

Chapter 7

Atomic Structure and Periodicity

Chapter 7: Atomic Structure and Periodicity

Objectives

- Describe the evidence for particle-wave duality.
- Describe the properties of electromagnetic radiation.
- Explain the relationship between energy, frequency, and wavelength.
- Describe the origin of light emitted by excited atoms and its relationship to atomic structure.
- Define the Bohr atomic model and explain how it is flawed.
- Identify the principles of the quantum mechanical model of the atom.
- Define the four quantum numbers(n , l , m_l , and m_s) and recognize their relationship to electronic structure.
- Write the electron configuration for atoms and monatomic ions.
- Explain trends in atom and ion sizes, ionization energy, and chemical properties.
- Describe the difference between ionic and covalent bonds.

Chapter 7: Atomic Structure and Periodicity

Table of Contents

- The Nature of Light
 - Properties of Waves and Light
 - Electromagnetic Spectrum
 - Plank's Quantum Theory
 - Photoelectric Effect
- Bohr's Theory
 - Emission Spectrum of the Hydrogen Atom
 - The Bohr Model
 - The de Broglie Wavelength
 - The Uncertainty Principle
- Quantum Mechanical Description of the Atomic Orbital
 - Description of the Hydrogen Atom
 - Indeterminacy and Probability Distribution Maps
 - Quantum Numbers
 - The Pauli Exclusion Principle
 - Wave Equation for the Hydrogen Atom
- Orbital Shapes
 - s, p, d, and f Orbitals
 - The Phase of Orbitals
- History of the Periodic Table
 - Development of the Periodic Table
 - Periods 1 Through 3
 - Transition Metals
 - Periodic Table Position and Electron Configuration
- Electron Configuration of Cations and Anions
- Electron Configuration
 - Rules for Assigning Electrons to Atomic Orbitals
 - The Aufbau Principle
 - Hund's Rule
 - Shielding Effect and Effective Nuclear Charge
 - Diamagnetism and Paramagnetism
- Periodic Trends
 - Atomic Radius
 - Ionic Radius
 - Variation of Physical Properties
 - Ionization Energy
 - Electron Affinity
 - Electron Configuration and Magnetic Properties of Ions
- Variation in Chemical Properties
 - General Trends
 - The Halogens
 - The Noble Gases

Section 7.1

Electromagnetic Radiation

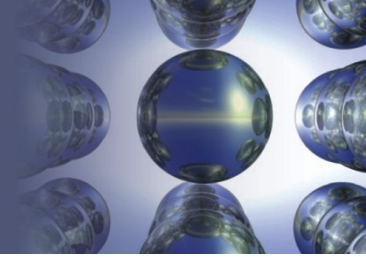
Different Colored Fireworks



PhotoDisc/Getty Images

Section 7.1

Electromagnetic Radiation

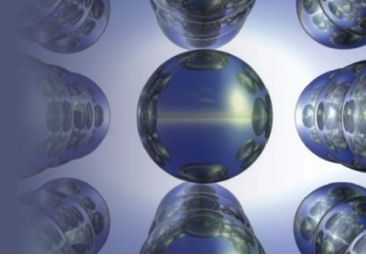


Questions to Consider

- Why do we get colors?
- Why do different chemicals give us different colors?

Section 7.1

Electromagnetic Radiation

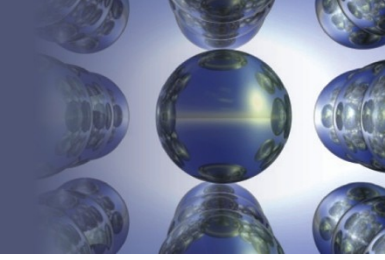


Electromagnetic Radiation

- One of the ways that energy travels through space.
- Three characteristics:
 - Wavelength
 - Frequency
 - Speed

Section 7.1

Electromagnetic Radiation



Characteristics

- Wavelength (λ) – distance between two consecutive peaks or troughs in a wave.
- Frequency (ν) – number of waves (cycles) per second that pass a given point in space
- Speed (c) – speed of light (2.9979×10^8 m/s)

$$c = \lambda \nu$$

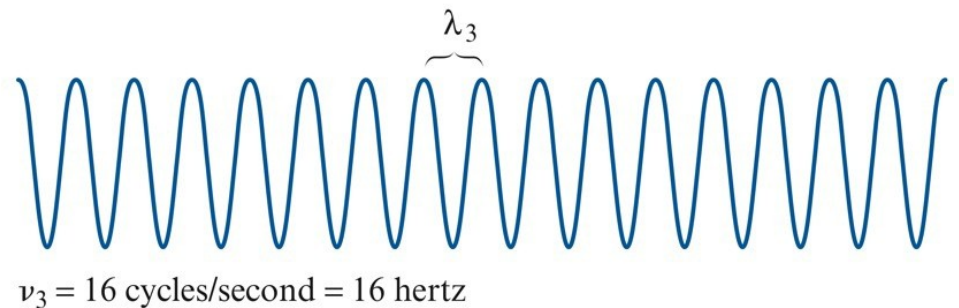
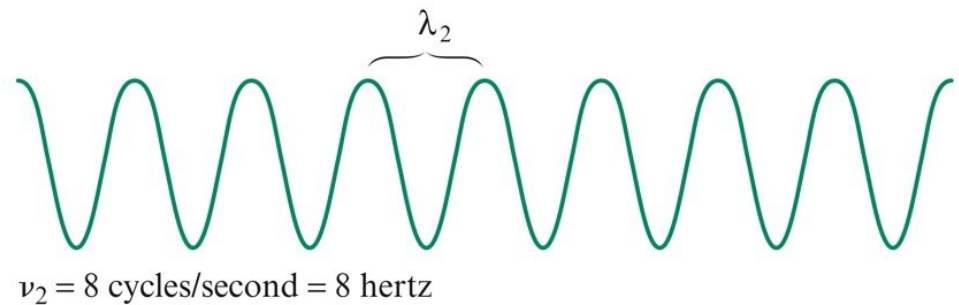
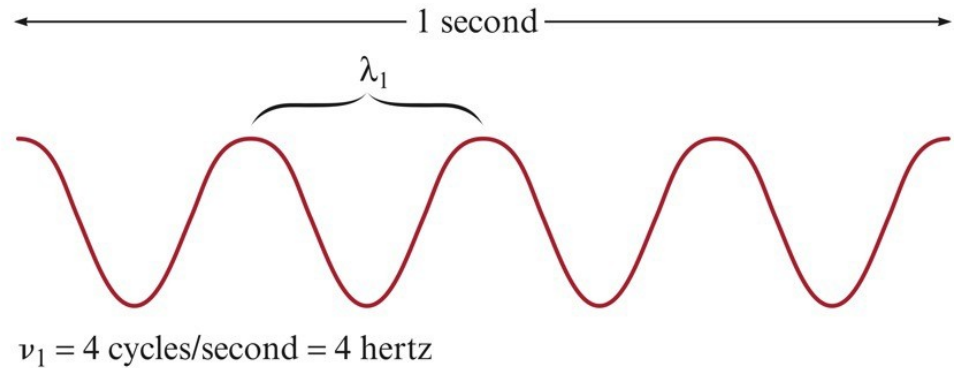
Section 7.1

Electromagnetic Radiation



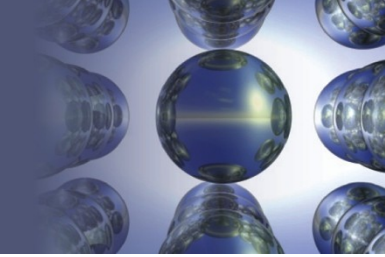
Kertlis/iStockphoto.com

The Nature of Waves

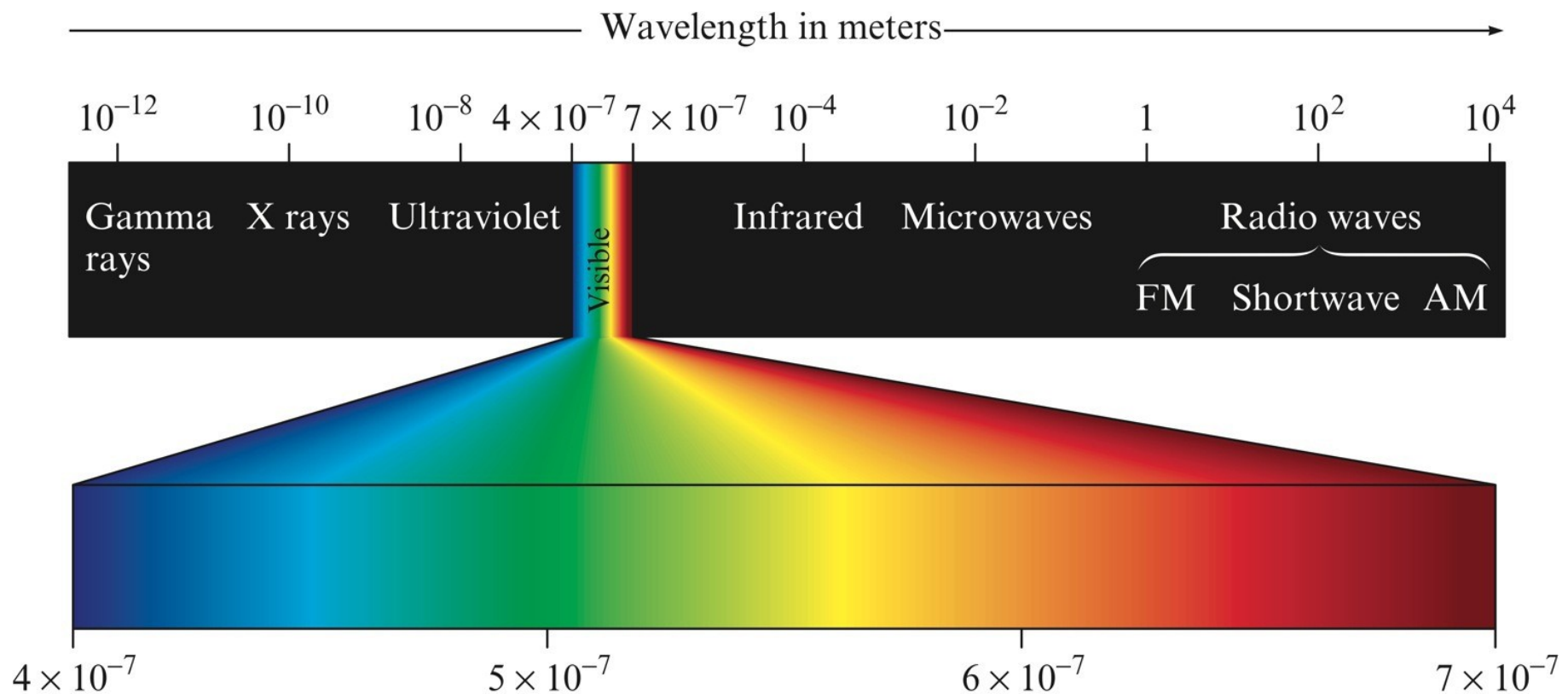


Section 7.1

Electromagnetic Radiation



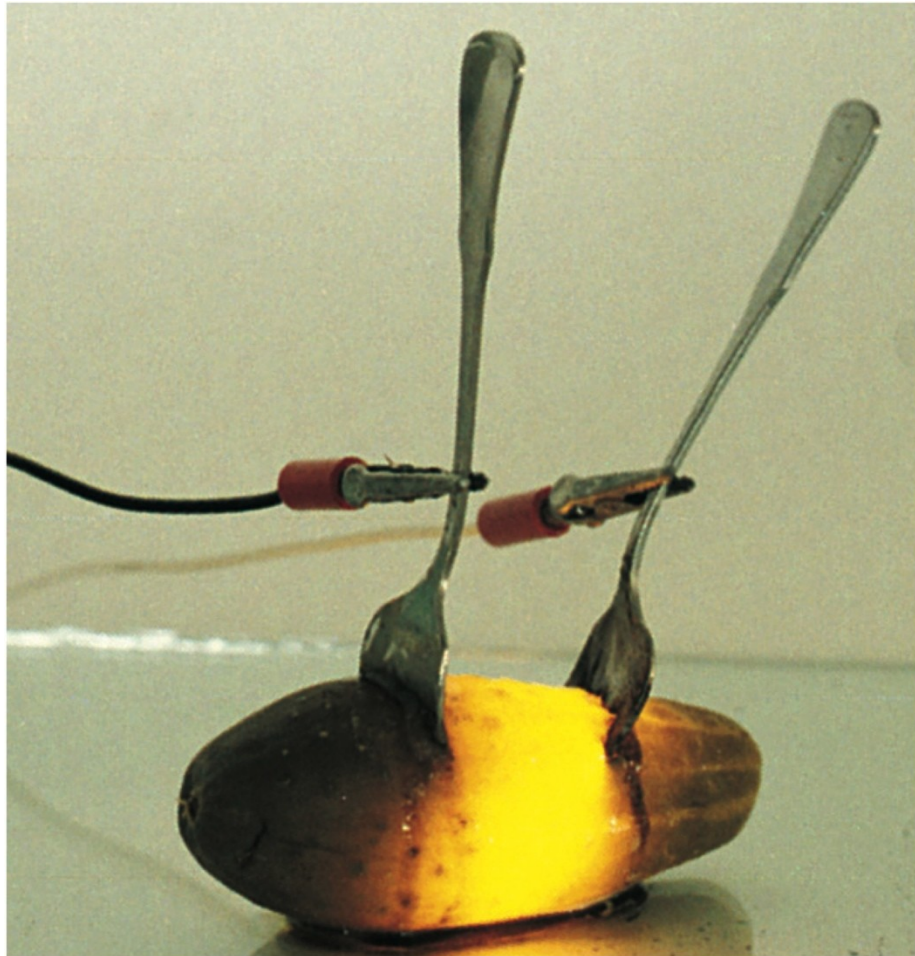
Classification of Electromagnetic Radiation



