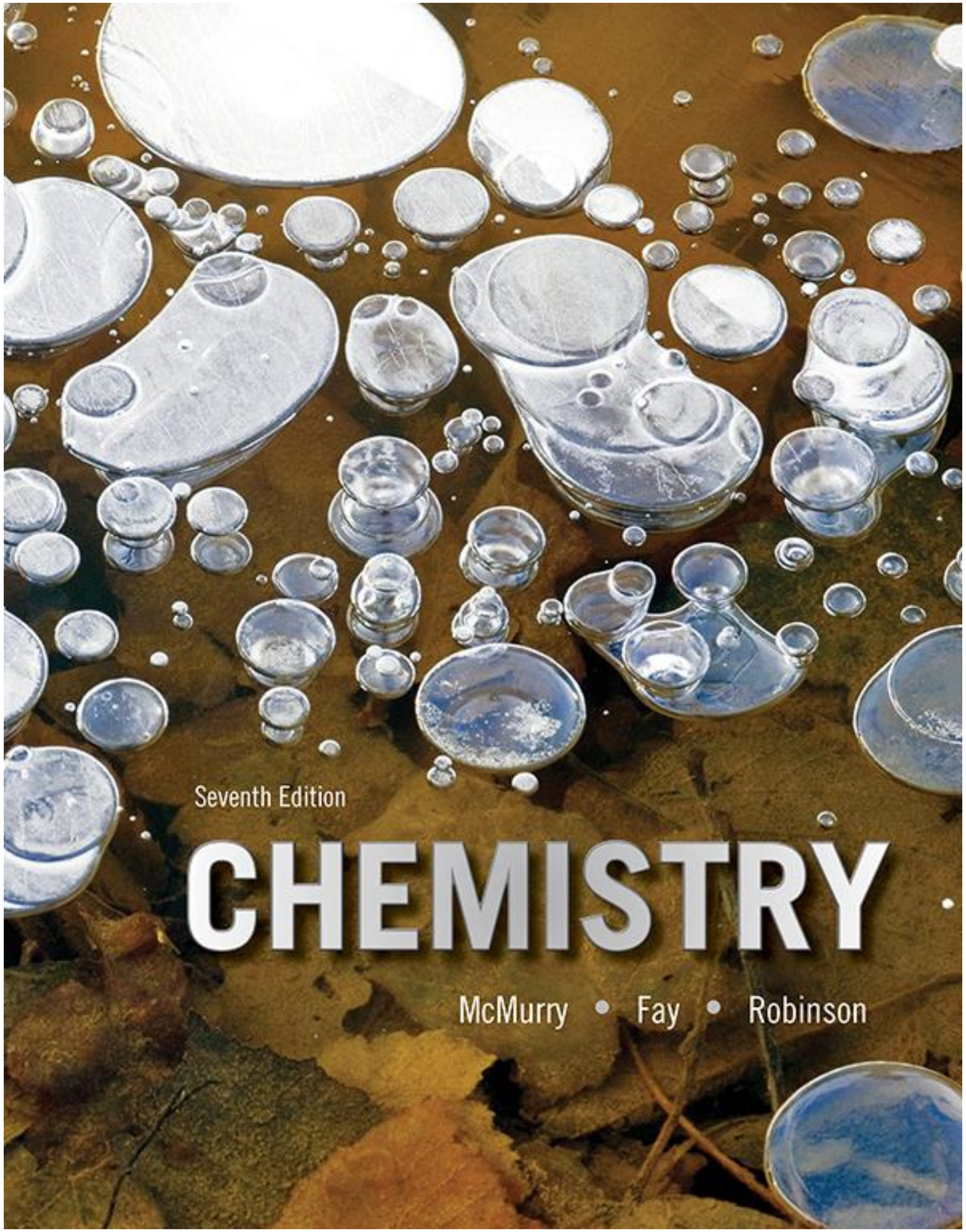


Lecture Presentation

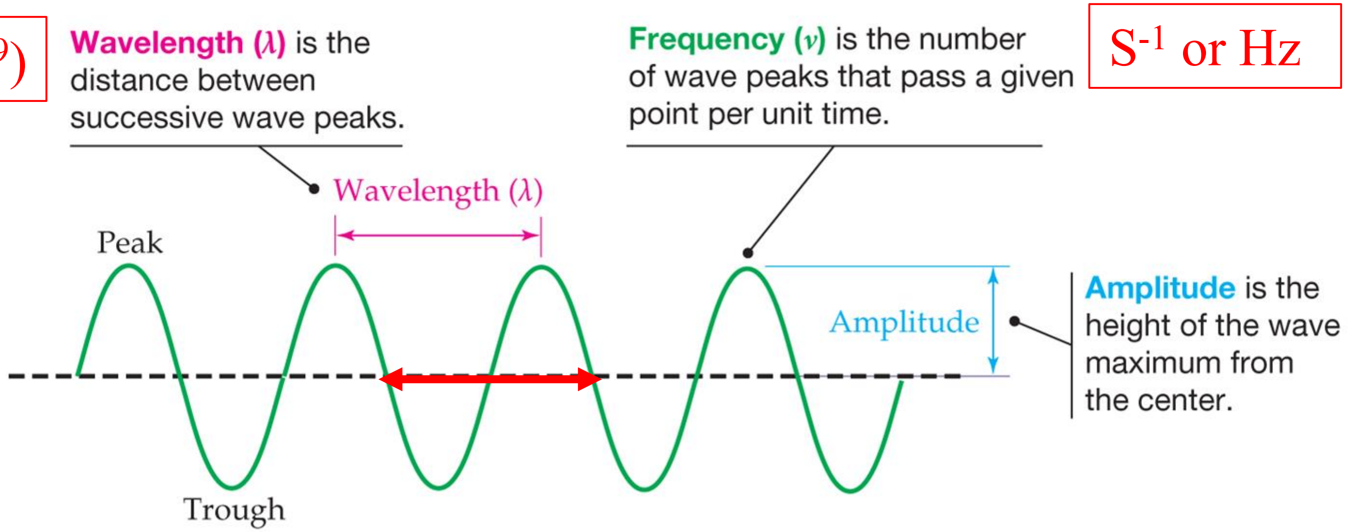
Chapter 5 Periodicity and the Electronic Structure of Atoms

HW: 5.1, 5.2, 5.3, 5.13, 5.15, 5.16, 5.17, 5.18, 5.26, 5.28, 5.36, 5.70, 5.72, 5.74, 5.82, 5.90, 5.94, 5.96, 5.106, 5.108

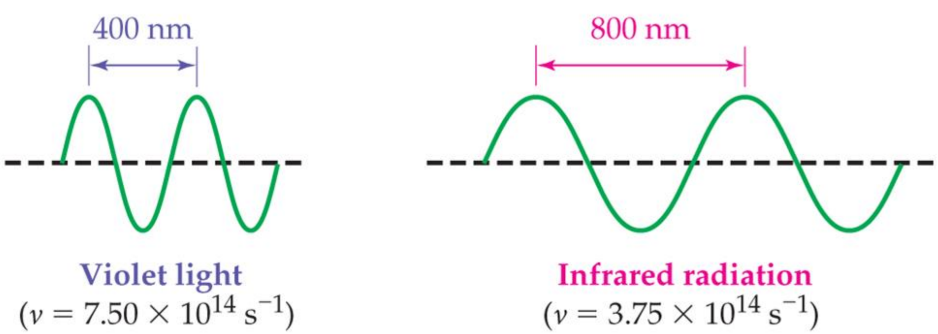
John E. McMurry
Robert C. Fay

The Nature of Radiant Energy and the Electromagnetic Spectrum (wave property)

(nm, 10^{-9})

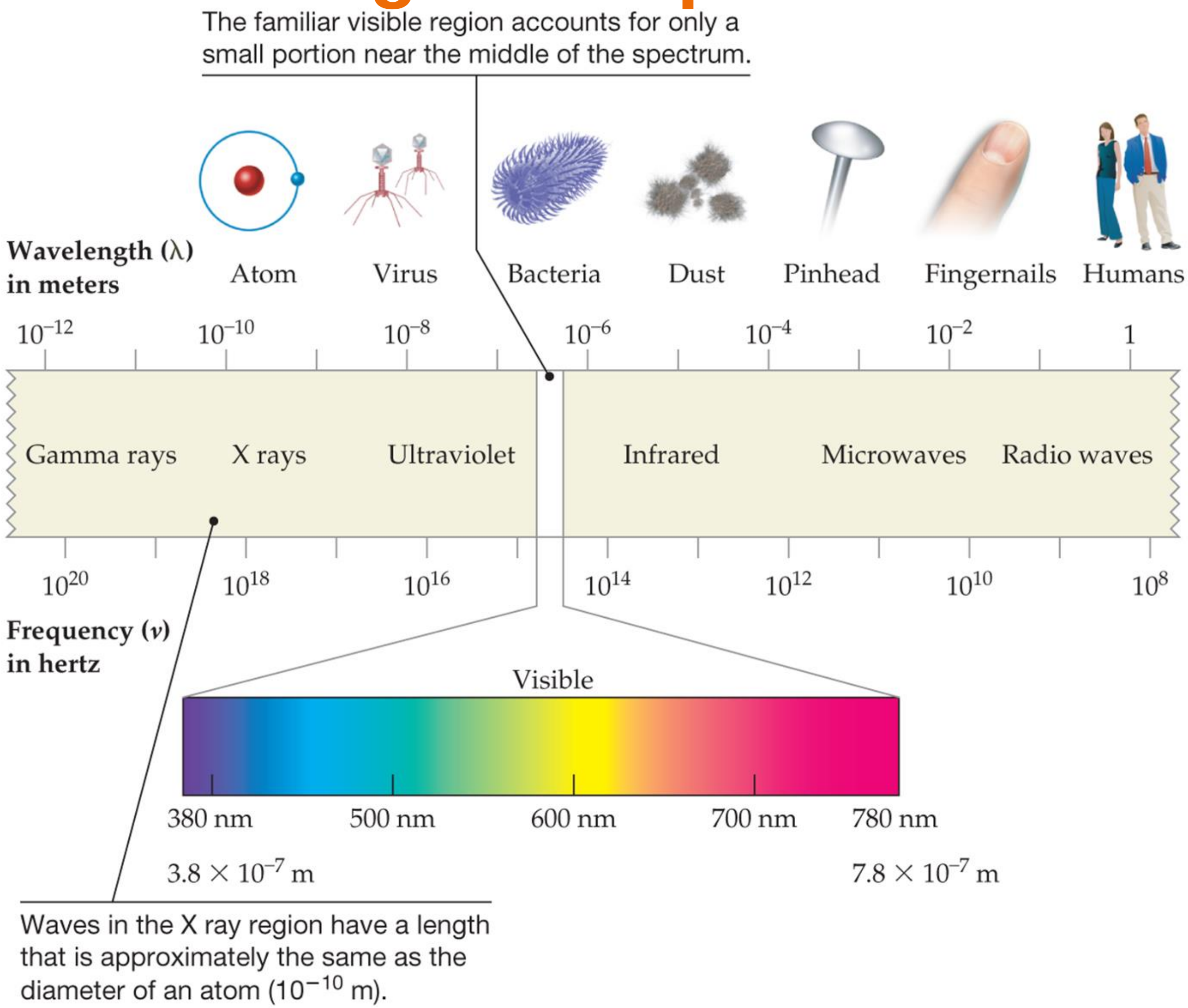


S^{-1} or Hz



higher frequency = lower wavelength = higher energy

The Nature of Radiant Energy and the Electromagnetic Spectrum



higher frequency = lower wavelength

$\nu \uparrow \quad \lambda \downarrow$

The Nature of Radiant Energy and the Electromagnetic Spectrum

Wavelength × Frequency = Speed

$$\lambda \times \nu = c$$

$$\text{m} \qquad \frac{1}{\text{s}} \text{ or Hz} \qquad \frac{\text{m}}{\text{s}}$$

c - speed of light (speed of all electromagnetic radiation), constant

$$c = 3.00 \times 10^{8}\ \frac{\text{m}}{\text{s}}$$

The Nature of Radiant Energy and the Electromagnetic Spectrum

The light blue glow given off by mercury streetlamps has a frequency of 6.88×10^{14} s^{-1} (or Hz). What is the wavelength in nanometers? ($\lambda \nu = c$)

