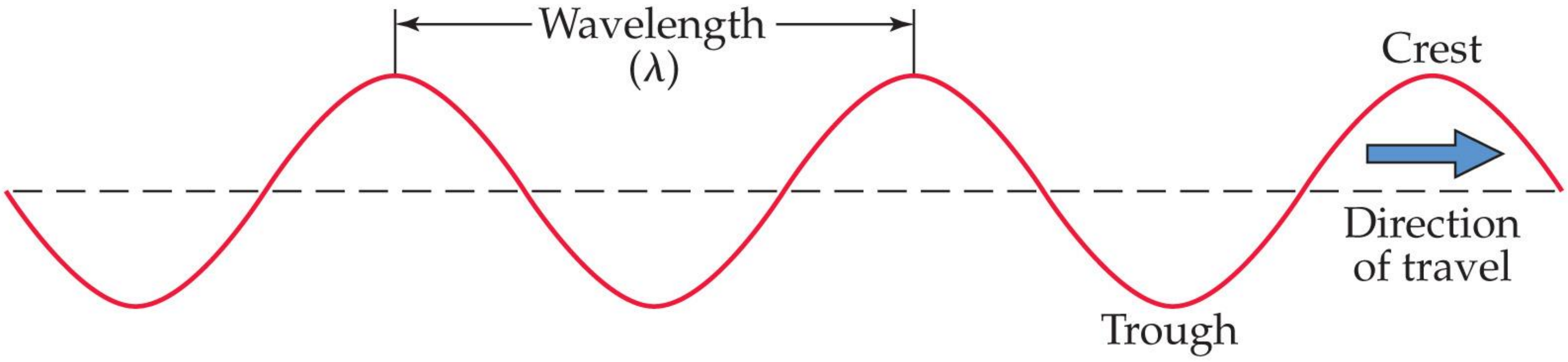
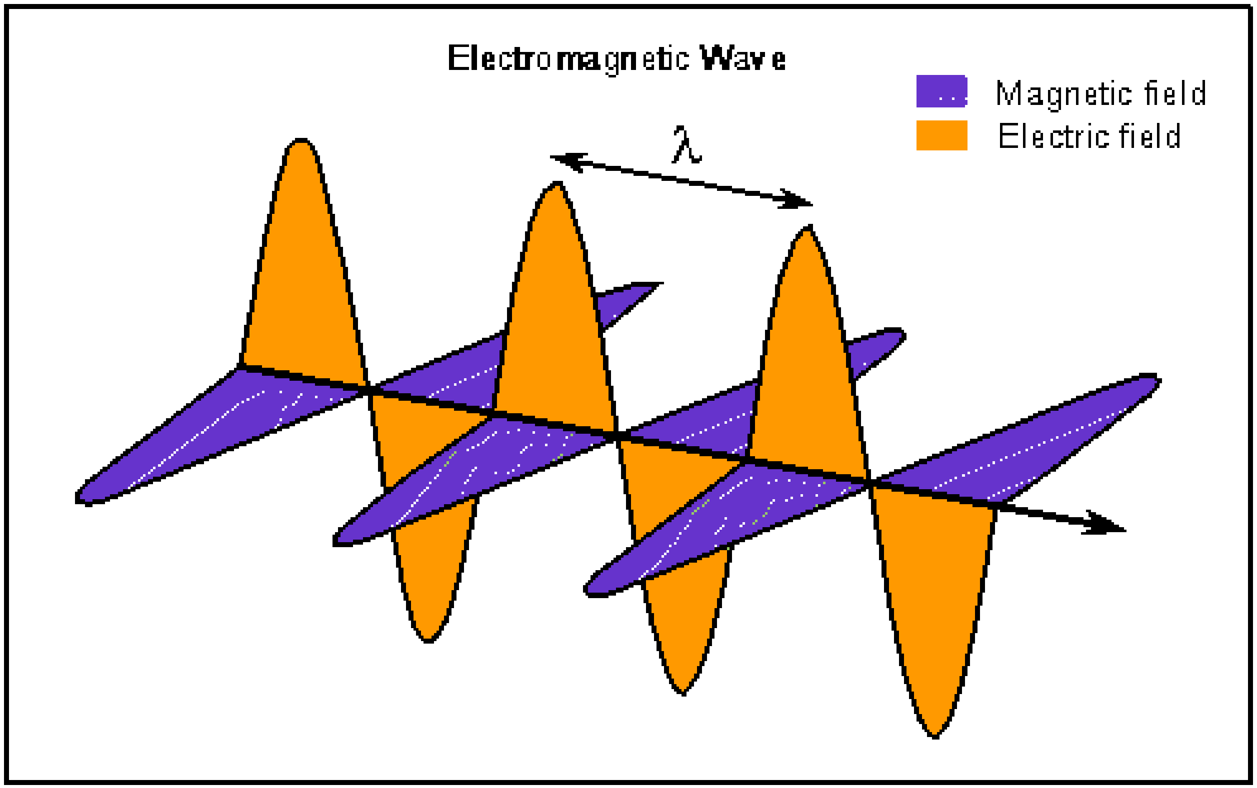
Module 5
Part 1: Light and Quantum Mechanics

The Wave Nature of Light



Frequency



$$c = \lambda \nu$$

Example

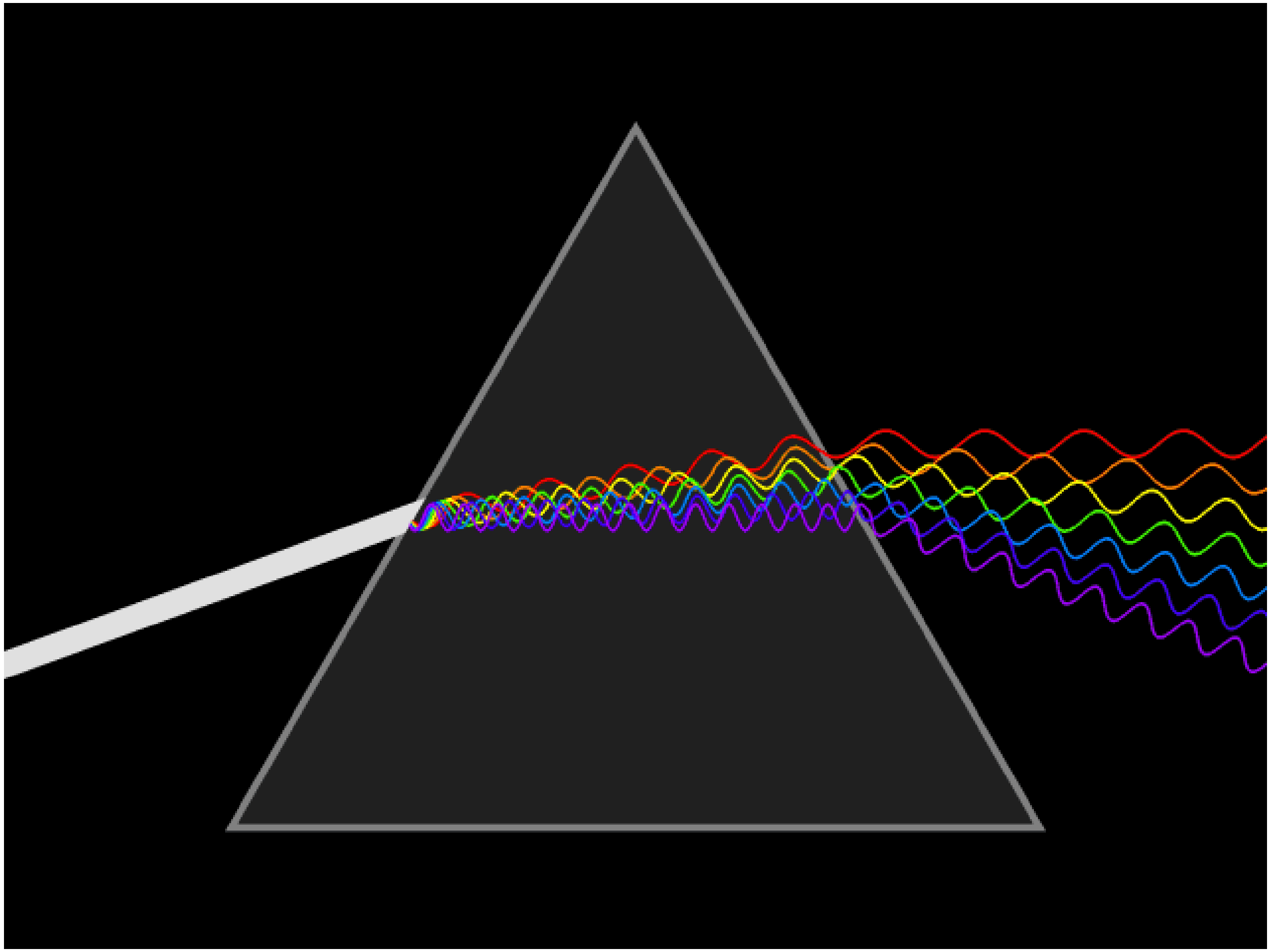
- Calculate the wavelength (in nm) of the red light emitted from a barcode scanner that has a frequency of 4.62×10^{14} Hz.

Example

- A laser emits UV light at 355nm. What is the frequency in Hz?