Section 7.2

The Nature of Matter

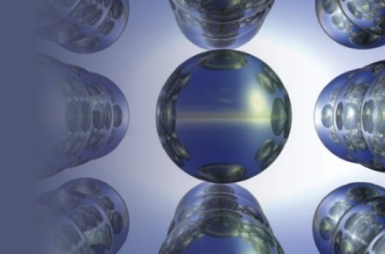
Pickle Light (starts at 2min 30sec)



Donald W. Clegg

Section 7.2

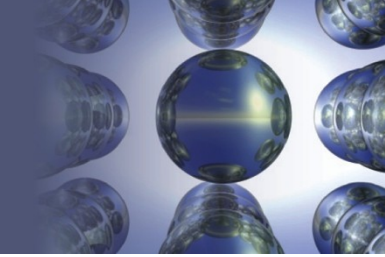
The Nature of Matter



- Energy can be gained or lost only in whole number multiples of $h\nu$.
- A system can transfer energy only in whole quanta (or “packets”).
- Energy seems to have particulate properties too.

Section 7.2

The Nature of Matter



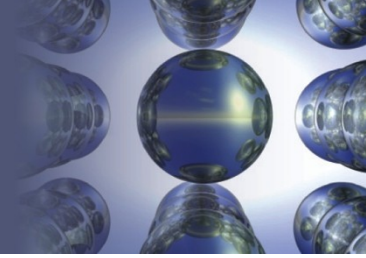
- Energy is quantized.
- Electromagnetic radiation is a stream of “particles” called photons.

$$E_{\text{photon}} = h\nu = \frac{hc}{\lambda}$$

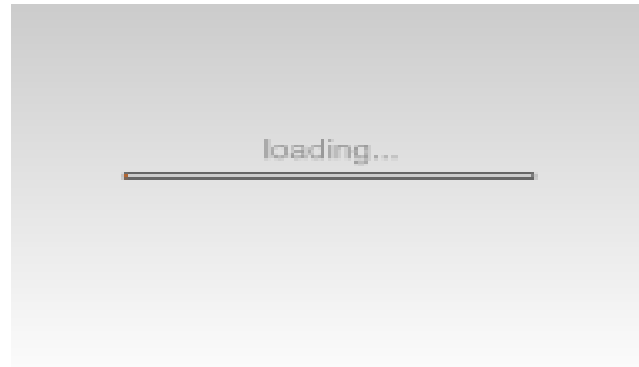
- Planck's constant = $h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$

Section 7.2

The Nature of Matter



The Photoelectric effect



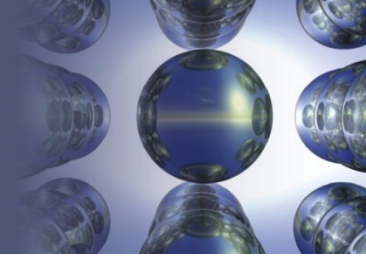
To play movie you must be in Slide Show Mode

PC Users: Please wait for content to load, then click to play

Mac Users: [CLICK HERE](#)

Section 7.2

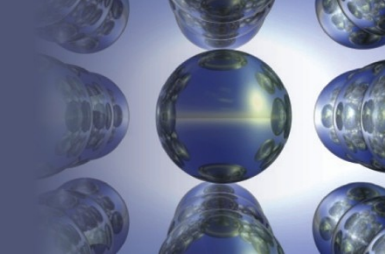
The Nature of Matter



- Energy has mass $E = mc^2$
- Dual nature of light:
 - Electromagnetic radiation (and all matter) exhibits wave properties and particulate properties.

Section 7.3

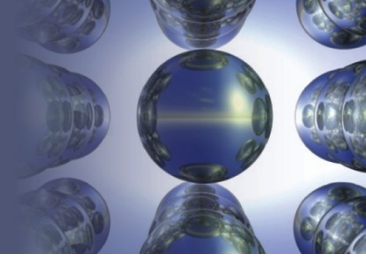
The Atomic Spectrum of Hydrogen



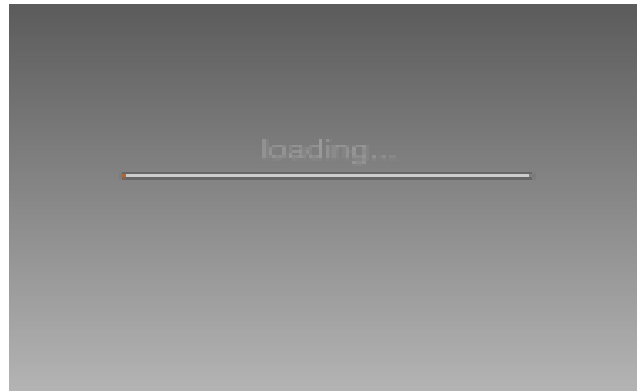
- Continuous spectrum (results when white light is passed through a prism) – contains all the wavelengths of visible light
- Line spectrum – each line corresponds to a discrete wavelength:
 - Hydrogen emission spectrum

Section 7.3

The Atomic Spectrum of Hydrogen



Refraction of White Light



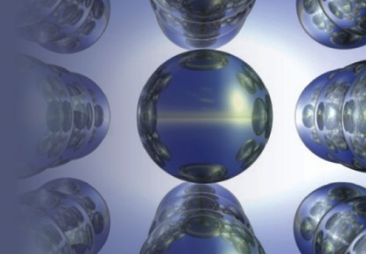
To play movie you must be in Slide Show Mode

PC Users: Please wait for content to load, then click to play

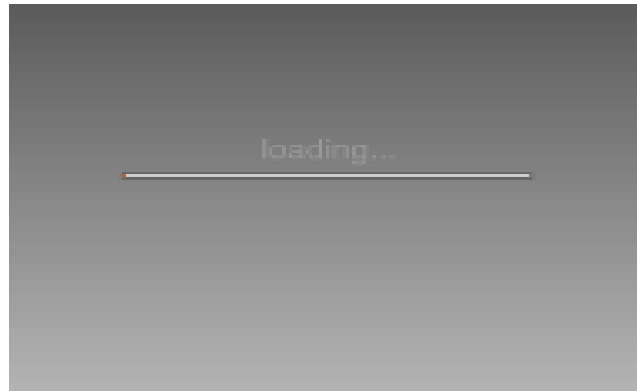
Mac Users: [CLICK HERE](#)

Section 7.3

The Atomic Spectrum of Hydrogen



The Line Spectrum of Hydrogen



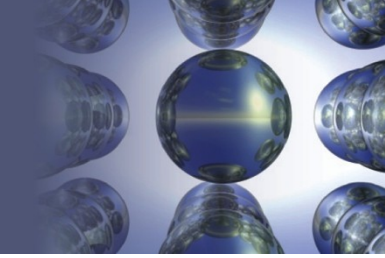
To play movie you must be in Slide Show Mode

PC Users: Please wait for content to load, then click to play

Mac Users: [CLICK HERE](#)

Section 7.3

The Atomic Spectrum of Hydrogen

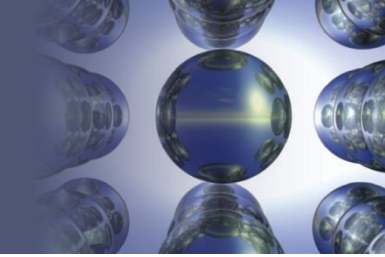


Significance

- Only certain energies are allowed for the electron in the hydrogen atom.
- Energy of the electron in the hydrogen atom is quantized.

Section 7.3

The Atomic Spectrum of Hydrogen



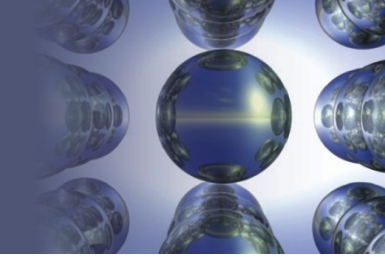
CONCEPT CHECK!

Why is it significant that the color emitted from the hydrogen emission spectrum is **not white**?

How does the emission spectrum support the idea of **quantized** energy levels?

Section 7.4

The Bohr Model



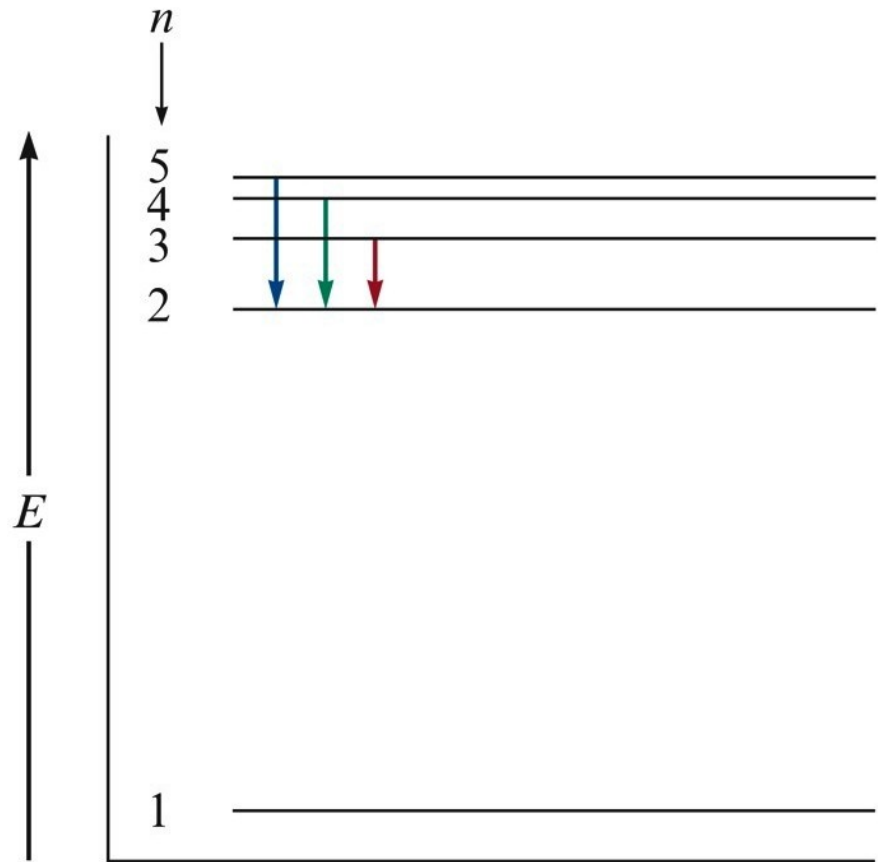
- Electron in a hydrogen atom moves around the nucleus only in certain allowed circular orbits.
- Bohr's model gave hydrogen atom energy levels consistent with the hydrogen emission spectrum.
- Ground state – lowest possible energy state ($n = 1$)