10/18 Friday
G section

$$\lambda = \frac{c}{\nu} = \frac{\left(3.00 \times 10^{8} \frac{m}{s}\right)}{\left(6.88 \times 10^{14} \frac{1}{s}\right)}$$

$$= 4.36 \times 10^{-7} \text{ m}$$

$$= 4.36 \times 10^{-7} \text{ m} * \frac{1 \times 10^{9} \text{ nm}}{\text{m}}$$

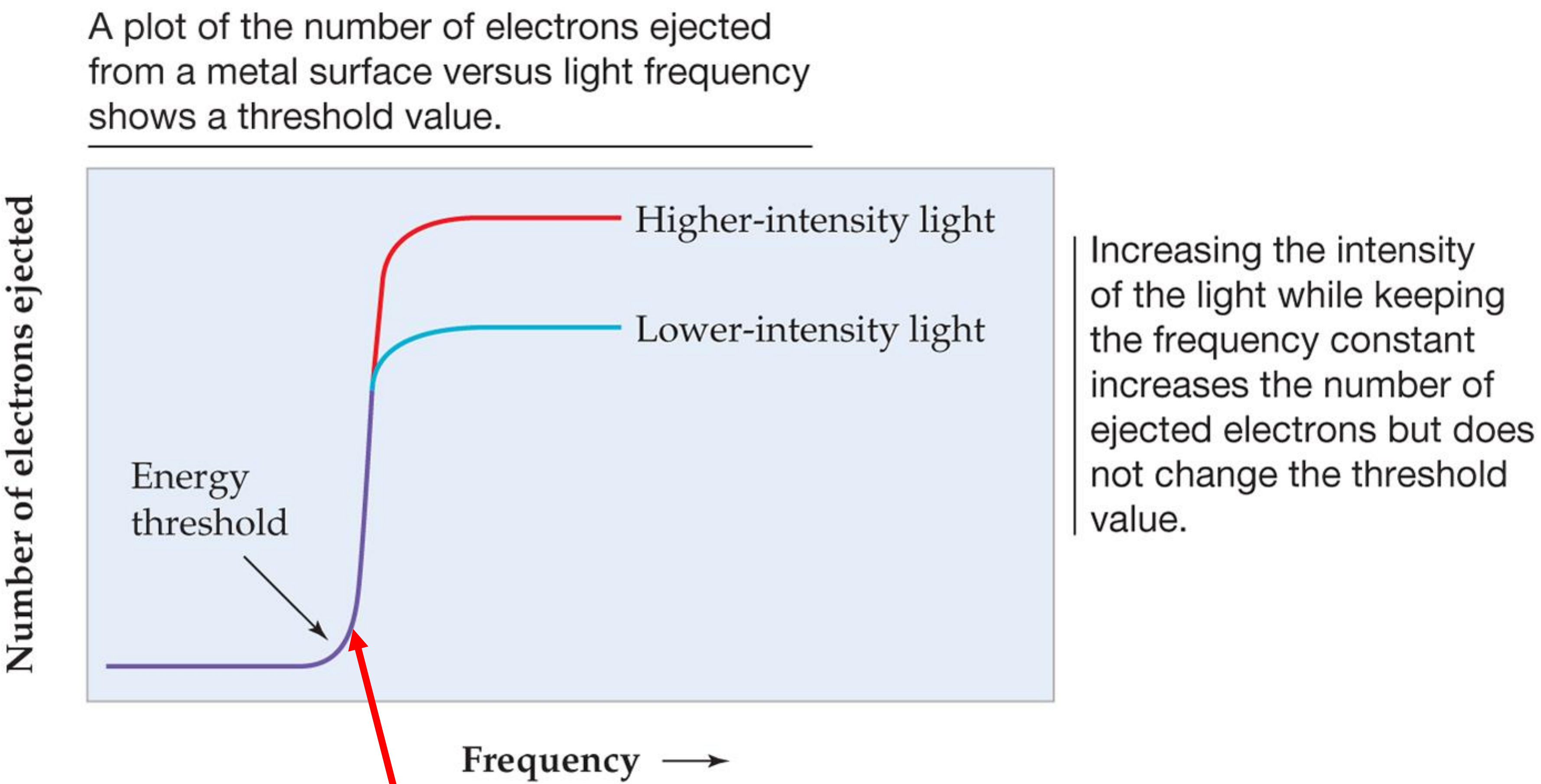
$$= 436 \text{ nm}$$

Particlelike Properties of Radiant Energy: The Photoelectric Effect and Planck's Postulate (particle property)

Photoelectric Effect: hit metal surface with light of certain energy kicks out electron from the metal. (frequency of the light used for photoelectric effect is different for each metal) (by analogy - picture shows ping pong balls don't break glass but baseball does)

End D section 10/21 Monday

Particlelike Properties of Radiant Energy: The Photoelectric Effect and Planck's Postulate



Frequency required to kick out electron

10/18 Friday F section

Particlelike Properties of Radiant Energy: The Photoelectric Effect and Planck's Postulate

$$E = h\nu = \frac{hc}{\lambda}$$

Energy of one photon

h (Planck's constant) = 6.626×10^{-34} J s

(above equation gives energy of one photon – so- multiply energy by Avogadro's # to get energy per mole)

higher energy = higher frequency (shorter wavelength)

$E \uparrow \quad \nu \uparrow$

Particlelike Properties of Radiant Energy: The Photoelectric Effect and Planck's Postulate

$$E = h\nu = \frac{hc}{\lambda}$$

h (Planck's constant) = 6.626×10^{-34} J s

(above equation gives energy of one photon – so- multiply energy by Avogadro's # to get energy per mole)

Example: What is the energy in J/mole for $\nu = 3.35 \times 10^{8}$ Hz (radar waves) ? (given frequency or wavelength – can get energy)

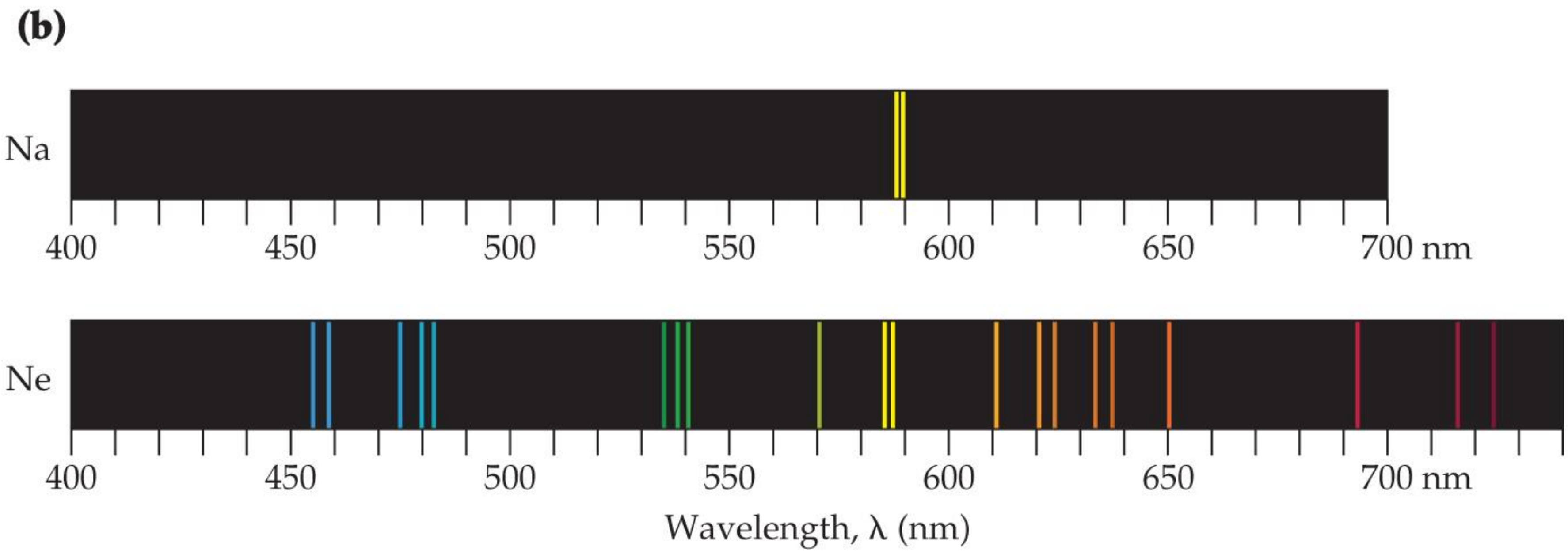
$E = (6.626 \times 10^{-34}\ \text{J s})(3.35 \times 10^{8}/\text{s}) = 2.22 \times 10^{-25}$ J for one photon

energy for one mole of photons

$= 2.22 \times 10^{-25}$ J/photon * 6.022×10^{23} photon/mol = 0.134 J/mol

The Interaction of Radiant Energy with Atoms: Line Spectra – hit atom with energy, get line spectrum

Line Spectrum: A series of discrete lines on an otherwise dark background as a result of light emitted by an excited atom



The Quantum Mechanical Model of the Atom: Orbitals and Quantum Numbers

Wave equation —Solve→ Wave function or orbital (Ψ) ⟶ Probability of finding electron in a region of space (Ψ^2)

A wave function is characterized by three parameters called **quantum numbers**: n, l, and m_l.