Light can Bounce (Reflection)

Light can Bend (Refraction)



The Particle Nature of Light

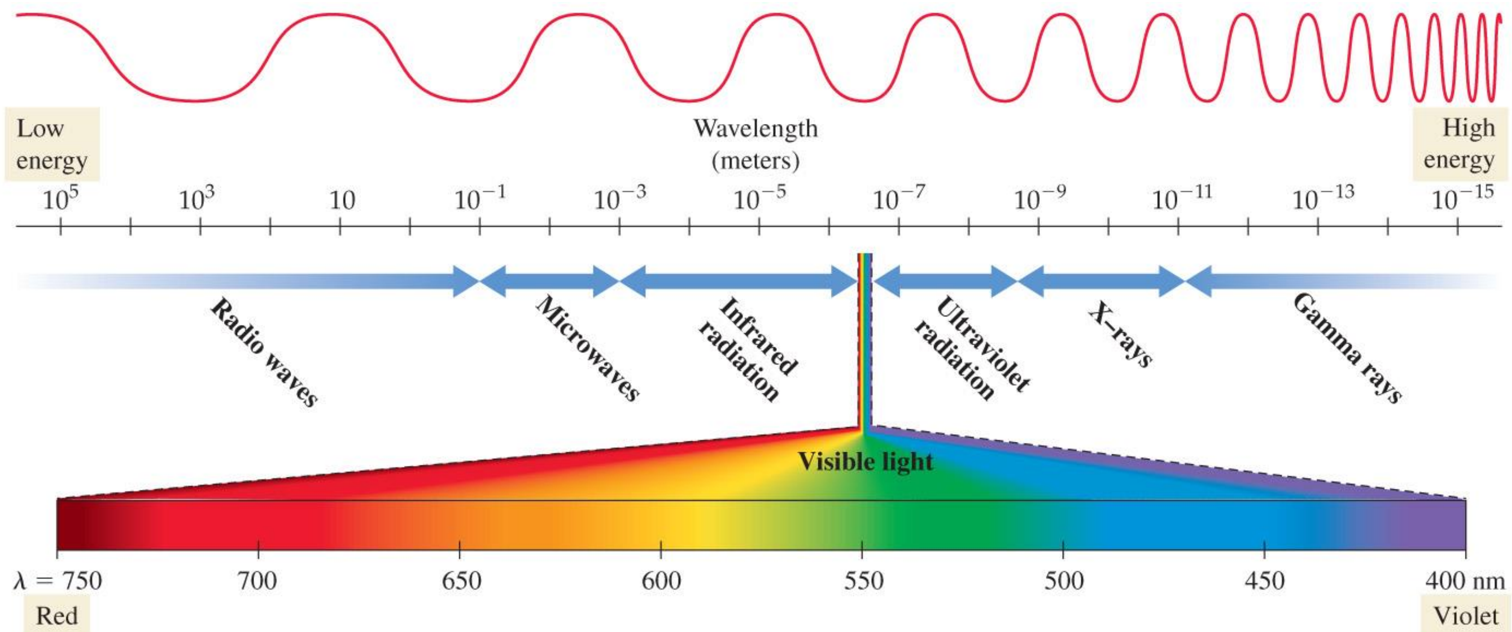
- Light is **quantized.** This means that light energy comes in small packets.

Energy of a Photon

$$E=hv$$

9.3 The Electromagnetic Spectrum

- The entire electromagnetic spectrum, with short-wavelength, high-frequency radiation on the right and long-wavelength, low-frequency radiation on the left. Visible light is the small sliver in the middle.



Example

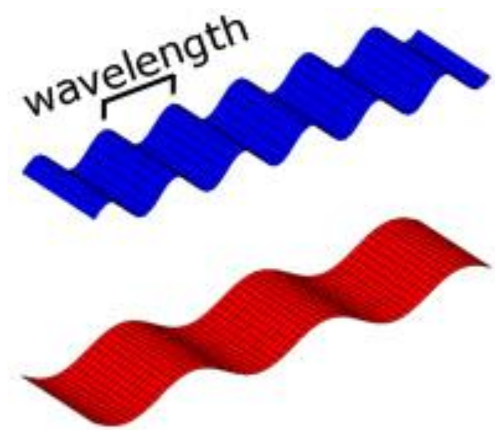
- A molecule's bond dissociates at 350 kJ/mol. What is the wavelength (in nm) of the light needed to break this bond?

Why is the Sky Blue?

Why is the Sky Blue?

FACT:

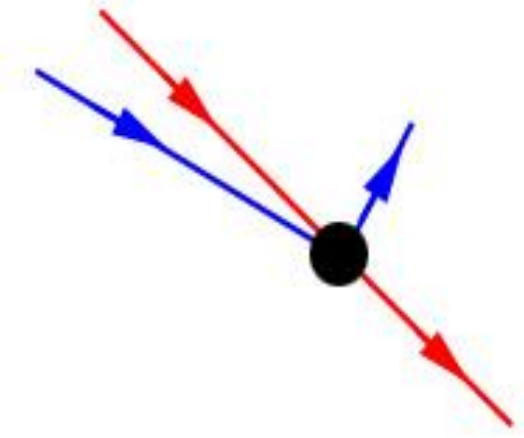
Light travels in waves. The light's wavelength determines its color. Short wavelength light, for example, appears blue, and long wavelength light appears red.



Why is the Sky Blue?

FACT:

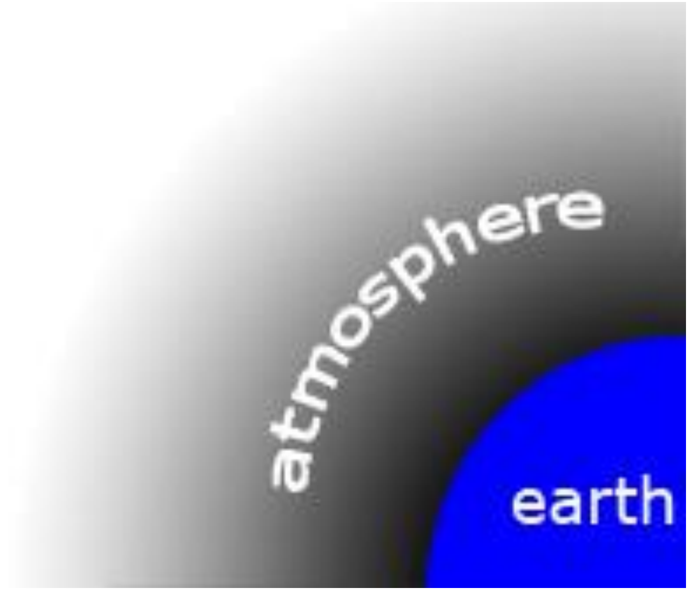
When light strikes particles that are smaller than its wavelength, the some of the light may be scattered. The shorter the wavelength, the more this scattering occurs.



Why is the Sky Blue?

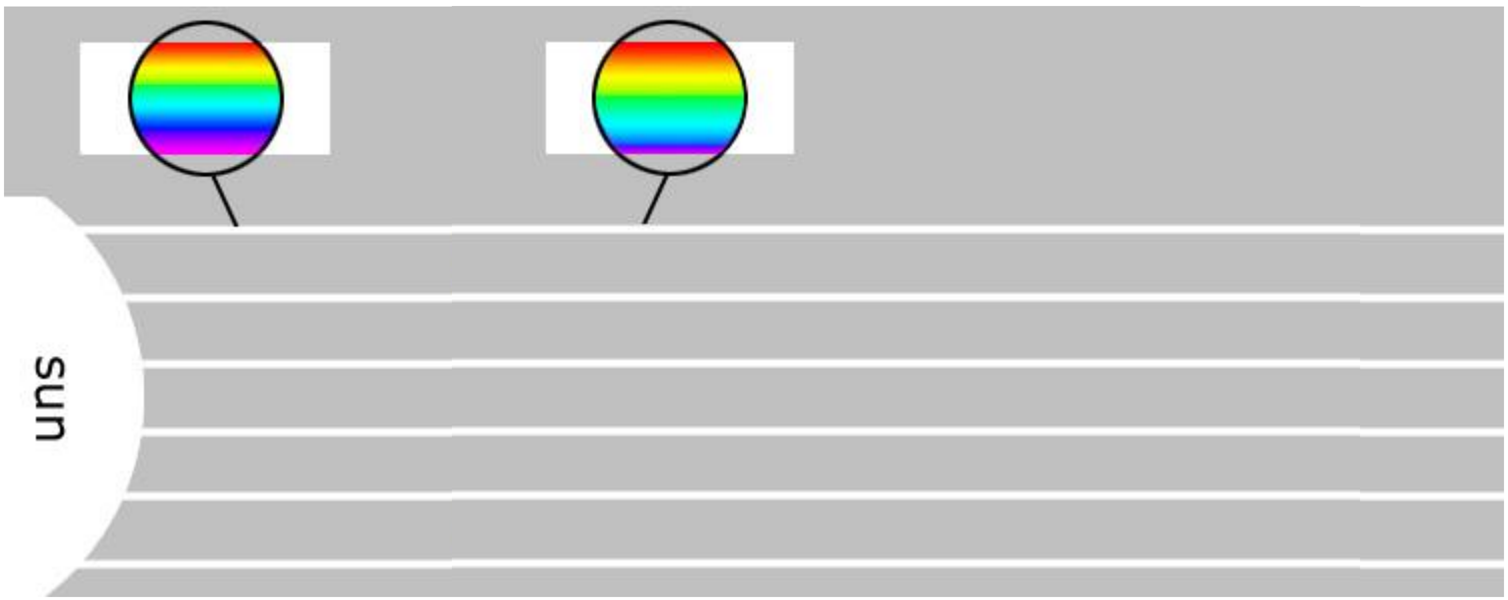
FACT:

The atmosphere contains many particles and gases, mainly nitrogen and oxygen.



