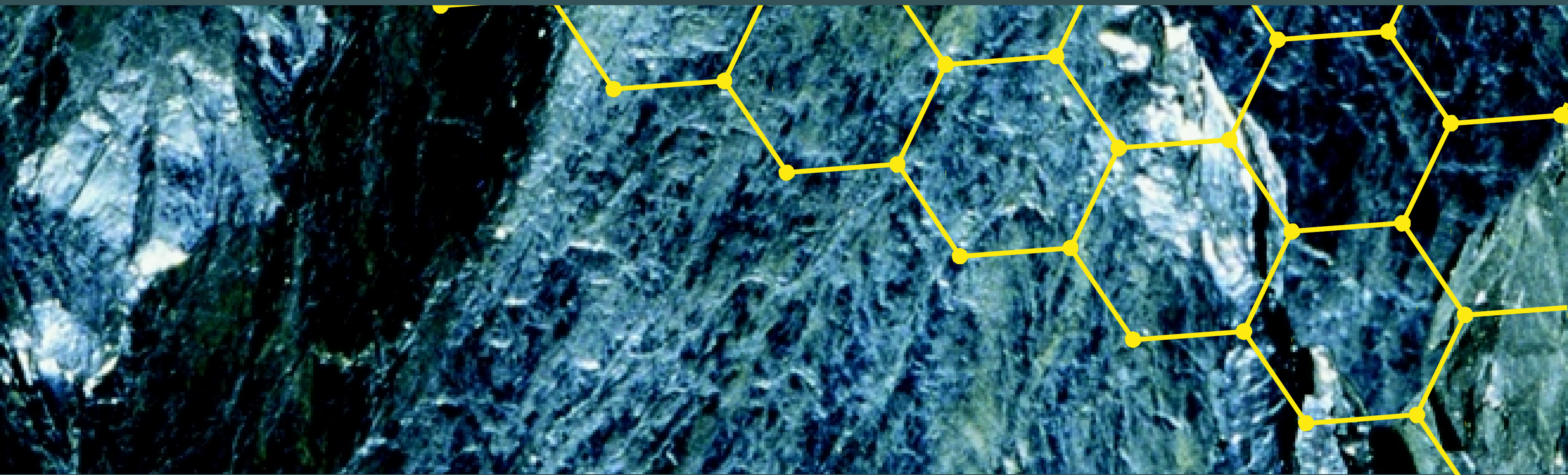
CHEMISTRY

an atoms-focused approach

Gilbert
Kirss
Foster



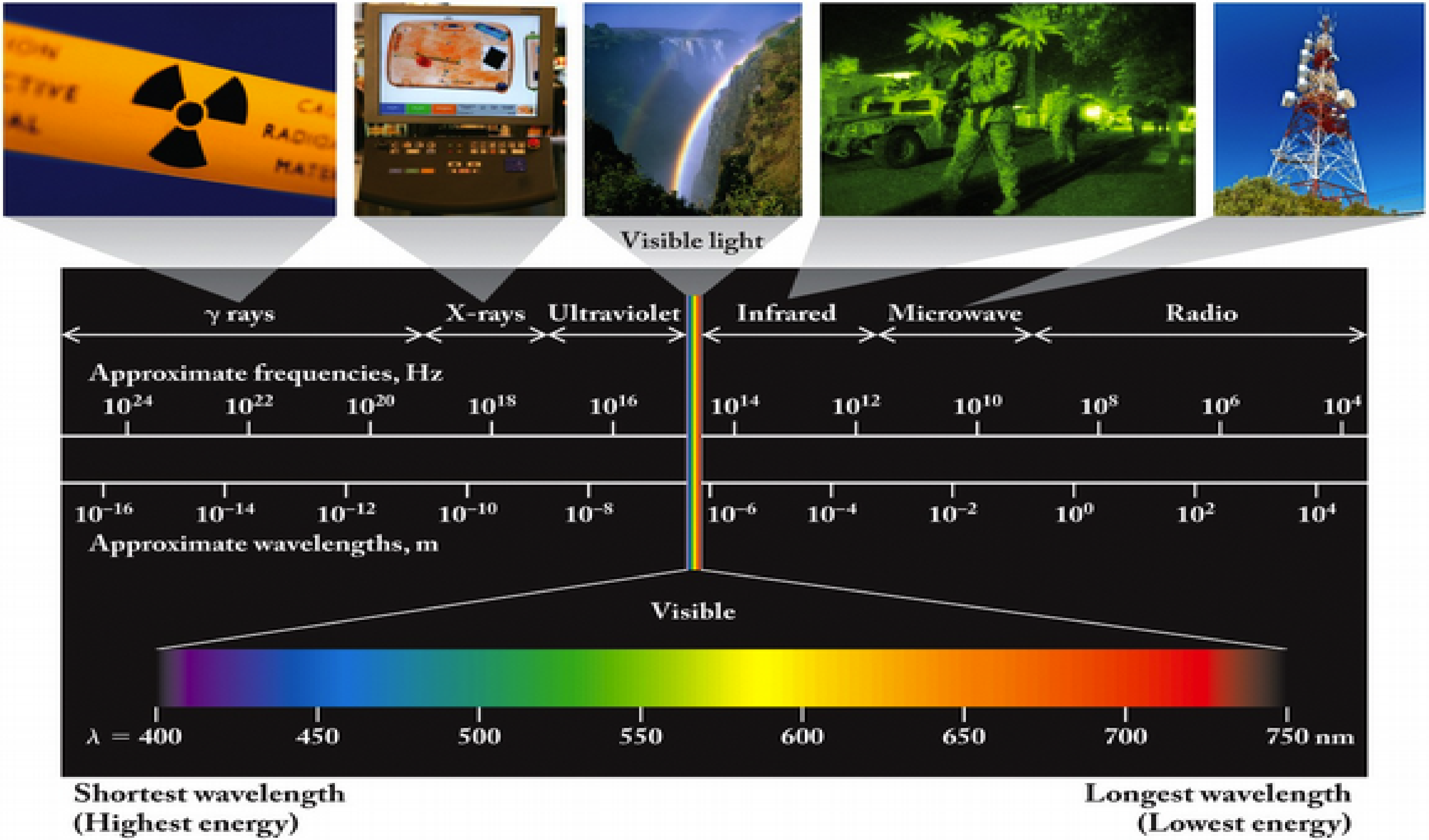
Chapter 3

Atomic Structure

Explaining the Properties of Elements

Chapter Outline

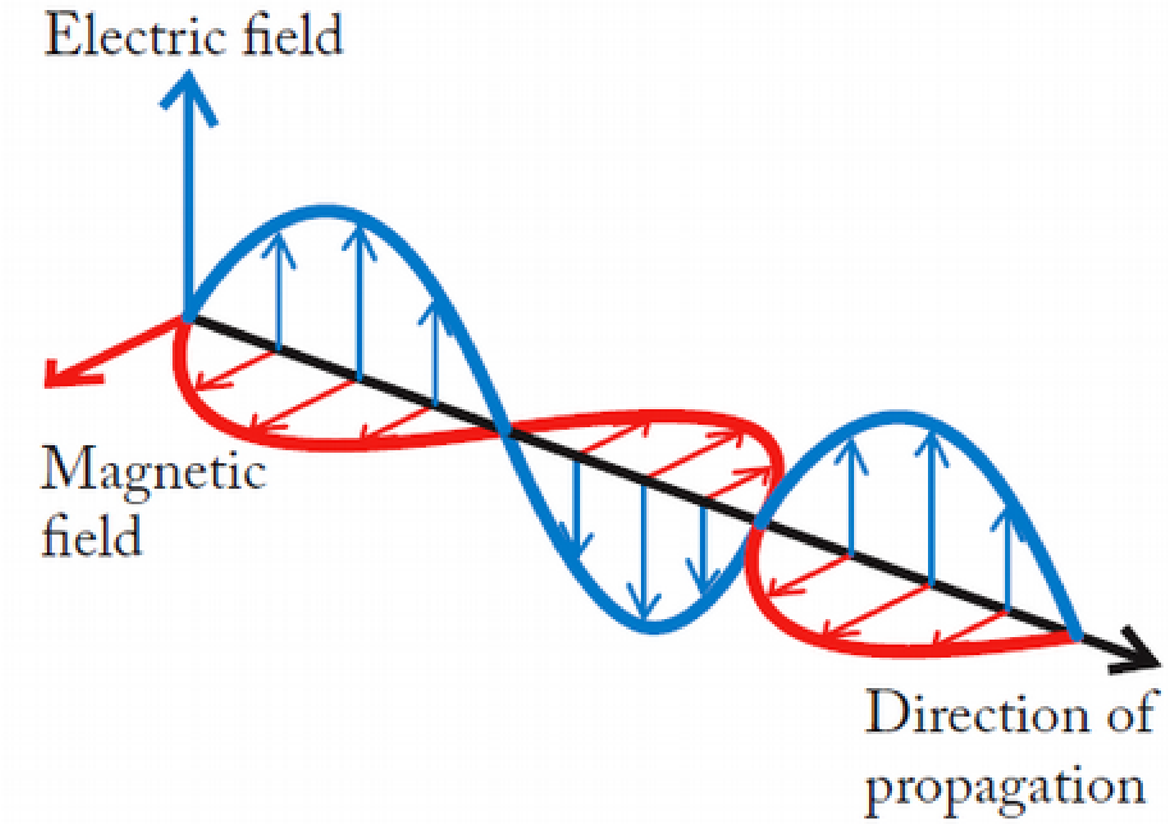
Electromagnetic Spectrum



Electromagnetic Radiation

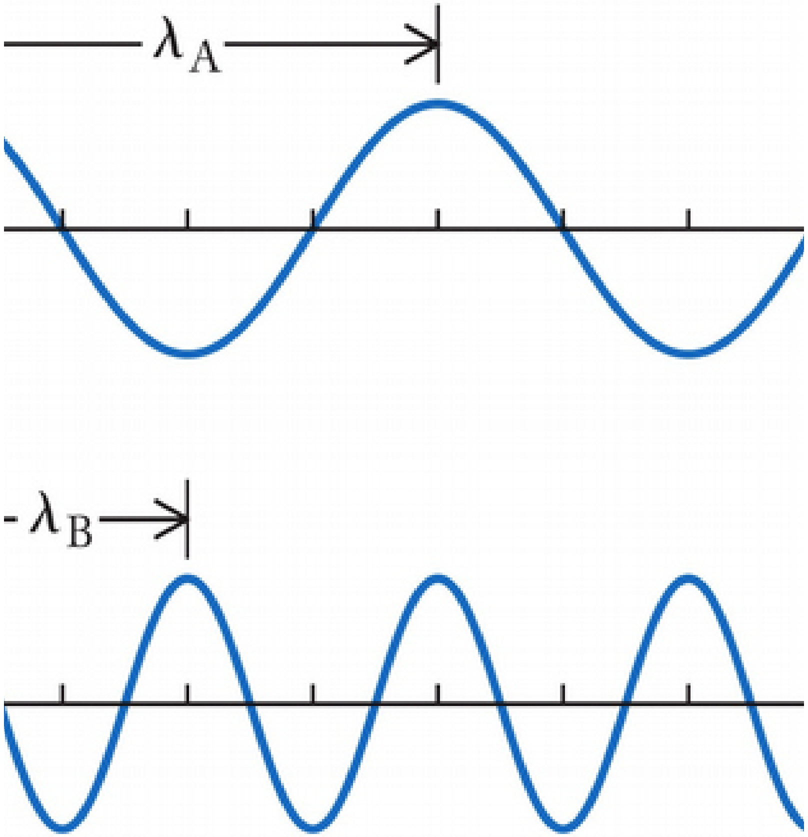
Perpendicular oscillating fields:

a) Electric
b) Magnetic



Wave Properties of Light

- Wavelength (λ):
  - Distance from crest to crest or trough to trough
- Frequency (ν):
  - The number of times a wave passes a point per unit of time
- Amplitude:
  - The height of the crest or depth of the trough



