

## Lecture Outlines

### Chapter 31

*Physics, 3<sup>rd</sup> Edition*

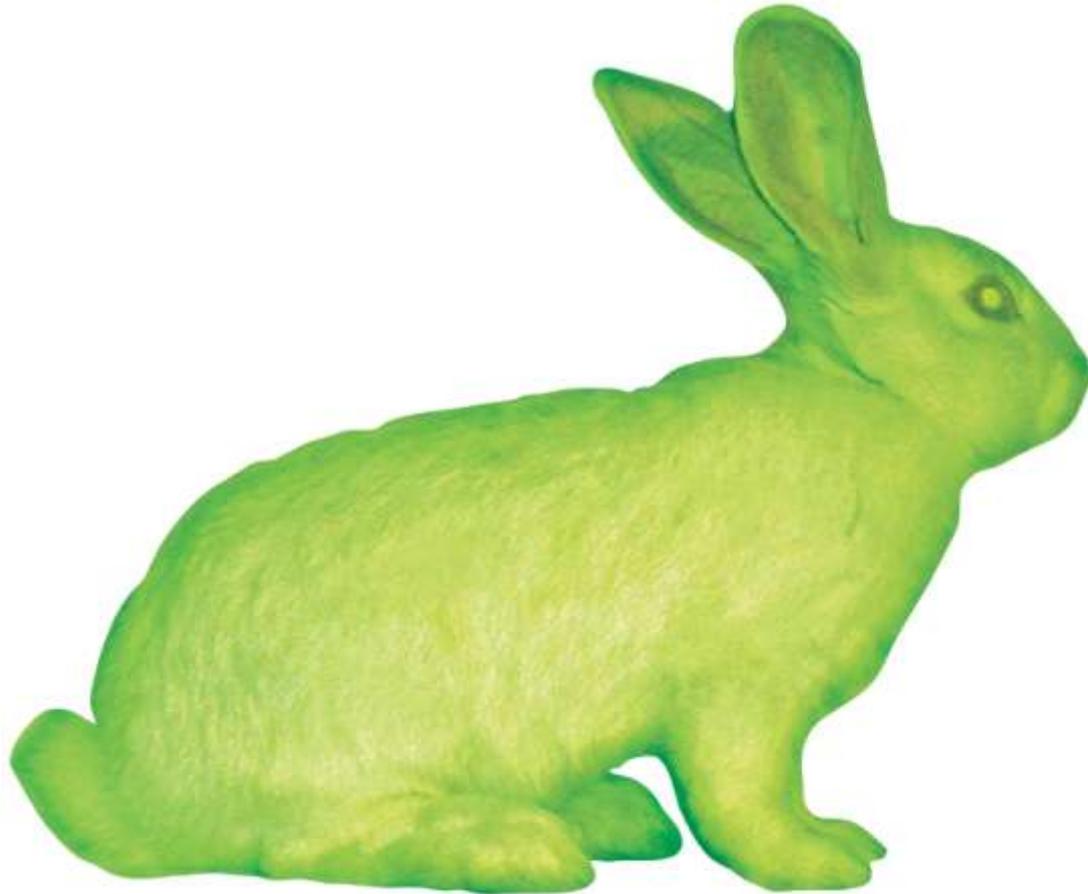
James S. Walker

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# Chapter 31

## Atomic Physics

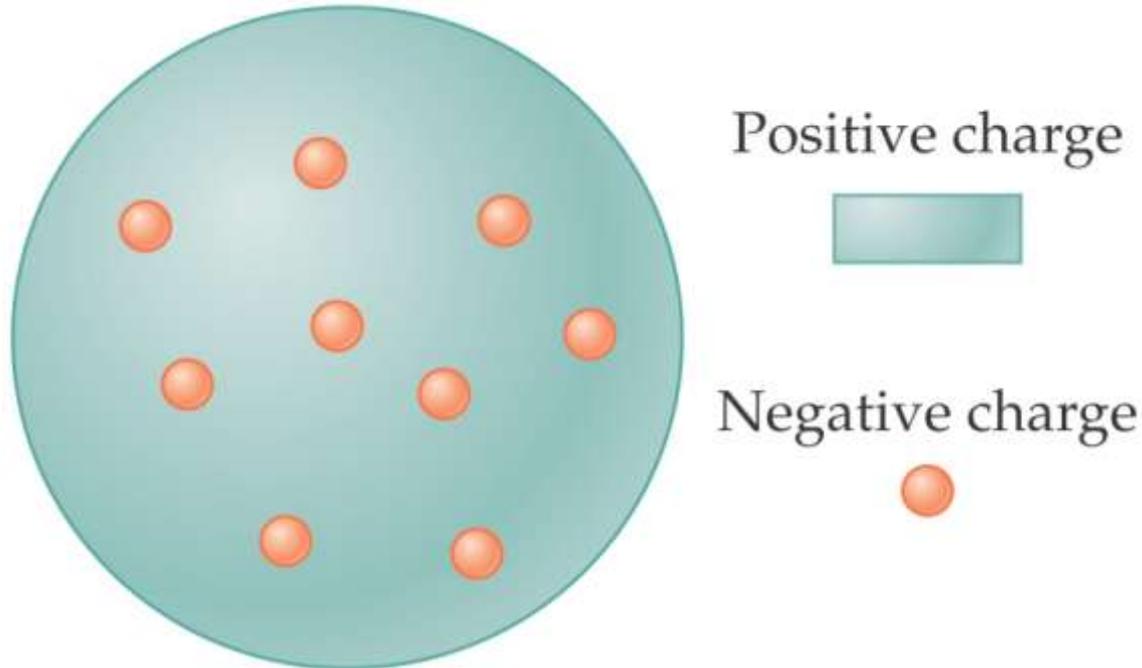


# Units of Chapter 31

- **Early Models of the Atom**
- **The Spectrum of Atomic Hydrogen**
- **Bohr's Model of the Hydrogen Atom**
- **de Broglie Waves and the Bohr Model**
- **The Quantum Mechanical Hydrogen Atom**
- **Multielectron atoms and the Periodic Table**
- **Atomic Radiation**

## 31-1 Early Models of the Atom

The electron was discovered in 1897, and was observed to be much smaller than the atom. It was known that atoms are electrically neutral; the first modern model of the atom was therefore the “plum pudding” model – tiny electrons embedded in a mass of positive charge.

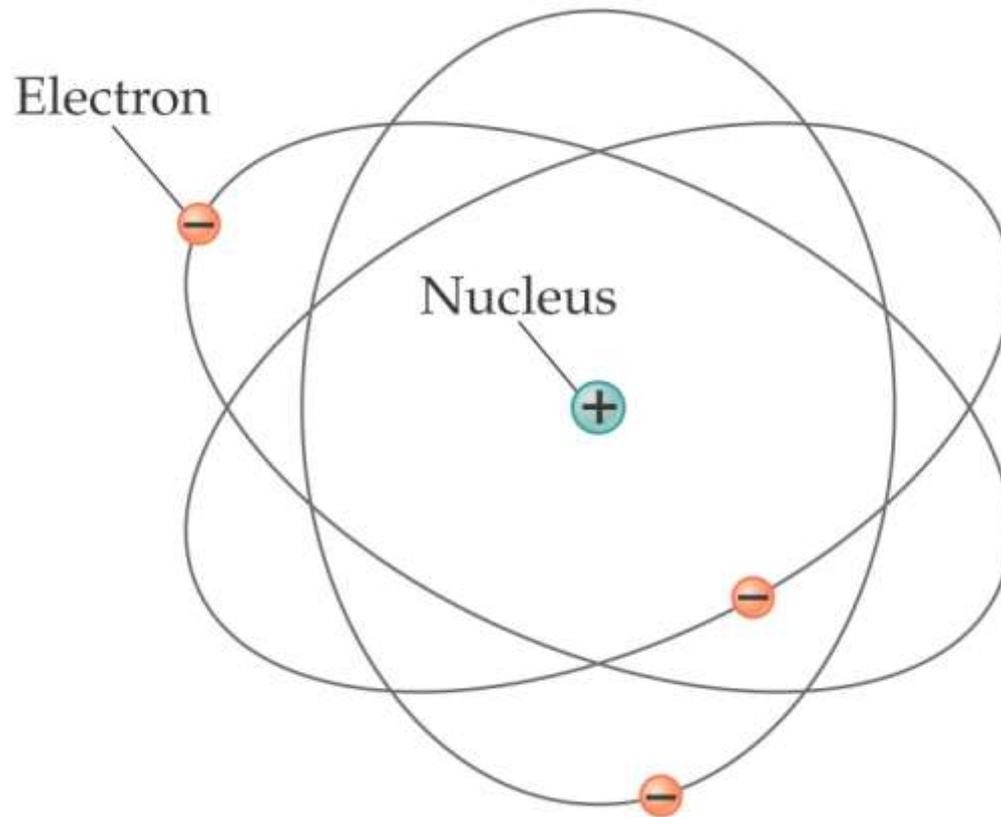


# 31-1 Early Models of the Atom

**Experiments were done to confirm this model, looking at how alpha particles (helium nuclei) scattered from a gold foil. They found many more large-angle scatters than expected – something that could only happen if the positive charge were concentrated in a tiny volume, rather than spread throughout the atom.**

# 31-1 Early Models of the Atom

This led to the “solar system” model of the atom – electrons orbiting a small, positively charged nucleus.



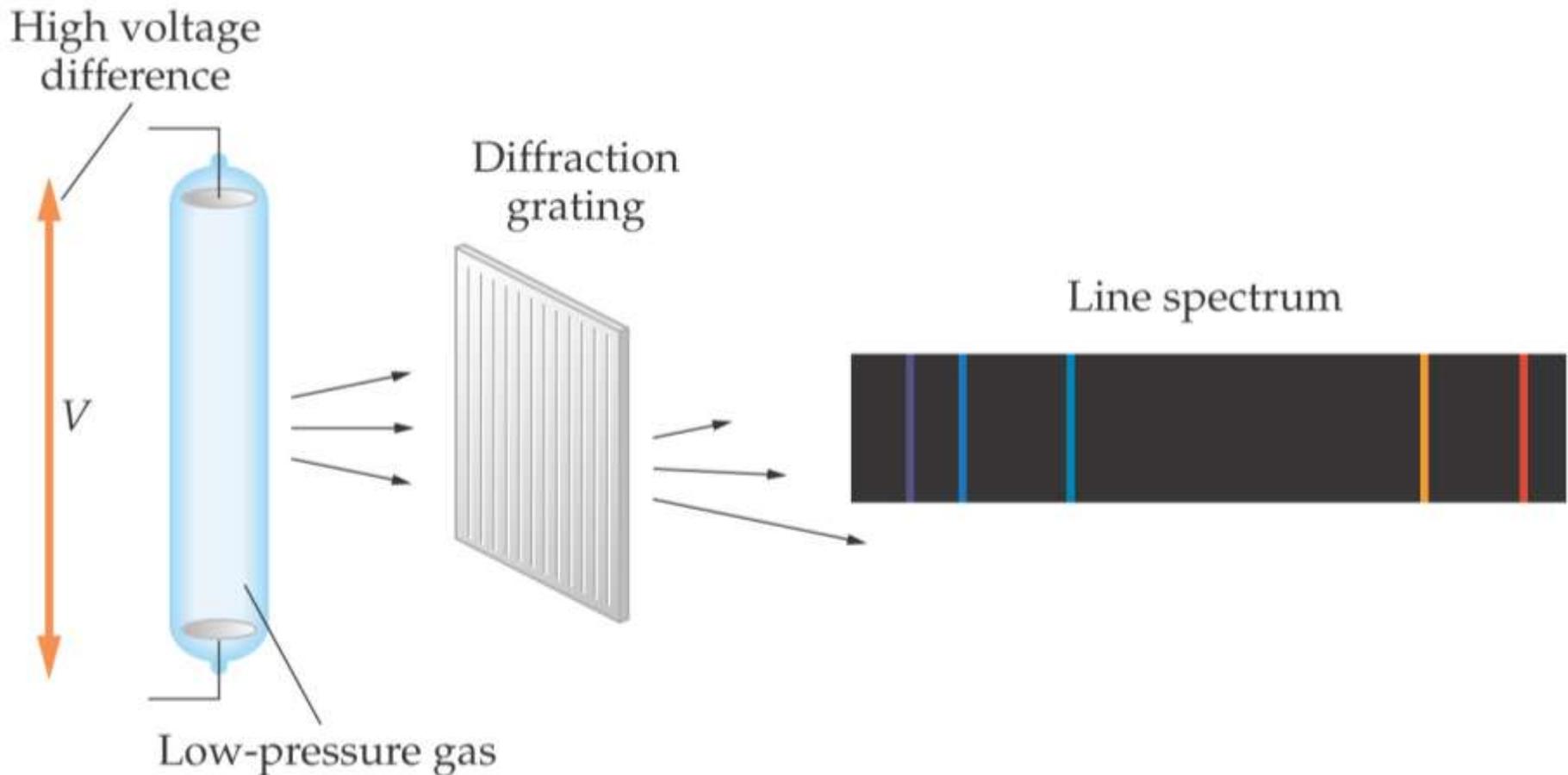
# 31-1 Early Models of the Atom

**This is still the mental picture many people have of the atom; what is wrong with it?**

- Orbiting electrons are accelerating and should radiate.**
- The frequency of radiation emitted by a continuously radiating electron would have a continuous spectrum, rather than the individual frequencies that are actually observed.**

# 31-2 The Spectrum of Atomic Hydrogen

If we look at the light emitted by a low-pressure gas when subjected to a large electric field, we find a series of individual lines, called a line spectrum. This is the line spectrum of hydrogen.