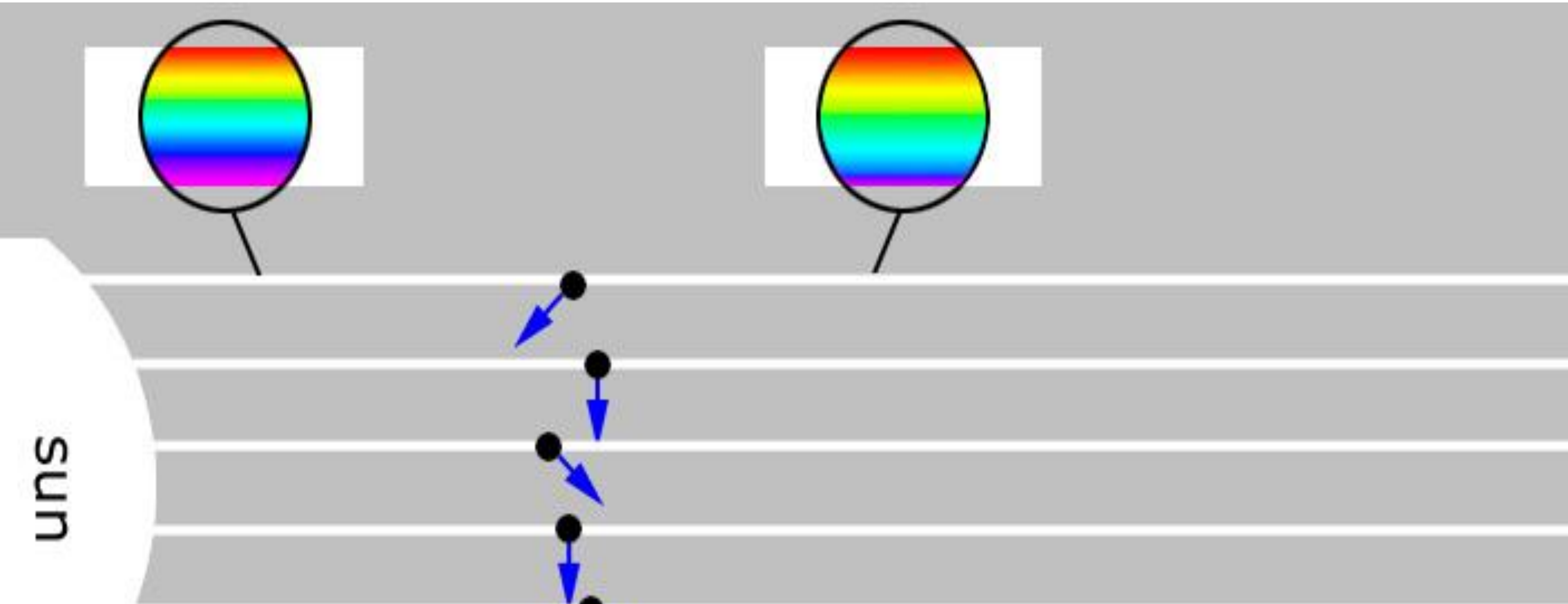
Why is the Sky Blue?

- Sunlight is composed of light of many different wavelengths. Longer wavelength light appears red, orange, and yellow, while shorter wavelength light appears blue, indigo and violet.



Why is the Sky Blue?

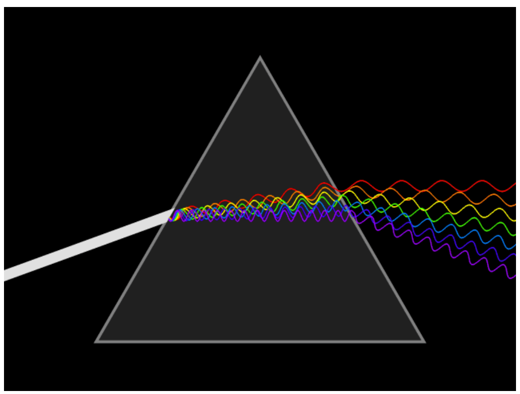
- Gas molecules in the atmosphere scatter, in all directions, the short wavelength light that appears blue to us. Longer wavelength light is largely unaffected as it passes through the atmosphere. As a result, when you look at the sky, you see blue everywhere.



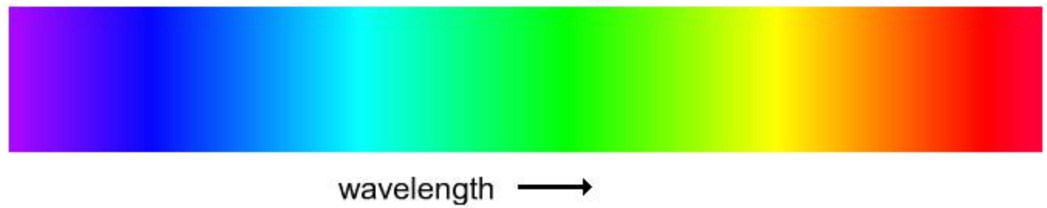
Why not Violet?

- There is less violet light than blue in sunlight.
- However, the primary reason for this is that our eyes are better at detecting blue light than they are at detecting violet light.

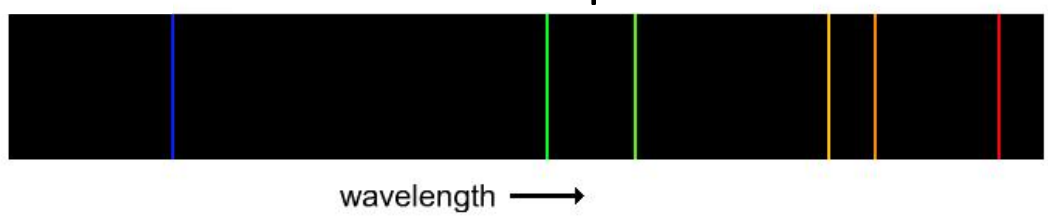
What is a Spectrum?



