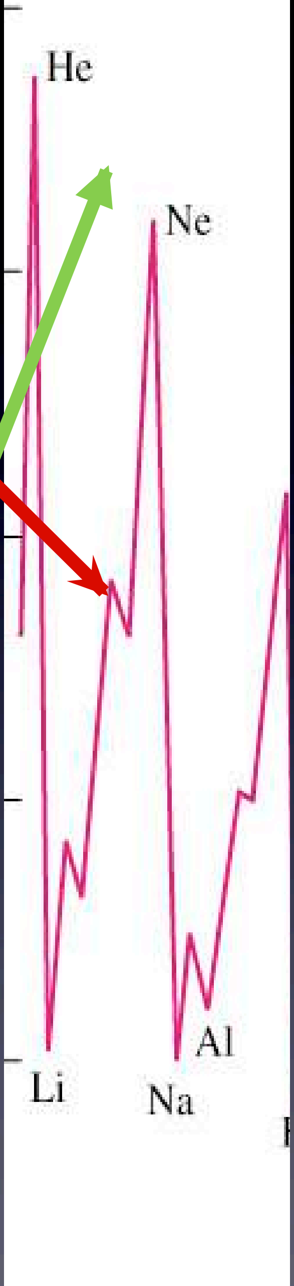
s electron removed

$$\mathrm{Be}\ (1s^2\ 2s^2) \longrightarrow \mathrm{Be}^+\ (1s^2\ 2s) + e^- \qquad E_\mathrm{i} = 899.4\ \mathrm{kJ/mol}$$

p electron removed

$$\mathrm{B}\ (1s^2\ 2s^2\ 2p^1) \longrightarrow \mathrm{B}^+\ (1s^2\ 2s^2) + e^- \qquad E_\mathrm{i} = 800.6\ \mathrm{kJ/mol}$$

- why the mini valley between O and N?
- N charges more b/c it is half-filled; O gets to be half-filled so charges less
- why the decrease from He to Ne?...



- **IE decreases going down a group**
- each row down adds a shell
- the shells shield the valence e-’s
- therefore easier to knock off
- (more protons? true, but not enough to cancel shielding)