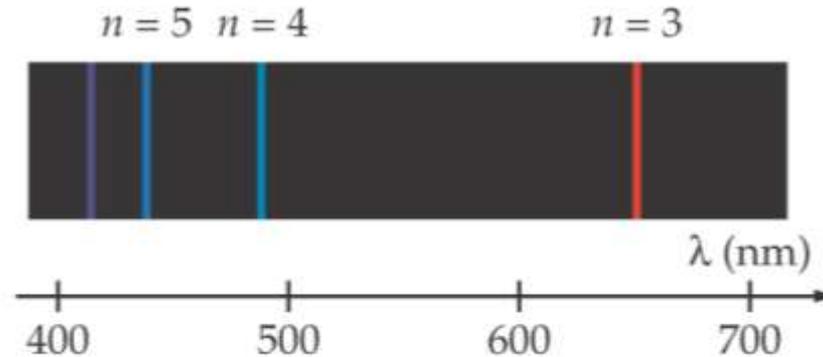
# 31-2 The Spectrum of Atomic Hydrogen

Hydrogen is responsible for the red color of this emission nebula.

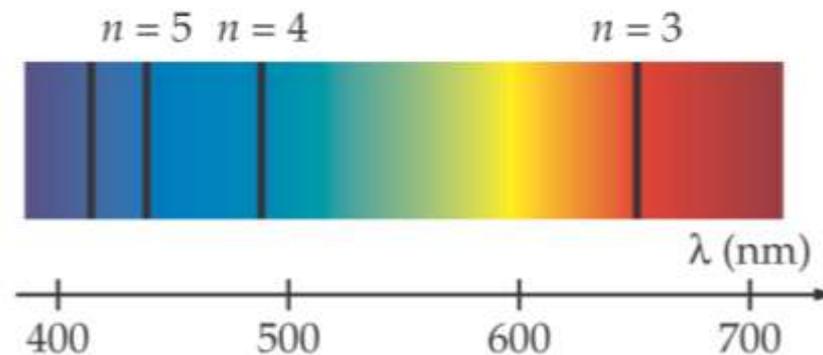


# 31-2 The Spectrum of Atomic Hydrogen

Each atom has its own particular pattern of emission lines. If white light passes through, it absorbs at the same frequencies seen in the emission spectrum.



(a)



(b)

## 31-2 The Spectrum of Atomic Hydrogen

The wavelengths in the visible spectrum of hydrogen are given by:

$$\frac{1}{\lambda} = R \left( \frac{1}{2^2} - \frac{1}{n^2} \right) \quad n = 3, 4, 5, \dots$$

**$R$  is the Rydberg constant:**  $R = 1.097 \times 10^7 \text{ m}^{-1}$

**This is called the Balmer series of spectral lines.**

# 31-2 The Spectrum of Atomic Hydrogen

To find all the series of hydrogen, the following general formula can be used:

$$\frac{1}{\lambda} = R \left( \frac{1}{n'^2} - \frac{1}{n^2} \right) \quad n' = 1, 2, 3, \dots$$
$$n = n' + 1, n' + 2, n' + 3, \dots$$

**TABLE 31-1**

**Common Spectral Series of Hydrogen**

$n'$	Series name
1	Lyman
2	Balmer
3	Paschen
4	Brackett
5	Pfund

# 31-3 Bohr's Model of the Hydrogen Atom

## Assumptions of the Bohr model:

- The electron in a hydrogen atom moves in a circular orbit around the nucleus.
- Only certain orbits are allowed, where the angular momentum in the  $n$ th allowed orbit is

$$L_n = nh/2\pi$$

- Electrons in allowed orbits do not radiate. Radiation is emitted when an electron changes from one orbit to another, with frequency given by:  $|\Delta E| = hf$ .

## 31-3 Bohr's Model of the Hydrogen Atom

In order for an electron to move in a circle of radius  $r$  at speed  $v$ , the electrostatic force must provide the required centripetal force.

Hence:

$$mv^2 = k \frac{e^2}{r}$$

Adding the angular momentum requirement gives:

$$v_n = \frac{nh}{2\pi mr_n} \quad n = 1, 2, 3, \dots$$

# 31-3 Bohr's Model of the Hydrogen Atom

Solving these equations for the allowed radii gives:

$$r_n = \left( \frac{h^2}{4\pi^2 m k e^2} \right) n^2 \quad n = 1, 2, 3, \dots$$

The quantity in parentheses is the radius of the smallest orbit:

$$r_1 = 5.29 \times 10^{-11} \text{ m}$$

## 31-3 Bohr's Model of the Hydrogen Atom

Bohr's model gives correct predictions for any single-electron atom, when the different Coulomb force is taken into account. Here,  $Z$  is the number of protons in the nucleus:

$$r_n = \left( \frac{h^2}{4\pi^2 m k Z e^2} \right) n^2 \quad n = 1, 2, 3, \dots$$

## 31-3 Bohr's Model of the Hydrogen Atom

The total mechanical energy of an electron in a Bohr orbit is the sum of its kinetic and potential energies. After some algebraic manipulation, and substituting known values of constants, we find:

$$E_n = -(13.6 \text{ eV}) \frac{Z^2}{n^2} \quad n = 1, 2, 3, \dots$$

The lowest energy is called the ground state. Most atoms at room temperature are in the ground state.

## 31-3 Bohr's Model of the Hydrogen Atom

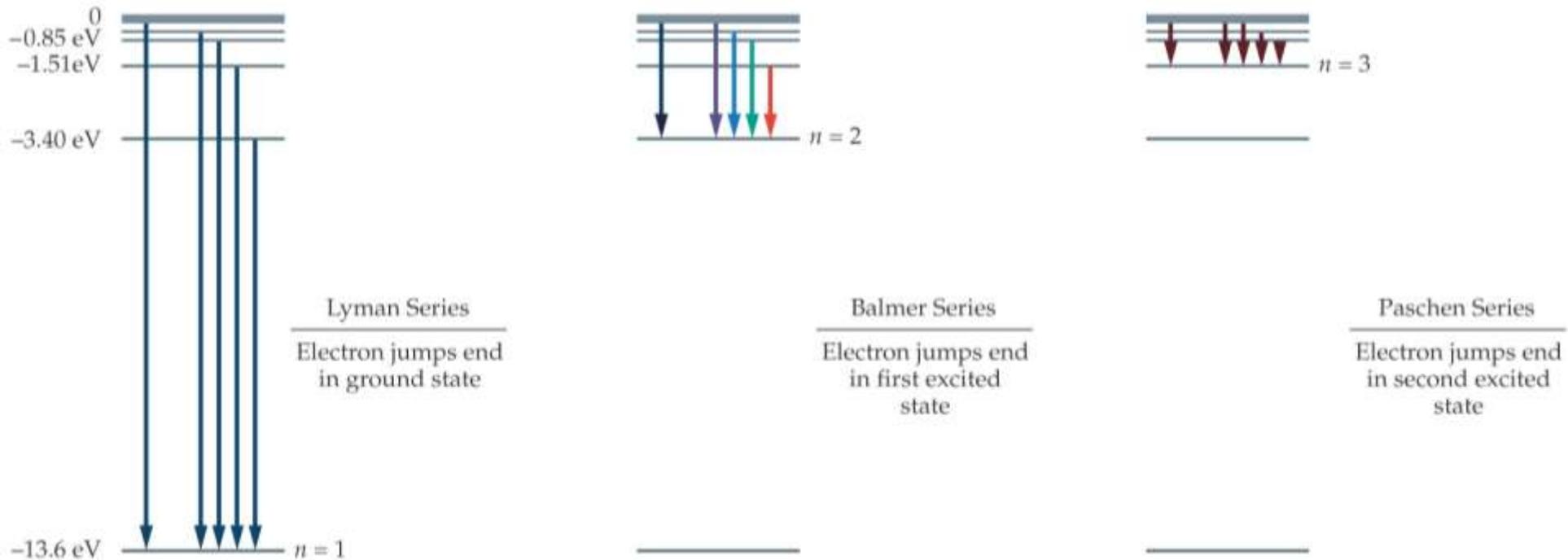
Now, if we assume that the radiation emitted from excited atoms corresponds to the energy difference between two levels, we find:

$$\frac{1}{\lambda} = \left( \frac{2\pi^2 m k^2 e^4}{h^3 c} \right) \left( \frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$$

This is the same as the observation made earlier; when numerical values are substituted for the constants in the parentheses, they yield Rydberg's constant.

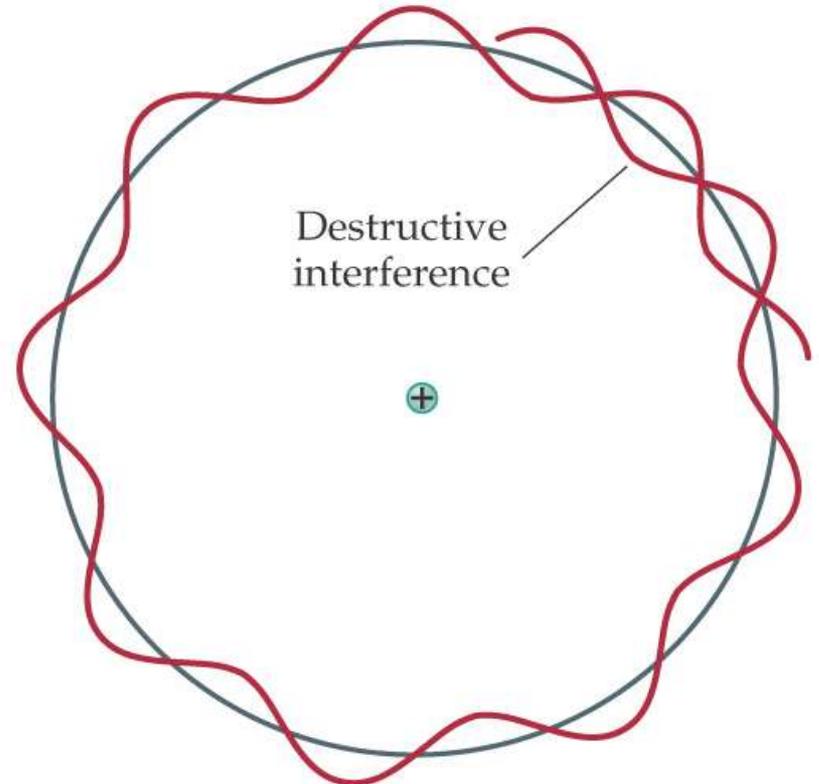
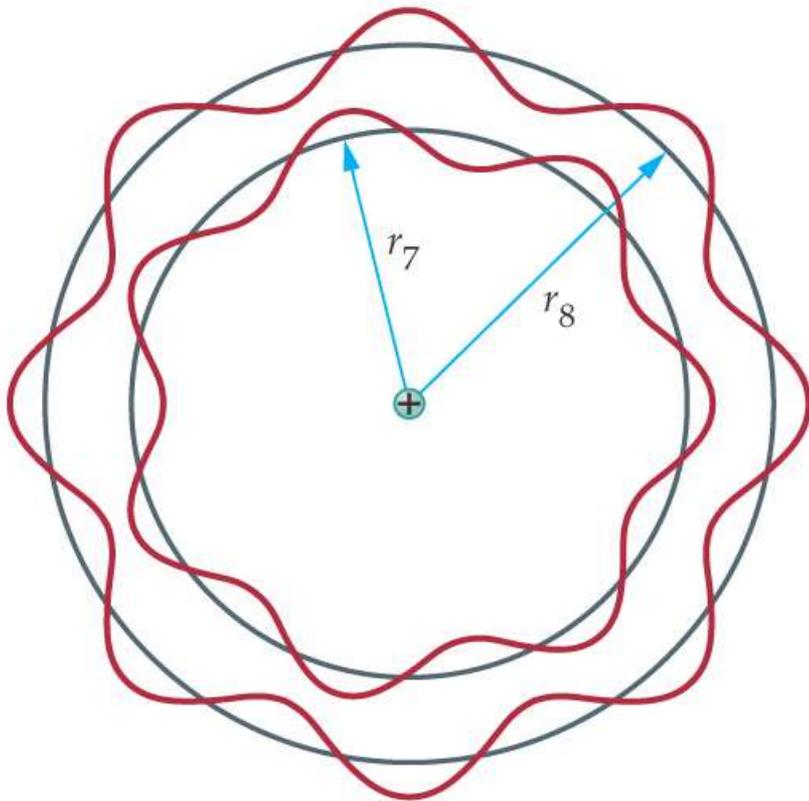
# 31-3 Bohr's Model of the Hydrogen Atom