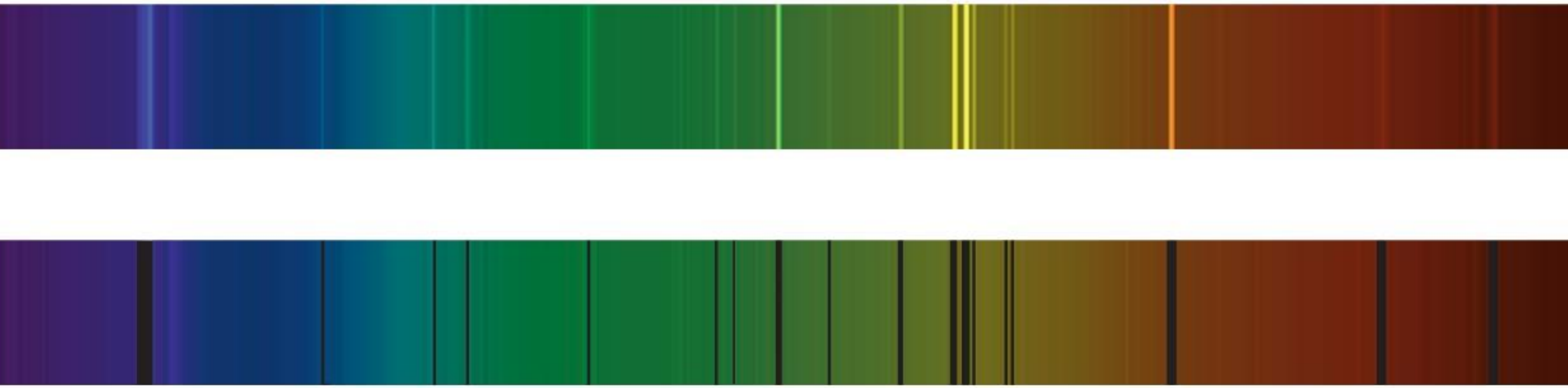
Continuous Spectrum



Discrete Line Spectrum



Emission Spectrum



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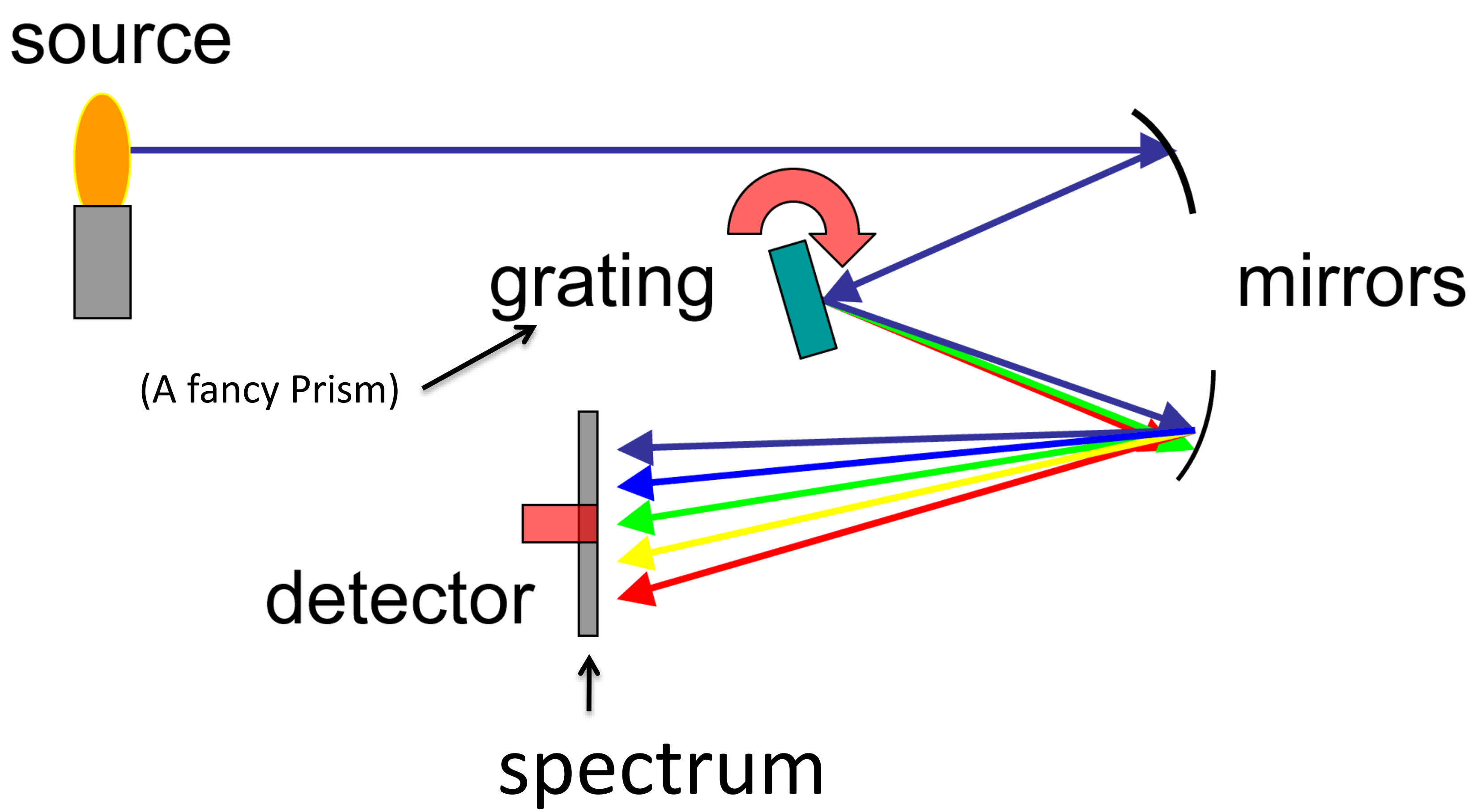
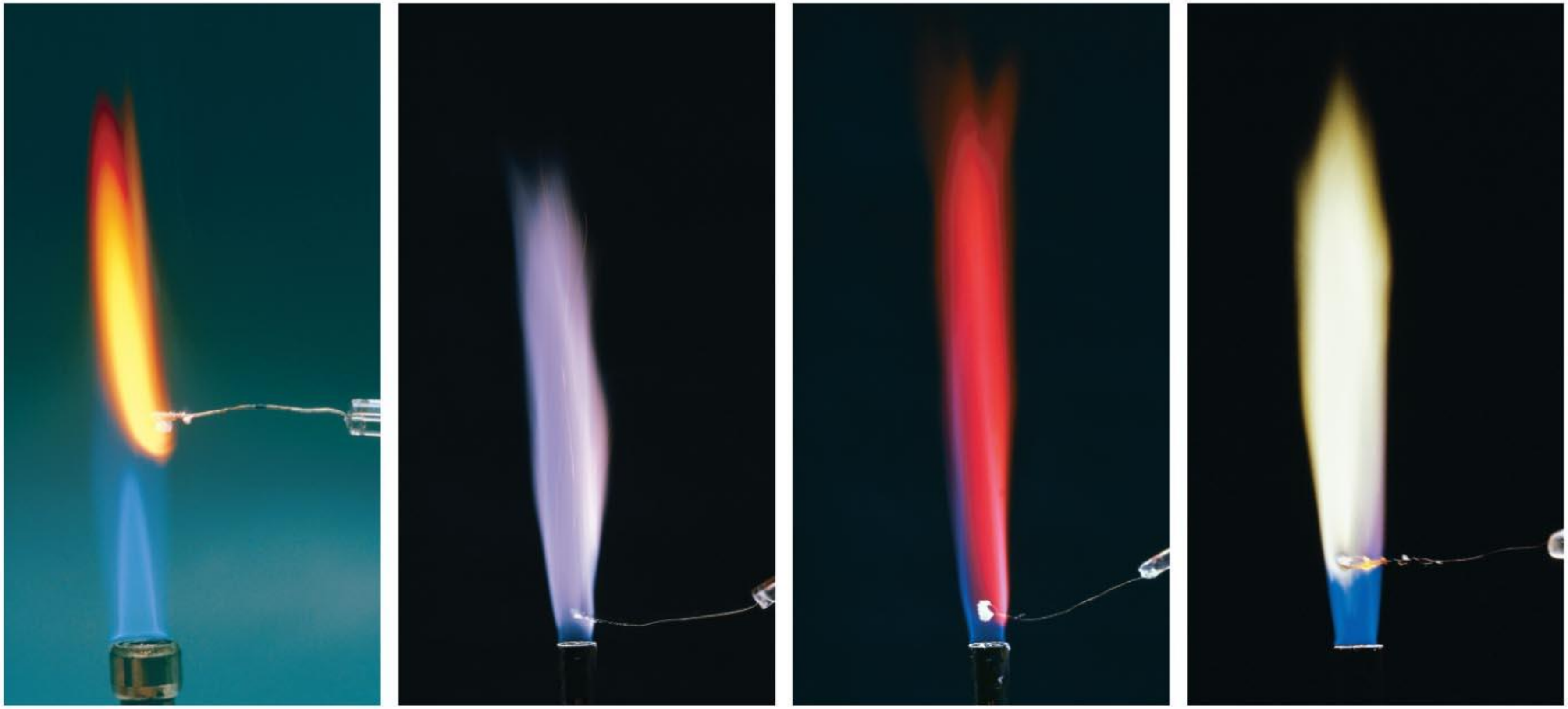
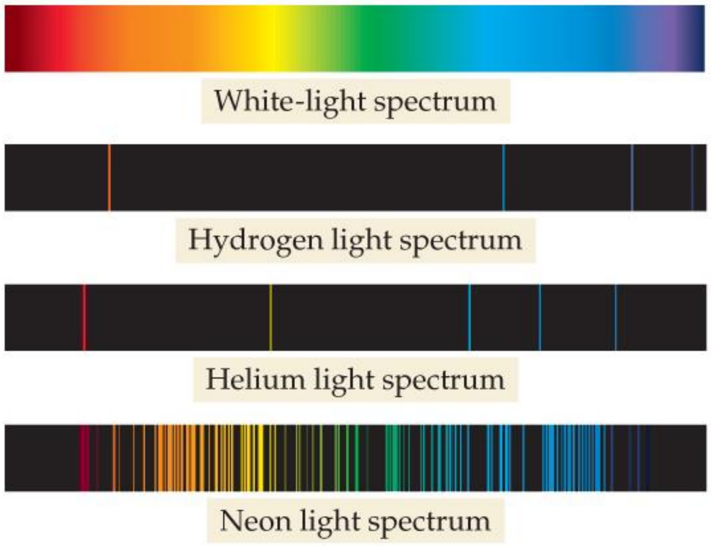
Absorption Spectrum

