Table of Contents

Objectives

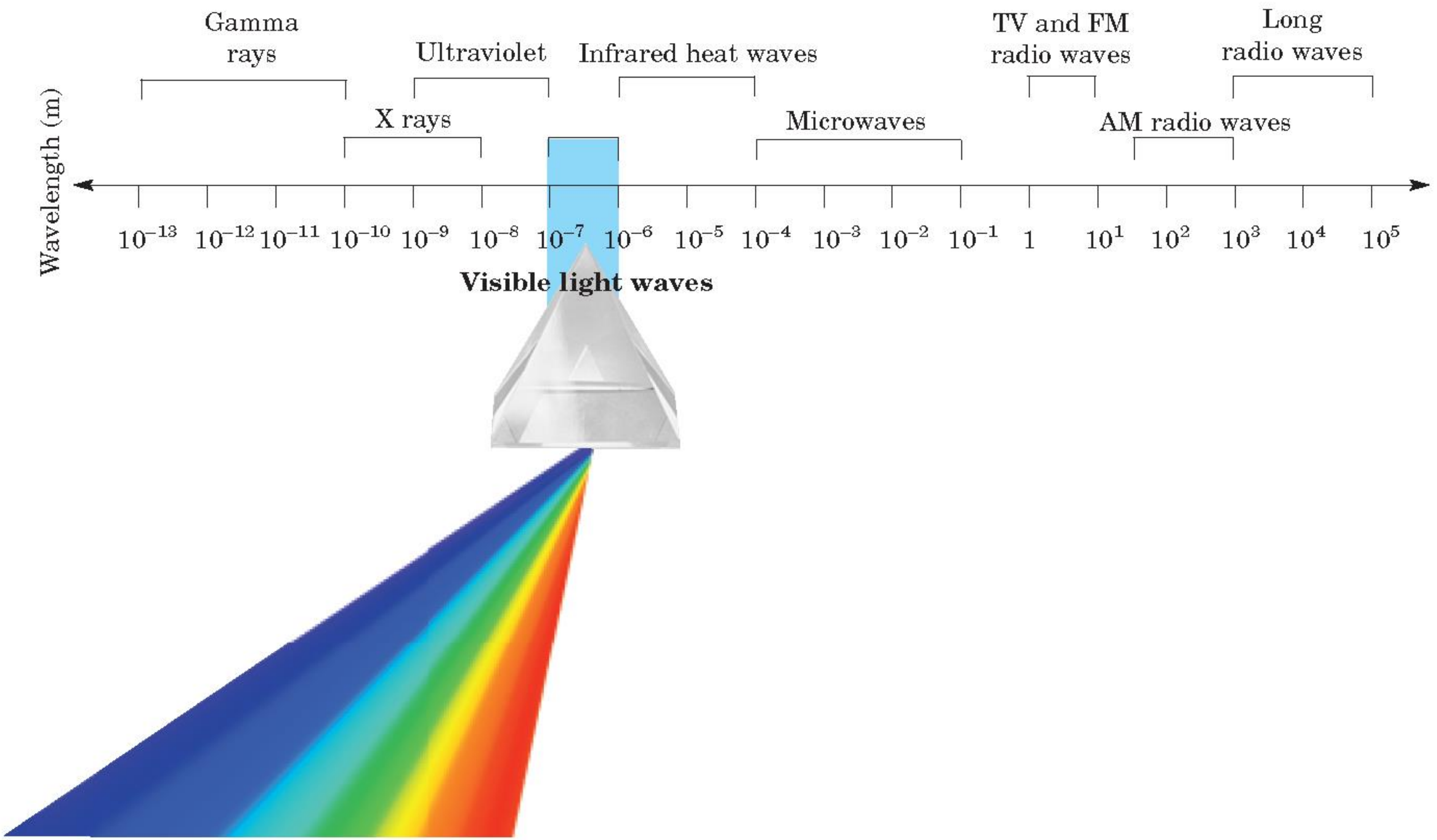
- **Explain** the mathematical relationship among the speed, wavelength, and frequency of electromagnetic radiation.
- **Discuss** the dual wave-particle nature of light.
- **Discuss** the significance of the photoelectric effect and the line-emission spectrum of hydrogen to the development of the atomic model.
- **Describe** the Bohr model of the hydrogen atom.

Properties of Light

The Wave Description of Light

- **Electromagnetic radiation** is a form of energy that exhibits wavelike behavior as it travels through space.
- Together, all the forms of electromagnetic radiation form the **electromagnetic spectrum.**

Electromagnetic Spectrum



Properties of Light, *continued*

- **Wavelength (λ)** is the distance between corresponding points on adjacent waves.

- **Frequency (v)** is defined as the number of waves that pass a given point in a specific time, usually one second.