## The Lyman, Balmer, and Paschen series:



# 31-4 de Broglie Waves and the Bohr Model

De Broglie proposed that the allowed orbits were those which comprised an integral number of wavelengths of the electron – a kind of standing wave.



## 31-4 de Broglie Waves and the Bohr Model

Therefore, the momentum of an electron in the  $n$ th orbit is:

$$p = mv = \frac{h}{\lambda} = \frac{h}{(2\pi r/n)} = \frac{nh}{2\pi r} \quad n = 1, 2, 3, \dots$$

And the angular momentum:

$$L = rmv = \frac{nh}{2\pi} \quad n = 1, 2, 3, \dots$$

**Exactly as predicted.**

# 31-5 The Quantum Mechanical Hydrogen Atom

The quantum mechanical description of the atom requires four numbers:

1. The principal quantum number,  $n$ , gives the energy of the electron.

$$E = (-13.6 \text{ eV})/n^2$$

2. The orbital angular momentum quantum number  $l$ :

$$l = 0, 1, 2, \dots, (n - 1)$$

$$L = \sqrt{l(l + 1)} \frac{h}{2\pi}$$

# 31-5 The Quantum Mechanical Hydrogen Atom

**3. The magnetic quantum number, which comes into play if there is a magnetic field:**

$$m_\ell = -\ell, -\ell + 1, -\ell + 2, \dots, -1, 0, 1, \dots, \ell - 2, \ell - 1, \ell$$

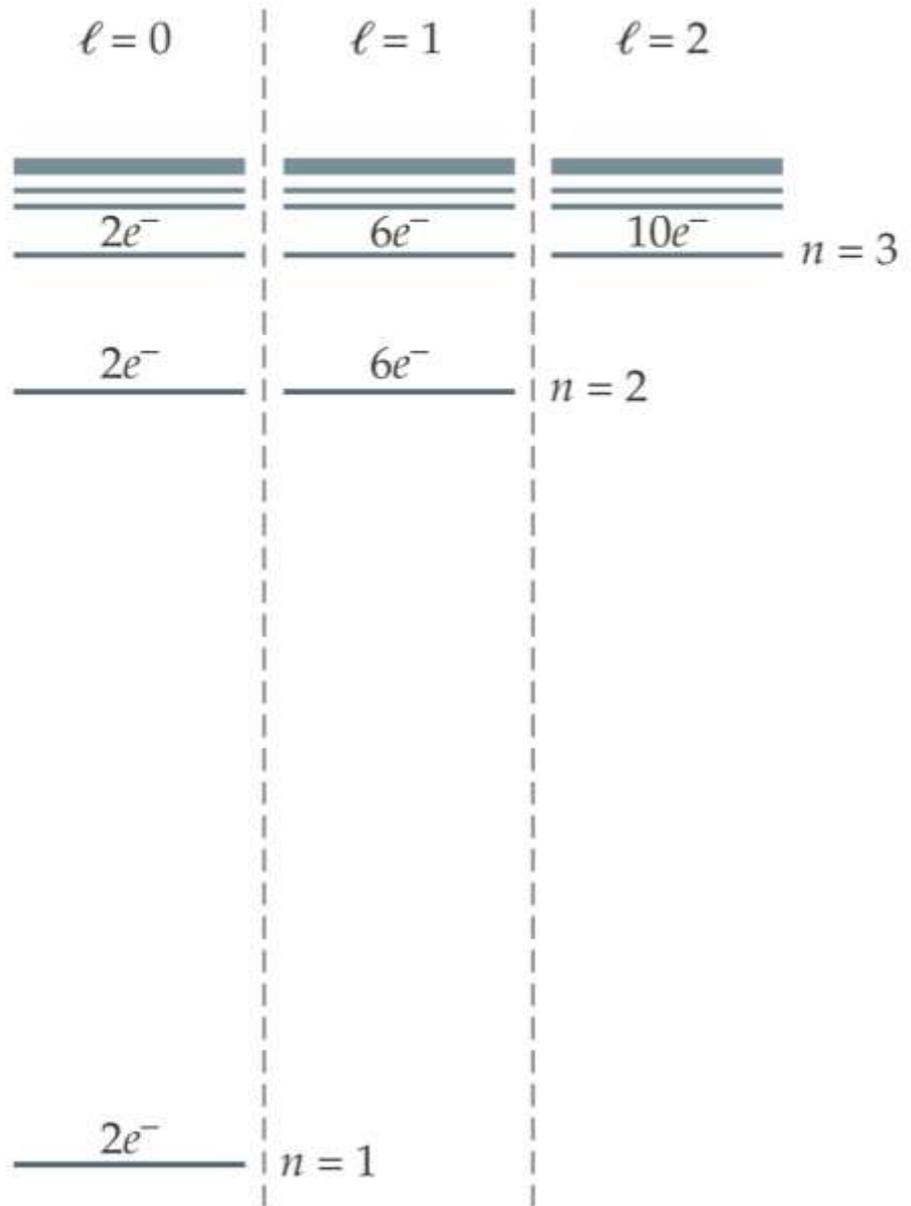
$$L_z = m_\ell \frac{h}{2\pi}$$

**4. The electron spin number, which is an intrinsic property of the electron:**

$$m_s = -\frac{1}{2}, \frac{1}{2}$$

# 31-5 The Quantum Mechanical Hydrogen Atom

In the absence of a magnetic field, the energy-level structure is the same as given by the Bohr model.



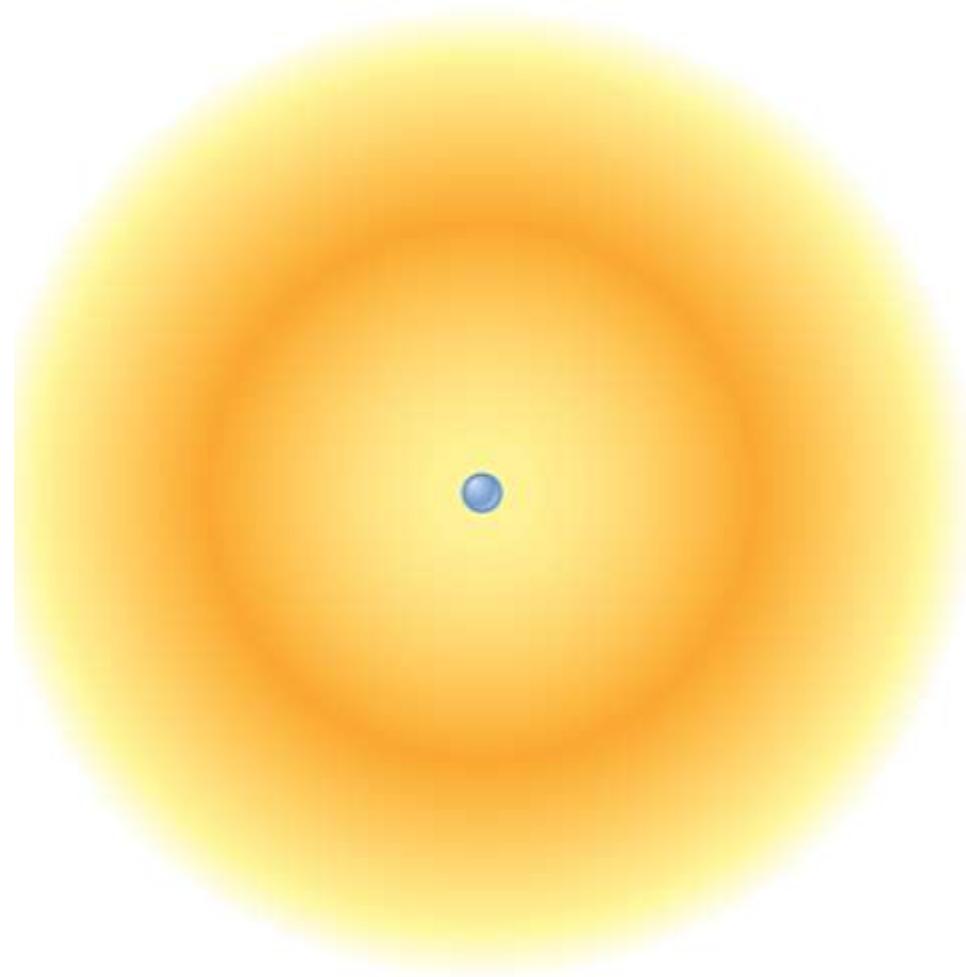
# 31-5 The Quantum Mechanical Hydrogen Atom

A state is defined as a particular assignment of the four quantum numbers:

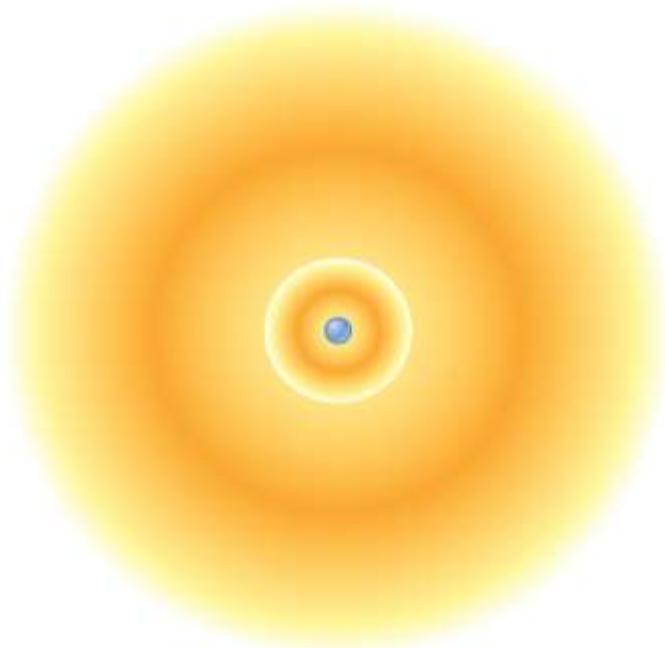
TABLE 31-2 States of Hydrogen for $n = 1$ and $n = 2$			
$n = 1, \ell = 0$			Two states
$n = 1$	$\ell = 0$	$m_\ell = 0$	$m_s = \frac{1}{2}$
$n = 1$	$\ell = 0$	$m_\ell = 0$	$m_s = -\frac{1}{2}$
$n = 2, \ell = 0$			Two states
$n = 2$	$\ell = 0$	$m_\ell = 0$	$m_s = \frac{1}{2}$
$n = 2$	$\ell = 0$	$m_\ell = 0$	$m_s = -\frac{1}{2}$
$n = 2, \ell = 1$			Six states
$n = 2$	$\ell = 1$	$m_\ell = 1$	$m_s = \frac{1}{2}$
$n = 2$	$\ell = 1$	$m_\ell = 1$	$m_s = -\frac{1}{2}$
$n = 2$	$\ell = 1$	$m_\ell = 0$	$m_s = \frac{1}{2}$
$n = 2$	$\ell = 1$	$m_\ell = 0$	$m_s = -\frac{1}{2}$
$n = 2$	$\ell = 1$	$m_\ell = -1$	$m_s = \frac{1}{2}$
$n = 2$	$\ell = 1$	$m_\ell = -1$	$m_s = -\frac{1}{2}$