Section 7.4

The Bohr Model

Electronic Transitions in the Bohr Model for the Hydrogen Atom

a) An Energy-Level Diagram for Electronic Transitions

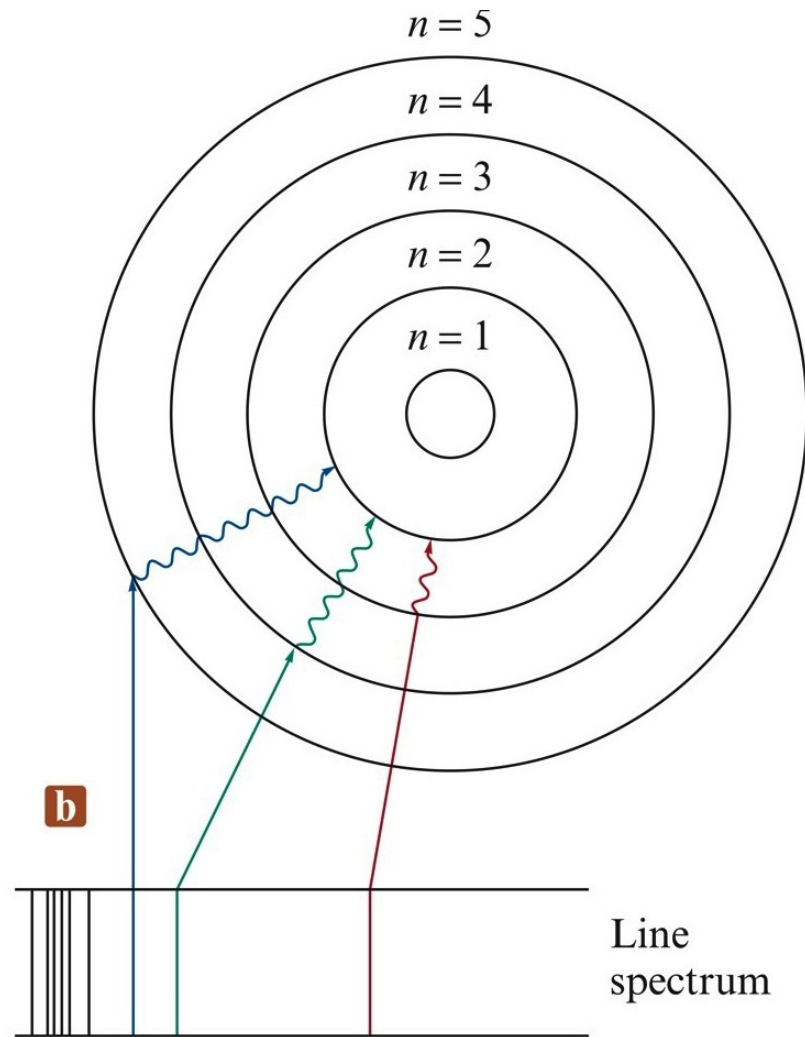


Section 7.4

The Bohr Model

Electronic Transitions in the Bohr Model for the Hydrogen Atom

b) An Orbit-Transition Diagram, Which Accounts for the Experimental Spectrum



Section 7.4

The Bohr Model

- For a single electron transition from one energy level to another:

$$\Delta E = -2.178 \times 10^{-18} \text{ J} \left(\frac{1}{n_{\text{final}}^2} - \frac{1}{n_{\text{initial}}^2} \right)$$

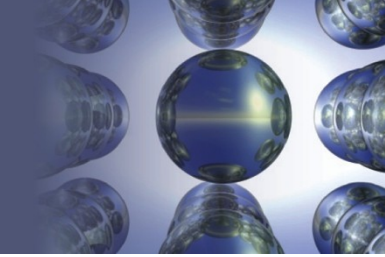
ΔE = change in energy of the atom (energy of the emitted photon)

n_{final} = integer; final distance from the nucleus

n_{initial} = integer; initial distance from the nucleus

Section 7.4

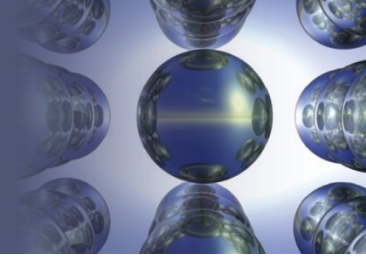
The Bohr Model



- The model correctly fits the quantized energy levels of the hydrogen atom and postulates only certain allowed circular orbits for the electron.
- As the electron becomes more tightly bound, its energy becomes more negative relative to the zero-energy reference state (free electron). As the electron is brought closer to the nucleus, energy is released from the system.

Section 7.4

The Bohr Model



- Bohr's model is incorrect. This model only works for hydrogen.
- Electrons move around the nucleus in circular orbits.

Section 7.4

The Bohr Model

EXERCISE!

What color of light is emitted when an excited electron in the hydrogen atom falls from:

a) $n = 5$ to $n = 2$

blue, $\lambda = 434 \text{ nm}$

b) $n = 4$ to $n = 2$

green, $\lambda = 486 \text{ nm}$

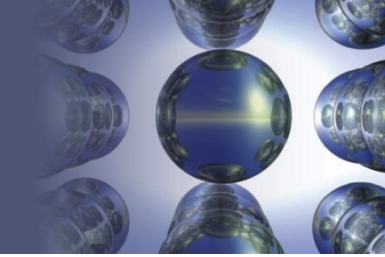
c) $n = 3$ to $n = 2$

orange/red, $\lambda = 657 \text{ nm}$

Which transition results in the **longest** wavelength of light?

Section 7.5

The Quantum Mechanical Model of the Atom



- We do not know the detailed pathway of an electron.
- Heisenberg uncertainty principle:
 - There is a fundamental limitation to just how precisely we can know both the position and momentum of a particle at a given time.

$$\Delta x \times \Delta (mv) \geq \frac{h}{4\pi}$$

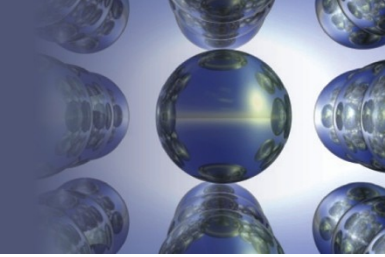
Δx = uncertainty in a particle's position

$\Delta(mv)$ = uncertainty in a particle's momentum

h = Planck's constant

Section 7.5

The Quantum Mechanical Model of the Atom



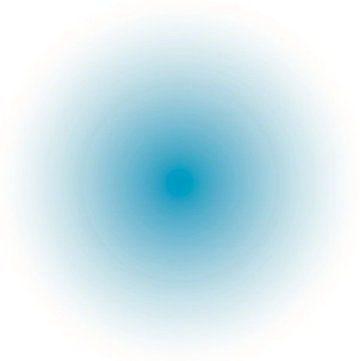
Physical Meaning of a Wave Function (Ψ)

- The square of the function indicates the probability of finding an electron near a particular point in space.
 - Probability distribution – intensity of color is used to indicate the probability value near a given point in space.

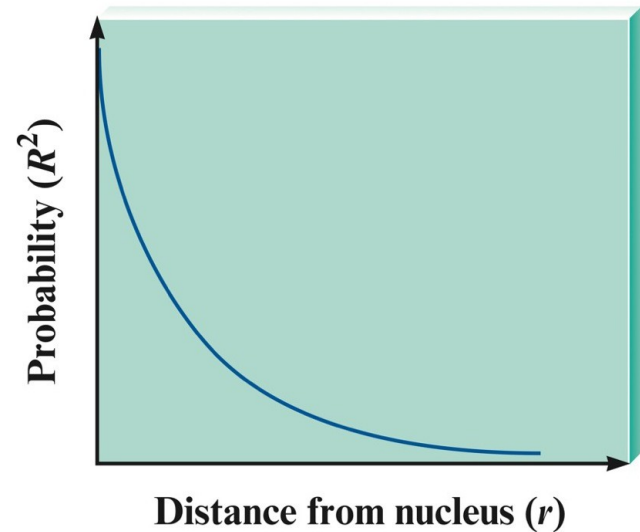
Section 7.5

The Quantum Mechanical Model of the Atom

Probability Distribution for the 1s Wave Function



a

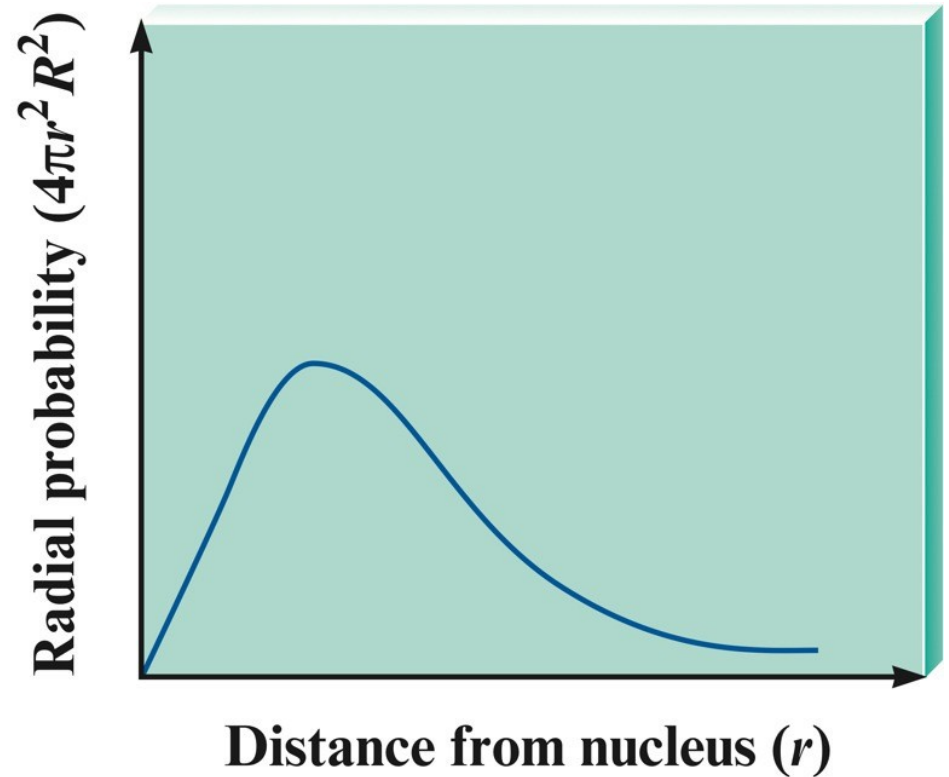
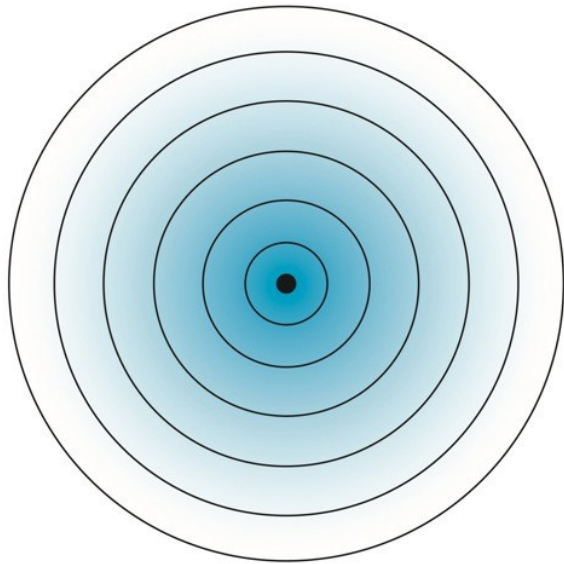


b

Section 7.5

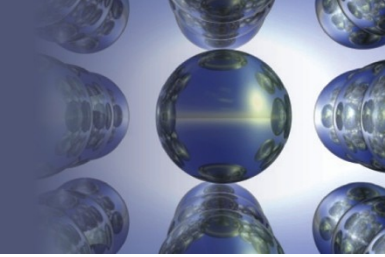
The Quantum Mechanical Model of the Atom

Radial Probability Distribution



Section 7.5

The Quantum Mechanical Model of the Atom

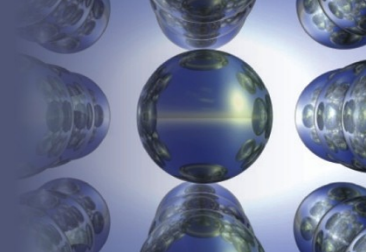


Relative Orbital Size

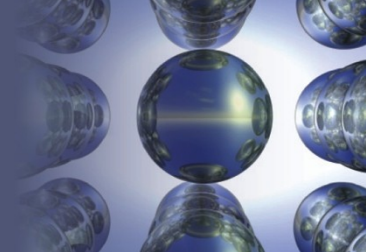
- Difficult to define precisely.
- Orbital is a wave function.
- Picture an orbital as a three-dimensional electron density map.
- Hydrogen 1s orbital:
 - Radius of the sphere that encloses 90% of the total electron probability.

Section 7.6

Quantum Numbers



- Principal quantum number (n) – size and energy of the orbital.
- Angular momentum quantum number (l) – shape of atomic orbitals (sometimes called a subshell).
- Magnetic quantum number (m_l) – orientation of the orbital in space relative to the other orbitals in the atom.



