Chapter 7
Electronic Structure of Atoms

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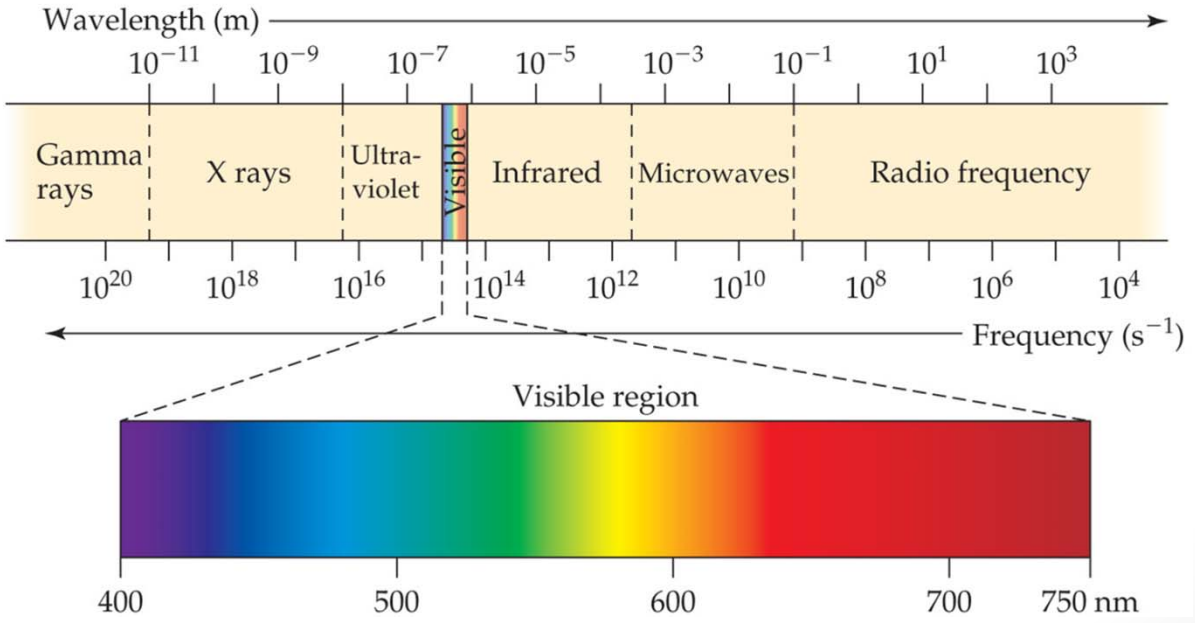
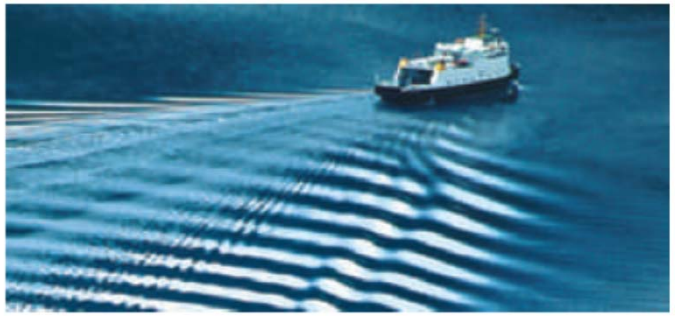
The Wave Nature of Light

- The light we see with our eyes, visible light, is one type of electromagnetic radiation.
  - ➢ electromagnetic radiation carries energy through space, it is also known as radiant energy.
- There are many types of electromagnetic radiation in addition to visible light.



Electromagnetic radiation

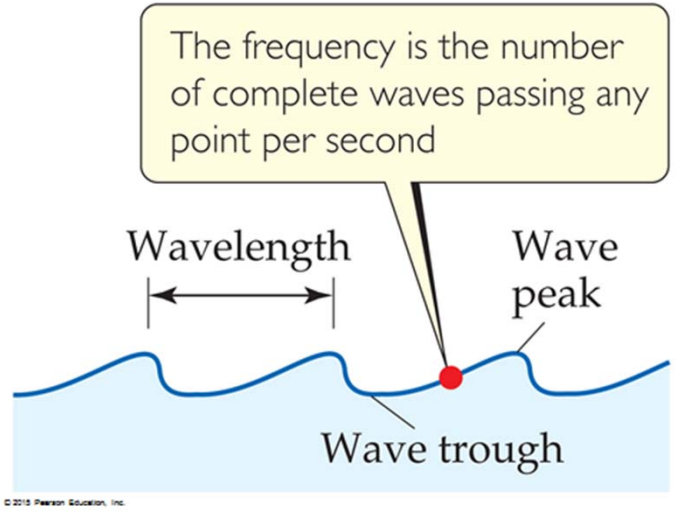
- All types of electromagnetic radiation move through a vacuum at $2.998 * 10^8$ m/s, the speed of light.
- All types of electromagnetic radiation have wave-like characteristics similar to those of waves that move through water.



▲ Figure 6.1 **Water waves.** The movement of a boat through the water forms waves. The regular variation of peaks and troughs enables us to sense the motion of the waves away from the boat.

Waves

- A cross section of a water wave shows that it is periodic, which means that the pattern of peaks and troughs repeats itself at regular intervals.
  - The distance between two adjacent peaks (or between two adjacent troughs) is called the wavelength.
  - The number of complete wavelengths, or cycles, that pass a given point each second is the frequency of the wave.

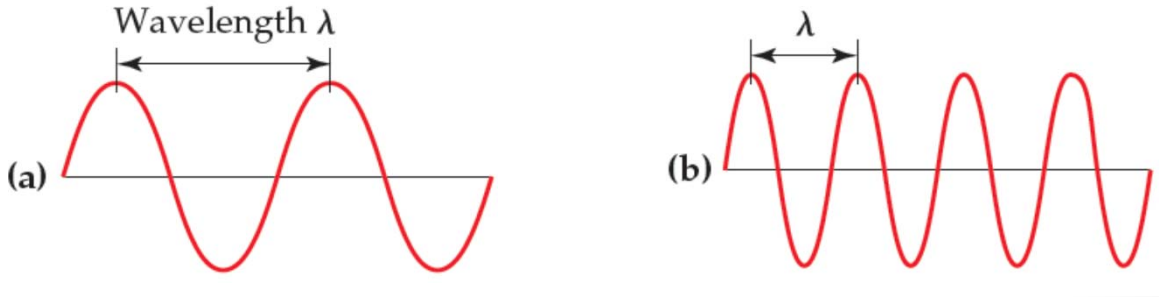


Waves

- The number of waves passing a given point per unit of time is the frequency (ν).
- For waves traveling at the same velocity, the longer the wavelength, the smaller the frequency.
- This inverse relationship between the frequency and wavelength of electromagnetic radiation is expressed by the equation

$$\lambda\nu = c$$

λ (lambda) is wavelength, ν (nu) is frequency, and c is the speed of light.