# 31-5 The Quantum Mechanical Hydrogen Atom

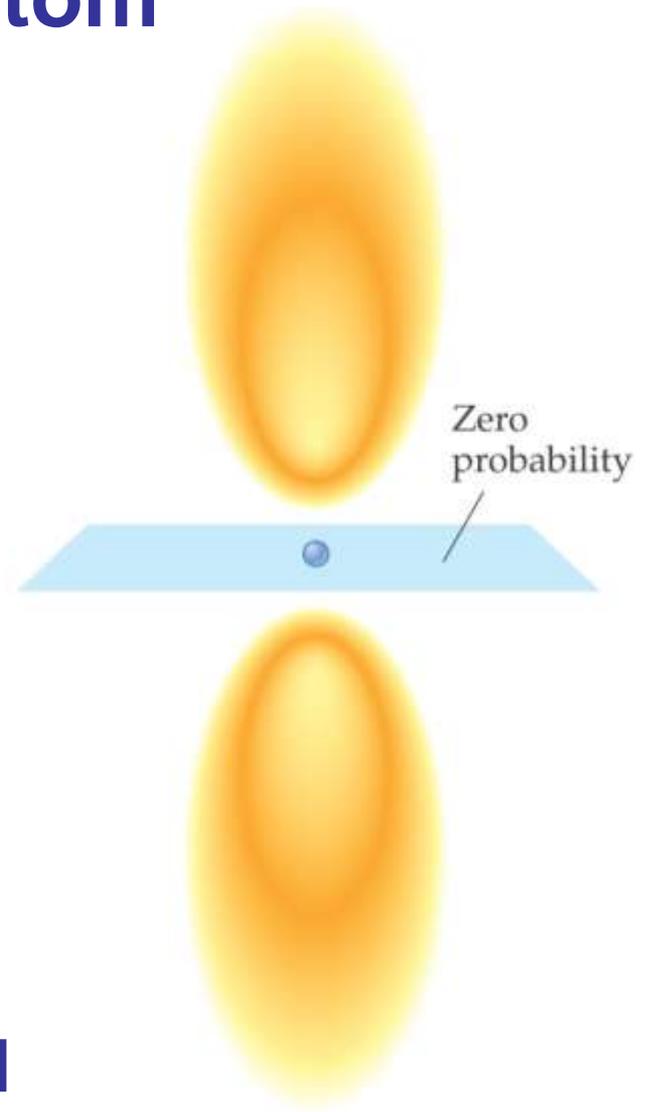
In the Schrödinger theory of quantum mechanics, the electron is described by a wave function, which gives the probability of finding the electron at a given position. This is the ground state of hydrogen.



# 31-5 The Quantum Mechanical Hydrogen Atom



(a)

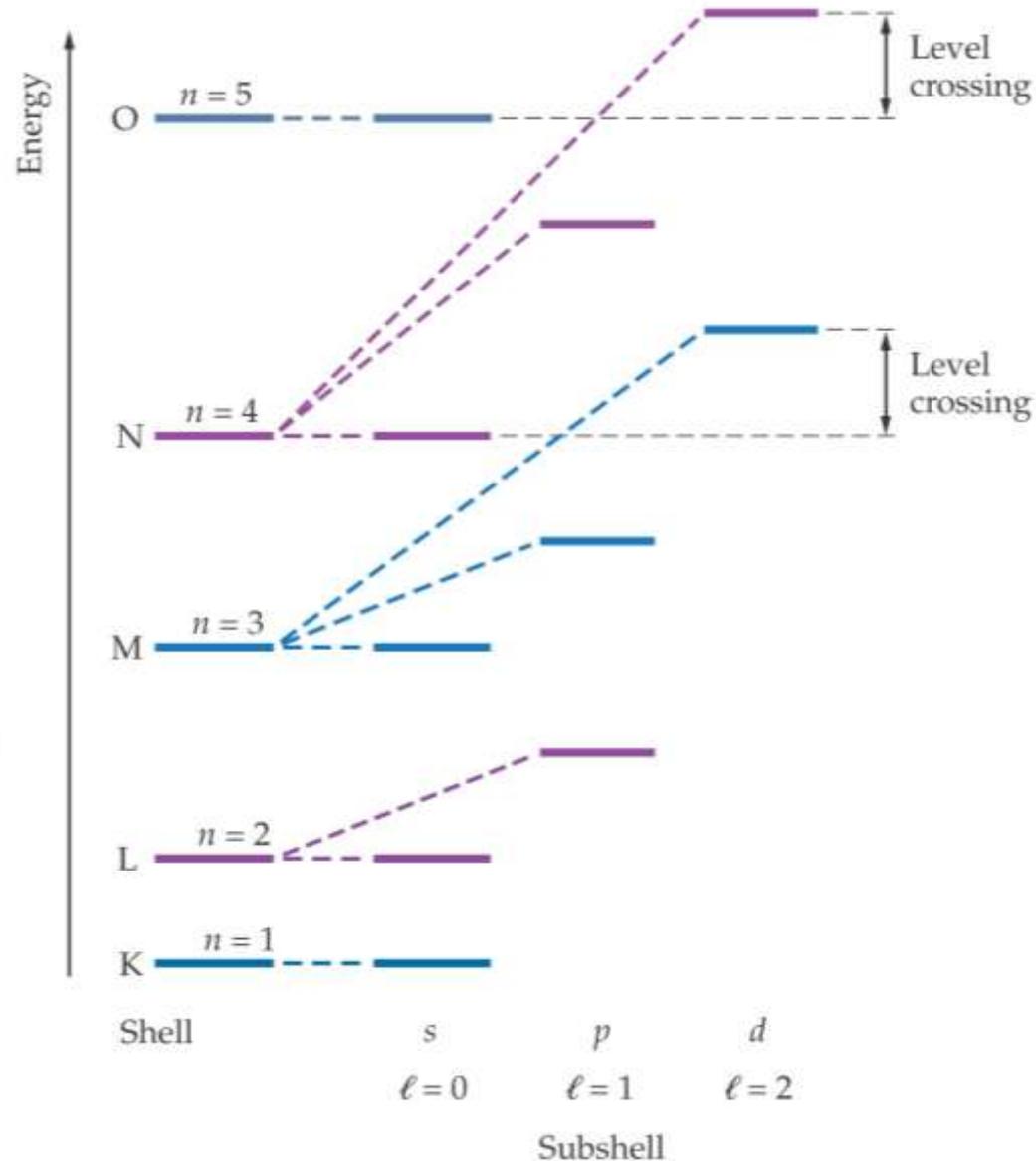


(b)

**(a)  $n = 2, l = 0$ ; (b)  $n = 2, l = 1$**

# 31-6 Multielectron Atoms and the Periodic Table

Atoms with multiple electrons are more complex, and there are no simple formulas for the energy levels. They are still characterized by the same quantum numbers.



# 31-6 Multielectron Atoms and the Periodic Table

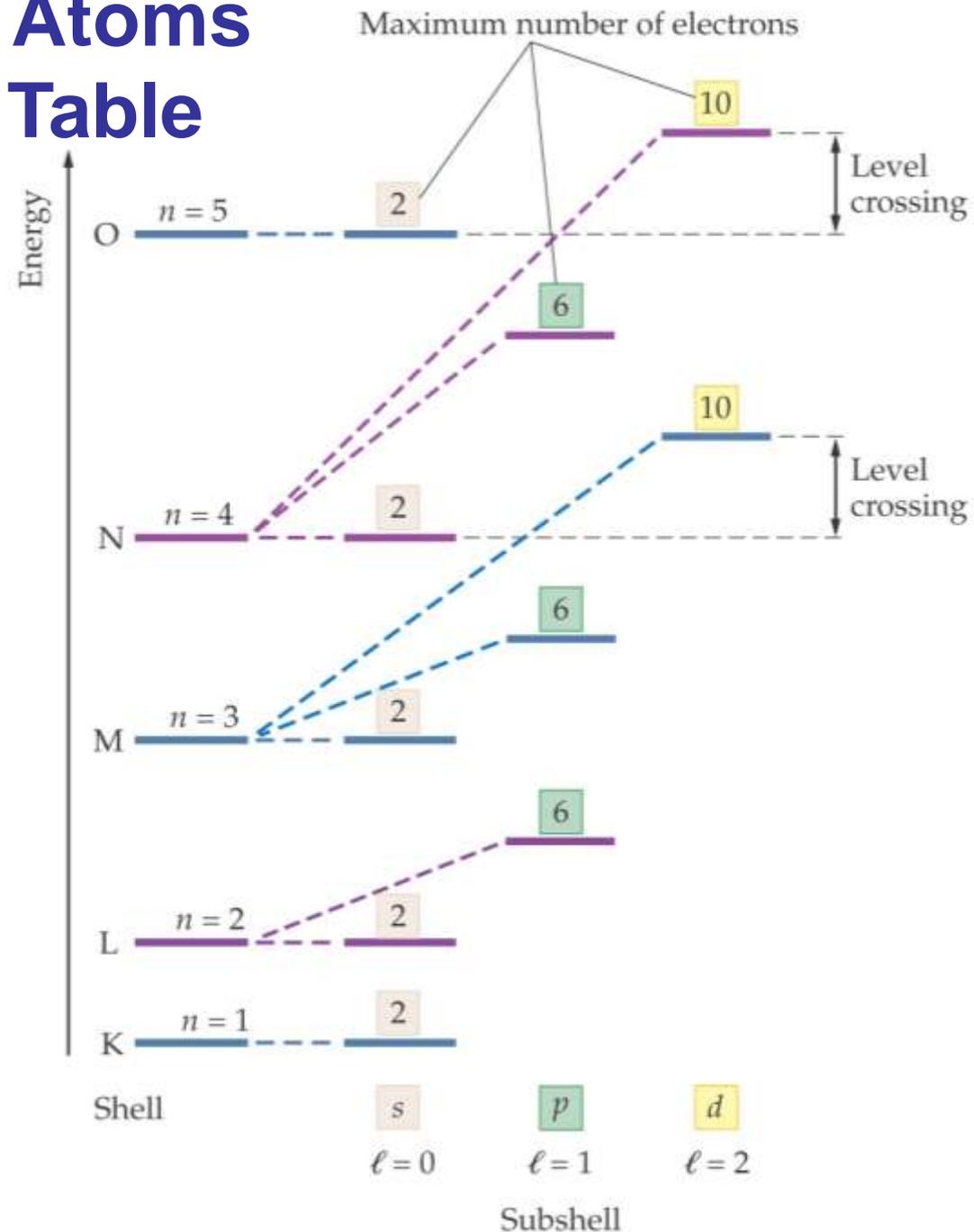
The Pauli exclusion principle states that only one electron may be in each quantum state:

Only one electron at a time may have a particular set of quantum numbers,  $n$ ,  $l$ ,  $m_l$ , and  $m_s$ . Once a particular state is occupied, other electrons are excluded from that state.

Therefore, if electrons are added to an atom, they must go into higher and higher energy states.