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Properties of Light, *continued*

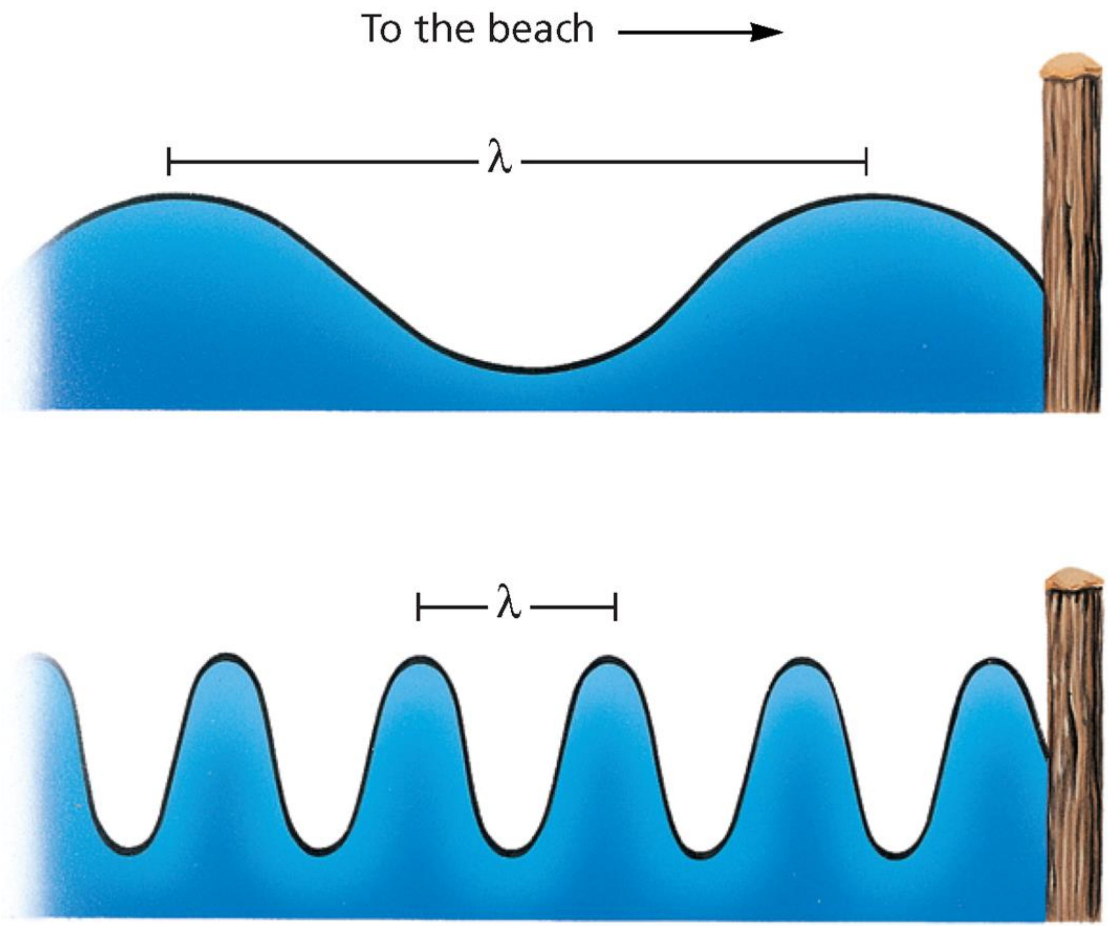
- Frequency and wavelength are mathematically related to each other:

$$c = \lambda \nu$$

- In the equation, c is the speed of light (in m/s), λ is the wavelength of the electromagnetic wave (in m), and ν is the frequency of the electromagnetic wave (in s^{-1}).

Wavelength and Frequency

The distance between any two corresponding points on one of these water waves, such as from crest to crest, is the wave's wavelength, λ. We can measure the wave's frequency, ν, by observing how often the water level rises and falls at a given point, such as at the post.



The Photoelectric Effect

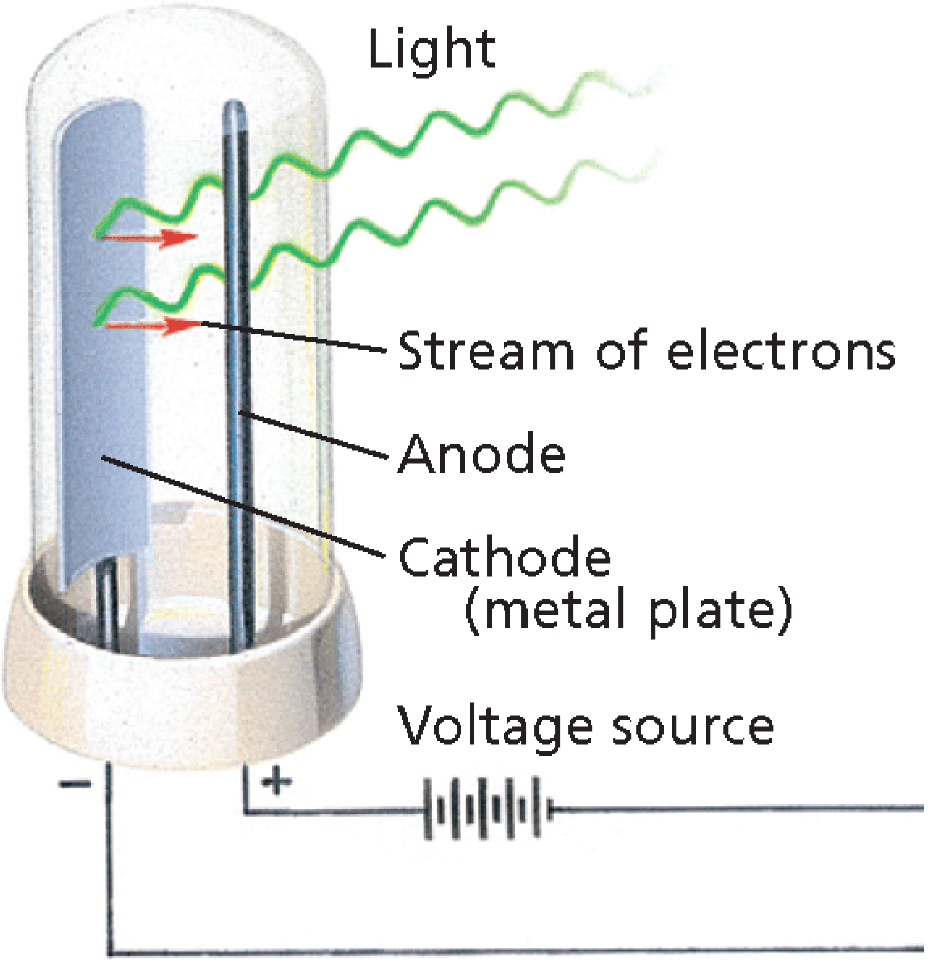
- The **photoelectric effect** refers to the emission of electrons from a metal when light shines on the metal.