Absorption Spectrometer

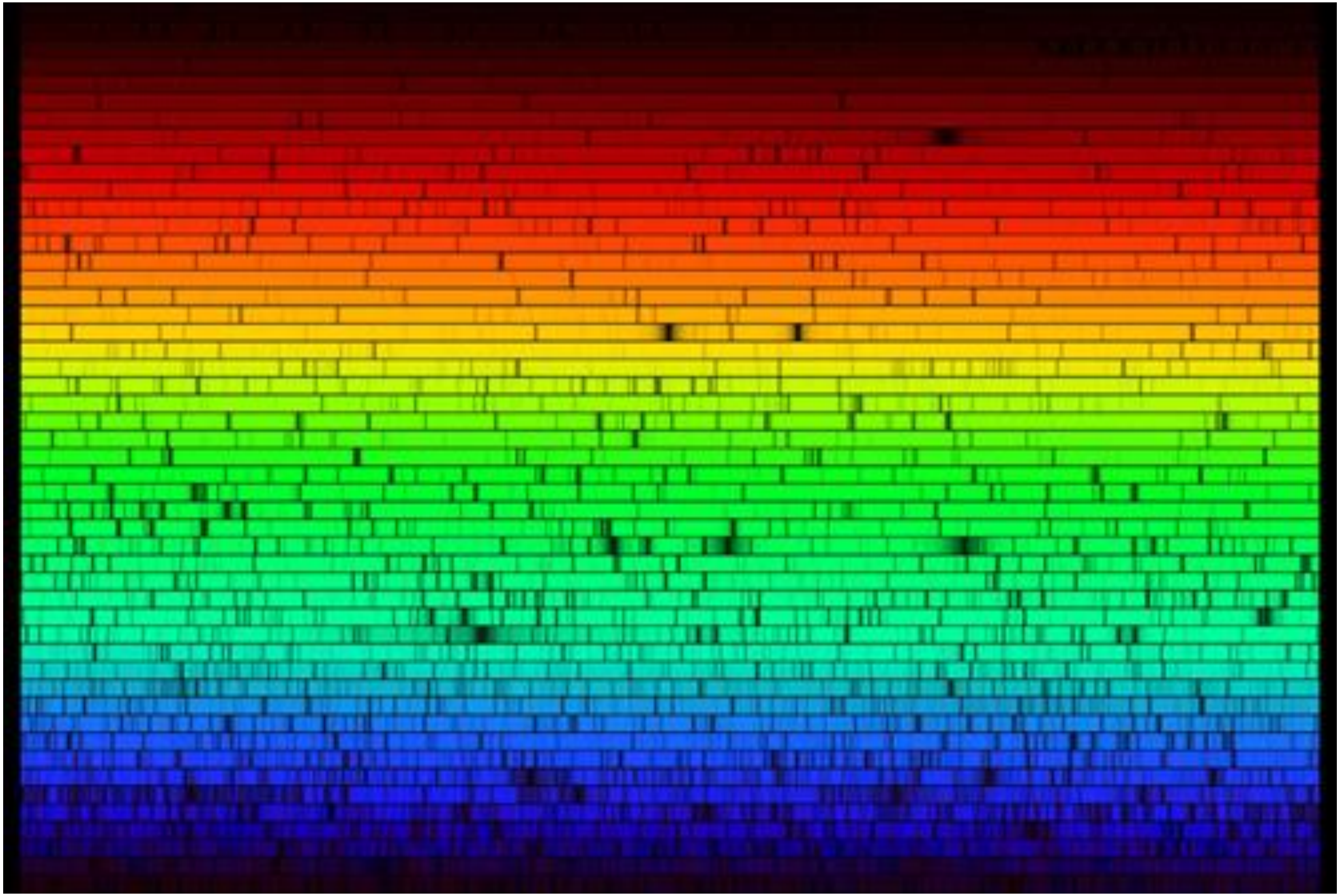




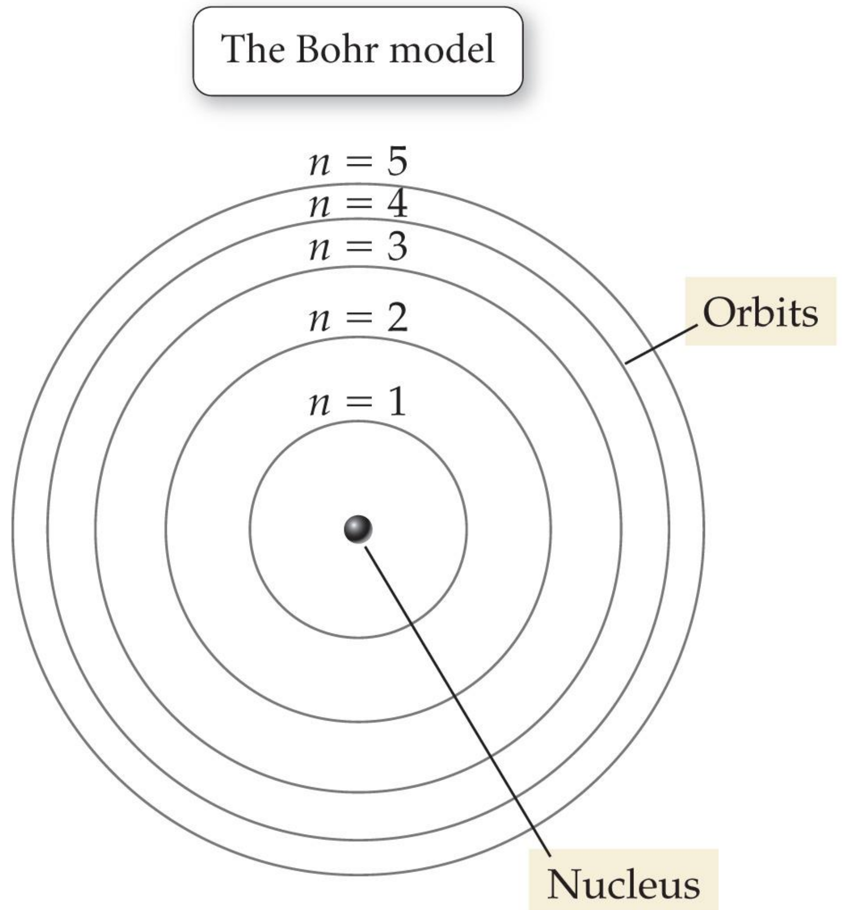
A white-light spectrum is continuous, with some radiation emitted at every wavelength.
The emission spectrum of an individual element includes only certain specific wavelengths.



The Spectrum of the Sun



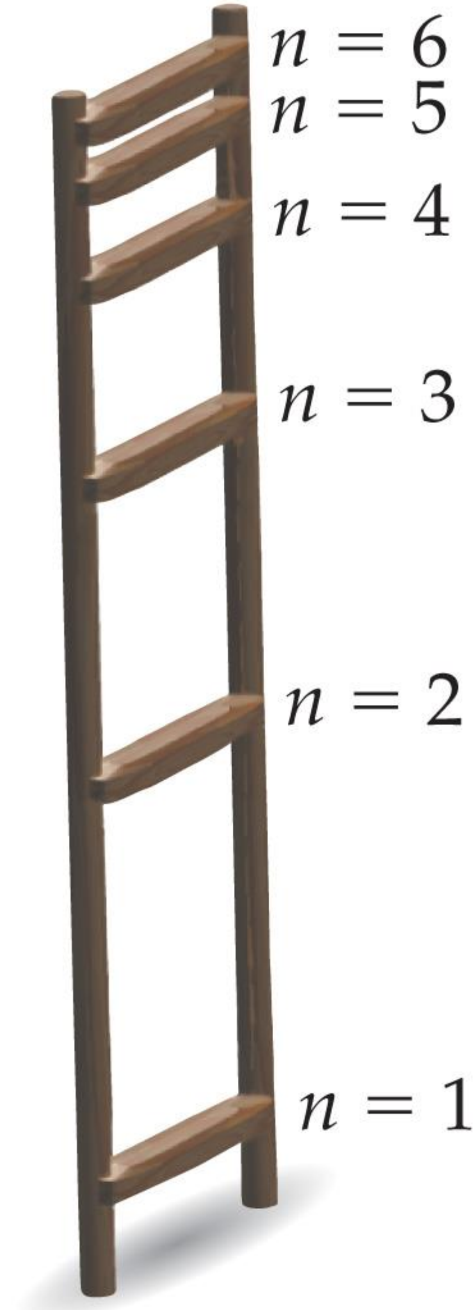
Niels Bohr developed a simple model to explain these results.



The *energy* of each Bohr orbit, specified by a **quantum number** *n = 1, 2, 3* is fixed, or **quantized**.

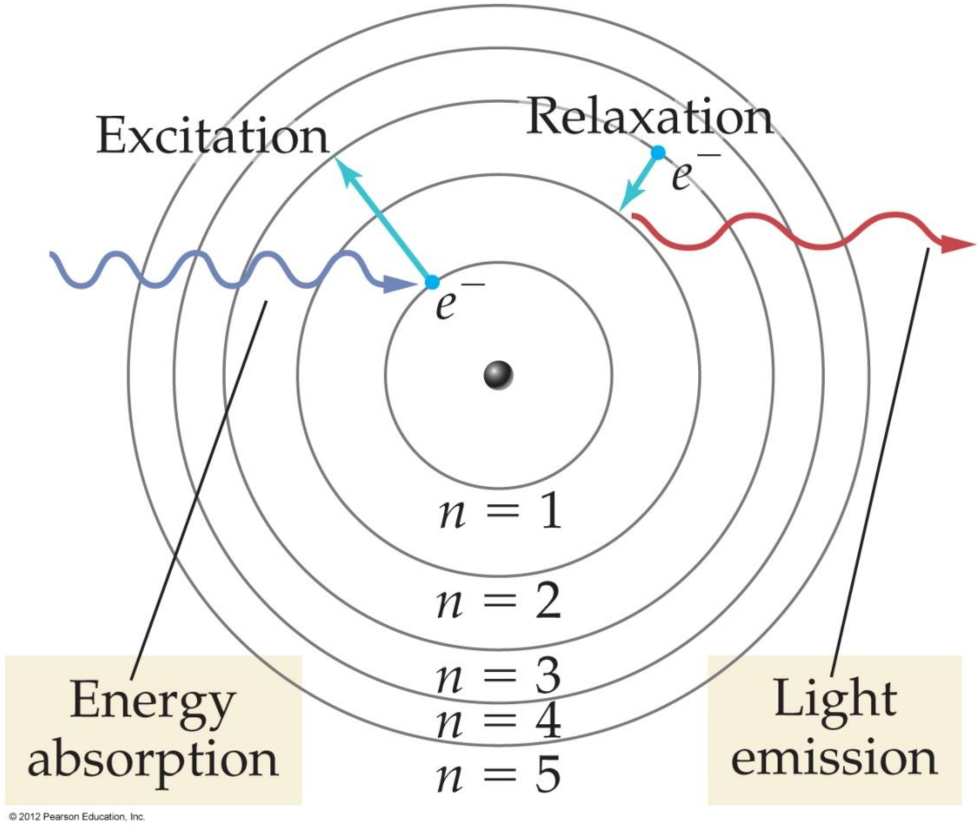
Bohr orbits are like steps of a ladder, each at a specific distance from the nucleus and each at a specific energy.

It is impossible for an electron to exist *between orbits* in the Bohr model.

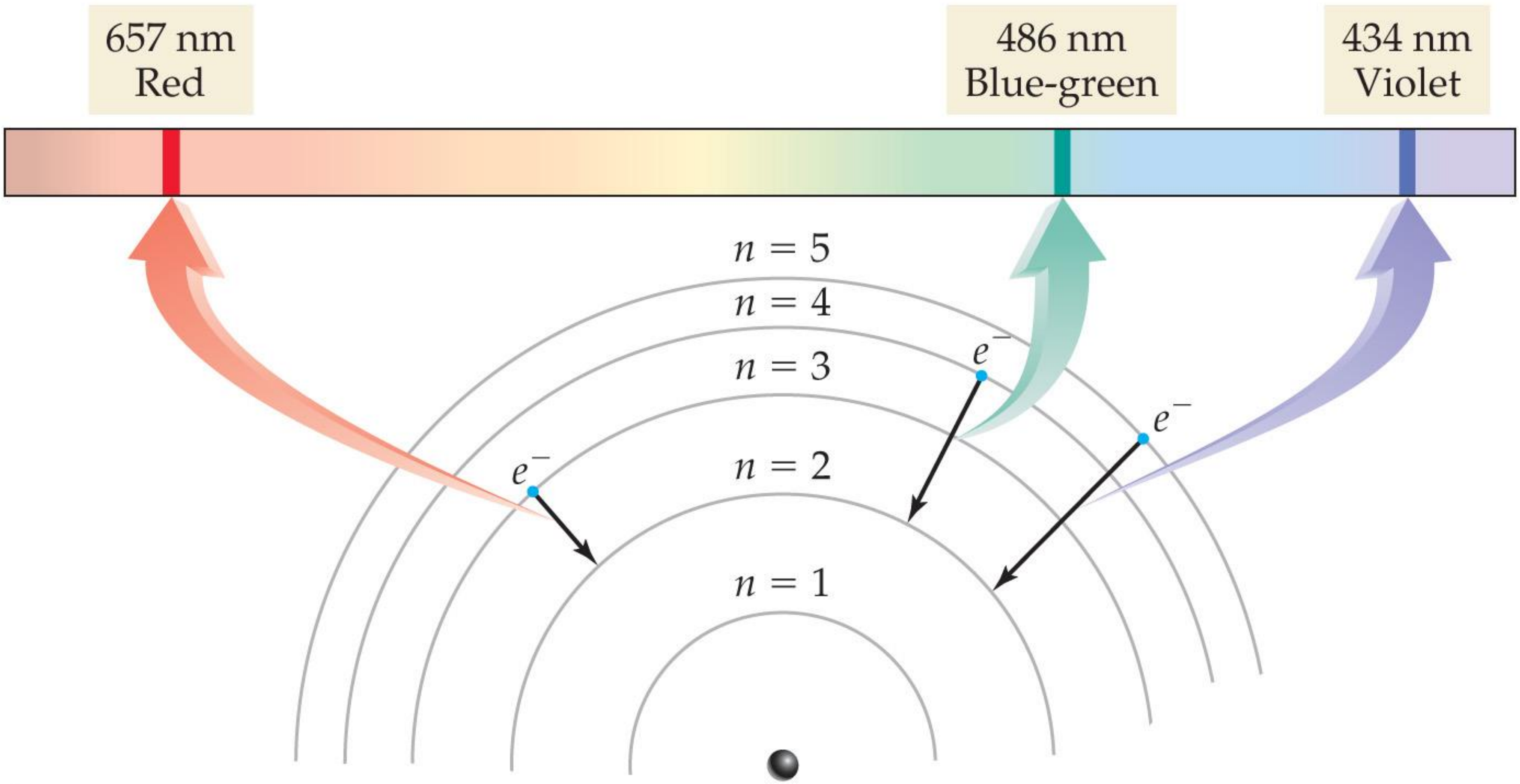


Excitation and Emission

- When a hydrogen atom absorbs energy, an electron is excited to a higher-energy orbit. The electron then relaxes back to a lower-energy orbit, emitting a photon of light.