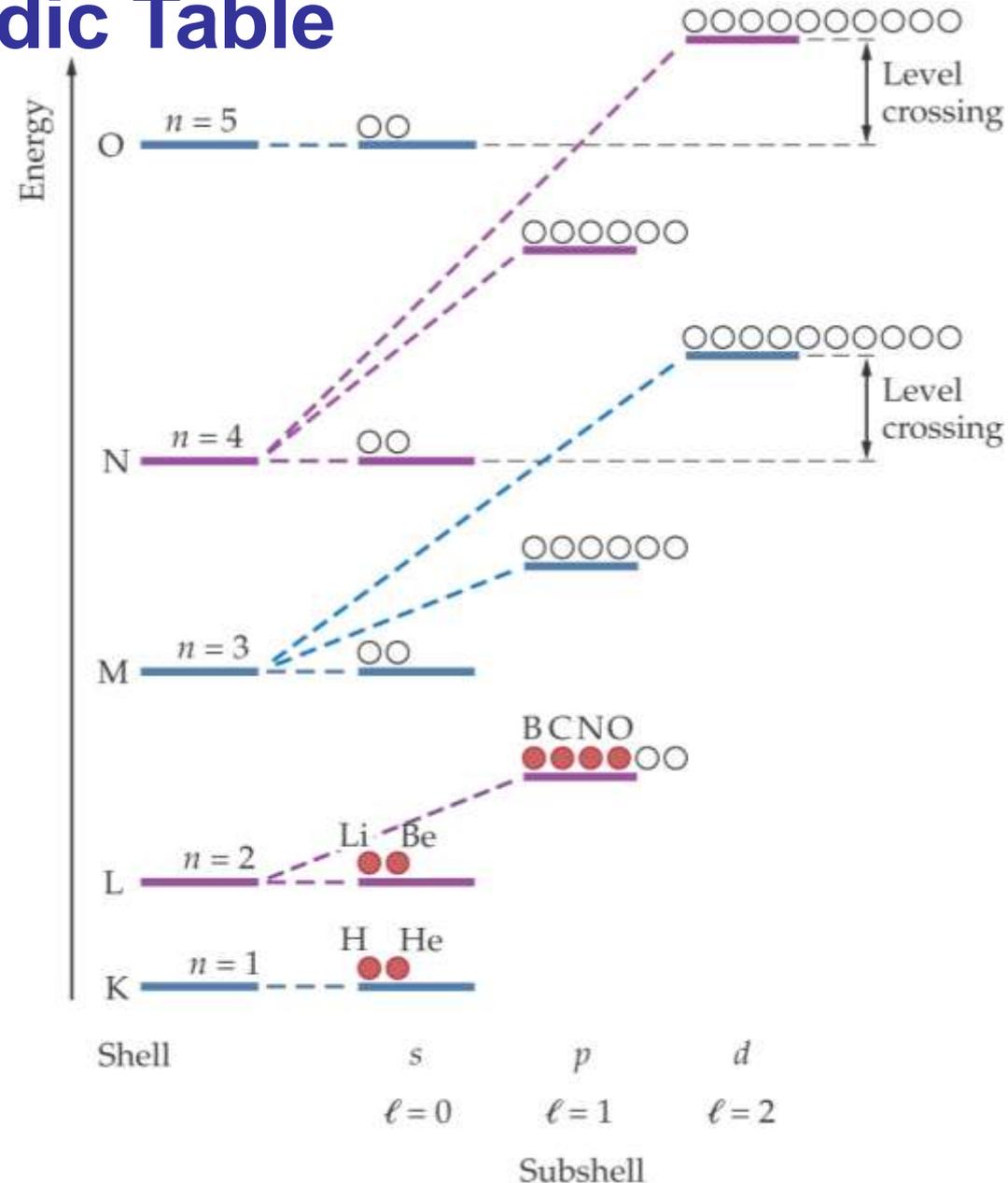
# 31-6 Multielectron Atoms and the Periodic Table

Each energy level can accommodate a certain number of electrons, depending on  $n$ .



# 31-6 Multielectron Atoms and the Periodic Table

The electron arrangements are given here for the first eight elements of the periodic table:



# 31-6 Multielectron Atoms and the Periodic Table

Principal quantum number  
 $n = 3$

Number of electrons in subshell = 6

$3p^6$

Angular momentum quantum number  
 $\ell = 1$  ( $p$ )

A shorthand notation has been devised to describe the configuration of electrons in a particular atom. This diagram shows how to describe a single energy level.

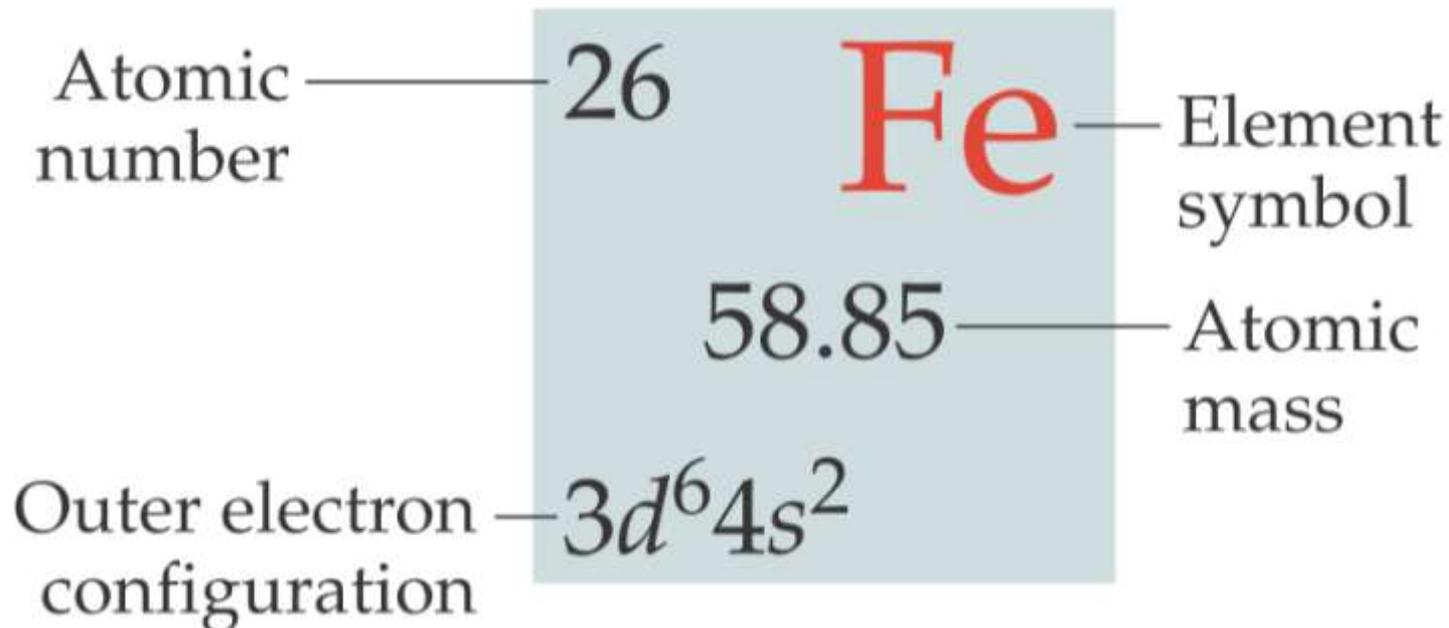
# 31-6 Multielectron Atoms and the Periodic Table

Once every occupied level is described in this way, we have a complete description of the electron configuration of the atom.

Atomic number	Element	Electronic configuration
1	Hydrogen (H)	$1s^1$
2	Helium (He)	$1s^2$
3	Lithium (Li)	$1s^2 2s^1$
4	Beryllium (Be)	$1s^2 2s^2$
5	Boron (B)	$1s^2 2s^2 2p^1$
6	Carbon (C)	$1s^2 2s^2 2p^2$
7	Nitrogen (N)	$1s^2 2s^2 2p^3$
8	Oxygen (O)	$1s^2 2s^2 2p^4$
9	Fluorine (F)	$1s^2 2s^2 2p^5$

# 31-6 Multielectron Atoms and the Periodic Table

The periodic table groups elements with the same structure in their outermost shells together in columns. The description of each element is included.