Quantum numbers define where electron is located with some probability. (because of Heisenburg Uncertainty Principal – can’t tell exactly the electron’s position & electron’s velocity, electron is a blur)

Electron can move to lower energy level inside the atom (& release light) OR can move to higher energy level when the electron is hit with light (& absorb light)

The Quantum Mechanical Model of the Atom: Orbitals and Quantum Numbers

Principal Quantum Number (*n*)

- Describes the size and energy level of the orbital
- ***shell (period #)***
- Positive integer (n = 1, 2, 3, 4, …)
- *larger n*
  - higher energy
  - larger distance of the e^- from the nucleus

The Quantum Mechanical Model of the Atom: Orbitals and Quantum Numbers

Angular-Momentum Quantum Number (*l*)
{for n, *l* = 0 to (n-1)}

- Defines the three-dimensional shape of the orbital
- **<u>subshell (s,p,d,f regions of the periodic table)</u>**
  - If $n = 1$, then $l = 0$.
  - If $n = 2$, then $l = 0$ or 1. (n-1 = 1)
  - If $n = 3$, then $l = 0, 1,$ or 2 (n-1 = 2)
- nickname letters are (subshell notation)
  - $l = 0$ *s* (sharp) (memorize spdf)
  - $l = 1$ *p* (principal) (do not memorize
  - $l = 2$ *d* (diffuse) s= sharp, etc.)
  - $l = 3$ *f* (fundamental)

The Quantum Mechanical Model of the Atom: Orbitals and Quantum Numbers

Magnetic Quantum Number (m_l)

- Defines the spatial orientation of the orbital
- For *l*, m_l = -*l* …. 0 ….. +*l*.
- **orbitals within subshell** (holds 2 electrons maximum in each orbital)

- If *l* = 1, then m_l = –1, 0, or 1.

(p subshell has 3 orbitals – p_x, p_y, p_z)

- If *l* = 2, then m_l = –2, –1, 0, 1, or 2.

(d subshell has 5 orbitals – d_{xy}, d_{xz}, d_{yz}, d_{z^2}, d_{x2-y2})

The Quantum Mechanical Model of the Atom: Orbitals and Quantum Numbers

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	

The Quantum Mechanical Model of the Atom: Orbitals and Quantum Numbers

Example: Identify the possible values for each of the three quantum numbers for a 4*p* orbital.
(4 is n, p is *l*, so question is asking what are m_l values given n & *l*)

The Quantum Mechanical Model of the Atom: Orbitals and Quantum Numbers

Example: Identify the possible values for each of the three quantum numbers for a 4*p* orbital.

$n = 4$ $l = 1$ (p)

For $l = 1$, what are possible m_l values ? ($-l \ldots 0 \ldots +l$)

$m_l = -1, 0,$ or 1 (3 orbitals within the p subshell within the n=4 shell)

HW: The Quantum Mechanical Model of the Atom: Orbitals and Quantum Numbers

For n = 3 (principal quantum #), what are all possible angular momentum quantum # (l values){ l = 0 to (n-1)}

____

For l = 2, what are all possible magnetic quantum # (m_l) values {$m_l = -l, \ldots 0 \ldots +l$}

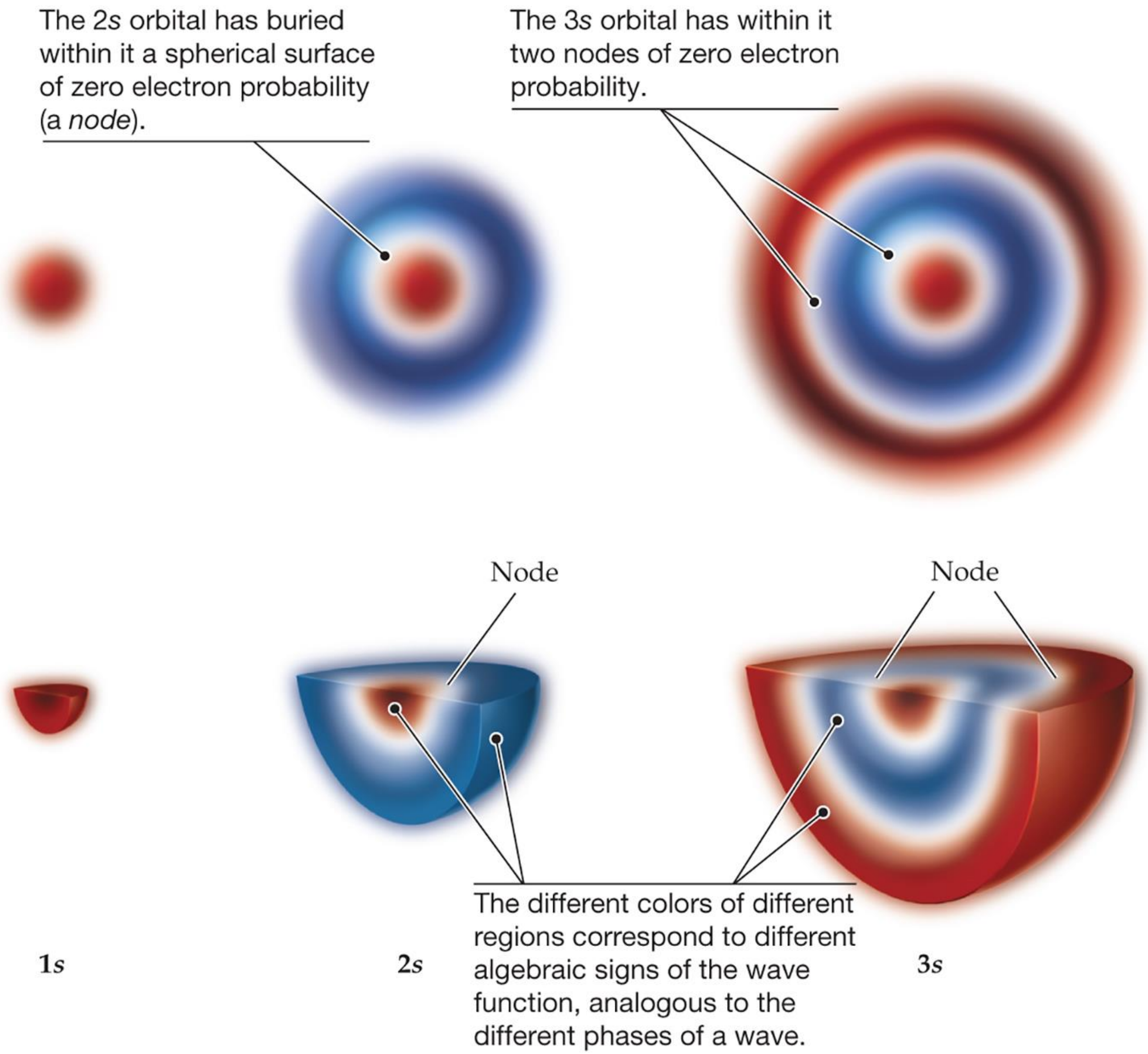
End F,G section 10/21Monday

____

For l = 0, 1, 2, 3 what is the letter nickname designation (question typo F,G class NOT m_l but l)

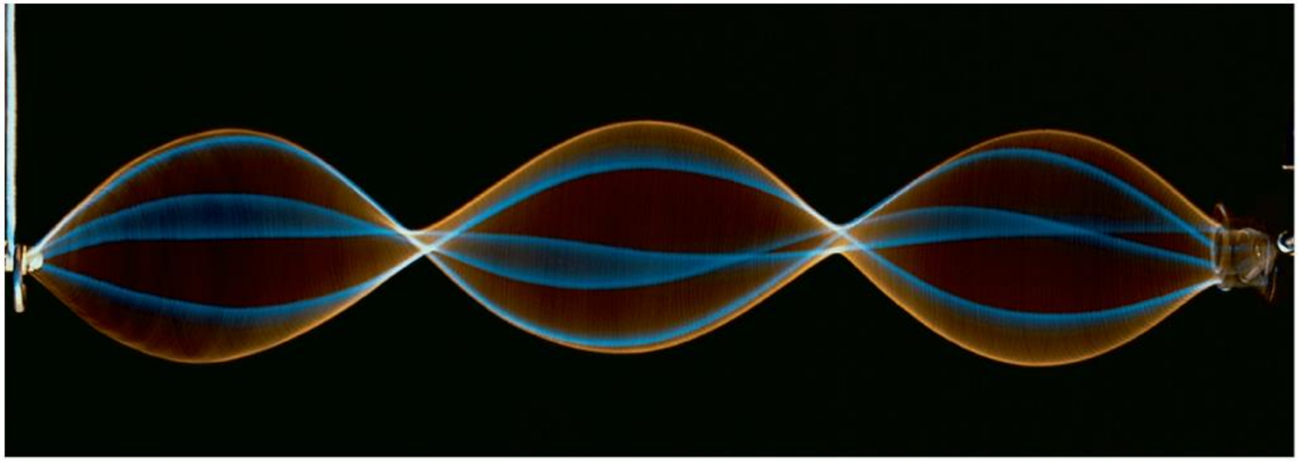
0 = _____ 1 = _____ 2 = _______ 3 = ________

The Shapes of Orbitals – s orbitals are all spherical (2s has one node, 3s has 2 nodes)

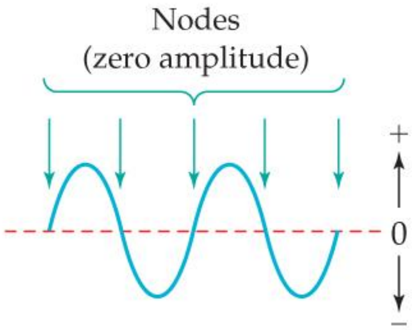


Node: A surface of zero probability for electron

The Shapes of Orbitals



When a rope is fixed at one end and vibrated rapidly at the other, a standing wave is generated.

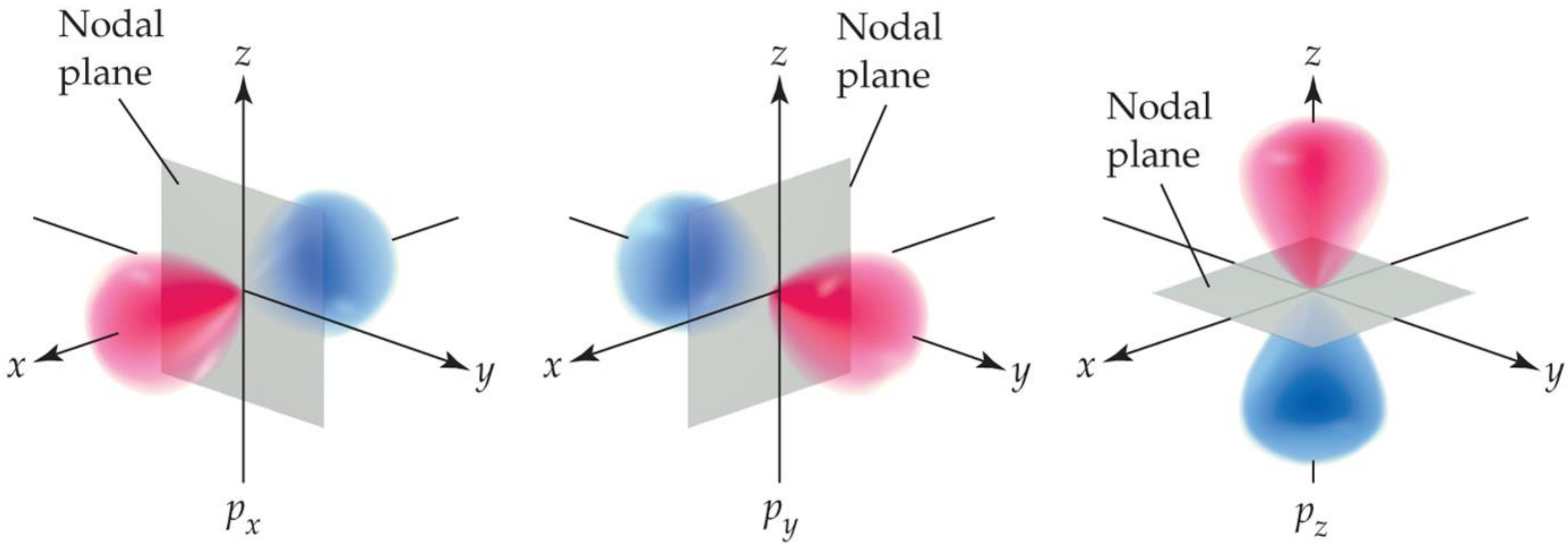


The wave has two phases with different algebraic signs, + and −, separated by zero-amplitude regions, called *nodes*.