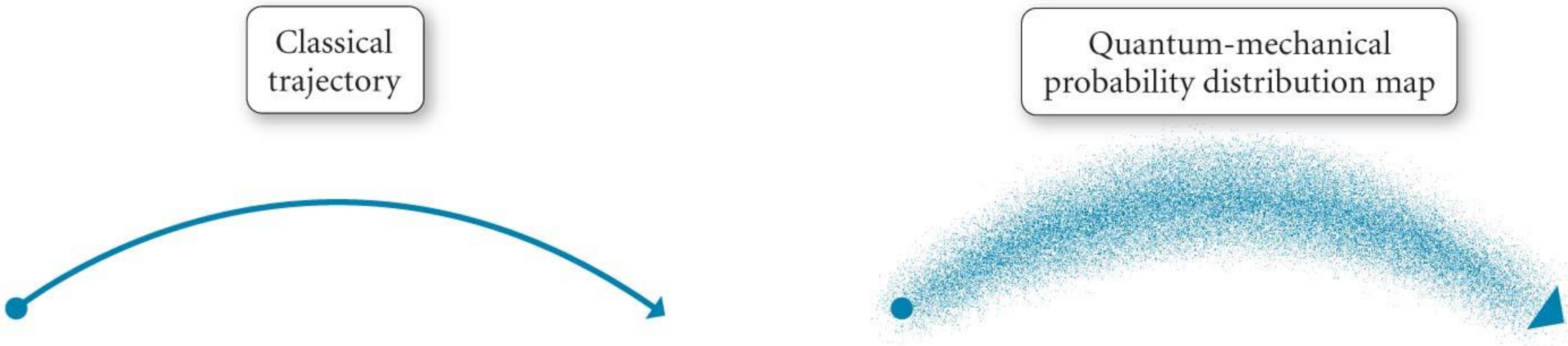


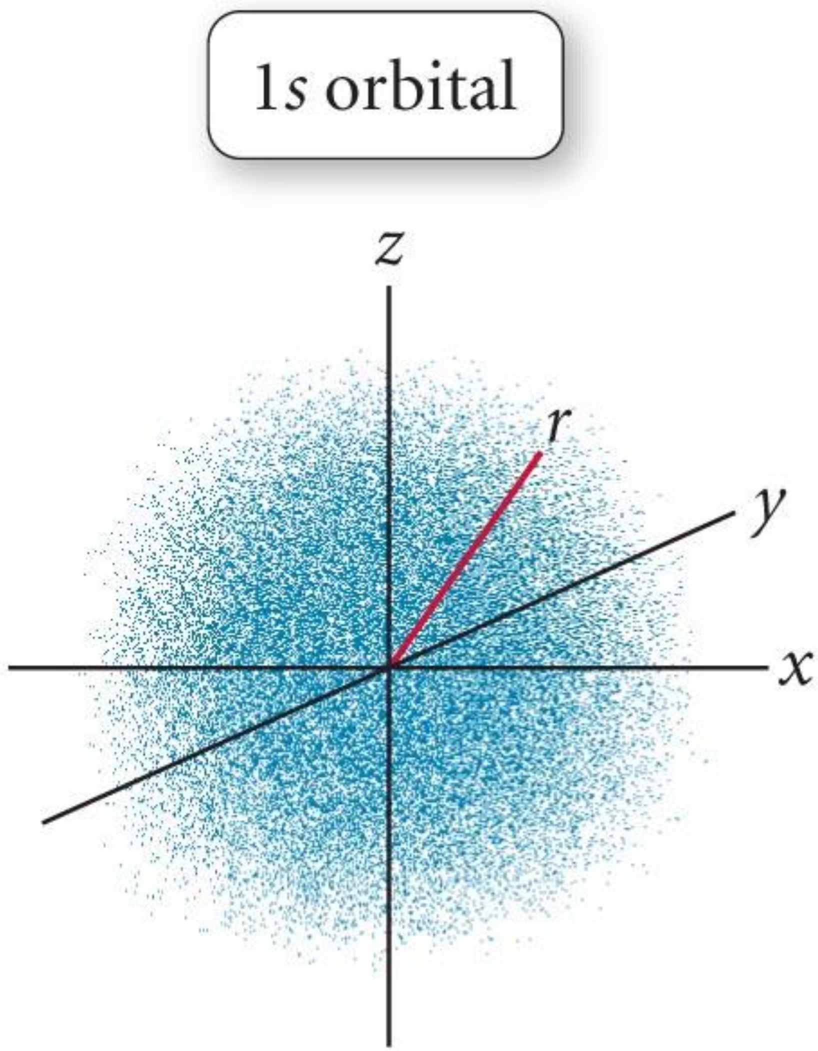
Hydrogen emission lines



9.4 The Bohr Model: Atoms with Orbits

- The great success of the Bohr model of the atom was that it predicted the lines of the hydrogen emission spectrum.
- However, it failed to predict the emission spectra of other elements that contained more than one electron.
- For this and other reasons, the Bohr model was replaced with a more sophisticated model called the quantum-mechanical or wave-mechanical model.

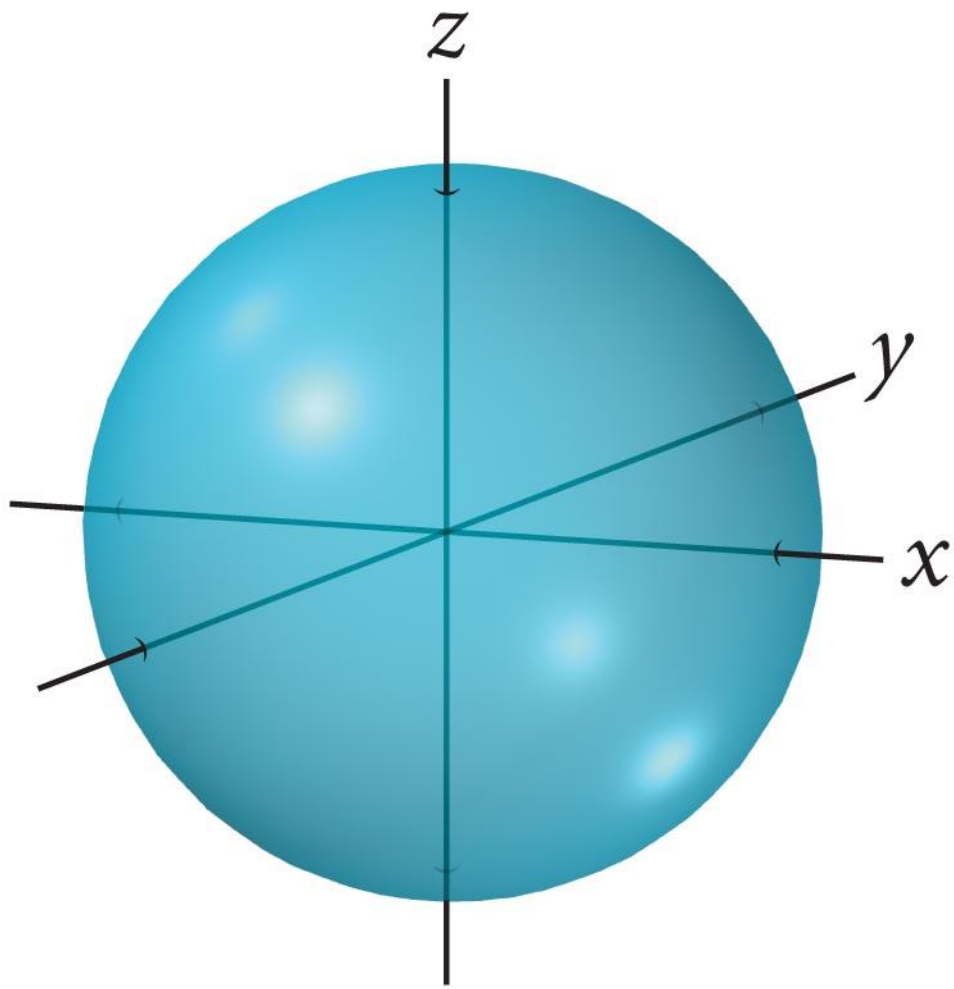

Classical trajectory
Quantum-mechanical probability distribution map



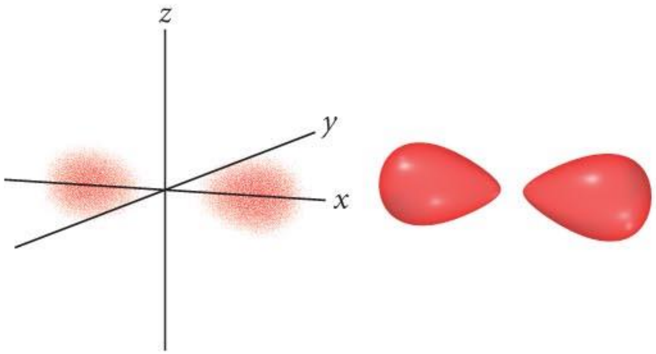
Density of dots proportional to probability density (ψ^2).

