Chapter 5: Electrons in Atoms

5.1 Models of the Atom

The Development of Atomic Models

- **Rutherford's atomic model could not explain the chemical properties of elements.**

- Rutherford's atomic model could not explain why objects change color when heated.

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The Bohr Model

- Bohr proposed that an electron is found only in specific circular paths, or orbits, around the nucleus.

The Bohr Model

- Each possible electron orbit in Bohr's model has a fixed energy.
  - The fixed energies an electron can have are called **<u>energy levels</u>.**
  - A **<u>quantum</u>** of energy is the amount of energy required to move an electron from one energy level to another energy level.

The Bohr Model

- Like the rungs of the strange ladder, the energy levels in an atom are not equally spaced.
- The higher the energy level occupied by an electron, the less energy it takes to move from that energy level to the next higher energy level.

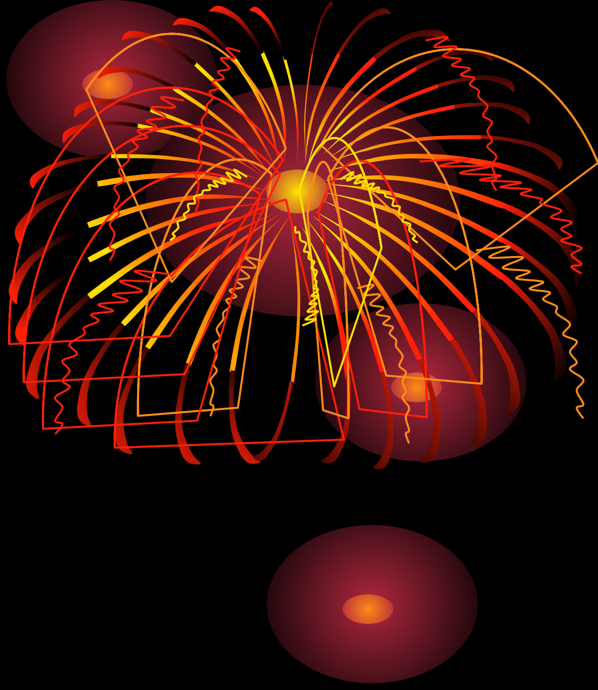
The Quantum Mechanical Model

- The quantum mechanical model determines the allowed energies an electron can have and how likely it is to find the electron in various locations around the nucleus.

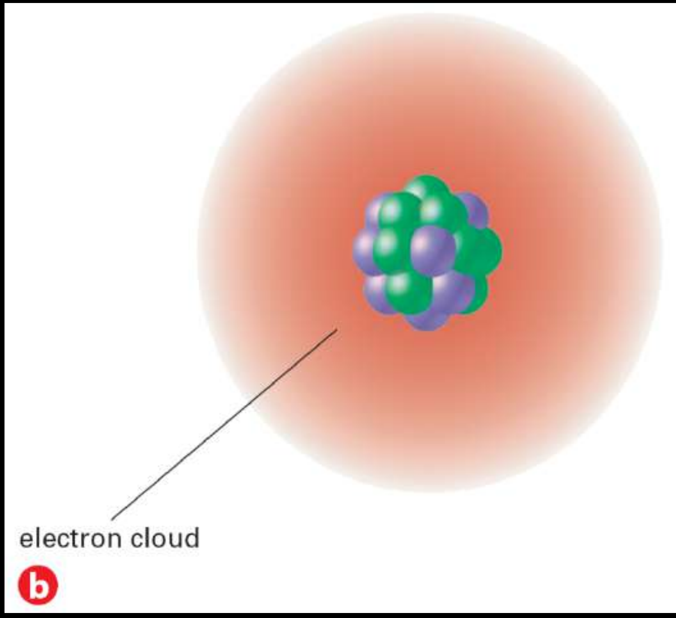
The Quantum Mechanical Model

- Austrian physicist Erwin Schrödinger (1887–1961) used new theoretical calculations and results to devise and solve a mathematical equation describing the behavior of the electron in a hydrogen atom.
- The modern description of the electrons in atoms, the **quantum mechanical model**, comes from the mathematical solutions to the Schrödinger equation.