The Particle Description of Light

- A **quantum** of energy is the minimum quantity of energy that can be lost or gained by an atom.

Photoelectric Effect

Electromagnetic radiation strikes the surface of the metal, ejecting electrons from the metal and causing an electric current.



The Photoelectric Effect, *continued*

The Particle Description of Light, *continued*

- German physicist Max Planck proposed the following relationship between a quantum of energy and the frequency of radiation:

$$E = h\nu$$

  - E is the energy, in joules, of a quantum of radiation, ν is the frequency, in s^{-1}, of the radiation emitted, and h is a fundamental physical constant now known as Planck's constant; $h = 6.626 \times 10^{-34}$ J• s.

End Of Slide

The Photoelectric Effect, *continued*

The Particle Description of Light, *continued*

- A **photon** is a particle of electromagnetic radiation having zero mass and carrying a quantum of energy.
- The energy of a particular photon depends on the frequency of the radiation.

$$E_{photon} = hv$$

End of Slide

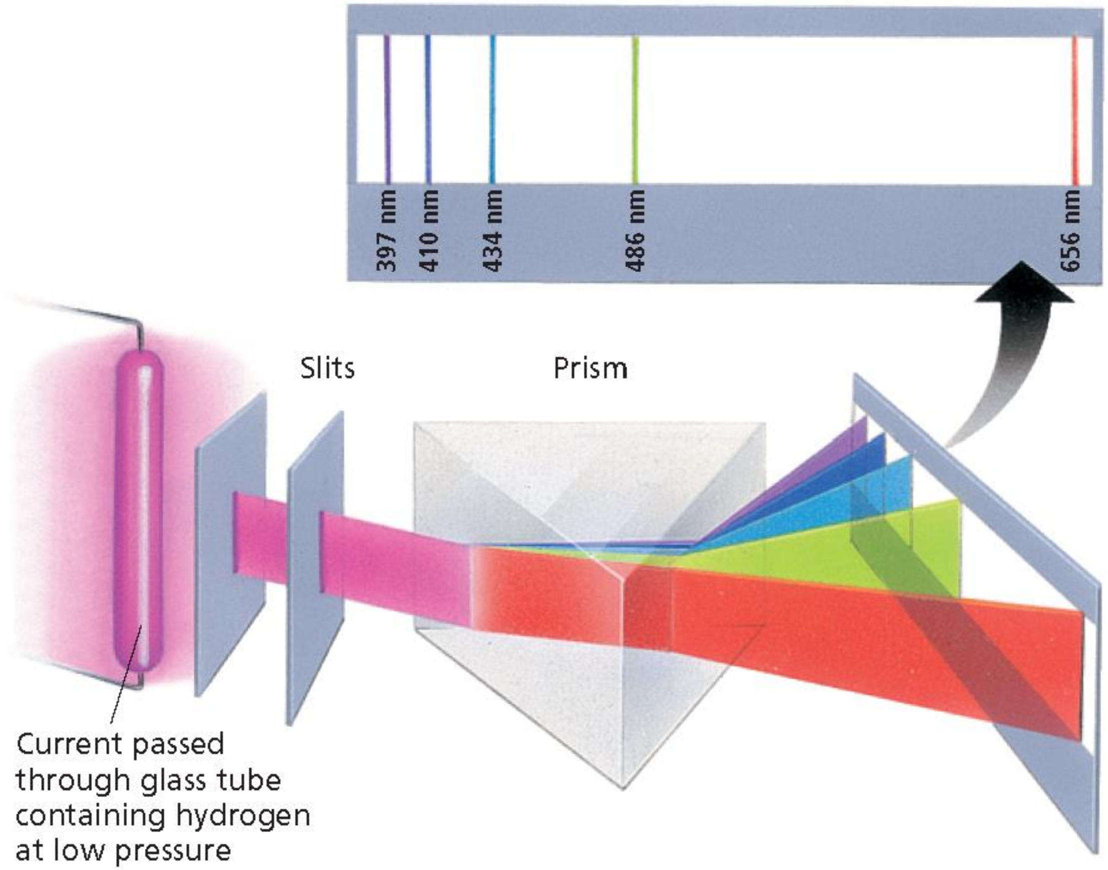
The Hydrogen-Atom Line-Emission Spectrum

- The lowest energy state of an atom is its **ground state.**
- A state in which an atom has a higher potential energy than it has in its ground state is an **excited state.**

The Hydrogen-Atom Line-Emission Spectrum, *continued*

- When investigators passed electric current through a vacuum tube containing hydrogen gas at low pressure, they observed the emission of a characteristic pinkish glow.
- When a narrow beam of the emitted light was shined through a prism, it was separated into four specific colors of the visible spectrum.
- The four bands of light were part of what is known as hydrogen's **line-emission spectrum.**

Hydrogen's Line-Emission Spectrum



Excited hydrogen atoms emit a pinkish glow. When the visible portion of the emitted light is passed through a prism, it is separated into specific wavelengths that are part of hydrogen's line-emission spectrum.