Node: A surface of zero probability for finding the electron

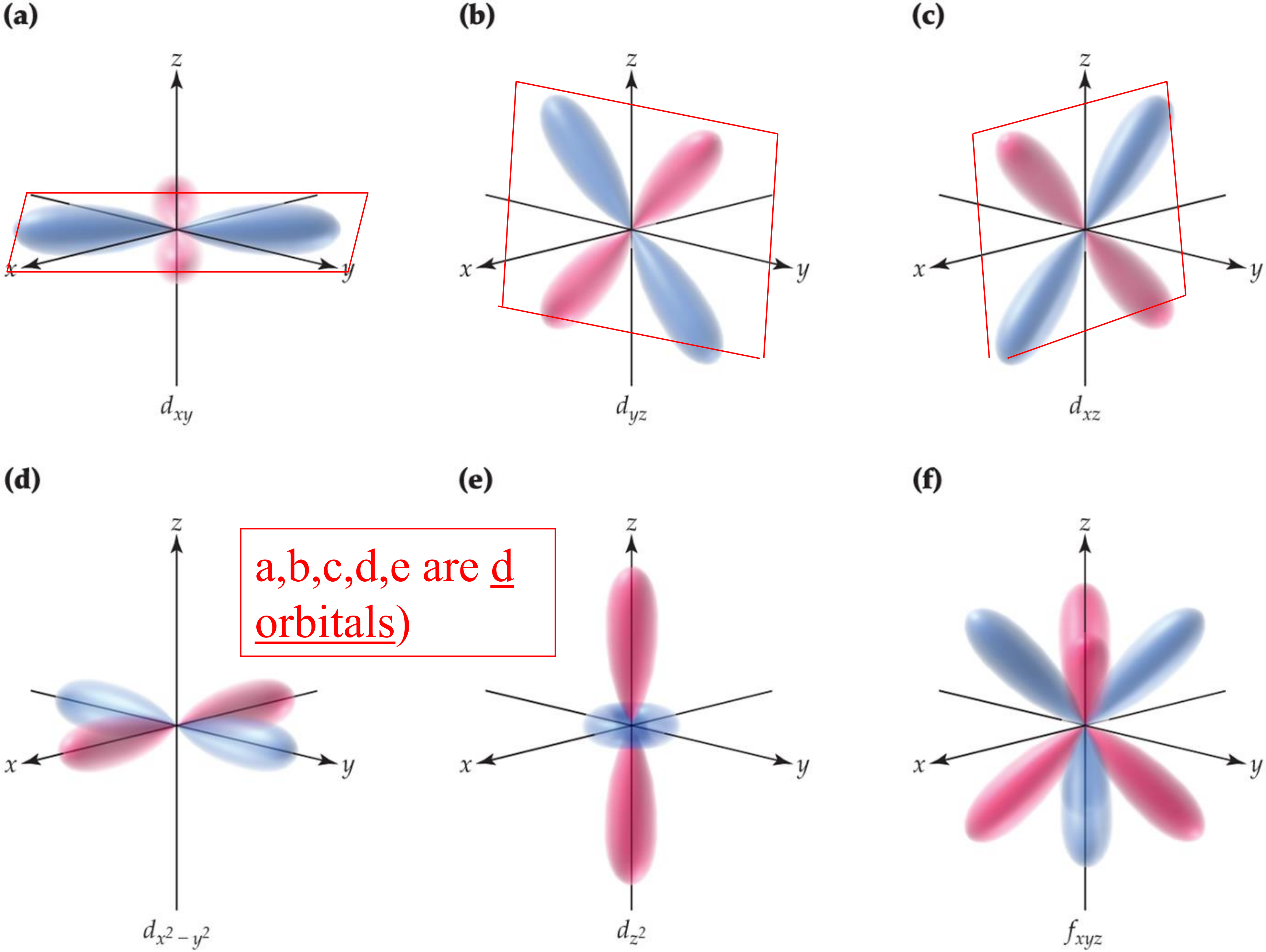
The Shapes of Orbitals – p orbital looks like dumbell

Each *p* orbital has two lobes of high electron probability separated by a nodal plane passing through the nucleus.



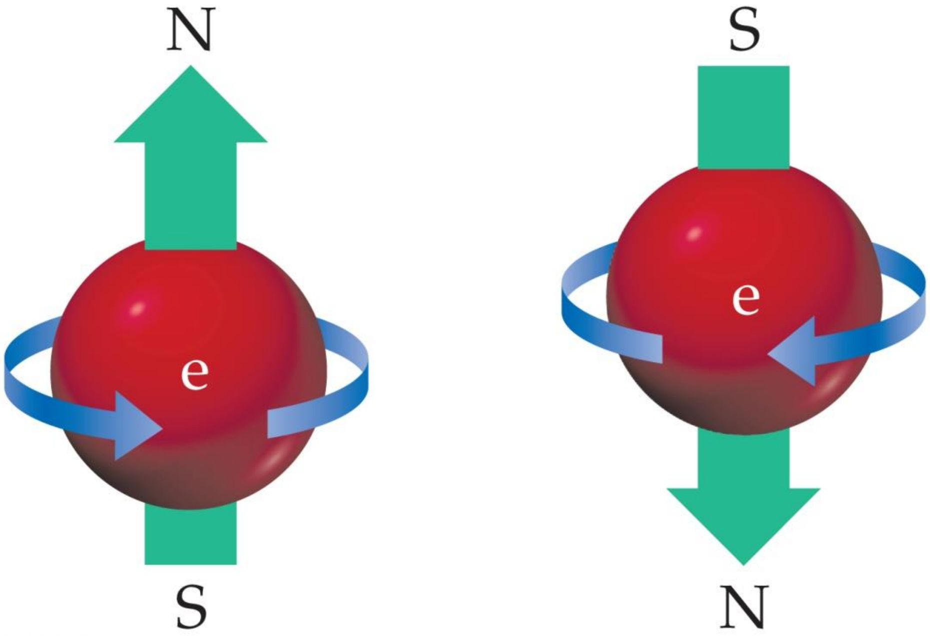
The **different colors** of the lobes represent different algebraic signs, analogous to the different phases of a wave.

cartesian coordinates – x,y,z


(a)
(b)
(c)
(d)
(e)
(f)
z
x
y
d_{xy}
d_{yz}
d_{xz}
$d_{x^2-y^2}$
d_{z^2}
f_{xyz}
a,b,c,d,e are d orbitals)

Electron Spin and the Pauli Exclusion Principle

Electrons have spin - results in tiny magnetic field & **spin quantum number (*m_s*)**. (+ ½ & - ½)

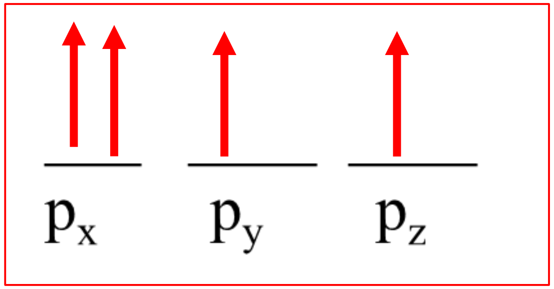


Pauli Exclusion Principle: No two electrons in an atom can have the same four quantum numbers.

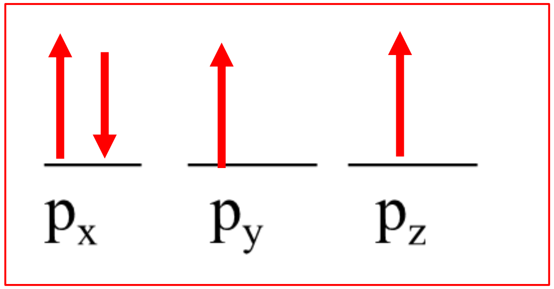
Electron Spin and the Pauli Exclusion Principle

Pauli Exclusion Principle: No two electrons in an atom can have the same four quantum numbers. (electron with 4 names – each e in atom has unique name)

Which figure violates Pauli ?

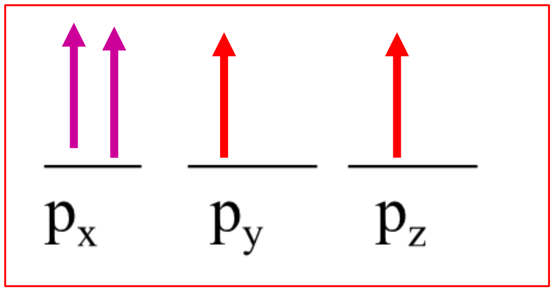


Let's say these are both 2p



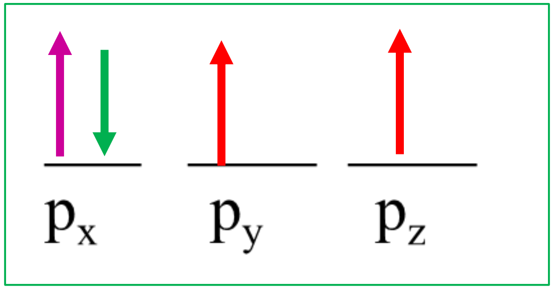
Electron Spin and the Pauli Exclusion Principle