Electromagnetic spectrum

- The wavelengths of electromagnetic radiation span an enormous range.
  - The wavelengths of gamma rays are comparable to the diameters of atomic nuclei. (10^{-11} m)
  - The wavelengths of radio waves can be longer than a football field. (km)

Table 6.1 Common Wavelength Units for Electromagnetic Radiation

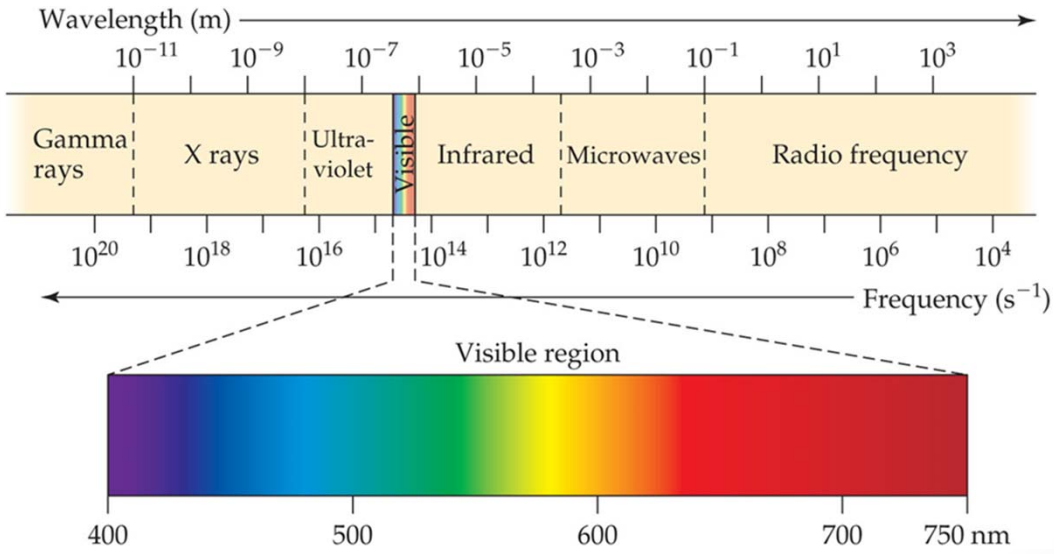
Unit	Symbol	Length (m)	Type of Radiation
Angstrom	Å	10^{-10}	X ray
Nanometer	nm	10^{-9}	Ultraviolet, visible
Micrometer	μm	10^{-6}	Infrared
Millimeter	mm	10^{-3}	Microwave
Centimeter	cm	10^{-2}	Microwave
Meter	m	1	Television, radio
Kilometer	km	1000	Radio

Frequency

- Frequency is expressed in cycles per second, a unit also called a hertz (Hz).
- The units of frequency are normally given simply as "per second," which is denoted by s^{-1} or /s.

EX: a frequency of 698 megahertz (MHz)

698 MHz, 698,000,000 Hz, 698,000,000 s^{-1}, or 698,000,000/s.



Exercise 6.2

- The yellow light given off by a sodium vapor lamp used for public lighting has a wavelength of 589 nm. What is the frequency of this radiation?

Sol:

$$\lambda\nu = c$$

$$\nu = \frac{c}{\lambda} = \left(\frac{3.00 \times 10^8\ \text{m/s}}{589\ \text{nm}}\right)\left(\frac{1\ \text{nm}}{10^{-9}\ \text{m}}\right) = 5.09 \times 10^{14}\ \text{s}^{-1}$$

Quantized Energy and Photons

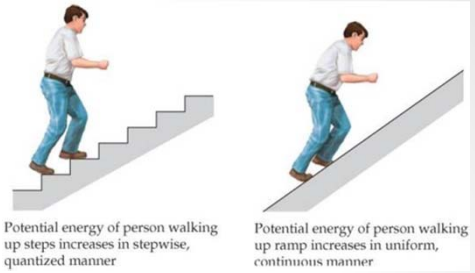
- Three of these are particularly pertinent to our understanding of how electromagnetic radiation and atoms interact:
  - the emission of light from hot objects (referred to as blackbody radiation because the objects studied appear black before heating)
  - the emission of electrons from metal surfaces on which light shines (the photoelectric effect)
  - the emission of light from electronically excited gas atoms (emission spectra)

Hot Objects

- When solids are heated, they emit radiation.
- The wavelength distribution of the radiation depends on temperature; a red-hot object, for instance, is cooler than a yellowish or white-hot one

▲ Figure 6.5 **Color and temperature.** The color and intensity of the light emitted by a hot object, such as this pour of molten steel, depend on the temperature of the object.

Quantization of Energy



Potential energy of person walking up steps increases in stepwise, quantized manner

Potential energy of person walking up ramp increases in uniform, continuous manner

- Max Planck (1858–1947) proposed that energy can be either released or absorbed by atoms only in discrete "chunks" of some minimum size.
  - Planck gave the name quantum (meaning "fixed amount") to the smallest quantity of energy that can be emitted or absorbed as electromagnetic radiation.

$$E = h\nu$$

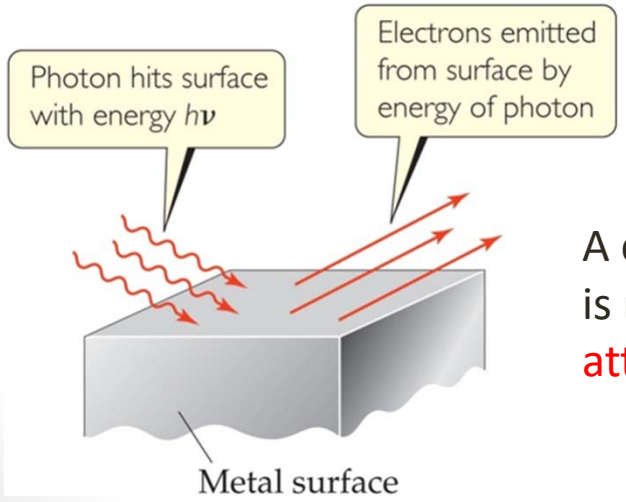
The constant h is called Planck constant and has a value of $6.626 * 10^{-34}$ joule-second (J-s).

- Matter can emit and absorb energy only in whole number multiples of $h\nu$, such as $h\nu$, $2h\nu$, $3h\nu$, and so forth.

The Photoelectric Effect

- Albert Einstein (1879– 1955) used Planck's theory to explain the photoelectric effect
- Light shining on a clean metal surface causes electrons to be emitted from the surface.
  - ➢ Each metal has a different energy at which it ejects electrons.

  Ex: light with a frequency of $4.60 * 10^{14}\ s^{-1}$ or greater causes cesium metal to emit electrons, but if the light has frequency less than that, no electrons are emitted.



$$\text{Energy of photon} = E = h\nu$$

A certain amount of energy—the work function—is required for the electrons to overcome the attractive forces holding them in the metal.

Exercise 6.3

- Calculate the energy of one photon of yellow light that has a wavelength of 589 nm.

Sol:

$\nu = c/\lambda = 3 \times 10^{8} / (589 \times 10^{-9}) = 5.09 * 10^{14}\ s^{-1}$

$E = h\nu = 6.626 * 10^{-34}\ (J\text{-}s)(5.09 * 10^{14}\ s^{-1}) = 3.37 * 10^{-19}$ J/photon

If one photon of radiant energy supplies $3.37 * 10^{-19}$ J, we calculate that one mole of these photons will supply:

$$(6.02 \times 10^{23}\ \text{photons/mol})(3.37 \times 10^{-19}\ \text{J/photon}) = 2.03 \times 10^{5}\ \text{J/mol}$$

Emission spectra

- Radiation composed of a single wavelength is monochromatic.

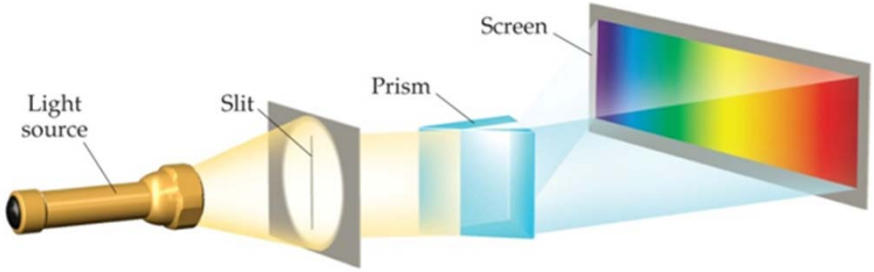
EX: light from a laser.

- Radiation sources produce radiation containing many different wavelengths and is polychromatic.