Bohr Model of the Hydrogen Atom

- Niels Bohr proposed a hydrogen-atom model that linked the atom's electron to photon emission.
- According to the model, the electron can circle the nucleus only in allowed paths, or *orbits.*
- The energy of the electron is higher when the electron is in orbits that are successively farther from the nucleus.

Bohr Model of the Hydrogen Atom, *continued*

- When an electron falls to a lower energy level, a photon is emitted, and the process is called *emission.*

- Energy must be added to an atom in order to move an electron from a lower energy level to a higher energy level. This process is called *absorption.*

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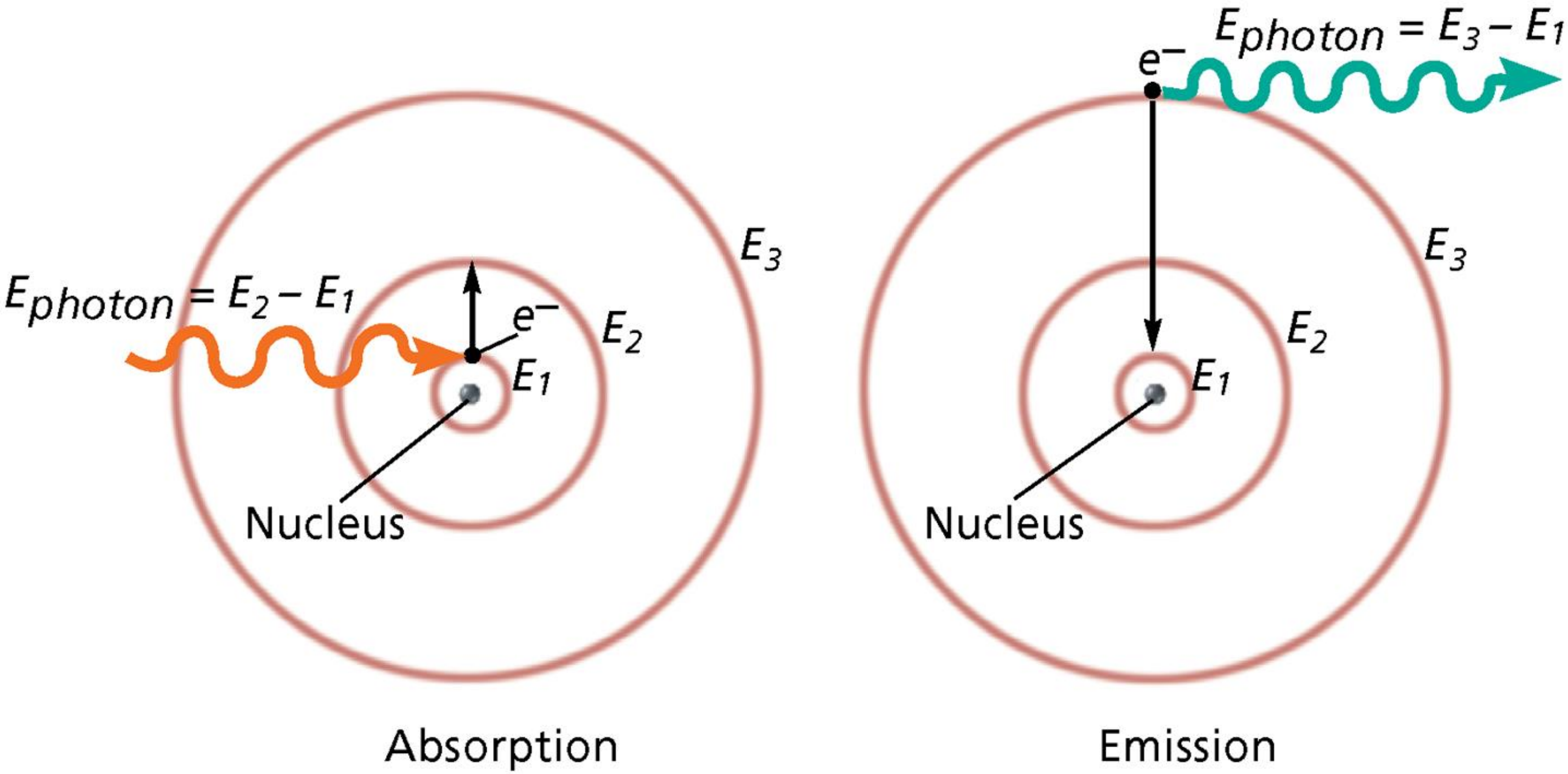
Photon Emission and Absorption



Objectives

- **Discuss** Louis de Broglie's role in the development of the quantum model of the atom.
- **Compare and contrast** the Bohr model and the quantum model of the atom.
- **Explain** how the Heisenberg uncertainty principle and the Schrödinger wave equation led to the idea of atomic orbitals.

Objectives, *continued*

- **List** the four quantum numbers and **describe** their significance.
- **Relate** the number of sublevels corresponding to each of an atom's main energy levels, the number of orbitals per sublevel, and the number of orbitals per main energy level.

