Chapter 6 Electronic Structure of Atoms

6.1 The Wave Nature of Light



- To understand the electronic structure of atoms, one must understand the nature of electromagnetic radiation.
- The distance between corresponding points on adjacent waves is the **wavelength** (λ).

Waves



- The number of waves passing a given point per unit of time is the **frequency** (v).
- For waves traveling at the same velocity, the longer the wavelength, the smaller the frequency.
- If the time associated with the lines to the left is one second, then the frequencies would be 2 s⁻¹ and 4 s⁻¹, respectively.

Waves



If wave (a) has a wavelength of 2.0 m and a frequency of 1.5 x 10⁸ cycles/s, what are the wavelength and frequency of wave (b)? Our bodies are penetrated by X rays but not by visible light. Is this because X rays travel faster than visible light?

- a. Yes: X-rays travel at at faster speeds than visible light.
- b. No: both X-rays and visible light travel at the speed of light, *c*.

Electromagnetic Radiation



All electromagnetic radiation travels at the same velocity: The speed of light (c) is 3.00 × 10⁸ m/s.

 $c = \lambda v$

Electromagnetic Radiation



Is the wavelength of a microwave longer or shorter than the wavelength of visible light? By how many orders of magnitude do the two waves differ in wavelength?

Two electromagnetic waves are represented below.

(a) Which wave has the higher frequency? (b) If one wave represents visible light and the other represents infrared radiation, which wave is which?
(c) If one of the waves in the margin represents blue light and the other red light, which is which?

Sample Exercise 6.2

1)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?

Practice Exercise 1

Consider the following three statements: (i) For any electromagnetic radiation, the product of the wavelength and the frequency is a constant. (ii) If a source of light has a wavelength of 3.0 Å, its frequency is 1.0×10^{18} Hz. (iii) The speed of ultraviolet light is greater than the speed of microwave radiation. Which of these three statements is or are true?

- (a) Only one statement is true.
- (b) Statements (i) and (ii) are true.
- (c) Statements (i) and (iii) are true.
- (d) Statements (ii) and (iii) are true.
- (e) All three statements are true.

Practice Exercise

- 1)A laser used in eye surgery to fuse
 - detached retinas produces radiation with a wavelength of 640.0 nm. Calculate the frequency of this radiation.
- 2)An FM radio station broadcasts electromagnetic radiation at a frequency of 103.4 MHz (megahertz). Calculate the wavelength of this radiation.

6.2 Quantized Energy and Photons

The Nature of Energy



The wave nature of light does not explain how an object can glow when its temperature increases.

The Nature of Energy



Which area in the photograph corresponds to the highest temperature?

The Nature of Energy—Quanta



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



Potential energy of person walking up steps increases in stepwise, quantized manner Max Planck explained it by assuming that energy comes in packets called quanta (singular: quantum).

Consider the notes that can be played on a piano. In what way is a piano an example of a quantized system?

- a. Multiple notes can be played at once.
- b. Only certain notes exist with none in between.

Consider the notes that can be played on a piano. In this analogy, would a violin be continuous or quantized?

- a. The violin would be continuous.
- b. The violin would be quantized.

The Photoelectric Effect



Einstein used quanta to explain the photoelectric effect.
Each metal has a different

energy at which it ejects electrons. At lower energy, electrons are not emitted.

He concluded that energy is proportional to frequency:

 $E = h_V$

where *h* is Planck's constant, 6.626×10^{-34} J·s.

The Photoelectric Effect



What is the source of energy that causes electrons to be emitted from the surface?

In Figure 6.7, will the kinetic energy of an emitted electron equal the energy of the photon that causes its emission?

- a. Yes
- b. No
- c. Can't tell from information given



Sample Exercise 6.3 (a) Calculate the energy of one photon of yellow light with a wavelength of 589 nm.

Practice Exercise 1

Which of the following expressions correctly gives the energy of a mole of photons with wavelength λ ?

(a)
$$E = \frac{h}{\lambda}$$
 (b) $E = N_A \frac{\lambda}{h}$ (c) $E = \frac{hc}{\lambda}$ (d) $E = N_A \frac{h}{\lambda}$ (e) $E = N_A \frac{hc}{\lambda}$

Practice Exercise

(a) A laser emits light with a frequency of 4.69 × 10¹⁴ s⁻¹. What is the energy of one photon of the radiation from this laser?
(b) If the laser emits a pulse of energy containing 5.0 × 10¹⁷ photons of this radiation, what is the total energy of that pulse?

(c) If the laser emits 1.3×10^{-2} J of energy during a pulse, how many photons are emitted during the pulse?

Do you think that the formation of a rainbow is more a demonstration of the wave-like or particle-like behavior of light?