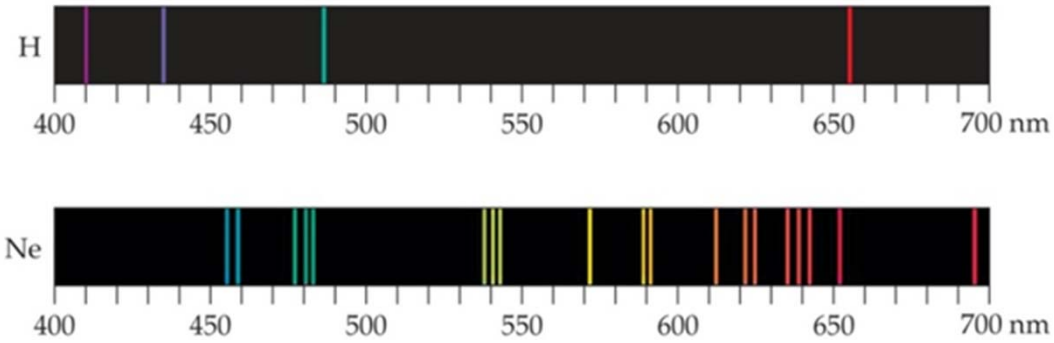
Ex: light bulbs and stars

Continuous vs. Line Spectra

- For atoms and molecules, one does not observe a continuous spectrum (the "rainbow"), as one gets from a white light source.

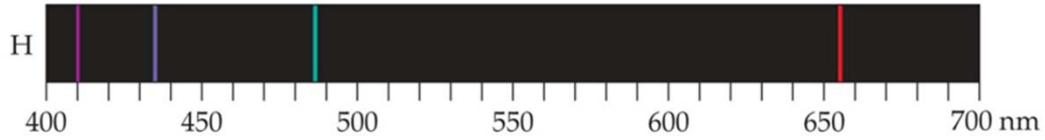


- Only a line spectrum of discrete wavelengths is observed. Each element has a unique line spectrum.



Line spectra of hydrogen

- The line spectrum of hydrogen is at wavelengths of 410 nm (violet), 434 nm (blue), 486 nm (blue-green), and 656 nm (red).



- Johannes Rydberg advanced this formula.

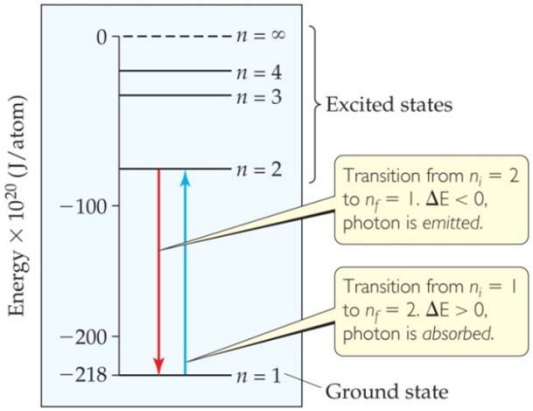
$$\frac{1}{\lambda} = (R_H)\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

➢ λ is the wavelength of a spectral line, R_H is the Rydberg constant ($1.096776 * 10^7\ m^{-1}$), n_1 and n_2 are positive integers, with n_2 being larger than n_1.

- Neils Bohr explained why this mathematical relationship works.

The Bohr Model

- Bohr based his model on three postulates:
  - Only orbits of certain radii, corresponding to certain specific energies, are permitted for the electron in a hydrogen atom.
  - An electron in a permitted orbit is in an "allowed" energy state. An electron in an allowed energy state does not radiate energy and, therefore, does not spiral into the nucleus.
  - Energy is emitted or absorbed by the electron only as the electron changes from one allowed energy state to another. This energy is emitted or absorbed as a photon that has energy $E = h\nu$.

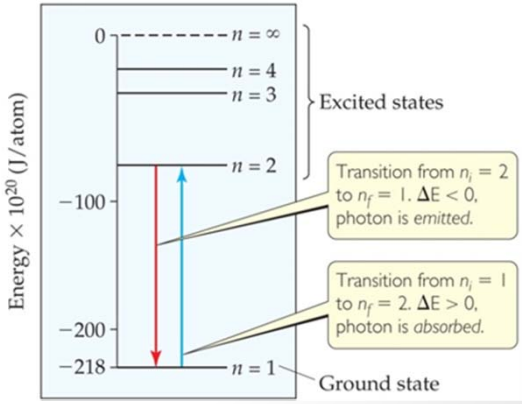


The Bohr Model

- Bohr calculated the energies corresponding to the allowed orbits for the electron in the hydrogen atom.

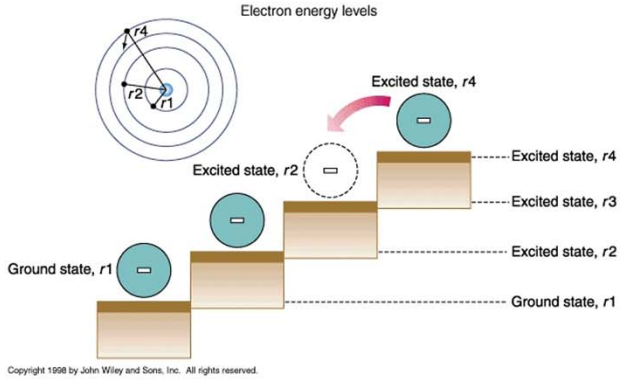
$$E = (-hcR_H)\left(\frac{1}{n^2}\right) = (-2.18 \times 10^{-18}\,\text{J})\left(\frac{1}{n^2}\right)$$

  - where h, c, and R_H are the Planck constant, the speed of light, and the Rydberg constant, respectively.
  - The integer n is called the principal quantum number (主量子數).
  - The lowest-energy state (n = 1) is called the ground state of the atom.
  - When the electron is in a higher-energy state (n = 2 or higher), the atom is said to be in an excited state.



The Bohr Model

- Bohr assumed that the electron can "jump" from one allowed orbit to another by either absorbing or emitting photons whose radiant energy corresponds exactly to the energy difference between the two orbits.



$$\Delta E = -hcR_H\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

$$\Delta E = E_f - E_i = (-2.18 \times 10^{-18}\ \text{J})\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

R_H is the Rydberg constant, $1.097 \times 10^7\ \text{m}^{-1}$,

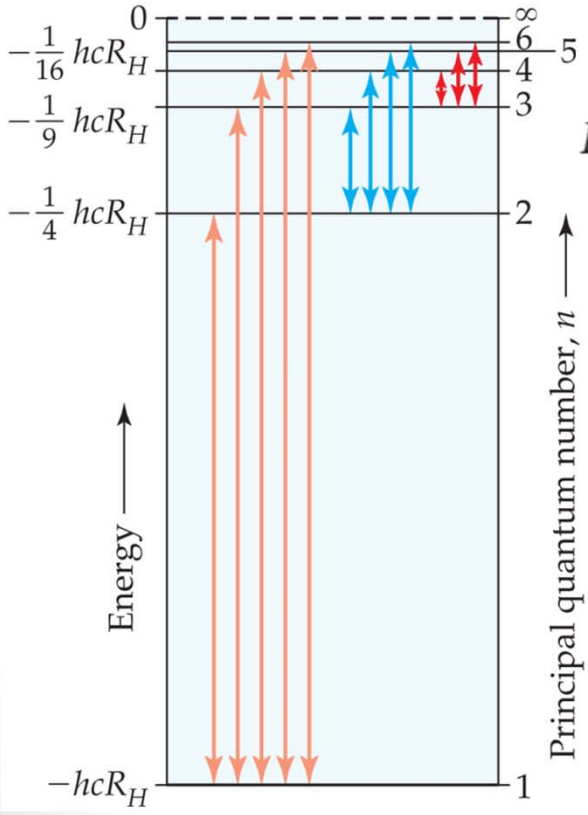
$\Delta E > 0\ (n_f > n_i)$: Photon *absorbed* with $E_{photon} = h\nu = \Delta E$

$\Delta E < 0\ (n_f < n_i)$: Photon *emitted* with $E_{photon} = h\nu = -\Delta E$

$$E_{photon} = \frac{hc}{\lambda} = -\Delta E = hcR_H\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right) \text{(for emission)}$$