The Quantum Mechanical Model



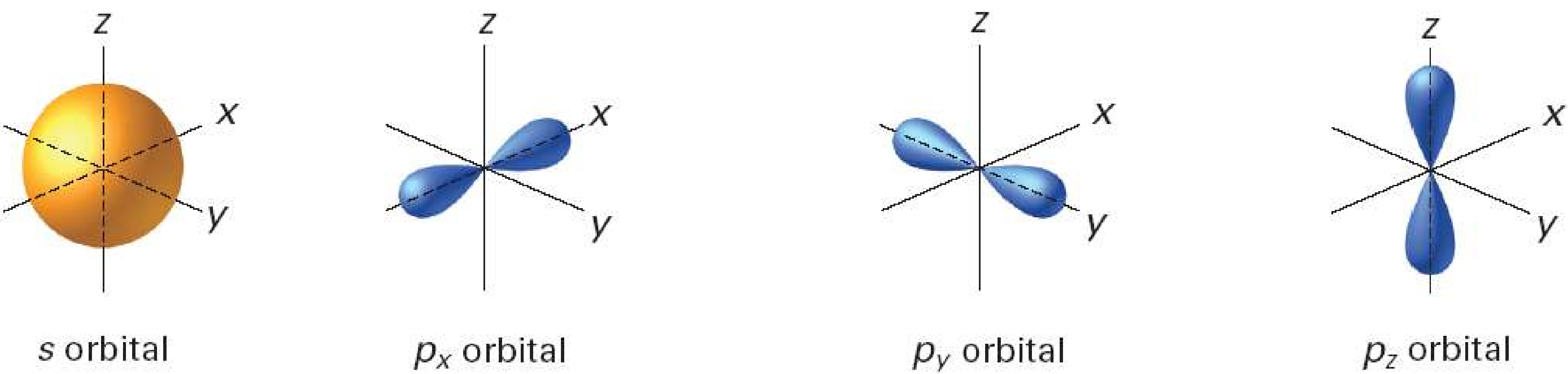
- In the quantum mechanical model, the probability of finding an electron within a certain volume of space surrounding the nucleus can be represented as a fuzzy cloud. The cloud is more dense where the probability of finding the electron is high.



Atomic Orbitals

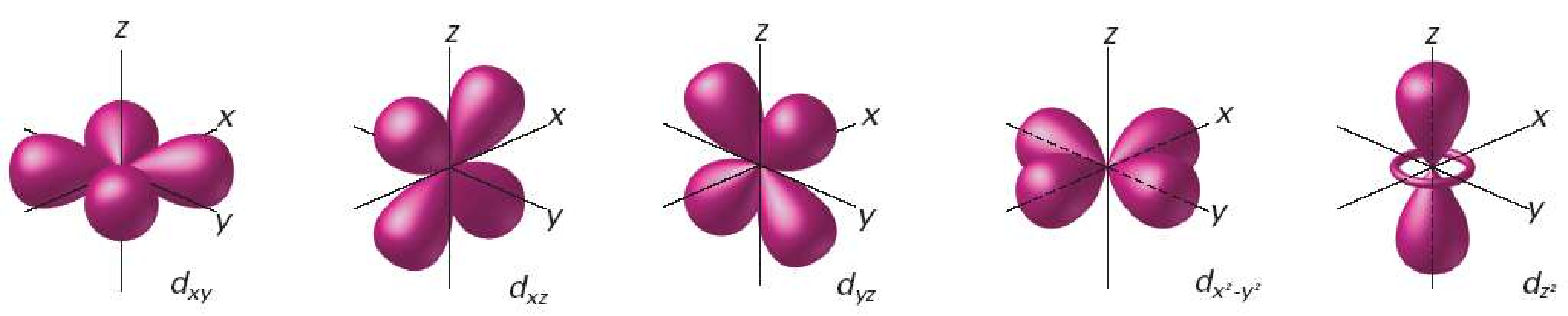
- An **atomic orbital** is often thought of as a region of space in which there is a high probability of finding an electron.
  - Each energy sublevel corresponds to an orbital of a different shape, which describes where the electron is likely to be found.