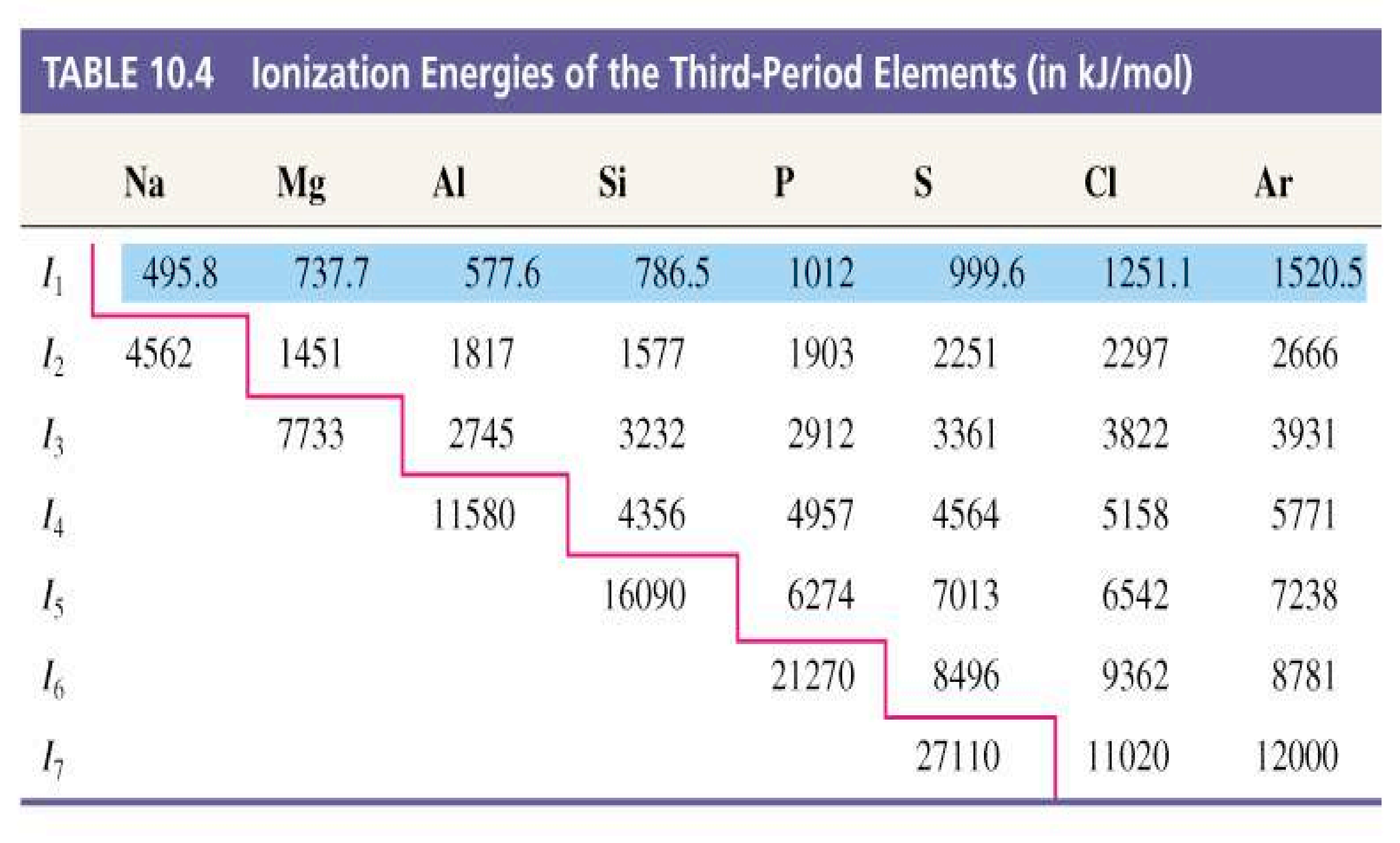


TABLE 10.4 Ionization Energies of the Third-Period Elements (in kJ/mol)

	Na	Mg	Al	Si	P	S	Cl	Ar
I_1	495.8	737.7	577.6	786.5	1012	999.6	1251.1	1520.5
I_2	4562	1451	1817	1577	1903	2251	2297	2666
I_3		7733	2745	3232	2912	3361	3822	3931
I_4			11580	4356	4957	4564	5158	5771
I_5				16090	6274	7013	6542	7238
I_6					21270	8496	9362	8781
I_7						27110	11020	12000