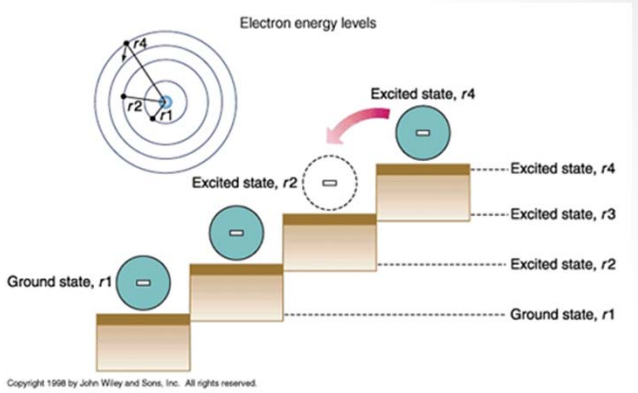
Exercise 6.4

- Using Figure 6.12, predict which of these electronic transitions produces the spectral line having the longest wavelength: n = 2 to n = 1, n = 3 to n = 2, or n = 4 to n = 3.



$$E_{\text{photon}} = \frac{hc}{\lambda} = -\Delta E = hcR_{\text{H}}\left(\frac{1}{n_{\text{f}}^2} - \frac{1}{n_{\text{i}}^2}\right) \text{(for emission)}$$

Bohr Model



Limitations

- The Bohr model explains the line spectrum of the hydrogen atom, it cannot explain the spectra of other atoms!
- Classical physics would result in an electron falling into the positively charged nucleus. Bohr simply assumed it would not!
- Circular motion is not wave-like in nature.

Important Ideas from the Bohr Model

- Electrons exist only in certain discrete energy levels, which are described by quantum numbers.
- Energy is involved in the transition of an electron from one level to another.

The Wave Behavior of Matter

- Louis de Broglie theorized that if light can have material properties, matter should exhibit wave properties.
- He demonstrated that the relationship between mass and wavelength was

$$\lambda = \frac{h}{mv}$$

- where h is the Planck constant.
- The quantity mv for any object is called its momentum (動量).

Exercise 6.5

- What is the wavelength of an electron moving with a speed of $5.97 * 10^6$ m/s? The mass of the electron is $9.11 * 10^{-31}$ kg.

Sol:

$h = 6.626 * 10^{-34}$ J-s

$$\lambda = \frac{h}{mv}$$

$\lambda = (6.626 * 10^{-34}\ \text{J-s})/[(9.11 * 10^{-31}\ \text{kg})(5.97 * 10^6\ \text{m/s})]$
$= 1.22 * 10^{-10}\ \text{m} = 0.122\ \text{nm} = 1.22\ \text{Å}$

Note: 1 J = 1 kg-m^2/s^2

The Uncertainty Principle

- Werner Heisenberg proposed that the dual nature of matter places a fundamental limitation on how precisely we can know both the location and the momentum of an object at a given instant.
  - Heisenberg's principle is called the uncertainty principle.

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

Δx: uncertainty in position

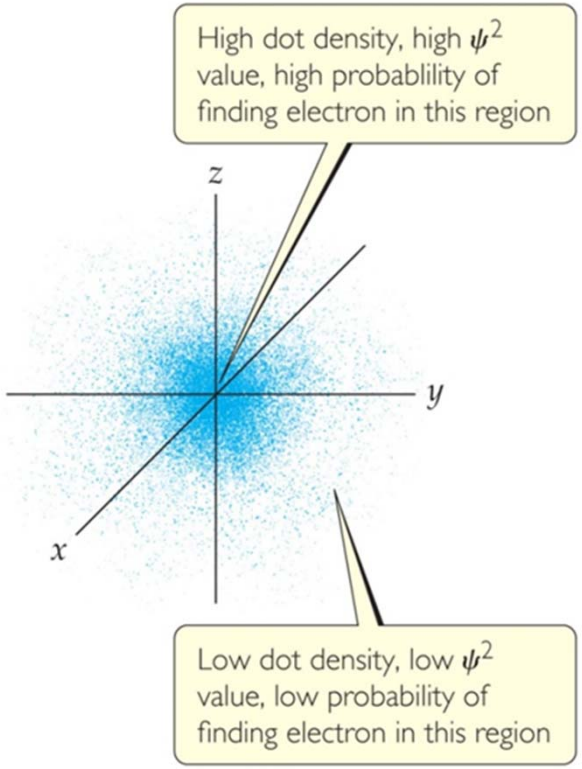
Δ(mv): uncertainty in momentum

Quantum Mechanics and Atomic Orbitals

- Solving Schrödinger's equation for the hydrogen atom leads to a series of mathematical functions called wave functions that describe the electron in the atom.
- Schrödinger's wave equation, that incorporates both the wave-like and particle-like behaviors of the electron.
  - This is known as quantum mechanics (量子力學)
- These wave functions are usually represented by the symbol ψ (lowercase Greek letter psi)

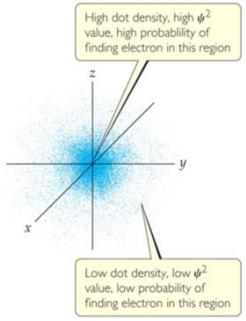
The square of the wave equation

- The square of the wave equation, ψ^2, gives the electron density, or probability of where an electron is likely to be at any given time. ψ^2 is called either the probability density or the electron density.



Orbitals and Quantum Numbers

- The solution to Schrödinger's equation for the hydrogen atom yields a set of wave functions called orbitals.
  - Each orbital has a characteristic shape and energy.
  - The lowest-energy orbital in the hydrogen atom has the spherical shape, and an energy of $-2.18 * 10^{-18}$ J.
  - Quantum mechanical Model, which describes electrons in terms of probabilities, visualized as electron clouds
  - The Bohr model introduced a single quantum number (principal quantum number), n, to describe an orbit.
  - The quantum mechanical model uses three quantum numbers, n, ℓ, and m_ℓ, which result naturally from the mathematics used to describe an orbital.



The principal quantum number

- The principal quantum number, n, can have positive integral values 1, 2, 3, ……..
- As n increases, the orbital becomes larger.
- An increase in n also means that the electron has a higher energy and is therefore less tightly bound to the nucleus.
- For the hydrogen atom, $E_n = -(2.18 * 10^{-18} J)(1/n^2)$, as in the Bohr model.

$$E = (-hcR_H)\left(\frac{1}{n^2}\right) = (-2.18 \times 10^{-18}\,\text{J})\left(\frac{1}{n^2}\right)$$

The second quantum number

- The second quantum number—the angular momentum quantum number (角量子數), ℓ — can have integral values from 0 to (n – 1) for each value of n.
- The angular momentum quantum number defines the shape of the orbital.
- The value of ℓ for a particular orbital is generally designated by the letters s, p, d, and f, corresponding to ℓ values of 0, 1, 2, and 3

Value of l	0	1	2	3
Letter used	s	p	d	f

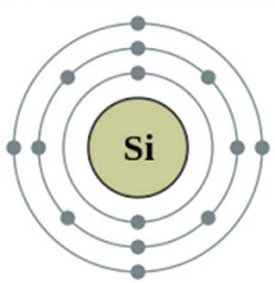
The magnetic quantum number

- The magnetic quantum number (磁量子數), m_ℓ, can have integral values between $-\ell$ and ℓ, including zero.
- The magnetic quantum number describes the orientation of the orbital in space.

Table 6.2 Relationship among Values of n, l, and m_l through $n = 4$

n	Possible Values of l	Subshell Designation	Possible Values of m_l	Number of Orbitals in Subshell	Total Number of Orbitals in Shell
1	0	1*s*	0	1	1
2	0	2*s*	0	1	
	1	2*p*	1, 0, −1	3	4
3	0	3*s*	0	1	
	1	3*p*	1, 0, −1	3	
	2	3*d*	2, 1, 0, −1, −2	5	9
4	0	4*s*	0	1	
	1	4*p*	1, 0, −1	3	
	2	4*d*	2, 1, 0, −1, −2	5	
	3	4*f*	3, 2, 1, 0, −1, −2, −3	7	16

Electron shell and subshell



- The collection of orbitals with the same value of n is called an electron shell.
  - For example: all the orbitals that have n = 3, are said to be in the third shell.
- The set of orbitals that have the same n and ℓ values is called a subshell. Each subshell is designated by a number (the value of n) and a letter (s, p, d, or f, corresponding to the value of ℓ).
  - For example, the orbitals that have n = 3 and ℓ = 2 are called 3d orbitals and are in the 3d subshell.