- Different atomic orbitals are denoted by letters. The *s* orbitals are spherical, and *p* orbitals are dumbbell-shaped.



Atomic Orbitals

- Four of the five *d* orbitals have the same shape but different orientations in space.



- The numbers and kinds of atomic orbitals depend on the energy sublevel.

Table 5.1

Summary of Principal Energy Levels, Sublevels, and Orbitals

Principal energy level	Number of sublevels	Type of sublevel
$n = 1$	1	1*s* (1 orbital)
$n = 2$	2	2*s* (1 orbital), 2*p* (3 orbitals)
$n = 3$	3	3*s* (1 orbital), 3*p* (3 orbitals), 3*d* (5 orbitals)
$n = 4$	4	4*s* (1 orbital), 4*p* (3 orbitals), 4*d* (5 orbitals), 4*f* (7 orbitals)

5.1

Table 5.2

Maximum Numbers of Electrons

Energy level *n*	Maximum number of electrons
1	2
2	8
3	18
4	32

- The number of electrons allowed in each of the first four energy levels are shown here.

Quantum numbers

- Principal quantum number (*n*) – main energy level occupied by the electron
  - Positive integers (1, 2, 3, . . .)
  - As *n* increases, the electron's energy and its average distance from the nucleus increase.

Quantum Numbers

- <u>Angular momentum quantum number (l)</u> indicates the shape of the orbital
  - the number of orbital shapes possible is equal to *n*
  - Can be zero up to all positive integers less than or equal to *n* - 1

l	Letter
0	*s*
1	*p*
2	*d*
3	*f*

Quantum Numbers

- Magnetic quantum number (*m*) – orientation of an orbital around the nucleus
    - Values are whole numbers from –*l* to +*l* including zero

Orbital shape	*m* values
s	0
p	-1, 0, +1
d	-2, -1, 0, +1, +2
f	-3, -2, -1, 0, +1, +2, +3

Quantum Numbers

- Spin quantum number (m_s) – indicates the two fundamental spin states of an electron in an orbital
  - Values of +1/2 or -1/2
  - A single orbital can hold a maximum of two electrons

5.1 Section Quiz.

- 1.Rutherford's planetary model of the atom could not explain

  a) any properties of elements.

  b) the chemical properties of elements.

  c) the distribution of mass in an atom.

  d) the distribution of positive and negative charges in an atom.

5.1 Section Quiz.

- 2.Bohr's model of the atom proposed that electrons are found

  a) embedded in a sphere of positive charge.

  b) in fixed positions surrounding the nucleus.

  c) in circular orbits at fixed distances from the nucleus.

  d) orbiting the nucleus in a single fixed circular path.

5.1 Section Quiz.

- 3.What is the lowest-numbered principal energy level in which *p* orbitals are found?

  a) 1

  b) 2

  c) 3

  d) 4

5.2 Electron Arrangement in Atoms

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Electron Configurations

- The ways in which electrons are arranged in various orbitals around the nuclei of atoms are called **electron configurations.**