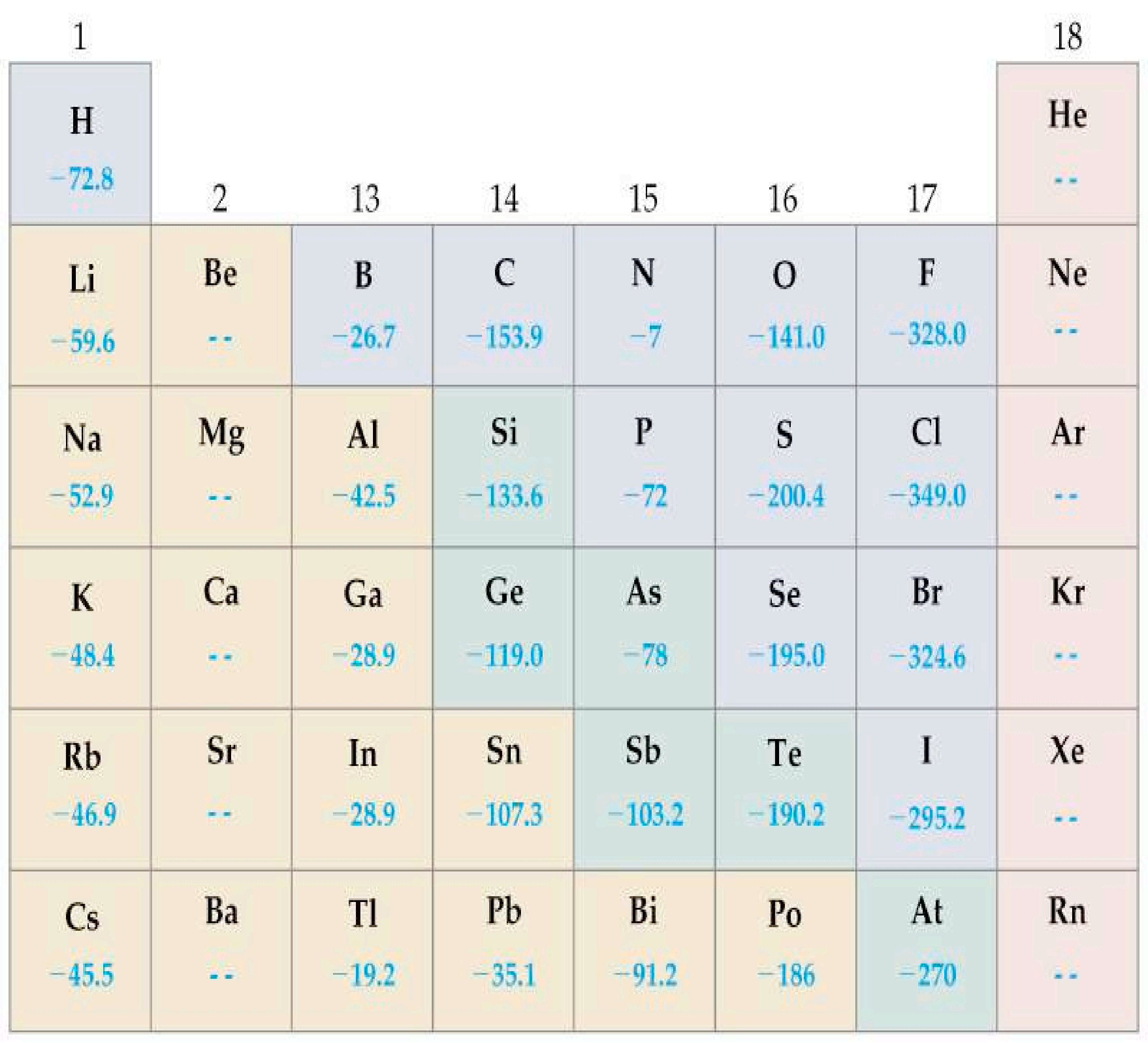
- see any patterns???

- **2nd IE is greater than the 1st IE**
- take an e- away and there is less e-/e- repulsion and the remaining e-’s will move in tighter
- and trying to take an e- from a cation is a pain