Wave Properties of Light

- Wavelength and frequency related by

  $\lambda\nu = c$

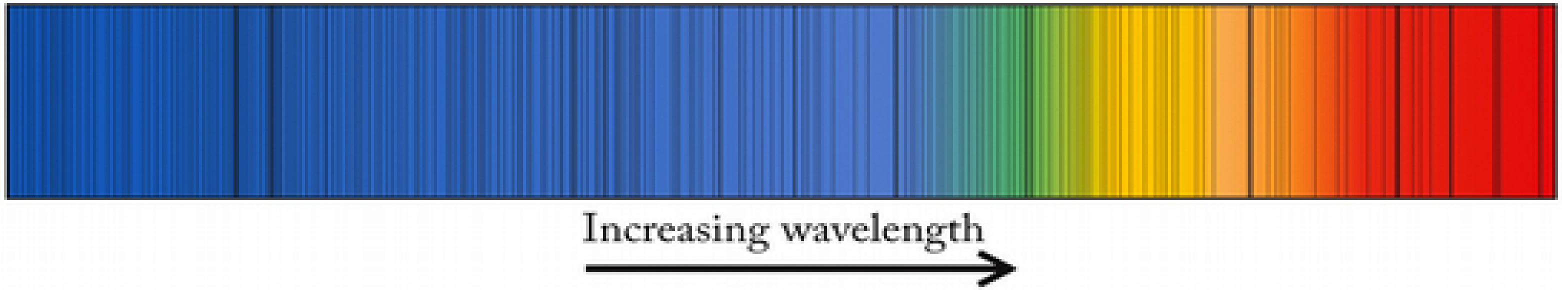
  - Where λ is wavelength (in meters), ν is frequency (in hertz, or s^{-1}), and *c* is the speed of light in vacuum
  - *c* is a constant and in a vacuum equals 2.99792458×10^8 m/s

Chapter Outline

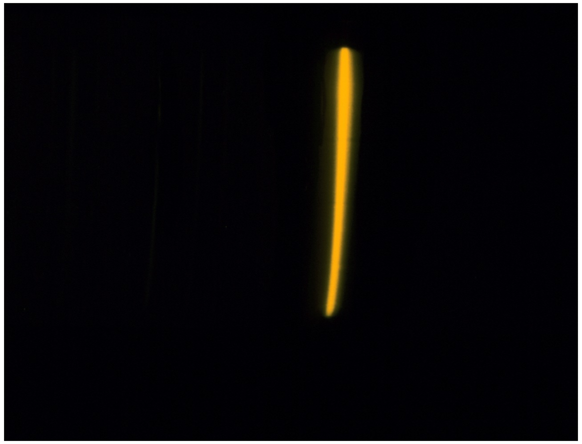
3.1 Waves of Light
3.2 Atomic Spectra
3.3 Particles of Light: Quantum Theory
3.4 The Hydrogen Spectrum and the Bohr Model
3.5 Electrons as Waves
3.6 Quantum Numbers
3.7 The Sizes and Shapes of Atomic Orbitals
3.8 The Periodic Table and Filling Orbitals
3.9 Electron Configurations of Ions
3.10 The Sizes of Atoms and Ions
3.11 Ionization Energies
3.12 Electron Affinities

Atomic Spectra

a) Fraunhofer lines (dark spectra)

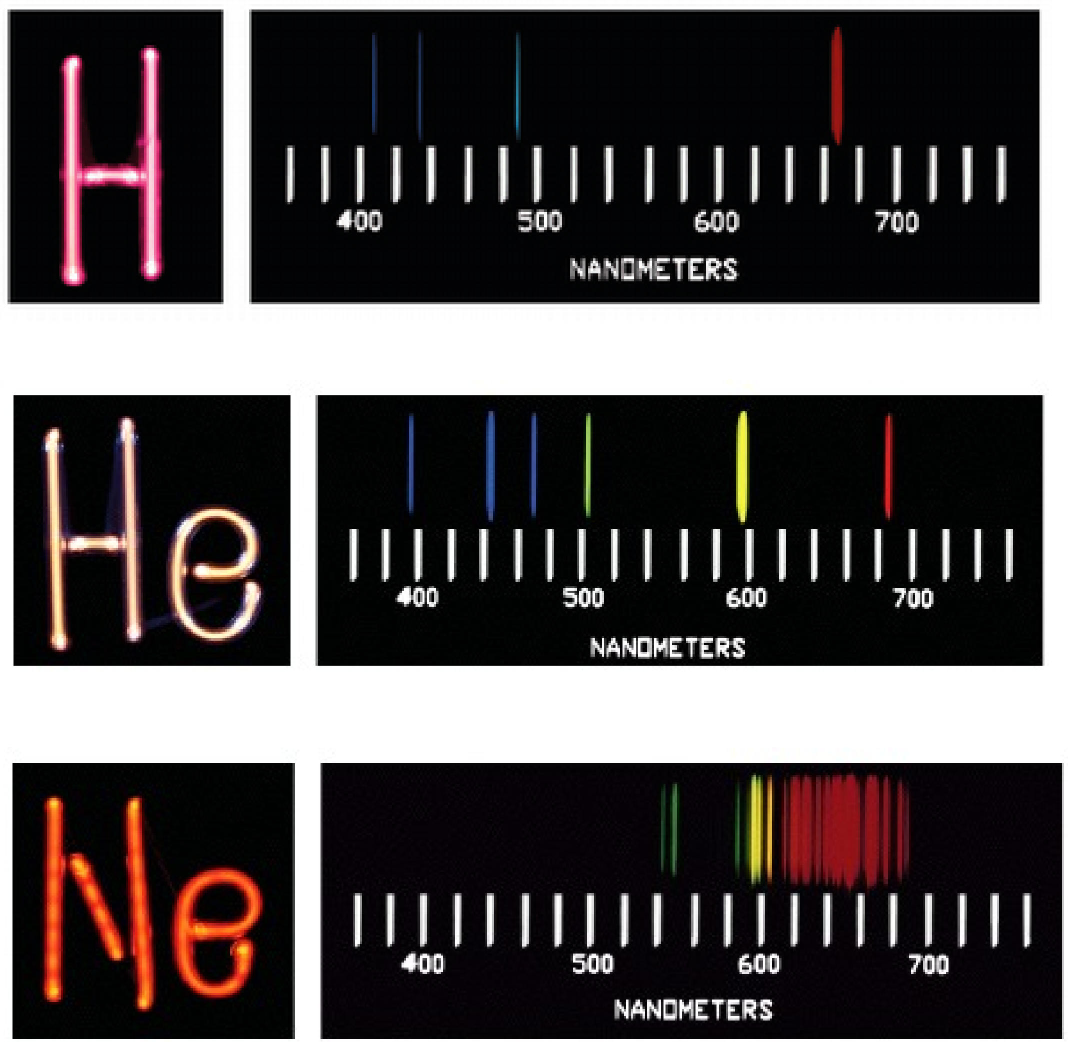


b) Sodium emission (bright line spectra)

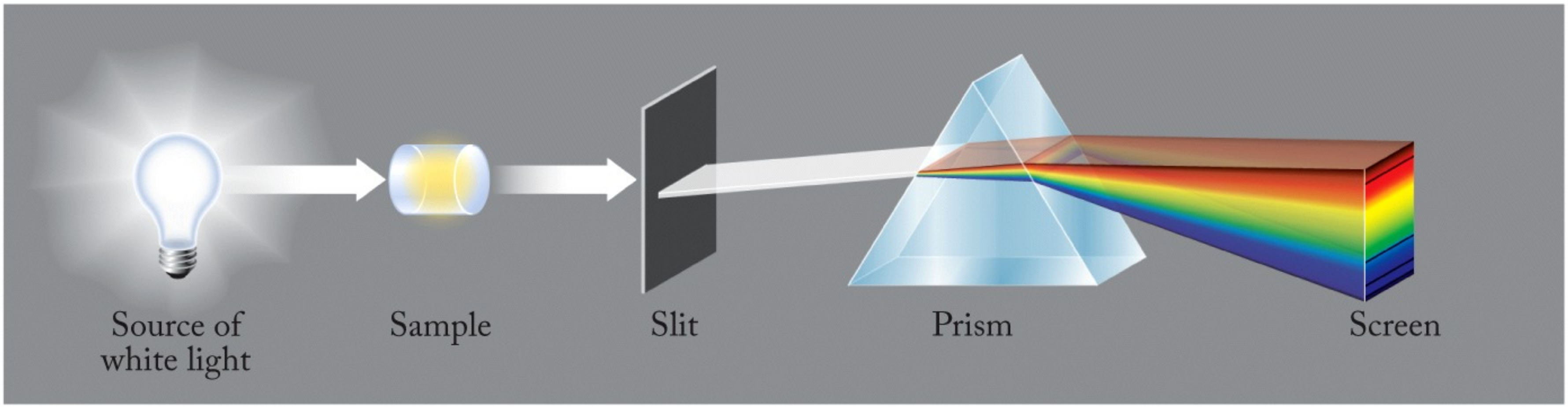


Atomic Emission Spectra

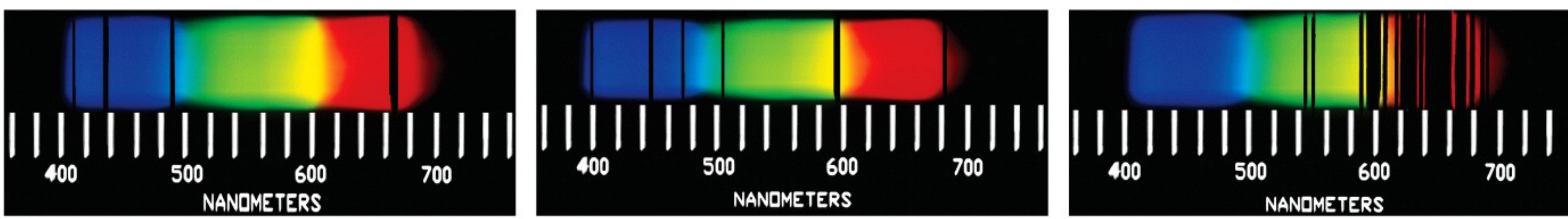
- Bright line emission spectra of other elements.



Atomic Spectra: Dark Lines



(a)



Absorption spectrum of hydrogen   Absorption spectrum of helium   Absorption spectrum of neon

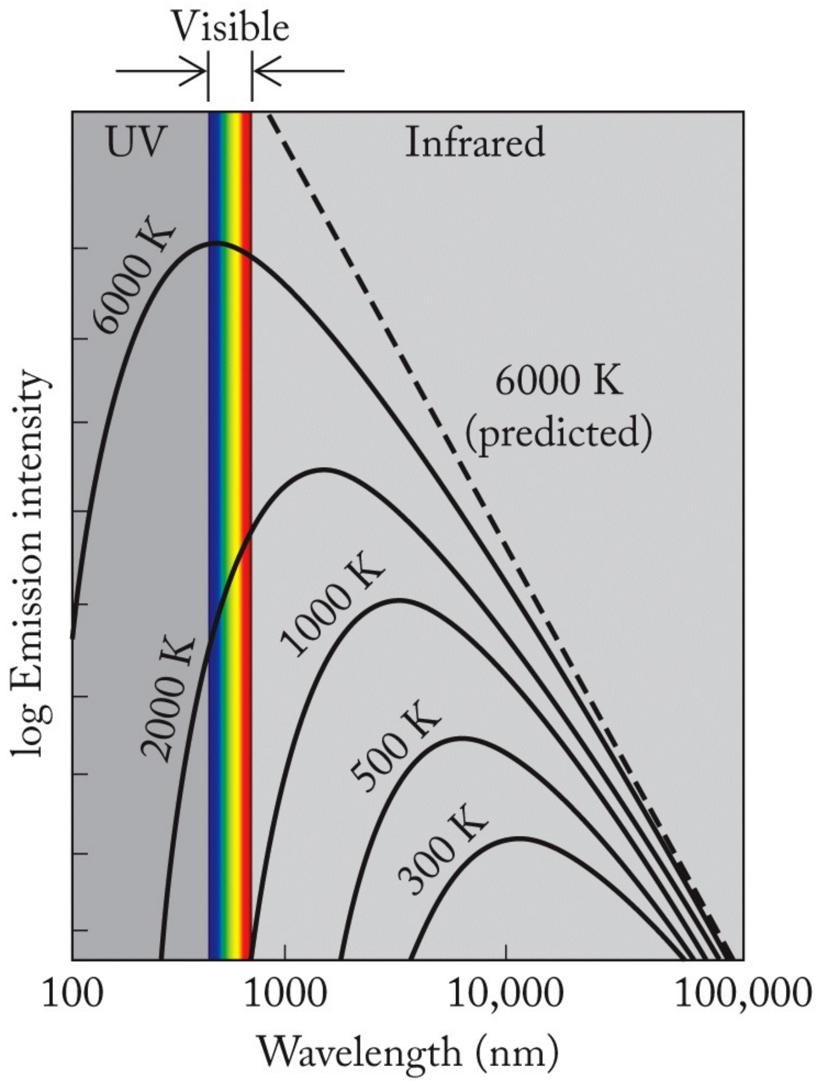
(b)

Hydrogen   Helium   Neon

Chapter Outline

Particles vs. Waves

- Incandescence not consistent with Maxwell equations, wave theory.
  - Most intense emissions in the infrared region.
  - Little or no emissions in UV.