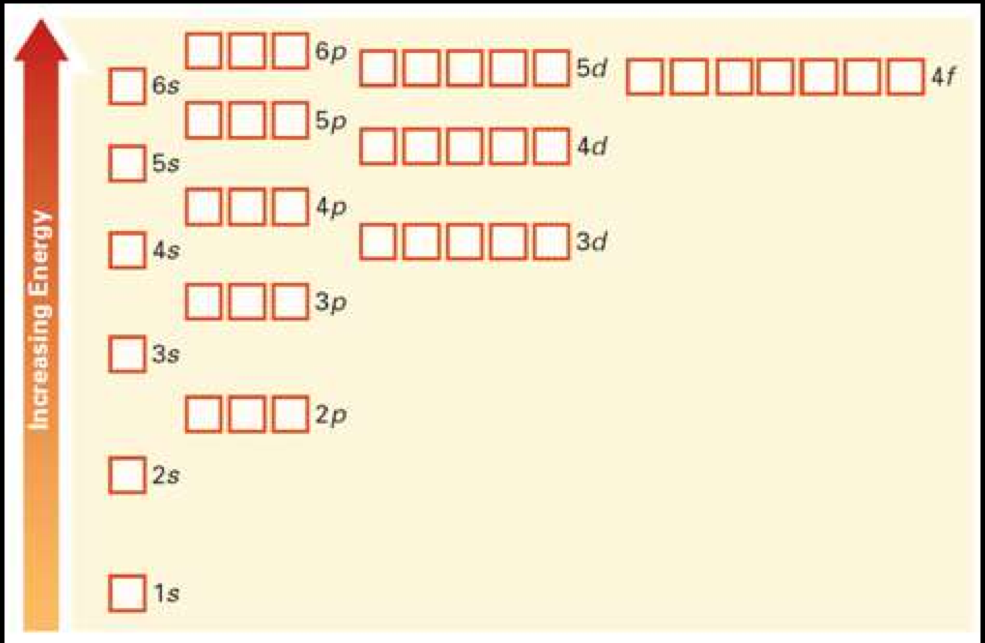
Three rules—the aufbau principle, the Pauli exclusion principle, and Hund's rule—tell you how to find the electron configurations of atoms.

- Aufbau Principle
  - According to the **aufbau principle**, electrons occupy the orbitals of lowest energy first. In the aufbau diagram below, each box represents an atomic orbital.