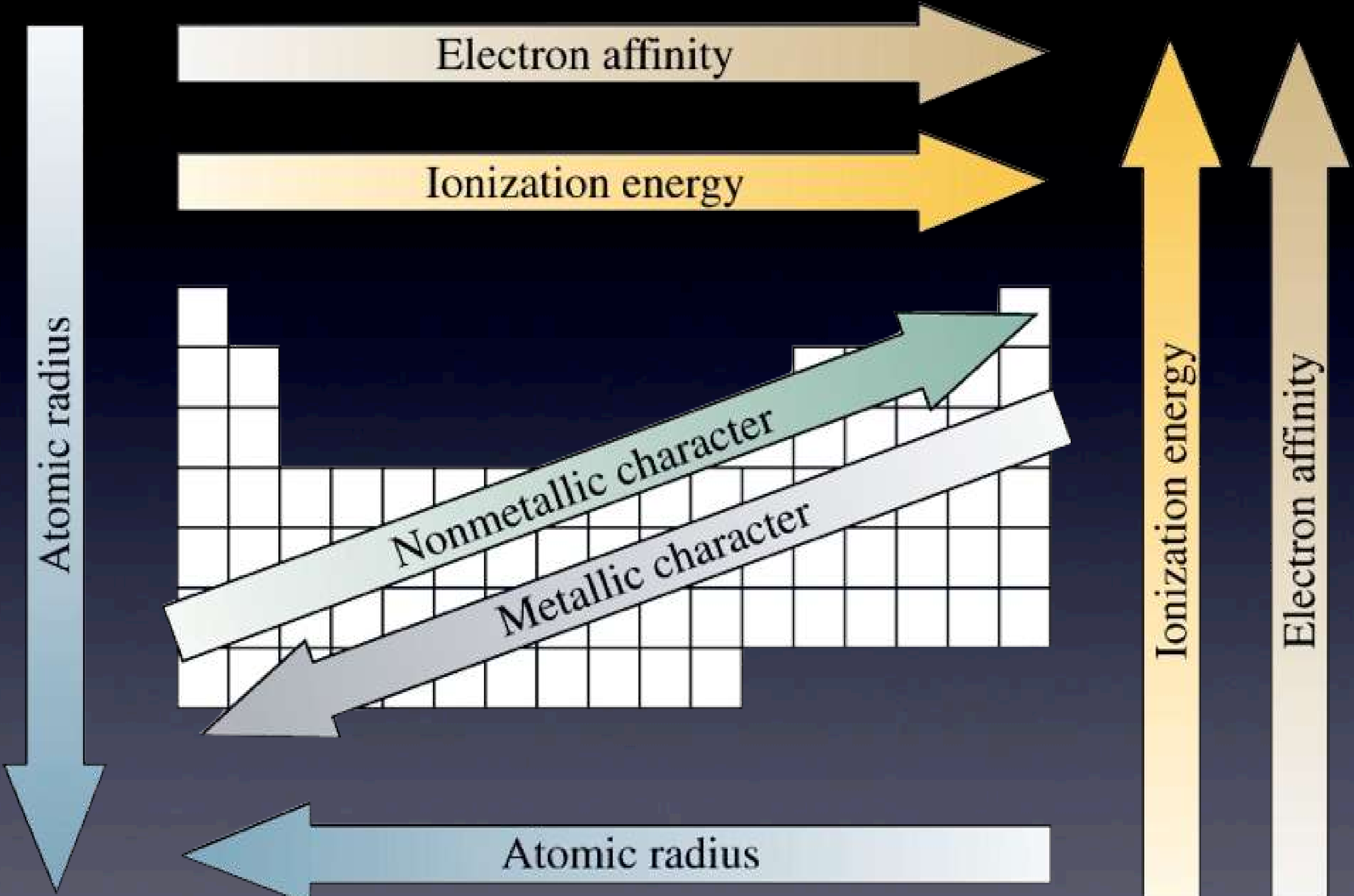
- remove enough e-s and ba-da-bing! suddenly it becomes almost impossible
- removing e-’s gets tougher until you lose a whole shell, then it ain’t happening
- for Na, the 1st EZ, 2nd fuggedaboutit
- for Mg, 1st and 2nd EZ, 3rd, no way
- etc.