Quantum Theory

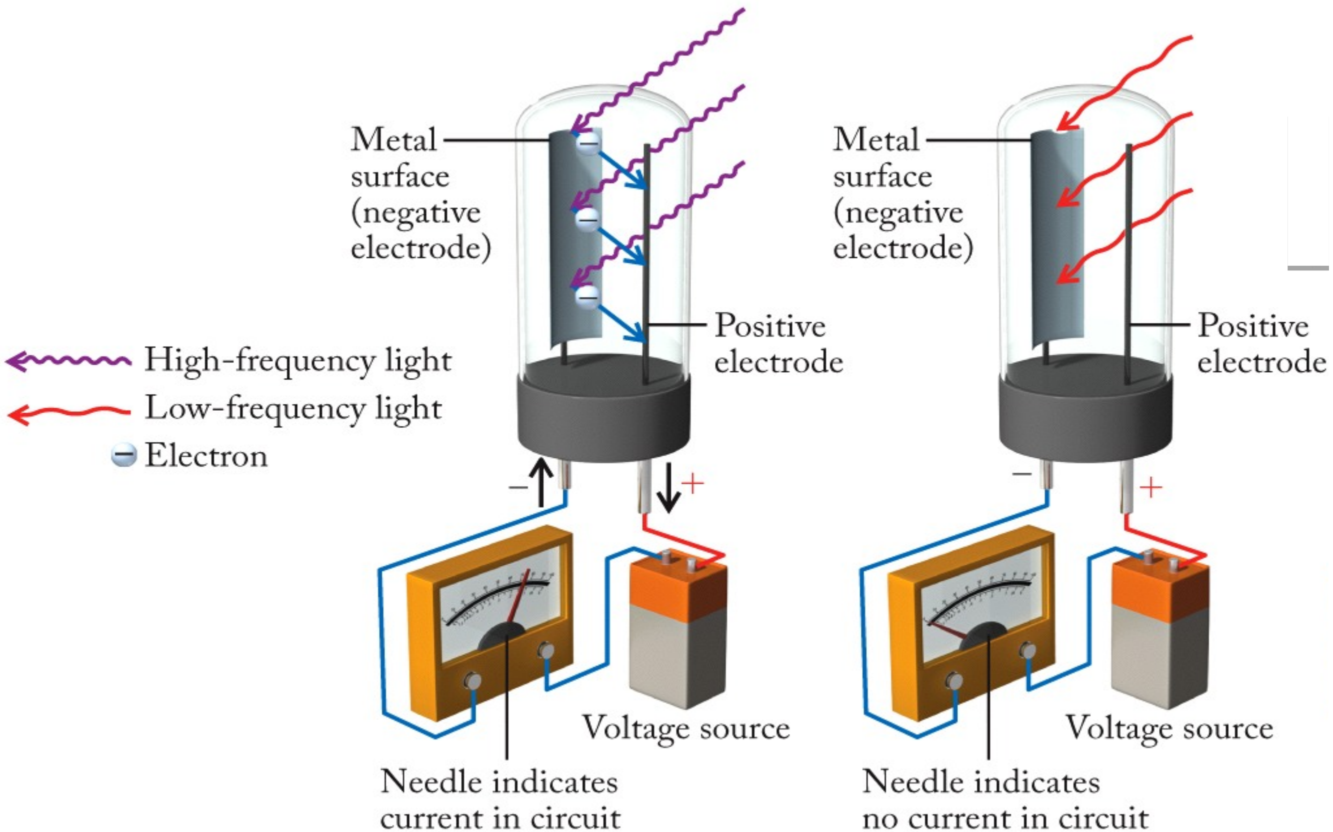
- Radiant energy is "quantized"
  - Having values restricted to whole-number multiples of a specific base value
- Quantum = smallest discrete quantity of energy
- Photon = a quantum of electromagnetic radiation
- Energy of photon: $E = h\nu$ or $E = hc/\lambda$
  - $h = 6.6260755 \times 10^{-34}$ J•s (Planck's constant)

Quantized States

- Quantized states: Discrete energy levels (e.g., steps)
- Continuum states: Smooth transition between levels (e.g., ramp)

Photoelectric Effect

- Photoelectric Effect:
  - Phenomenon of light striking a metal surface and producing an electric current (flow of electrons)



If radiation below threshold energy, no electrons released.

Photoelectric Effect

- Explained by quantum theory
  - Photons of sufficient energy ($h\nu$) dislodge e^- from metal surface
  - Work function (Φ) – amount of energy needed to dislodge an electron from the surface of a metal:

$$\Phi = h\nu_o$$

where Φ = work function; ν_o = threshold frequency

  - Kinetic energy of ejected electrons: $KE_{electron} = h\nu - \Phi$

Chapter Outline

Practice: Rydberg Equation

What is the wavelength of the line in the visible spectrum corresponding to $n_1 = 2$ and $n_2 = 4$?

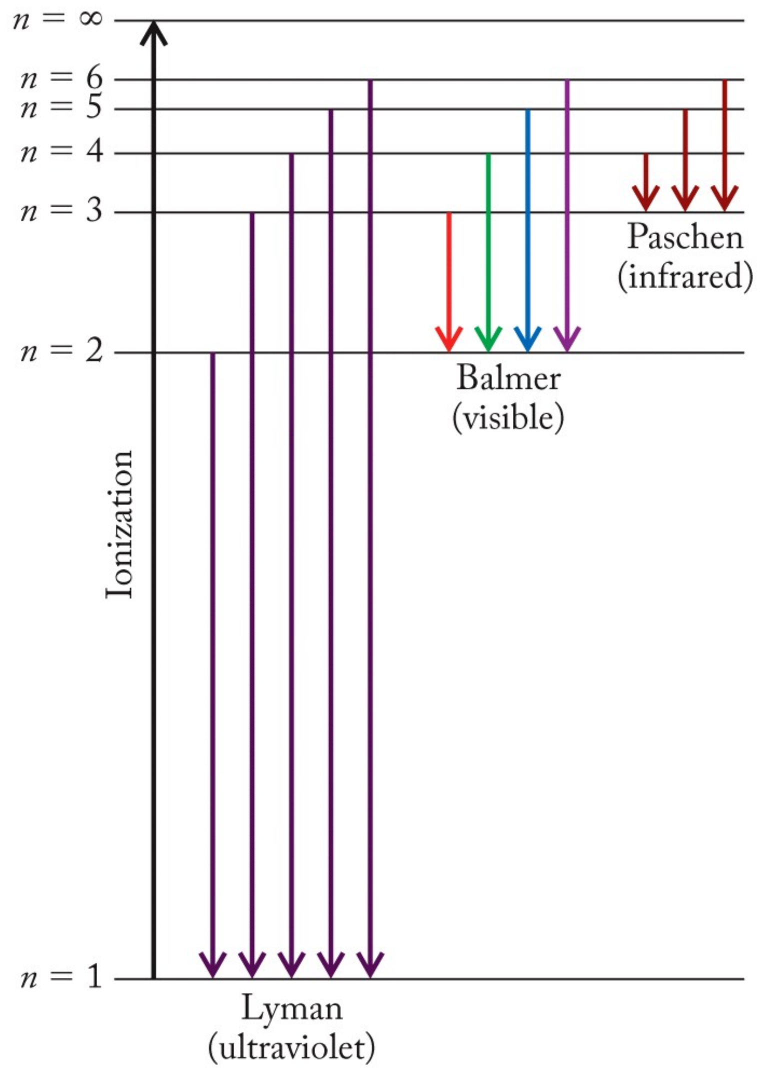
- Collect and Organize:
- Analyze:
- Solve:
- Think about It:

Electronic States

- Energy Level: An allowed energy band that an electron can occupy in an atom
- Ground State: Lowest energy state available to an electron in an atom
- Excited State: Any energy state above the ground state
- Electron Transition: Movement of an electron between energy states

Energy Level Diagram

- Shows electron **transitions** (i.e., movement of electrons between energy levels)
- ↑ = absorption of energy
- ↓ = emission of energy



Bohr Model

- Strengths:
  - Accurately predicts energy needed to remove an electron from an atom (ionization)
  - Allowed scientists to begin using quantum theory to explain matter at atomic level
- Limitations:
  - Does not account for spectra of multielectron atoms
  - Movement of electrons in atoms is less clearly defined than Bohr allowed

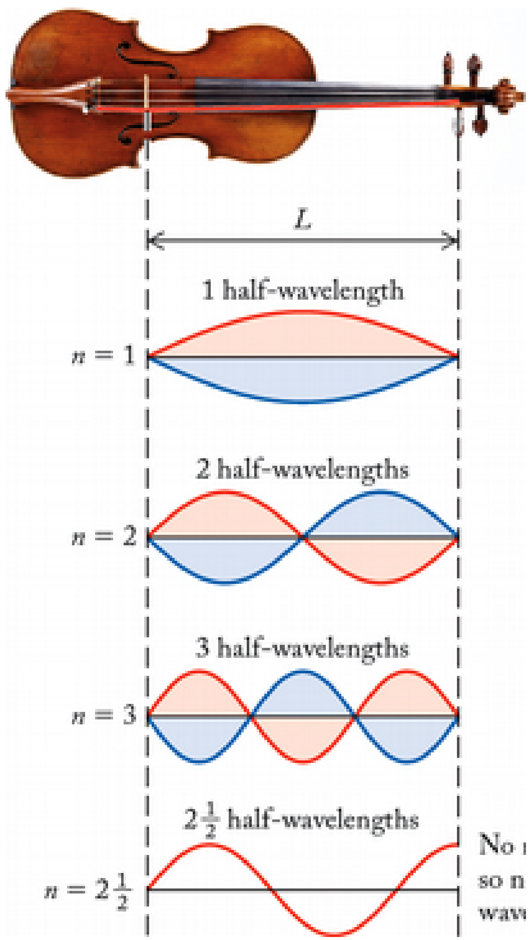
Chapter Outline

Particles or Waves?

- De Broglie (1892–1987)
  - If electromagnetic radiation behaves as a particle, could a particle in motion, such as an electron, behave as a wave?
- De Broglie's Equation:
  - $\lambda = hc/E = hc/mc^2 = h/mc$ (or $= h/mv$)
  - λ = De Broglie wavelength, m = mass of electron (in kg), v = velocity (in m/s), and h = Plank's constant.