Section 7.6

Quantum Numbers

Quantum Numbers for the First Four Levels of Orbitals in the Hydrogen Atom

Table 7.2 | Quantum Numbers for the First Four Levels of Orbitals in the Hydrogen Atom

| n | ℓ | Sublevel Designation | m_ℓ | Number of Orbitals |
|-----|--------|----------------------|------------------------|--------------------|
| 1 | 0 | 1s | 0 | 1 |
| 2 | 0 | 2s | 0 | 1 |
| | 1 | 2p | -1, 0, +1 | 3 |
| 3 | 0 | 3s | 0 | 1 |
| | 1 | 3p | -1, 0, 1 | 3 |
| | 2 | 3d | -2, -1, 0, 1, 2 | 5 |
| 4 | 0 | 4s | 0 | 1 |
| | 1 | 4p | -1, 0, 1 | 3 |
| | 2 | 4d | -2, -1, 0, 1, 2 | 5 |
| | 3 | 4f | -3, -2, -1, 0, 1, 2, 3 | 7 |

Section 7.6

Quantum Numbers

EXERCISE!

For principal quantum level $n = 3$, determine the number of allowed subshells (different values of l), and give the designation of each.

of allowed subshells = 3

$l = 0, 3s$

$l = 1, 3p$

$l = 2, 3d$

Section 7.6

Quantum Numbers

EXERCISE!

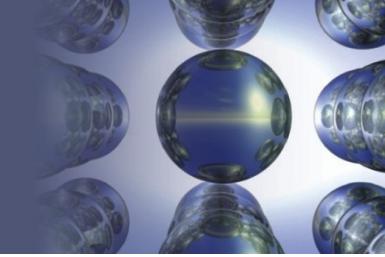
For $l = 2$, determine the magnetic quantum numbers (m_l) and the number of orbitals.

magnetic quantum numbers = $-2, -1, 0, 1, 2$

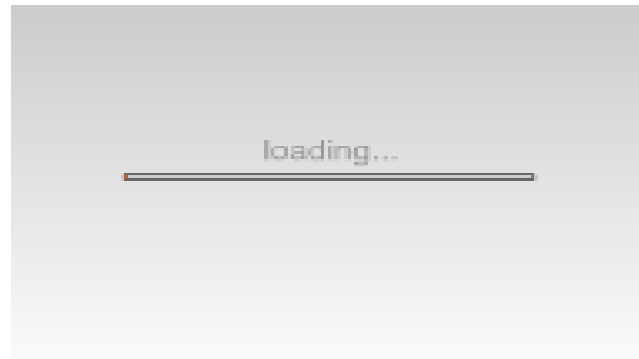
number of orbitals = 5

Section 7.7

Orbital Shapes and Energies



1s Orbital



To play movie you must be in Slide Show Mode

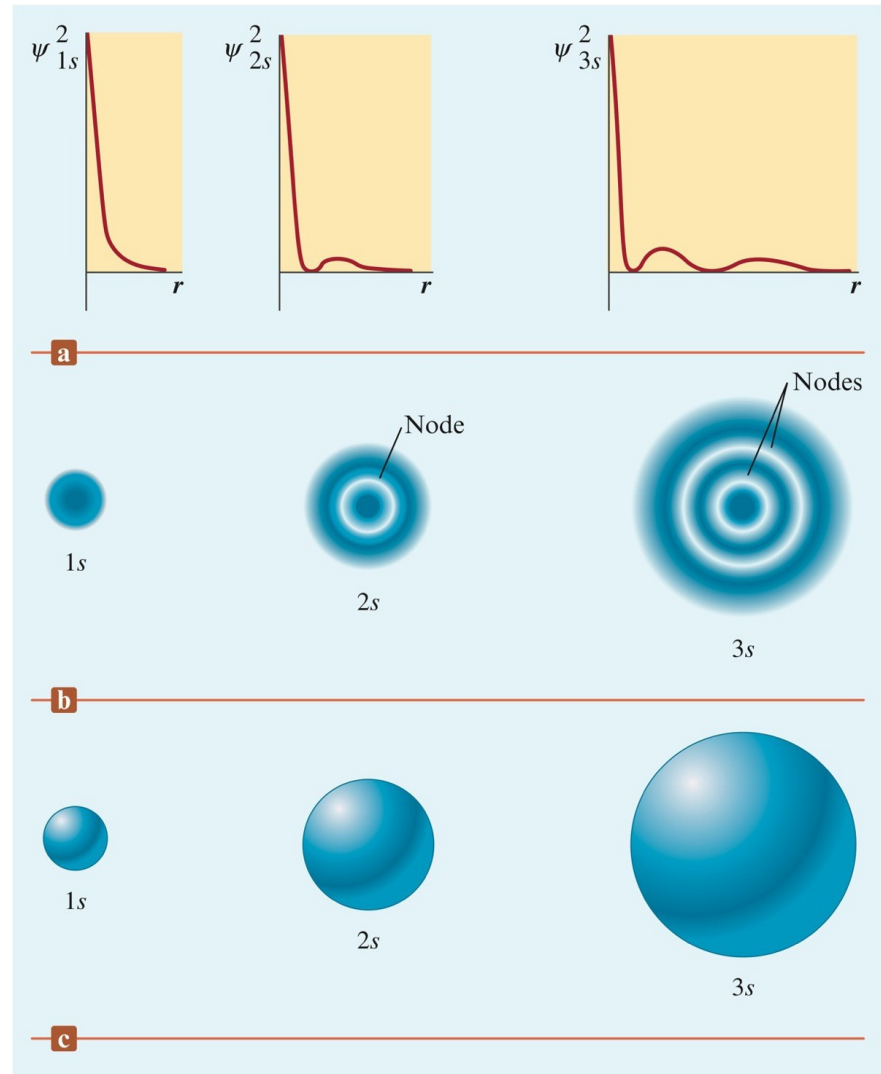
PC Users: Please wait for content to load, then click to play

Mac Users: [CLICK HERE](#)

Section 7.7

Orbital Shapes and Energies

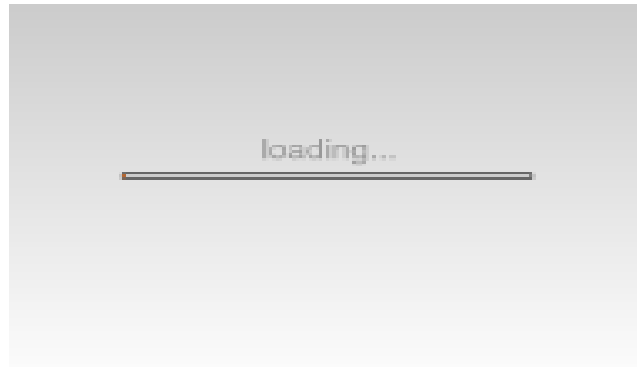
Three Representations of the Hydrogen 1s, 2s, and 3s Orbitals



Section 7.7

Orbital Shapes and Energies

2p_x Orbital



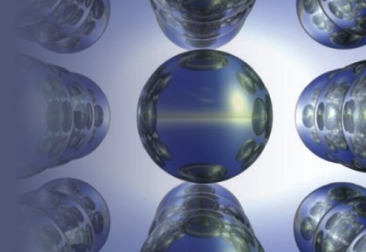
To play movie you must be in Slide Show Mode

PC Users: Please wait for content to load, then click to play

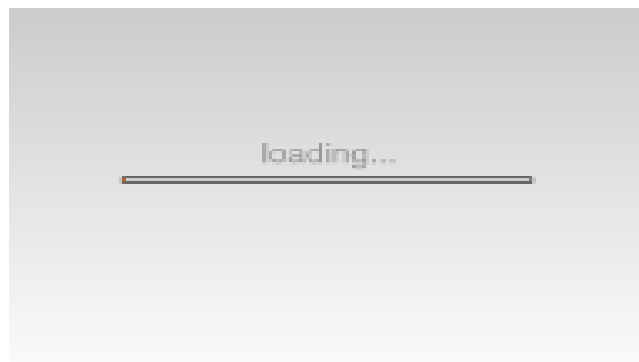
Mac Users: [CLICK HERE](#)

Section 7.7

Orbital Shapes and Energies



2p_y Orbital



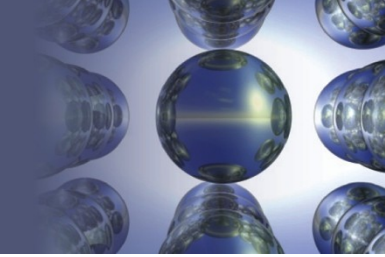
To play movie you must be in Slide Show Mode

PC Users: Please wait for content to load, then click to play

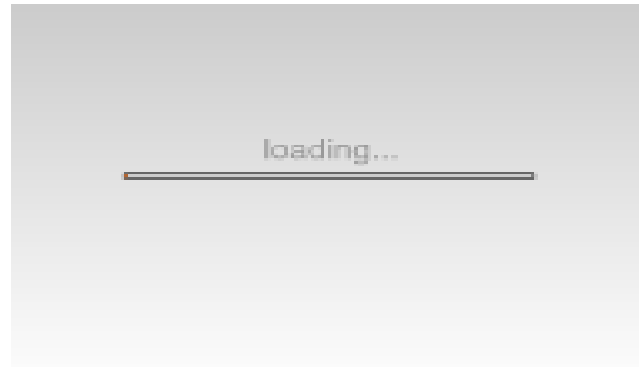
Mac Users: [CLICK HERE](#)

Section 7.7

Orbital Shapes and Energies



2p_z Orbital



To play movie you must be in Slide Show Mode

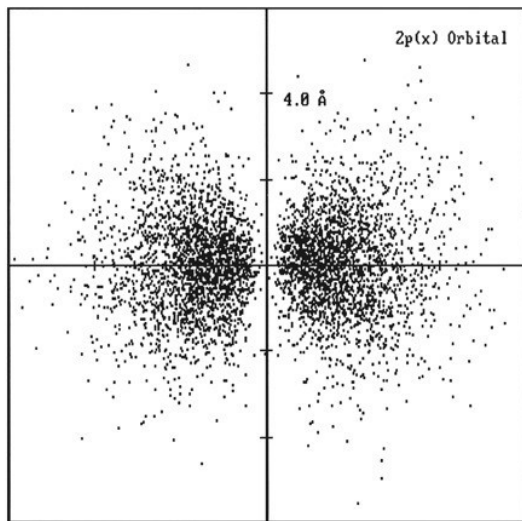
PC Users: Please wait for content to load, then click to play

Mac Users: [CLICK HERE](#)

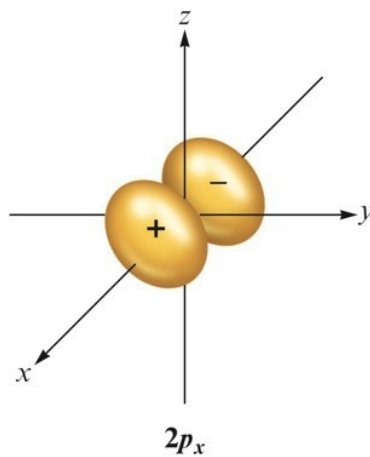
Section 7.7

Orbital Shapes and Energies

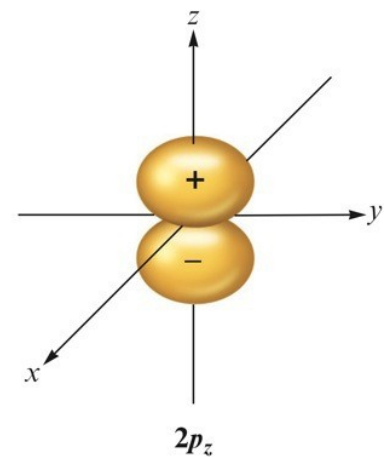
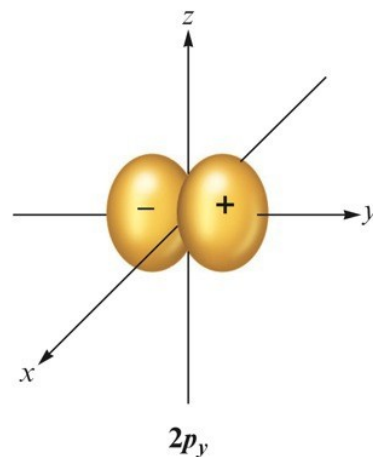
The Boundary Surface Representations of All Three $2p$ Orbitals



a



b



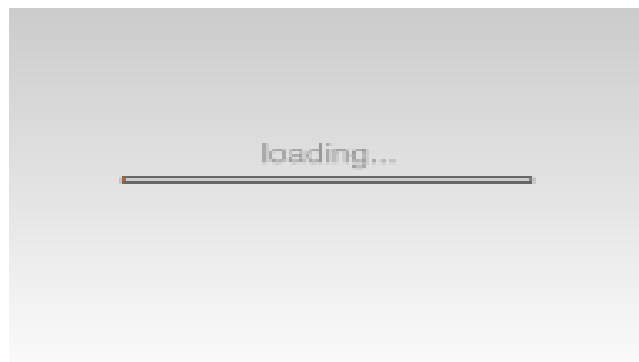
(Generated from a program by Robert Allendoerfer on Project SERAPHIM disk PC 2402; reprinted with permission.)