# 31-6 Multielectron Atoms and the Periodic Table

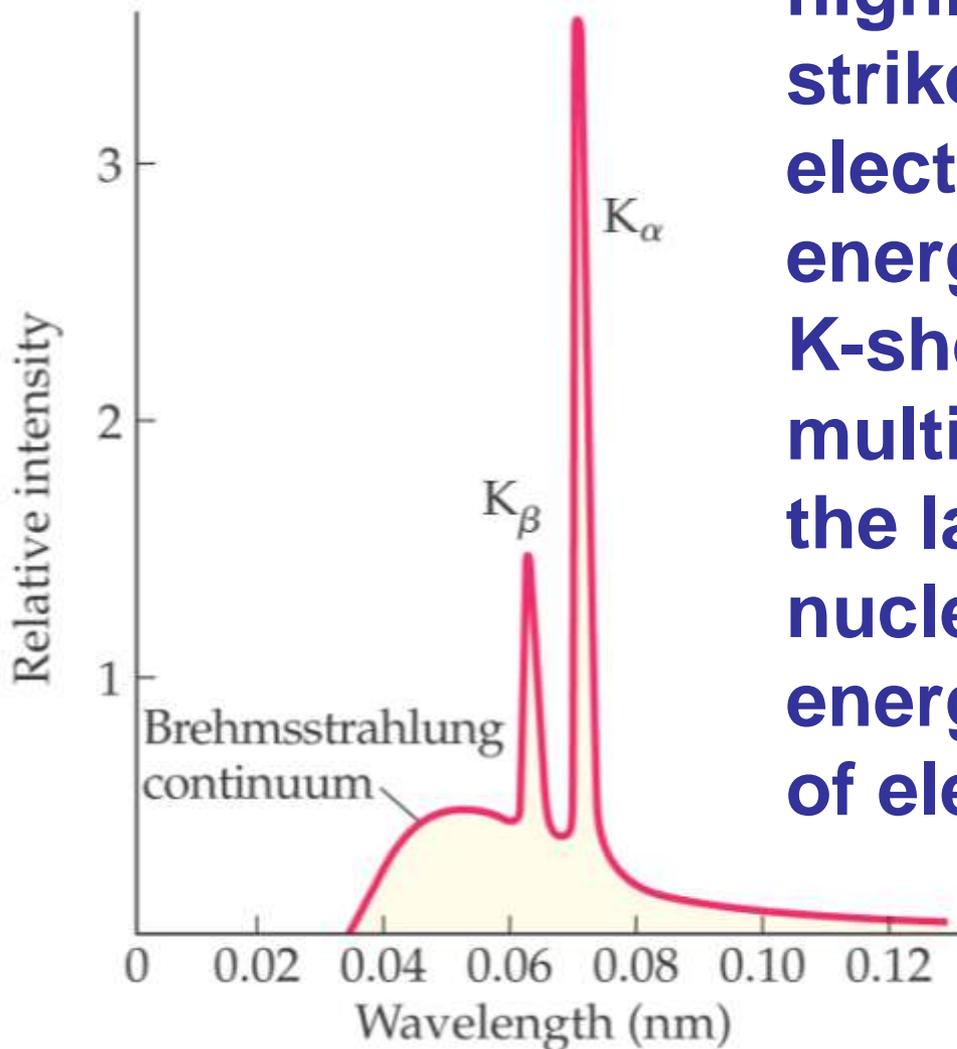
PERIODS	s		d										p					
	GROUP I	GROUP II	Transition elements										GROUP III	GROUP IV	GROUP V	GROUP VI	GROUP VII	GROUP VIII
1	1 <b>H</b> 1.01 1s <sup>1</sup>																	2 <b>He</b> 4.00 1s <sup>2</sup>
2	3 <b>Li</b> 6.94 2s <sup>1</sup>	4 <b>Be</b> 9.01 2s <sup>2</sup>											5 <b>B</b> 10.81 2p <sup>1</sup>	6 <b>C</b> 12.01 2p <sup>2</sup>	7 <b>N</b> 14.01 2p <sup>3</sup>	8 <b>O</b> 16.00 2p <sup>4</sup>	9 <b>F</b> 19.00 2p <sup>5</sup>	10 <b>Ne</b> 20.18 2p <sup>6</sup>
3	11 <b>Na</b> 22.99 3s <sup>1</sup>	12 <b>Mg</b> 24.31 3s <sup>2</sup>											13 <b>Al</b> 26.98 3p <sup>1</sup>	14 <b>Si</b> 28.09 3p <sup>2</sup>	15 <b>P</b> 30.97 3p <sup>3</sup>	16 <b>S</b> 32.07 3p <sup>4</sup>	17 <b>Cl</b> 35.45 3p <sup>5</sup>	18 <b>Ar</b> 39.95 3p <sup>6</sup>
4	19 <b>K</b> 39.10 4s <sup>1</sup>	20 <b>Ca</b> 40.08 4s <sup>2</sup>	21 <b>Sc</b> 44.96 3d <sup>1</sup> 4s <sup>2</sup>	22 <b>Ti</b> 47.88 3d <sup>2</sup> 4s <sup>2</sup>	23 <b>V</b> 50.94 3d <sup>3</sup> 4s <sup>2</sup>	24 <b>Cr</b> 52.00 3d <sup>5</sup> 4s <sup>1</sup>	25 <b>Mn</b> 54.94 3d <sup>5</sup> 4s <sup>2</sup>	26 <b>Fe</b> 55.85 3d <sup>6</sup> 4s <sup>2</sup>	27 <b>Co</b> 58.93 3d <sup>7</sup> 4s <sup>2</sup>	28 <b>Ni</b> 58.69 3d <sup>8</sup> 4s <sup>2</sup>	29 <b>Cu</b> 63.55 3d <sup>10</sup> 4s <sup>1</sup>	30 <b>Zn</b> 65.39 3d <sup>10</sup> 4s <sup>2</sup>	31 <b>Ga</b> 69.72 4p <sup>1</sup>	32 <b>Ge</b> 72.61 4p <sup>2</sup>	33 <b>As</b> 74.92 4p <sup>3</sup>	34 <b>Se</b> 78.96 4p <sup>4</sup>	35 <b>Br</b> 79.90 4p <sup>5</sup>	36 <b>Kr</b> 83.80 4p <sup>6</sup>
5	37 <b>Rb</b> 85.47 5s <sup>1</sup>	38 <b>Sr</b> 87.62 5s <sup>2</sup>	39 <b>Y</b> 88.96 4d <sup>1</sup> 5s <sup>2</sup>	40 <b>Zr</b> 91.22 4d <sup>2</sup> 5s <sup>2</sup>	41 <b>Nb</b> 92.91 4d <sup>4</sup> 5s <sup>1</sup>	42 <b>Mo</b> 95.94 4d <sup>5</sup> 5s <sup>1</sup>	43 <b>Tc</b> (98) 4d <sup>5</sup> 5s <sup>2</sup>	44 <b>Ru</b> 101.07 4d <sup>7</sup> 5s <sup>1</sup>	45 <b>Rh</b> 102.91 4d <sup>8</sup> 5s <sup>1</sup>	46 <b>Pd</b> 106.42 4d <sup>10</sup> 5s <sup>0</sup>	47 <b>Ag</b> 107.87 4d <sup>10</sup> 5s <sup>1</sup>	48 <b>Cd</b> 112.41 4d <sup>10</sup> 5s <sup>2</sup>	49 <b>In</b> 114.82 5p <sup>1</sup>	50 <b>Sn</b> 118.71 5p <sup>2</sup>	51 <b>Sb</b> 121.76 5p <sup>3</sup>	52 <b>Te</b> 127.60 5p <sup>4</sup>	53 <b>I</b> 126.90 5p <sup>5</sup>	54 <b>Xe</b> 131.29 5p <sup>6</sup>
6	55 <b>Cs</b> 132.91 6s <sup>1</sup>	56 <b>Ba</b> 137.33 6s <sup>2</sup>	57 <b>La</b> 138.91 5d <sup>1</sup> 6s <sup>2</sup>	72 <b>Hf</b> 178.49 5d <sup>2</sup> 6s <sup>2</sup>	73 <b>Ta</b> 180.95 5d <sup>3</sup> 6s <sup>2</sup>	74 <b>W</b> 183.85 5d <sup>4</sup> 6s <sup>2</sup>	75 <b>Re</b> 186.21 5d <sup>5</sup> 6s <sup>2</sup>	76 <b>Os</b> 190.2 5d <sup>6</sup> 6s <sup>2</sup>	77 <b>Ir</b> 192.22 5d <sup>7</sup> 6s <sup>2</sup>	78 <b>Pt</b> 195.08 5d <sup>9</sup> 6s <sup>1</sup>	79 <b>Au</b> 196.97 5d <sup>10</sup> 6s <sup>1</sup>	80 <b>Hg</b> 200.59 5d <sup>10</sup> 6s <sup>2</sup>	81 <b>Tl</b> 204.36 6p <sup>1</sup>	82 <b>Pb</b> 207.2 6p <sup>2</sup>	83 <b>Bi</b> 208.98 6p <sup>3</sup>	84 <b>Po</b> (209) 6p <sup>4</sup>	85 <b>At</b> (210) 6p <sup>5</sup>	86 <b>Rn</b> (222) 6p <sup>6</sup>
7	87 <b>Fr</b> (223) 7s <sup>1</sup>	88 <b>Ra</b> 226.03 7s <sup>2</sup>	89 <b>Ac</b> 227.03 6d <sup>1</sup> 7s <sup>2</sup>	104 <b>Rf</b> (261) 6d <sup>2</sup> 7s <sup>2</sup>	105 <b>Db</b> (262) 6d <sup>3</sup> 7s <sup>2</sup>	106 <b>Sg</b> (266) 6d <sup>4</sup> 7s <sup>2</sup>	107 <b>Bh</b> (264) 6d <sup>5</sup> 7s <sup>2</sup>	108 <b>Hs</b> (269) 6d <sup>6</sup> 7s <sup>2</sup>	109 <b>Mt</b> (268) 6d <sup>7</sup> 7s <sup>2</sup>	110 (271)	111 (272)	112 (277)		114 (289)		116 (289)		118 (293)

Atomic number — 26 **Fe** — Symbol  
 Outer electron configuration — 3d<sup>6</sup>4s<sup>2</sup> — Atomic mass — 58.85

	f														
*	58 <b>Ce</b> 140.12 5d <sup>1</sup> 4f <sup>1</sup> 6s <sup>2</sup>	59 <b>Pr</b> 140.91 4f <sup>3</sup> 6s <sup>2</sup>	60 <b>Nd</b> 144.24 4f <sup>4</sup> 6s <sup>2</sup>	61 <b>Pm</b> (145) 4f <sup>5</sup> 6s <sup>2</sup>	62 <b>Sm</b> 150.36 4f <sup>6</sup> 6s <sup>2</sup>	63 <b>Eu</b> 151.96 4f <sup>7</sup> 6s <sup>2</sup>	64 <b>Gd</b> 157.25 5d <sup>1</sup> 4f <sup>7</sup> 6s <sup>2</sup>	65 <b>Tb</b> 158.93 5d <sup>1</sup> 4f <sup>8</sup> 6s <sup>2</sup>	66 <b>Dy</b> 162.50 4f <sup>10</sup> 6s <sup>2</sup>	67 <b>Ho</b> 164.93 4f <sup>11</sup> 6s <sup>2</sup>	68 <b>Er</b> 167.26 4f <sup>12</sup> 6s <sup>2</sup>	69 <b>Tm</b> 168.93 4f <sup>13</sup> 6s <sup>2</sup>	70 <b>Yb</b> 173.04 4f <sup>14</sup> 6s <sup>2</sup>	71 <b>Lu</b> 174.97 5d <sup>1</sup> 4f <sup>14</sup> 6s <sup>2</sup>	Lanthanides
†	90 <b>Th</b> 232.04 6d <sup>2</sup> 7s <sup>2</sup>	91 <b>Pa</b> 231.04 5f <sup>2</sup> 6d <sup>1</sup> 7s <sup>2</sup>	92 <b>U</b> 238.03 5f <sup>3</sup> 6d <sup>1</sup> 7s <sup>2</sup>	93 <b>Np</b> 237.05 5f <sup>4</sup> 6d <sup>1</sup> 7s <sup>2</sup>	94 <b>Pu</b> (244) 5f <sup>6</sup> 6d <sup>0</sup> 7s <sup>2</sup>	95 <b>Am</b> (243) 5f <sup>7</sup> 6d <sup>0</sup> 7s <sup>2</sup>	96 <b>Cm</b> (247) 5f <sup>7</sup> 6d <sup>1</sup> 7s <sup>2</sup>	97 <b>Bk</b> (247) 5f <sup>8</sup> 6d <sup>1</sup> 7s <sup>2</sup>	98 <b>Cf</b> (251) 5f <sup>10</sup> 6d <sup>0</sup> 7s <sup>2</sup>	99 <b>Es</b> (252) 5s <sup>11</sup> 6d <sup>0</sup> 7s <sup>2</sup>	100 <b>Fm</b> (257) 5f <sup>12</sup> 6d <sup>0</sup> 7s <sup>2</sup>	101 <b>Md</b> (258) 5f <sup>13</sup> 6d <sup>0</sup> 7s <sup>2</sup>	102 <b>No</b> (259) 5f <sup>14</sup> 6d <sup>0</sup> 7s <sup>2</sup>	103 <b>Lr</b> (262) 5f <sup>14</sup> 6d <sup>1</sup> 7s <sup>2</sup>	Actinides

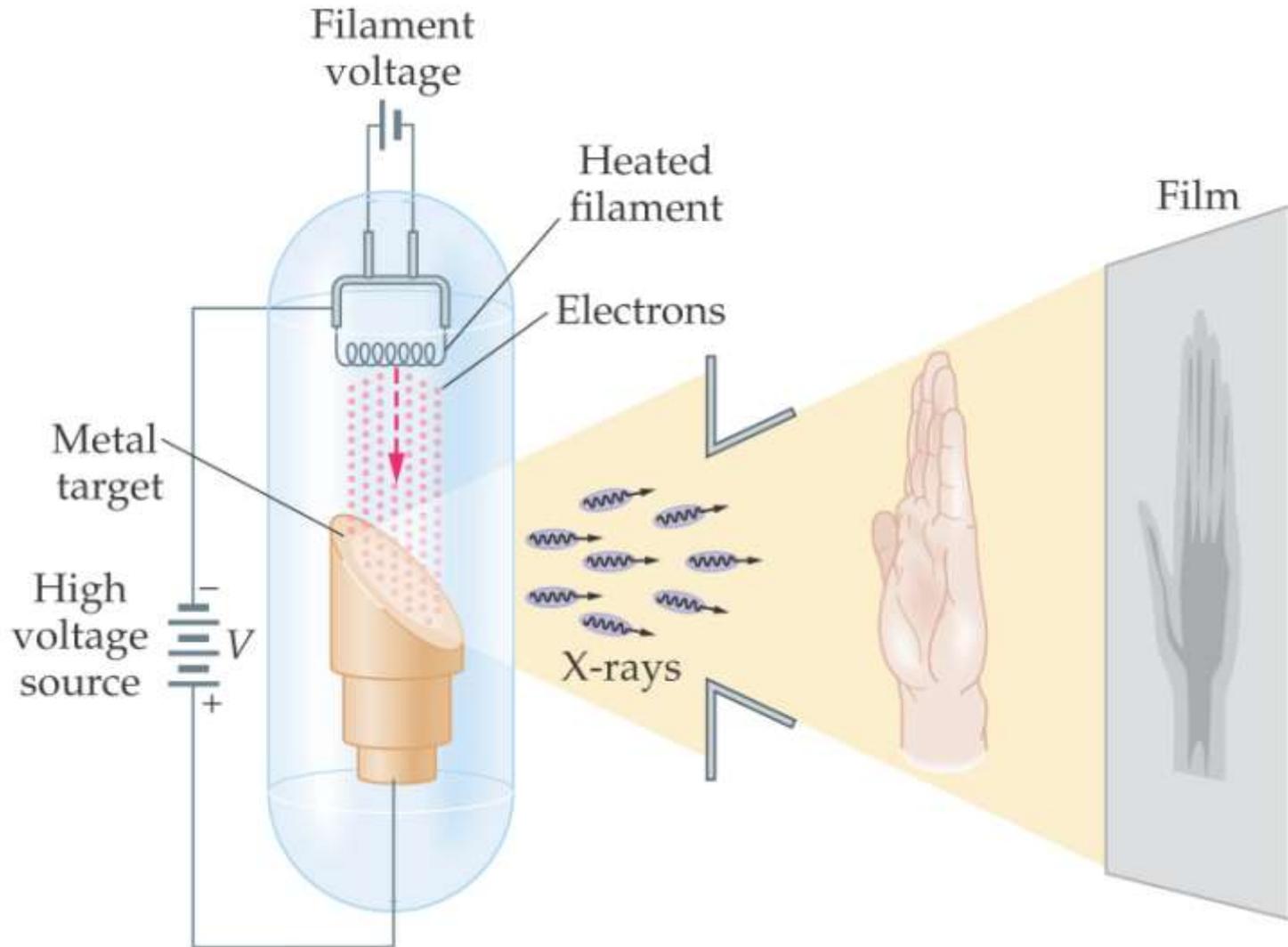
# 31-7 Atomic Radiation

X-rays are emitted when highly energetic electrons strike a metal target. If the electrons have enough energy, they can knock out a K-shell ( $n = 1$ ) electron in a multielectron atom. Due to the large charge of the nucleus, this can take an energy of tens of thousands of electron volts.