Electrons as Waves

- French scientist Louis de Broglie suggested that electrons be considered waves confined to the space around an atomic nucleus.
- It followed that the electron waves could exist only at specific frequencies.
- According to the relationship $E = h\nu$, these frequencies corresponded to specific energies—the quantized energies of Bohr's orbits.

End Of Slide

Electrons as Waves, *continued*

- Electrons, like light waves, can be bent, or diffracted.
- *Diffraction* refers to the bending of a wave as it passes by the edge of an object or through a small opening.
- Electron beams, like waves, can interfere with each other.
- *Interference* occurs when waves overlap.

The Heisenberg Uncertainty Principle

- German physicist Werner Heisenberg proposed that any attempt to locate a specific electron with a photon knocks the electron off its course.
- The **Heisenberg uncertainty principle** states that it is impossible to determine simultaneously both the position and velocity of an electron or any other particle.
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The Schrödinger Wave Equation

- In 1926, Austrian physicist Erwin Schrödinger developed an equation that treated electrons in atoms as waves.
- Together with the Heisenberg uncertainty principle, the Schrödinger wave equation laid the foundation for modern quantum theory.
- **Quantum theory** describes mathematically the wave properties of electrons and other very small particles.

The Schrödinger Wave Equation, *continued*

- Electrons do not travel around the nucleus in neat orbits, as Bohr had postulated.
- Instead, they exist in certain regions called orbitals.
- An **orbital** is a three-dimensional region around the nucleus that indicates the probable location of an electron.

Atomic Orbitals and Quantum Numbers

- **Quantum numbers** specify the properties of atomic orbitals and the properties of electrons in orbitals.
- The **principal quantum number,** symbolized by *n,* indicates the main energy level occupied by the electron.
- The **angular momentum quantum number,** symbolized by *l,* indicates the shape of the orbital.

Atomic Orbitals and Quantum Numbers, *continued*

- The **magnetic quantum number,** symbolized by *m,* indicates the orientation of an orbital around the nucleus.
- The **spin quantum number** has only two possible values—(+1/2 , −1/2)—which indicate the two fundamental spin states of an electron in an orbital.

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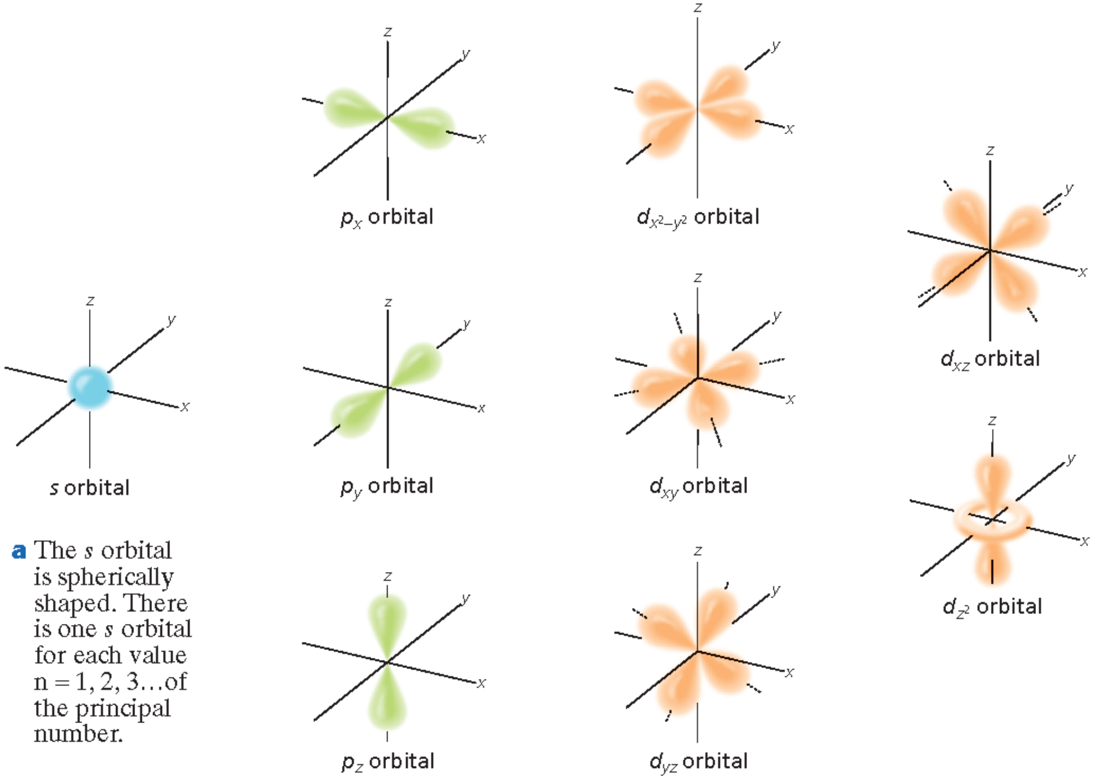
Shapes of *s*, *p*, and *d* Orbitals



a The *s* orbital is spherically shaped. There is one *s* orbital for each value n = 1, 2, 3…of the principal number.

b For each of the values n = 2, 3, 4…, there are three *p* orbitals. All are dumbbell shaped, but they differ in orientation.

c For each of the values n = 3, 4, 5…, there are five *d* orbitals. Four of the five have similar shapes, but differ in orientation.