© Cengage Learning. All Rights Reserved.

Section 7.7

Orbital Shapes and Energies

$3d_{x^2-y^2}$ Orbital



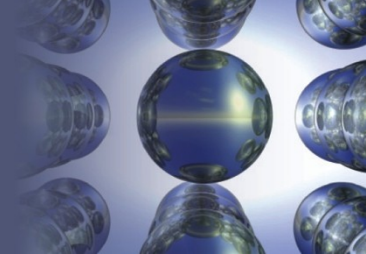
To play movie you must be in Slide Show Mode

PC Users: Please wait for content to load, then click to play

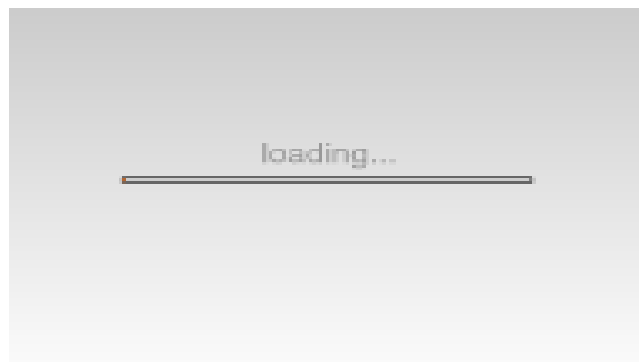
Mac Users: [CLICK HERE](#)

Section 7.7

Orbital Shapes and Energies



$3d_{xy}$ Orbital



To play movie you must be in Slide Show Mode

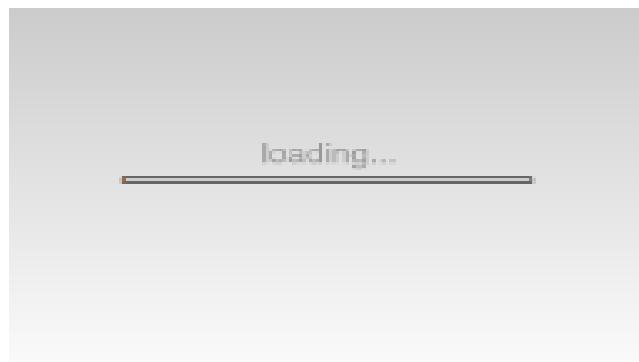
PC Users: Please wait for content to load, then click to play

Mac Users: [CLICK HERE](#)

Section 7.7

Orbital Shapes and Energies

$3d_{xz}$ Orbital



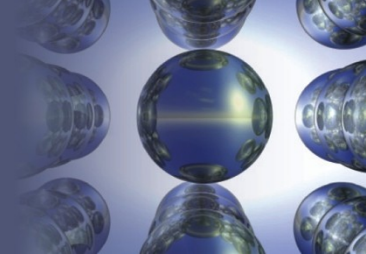
To play movie you must be in Slide Show Mode

PC Users: Please wait for content to load, then click to play

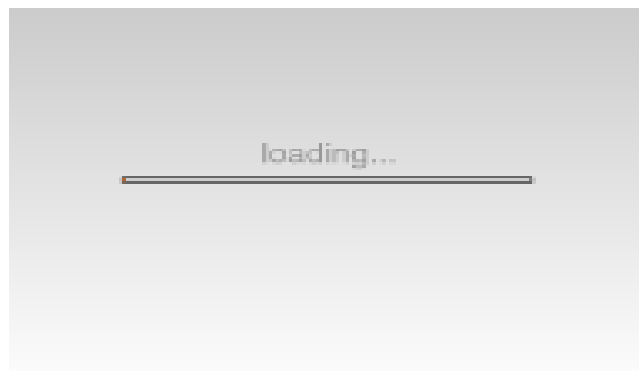
Mac Users: [CLICK HERE](#)

Section 7.7

Orbital Shapes and Energies



$3d_{yz}$ Orbital



To play movie you must be in Slide Show Mode

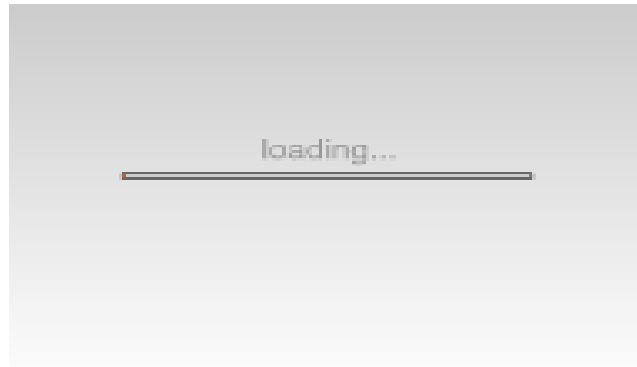
PC Users: Please wait for content to load, then click to play

Mac Users: [CLICK HERE](#)

Section 7.7

Orbital Shapes and Energies

$3d_z^2$



To play movie you must be in Slide Show Mode

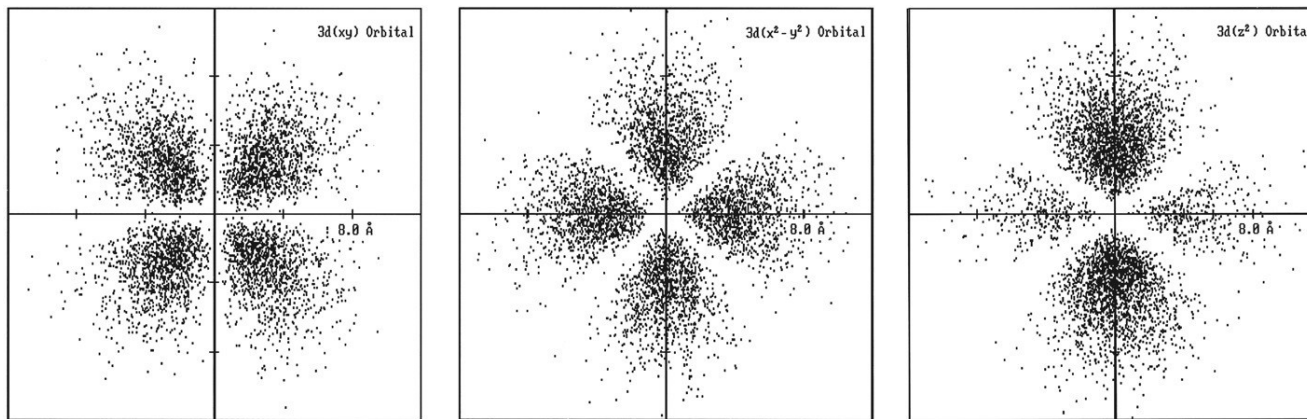
PC Users: Please wait for content to load, then click to play

Mac Users: [CLICK HERE](#)

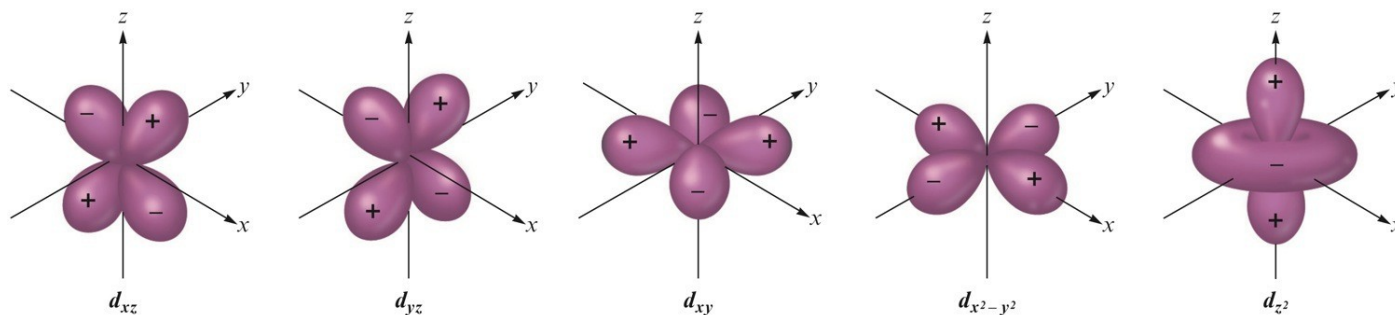
Section 7.7

Orbital Shapes and Energies

The Boundary Surfaces of All of the 3d Orbitals



a



b

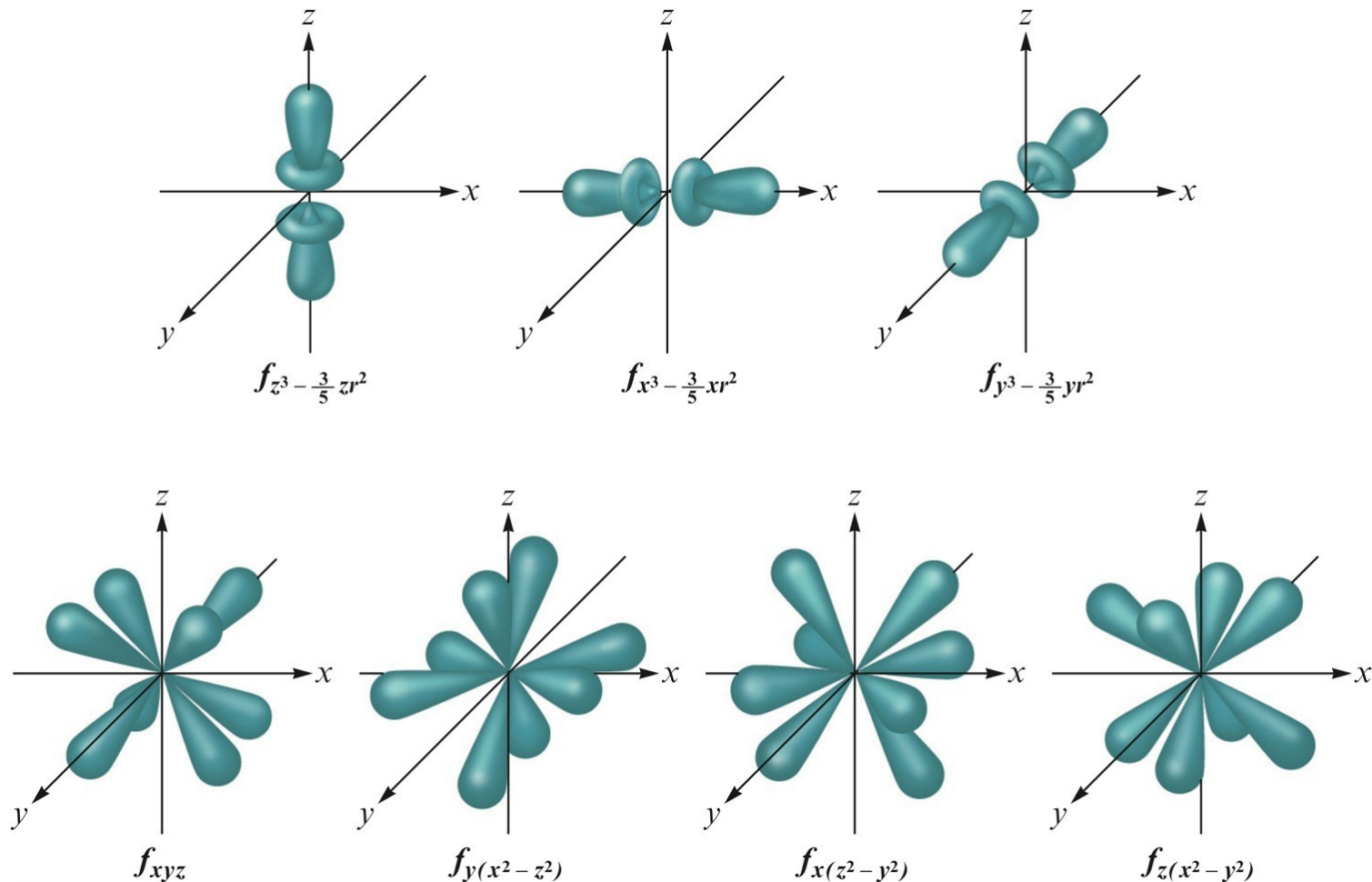
(Generated from a program by Robert Allendoerfer on Project SERAPHIM disk PC 2402; reprinted with permission.)