- a. Wave-like because the colors represent continuous wavelengths.
- b. Particle-like because the colors represent specific energies.

6.3 Line Spectra and the Bohr Model

Atomic Emissions

Another mystery in the early twentieth century involved the emission spectra observed from energy emitted by atoms and molecules.



Neon (Ne)



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.





The Hydrogen Spectrum



- Johann Balmer (1885) discovered a simple formula relating the four lines to integers.
- Johannes Rydberg advanced this formula. $\frac{1}{\lambda} = (R_{\rm H}) \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$

Neils Bohr explained why this mathematical relationship works.

The Bohr Model



Niels Bohr adopted Planck's assumption and explained these phenomena in this way:

 Electrons in an atom can only occupy certain orbits (corresponding to certain energies).

The Bohr Model



2.Electrons in permitted orbits have specific, "allowed" energies; these energies will not be radiated from the atom.

3.Energy is only absorbed or emitted in such a way as to move an electron from one "allowed" energy state to another; the energy is defined by

$$E = h_{\rm V}$$

With reference to Figure 6.6, in what way is the Bohr model for the H atom more like steps than a ramp?

- a. If the climber slips, he or she can fall to the base of the staircase.
- b. Only certain stairs are available.
- c. Two stairs can be occupied simultaneously.



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



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

The Bohr Model



The energy absorbed or emitted from the process of electron promotion or demotion can be calculated by the equation

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

where R_H is the Rydberg constant, 1.097×10^7 m⁻¹, and n_i and n_f are the initial and final energy levels of the electron. Why does it make sense that an orbit with a larger radius has a higher energy than one with a smaller radius?

- a. The distance from the nucleus is greater, so the attraction is less.
- b. The distance from the nucleus is less, so the attraction is less.
- c. The distance from the nucleus is greater, so the attraction is greater.
- d. The distance from the nucleus is less, so the attraction is greater.

What is the significance of the minus sign in front of ΔE in the following equation?

$$\lambda = \frac{c}{\nu} = \frac{hc}{E_{\text{photon}}} = \frac{hc}{-\Delta E} = \frac{(6.626 \times 10^{-34} \,\text{J}\text{-s})(2.998 \times 10^8 \,\text{m/s})}{+1.94 \times 10^{-18} \,\text{J}} = 1.02 \times 10^{-7} \,\text{m}$$

- a. The photon energy is absorbed by the electron transition.
- b. The photon energy is released by the electron transition.

Which transition will lead to the emission of light with longer wavelength, n = 3 to n = 2, or n = 4 to n = 3?


What is the relationship between $1/\lambda$ and ΔE for a transition of the electron from a lower value of *n* to a higher one?

- a. $1/\lambda = \Delta E/hc$ b. $1/\lambda = -\Delta E/hc$ c. $1/\lambda = hc/\Delta E$
- d. $1/\lambda = hc/-\Delta E$

Sample Exercise 6.4 Electronic Transitions in the Hydrogen Atom



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

Practice Exercise 1

In the top part of Figure 6.11, the four lines in the H atom spectrum are due to transitions from a level for which

 $n_i > 2$ to the $n_f = 2$ level. What is the value of n_i for the blue-green line in the spectrum?





Practice Exercise 2

For each of the following transitions, give the sign of ΔE and indicate whether a photon is emitted or absorbed:

Limitations of the Bohr Model

- It only works for hydrogen!
- 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

- Points that are incorporated into the current atomic model include the following:
- 1) Electrons exist only in certain discrete energy levels.
- 2) Energy is involved in the transition of an electron from one level to another.

6.4 The Wave Behavior of Matter

The Wave Nature 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

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

The wave nature of light is used to produce this electron micrograph. 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.

Practice Exercise 1

Consider the following three moving objects: (i) a golf ball with a mass of 45.9 g moving at a speed of 50.0 m/s, (ii) An electron moving at a speed of 3.50 \times 10⁵ m/s, (iii) A neutron moving at a speed of 2.3×10^2 m/s. List the three objects in order from shortest to longest de Broglie wavelength. (a) i < iii < ii (b) ii < iii < i (c) iii < ii < i

(d) i < ii < iii (e) iii < i < ii

Practice Exercise 2

Calculate the velocity of a neutron whose de Broglie wavelength is 505 pm. The mass of a neutron is $1.6749 \times 10^{-27} \text{ kg}$

A baseball pitcher throws a fastball that moves at 95 miles per hour. Does that moving baseball generate matter waves? If so, can we observe them?

- a. Yes, but they are not observable because the matter wavelength is too small for our eyes to detect.
- b. Yes and they are observable because matter waves of macroscopic objects are in the visible region of light.
- c. Matter waves are not produced for macroscopic objects and thus none are observable.