Electrons Accommodated in Energy Levels and Sublevels

Principal energy level	Sublevels available	Number of orbitals in sublevel $(2\ell + 1)$	Number of electrons possible in sublevel $[2(2\ell + 1)]$	Total electrons possible for energy level $(2n^2)$
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

Electrons Accommodated in Energy Levels and Sublevels

Principal energy level	Sublevels available	Number of orbitals in sublevel $(2\ell + 1)$	Number of electrons possible in sublevel $[2(2\ell + 1)]$	Total electrons possible for energy level $(2n^2)$
5	*s*	1	2	50
	p	3	6	
	d	5	10	
	f	7	14	
	*g**	9	18	
6	*s*	1	2	72
	p	3	6	
	d	5	10	
	*f**	7	14	
	*g**	9	18	
	*h**	11	22	

*These orbitals are not used in the ground state of any known element.

Quantum Numbers of the First 30 Atomic Orbitals

n	*l*	*m*	Orbital name	Number of orbitals
1	0	0	1*s*	1
2	0	0	2*s*	1
2	1	–1, 0, 1	2*p*	3
3	0	0	3*s*	1
3	1	–1, 0, 1	3*p*	3
3	2	–2, –1, 0, 1, 2	3*d*	5
4	0	0	4*s*	1
4	1	–1, 0, 1	4*p*	3
4	2	–2, –1, 0, 1, 2	4*d*	5
4	3	–3, –2, –1, 0, 1, 2, 3	4*f*	7

Objectives

- **List** the total number of electrons needed to fully occupy each main energy level.
- **State** the *Aufbau principle,* the *Pauli exclusion principle,* and *Hund's rule.*
- **Describe** the electron configurations for the atoms of any element using *orbital notation, electron-configuration notation,* and, when appropriate, *noble-gas notation.*

Electron Configurations

- The arrangement of electrons in an atom is known as the atom's **electron configuration.**
- The lowest-energy arrangement of the electrons for each element is called the element's *ground-state electron configuration.*

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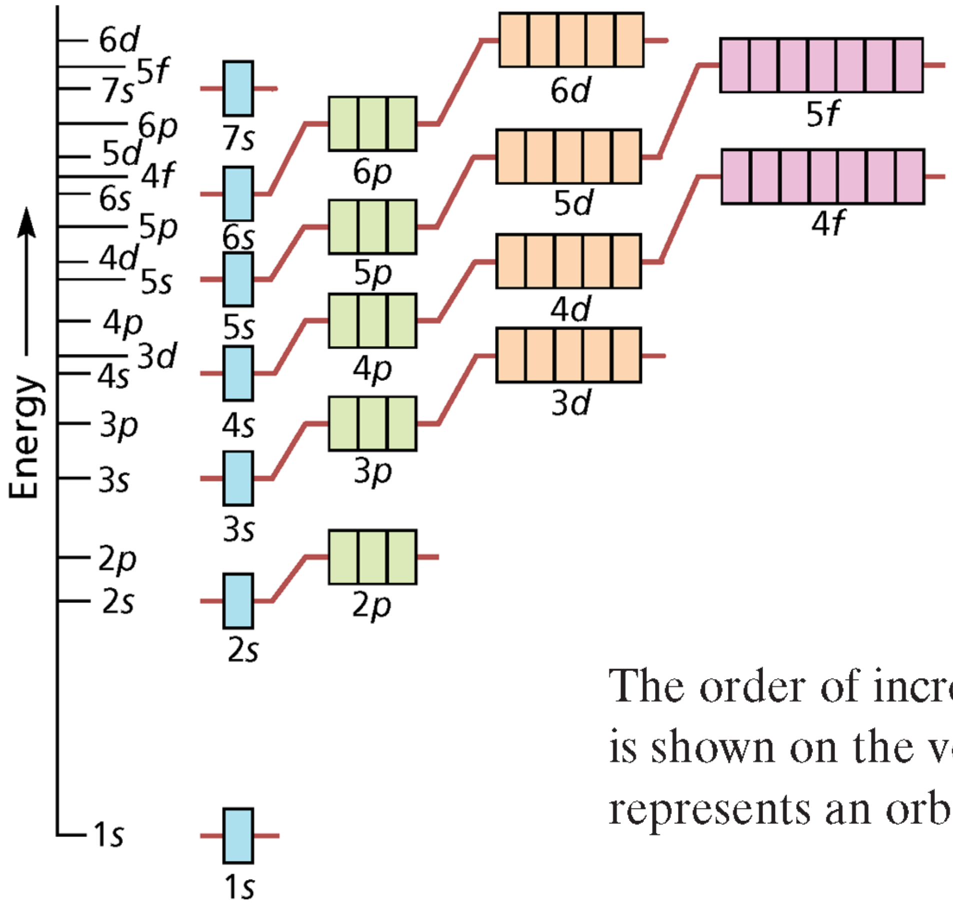
Relative Energies of Orbitals



The order of increasing energy for atomic sublevels is shown on the vertical axis. Each individual box represents an orbital.

Rules Governing Electron Configurations

- According to the **Aufbau principle,** an electron occupies the lowest-energy orbital that can receive it.
- According to the **Pauli exclusion principle,** no two electrons in the same atom can have the same set of four quantum numbers.

End Of Slide

Rules Governing Electron Configurations, *continued*

- According to **Hund's rule,** orbitals of equal energy are each occupied by one electron before any orbital is occupied by a second electron, and all electrons in singly occupied orbitals must have the same spin state.