Electron Affinity

- throw an e- onto an element
- if energy is given off, it is more stable
- if energy is needed to keep it there, it is less stable
- details not easy; just pretend it is the love an element has for electrons

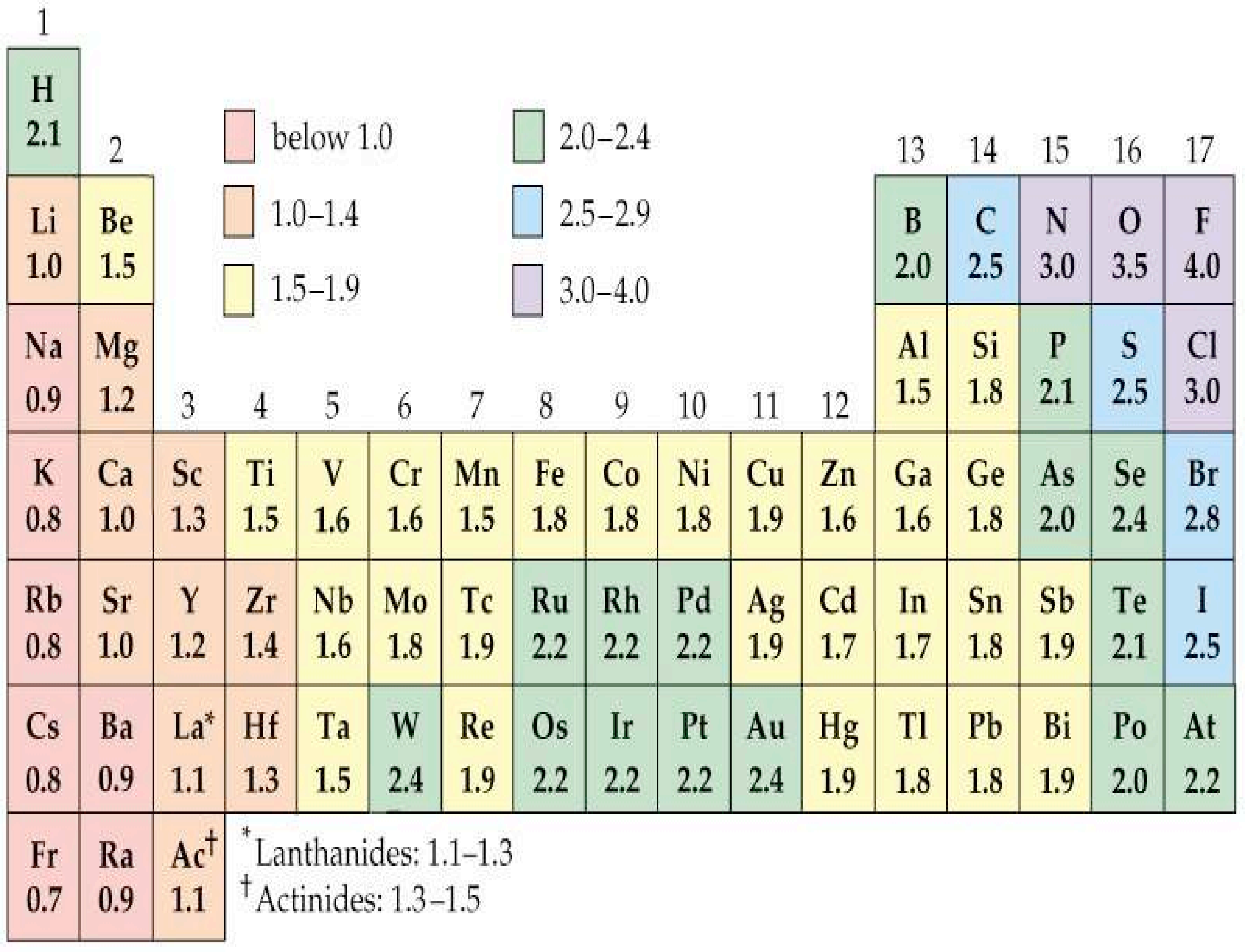


1	2	13	14	15	16	17	18
H −72.8							He - -
Li −59.6	Be - -	B −26.7	C −153.9	N −7	O −141.0	F −328.0	Ne - -
Na −52.9	Mg - -	Al −42.5	Si −133.6	P −72	S −200.4	Cl −349.0	Ar - -
K −48.4	Ca - -	Ga −28.9	Ge −119.0	As −78	Se −195.0	Br −324.6	Kr - -
Rb −46.9	Sr - -	In −28.9	Sn −107.3	Sb −103.2	Te −190.2	I −295.2	Xe - -
Cs −45.5	Ba - -	Tl −19.2	Pb −35.1	Bi −91.2	Po −186	At −270	Rn - -