The Uncertainty Principle

Heisenberg showed that the more precisely the momentum of a particle is known, the less precisely is its position is known:

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



What is the principal reason we must consider the uncertainty principle when discussing electrons and other subatomic particles but not when discussing our macroscopic world?

- a. The charge of subatomic particles compared to zero charge for macroscopic particles
- b. The small mass and size of subatomic particles compared to the mass and size of macroscopic particles
- c. The small volume occupied by subatomic particles compared to the volume occupied by macroscopic particles
- d. The slower speeds of subatomic particles compared to the speeds of macroscopic particles

6.5 Quantum Mechanics and Atomic Orbitals

Quantum Mechanics

Erwin Schrödinger developed a mathematical treatment into which both the wave and particle nature of matter could be incorporated.

This is known as quantum mechanics.



Quantum Mechanics

- The solution of Schrödinger's wave equation is designated with a lowercase Greek psi (\u03c6).
- The square of the wave equation, \u03c6², gives the electron density, or probability of where an electron is likely to be at any given time.



Quantum Mechanics

Where in the figure is the region of highest electron density?



What is the difference between stating "The electron is located at a particular point in space" and "There is a high probability that the electron is located at a particular point in space"?

- In the first statement, knowing the position of the electron enables us to know its momentum whereas in the second statement, the momentum is already known.
- Both statements suggest that the momentum and position are known.
- c. The first statement presents a known position whereas the second statement says we don't know exactly the position.

Quantum Numbers

- Solving the wave equation gives a set of wave functions, or orbitals, and their corresponding energies.
- Each orbital describes a spatial distribution of electron density.
- An orbital is described by a set of three quantum numbers.

Principal Quantum Number (n)

- The principal quantum number, n, describes the energy level on which the orbital resides.
- The values of *n* are integers \geq 1.
- These correspond to the values in the Bohr model.

Angular Momentum Quantum Number (/)

- This quantum number defines the shape of the orbital.
- Allowed values of *I* are integers ranging from 0 to *n* – 1.

We use letter designations to communicate the different values of / and, therefore, the shapes and types of orbitals.

Angular Momentum Quantum Number (/)

| Value of <i>l</i> | 0 | 1 | 2 | 3 |
|-------------------|---|---|---|---|
| Letter used | S | р | d | f |

Magnetic Quantum Number (m_l)

- The magnetic quantum number describes the three-dimensional orientation of the orbital.
- Allowed values of m_l are integers ranging from -/ to /:

 $-l \leq m_l \leq l$

Therefore, on any given energy level, there can be up to 1 s orbital, 3 p orbitals, 5 d orbitals, 7 f orbitals, and so forth. What is the difference between an *orbit* in the Bohr model of the hydrogen atom and an *orbital* in the quantum mechanical model?

- a. An orbital is composed of some integral number of orbits.
- b. An orbit is a well-defined circular path around the nucleus while an orbital is a wave function that gives the probability of finding the electron at any point in space.
- c. An orbit is a well-defined circular path around the nucleus while an orbital is the object (electron) that is moving around the nucleus.
- d. There is no difference between the definitions of the terms "orbit" and "orbital." They simply were proposed by different scientists.

If the fourth shell (the n = 4energy level) were shown, how many subshells would it contain? How would they be labeled?



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

Magnetic Quantum Number (*m*_l)

Orbitals with the same value of n form an electron shell.

Different orbital types within a shell are **subshells**.

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 | 1s | 0 | 1 | 1 |
| 2 | 0 | 2 <i>s</i> | 0 | 1 | |
| | 1 | 2 <i>p</i> | 1, 0, -1 | 3 | 4 |
| 3 | 0 | 3s | 0 | 1 | |
| | 1 | 3 <i>p</i> | 1, 0, -1 | 3 | |
| | 2 | 3 <i>d</i> | 2, 1, 0, -1, -2 | 5 | 9 |
| 4 | 0 | 4 <i>s</i> | 0 | 1 | |
| | 1 | 4 <i>p</i> | 1, 0, -1 | 3 | |
| | 2 | 4d | 2, 1, 0, -1, -2 | 5 | |
| | 3 | 4f | 3, 2, 1, 0, -1, -2, -3 | 7 | 16 |

6.6 Representation of Orbitals

s Orbitals



The value of *I* for *s* orbitals is 0.

They are spherical in shape.

The radius of the sphere increases with the value of n.

s Orbitals





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 is more spread out and there is a greater probability of finding an electron further from the nucleus.