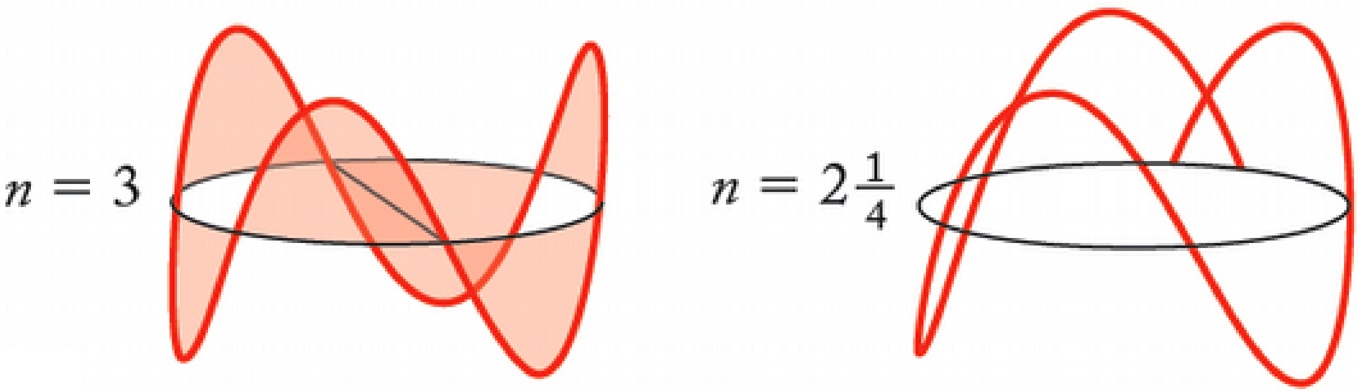
Linear Waves

- Nodes = regions of standing waves that experience no displacement
- Wavelength, $\lambda = 2L$
- Harmonic frequencies:
    - $L = n\lambda/2$
    - For $n = 2\frac{1}{2}$: no nodes, so no standing wave



Electrons as Waves

- Electrons behave like circular waves oscillating around nucleus.
  - No defined stationary ends
- Stable circular waves of circumference =



a) stable b) not stable

Chapter Outline

Electron Wave Equations

- Erwin Schrödinger (1925)
  - Developed mathematical equations to describe behavior of electron waves; became the basis of **quantum mechanics**
- Wave functions (ψ)
  - ψ = describes the motion of electron waves as they vary with location and time
  - ψ^2 defines an **orbital**, or region of high probability for locating an electron

Wave Equations and Quantum Numbers

- **Quantum Numbers:**
  - Principle quantum number (n); indicates the overall energy level (shell). $n = 1, 2, 3,\ldots$
  - Angular momentum quantum number (ℓ); shows number of sublevels (subshells). $\ell = 0 \rightarrow (n - 1)$
  - Magnetic quantum number (m_l); indicates number of orbitals in each sublevel. $m_l = -\ell \rightarrow +\ell$
  - Spin quantum number ();indicates that each orbital can only hold two electrons.

Quantum Numbers

Energy Shells (or Levels)

- 3D regions around the nucleus where electrons may be found
- Shells are distinguished by quantum numbers: $n = 1, 2, 3, 4, \ldots$
- All electrons in a particular shell have similar energies (shells may overlap slightly in energy)

Quantum Numbers

- Each Shell has a maximum # of electrons it can hold
- Lower # shells:
  - Are physically smaller
  - Have lower energy electrons
  - Are usually filled up before higher energy shells

Quantum Numbers

Each shell has at least one sub-shell (or sub-level)

Sub-shells are also labeled by quantum numbers:

l = 0, 1, 2, 3, 4…. (n-1)

but letters are usually used to avoid confusion:

l = s, p, d, f, g, ….

1st shell – only one sub-shell – s

2nd shell – only two sub-shells – s, p

3rd shell – only three sub-shells – s, p, d

4th shell – only four sub-shells – s, p, d, f

…..

Quantum Numbers

Each sub-shell has a fixed number of orbitals (from 3rd quantum number):

s – 1 orbital (spherically shaped)

p – 3 orbitals (lobed shaped, along each axis)

d – 5 orbitals

f – 7 orbitals

g – 9 orbitals

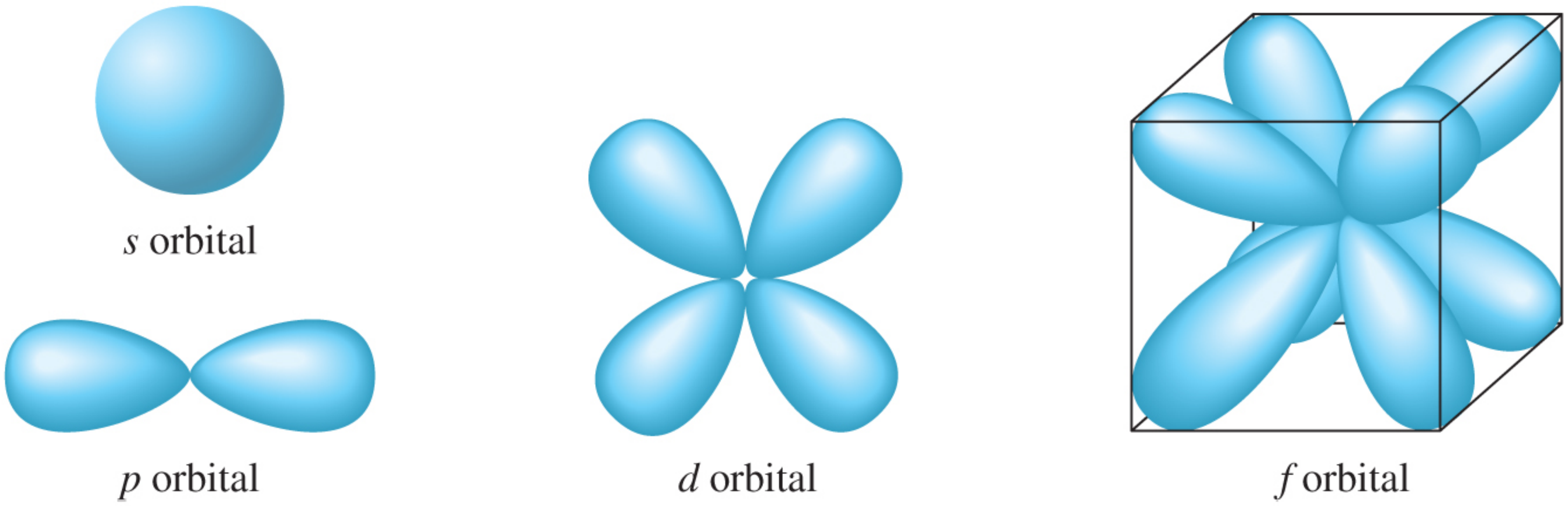
Quantum Numbers

TABLE 3.1 Quantum Numbers of the Orbitals in the First Four Shells

Value of n	Allowed Values of ℓ	Subshell Letter	Allowed Values of m_ℓ	Number of Orbitals in: Subshell	Number of Orbitals in: Shell
1	0	*s*	0	1	1
2	0	*s*	0	1	
	1	*p*	−1, 0, +1	3	4
3	0	*s*	0	1	
	1	*p*	−1, 0, +1	3	
	2	*d*	−2, −1, 0, +1, +2	5	9
4	0	*s*	0	1	
	1	*p*	−1, 0, +1	3	
	2	*d*	−2, −1, 0, +1, +2	5	
	3	*f*	−3, −2, −1, 0, +1, +2, +3	7	16

Chapter Outline

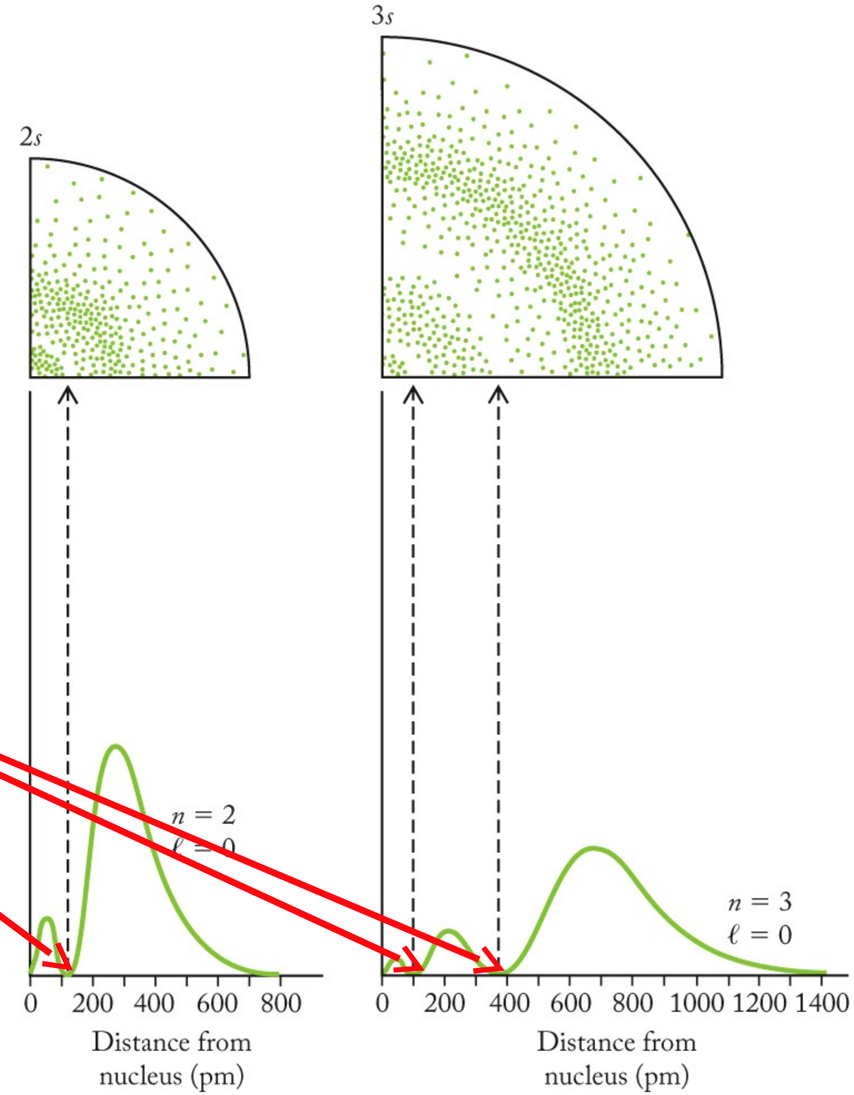
Shapes and Sizes of Orbitals



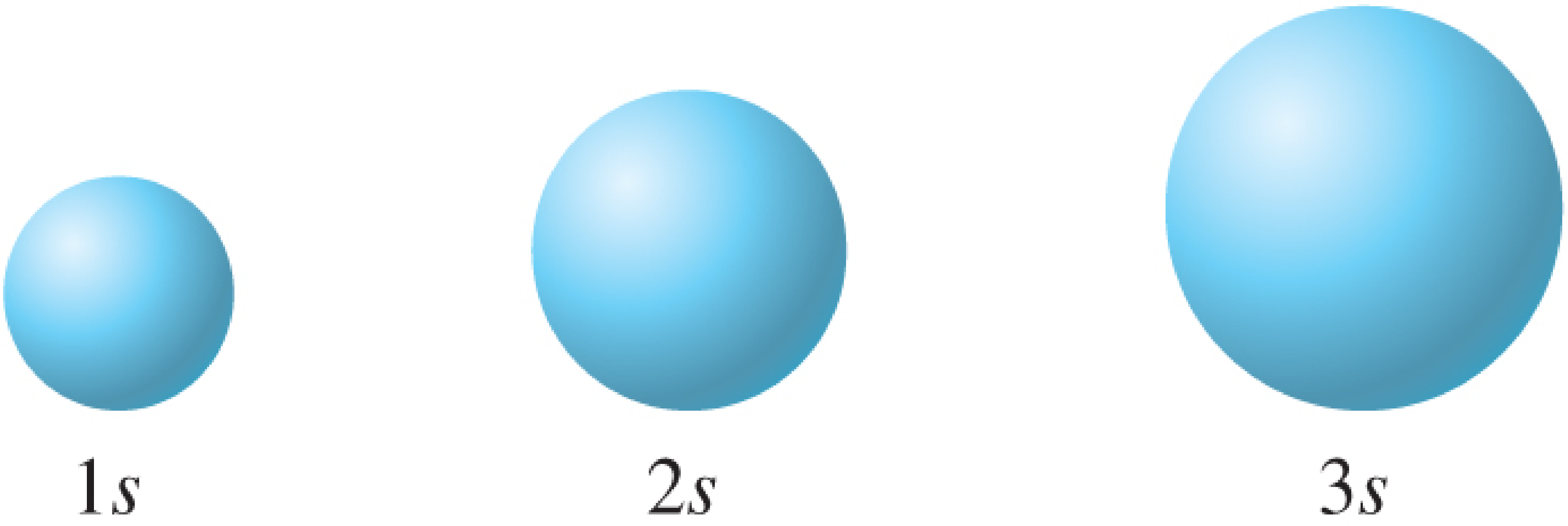
Comparison of *s* Orbital

One *s* orbital in each shell.