Pauli Exclusion Principle: No two electrons in an atom can have the same four quantum numbers. (electron with 4 names – each e in atom has unique name) (note: p_x is not defined to always be -1)



bad – violates Pauli

These 2 electrons both have quantum number $n = 2$, $l = 1$, $m_l = -1$, $m_s = +\frac{1}{2}$



good

These 2 electrons have different quantum number

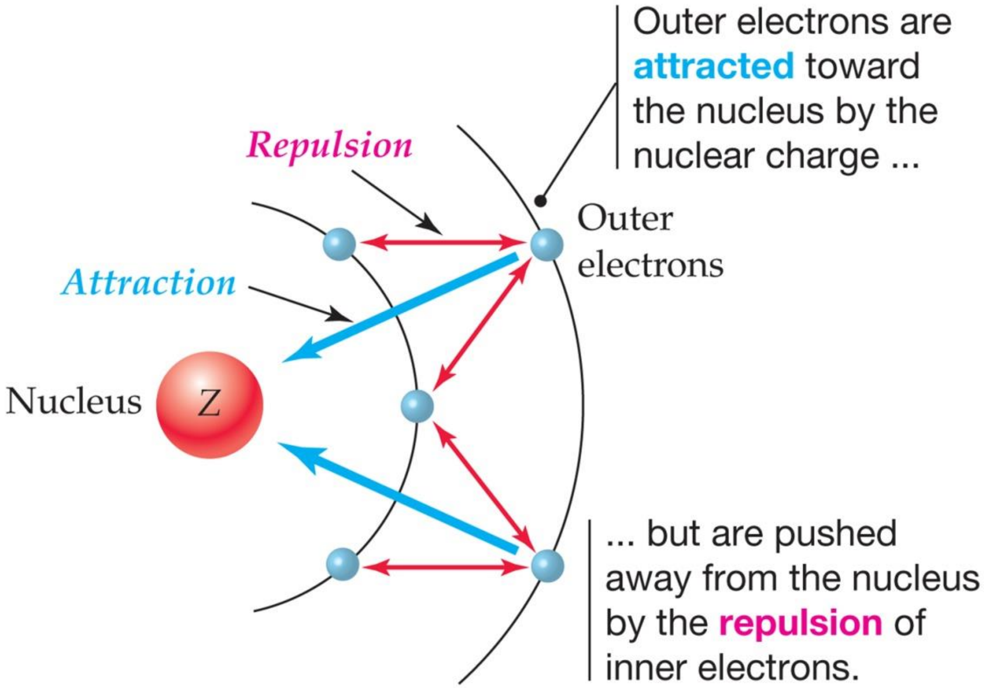
$n = 2$, $l = 1$, $m_l = -1$, $m_s = +\frac{1}{2}$

$n = 2$, $l = 1$, $m_l = -1$, $m_s = -\frac{1}{2}$

Orbital Energy Levels in Multielectron Atoms

Effective Nuclear Charge (Z_{eff}): The nuclear charge actually felt by an electron (Z actual is charge of the nucleus)

$$Z_{eff} = Z_{actual} - \text{electron shielding}$$



Electron Configurations of Multielectron Atoms

Electron Configuration: A description of which orbitals are occupied by electrons

End D section 10/23/19

Degenerate Orbitals: Orbitals that have the same energy level—for example, the 3 *p* orbitals in a given subshell, 5 d orbitals in a given subshell

Ground-State Electron Configuration: The lowest-energy electron configuration

Aufbau Principle ("building up"): A guide for determining the filling order of orbitals

Electron Configurations of Multielectron Atoms

Rule of the aufbau principle:

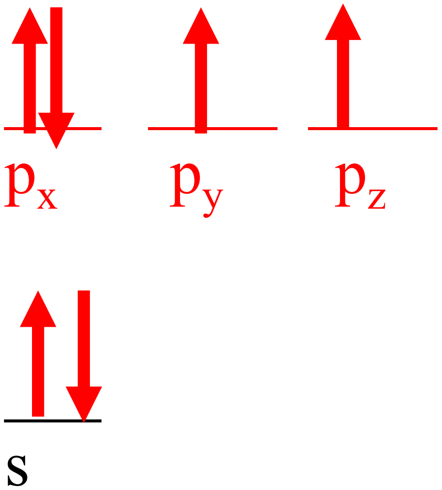
1. Lower-energy orbitals fill before higher-energy orbitals.
2. An orbital can hold only two electrons, which must have opposite spins (Pauli exclusion principle).
3. If two or more degenerate orbitals are available, follow Hund's rule.

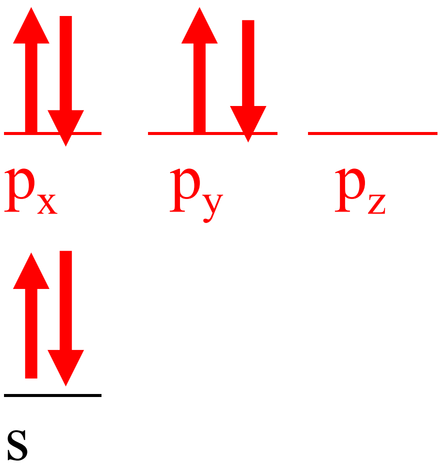
Hund's Rule: If two orbitals have the same energy (degenerate orbitals within the same subshell) – add e until all degenerate orbitals are half full and then add 2^{nd} electrons into orbitals (example 3 p degenerate orbitals)

Electron Configurations of Multielectron Atoms

Hund's Rule: If two orbitals have the same energy (degenerate orbitals within the same subshell) – add e until all degenerate orbitals are half full and then add 2nd electrons into orbitals (example 3 p degenerate orbitals)

Which violates Hund ?

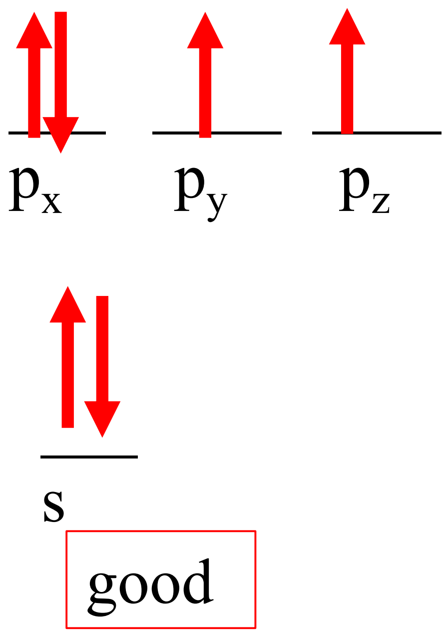


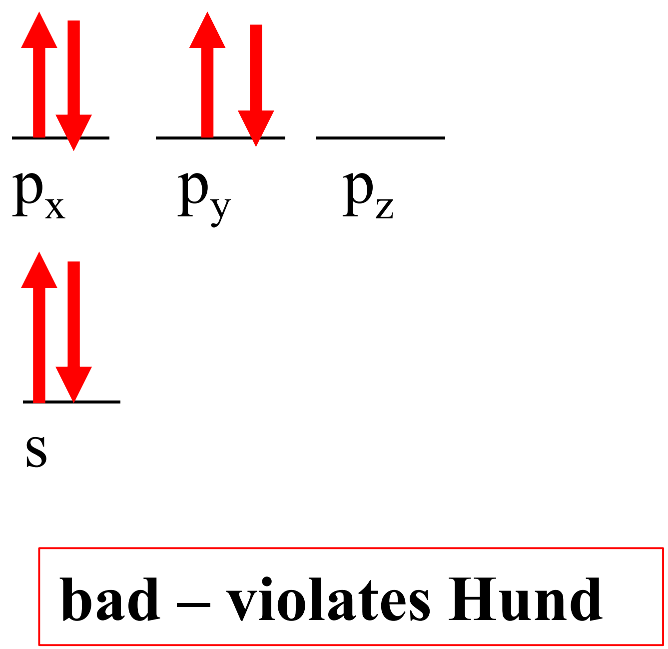


degenerate orbitals have the same energy

Electron Configurations of Multielectron Atoms

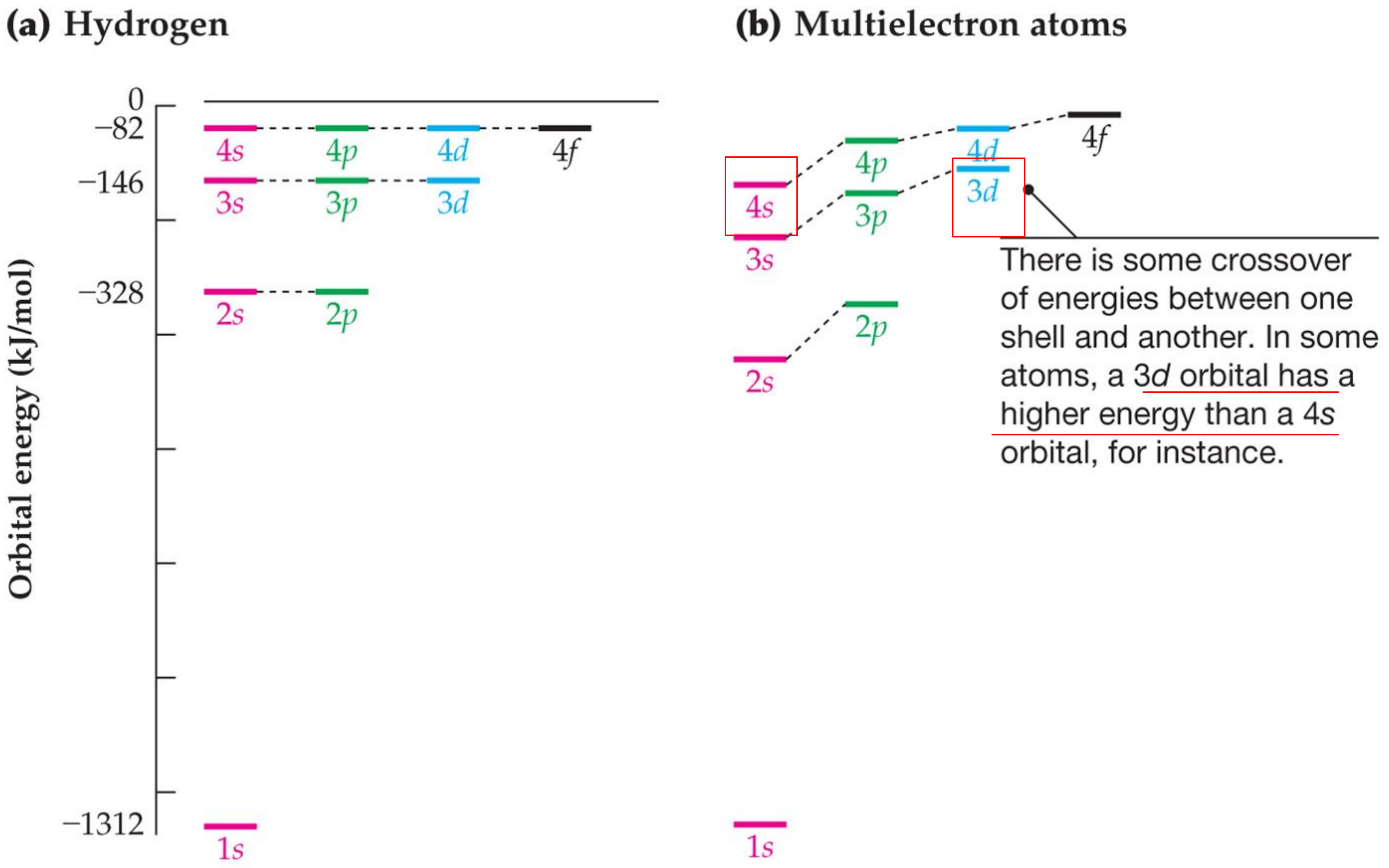
Hund's Rule: If two orbitals have the same energy (degenerate orbitals within the same subshell) – add e until all degenerate orbitals are half full and then add 2nd electrons into orbitals (example 3 p degenerate orbitals)





Electron Configurations of Multielectron Atoms

(this diagram does not show the individual degenerate orbitals within the subshells)



There is some crossover of energies between one shell and another. In some atoms, a 3*d* orbital has a higher energy than a 4*s* orbital, for instance.