How many maxima would you expect to find in the radial probability function for the 4s orbital of the hydrogen atom? How many nodes would you expect in this function?

p Orbitals

The value of *I* for *p* orbitals is 1.

They have two lobes with a node between them.



p Orbitals

Note on the left that the color is deep pink in the interior of each lobe bud fades to pale pink at the edges. What does this change in color represent? What label is applied to the 2p orbital aligned along the x axis?



d Orbitals



The value of *I* 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.

d Orbitals Which of the *d* orbitals most resembles a *p_z* orbital?





f Orbitals Very complicated shapes Seven equivalent orbitals in a sublevel I = 3 i i i i


6.7 Many-Electron Atoms

Energies of Orbitals—Hydrogen

For a oneelectron hydrogen atom, orbitals on the same energy level have the same energy. Chemists call them **degenerate** orbitals.



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 Atoms



- As the number of electrons increases, so does the repulsion between them.
- Therefore, in atoms with more than one electron, not all orbitals on the same energy level are degenerate.
- Orbital sets in the same sublevel are still degenerate.
- Energy levels start to overlap in energy (e.g., 4s is lower in energy than 3d.)

Energies of Orbitals— Many-electron Atoms



Not all of the orbitals in the n = 4shell are shown in this figure. Which subshells are missing? In a many-electron atom, can we predict unambiguously whether the 4*s* orbital is lower or higher in energy than the 3*d* orbitals?

- a. Yes, the figure shows that 4s is lower than 3d.
- b. Yes, the *n* quantum number 4 is higher in energy than the quantum number 3.
- c. No, because this will be different depending on the identity of the atom.

Spin Quantum Number, ms

- In the 1920s, 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 quantum number has only two allowed values, +½ and -½.



Spin Quantum Number, ms

From this figure, why are there only two possible values for the spin quantum number?



Pauli Exclusion Principle

- No two electrons in the same atom can have exactly the same energy.
- Therefore, 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.

6.8 Electron Configurations

Electron Configurations

- The way electrons are distributed in an atom is called its electron configuration.
- The most stable organization is the lowest possible energy, called the ground state.
- Each component consists of
 - □a number denoting the energy level;
 - \Box a letter denoting the type of orbital;
 - □ a superscript denoting the number of electrons in those orbitals.

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.



Hund's Rule



"For degenerate orbitals, the lowest energy is attained when the number of electrons with the same spin is maximized."

Ø 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.

Paramagnetism & Diamagnetism

- Paramagnetism the substance contains one or more unpaired electrons and is drawn into a magnetic field.
- Diamagnetism the substance has no unpaired electrons and is weakly repelled by a magnetic field.

Sample Exercise 6.7 Orbital Diagrams and Electron Configurations

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

Practice Exercise 1

How many of the elements in the second row of the periodic table (Li through Ne) will have at least one unpaired electron in their electron configurations?

- **(a)** 3
- **(b)** 4
- **(c)** 5
- **(d)** 6
- **(e)** 7

Practice Exercise 2

(a) Write the electron configuration for silicon, element 14, in its ground state.(b) How many unpaired electrons does a ground-state silicon atom possess?

Condensed Electron Configurations

1A 3 Li [He]2s 11 Na [Ne]3s¹ 19 Κ $[Ar]4s^1$ 37 Rb [Kr]5s 55 Cs $[Xe]6s^{1}$ 87 Fr $[Rn]7s^1$

Alkali metals

- Elements in the same group of the periodic table have the same number of electrons in the outer most shell. These are the valence electrons.
- The filled inner shell electrons are called core electrons. These include completely filled d or f sublevels.
- We write a shortened version of an electron configuration using brackets around a noble gas symbol and listing only valence electrons.

Based on the structure of the periodic table, which becomes occupied first, the 6s orbital or the 5d orbitals?

a. 5d

b. 6s

6.9 ElectronConfigurations andthe Periodic Table

Periodic Table

We fill orbitals in increasing order of energy.

Different blocks on the periodic table correspond to different types of orbitals: s = blue, p = pink (s and p are representative elements); d = orange (transition elements); f = tan (lanthanides and actinides, or inner transition elements)



Sample Exercise 6.8 Electron Configurations for a Group What is the characteristic valence electron configuration of the group 7A elements, the halogens?

Practice Exercise

Which family of elements is characterized by an ns^2np^2 electron configuration in the outermost occupied shell?

Practice Exercise 1

A certain atom has an ns^2np^6 electron configuration in its outermost occupied shell. Which of the following elements could it be?

(a) Be
(b) Si
(c) I
(d) Kr
(e) Rb

Sample Exercise 6.9 Electron Configurations from the Periodic Table (a) Write the condensed electron configuration for bismuth, element number 83.