Table 6.2 Relationship among Values of n, l, and m_l through $n = 4$

n	Possible Values of l	Subshell Designation	Possible Values of m_l	Number of Orbitals in Subshell	Total Number of Orbitals in Shell
3	0	$3s$	0	1	
	1	$3p$	$1, 0, -1$	3	
	2	$3d$	$2, 1, 0, -1, -2$	5	9

Exercise 6.6

- (a) Without referring to Table 6.2, predict the number of subshells in the fourth shell, that is, for n = 4. (b) Give the label for each of these subshells.

Sol:

(a) There are four subshells in the fourth shell, corresponding to the four possible values of ℓ (0, 1, 2, and 3). These subshells are labeled 4s, 4p, 4d, and 4f.

(b) The number given in the designation of a subshell is the principal quantum number, n; the letter designates the value of the angular momentum quantum number, ℓ : for $\ell = 0$, s; for $\ell = 1$, p; for $\ell = 2$, d; for $\ell = 3$, f.

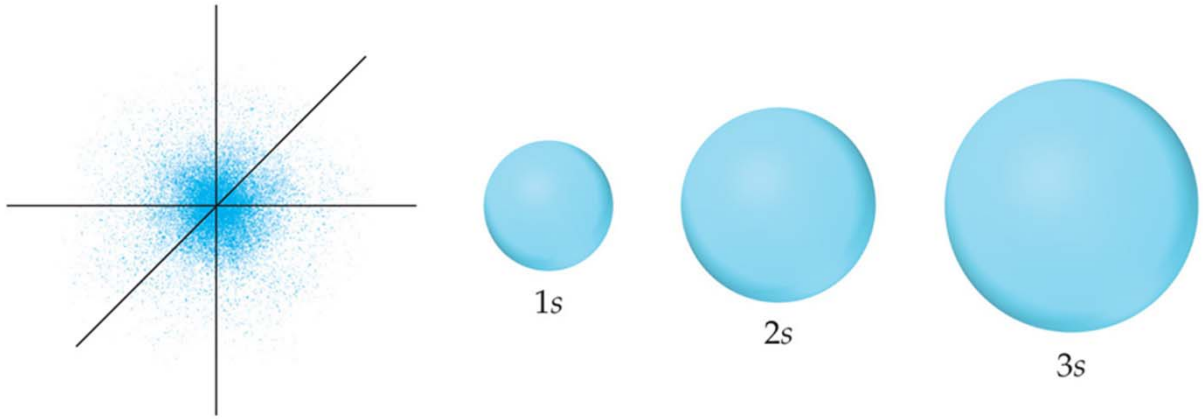
- (c) How many orbitals are in each of these subshells?

Sol:

- There is one 4s orbital (when $\ell = 0$, there is only one possible value of m_ℓ : 0).
- There are three 4p orbitals (when $\ell = 1$, there are three possible values of m_ℓ: 1, 0, -1).
- There are five 4d orbitals (when $\ell = 2$, there are five allowed values of m_ℓ: 2, 1, 0, -1, -2).
- There are seven 4f orbitals (when $\ell = 3$, there are seven permitted values of m_ℓ: 3, 2, 1, 0, -1, -2, -3).

The s Orbitals

- The value of ℓ for s orbitals is 0.
- They are spherical in shape.
- The radius of the sphere increases with the value of n.

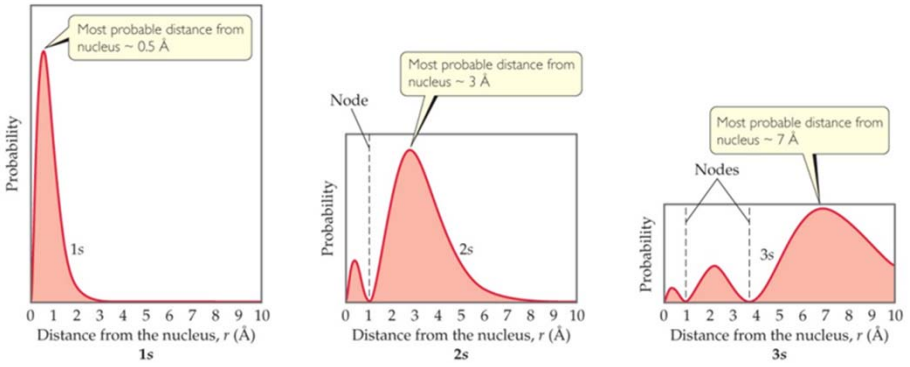


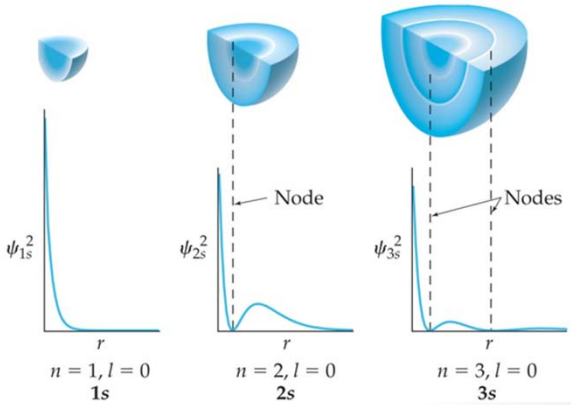
(a) An electron density model **(b)** Contour models

Radial probability function

- For an ns orbital, the number of peaks is n.
- For an ns orbital, the number of nodes (where there is zero probability of finding an electron) is n–1.
- As n increases, the electron density becomes more spread out, that is, there is a greater probability of finding the electron further from the nucleus.