© Cengage Learning. All Rights Reserved.

Section 7.7

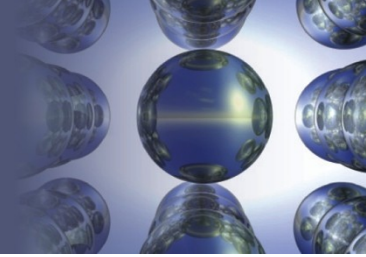
Orbital Shapes and Energies

Representation of the 4f Orbitals in Terms of Their Boundary Surfaces



Section 7.8

Electron Spin and the Pauli Principle

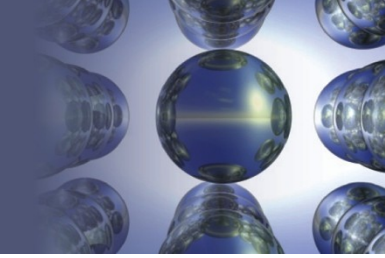


Electron Spin

- Electron spin quantum number (m_s) – can be $+\frac{1}{2}$ or $-\frac{1}{2}$.
- Pauli exclusion principle - in a given atom no two electrons can have the same set of four quantum numbers.
- An orbital can hold only two electrons, and they must have opposite spins.

Section 7.9

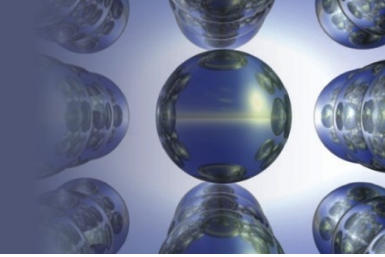
Polyelectronic Atoms



- Atoms with more than one electron.
- Electron correlation problem:
 - Since the electron pathways are unknown, the electron repulsions cannot be calculated exactly.
- When electrons are placed in a particular quantum level, they “prefer” the orbitals in the order s , p , d , and then f .

Section 7.9

Polyelectronic Atoms

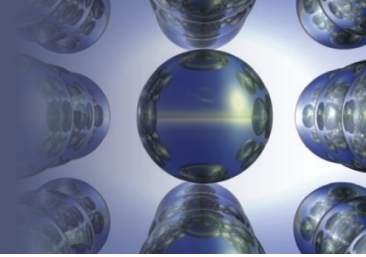


Penetration Effect

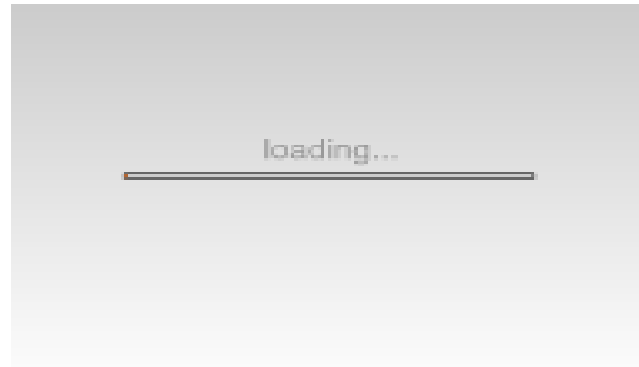
- A $2s$ electron penetrates to the nucleus more than one in the $2p$ orbital.
- This causes an electron in a $2s$ orbital to be attracted to the nucleus more strongly than an electron in a $2p$ orbital.
- Thus, the $2s$ orbital is lower in energy than the $2p$ orbitals in a polyelectronic atom.

Section 7.9

Polyelectronic Atoms



Orbital Energies



To play movie you must be in Slide Show Mode

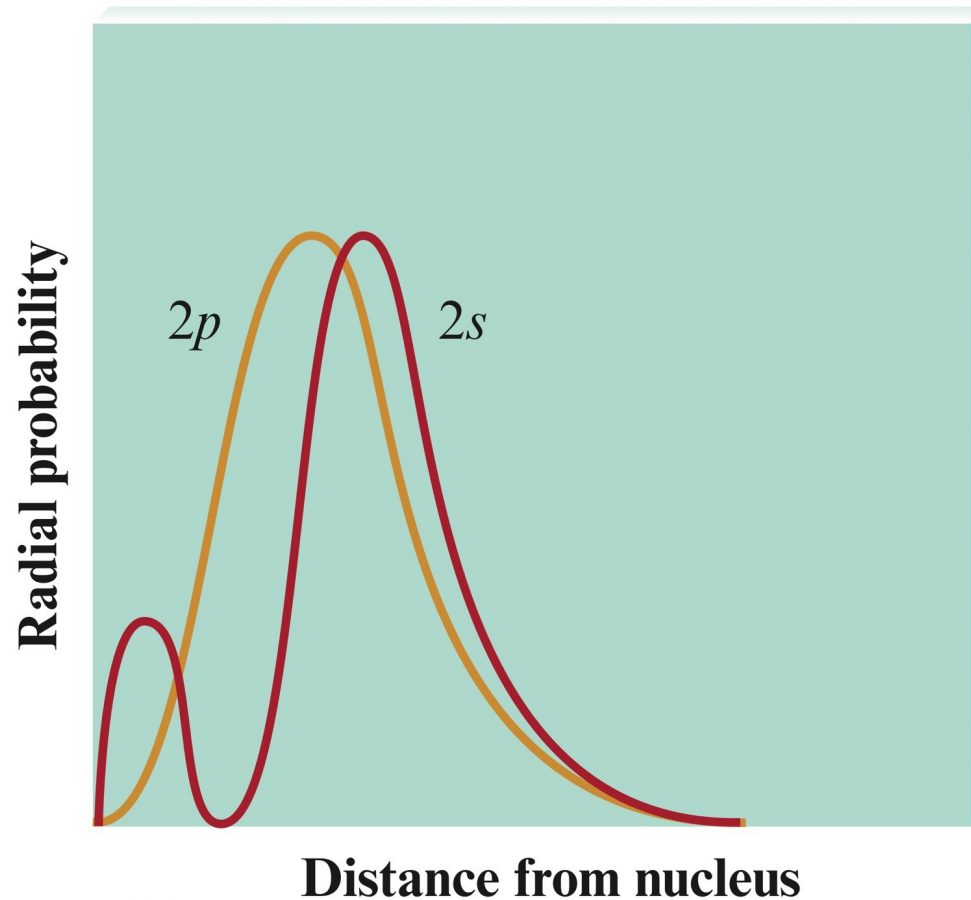
PC Users: Please wait for content to load, then click to play

Mac Users: [CLICK HERE](#)

Section 7.9

Polyelectronic Atoms

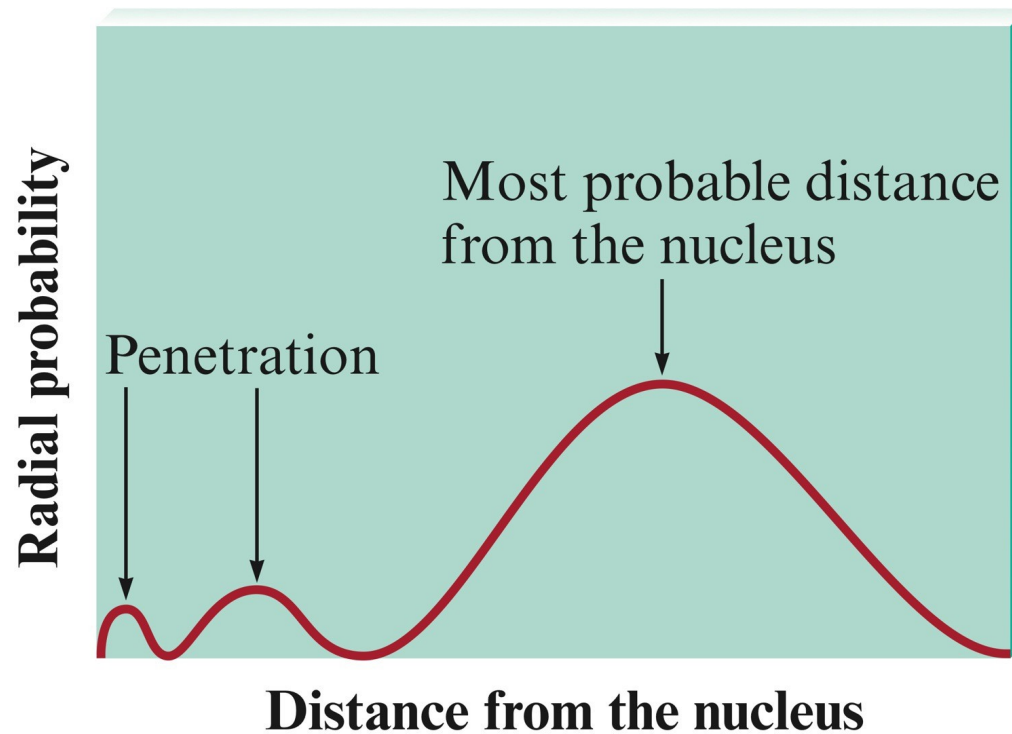
A Comparison of the Radial Probability Distributions of the $2s$ and $2p$ Orbitals



Section 7.9

Polyelectronic Atoms

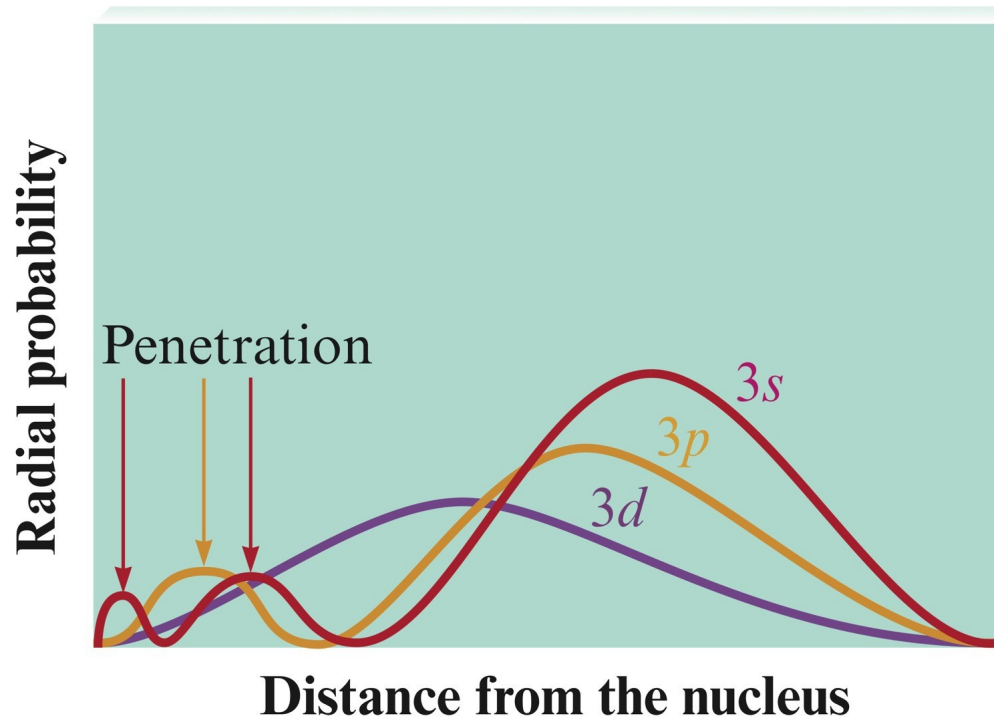
The Radial Probability Distribution of the 3s Orbital



Section 7.9

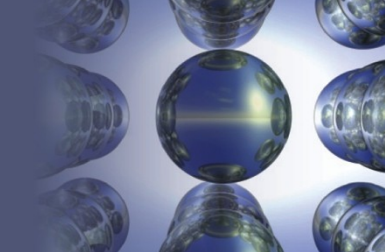
Polyelectronic Atoms

A Comparison of the Radial Probability Distributions of the $3s$, $3p$, and $3d$ Orbitals



Section 7.10

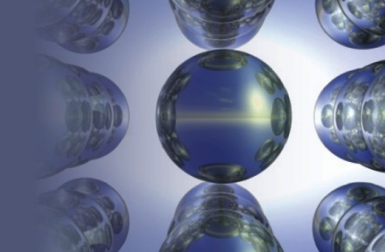
The History of the Periodic Table



- Originally constructed to represent the patterns observed in the chemical properties of the elements.
- Mendeleev is given the most credit for the current version of the periodic table because he emphasized how useful the periodic table could be in predicting the existence and properties of still unknown elements.