(b) How many unpaired electrons does each atom of bismuth have?

Practice Exercise 1

A certain atom has an [noble gas] $5s^24d^{10}5p^4$ electron configuration. Which element is it?

- **(a)** Cd
- **(b)** Te
- **(c)** Sm
- **(d)** Hg
- (e) More information is needed

Sample Exercise 6.9 Electron Configurations from the Periodic Table

Use the periodic table to write the condensed electron configurations for

(a) Co (atomic number 27)(b) In (atomic number 49)

Some Anomalies

| | 1A 1 | 1 | | | | | | | | | | | | | | | | 8A 18 |
|--------------|------------------------------------|------------------------------------|--------------------------------------|--|--|--|--|--|--|---|---|--|---|---|--|---|--|---|
| 122 | $\hat{\mathbf{H}}_{1s^1}$ | 2A 2 | | | | | | | | | | | 3A 13 | 4A 14 | 5A 15 | 6A 16 | 7A 17 | $\frac{2}{1s^2}$ |
| Core [He] | 3 Li 2s ¹ | 4 Be 2s ² | | | | | | | | | | | $5 \\ \mathbf{B} \\ 2s^2 2p^1$ | $ \begin{array}{c} 6 \\ C \\ 2s^2 2p^2 \end{array} $ | | | 9 F $2s^22p^5$ | 10 Ne 2s ² 2p |
| [Ne] | 11 Na 3s ¹ | 12 Mg 3s ² | 3B 3 | 4B 4 | 5B 5 | 6B 6 | 7B 7 | 8 | 8B 9 | 10 | 1 B | 2B 12 | $13 \\ Al \\ 3s^2 3p^1$ | $14 \\ Si \\ 3s^2 3p^2$ | | | | 18 Ar 3s ² 3p |
| [Ar] | 19 K 4s ¹ | $20 \\ Ca \\ 4s^2$ | $21 \\ Sc \\ 4s^2 3d^1$ | $\begin{array}{c} 22\\ \mathbf{Ti}\\ 4s^23d^2 \end{array}$ | $23 \\ V \\ 4s^2 3d^3$ | $24 \\ Cr \\ 4s^1 3d^5$ | 25 Mn 4s ² 3d ⁵ | $26 \\ Fe \\ 4s^2 3d^6$ | $27 \\ Co \\ 4s^2 3d^7$ | $28 \\ Ni \\ 4s^2 3d^8$ | $29 \\ Cu \\ 4s^1 3d^{10}$ | 30 Zn 4s ² 3d ¹⁰ | 31 Ga $4s^23d^{10}$ $4v^1$ | 32 Ge 4s ² 3d ¹⁰ 4p ² | 33 As 4s ² 3d ¹⁰ 4v ³ | 34 Se 4s ² 3d ¹⁰ 4p ⁴ | 35 Br $4s^23d^{10}$ $4v^5$ | 36 Kr 4s ² 3d 4p ⁶ |
| [Kr] | 37 Rb 5s ¹ | 38 Sr 5s ² | $39 \\ Y \\ 5s^2 4d^1$ | $\begin{array}{c} 40 \\ \mathbf{Zr} \\ 5s^2 4d^2 \end{array}$ | $ \begin{array}{c} 41 \\ Nb \\ 5s^2 4d^3 \end{array} $ | 42 Mo 5s ¹ 4d ⁵ | 43 Tc 5s ² 4d ⁵ | $\begin{array}{c} 44 \\ \mathbf{Ru} \\ 5\mathrm{s}^1 4d^7 \end{array}$ | $45 \\ Rh \\ 5s^1 4d^8$ | $46 \\ Pd \\ 4d^{10}$ | $47 \\ Ag \\ 5s^1 4d^{10}$ | $48 \\ Cd \\ 5s^2 4d^{10}$ | 49 In $5s^24d^{10}$ $5n^1$ | 50 Sn $5s^24d^{10}$ $5n^2$ | 51 Sb 5s ² 4d ¹⁰ 5n ³ | $52 \\ Te \\ 5s^2 4d^{10} \\ 5w^4$ | 53 I 5s ² 4d ¹⁰ 5n ⁵ | 54 Xe 5s ² 4d 5n ⁶ |