# 31-7 Atomic Radiation

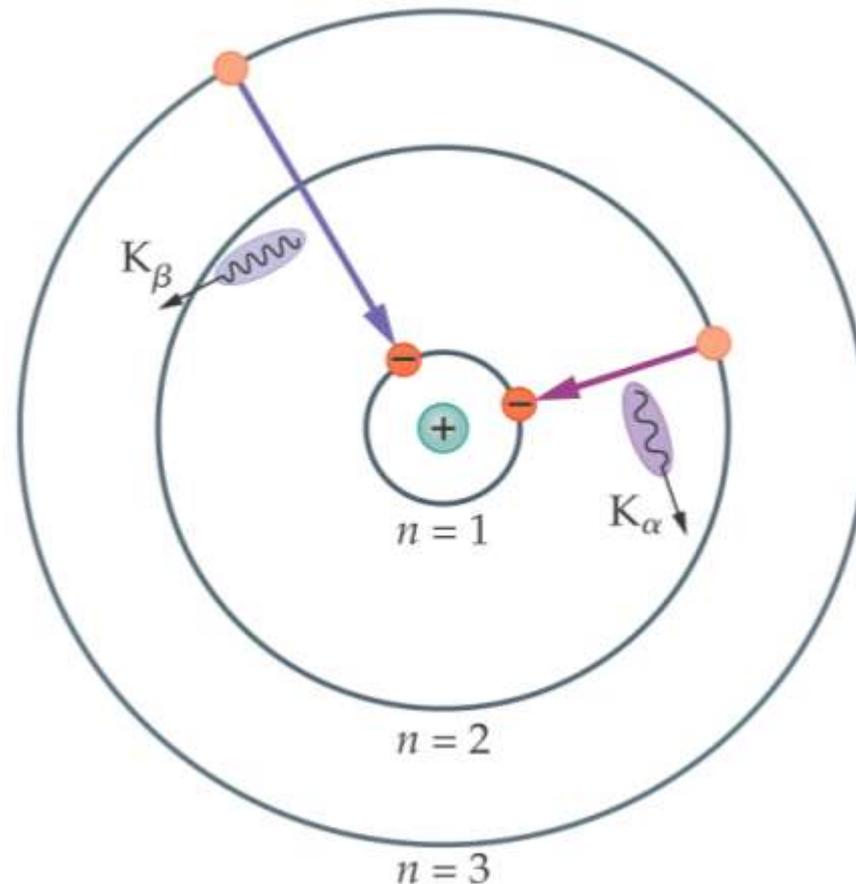
**X-rays are used for medical imaging, among other applications.**



# 31-7 Atomic Radiation

The energy of a K-shell electron is given by:

$$E_K = -(13.6 \text{ eV}) \frac{(Z - 1)^2}{1^2}$$



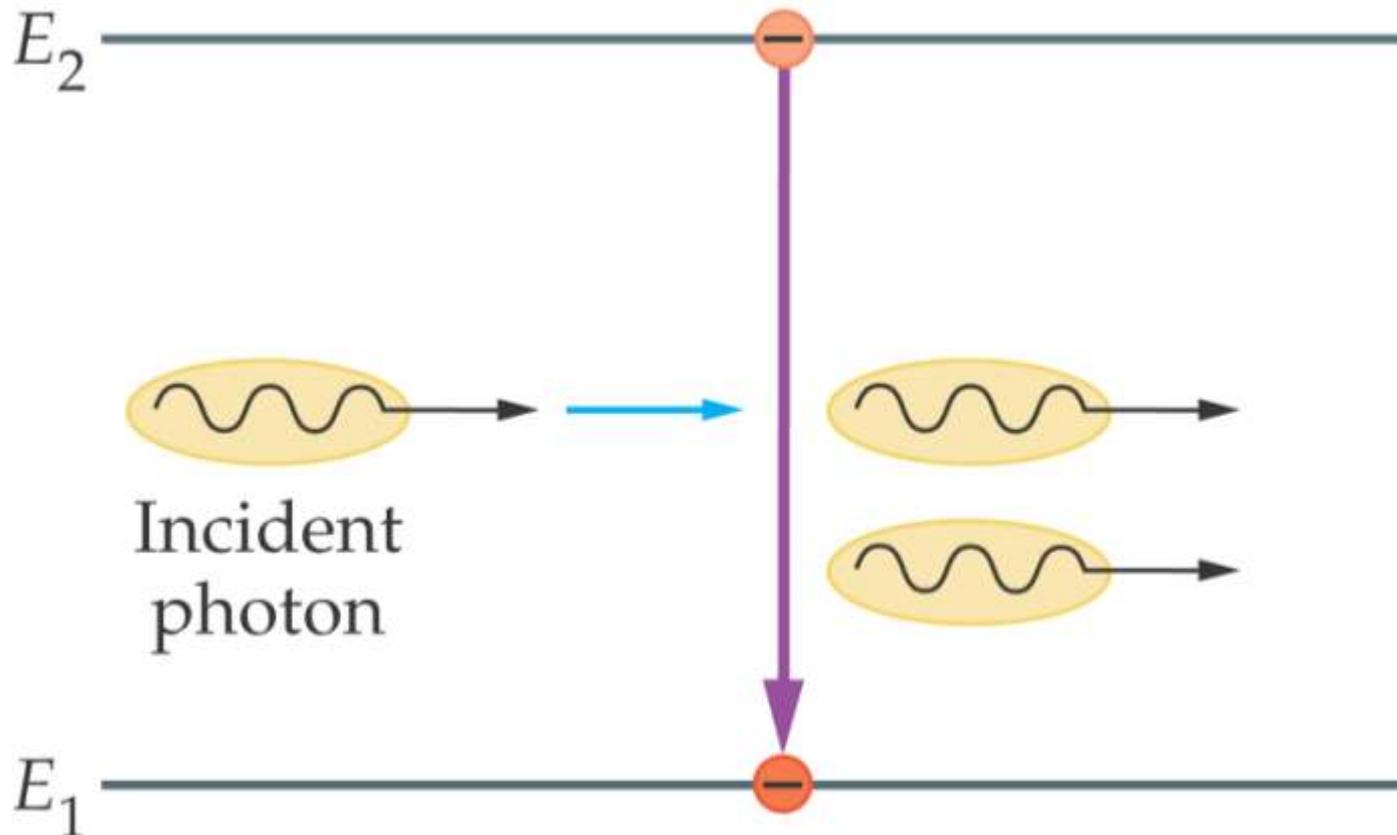
# 31-7 Atomic Radiation

The word “laser” is an acronym for **l**ight **a**mplification by the **s**timulated **e**mission of **r**adiation.

Spontaneous emission occurs when an electron in an excited state drops to a lower state, emitting a photon in the process. The photons are emitted in random directions.

# 31-7 Atomic Radiation

Stimulated emission occurs when the atom is bombarded with photons of the emission wavelength; this stimulates the transition to occur.



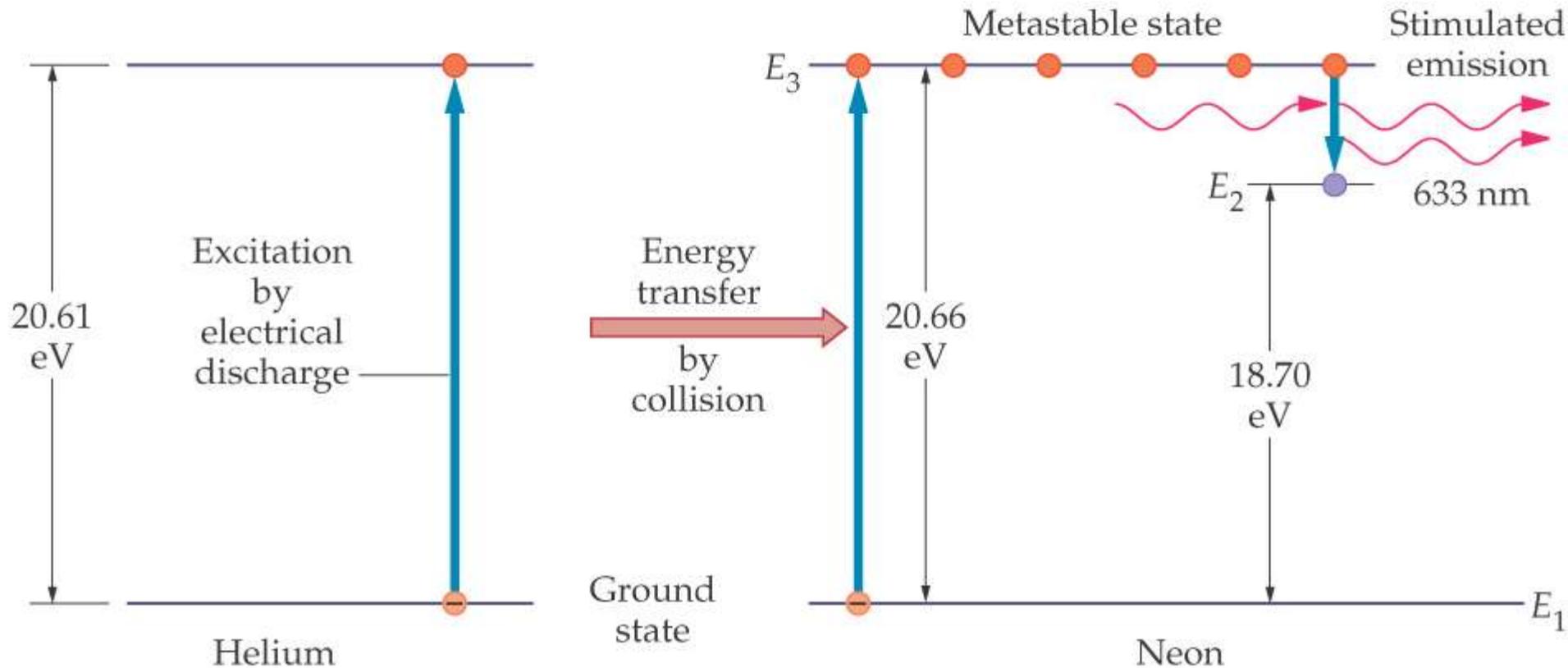
## **31-7 Atomic Radiation**

**The stimulated emission is in the same direction as the incident photon, and the photons have the same phase. Since one incoming photon produces two outgoing photons, the light intensity is amplified.**

**This process will not sustain itself unless there is a continuing supply of atoms in excited states. This is ensured by exciting electrons into metastable states – excited states that cannot decay directly to the ground state and therefore have relatively long lifetimes.**