Lazy

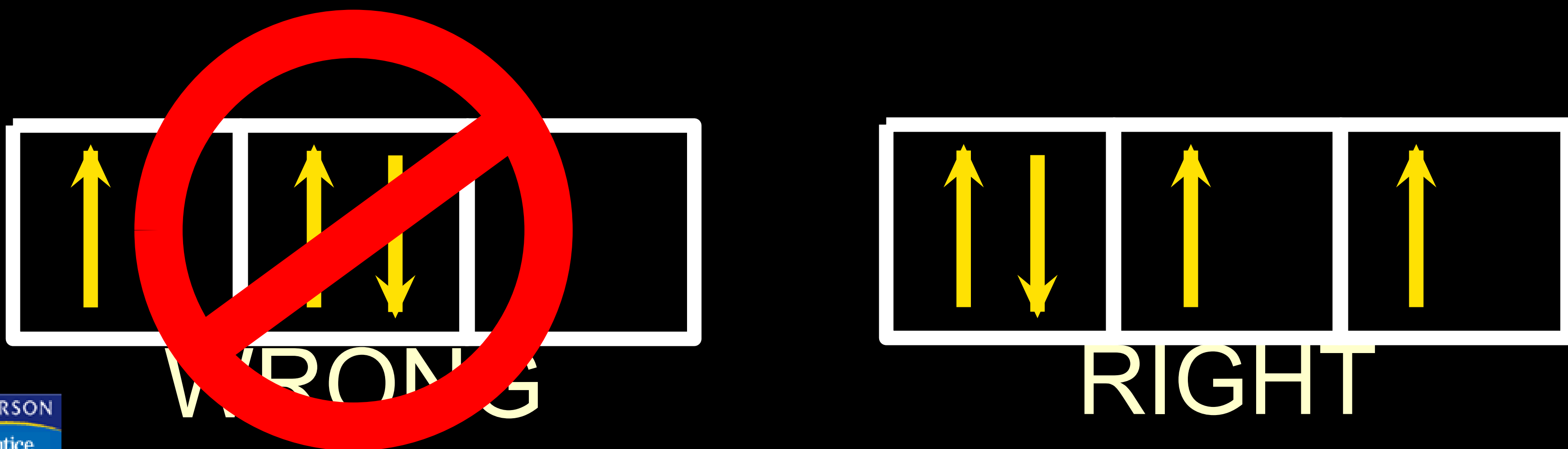
Tenant

Rule



- Pauli Exclusion Principle
  - According to the **<u>Pauli exclusion principle</u>**, an atomic orbital may describe at most two electrons. To occupy the same orbital, two electrons must have opposite spins; that is, the electron spins must be paired.

- **<u>Hund's Rule</u>**
  - Within a sublevel, place one e$^-$ per orbital before pairing them.
  - "Empty Bus Seat Rule"



Conceptual Problem 5.1

Writing Electron Configurations

Phosphorus, an element used in matches, has an atomic number of 15. Write the electron configuration of a phosphorus atom.

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for Conceptual Problem 5.1

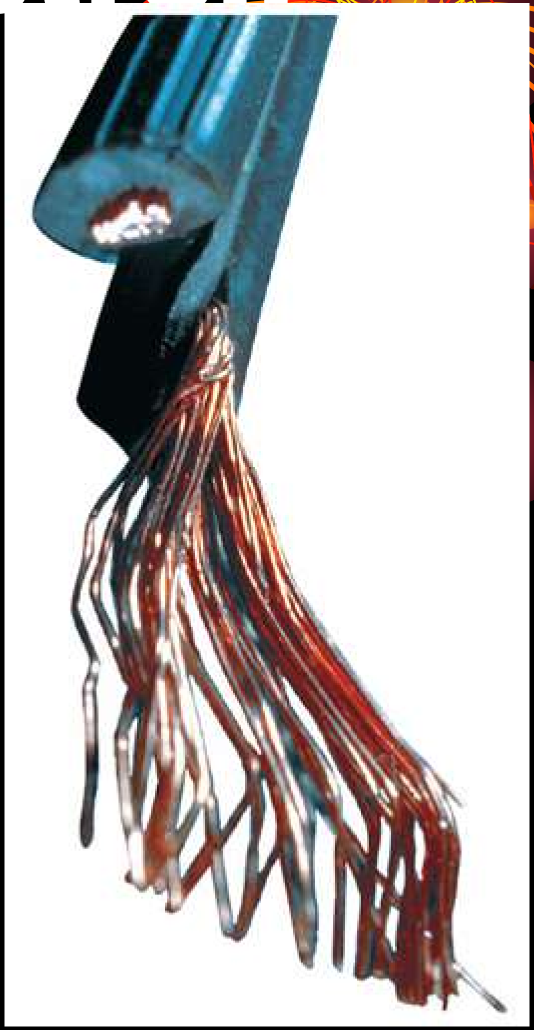
9. Write the electron configuration for each atom. How many unpaired electrons does each atom have?

a. boron **b.** silicon

Exceptional Electron Configurations

- Some actual electron configurations differ from those assigned using the aufbau principle because half-filled sublevels are not as stable as filled sublevels, but they are more stable than other configurations.

Exceptional Electron Configurations



- Exceptions to the aufbau principle are due to subtle electron-electron interactions in orbitals with very similar energies.
- Copper has an electron configuration that is an exception to the aufbau principle.