特定半徑發現電子的機率





Radial probability function

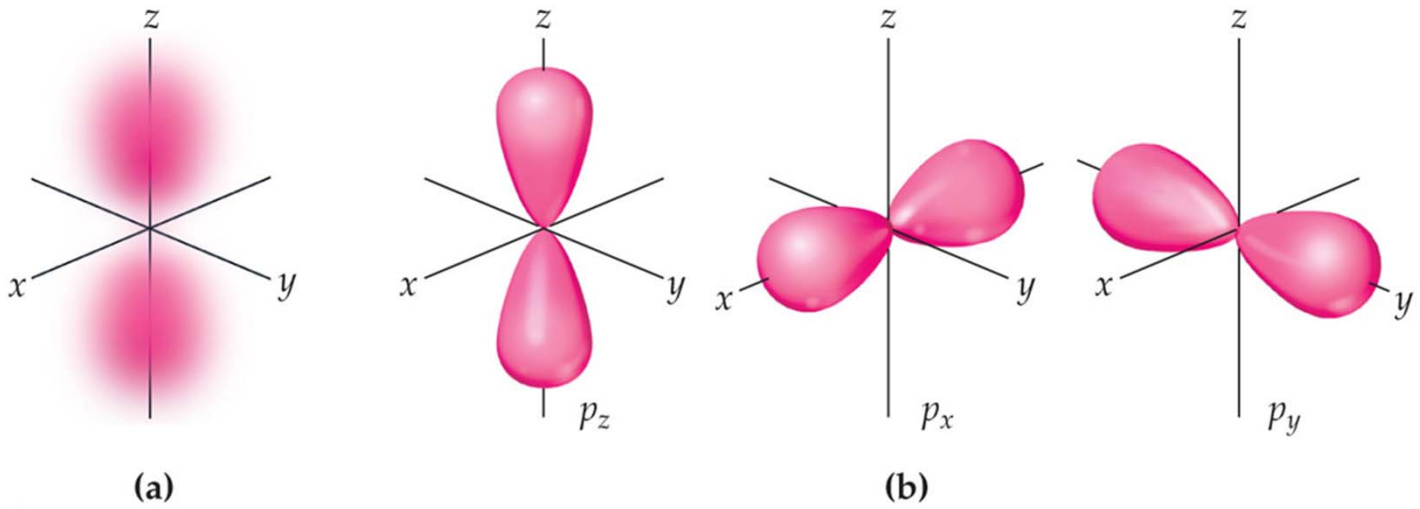
$4\pi r^2[\psi(r)]^2$

probability density

$[\psi(r)]^2$

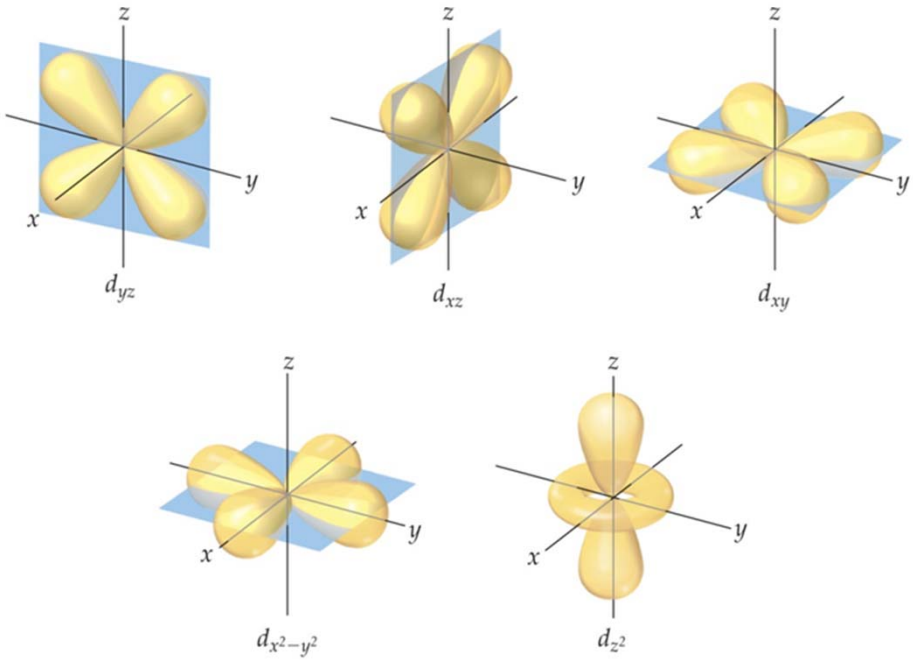
The p Orbitals

- The value of ℓ for p orbitals is 1.
- They have two lobes with a node between them.



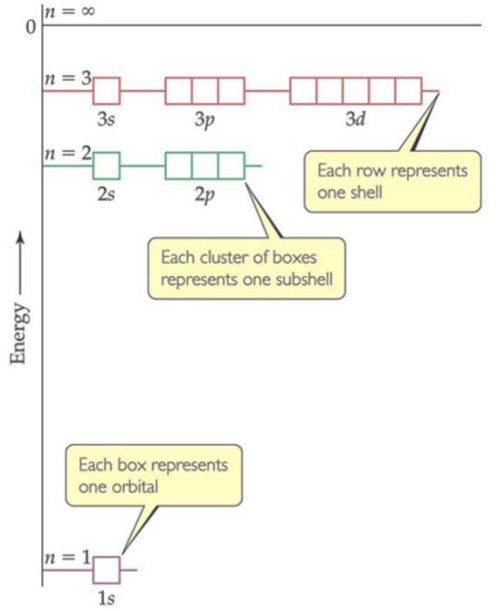
The d Orbitals

- The value of ℓ for a d orbital is 2.
- Four of the five d orbitals have four lobes; the other resembles a p orbital with a doughnut around the center.



Energies of Orbitals—Hydrogen

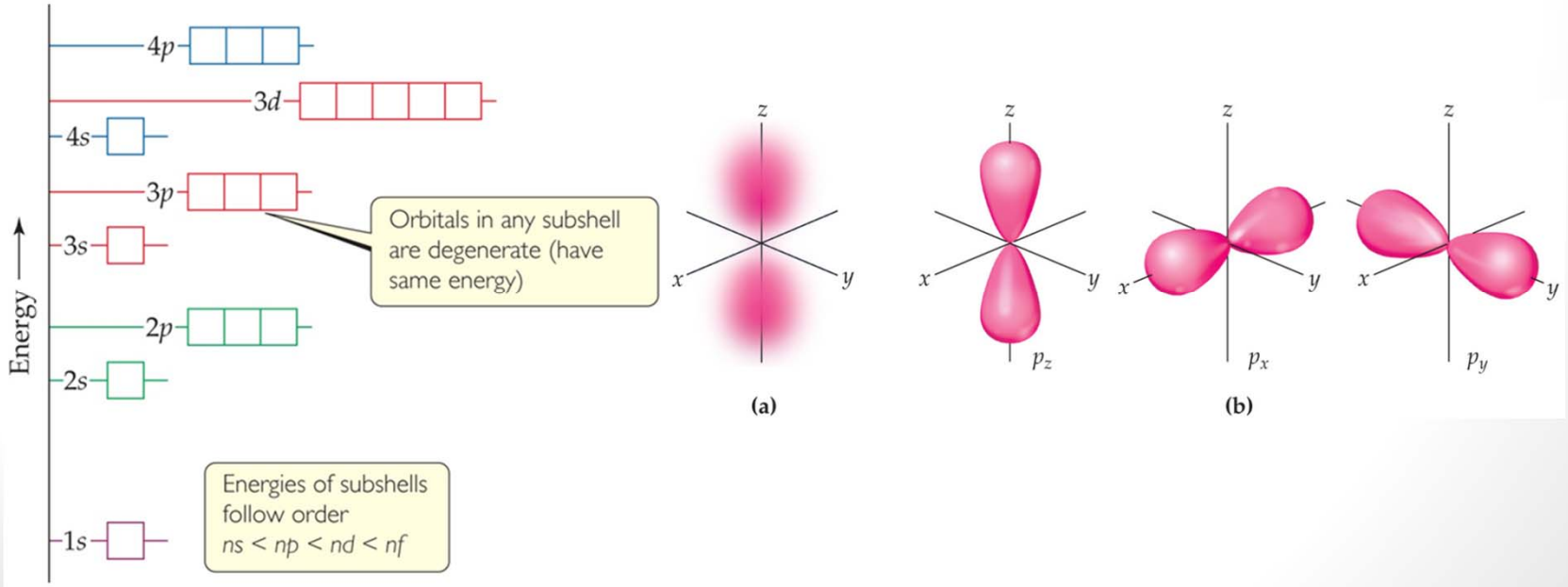
- For a one-electron hydrogen atom, orbitals on the same energy level have the same energy.
  - In a hydrogen atom the 3s, 3p, and 3d subshells all have the same energy.
  - Orbitals with the same energy are said to be degenerate.