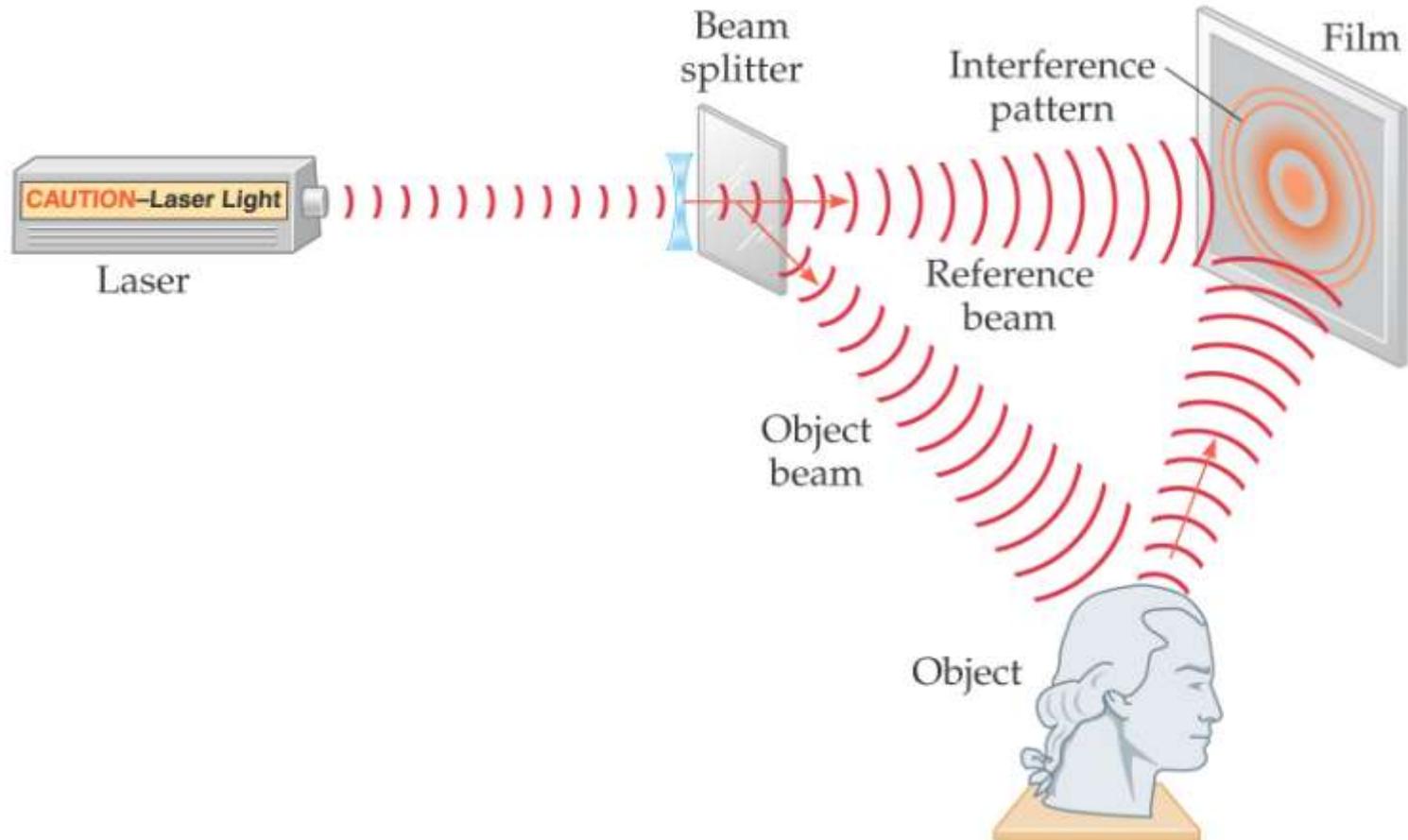
# 31-7 Atomic Radiation

This diagram shows the energy levels involved in a helium-neon laser.



# 31-7 Atomic Radiation

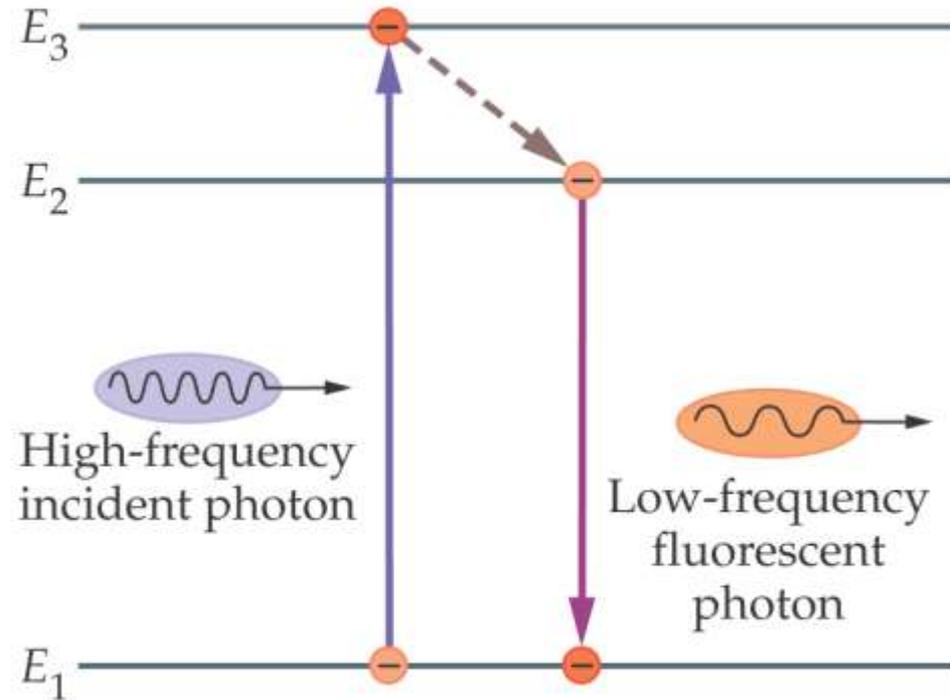
The coherence of laser light is critical in the manufacture of holograms, as information is contained in the interference pattern.



# 31-7 Atomic Radiation

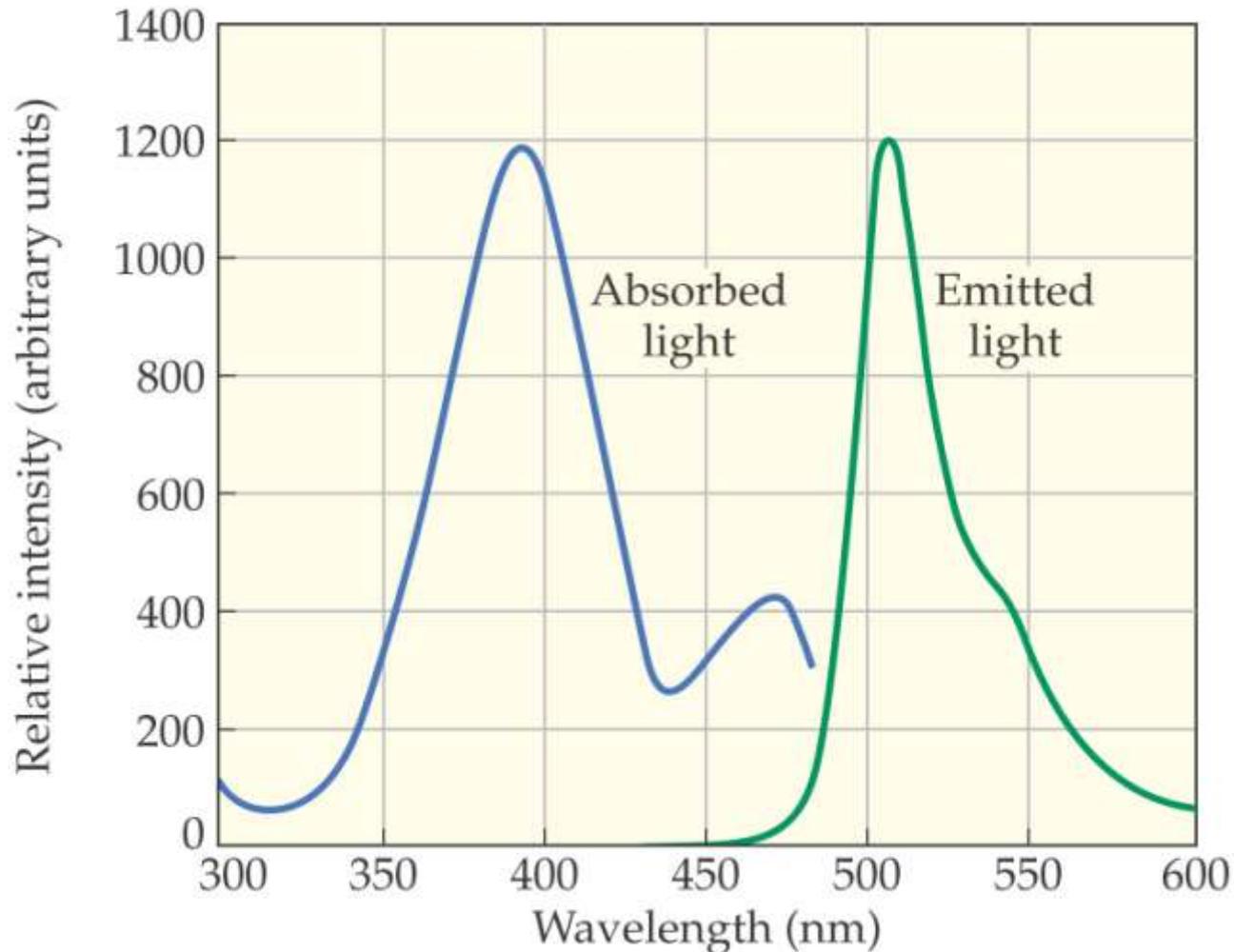
Fluorescence and phosphorescence occur when electrons emit photons of various frequencies when returning to the ground state.

In fluorescence, ultraviolet light excites atoms into metastable states. They then decay in two or more steps, emitting visible light. This is how “black lights” work.



# 31-7 Atomic Radiation

This is an example of a fluorescent compound, green fluorescent protein (GFP).



# 31-7 Atomic Radiation

**Phosphorescence is similar, except that the excited state is metastable, allowing the object to continue to glow long after light has ceased falling on it. This is how glow-in-the-dark paints and plastics work.**

# Summary of Chapter 31

- The plum pudding model had a positively charged “pudding” embedded with electrons.
- Rutherford showed that the atom was mostly empty space, leading to the solar system model.
- Excited atoms of hydrogen emit light at specific wavelengths:

$$\frac{1}{\lambda} = R \left( \frac{1}{n'^2} - \frac{1}{n^2} \right) \quad n' = 1, 2, 3, \dots$$
$$n = n' + 1, n' + 2, n' + 3, \dots$$

# Summary of Chapter 31

- **Bohr model: electrons move in circular orbits around the nucleus; only certain angular momentum values are allowed; electrons do not radiate when in allowed orbits; radiation is emitted when electrons jump from one energy level to another.**
- **Radii of allowed orbits:**

$$r_n = \left( \frac{h^2}{4\pi^2 m k Z e^2} \right) n^2 = (5.29 \times 10^{-11} \text{ m}) n^2 \quad n = 1, 2, 3, \dots$$

# Summary of Chapter 31

- **Energy of allowed orbits:**

$$E_n = -(13.6 \text{ eV}) \frac{Z^2}{n^2} \quad n = 1, 2, 3, \dots$$

- **Allowed orbits correspond to standing matter waves of the electrons.**
- **Quantum mechanically, atomic electrons are described by four quantum numbers,  $n$ ,  $l$ ,  $m_l$ , and  $m_s$ .**
- **Electrons do not have a precise position, and are described by a probability cloud.**

# Summary of Chapter 31

- Only one electron in an atom can have any given set of quantum numbers.
- Energy levels increase with  $n$  and  $l$ .
- Elements of increasing  $Z$  add electrons into the lowest available level.
- Atoms with the same configuration of outer shell electrons have similar chemical properties.
- Atoms emit X-rays when an inner-shell electron is knocked out.

# Summary of Chapter 31

- **A laser produces light amplification through the stimulated emission of radiation.**
- **An electron in an excited state may return to the ground state in two or more steps, emitting longer-wavelength light. This is the source of fluorescence and phosphorescence.**