| [Xe] | 55 Cs 6s ¹ | 56 Ba 6s ² | 71 Lu $6s^24f^{14}$ $5d^1$ | 72 Hf $6s^24f^{14}$ $5d^2$ | 73 Ta $6s^24f^{14}$ $5d^3$ | $74 \\ W \\ 6s^2 4f^{14} \\ 5d^4$ | 75 Re $6s^24f^{14}$ $5d^5$ | 76 Os $6s^24f^{14}$ $5d^6$ | 77 Ir $6s^24f^{14}$ $5d^7$ | 78 Pt $6s^{1}4f^{14}$ $5d^{9}$ | 79 Au ${}^{6s^14f^{14}}_{5d^{10}}$ | $80 \\ Hg \\ 6s^2 4f^{14} \\ 5d^{10}$ | 81 Tl 6s ² 4f ¹⁴ 5d ¹⁰ 6v ¹ | 82 Pb 6s ² 4f ¹⁴ 5d ¹⁰ 6p ² | 83 Bi 6s ² 4f ¹⁴ 5d ¹⁰ 6n ³ | 84 Po 6s ² 4f ¹⁴ 5d ¹⁰ 6p ⁴ | 85 At 6s ² 4f ¹⁴ 5d ¹⁰ 6n ⁵ | 86 Rn 6s ² 4f 5d ¹⁰ 62 |
| [Rn] | 87 Fr 7s ¹ | 88 Ra 7s ² | 103 Lr $7s^25f^{14}$ $6d^8$ | 104 Rf 7s ² 5f ¹⁴ 6d ² | 105 Db 7s ² 5f ¹⁴ 6d ³ | 106 Sg 7s ² 5f ¹⁴ 6d ⁴ | 107 Bh 7s ² 5f ¹⁴ 6d ⁵ | 108 Hs 7s ² 5f ¹⁴ 6d ⁶ | 109 Mt 7s ² 5f ¹⁴ 6d ⁷ | $110 \\ Ds \\ 7s^2 5f^{14} \\ 6d^8$ | 111 Rg 7s ² 5f ¹⁴ 6d ⁹ | $112 \\ Cn \\ 7s^2 5f^{14} \\ 6d^{10}$ | 113 Fl 7s ² 5f ¹⁴ 6d ¹⁰ 7p ¹ | 114 Cn 7s ² 5f ¹⁴ 6d ¹⁰ 7p ² | 115 7s ² 5f ¹⁴ 6d ¹⁰ 7p ³ | 116 Lv 7s ² 5f ¹⁴ 6d ¹⁰ 7p ⁴ | 117 7s ² 5f ¹⁴ 6d ¹⁰ 7p ⁵ | 118 7s ² 5f 6d ¹⁰ 7j |
| [Xe] | Lanthanide series | | | 57 La $6s^25d^1$ | $58 \\ Ce \\ 6s^2 4f^1 \\ 5d^2$ | $59 \\ Pr \\ 6s^2 4f^3$ | $60 \\ Nd \\ 6s^2 4f^4$ | $\mathop{\mathbf{Pm}}_{6s^24f^5}$ | 62 Sm 6s ² 4f ⁶ | 63 Eu _{6s²4f⁷} | 64 Gd 65 ² 4f ⁷ 5d ¹ | 65 Tb 6s ² 4f ⁹ | 66 Dy 6s ² 4f ¹⁰ | 67 Ho 6s ² 4f ¹¹ | $ \begin{array}{c} 68 \\ \mathbf{Er} \\ 6s^2 4 f^{12} \end{array} $ | $69 \\ Tm \\ 6s^2 4f^{13}$ | 70 Yb 6s ² 4f ¹⁴ | 1 |
| [Rn] | Actinide series | | | 89 Ac 7s ² 6d ¹ | 90 Th $7s^{2}6d^{2}$ | 91 Pa $7s^25f^2$ $6d^1$ | 92 $U_{7s^25f^3}_{6d^1}$ | 93 Np 7s ² 5f ⁴ 6d ¹ | 94 Pu 7s ² 5f ⁶ | 95 Am 7s ² 5f ⁷ | 96 Cm $7s^25f^7$ $6d^1$ | 97 Bk 7s ² 5f ⁹ | 98 Cf 7s ² 5f ¹⁰ | 99 Es 7s ² 5f ¹¹ | 100 Fm 7s ² 5f ¹² | $101 \\ Md \\ 7s^25f^{13}$ | 102 No 7s ² 5f ¹⁴ | |
| | | | | | Aetals | | Meta | illoids | | Non | metals | | A. | | | | | t. |