$n = 1$ shell has one orbital
$n = 2$ shell has two subshells composed of four orbitals
$n = 3$ shell has three subshells composed of nine orbitals

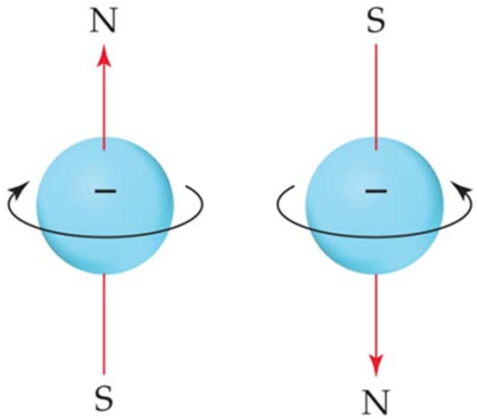
Energies of Orbitals—many-electron atom

- However, in a many-electron atom, the energies of the various subshells in a given shell are different because of electron–electron repulsions.
- Orbitals in any subshell are degenerate (have same energy)
- Energies of subshells follow order ns < np < nd < nf



Electron Spin

- It was discovered that two electrons in the same orbital do not have exactly the same energy.
- The "spin" of an electron describes its magnetic field, which affects its energy.
- The spin magnetic quantum number 自旋磁量子數), is denoted m_s (the subscript s stands for spin). Two possible values are allowed for m_s, + ½ or - ½ , which were first interpreted as indicating the two opposite directions in which the electron can spi



Pauli exclusion principle

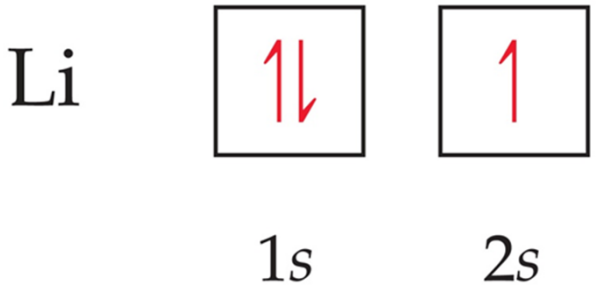
- No two electrons in the same atom can have exactly the same energy.
  - No two electrons in the same atom can have identical sets of quantum numbers.
  - This means that every electron in an atom must differ by at least one of the four quantum number values: n, ℓ, m_ℓ, and m_s.

Electron Configurations

- The way electrons are distributed in an atom is called its electron configuration (電子組態).
- The most stable organization is the lowest possible energy, called the ground state.
- Each component consists of
  - a number denoting the energy level.
  - a letter denoting the type of orbital.
  - a superscript denoting the number of electrons in those orbitals.
- EX: Boron, atomic number 5, has the electron configuration $1s^2 2s^2 2p^1$. The fifth electron must be placed in a 2p orbital because the 2s orbital is filled.

Orbital Diagrams

- Each box in the diagram represents one orbital.
- Half-arrows represent the electrons.
- The direction of the arrow represents the relative spin of the electron.
  - A half arrow pointing up (↑) represents an electron with a positive spin magnetic quantum number ($m_s = +1/2$)
  - A half arrow pointing down (↓) represents an electron with a negative spin magnetic quantum number ($m_s = -1/2$)



Hund's Rule

- "For degenerate orbitals, the lowest energy is attained when the number of electrons with the same spin is maximized."
  - This means that, for a set of orbitals in the same sublevel, there must be one electron in each orbital before pairing and the electrons have the same spin, as much as possible.

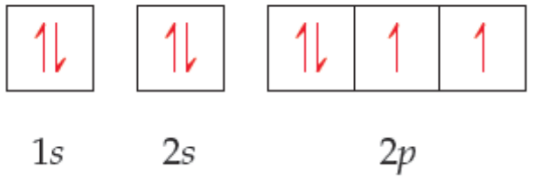
Table 6.3 Electron Configurations of Several Lighter Elements

Element	Total Electrons	Orbital Diagram					Electron Configuration
		1s	2s	2p		3s	
Li	3	↑↓	↑				$1s^2 2s^1$
Be	4	↑↓	↑↓				$1s^2 2s^2$
B	5	↑↓	↑↓	↑			$1s^2 2s^2 2p^1$
C	6	↑↓	↑↓	↑	↑		$1s^2 2s^2 2p^2$
N	7	↑↓	↑↓	↑	↑	↑	$1s^2 2s^2 2p^3$
Ne	10	↑↓	↑↓	↑↓	↑↓	↑↓	$1s^2 2s^2 2p^6$
Na	11	↑↓	↑↓	↑↓	↑↓	↑↓	$1s^2 2s^2 2p^6 3s^1$

Exercise 6.7

- Draw the orbital diagram for the electron configuration of oxygen, atomic number 8. How many unpaired electrons does an oxygen atom possess?

Sol: Two electrons each go into the 1s and 2s orbitals with their spins paired. This leaves four electrons for the three degenerate 2p orbitals. Following Hund's rule, we put one electron into each 2p orbital until all three orbitals have one electron each. The fourth electron is then paired up with one of the three electrons already in a 2p orbital, so that the orbital diagram is



The corresponding electron configuration is written $1s^2 2s^2 2p^4$. The atom has two unpaired electrons.

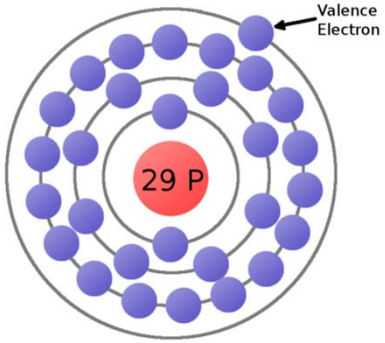
Condensed Electron Configurations

- Condensed electron configuration of an element: the electron configuration of the nearest noble-gas element of lower atomic number is represented by its chemical symbol in brackets.

Ex:

lithium → Li: $[He]2s^1$

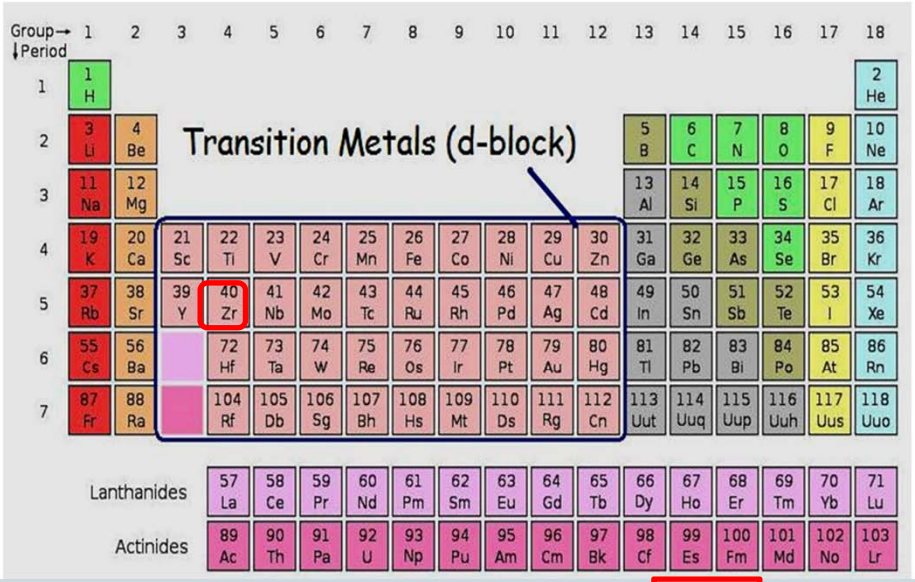
- The Inner-shell electrons are referred to as the core electrons (內層電子).
- The outer-shell electrons are called the valence electrons (價電子).



1A
3 **Li** $[He]2s^1$
11 **Na** $[Ne]3s^1$
19 **K** $[Ar]4s^1$
37 **Rb** $[Kr]5s^1$
55 **Cs** $[Xe]6s^1$
87 **Fr** $[Rn]7s^1$
Alkali metals

Transition Metals

- In writing the electron configurations of the transition elements, we fill orbitals in accordance with Hund's rule
  - we add them to the 3d orbitals singly until all five orbitals have one electron each and then place additional electrons in the 3d orbitals with spin pairing until the shell is completely filled



Zr has four more electrons than Kr.

$[Kr]\ 5s^2 4d^2$

	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
Atomic number	21	22	23	24	25	26	27	28	29	30
Electron config.[a]	$3d^1 4s^2$	$3d^2 4s^2$	$3d^3 4s^2$	$3d^5 4s^1$	$3d^5 4s^2$	$3d^6 4s^2$	$3d^7 4s^2$	$3d^8 4s^2$	$3d^{10} 4s^1$	$3d^{10} 4s^2$

Anomalous Electron Configurations

- Some irregularities occur when there are enough electrons to half-fill s and d orbitals on a given row.

EX:

Chromium is [Ar] $3d^5 4s^1$

rather than the expected [Ar] $3d^4 4s^2$.