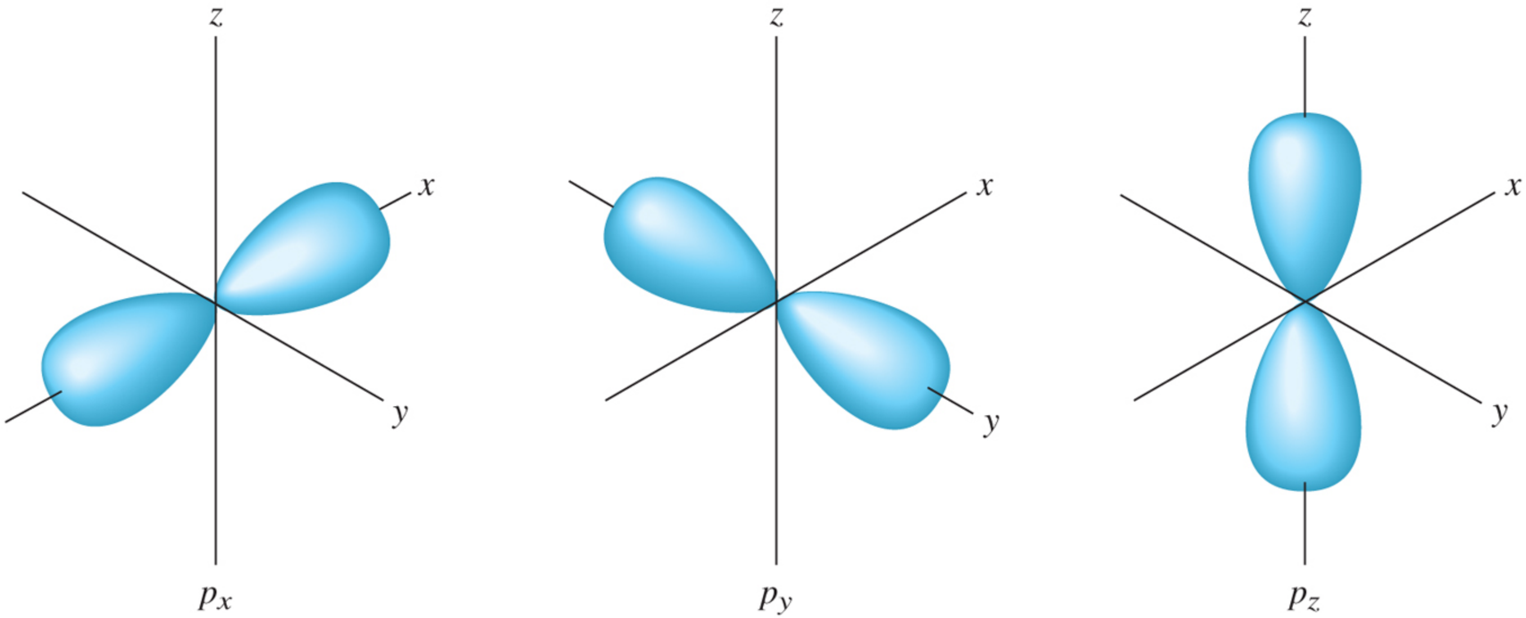
nodes



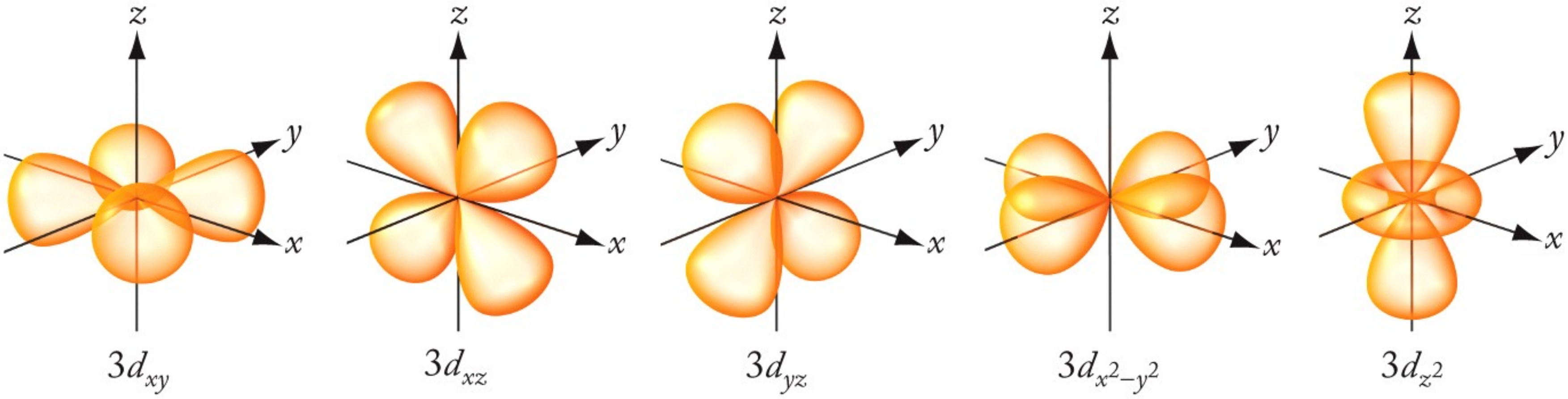
Comparison of *s* Orbital



The Three 2*p* Orbitals



The Five 3*d* Orbitals



Five *d* orbitals in each shell with $n \geq 3$; nodes at nucleus.

Chapter Outline

Aufbau Principle

- Aufbau Principle:
  - Method of building electron configurations by adding one electron at a time as atomic number increases
- When adding electrons to an atom:
  - Electrons always go in lowest energy orbitals available
  - Maximum of two electrons per orbital

Electron Configuration

Each orbital can hold a maximum of two electrons (from 4th quantum number)

- Electrons in an orbital must have opposite spins
  - “Spin up” or “spin down” (+1/2 or -1/2)

- This determines the maximum # of electrons a sub-shell and shell can hold:
  - s – 1 orbital – 2 electrons
  - p – 3 orbitals – 6 electrons
  - d – 5 orbitals – 10 electrons
  - f – 7 orbitals – 14 electrons

Electron Configuration

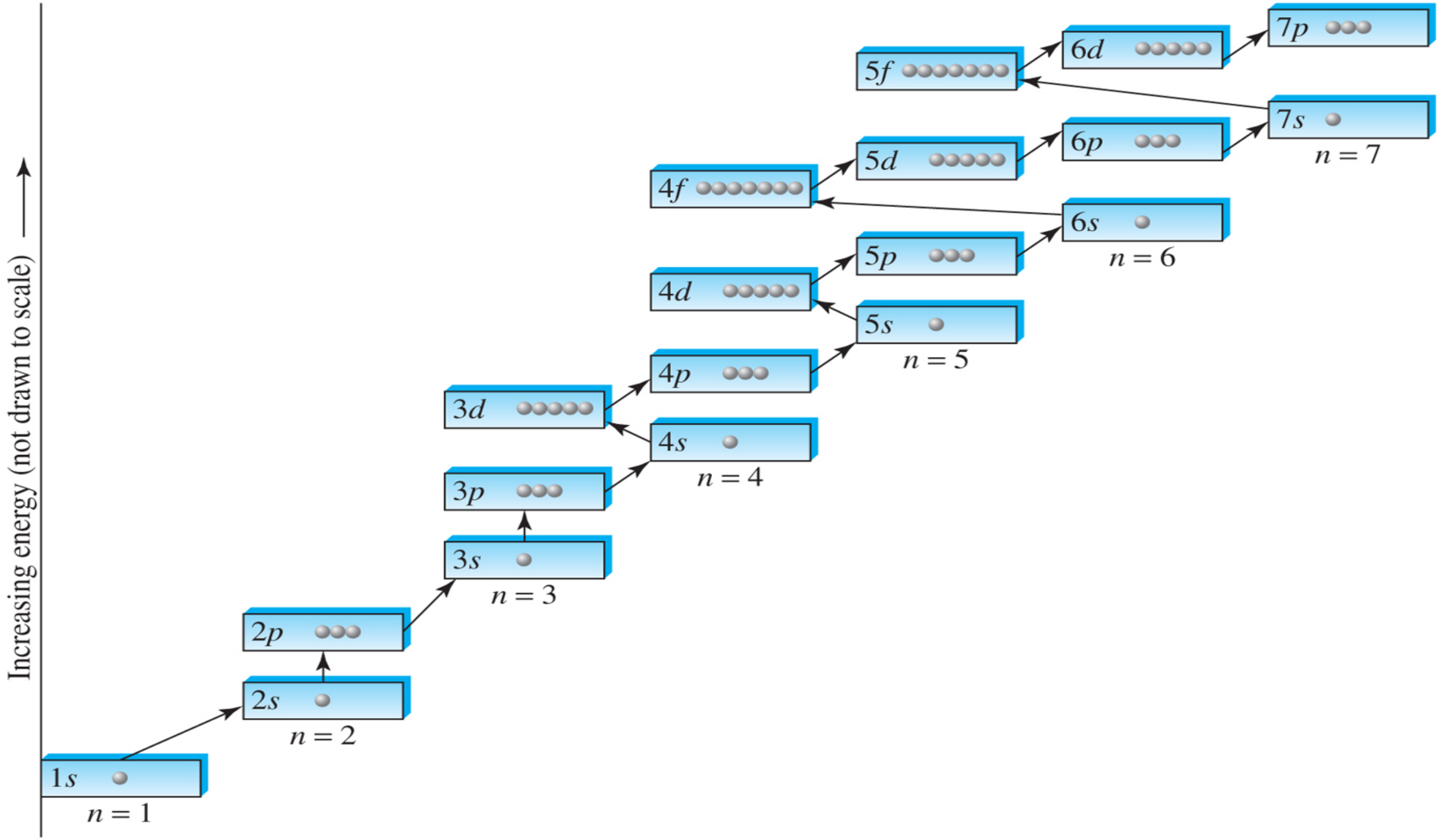
shell	sub-shells	orbitals per subshell	electrons per sub-shell	maximum possible electrons
1	s	1	2	2
2	s p	1 3	2 6	8
3	s p d	1 3 5	2 6 10	18
4	s p d f	1 3 5 7	2 6 10 14	32

Electron Configuration

- Electrons want to occupy the lowest available energy state.

- Some shells overlap, so the lowest state is not always in the lowest shell (level).