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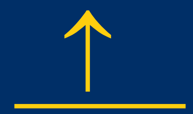
Representing Electron Configurations

Orbital Notation

- An unoccupied orbital is represented by a line, with the orbital’s name written underneath the line.

- An orbital containing one electron is represented as:



Representing Electron Configurations, *continued*

Orbital Notation

- An orbital containing two electrons is represented as:

$$\underline{\uparrow\downarrow}$$

- The lines are labeled with the principal quantum number and sublevel letter. For example, the orbital notation for helium is written as follows:

$$\text{He}\ \underset{1s}{\underline{\uparrow\downarrow}}$$

Representing Electron Configurations, *continued*

Electron-Configuration Notation

- Electron-configuration notation eliminates the lines and arrows of orbital notation.
- Instead, the number of electrons in a sublevel is shown by adding a superscript to the sublevel designation.
- The helium configuration is represented by $1s^2$.
- The superscript indicates that there are two electrons in helium's 1*s* orbital.

End Of Slide

Representing Electron Configurations, *continued*

Sample Problem A

The electron configuration of boron is $1s^22s^22p^1$. How many electrons are present in an atom of boron? What is the atomic number for boron? Write the orbital notation for boron.

End of Slide

Representing Electron Configurations, *continued*

Sample Problem A Solution

The number of electrons in a boron atom is equal to the sum of the superscripts in its electron-configuration notation: 2 + 2 + 1 = 5 electrons. The number of protons equals the number of electrons in a neutral atom. So we know that boron has 5 protons and thus has an atomic number of 5. To write the orbital notation, first draw the lines representing orbitals.

Representing Electron Configurations, *continued*

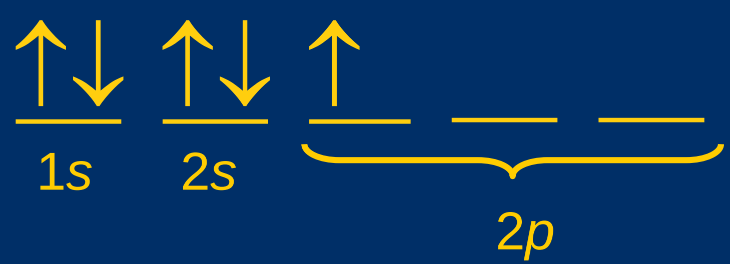
Sample Problem A Solution, *continued*

Next, add arrows showing the electron locations. The first two electrons occupy n = 1 energy level and fill the 1*s* orbital.

Representing Electron Configurations, *continued*

Sample Problem A Solution, *continued*

The next three electrons occupy the *n* = 2 main energy level. Two of these occupy the lower-energy 2*s* orbital. The third occupies a higher-energy *p* orbital.



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Elements of the Second Period

- In the first-period elements, hydrogen and helium, electrons occupy the orbital of the first main energy level.
- According to the Aufbau principle, after the 1*s* orbital is filled, the next electron occupies the *s* sublevel in the second main energy level.

Elements of the Second Period, *continued*

- The *highest-occupied energy level* is the electron-containing main energy level with the highest principal quantum number.
- *Inner-shell electrons* are electrons that are not in the highest-occupied energy level.

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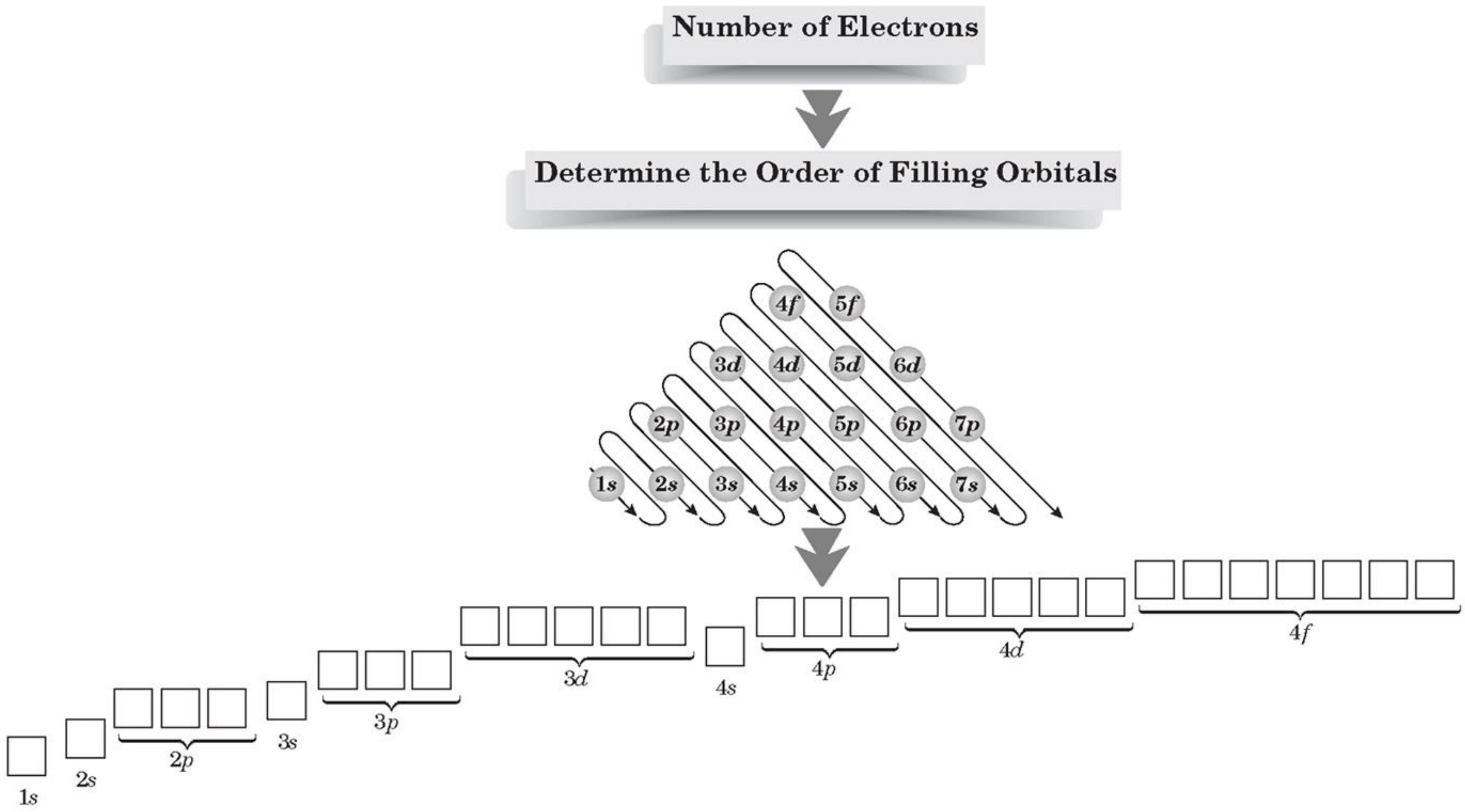
Writing Electron Configurations