5.2 Section Quiz.

- 1. Identify the element that corresponds to the following electron configuration: $1s^2 2s^2 2p^5$.

   a) F

   b) Cl

   c) Ne

   d) O

5.2 Section Quiz.

- 2. Write the electron configuration for the atom N.

  a) $1s^2 2s^2 2p^5$

  b) $1s^2 2s^2 2p^3$

  c) $1s^2 2s 1p^2$

  d) $1s^2 2s^2 2p^1$

5.3 Physics and the Quantum Mechanical Model

Light

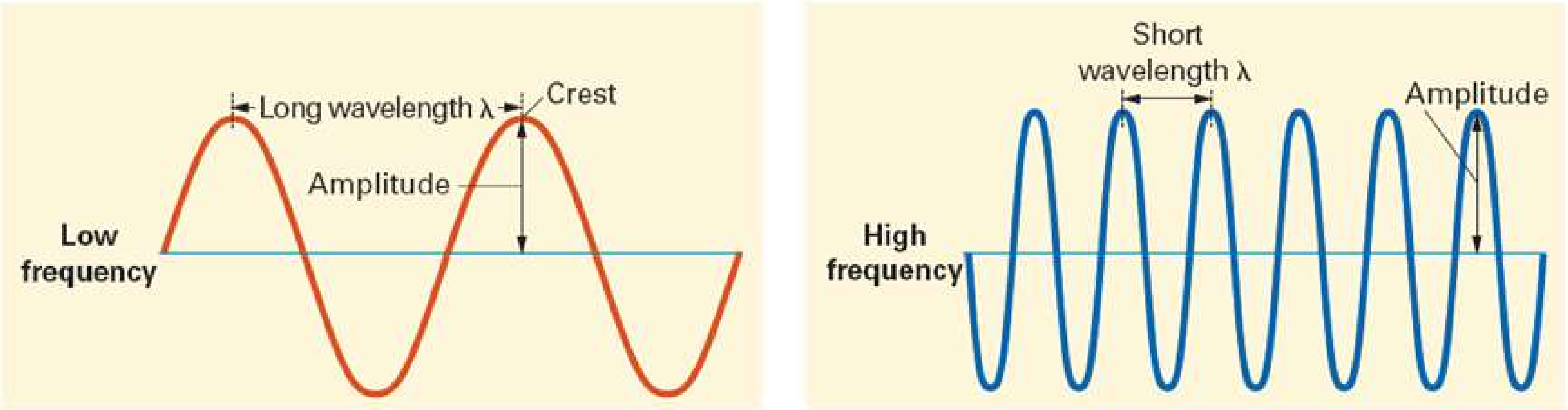
- The **amplitude** of a wave is the wave's height from zero to the crest.
- The **wavelength**, represented by λ (the Greek letter lambda), is the distance between the crests.

Light

- The **frequency**, represented by ν (the Greek letter nu), is the number of wave cycles to pass a given point per unit of time.
- The SI unit of cycles per second is called a **hertz** (Hz).

Light

- The wavelength and frequency of light are inversely proportional to each other.



Light

- The product of the frequency and wavelength always equals a constant (*c*), the speed of light.

$$c = \lambda \nu$$

Light

- According to the wave model, light consists of electromagnetic waves.
  - <u>**Electromagnetic radiation**</u> includes radio waves, microwaves, infrared waves, visible light, ultraviolet waves, X-rays, and gamma rays.
  - All electromagnetic waves travel in a vacuum at a speed of 2.998×10^8 m/s.

Light

- Sunlight consists of light with a continuous range of wavelengths and frequencies.
  - When sunlight passes through a prism, the different frequencies separate into a **spectrum** of colors.
  - In the visible spectrum, red light has the longest wavelength and the lowest frequency.

Sample Problem 5.1

Calculating the Wavelength of Light

Calculate the wavelength of the yellow light emitted by the sodium lamp shown above if the frequency of the radiation is 5.10×10^{14} Hz (5.10×10^{14}/s).

Practice Problem 14

- What is the wavelength of radiation with a frequency of 1.50×10^{13} Hz? Does this radiation have a longer or shorter wavelength than red light?

for Sample Problem 5.1

15. What is the frequency of radiation with a wavelength of 5.00×10^{-8} m? In what region of the electromagnetic spectrum is this radiation?