Copper (element 29) is [Ar]$3d^{10} 4s^1$

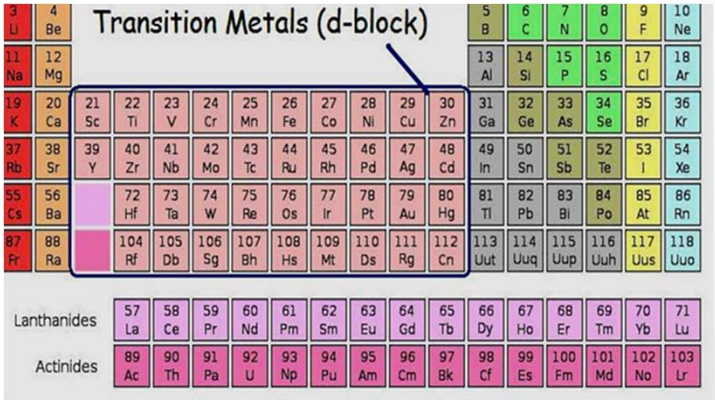
rather than the expected [Ar]$3d^9 4s^2$.

- This occurs because the 4s and 3d orbitals are very close in energy.
- These anomalies occur in f-block atoms with f and d orbitals, as well.

The Lanthanides

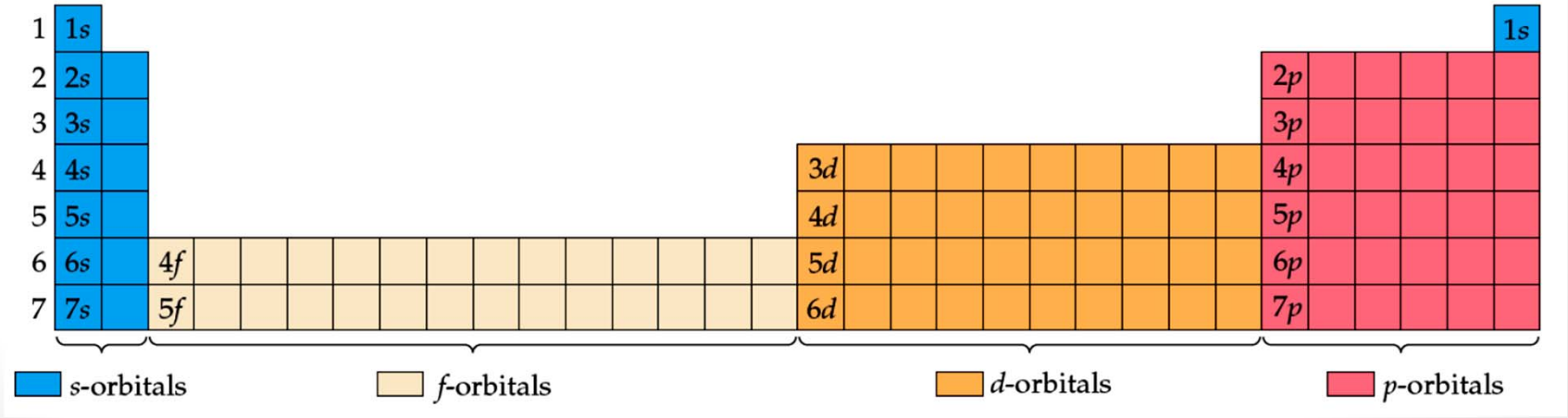
- The sixth row of the periodic table begins by filling the 6s orbitals.
  - There are seven degenerate 4f orbitals, corresponding to the seven allowed values of m_ℓ, ranging from 3 to -3. Thus, it takes 14 electrons to fill the 4f orbitals completely.
  - The 14 elements corresponding to the filling of the 4f orbitals are known as either the lanthanide elements or the rare earth elements.
  - the energies of the 4f and 5d orbitals are very close to each other, the electron configurations of some of the lanthanides involve 5d electrons.

$[Xe]6s^25d^1$	$[Xe]6s^25d^14f^1$	$[Xe]6s^24f^3$
Lanthanum	Cerium	Praseodymium



The actinide elements

- The final row of the periodic table begins by filling the 7s orbitals.
  - The actinide elements are then built up by completing the 5f orbitals



Exercise 6.8

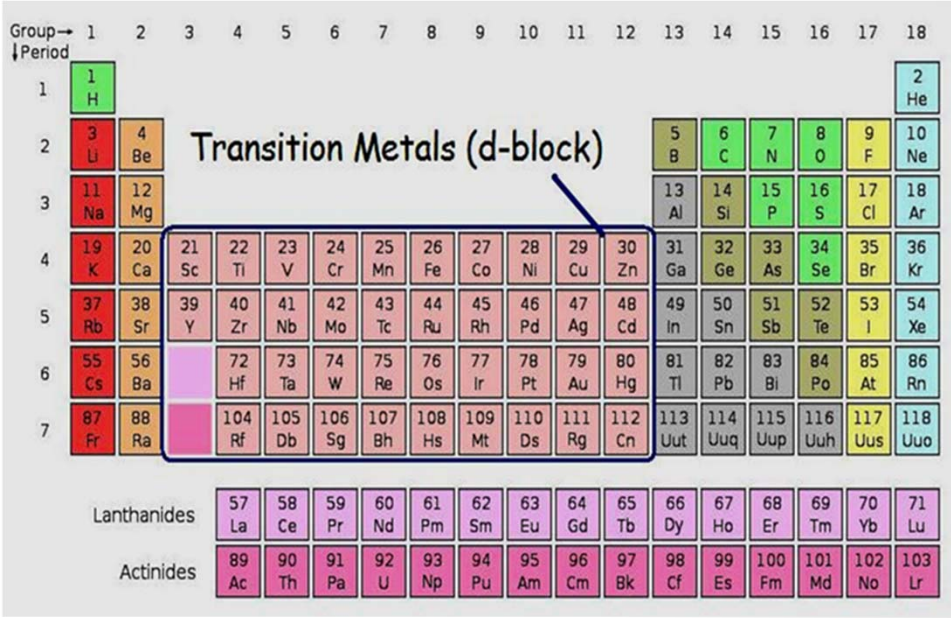
- What is the characteristic valence electron configuration of the group 7A elements, the halogens?

Sol:

F: $[He]2s^2 2p^5$

Cl: $[Ne]3s^2 3p^5$

The characteristic valence electron configuration of a halogen is ns^2np^5



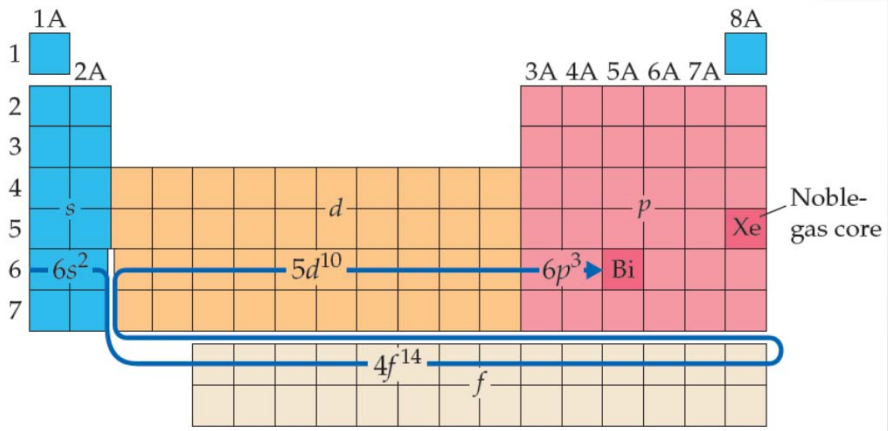
Exercise 6.9

- (a) Based on its position in the periodic table, write the condensed electron configuration for bismuth, element 83. (b) How many unpaired electrons does a bismuth atom have?

Sol:

(a) Bi → $[Xe]6s^2 4f^{14} 5d^{10} 6p^3$.

(b) the only partially occupied subshell is 6p. The orbital diagram representation for this subshell is



Outer-shell electron configurations of the elements.