Elements of the Third Period

- After the outer octet is filled in neon, the next electron enters the *s* sublevel in the *n* = 3 main energy level.

Noble-Gas Notation

- The Group 18 elements (helium, neon, argon, krypton, xenon, and radon) are called the **noble gases.**
- A **noble-gas configuration** refers to an outer main energy level occupied, in most cases, by eight electrons.

End Of Slide

Orbital Notation for Three Noble Gases

Noble gas	Helium, He	Neon, Ne	Argon, Ar
Third energy level			↑↓ 3*s*; ↑↓ ↑↓ ↑↓ 3*p*
Second energy level		↑↓ 2*s*; ↑↓ ↑↓ ↑↓ 2*p*	↑↓ 2*s*; ↑↓ ↑↓ ↑↓ 2*p*
First energy level	↑↓ 1*s*	↑↓ 1*s*	↑↓ 1*s*

Elements of the Fourth Period

- The period begins by filling the 4*s* orbital, the empty orbital of lowest energy.
- With the 4*s* sublevel filled, the 4*p* and 3*d* sublevels are the next available vacant orbitals.
- The 3*d* sublevel is lower in energy than the 4*p* sublevel. Therefore, the five 3*d* orbitals are next to be filled.

Orbital Notation for Argon and Potassium

Element	Argon, Ar	Potassium, K
Fourth energy level		↑ 4s
Third energy level	↑↓ 3s; ↑↓ ↑↓ ↑↓ 3p	↑↓ 3s; ↑↓ ↑↓ ↑↓ 3p
Second energy level	↑↓ 2s; ↑↓ ↑↓ ↑↓ 2p	↑↓ 2s; ↑↓ ↑↓ ↑↓ 2p
First energy level	↑↓ 1s	↑↓ 1s

Elements of the Fifth Period

- In the 18 elements of the fifth period, sublevels fill in a similar manner as in elements of the fourth period.
- Successive electrons are added first to the 5*s* orbital, then to the 4*d* orbitals, and finally to the 5*p* orbitals.

Sample Problem B

a. Write both the complete electron-configuration notation and the noble-gas notation for iron, Fe.

b. How many electron-containing orbitals are in an atom of iron? How many of these orbitals are completely filled? How many unpaired electrons are there in an atom of iron? In which sublevel are the unpaired electrons located?

Sample Problem B Solution

a. The complete electron-configuration notation of iron is $1s^22s^22p^63s^23p^63d^64s^2$. Iron's noble-gas notation is $[Ar]3d^64s^2$.

b. An iron atom has 15 orbitals that contain electrons.

They consist of one 1*s* orbital, one 2*s* orbital, three 2*p* orbitals, one 3*s* orbital, three 3*p* orbitals, five 3*d* orbitals, and one 4*s* orbital.

Eleven of these orbitals are filled, and there are four unpaired electrons.

They are located in the 3*d* sublevel.

The notation $3d^6$ represents 3*d* $\underline{\uparrow\downarrow}$ $\underline{\uparrow}$ $\underline{\uparrow}$ $\underline{\uparrow}$ $\underline{\uparrow}$.

Sample Problem C

a. Write both the complete electron-configuration notation and the noble-gas notation for a rubidium atom.

b. Identify the elements in the second, third, and fourth periods that have the same number of highest-energy-level electrons as rubidium.

Sample Problem C Solution

a. $1s^22s^22p^63s^23p^63d^{10}4s^24p^65s^1$, [Kr]$5s^1$

b. Rubidium has one electron in its highest energy level (the fifth). The elements with the same outermost configuration are,

in the second period, lithium, Li;

in the third period, sodium, Na;

and in the fourth period, potassium, K.

End of Chapter 4 Show