Orbital Energies in Multielectron Atoms



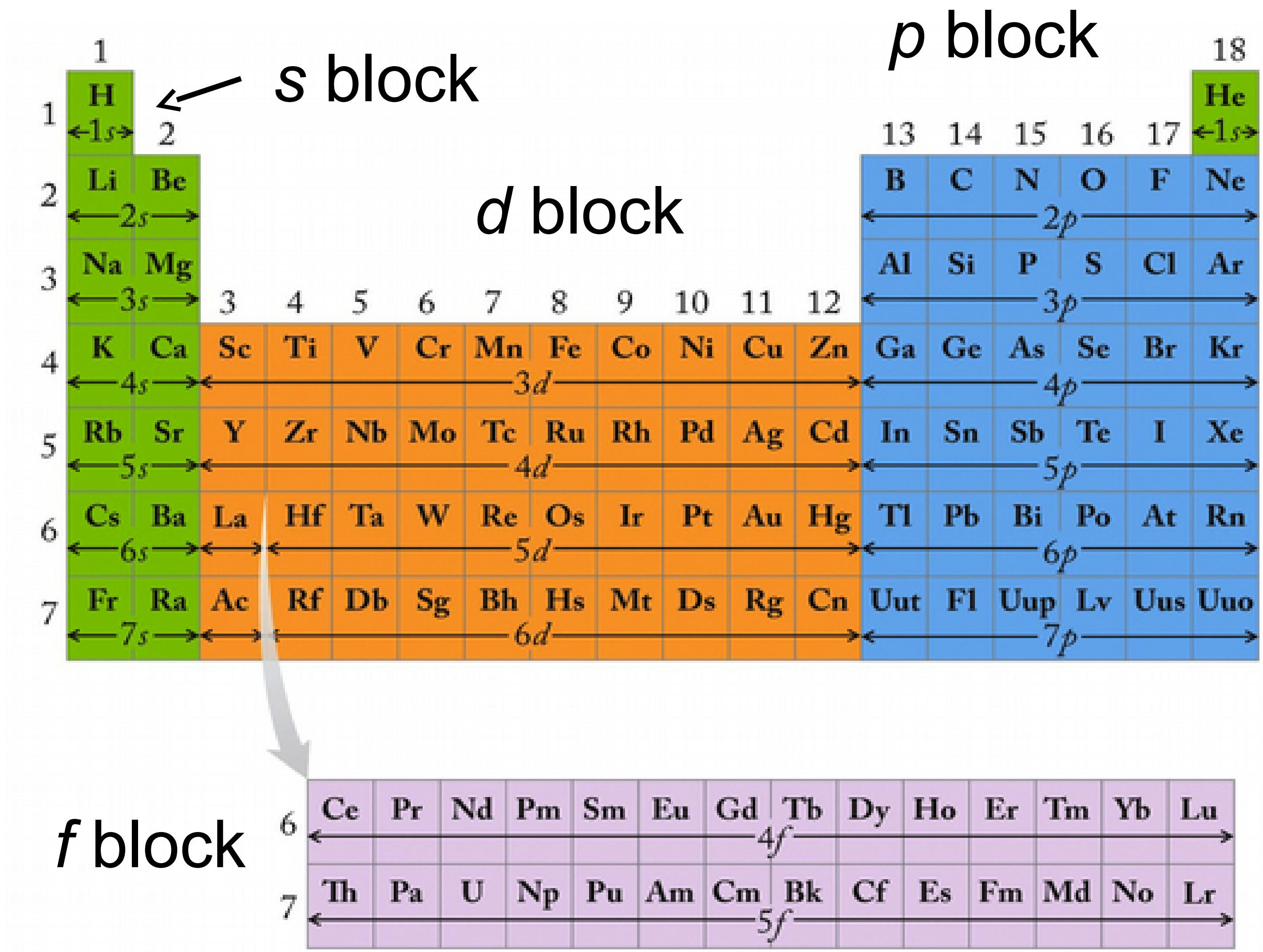
Electron Configuration

- We can use the periodic table to determine which sub-shells should be filled up first.

- The periodic table also shows how many electrons each sub-shell can hold.

- The periodic table can be split into 4 blocks:
  s, p, d, and f blocks

Periodic Table: *s*, *p*, *d*, and *f* Orbitals



Electron Configuration

- Each box can hold one electron, so fill up the boxes, starting in the upper left, until all electrons have been used

- The sub-shells are labeled with the energy level (shell) first, then the sub-shell: e.g. 1s, 2p, 4d....

- Superscripts after each sub-shell tell how many electrons are in that sub-shell: e.g. $1s^2$, $3p^4$,....

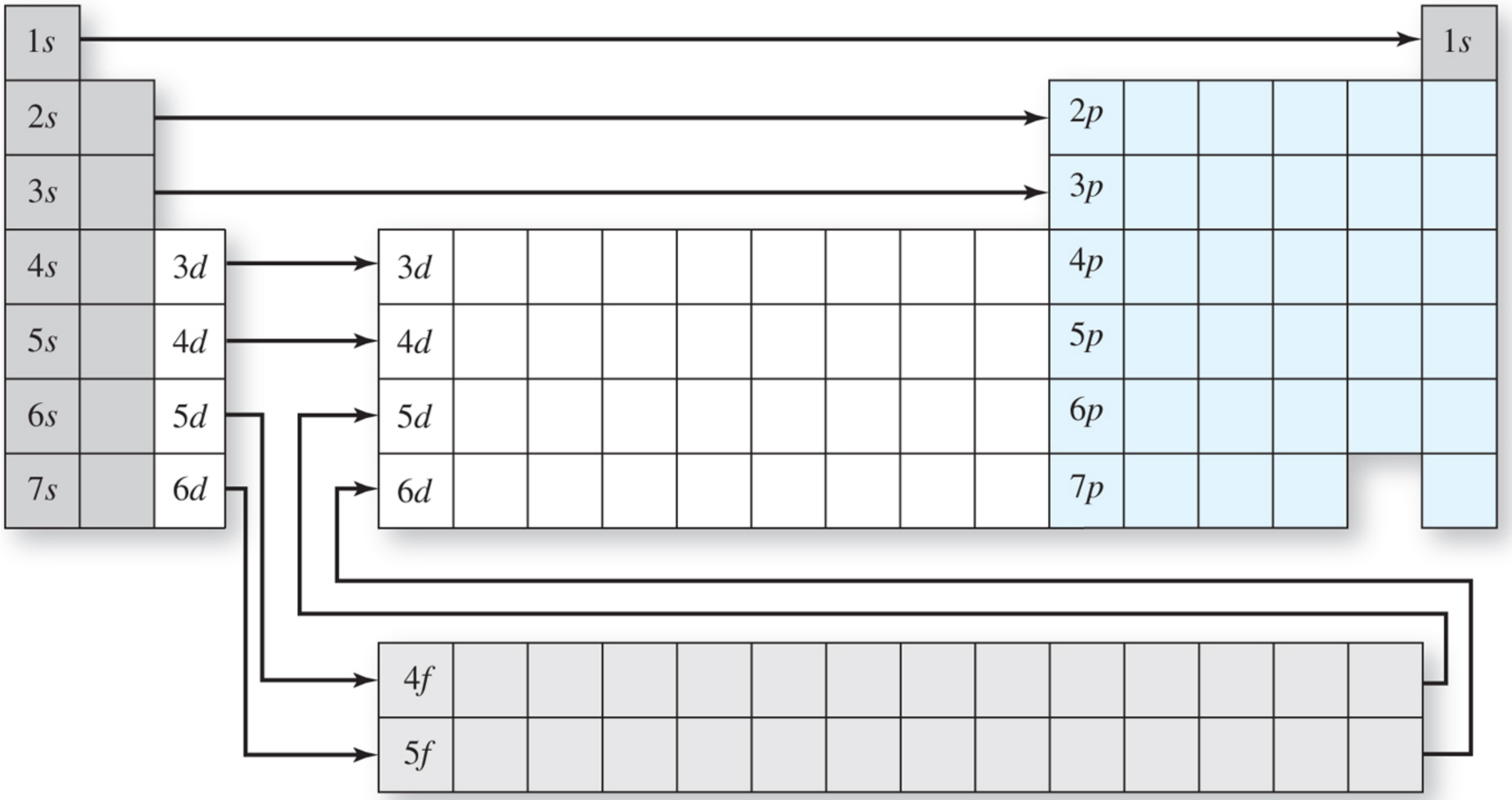
Electron Configuration

- The shell of each s and p sub-shell is the same as the period on the PT, but not so for d and f:

- The simplest way to determine energy shell is:

  - The first s is 1s
  - The first p is 2p
  - The first d is 3d
  - The first f is 4f

Electron Configuration



Electron Configuration

- Electron Configuration
  - Use shells, sub-shells, and superscripts to show where all electrons are positioned
    - Ground state – electrons occupy lowest available states
    - Excited state – electrons are at higher energy states due to added energy

1 Group IA	2 Group IIA	3 Group IIIB	4 Group IVB	5 Group VB	6 Group VIB	7 Group VIIB	8 Group ←	9 Group VIIIB	10 Group →	11 Group IB	12 Group IIB	13 Group IIIA	14 Group IVA	15 Group VA	16 Group VIA	17 Group VIIA	18 Group VIIIA
1 **H** 1.01																	2 **He** 4.00
3 **Li** 6.94	4 **Be** 9.01											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
11 **Na** 22.99	12 **Mg** 24.31											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.07	17 **Cl** 35.45	18 **Ar** 39.95
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.87	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.38	31 **Ga** 69.72	32 **Ge** 72.64	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.96	43 **Tc** (98)	44 **Ru** 101.07	45 **Rh** 102.91	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.76	52 **Te** 127.60	53 **I** 126.90	54 **Xe** 131.29
55 **Cs** 132.91	56 **Ba** 137.33	57 **La** 138.91	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.84	75 **Re** 186.21	76 **Os** 190.23	77 **Ir** 192.22	78 **Pt** 195.08	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.38	82 **Pb** 207.2	83 **Bi** 208.98	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** (226)	89 **Ac** (227)	104 **Rf** (267)	105 **Db** (268)	106 **Sg** (271)	107 **Bh** (272)	108 **Hs** (270)	109 **Mt** (276)	110 **Ds** (281)	111 **Rg** (280)	112 **–** (285)	113 **–** (284)	114 **–** (289)	115 **–** (288)	116 **–** (293)		118 **–** (294)

Metals ← → Non-metals

58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.36	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** 173.05	71 **Lu** 174.97
90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (262)

Electron Configuration

- There are exceptions to expected config.
  - Electrons may shift from one subshell to another subshell of similar energy
    - Fully filled or half filled subshells may be preferable to other fractions.

Electron Configuration

- Noble Gas configuration
  - Shorthand – only shows outer electrons
  - Brackets around Noble Gas represents the <u>electron configuration</u> of that gas
  - e.g. [Ar] = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Orbital Diagrams

- Use boxes, lines, or circles to show individual orbitals
- Usually only orbitals of outer sub-shells are shown
- In a given sub-shell:
  - Place one electron per orbital before pairing
  - All single electrons must have same spin
  - Paired electrons must have opposite spin

Orbital Diagrams for Hydrogen and Helium

	Orbital diagram			Electron configuration	Condensed electron configuration
	1*s*	2*s*	2*p*		
H	↑			$1s^1$	
He	↑↓			$1s^2$	

(filled *n* = 1 shell!)

Orbital Diagrams for Multielectron Atoms

	Orbital diagram					Electron configuration	Condensed electron configuration
	1*s*	2*s*	2*p*				
Li	↑↓	↑				$1s^2 2s^1$	$[He]2s^1$
Be	↑↓	↑↓				$1s^2 2s^2$	$[He]2s^2$
B	↑↓	↑↓	↑			$1s^2 2s^2 2p^1$	$[He]2s^2 2p^1$
C	↑↓	↑↓	↑	↑		$1s^2 2s^2 2p^2$	$[He]2s^2 2p^2$
N	↑↓	↑↓	↑	↑	↑	$1s^2 2s^2 2p^3$	$[He]2s^2 2p^3$
O	↑↓	↑↓	↑↓	↑	↑	$1s^2 2s^2 2p^4$	$[He]2s^2 2p^4$
F	↑↓	↑↓	↑↓	↑↓	↑	$1s^2 2s^2 2p^5$	$[He]2s^2 2p^5$
Ne	↑↓	↑↓	↑↓	↑↓	↑↓	$1s^2 2s^2 2p^6$	$[He]2s^2 2p^6 = [Ne]$

Chapter Outline

Electron Configuration of Ions

- Formation of Ions:
  - Gain/loss of valence electrons to achieve stable electron configuration (filled shell)
  - Cations:
    - $Na(g) \rightarrow Na^{+}(g) + e^{-}$
    - $[He]3s^{1} \rightarrow [He] + e^{-}$
  - Anions:
    - $Cl(g) + e^{-} \rightarrow Cl^{-}(g)$
    - $[Ne]3s^{2}3p^{5} + e^{-} \rightarrow [Ne]3s^{2}3p^{6} = [Ar]$

Isoelectronic Atoms/Ions

- Main Group Elements:
  - Form ions by gain/loss of e^- to obtain noble gas configuration:
  - $Mg \rightarrow Mg^{2+} + 2e^-$ = [Ar]
  - $O + 2e^- \rightarrow O^{2-}$ = [Ne]
- **Isoelectronic**: Describes atoms/ions having identical electron configurations
  - Na^+, Mg^{2+}, O^{2-}, F^-, Ne = $1s^22s^22p^6$ (= [Ne])
  - K^+, Cl^-, Ca^{2+}, Ar = $1s^22s^22p^63s^23p^6$ (= [Ar])

Cations of Transition Metals

- Transition Metal Cations:
  - Loss of valence electrons (*s*) and, in some cases, *d* electrons:
  - Fe: [Ar]$3d^6\,4s^2$
  - Fe^{2+}: [Ar]$3d^6 4s^2$ = [Ar]$3d^6$ (loss of valence e^-)
  - Fe^{3+}: [Ar]$3d^6$ = [Ar]$3d^5$ (loss of one 3*d* = half-filled *d* subshell!)