Electron Configurations of Multielectron Atoms

Electron Configuration

H: $1s^1$

→ 1 electron (# of electron)

→ *s* orbital (l = 0)

→ n = 1

Get electron configuration from position of element in periodic table.

n = period number (shell)

s (or choose s,p,d,f) is from regions of periodic table (subshell)

electrons - one move to right in periodic table = one more electron (what kind of electron ? from position of element in periodic table)

What is outermost electron orbital ?

S block

d block n-1

p block

	1A	2A	3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	8A
1	1 H 1.008																	2 He 4.003
2	3 Li 6.94	4 Be 9.01											5 B 10.81	6 C 12.011	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
3	11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
4	19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.90	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.71	29 Cu 63.55	30 Zn 65.37	31 Ga 69.72	32 Ge 72.59	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
5	37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc 98.91	44 Ru 101.07	45 Rh 102.91	46 Pd 106.4	47 Ag 107.87	48 Cd 112.40	49 In 114.82	50 Sn 118.69	51 Sb 121.75	52 Te 127.60	53 I 126.90	54 Xe 131.30
6	55 Cs 132.91	56 Ba 137.34	57 La 138.91	72 Hf 178.49	73 Ta 180.95	74 W 183.85	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.09	79 Au 196.97	80 Hg 200.59	81 Tl 204.37	82 Pb 207.19	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
7	87 Fr (223)	88 Ra 226.03	89 Ac (227)	104 Unq* (261)	105 Unp* (262)	106 Unh* (263)	107 Uns* (262)	108 Uno* (265)	109 Una* (266)									

f block

n-2

Lanthanides	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.35	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
Actinides	90 Th 232.04	91 Pa (231)	92 U 238.03	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (249)	98 Cf (249)	99 Es (254)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)

*Symbol (and name) provisional.

Numbers in parentheses: available radioactive isotope of longest half-lif

Electron Configurations of Multielectron Atoms

	Electron Configuration
H:	$1s^1$
He:	$1s^2$

$1s^2$ → 2 electrons

→ s orbital (l = 0)

→ n = 1

Electron Configurations of Multielectron Atoms

Electron Configuration

H: $1s^1$

He: $1s^2$

Lowest energy to **highest** energy

Li: $1s^2\ 2s^1$ → 1 electron

s orbital ($l = 0$)

$n = 2$

Electron Configurations of Multielectron Atoms

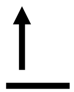
	Electron Configuration
H:	$1s^1$
He:	$1s^2$
Li:	$1s^2\ 2s^1$
N:	$1s^2\ 2s^2\ 2p^3$

$2p^3$ → 3 electrons

→ p orbital ($l = 1$)

→ $n = 2$

Electron Configurations of Multielectron Atoms

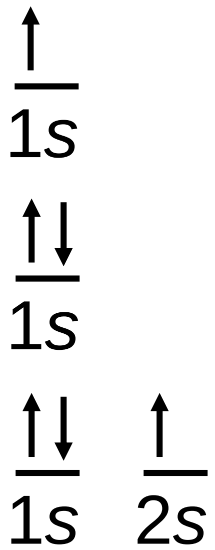
	Electron Configuration	Orbital-Filling Diagram
H:	$1s^1$	 $1s$
He:	$1s^2$	
Li:	$1s^2\ 2s^1$	
N:	$1s^2\ 2s^2\ 2p^3$	

Electron Configurations of Multielectron Atoms

	Electron Configuration	Orbital-Filling Diagram
H:	$1s^1$	↑ $1s$
He:	$1s^2$	↑↓ $1s$
Li:	$1s^2\ 2s^1$	
N:	$1s^2\ 2s^2\ 2p^3$	

Electron Configurations of Multielectron Atoms

	Electron Configuration	Orbital-Filling Diagram
H:	$1s^1$	
He:	$1s^2$	
Li:	$1s^2\ 2s^1$	
N:	$1s^2\ 2s^2\ 2p^3$	



Electron Configurations of Multielectron Atoms

	Electron Configuration	Orbital-Filling Diagram
H:	$1s^1$	↑ (1*s*)
He:	$1s^2$	↑↓ (1*s*)
Li:	$1s^2\ 2s^1$	↑↓ (1*s*) ↑ (2*s*)
N:	$1s^2\ 2s^2\ 2p^3$	↑↓ (1*s*) ↑↓ (2*s*) ↑ ↑ ↑ (2*p*)

HW: Electron Configurations of Multielectron Atoms

Electron Configuration

He: $1s^2$

- → 2 electrons
- → s orbital (l = 0)
- → n = 1

Give the electron configuration of the element Ca, S and Ti (starting from $1s^2$ electrons in the format shown above)

End 10/23 Wednesday F & G section

HW: Electron Configurations of Multielectron Atoms

Electron Configuration

He: $1s^2$ → 2 electrons
→ s orbital ($l = 0$)
→ $n = 1$

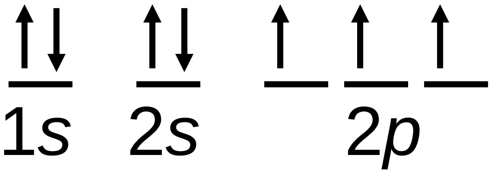
Give the electron configuration of the element Ca, S and Ti (starting from $1s^2$ electrons in the format shown above)

Ca $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2$ (Ar)

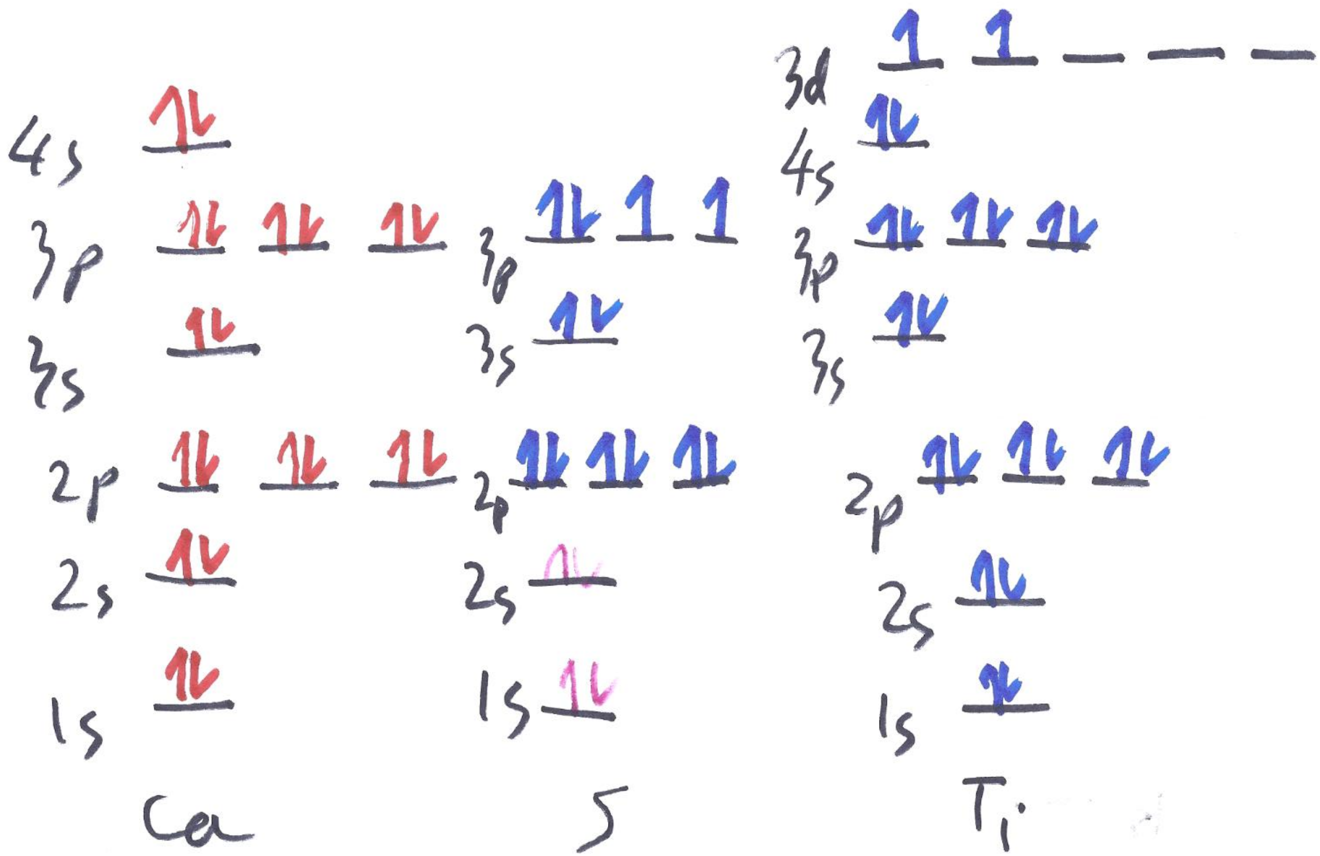
S $1s^2, 2s^2, 2p^6, 3s^2, 3p^4$ (Ne)

Ti $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^2$ (Ar)

HW: Electron Configurations of Multielectron Atoms

	Electron Configuration	Orbital-Filling Diagram
N:	$1s^2\ 2s^2\ 2p^3$	

Give the orbital filling diagram for the element Ca, S and Ti (starting from $1s^2$ electrons in the format shown above)


3d
4s
3p
3s
2p
2s
1s
Ca
S
Ti

Electron Configurations of Multielectron Atoms

	Electron Configuration	Shorthand Configuration
Na:	$1s^2\ 2s^2\ 2p^6\ 3s^1$	

Electron Configurations of Multielectron Atoms

	Electron Configuration	Shorthand Configuration
Na:	$1s^2\ 2s^2\ 2p^6\ 3s^1$	$[Ne]\ 3s^1$

Ne configuration

Electron Configurations of Multielectron Atoms

	Electron Configuration	Shorthand Configuration
Na:	$1s^2\ 2s^2\ 2p^6\ 3s^1$	$[Ne]\ 3s^1$
P:	$\underbrace{1s^2\ 2s^2\ 2p^6}_{\text{Ne configuration}}\ 3s^2\ 3p^3$	