Wavelength

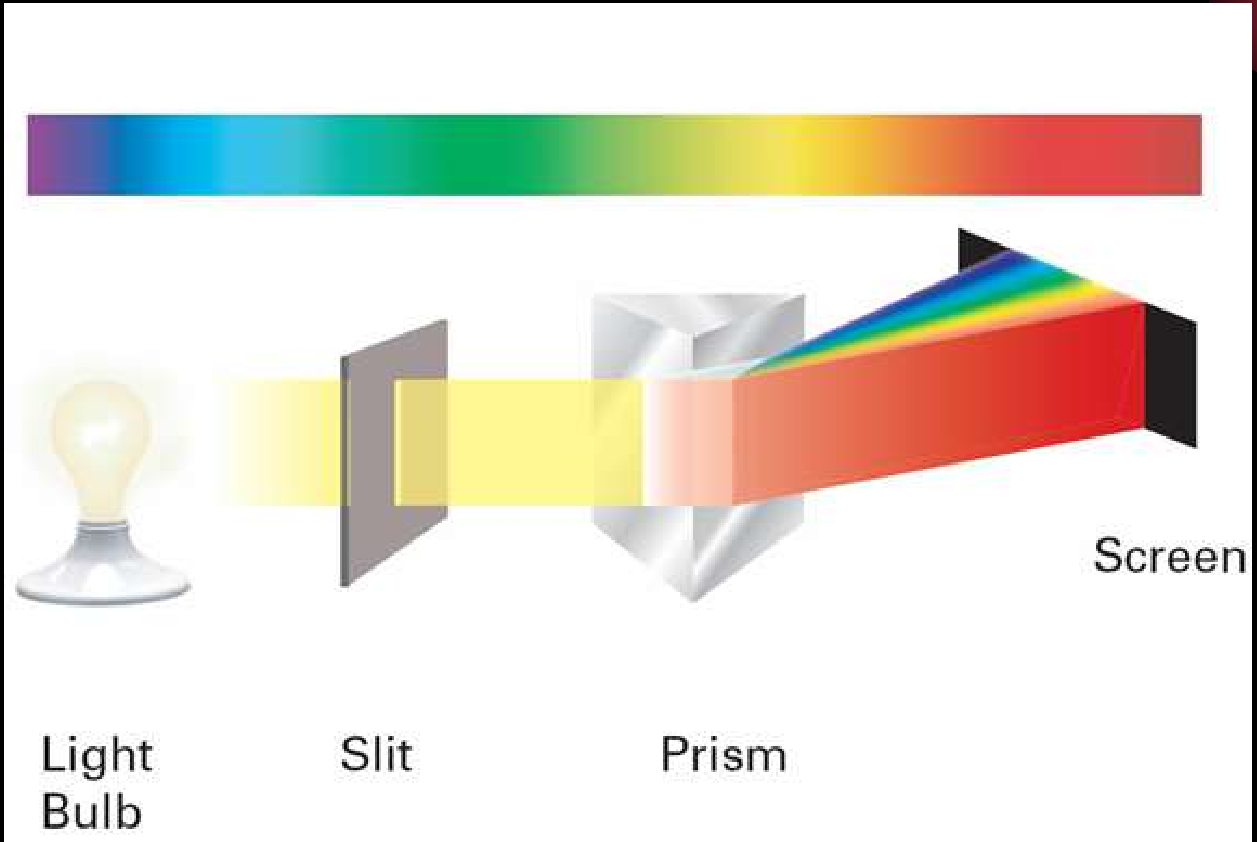
- While an FM radio station broadcasts at a frequency of 97.1 MHz, an AM station broadcasts at a frequency of 750 kHz. What are the wavelengths of the two broadcasts?

Atomic Spectra

- When atoms absorb energy, electrons move into higher energy levels. These electrons then lose energy by emitting light when they return to lower energy levels.

Atomic Spectra

- A prism separates light into the colors it contains. When white light passes through a prism, it produces a rainbow of colors.