Section 7.11

The Aufbau Principle and the Periodic Table



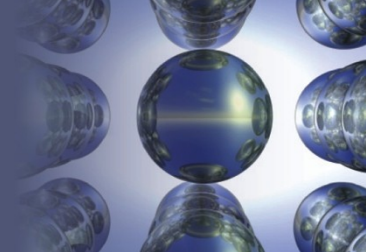
Aufbau Principle

- As protons are added one by one to the nucleus to build up the elements, electrons are similarly added to hydrogen-like orbitals.
- An oxygen atom has an electron arrangement of two electrons in the $1s$ subshell, two electrons in the $2s$ subshell, and four electrons in the $2p$ subshell.

Oxygen: $1s^2 2s^2 2p^4$

Section 7.11

The Aufbau Principle and the Periodic Table



Hund's Rule

- The lowest energy configuration for an atom is the one having the maximum number of unpaired electrons allowed by the Pauli principle in a particular set of degenerate (same energy) orbitals.

Section 7.11

The Aufbau Principle and the Periodic Table

Orbital Diagram

- A notation that shows how many electrons an atom has in each of its occupied electron orbitals.

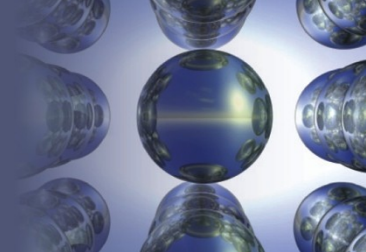
Oxygen: $1s^2 2s^2 2p^4$

Oxygen: $1s$ $2s$ $2p$



Section 7.11

The Aufbau Principle and the Periodic Table



Valence Electrons

- The electrons in the outermost principal quantum level of an atom.



- The elements in the same group on the periodic table have the same valence electron configuration.

Section 7.11

The Aufbau Principle and the Periodic Table

The Orbitals Being Filled for Elements in Various Parts of the Periodic Table

| | | Group | | | | | | | | | | | | | | | | | | | | | | | |
|--------|---|-------|----|--|--|--|----|----|--|--|--|--|----|--|--|--|----|--|----|----|----|----|----|----|----|
| | | 1A | | | | | | | | | | | | | | | | | 8A | | | | | | |
| Period | 1 | 1s | 2A | | | | | | | | | | | | | | | | | 3A | 4A | 5A | 6A | 7A | 1s |
| | 2 | 2s | | | | | | | | | | | | | | | | | | | 2p | | | | |
| | 3 | 3s | | | | | | | | | | | | | | | | | | | 3p | | | | |
| | 4 | 4s | | | | | 3d | | | | | | | | | | 4p | | | | | | | | |
| | 5 | 5s | | | | | 4d | | | | | | | | | | 5p | | | | | | | | |
| | 6 | 6s | La | | | | | 5d | | | | | | | | | 6p | | | | | | | | |
| | 7 | 7s | Ac | | | | | 6d | | | | | | | | | 7p | | | | | | | | |
| | | | | | | | | | | | | | | | | | | | | | | | | | |
| | | | | | | | | | | | | | 4f | | | | | | | | | | | | |
| | | | | | | | | | | | | | 5f | | | | | | | | | | | | |

Section 7.11

The Aufbau Principle and the Periodic Table

EXERCISE!

Determine the expected electron configurations for each of the following.

a) S



b) Ba

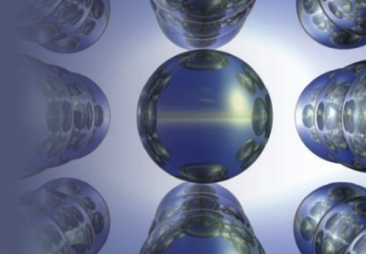


c) Eu



Section 7.12

Periodic Trends in Atomic Properties

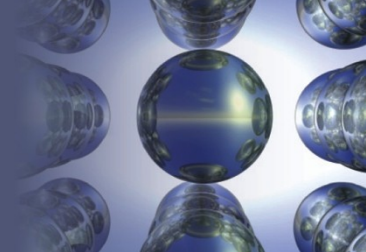


Periodic Trends

- Ionization Energy
- Electron Affinity
- Atomic Radius

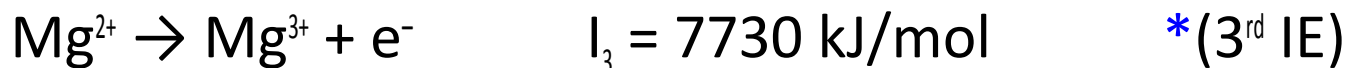
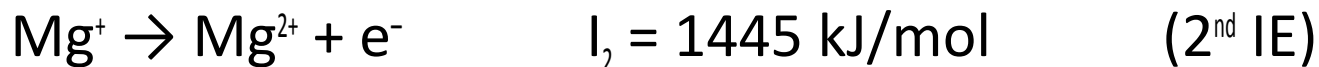
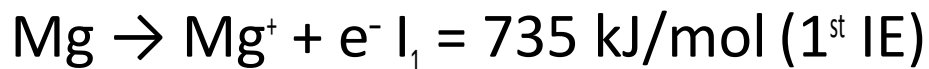
Section 7.12

Periodic Trends in Atomic Properties



Ionization Energy

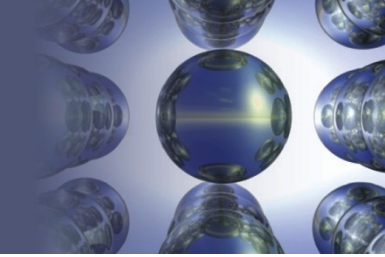
- Energy required to remove an electron from a gaseous atom or ion.
 - $X(g) \rightarrow X^+(g) + e^-$



*Core electrons are bound much more tightly than valence electrons.

Section 7.12

Periodic Trends in Atomic Properties

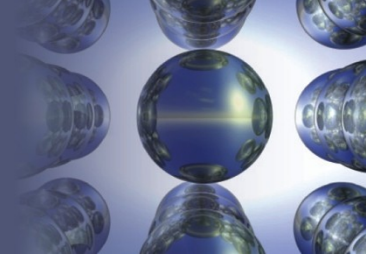


Ionization Energy

- In general, as we go across a period from left to right, the first ionization energy increases.
- Why?
 - Electrons added in the same principal quantum level do not completely shield the increasing nuclear charge caused by the added protons.
 - Electrons in the same principal quantum level are generally more strongly bound from left to right on the periodic table.

Section 7.12

Periodic Trends in Atomic Properties



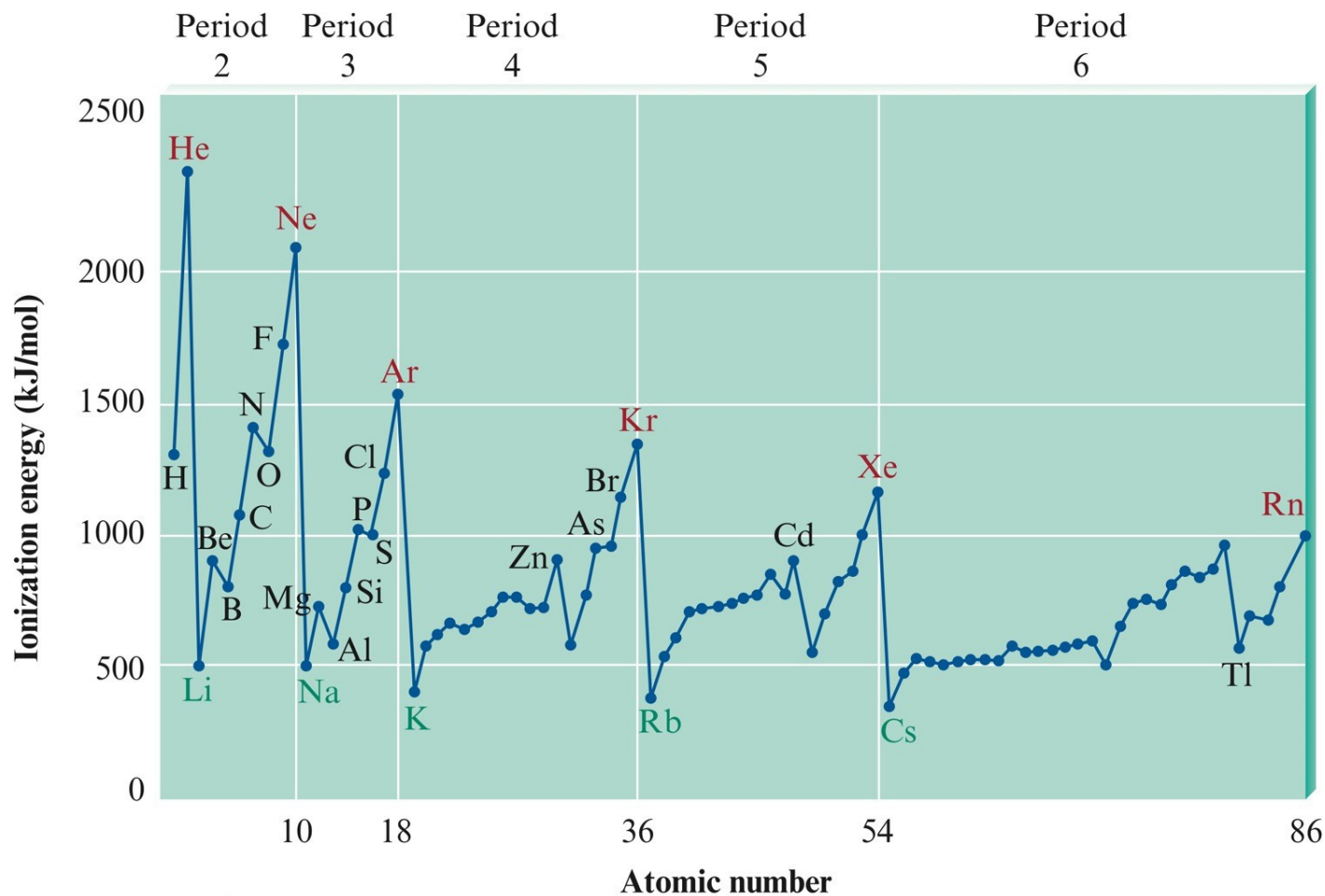
Ionization Energy

- In general, as we go down a group from top to bottom, the first ionization energy decreases.
- Why?
 - The electrons being removed are, on average, farther from the nucleus.

Section 7.12

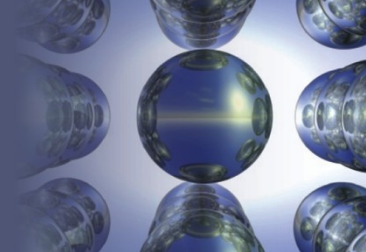
Periodic Trends in Atomic Properties

The Values of First Ionization Energy for the Elements in the First Six Periods



Section 7.12

Periodic Trends in Atomic Properties



CONCEPT CHECK!

Explain why the graph of ionization energy versus atomic number (across a row) is not linear.

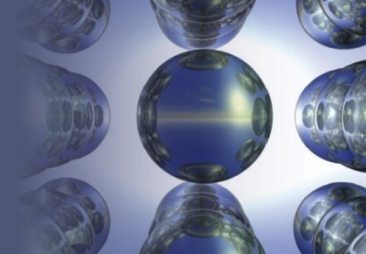
electron repulsions

Where are the exceptions?

some include from Be to B and N to O

Section 7.12

Periodic Trends in Atomic Properties



CONCEPT CHECK!

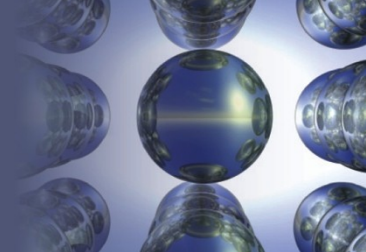
Which atom would require **more** energy to remove an electron? Why?

Na

Cl

Section 7.12

Periodic Trends in Atomic Properties



CONCEPT CHECK!

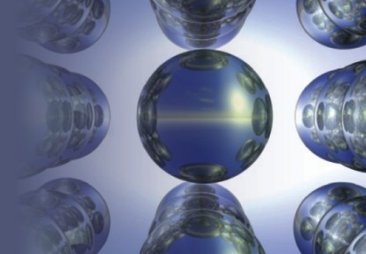
Which atom would require **more** energy to remove an electron? Why?