Electron Configurations of Multielectron Atoms

	Electron Configuration	Shorthand Configuration
Na:	$1s^2\ 2s^2\ 2p^6\ 3s^1$	$[\mathrm{Ne}]\ 3s^1$
P:	$\underbrace{1s^2\ 2s^2\ 2p^6}_{\text{Ne configuration}}\ 3s^2\ 3p^3$	$[\mathrm{Ne}]\ 3s^2\ 3p^3$

Electron Configurations of Multielectron Atoms

	Electron Configuration	**Shorthand Configuration**
Na:	$1s^2\ 2s^2\ 2p^6\ 3s^1$	[Ne] $3s^1$
P:	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$	[Ne] $3s^2\ 3p^3$
K:	$\underbrace{1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6}_{\text{Ar configuration}}\ 4s^1$	

Electron Configurations of Multielectron Atoms

	Electron Configuration	**Shorthand Configuration**
Na:	$1s^2\ 2s^2\ 2p^6\ 3s^1$	$[Ne]\ 3s^1$
P:	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$	$[Ne]\ 3s^2\ 3p^3$
K:	$\underbrace{1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6}_{\text{Ar configuration}}\ 4s^1$	$[Ar]\ 4s^1$

Electron Configurations of Multielectron Atoms

	Electron Configuration	Shorthand Configuration
Na:	$1s^2\ 2s^2\ 2p^6\ 3s^1$	$[Ne]\ 3s^1$
P:	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$	$[Ne]\ 3s^2\ 3p^3$
K:	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$	$[Ar]\ 4s^1$
Sc:	$\underbrace{1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6}_{\text{Ar configuration}}\ 4s^2\ 3d^1$	$[Ar]\ 4s^2\ 3d^1$

Ar configuration

Anomalous Electron Configurations
(for transition metals near d^5 or d^{10})

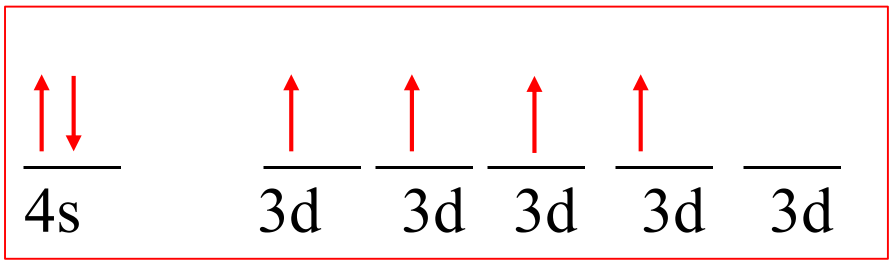
	Expected Configuration	Actual Configuration
Cr:	[Ar] $4s^2\ 3d^4$	[Ar] $4s^1\ 3d^5$
Cu:	[Ar] $4s^2\ 3d^9$	[Ar] $4s^1\ 3d^{10}$

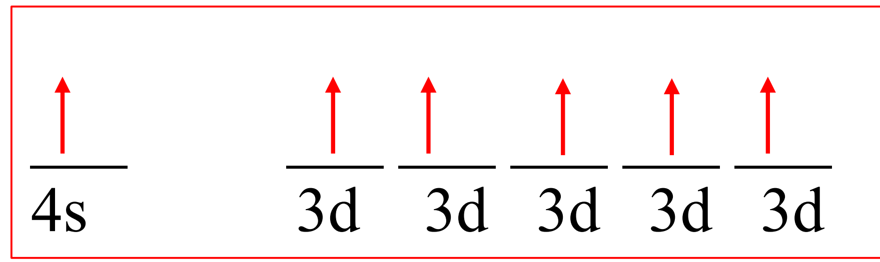
End 10/24 R D section

Anomalous Electron Configurations

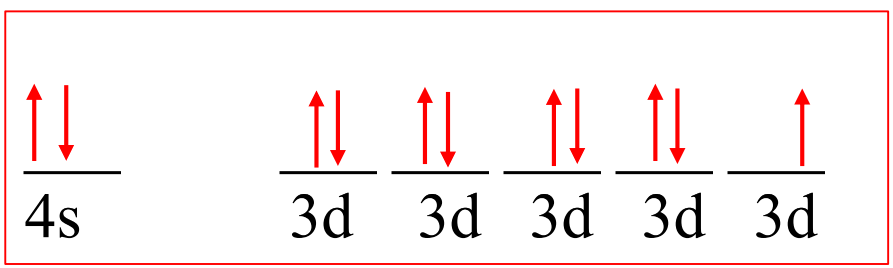
(transition metals near d^5 or d^{10}) (full or half filled d is stable)

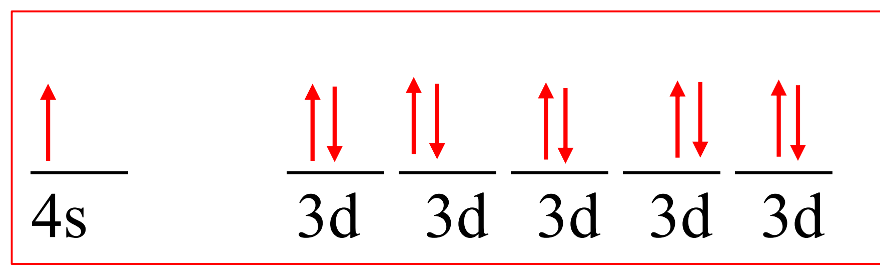
	Expected Configuration	Actual Configuration
Cr:	[Ar] $4s^2\ 3d^4$	[Ar] $4s^1\ 3d^5$



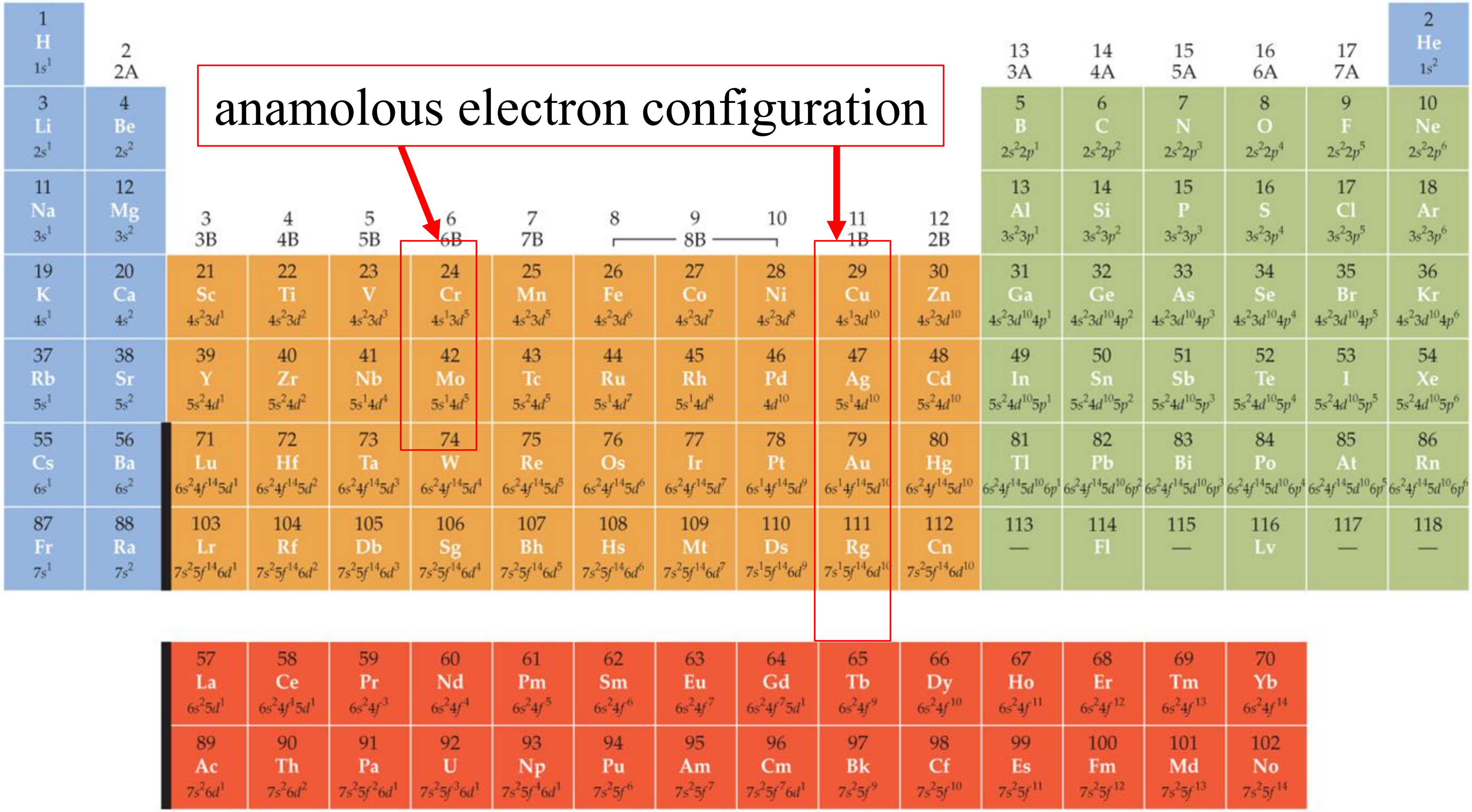


	Expected Configuration	Actual Configuration
Cu:	[Ar] $4s^2\ 3d^9$	[Ar] $4s^1\ 3d^{10}$