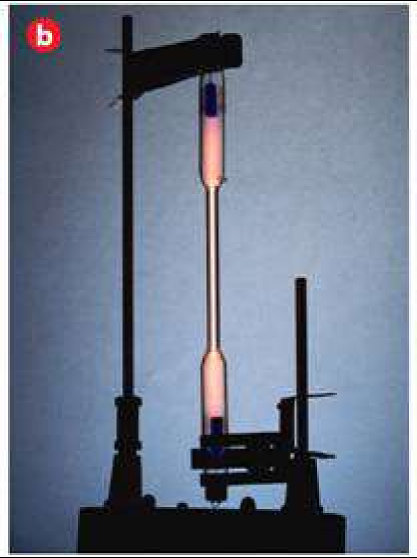


Atomic Spectra

- The frequencies of light emitted by an element separate into discrete lines to give the **atomic emission spectrum** of the element.

Mercury

Nitrogen



An Explanation of Atomic Spectra

- In the Bohr model, the lone electron in the hydrogen atom can have only certain specific energies.
  - When the electron has its lowest possible energy, the atom is in its **ground state.**
  - Excitation of the electron by absorbing energy raises the atom from the ground state to an **excited state**.
  - A quantum of energy in the form of light is emitted when the electron drops back to a lower energy level.

An Explanation of Atomic Spectra

- The light emitted by an electron moving from a higher to a lower energy level has a frequency directly proportional to the energy change of the electron.
  - $E = h\nu$ where h is Planck's constant (6.626×10^{-34} J·s)

Practice Problems

1) What is the energy of a photon from the violet portion of the Sun's light if it has a frequency of 7.23×10^{14} Hz?

2) How much energy is in a photon with the following frequencies:

- A) 6.32×10^{20} Hz
- B) 9.50×10^{13} Hz

Energy of a Photon

- The blue color in some fireworks occur when copper (I) chloride is heated to 1500 K and emits blue light of wavelength 4.50×10^2 nm. How much energy does one photon of this light carry?
- The microwaves used to heat food have a wavelength of 0.125 m. What is the energy of one photon of microwave radiation?

Energy of a photon

- It takes 8.17×10^{-19} J of energy to remove one electron from a gold surface. What is the maximum wavelength of light capable of causing this effect?

Quantum Mechanics

- In 1905, Albert Einstein successfully explained experimental data by proposing that light could be described as quanta of energy.
  - The quanta behave as if they were particles.
  - Light quanta are called **<u>photons</u>.**
- In 1924, De Broglie developed an equation that predicts that all moving objects have wavelike behavior.

Quantum Mechanics

- Classical mechanics adequately describes the motions of bodies much larger than atoms, while quantum mechanics describes the motions of subatomic particles and atoms as waves.

Quantum Mechanics

- The **Heisenberg uncertainty principle** states that it is impossible to know exactly both the velocity and the position of a particle at the same time.
  - This limitation is critical in dealing with small particles such as electrons.
  - This limitation does not matter for ordinary-sized object such as cars or airplanes.

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5.3 Section Quiz.

- 1.Calculate the frequency of a radar wave with a wavelength of 125 mm.

  a) $2.40 \quad 10^{9}$ Hz

  b) $2.40 \quad 10^{24}$ Hz

  c) $2.40 \quad 10^{6}$ Hz

  d) $2.40 \ 10^{2}$ Hz

5.3 Section Quiz.

- 2.The lines in the emission spectrum for an element are caused by

  a) the movement of electrons from lower to higher energy levels.

  b) the movement of electrons from higher to lower energy levels.

  c) the electron configuration in the ground state.

  d) the electron configuration of an atom.

5.3 Section Quiz.

- 3.Spectral lines in a series become closer together as n increases because the

  a) energy levels have similar values.

  b) energy levels become farther apart.

  c) atom is approaching ground state.

  d) electrons are being emitted at a slower rate.