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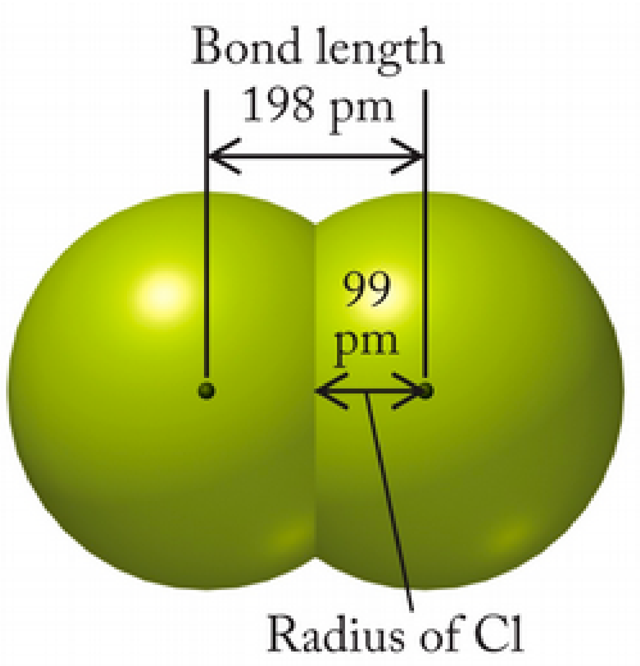
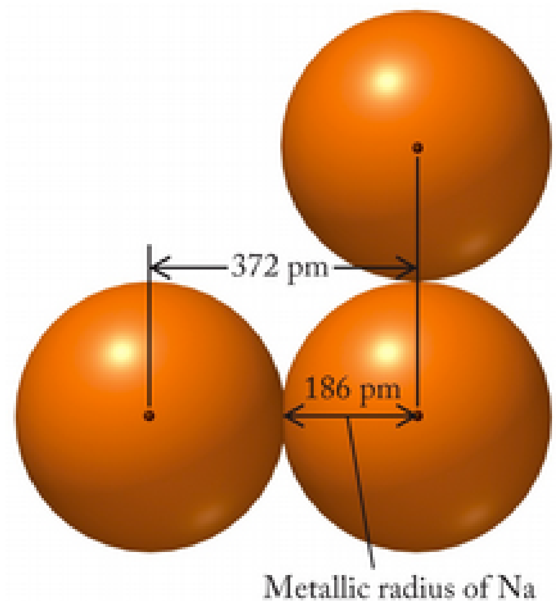
Sizes of Atoms/Ions

- **Atomic Radius** (i.e., covalent radius)
  - Half the distance between identical nuclear centers in a molecule
- **Metallic Radius**
  - half the distance between nuclear centers in the crystal of a metal
- **Ionic Radius**
  - Derived from the distance between nuclear centers in ionic crystals

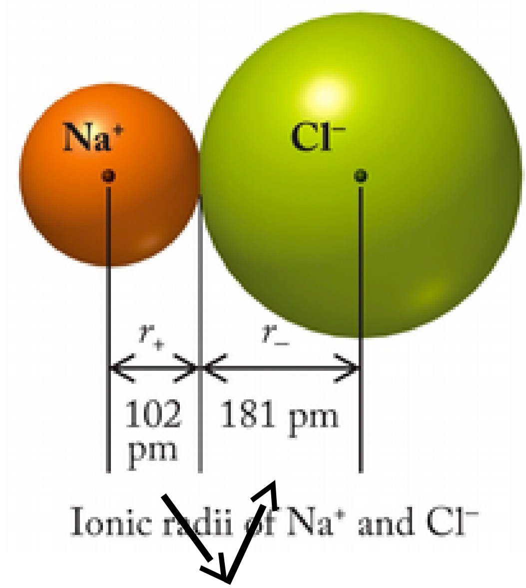
Atomic, Metallic, Ionic Radii



Covalent radius of Cl.



Metallic radius of Na.



Ionic radii of Na^+, Cl^-.

Atomic Radius Trends

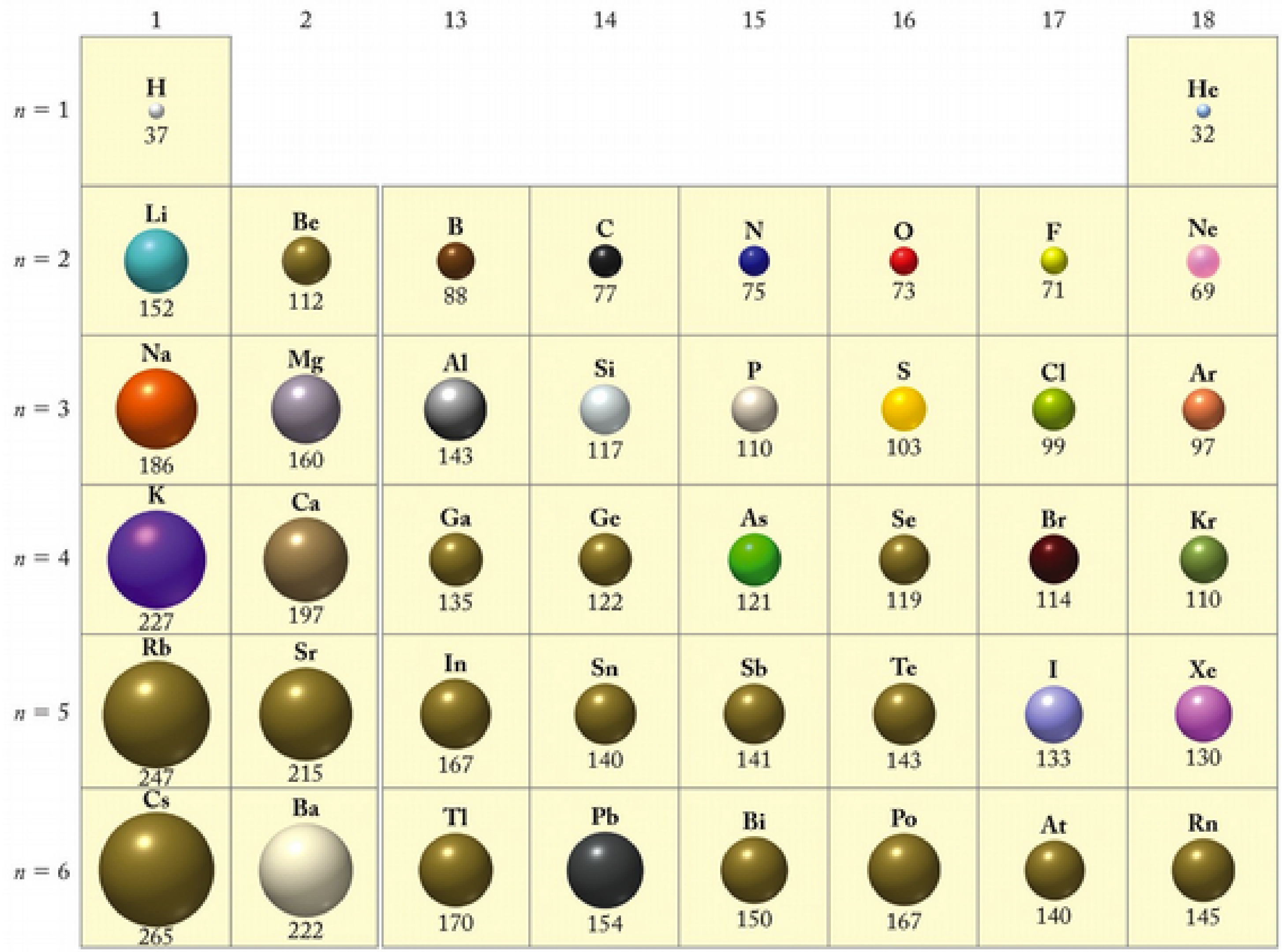
- Atomic Radii:
  - Increase going down a family
    - "Shielding" by inner shell electrons decreases effective nuclear charge (z_{eff})
  - Decrease going across a row
    - Increased nuclear charge (z) moving across row
    - Increased attraction for electrons in inner orbitals → atomic size decreases

1 Group IA	2 Group IIA	3 Group IIIB	4 Group IVB	5 Group VB	6 Group VIB	7 Group VIIB	8 Group ←	9 Group VIIIB	10 Group →	11 Group IB	12 Group IIB	13 Group IIIA	14 Group IVA	15 Group VA	16 Group VIA	17 Group VIIA	18 Group VIIIA
1 **H** 1.01																	2 **He** 4.00
3 **Li** 6.94	4 **Be** 9.01											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
11 **Na** 22.99	12 **Mg** 24.31											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.07	17 **Cl** 35.45	18 **Ar** 39.95
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.87	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.38	31 **Ga** 69.72	32 **Ge** 72.64	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.96	43 **Tc** (98)	44 **Ru** 101.07	45 **Rh** 102.91	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.76	52 **Te** 127.60	53 **I** 126.90	54 **Xe** 131.29
55 **Cs** 132.91	56 **Ba** 137.33	57 **La** 138.91	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.84	75 **Re** 186.21	76 **Os** 190.23	77 **Ir** 192.22	78 **Pt** 195.08	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.38	82 **Pb** 207.2	83 **Bi** 208.98	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** (226)	89 **Ac** (227)	104 **Rf** (267)	105 **Db** (268)	106 **Sg** (271)	107 **Bh** (272)	108 **Hs** (270)	109 **Mt** (276)	110 **Ds** (281)	111 **Rg** (280)	112 **–** (285)	113 **–** (284)	114 **–** (289)	115 **–** (288)	116 **–** (293)		118 **–** (294)

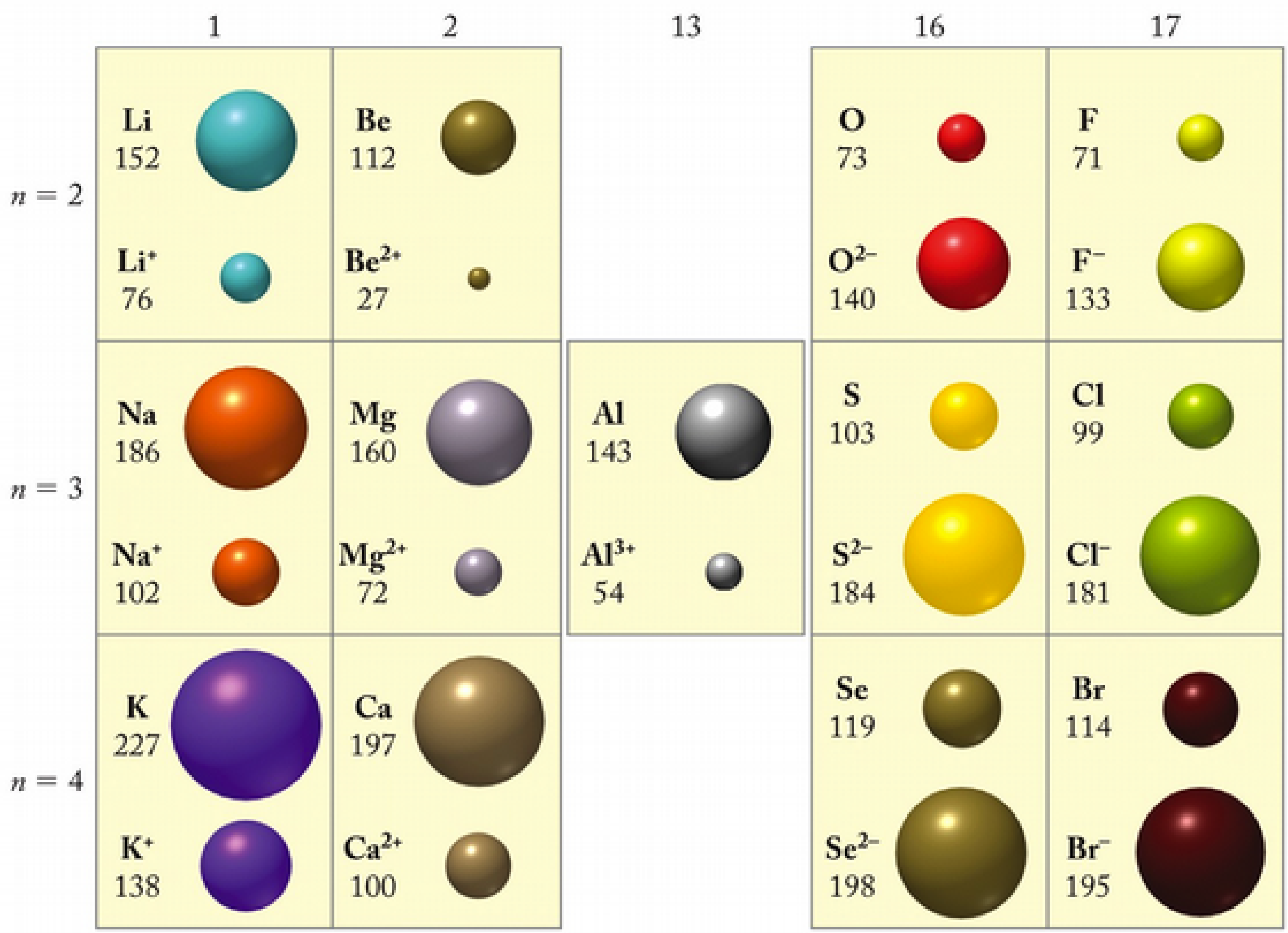
Metals ← → Non-metals

58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.36	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** 173.05	71 **Lu** 174.97
90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (262)