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

Some Anomalies

| | 1A 1 H 1s ¹ | 2A 2 | | | | | | | | | | | 3A 13 | 4A 14 | 5A 15 | 6A 16 | 7A 17 | 8A 18 2 He 1s ² |
|--------------|---------------------------------------|-----------------------------|--|--|---|---|---|---|--|--|---|--|---|--|---|---|---|---|
| Core [He] | 3 Li 2s ¹ | 4 Be 2s ² | | | | | | | | | | | 5 B $2s^22p^1$ | $\begin{array}{c} 6 \\ \mathbf{C} \\ 2s^2 2p^2 \end{array}$ | 7 N $2s^22p^3$ | 8 0 2s ² 2p ⁴ | 9 \mathbf{F} $2s^22p^5$ | 10 Ne 2s ² 2p |
| [Ne] | 11 Na 3s ¹ | 12 Mg 3s ² | 3B 3 | 4B 4 | 5B 5 | 6B 6 | 7B 7 | 8 | 8B 9 | 10 | 1 B 11 | 2B 12 | $13 \\ Al \\ 3s^2 3p^1$ | $ \begin{array}{c} 14 \\ Si \\ 3s^2 3p^2 \end{array} $ | 15 P 3s ² 3p ³ | | $17 \\ Cl \\ 3s^2 3p^5$ | 18 Ar 3s ² 3p |
| [Ar] | 19 K 4s ¹ | 20 Ca 4s ² | 21 Sc $4s^23d^1$ | $22 \\ Ti \\ 4s^2 3d^2$ | 23 V 4s ² 3d ³ | $24 \\ Cr \\ 4s^1 3d^5$ | 25 Mn 4s ² 3d ⁵ | 26 Fe 4s ² 3d ⁶ | 27 Co 4s ² 3d ⁷ | 28 Ni 4s ² 3d ⁸ | $29 \\ Cu \\ 4s^1 3d^{10}$ | 30 Zn 4s ² 3d ¹⁰ | $31 \\ Ga \\ 4s^2 3d^{10} \\ 4p^1 $ | 32 Ge 4s ² 3d ¹⁰ 4p ² | $33 \\ As \\ 4s^2 3d^{10} \\ 4p^3 $ | 34 Se $4s^23d^{10}$ $4p^4$ | $35 \\ Br \\ 4s^2 3d^{10} \\ 4p^5 \\ 4p^5 \\$ | 36 Kr 4s ² 3d 4p ⁶ |
| [Kr] | 37 Rb 5s ¹ | 38 Sr 5s ² | 39 Y 5s ² 4d ¹ | $40 \\ \mathbf{Zr} \\ 5s^2 4d^2$ | $\begin{array}{c} 41 \\ \mathbf{Nb} \\ 5s^2 4d^3 \end{array}$ | 42 Mo 5s ¹ 4d ⁵ | 43 Tc 5s ² 4d ⁵ | $\begin{array}{c} 44 \\ \mathbf{Ru} \\ 5\mathrm{s}^{1}4d^{7} \end{array}$ | 45 Rh _{5s¹4d⁸} | $46 \\ Pd \\ 4d^{10}$ | $47 \\ Ag \\ 5s^1 4d^{10}$ | 48 Cd 5s ² 4d ¹⁰ | $ \begin{array}{c} 49 \\ In \\ 5s^2 4d^{10} \\ 5p^1 \end{array} $ | 50 Sn $5s^24d^{10}$ $5p^2$ | $51 \\ Sb \\ 5s^24d^{10} \\ 5p^3 $ | $52 \\ Te \\ 5s^2 4d^{10} \\ 5p^4$ | 53 I 5s ² 4d ¹⁰ 5p ⁵ | 54 Xe 5s ² 4d 5p ⁶ |
| [Xe] | 55 Cs _{6s¹} | 56 Ba 6s ² | $71 \\ Lu \\ 6s^2 4f^{14} \\ 5d^3$ | $72 \\ Hf \\ 6s^2 4f^{14} \\ 5d^2$ | $73 \\ Ta \\ 6s^2 4f^{14} \\ 5d^3$ | $74 \\ W \\ 6s^2 4f^{14} \\ 5d^4$ | 75 Re 6s ² 4f ¹⁴ 5d ⁵ | 76 Os $6s^24f^{14}$ $5d^6$ | 77 Ir $6s^24f^{14}$ $5d^7$ | $78 \\ Pt \\ 6s^1 4f^{14} \\ 5d^9$ | $79 \\ Au \\ 6s^1 4f^{14} \\ 5d^{10}$ | $80 \\ Hg \\ 6s^2 4f^{14} \\ 5d^{10}$ | ${ {81}\atop{{16}}\atop{{6s}^{2}4f^{14}}\atop{{5d}^{10}6p^{1}}}$ | $82 \\ Pb \\ 6s^2 4f^{14} \\ 5d^{10} 6p^2$ | $83 \\ Bi \\ 6s^2 4f^{14} \\ 5d^{10} 6p^3$ | $84 \\ \textbf{Po} \\ \frac{6s^2 4f^{14}}{5d^{10}6p^4}$ | $85 \\ At \\ 6s^2 4f^{14} \\ 5d^{10} 6p^5$ | 86 Rn 6s ² 4f 5d ¹⁰ 6j |