Li

Cs

Section 7.12

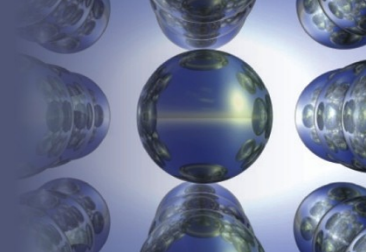
Periodic Trends in Atomic Properties



CONCEPT CHECK!

Which has the larger **second** ionization energy? Why?

Lithium or Beryllium



Section 7.12

Periodic Trends in Atomic Properties

Successive Ionization Energies (KJ per Mole) for the Elements in Period 3

Table 7.5 | Successive Ionization Energies (kJ/mol) for the Elements in Period 3

↑ General decrease

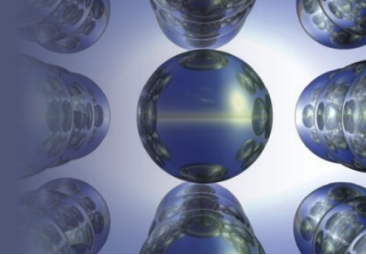
| Element | I_1 | I_2 | I_3 | I_4 | I_5 | I_6 | I_7 |
|---------|-------|-------|-------|-----------------|--------|--------|--------|
| Na | 495 | 4560 | | | | | |
| Mg | 735 | 1445 | 7730 | Core electrons* | | | |
| Al | 580 | 1815 | 2740 | 11,600 | | | |
| Si | 780 | 1575 | 3220 | 4350 | 16,100 | | |
| P | 1060 | 1890 | 2905 | 4950 | 6270 | 21,200 | |
| S | 1005 | 2260 | 3375 | 4565 | 6950 | 8490 | 27,000 |
| Cl | 1255 | 2295 | 3850 | 5160 | 6560 | 9360 | 11,000 |
| Ar | 1527 | 2665 | 3945 | 5770 | 7230 | 8780 | 12,000 |

General increase →

*Note the large jump in ionization energy in going from removal of valence electrons to removal of core electrons.

Section 7.12

Periodic Trends in Atomic Properties

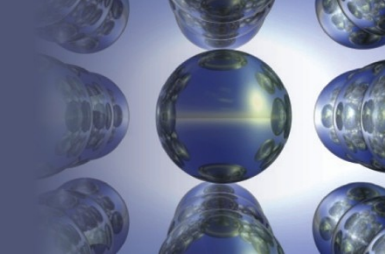


Electron Affinity

- Energy change associated with the addition of an electron to a gaseous atom.
 - $X(g) + e^- \rightarrow X^-(g)$
- In general as we go across a period from left to right, the electron affinities become more negative.
- In general electron affinity becomes more positive in going down a group.

Section 7.12

Periodic Trends in Atomic Properties



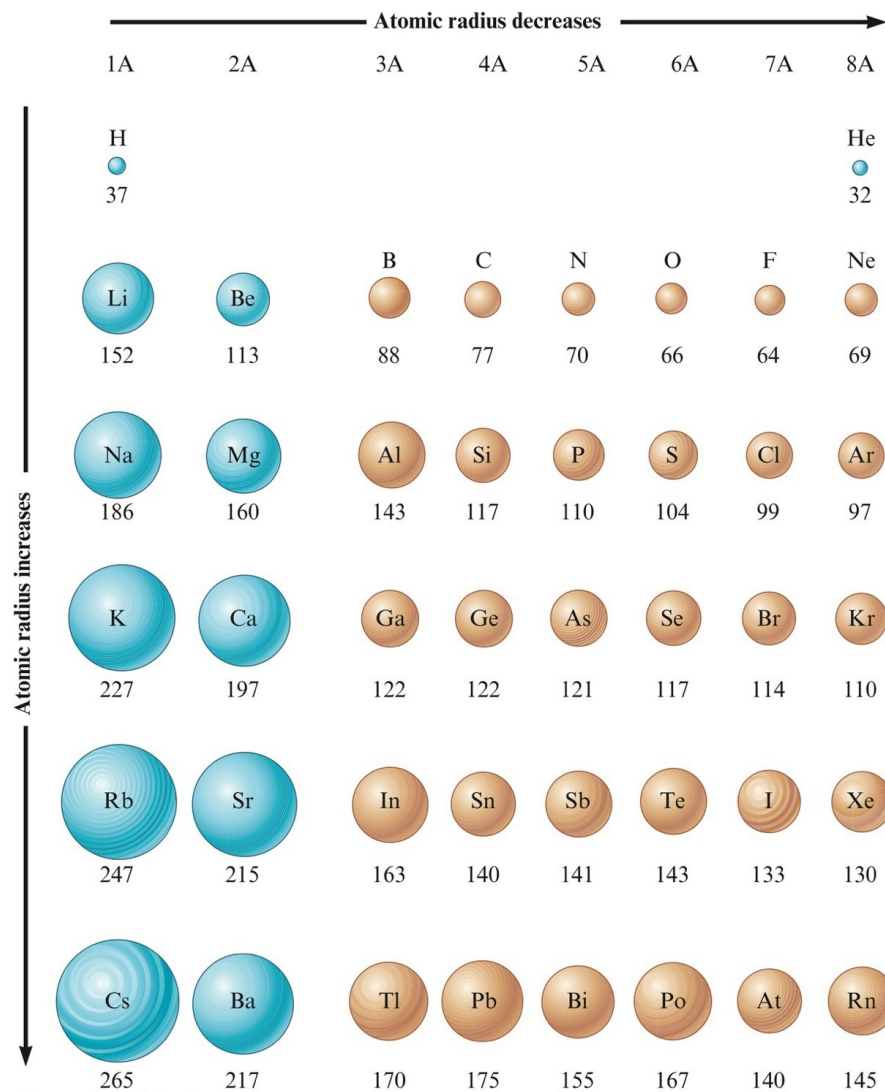
Atomic Radius

- In general as we go across a period from left to right, the atomic radius decreases.
 - Effective nuclear charge increases, therefore the valence electrons are drawn closer to the nucleus, decreasing the size of the atom.
- In general atomic radius increases in going down a group.
 - Orbital sizes increase in successive principal quantum levels.

Section 7.12

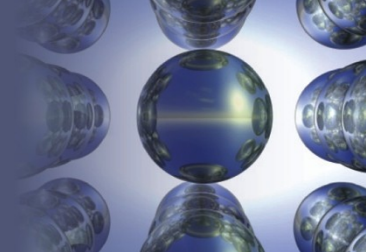
Periodic Trends in Atomic Properties

Atomic Radii for Selected Atoms



Section 7.12

Periodic Trends in Atomic Properties



CONCEPT CHECK!

Which should be the **larger** atom? Why?

Na

Cl

Section 7.12

Periodic Trends in Atomic Properties

CONCEPT CHECK!

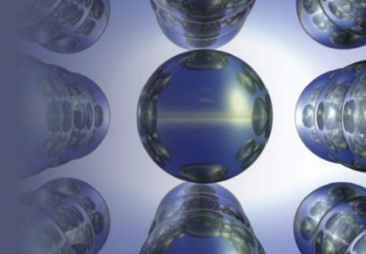
Which should be the **larger** atom? Why?

Li

Cs

Section 7.12

Periodic Trends in Atomic Properties



CONCEPT CHECK!

Which is larger?

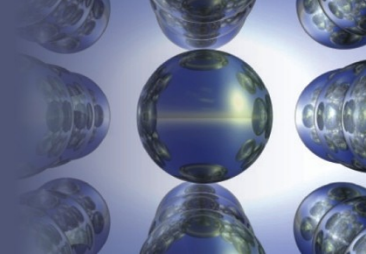
- The hydrogen 1s orbital
- The lithium 1s orbital

Which is lower in energy?

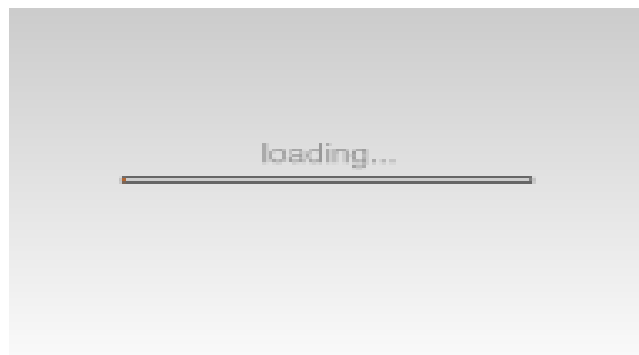
- The hydrogen 1s orbital
- The lithium 1s orbital

Section 7.12

Periodic Trends in Atomic Properties



Atomic Radius of a Metal



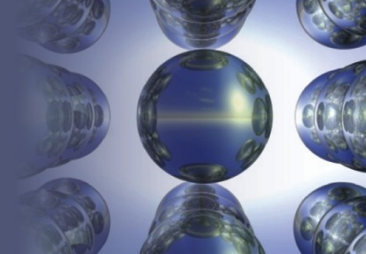
To play movie you must be in Slide Show Mode

PC Users: Please wait for content to load, then click to play

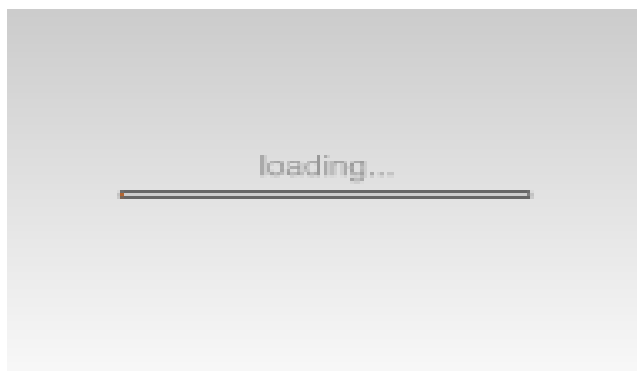
Mac Users: [CLICK HERE](#)

Section 7.12

Periodic Trends in Atomic Properties



Atomic Radius of a Nonmetal



To play movie you must be in Slide Show Mode

PC Users: Please wait for content to load, then click to play

Mac Users: [CLICK HERE](#)

Section 7.12

Periodic Trends in Atomic Properties

EXERCISE!

Arrange the elements oxygen, fluorine, and sulfur according to increasing:

- Ionization energy

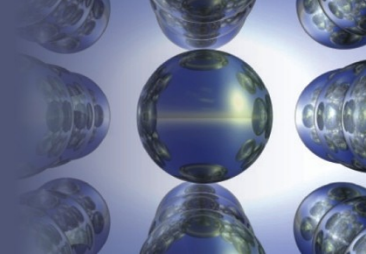
S, O, F

- Atomic size

F, O, S

Section 7.13

The Properties of a Group: The Alkali Metals



The Periodic Table – Final Thoughts

1. It is the number and type of valence electrons that primarily determine an atom's chemistry.
2. Electron configurations can be determined from the organization of the periodic table.
3. Certain groups in the periodic table have special names.

Section 7.13

The Properties of a Group: The Alkali Metals

Special Names for Groups in the Periodic Table

Alkali metals

1A
H

2A

Alkaline earth metals

Halogens

3A 4A 5A 6A 7A

Noble gases
↓
8A

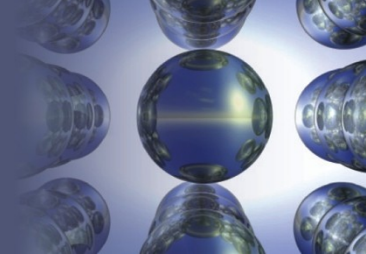
Transition elements

Lanthanides

Actinides

Section 7.13

The Properties of a Group: The Alkali Metals

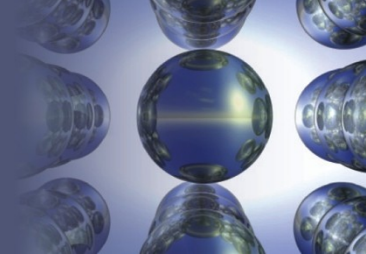


The Periodic Table – Final Thoughts

4. Basic division of the elements in the periodic table is into metals and nonmetals.

Section 7.13

The Properties of a Group: The Alkali Metals



The Alkali Metals

- Li, Na, K, Rb, Cs, and Fr
 - Most chemically reactive of the metals
 - React with nonmetals to form ionic solids
 - Going down group:
 - Ionization energy decreases
 - Atomic radius increases
 - Density increases
 - Melting and boiling points smoothly decrease