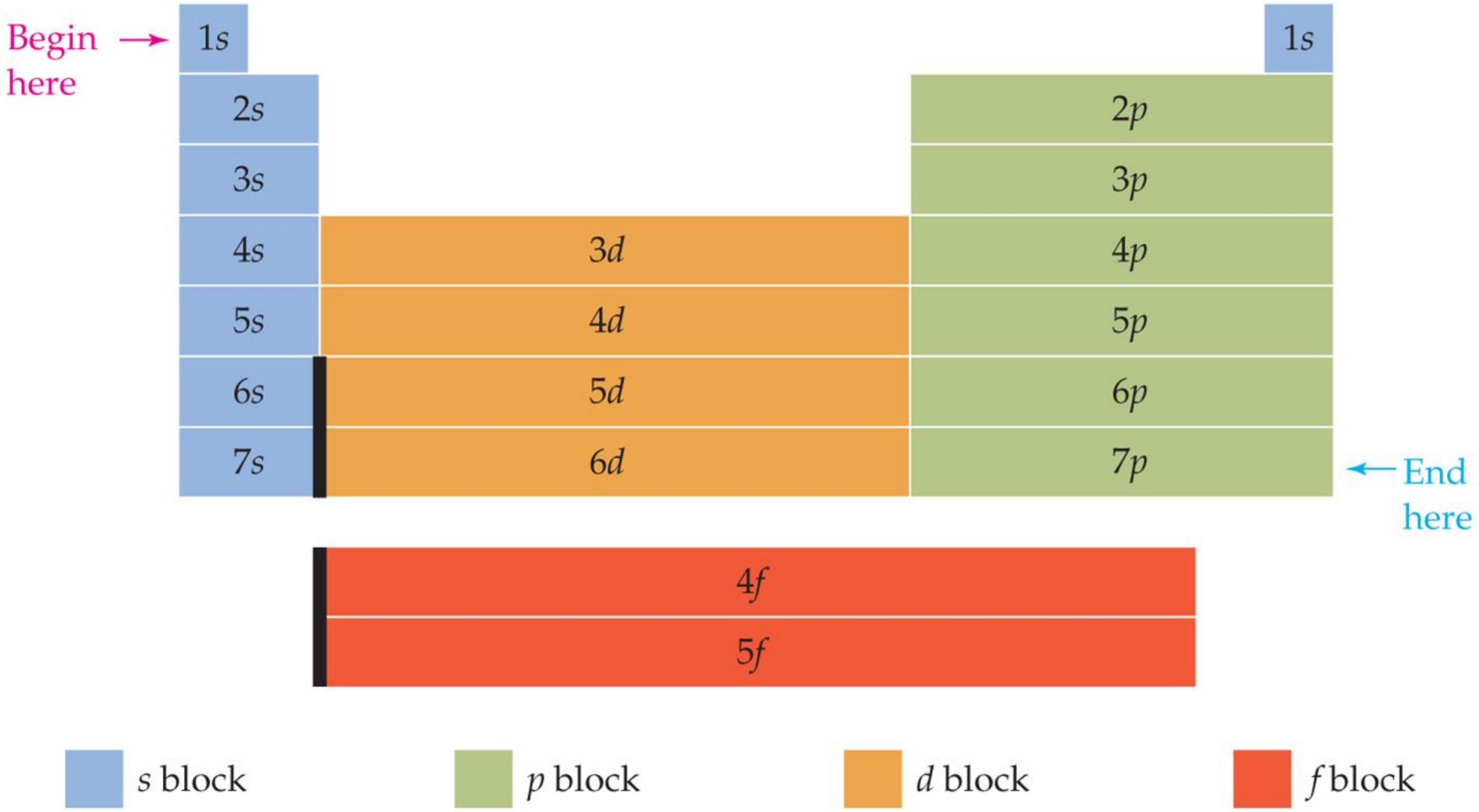
Electron configuration for all elements for last set of electrons (for main group elements – NOT valence electrons)



inner electrons are noble gas configurations

Electron Configurations and the Periodic Table

The arrangement of the periodic table provides a method for remembering the order of orbital filling. Beginning at the top left and moving across successive rows, the order is 1*s* → 2*s* → 2*p* → 3*s* → 3*p* → 4*s* → 3*d* → 4*p* and so on.



Electron Configurations and the Periodic Table

Valence Shell: Outermost shell

TABLE 5.3 Valence-Shell Electron Configurations of Main-Group Elements

Group	Valence-Shell Electron Configuration	
1A	ns^1	(1 total)
2A	ns^2	(2 total)
3A	ns^2np^1	(3 total)
4A	ns^2np^2	(4 total)
5A	ns^2np^3	(5 total)
6A	ns^2np^4	(6 total)
7A	ns^2np^5	(7 total)
8A	ns^2np^6	(8 total)

Li: $2s^1$

Na: $3s^1$

Cl: $3s^2\ 3p^5$

Br: $4s^2\ 4p^5$

valence electrons = group number for main group elements

HW: Electron Configurations and the Periodic Table

End 10/25
Friday F,G
section

End 10/28 D

Valence Shell: Outermost shell

valence electrons = group number for main group elements (leave out d electrons for valence electrons)

Li: $2s^1$

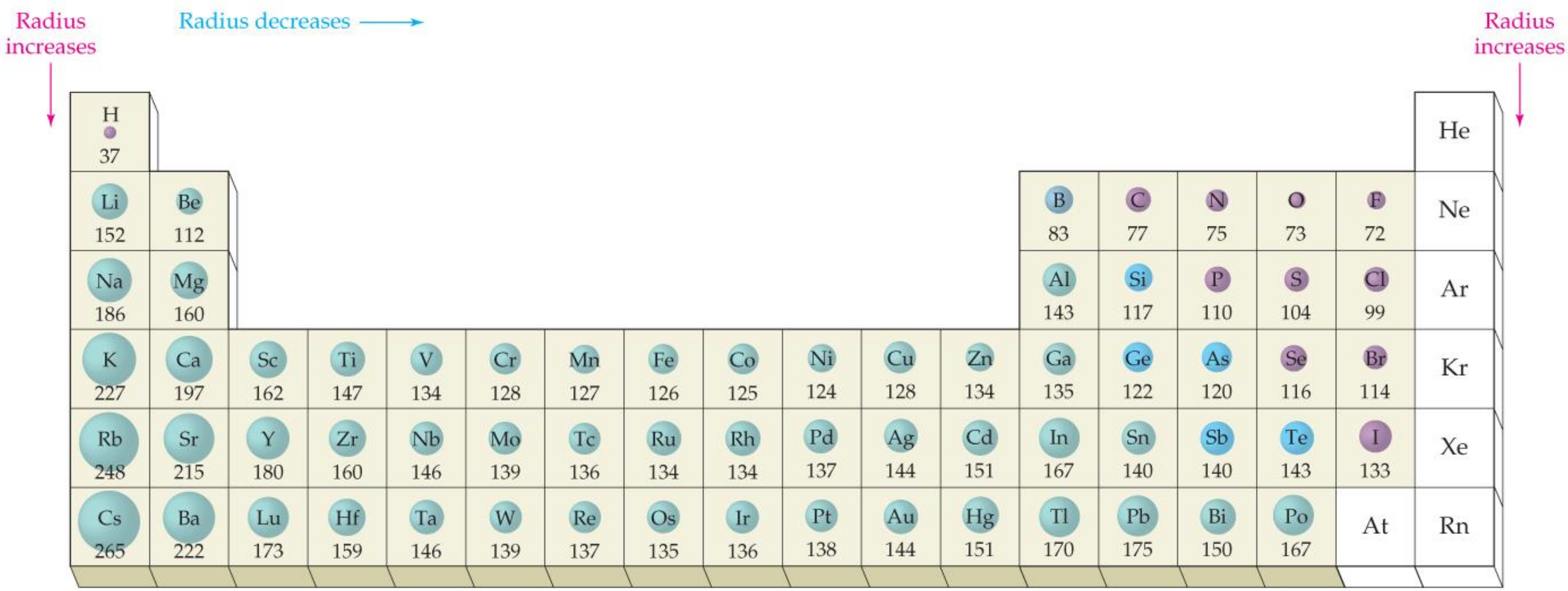
Cl: $3s^2\ 3p^5$

Give electron configuration of just the valence electrons for the main group elements:

Se

Ba

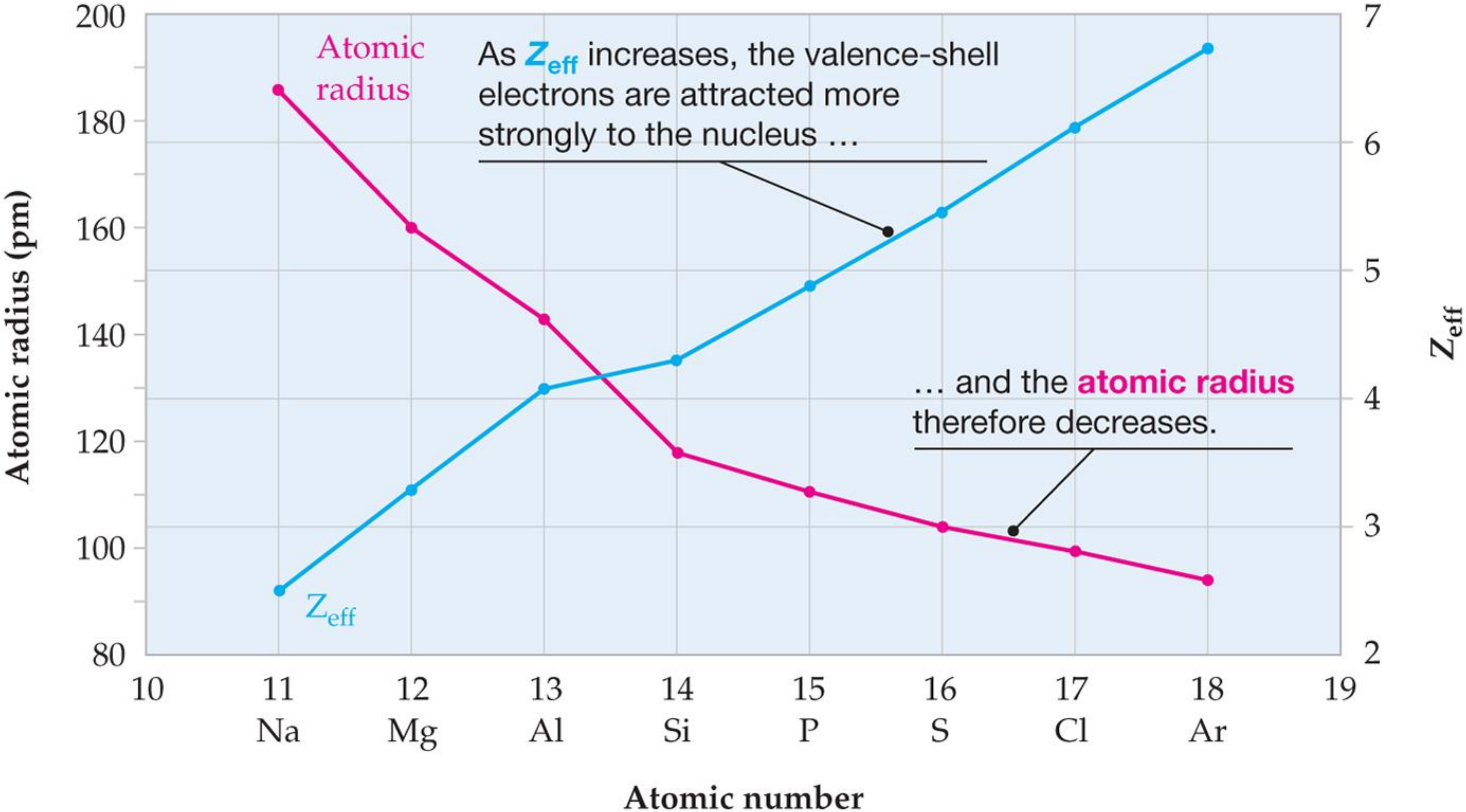
Electron Configurations and Periodic Properties: Atomic Radii



↓ Column ↑ Radius (size)

→ Row ↓ Radius

Electron Configurations and Periodic Properties: Atomic Radii



Electron Configurations and Periodic Properties: Atomic Radii

Which of the following pairs of atoms is LARGER ?

Smaller → across period

Larger ↓ down group

in periodic table

N vs F

N vs As

Na vs Ar

F vs I

Quiz 3 & Test 3 ends at the end of Chapter 5

.(We finished Chapter 6 in my F&G sections.   Section D is about 1.5 lectures away from finishing Chapter 6.   No part of Chapter 6 will be on either Quiz 3 nor Test 3)