Electron affinity
Ionization energy
Atomic radius
Nonmetallic character
Metallic character
Ionization energy
Electron affinity
Atomic radius

Electronegativity

- really important this one
- when atoms share electrons how badly does this guy want them? = en
- F rules here
- the closer you get to F the more en it is...

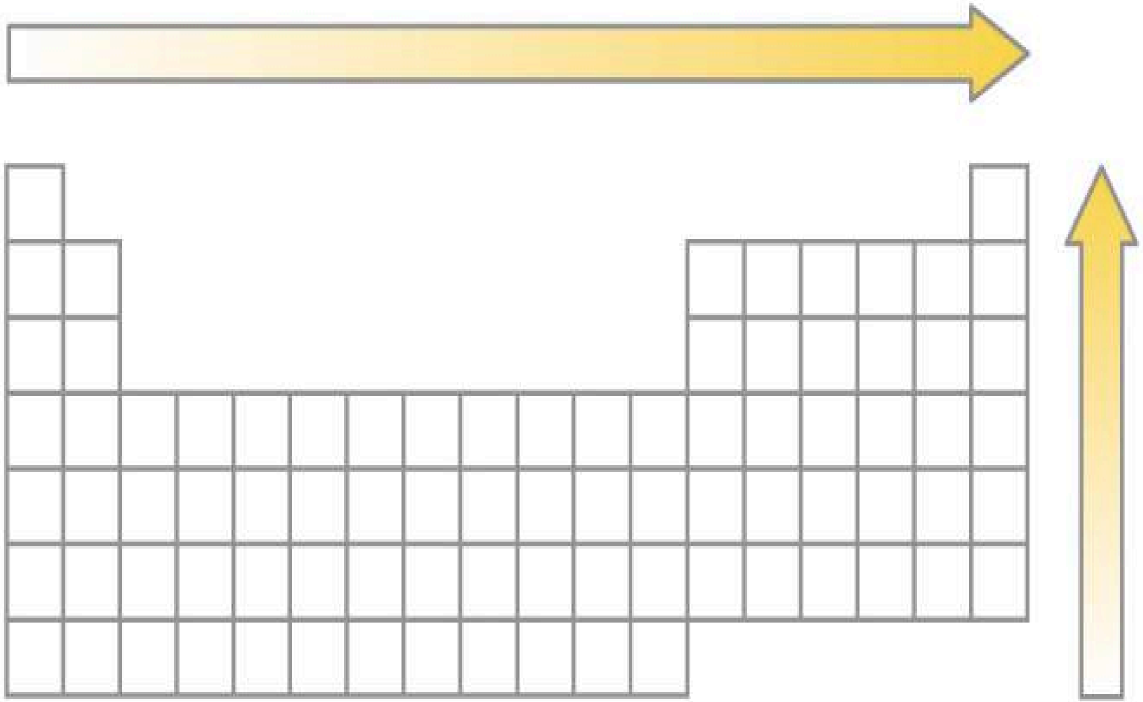


- 🟥 below 1.0
- 🟧 1.0–1.4
- 🟨 1.5–1.9
- 🟩 2.0–2.4
- 🟦 2.5–2.9
- 🟪 3.0–4.0

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17
H 2.1																
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.8	Ba 0.9	La* 1.1	Hf 1.3	Ta 1.5	W 2.4	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2
Fr 0.7	Ra 0.9	Ac† 1.1														

*Lanthanides: 1.1–1.3

†Actinides: 1.3–1.5



Increasing electronegativity

Increasing electro-negativity

eocs

- 7.87, 88, 121-123