Atomic Radii (pm)



Radii of Atoms and Ions



Chapter Outline

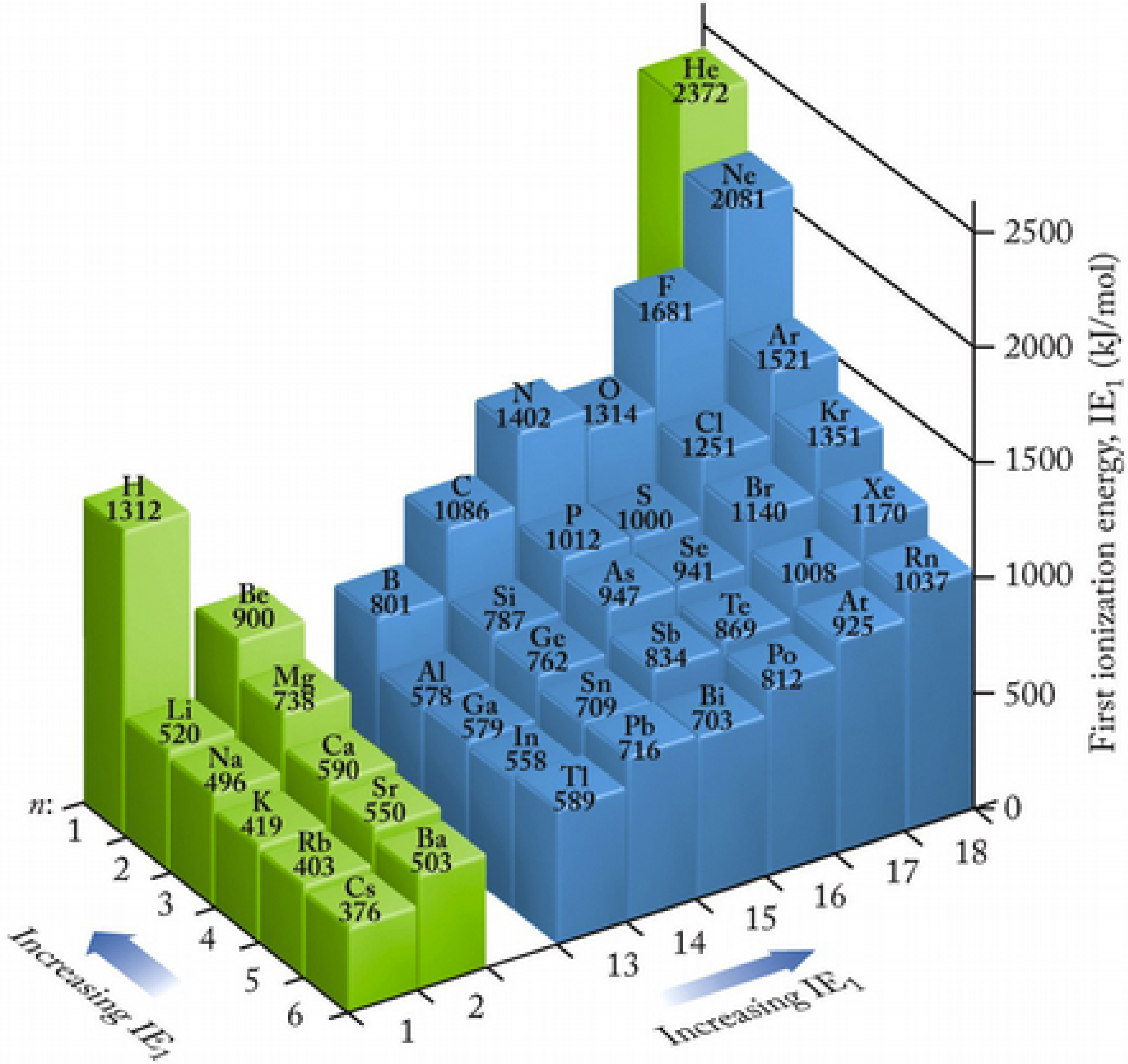
Ionization Energy (IE)

- **Ionization Energy:**
  - Amount of energy needed to remove 1 mole of e^- from 1 mole of the ground-state atoms or ions in the gas phase (kJ/mole)
    - $X(g) \rightarrow X^+(g) + e^-(g)$
  - 1st Ionization Energy (IE_1): $Mg(g) \rightarrow Mg^+(g) + e^-$
  - 2nd Ionization Energy (IE_2): $Mg^+(g) \rightarrow Mg^{2+}(g) + e^-$
  - Note: $IE_2 > IE_1$

Ionization Energy (IE)

- **Periodic Trend:**
  - Ionization energy increases going up a group
  - Ionization energy tends to increase going across a period
  - Note: Lower ionization energy indicates that the atom more easily forms a cation

First Ionization Energies



Successive Ionization Energies (kJ/mol)

TABLE 3.2 Successive Ionization Energies[a] of the First 10 Elements

Element	Z	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7	IE_8	IE_9	IE_{10}
H	1	1312									
He	2	2372	5249								
Li	3	520	7296	12040							
Be	4	900	1758	15050	21070						
B	5	801	2426	3660	24682	32508					
C	6	1086	2348	4617	6201	37926	46956				
N	7	1402	2860	4581	7465	9391	52976	64414			
O	8	1314	3383	5298	7465	10956	13304	71036	84280		
F	9	1681	3371	6020	8428	11017	15170	17879	92106	106554	
Ne	10	2081	3949	6140	9391	12160	15231	19986	23057	115584	131236

[a] In kJ/mol.

Chapter Outline

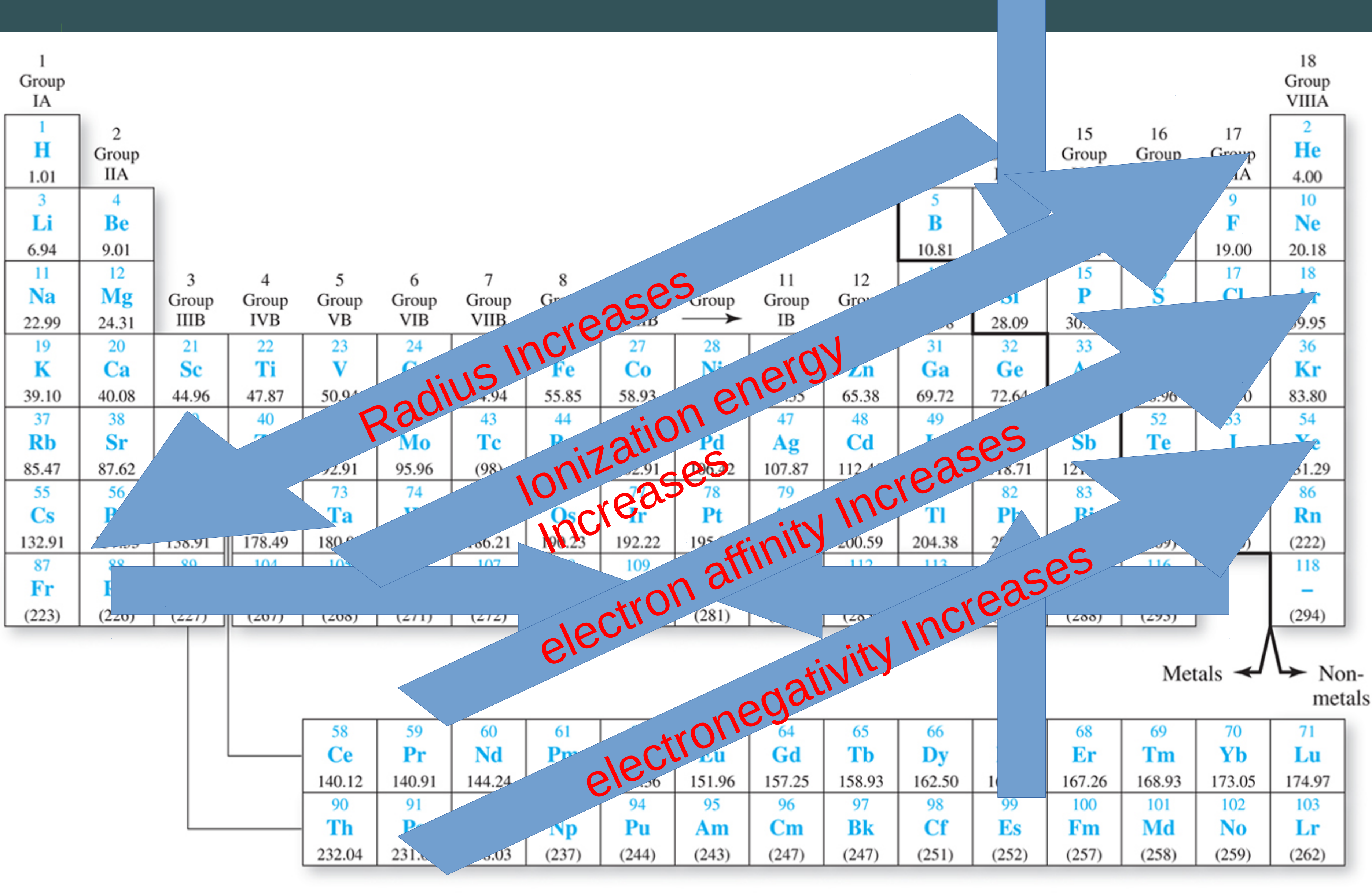
Electron Affinities (EA)

- **Electron Affinity**:
  - Energy change that occurs when 1 mole of electrons combine with 1 mole of atoms or ions in the gas phase
  - $Cl(g) + e^- \rightarrow Cl^-(g) \quad EA_1 = –349$ kJ/mol
- Periodic Trends
  - EA values become more negative moving to the right and up in the periodic table
    - A more negative EA means a stronger attraction for electrons

Periodic Trends in EA

1	2	13	14	15	16	17	18
H −72.6							He (0.0)[a]
Li −59.6	Be >0	B −26.7	C −122	N +7	O −141	F −328	Ne (+29)[a]
Na −52.9	Mg >0	Al −42.5	Si −134	P −72.0	S −200	Cl −349	Ar (+35)[a]
K −48.4	Ca −2.4	Ga −28.9	Ge −119	As −78.2	Se −195	Br −325	Kr (+39)[a]
Rb −46.9	Sr −5.0	In −28.9	Sn −107	Sb −103	Te −190	I −295	Xe (+41)[a]
Cs −45.5	Ba −14	Tl −19.2	Pb −35.2	Bi −91.3	Po −183.3	At −270[a]	Rn (+41)[a]

[a]Calculated values.


Radius Increases
Ionization energy Increases
electron affinity Increases
electronegativity Increases
Metals
Non-metals

This concludes the Lecture PowerPoint presentation for Chapter 3

CHEMISTRY
an atoms-focused approach

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