| [Rn] | 87 Fr _{7s¹} | 88 Ra 7s ² | $ \begin{array}{r} 103 \\ Lr \\ 7s^2 5f^{14} \\ 6d^3 \end{array} $ | $104 \\ \mathbf{Rf} \\ 7s^2 5f^{14} \\ 6d^2$ | $105 \\ \textbf{Db} \\ 7s^2 5f^{14} \\ 6d^3$ | $106 \\ Sg \\ 7s^25f^{14} \\ 6d^4$ | $107 \\ {Bh} \\ 7s^2 5f^{14} \\ 6d^5$ | $108 \\ Hs \\ 7s^2 5f^{14} \\ 6d^6$ | 109 Mt 7s ² 5f ¹⁴ 6d ⁷ | $110 \\ \textbf{Ds} \\ 7s^2 5f^{14} \\ 6d^8$ | 111 Rg 7s ² 5f ¹⁴ 6d ⁹ | $112 \\ Cn \\ 7s^2 5f^{14} \\ 6d^{10}$ | $113 \\ F1 \\ 7s^25f^{14} \\ 6d^{10}7p^1$ | $114 \\ Cn \\ 7s^2 5f^{14} \\ 6d^{10} 7p^2$ | 115 7s ² 5f ¹⁴ 6d ¹⁰ 7p ³ | $116 \\ Lv \\ 7s^25f^{14} \\ 6d^{10}7p^4$ | 117 7s ² 5f ¹⁴ 6d ¹⁰ 7p ⁵ | 118 7s ² 5f 6d ¹⁰ 7j |
| [Xe] | Lanthanide series | | | 57 La $6s^25d^1$ | $58 \\ Ce \\ 6s^2 4f^1 \\ 5d^1$ | 59 Pr 6s ² 4f ³ | $60 \\ Nd \\ 6s^2 4f^4$ | 61 Pm 6s ² 4f ⁵ | 62 Sm 65 ² 4f ⁶ | 63 Eu 6s ² 4f ⁷ | $64 \\ Gd \\ 6s^2 4f^7 \\ 5d^1$ | 65 Tb 6s ² 4f ⁹ | 66 Dy 6s ² 4f ¹⁰ | 67 Ho $6s^24f^{11}$ | $68 \\ Er \\ 6s^2 4 f^{12}$ | $69 \\ Tm \\ 6s^2 4f^{13}$ | 70 Yb 6s ² 4f ¹⁴ | |
| [Rn] | Actinide series | | | 89 Ac $7s^26d^1$ | 90 Th $7s^{2}6d^{2}$ | 91 Pa $7s^25f^2$ $6d^1$ | 92 U $7s^{2}5f^{3}$ $6d^{1}$ | 93 Np $7s^25f^4$ $6d^1$ | 94 Pu 7s ² 5f ⁶ | 95 Am 7s ² 5f ⁷ | 96 Cm $7s^25f^7$ $6d^1$ | 97 Bk 7s ² 5f ⁹ | 98 Cf 7s ² 5f ¹⁰ | 99 Es 7s ² 5f ¹¹ | 100 Fm 7s ² 5f ¹² | $101 \\ Md \\ 7s^25f^{13}$ | 102 No 7s ² 5f ¹⁴ | |
| | | | | N | Aetals | | Meta | lloids | | Non | metals | | | | | | | |

A friend tells you that her favorite element has an electron configuration of [noble gas] $6s^24f^{14}5d^6$. Which element is it?

Chromium as an Anomaly

- For instance, the electron configuration for chromium is
- [Ar] 4s¹ 3d⁵

rather than the expected

- $[Ar] 4s^2 3d^4.$
- 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.

Boron, atomic number 5, occurs naturally as two isotopes, ¹⁰B and ¹¹B, with natural abundances of 19.9% and 80.1%, respectively. (a) In what ways do the two isotopes differ from each other? Does the electronic configuration of ¹⁰B differ from that of ${}^{11}B?$ (b) Draw the orbital diagram for an atom of ¹¹B. Which electrons are the valence electrons? (c) Indicate three major ways in which the 1s electrons in boron differ from its 2s electrons. (d) Elemental boron reacts with fluorine to form BF₃, a gas. Write a balanced chemical equation for the reaction of solid boron with fluorine gas. (e) ΔH°_{f} for BF₃(g) is -1135.6 kJ mol⁻¹. Calculate the standard enthalpy change in the reaction of boron with fluorine. (f) Will the mass percentage of F be the same in ¹⁰BF₃ and ¹¹BF₃? If not, why is that the case?