(a)

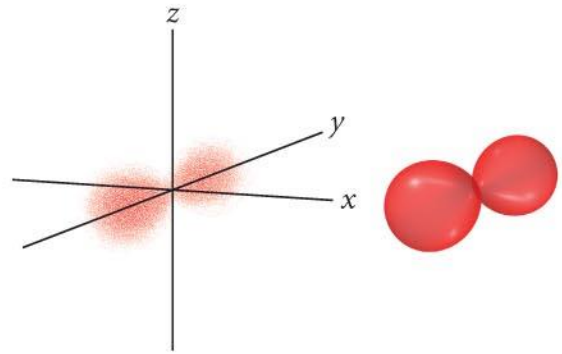
1s orbital surface



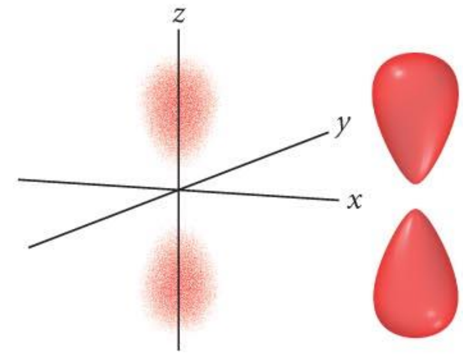
p_x orbital



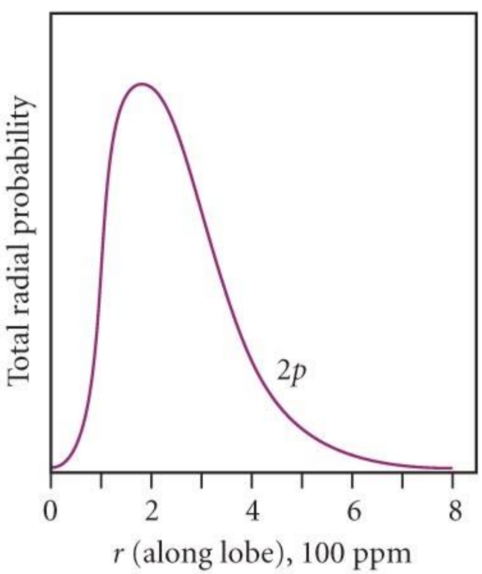
p_y orbital

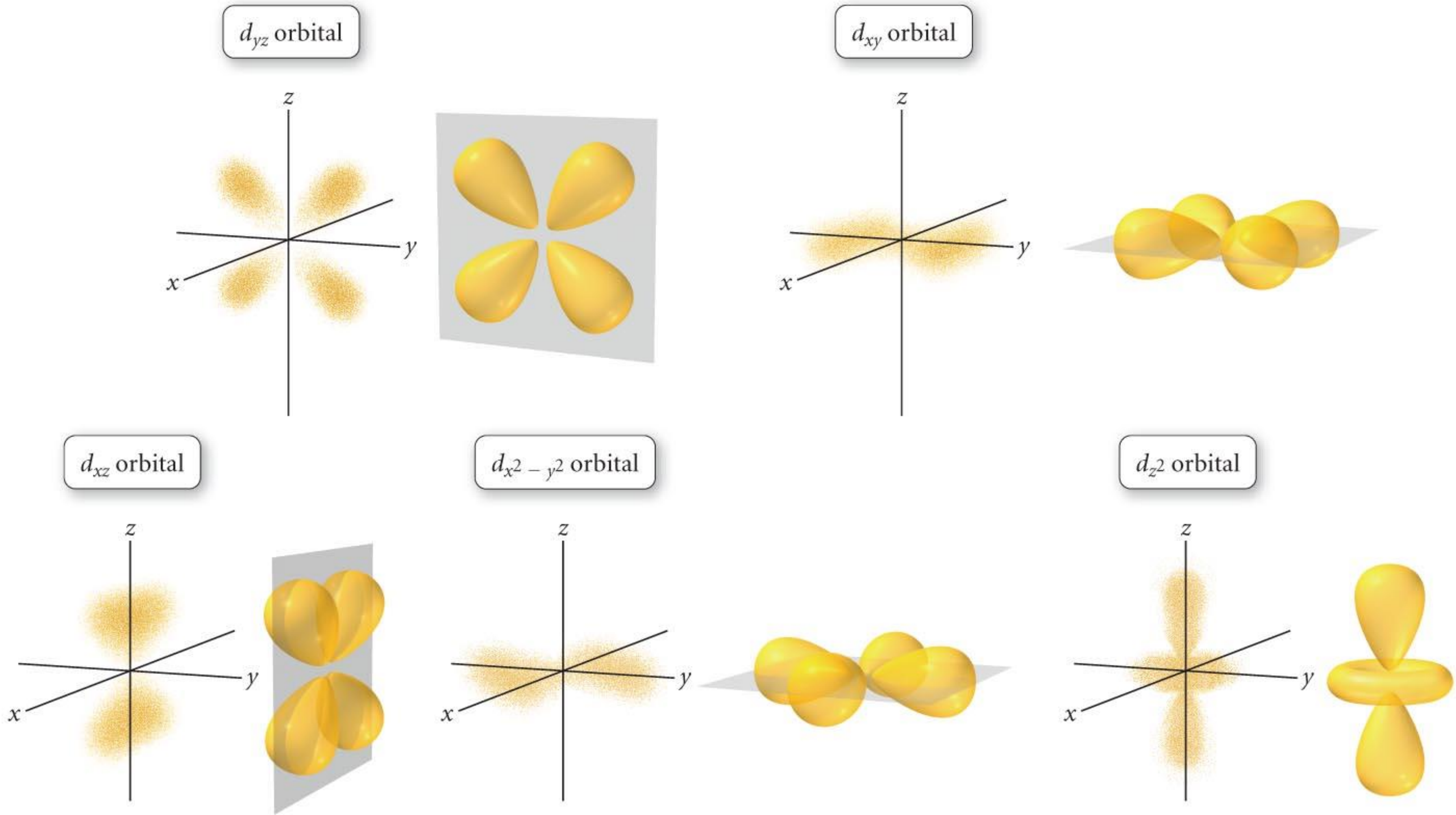


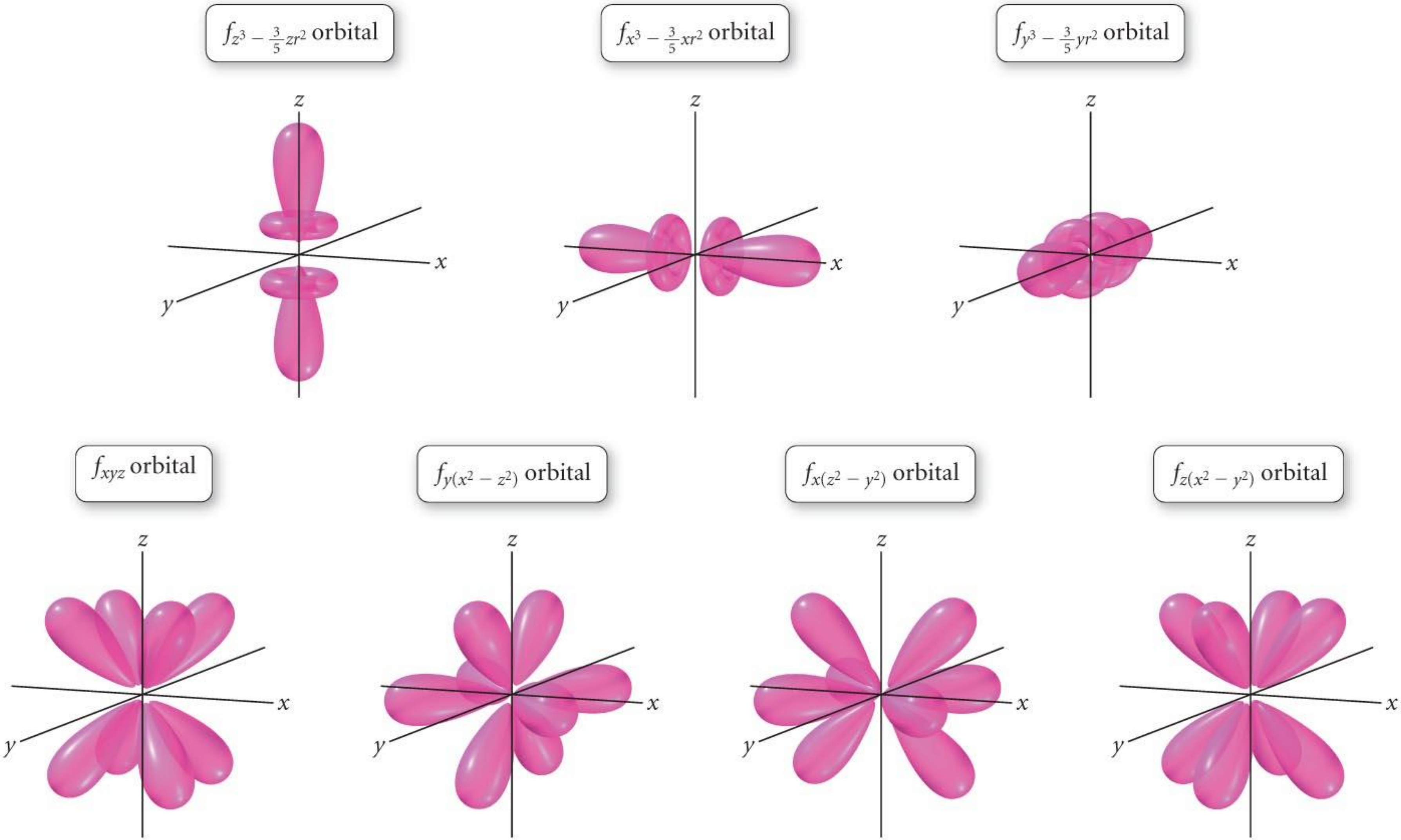
p_z orbital



Radial Distribution Function



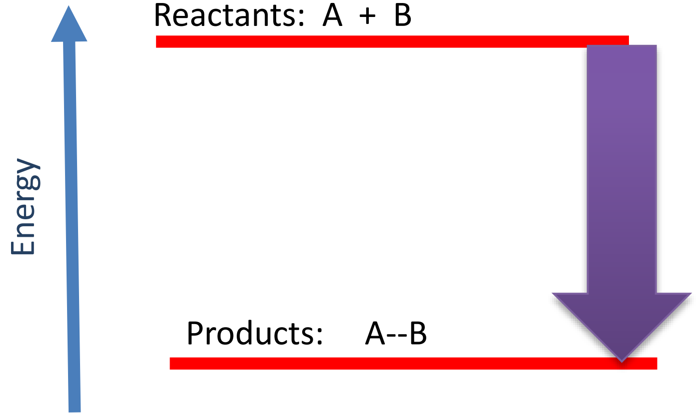

d_{yz} orbital
d_{xy} orbital
d_{xz} orbital
$d_{x^2-y^2}$ orbital
d_{z^2} orbital
z
y
x


$f_{z^3 - \frac{3}{5}zr^2}$ orbital
$f_{x^3 - \frac{3}{5}xr^2}$ orbital
$f_{y^3 - \frac{3}{5}yr^2}$ orbital
f_{xyz} orbital
$f_{y(x^2 - z^2)}$ orbital
$f_{x(z^2 - y^2)}$ orbital
$f_{z(x^2 - y^2)}$ orbital
z
x
y

Module 5
Part 2: Lewis Structures

Why do atoms bond?

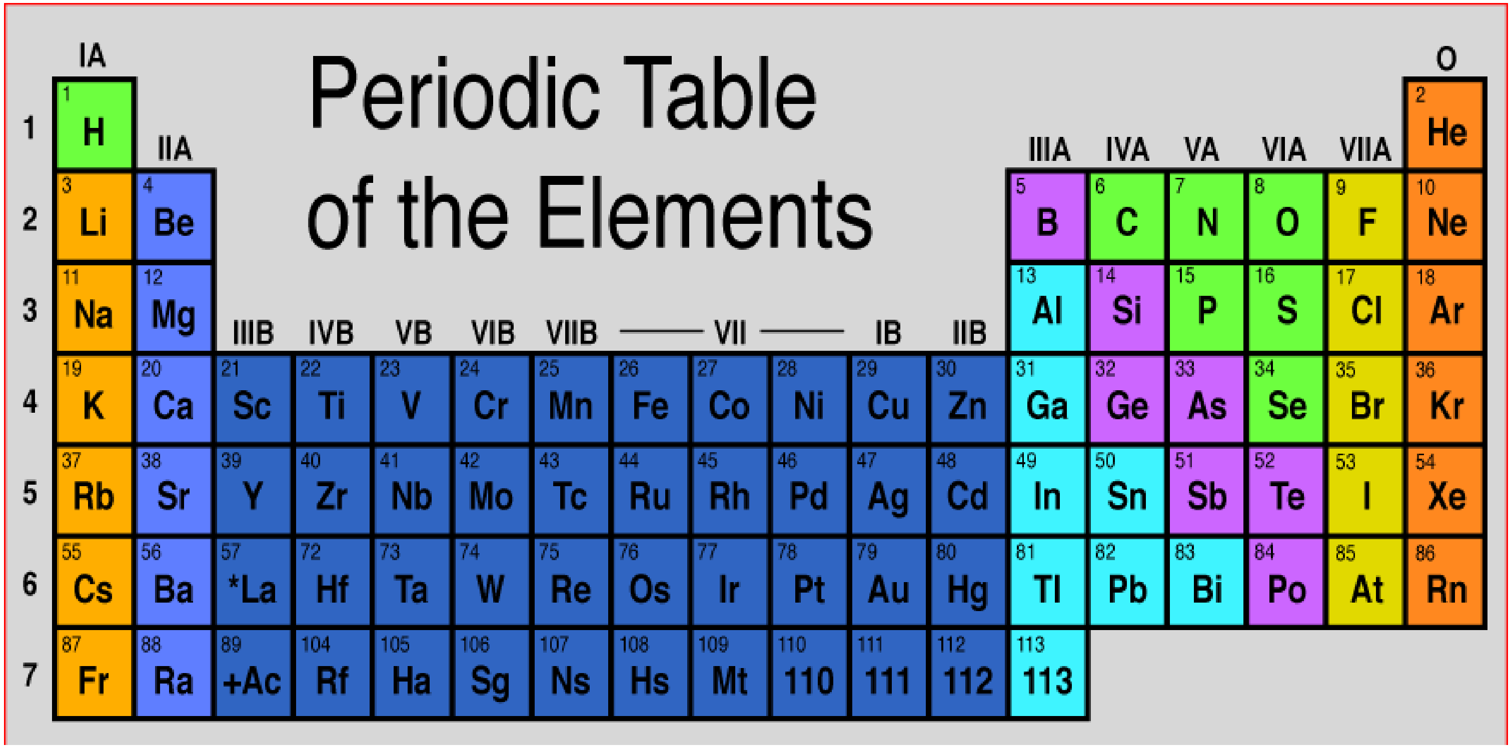
- Atoms bond primarily through transferring electrons (ionic bonds) or sharing electrons (covalent bonds).
  - A bond will form in order to put the electrons and nuclei in a lower potential energy state.



Valence Electrons

- Electrons are located outside of a nucleus in layers. Because it is lower in energy for the electron, the inner shells will fill first.
- Core electrons are those electrons that populate the inner shells of an atom.
- Valence Electrons are the electrons located in the outermost shell of the atom.
  – These are the electrons responsible for all the chemistry.

How many electrons are in the valence shell?



Octets and Duets

- Filled outer shell leads to stability
  - 2 for the innermost shell (duet)
  - 8 for outermost shell (octet)
- Atoms strive to have a filled valence shell
  - Lose or gain electrons to achieve this
  - Share electrons so that valence is filled
  - Note that atoms that are one away from achieving a filled valence are the most reactive while atoms that have a filled valence are non-reactive.

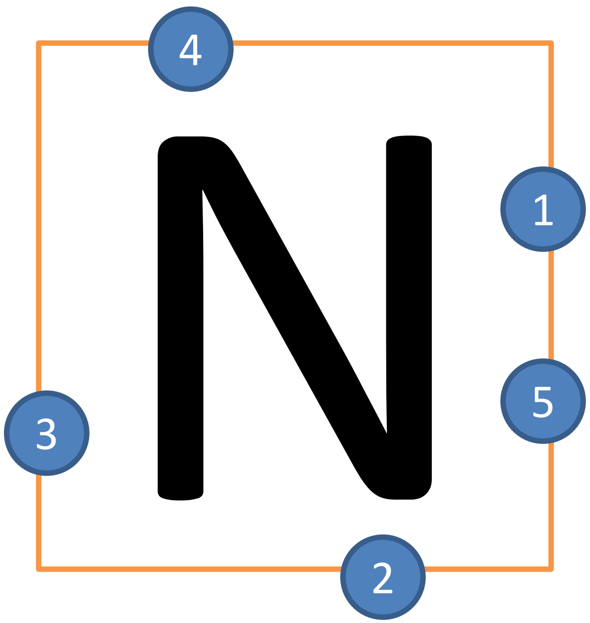
Lewis Symbols

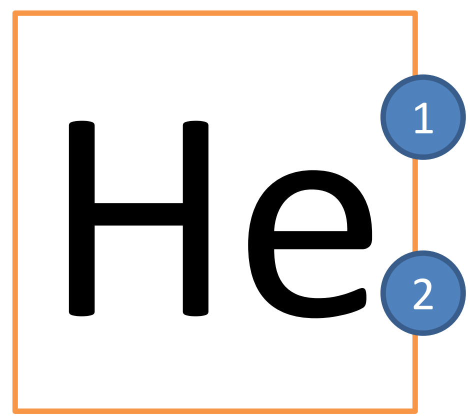
- Represent electrons as dots.
- Write the elemental symbol, then treat each side as a place to add electrons.
- Place electrons on separate sides first, pair when needed. When each side is filled, you have your octet:

N

He

Don’t actually draw the orange box! It is just there to emphasize the places to put electrons!





Use Lewis structures to illustrate an ionic bond:

1. Draw Lewis symbols for the elements involved in ionic compound.
2. Show electron transfer FROM METAL to NONMETAL.
3. Continue until each metal and nonmetal have complete octets, and note numbers of cations and anions show ratio of atoms in correct formula.

Use Lewis structures to illustrate covalent bonds:

- Each covalent bond is a “shared pair” of electrons. We represent them using a line.
- When more electrons are shared, we have a double (4 electrons) or triple (6 electrons) bond.

- Use Lewis theory to demonstrate why diatomic elements exist:
  - Examples:
    - Cl_2
    - O_2
    - N_2

Lewis structures for multi-atom molecules:

- CH_4
- H_2O

Steps for Determining a structure for a multi-atom molecule:

1. Count total valence in entire molecule.
2. Draw skeletal structure:
    1. Place “most-metallic” atom in center.
    2. Place other atoms as “terminal”.
    3. Connect using single bonds.
3. Distribute the rest of the valence electrons to complete octets. Start on terminal atoms, add electrons until you run out.
4. Check octets!
    1. Use double/triple bonds where needed to satisfy octets.
5. Add brackets and charges if polyatomic ion.